OCR A | Year 12 chemistry

Amount of substance

Complete notes

Clear explanations, precise definitions and worked examples for the whole topic. Use the contents or PDF bookmarks to find a section. TEST IT references point to the matching questions in the separate topic Questions & answers pack.

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Relative masses & the mole

1. Relative isotopic and atomic masses

SPEC 2.1.1(c)

CRACK IT

Atoms are too small for ordinary gram measurements to be convenient. Relative masses compare them with a shared reference: **one-twelfth of the mass of one carbon-12 atom**.

Relative isotopic mass describes an atom of one particular isotope. Relative atomic mass, A_r, is a weighted mean for the element, allowing for its isotope abundances.

Because each relative mass is a ratio of two masses, it has no unit. Keep that separate from molar mass, which is normally expressed in g mol^{-1}. The definitions below give the precise exam wording.

Use the A_r values on your supplied Periodic Table, which gives values to one decimal place, for example Cl = 35.5. Use the exact data given in a question, including any more precise isotopic masses.

DEFINITIONS & KEY TERMS

Relative isotopic mass

The mass of an atom of an isotope compared with one-twelfth of the mass of an atom of carbon-12.

Relative atomic mass, A_r

The weighted mean mass of an atom of an element compared with one-twelfth of the mass of an atom of carbon-12.

EXAM TRAP For A_r, "average" alone misses the weighting. Include both weighted mean and the carbon-12 comparison.

TEST IT

Try question 1 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Define the two required relative masses and explain the carbon-12 standard.

2. Calculate relative molecular and formula masses

SPEC 2.1.1(e)

CRACK IT

A formula is an atom-counting instruction. To find a relative molecular or formula mass, add the A_r contribution of **every atom in the formula**.

1. Write down how many atoms of each element are present.

2. Multiply the count of each element by its A_r.

3. Add the contributions.

Brackets are the usual trap. In $Al_2(SO_4)_3$, the outside 3 multiplies both S and O, giving $Al_2S_3O_{12}$. It does not multiply the Al.

Use relative molecular mass for separate molecules and relative formula mass for a formula describing a giant structure. Give no unit for a relative mass. The same numerical value used as a molar mass has units of g mol^{-1}.

Relative mass = sum of all A_r values in the formula

A_r relative atomic mass; no unit.

WORKED EXAMPLE

Brackets and subscripts

Calculate the relative formula mass of $Ca(NO_3)_2$. Use Ca 40.1, N 14.0 and O 16.0.

Worked answer

Relative molecular/formula mass = sum of the A_r contributions of every atom

Quantity	Ca	N	O
Number of atoms	1	2	6
A_r	40.1	14.0	16.0
Contribution	Contribution = atom count × A_r 40.1	Contribution = atom count × A_r 28.0	Contribution = atom count × A_r 96.0

40.1 + 28.0 + 96.0 = 164.1

Relative formula mass = 164.1.

EXAM TRAP The outside 2 multiplies both N and O. It does not multiply Ca.

TEST IT

Try question 2 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Count every atom in a formula, including brackets, and leave relative values unitless.

3. Relative mass and molar mass

SPEC 2.1.3(a)(iv); 2.1.1(e)

CRACK IT

Relative mass and molar mass often have the same numerical value, but they describe different quantities.

KEY DEFINITION

Relative molecular or formula mass is a comparison with the carbon-12 reference and has no unit. Molar mass, M, is the mass of one mole and is usually given in g mol^{-1}.

For water, M_r = 18.0 and M = 18.0 g mol^{-1}. Use the molar mass when calculating n = m/M, with mass in grams. Do not put g mol^{-1} on M_r itself.

Quantity	Meaning	Unit
A_r	Relative mass for an element's atoms	None
M_r / relative formula mass	Relative value for a molecule or formula unit	None
M	Mass per mole of the stated substance	g mol^{-1}

WORKED EXAMPLE

Water: the same number, different quantities

Use H 1.0 and O 16.0 to compare $M_r(H_2O)$, $M(H_2O)$ and the mass of 2.00 mol H_2O.

Worked answer

$M_r(H_2O) = 2(1.0) + 16.0 = 18.0$

$M(H_2O) = 18.0$ g mol^{-1}

$m = n \times M$
$m = 2.00 \times 18.0 = 36.0$ g

The relative mass has no unit; the molar mass is in g mol^{-1}; the sample mass is in g.

EXAM TRAP Read the symbol before choosing a unit. M_r = 18.0, but M = 18.0 g mol^{-1}.

TEST IT

Try question 3 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Distinguish relative mass from molar mass and use the correct units.

4. The mole and Avogadro constant

SPEC 2.1.3(a)(i–iii)

CRACK IT

KEY DEFINITION

THE EXACT DEFINITION

One mole contains exactly $6.02214076 \times 10^{23}$ specified entities. The exact Avogadro constant is $6.02214076 \times 10^{23}$ mol^{-1}. For calculations, use the supplied rounded value, $N_A = 6.02 \times 10^{23}$ mol^{-1}.

A mole is a unit for measuring amount of substance, rather like a dozen is a unit for counting items. You can have one mole, two moles or half a mole.

KEY DEFINITION

The mole and the Avogadro constant do different jobs. **mol is the unit; N_A tells you how many specified particles there are per mole.** Think of “dozen” and 12: the word names the counting unit, while the number tells you how many items it contains.

Use **$N_A = 6.02 \times 10^{23}$ mol^{-1}**, rounded to three significant figures. So 1 mol contains approximately 6.02×10^{23} specified particles. Use a different value if a question explicitly supplies one.

The amount of substance has symbol n and unit mol. Name what you are counting: atoms, molecules, formula units, ions or electrons. One mole of each contains the same number of entities, but their masses differ. One mole of O_2 molecules contains two moles of oxygen atoms.

Substance	Basic entity
Ne	Atom
O_2	Molecule
NaCl	Formula unit
Al^{3+}	Ion

$$N = nN_A; \; n = N/N_A$$

N number of the specified particles; a count with no unit; **N_A** Avogadro constant, 6.02×10^{23} mol^{-1}; **n** amount of the specified substance or particles, in mol.

EXAM TRAP Write the formula or entity beside n and N. “Particles” alone is too vague for a final answer.

TEST IT

Try question 4 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Explain amount of substance and name the entity counted by the Avogadro constant.

Mass, moles & particles

5. Choose the entity and formula multiplier

SPEC 2.1.3(a)

CRACK IT

Before using a formula, decide **what the question wants you to count**. A molecule, an atom and an ion are not interchangeable labels.

Use molecules for substances made of separate molecules, such as water. For an ionic compound such as NaCl, use formula units rather than molecules.

Then look inside the formula. One H_2O molecule contains two H atoms and one O atom. One formula unit of an ionic compound may contain more than one of a particular ion. Apply that multiplier only when the question asks for those component particles.

Formula	One-entity statement
CO_2	One molecule contains one C atom and two O atoms.
$Al_2(SO_4)_3$	One formula unit represents two Al^{3+} and three SO_4^{2-} ions, five ions in total.
Na^+	With atomic number 11, a Na^+ ion has 11 −1 =10 electrons.

Amount of requested particles = amount of substance × number per entity

WORKED EXAMPLE

Sulfate ions from a formula unit

Find the amount in moles of sulfate ions represented by 0.120 mol $Al_2(SO_4)_3$.

Worked answer

There are three sulfate ions per formula unit.

Moles of sulfate ions = moles of $Al_2(SO_4)_3$ × sulfate ions per formula unit
$n(SO_4^{2-}) = 0.120 \times 3$

0.360 mol sulfate ions.

EXAM TRAP For total ions, add the cation and anion subscripts. For $Al_2(SO_4)_3$ use 2 +3 =5, not 2 ×3. A sulfate ion remains one ion.

TEST IT

Try question 5 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Translate a formula into the exact atoms, ions or electrons requested.

6. Mass and amount

SPEC 2.1.3(e)(i)

CRACK IT

KEY DEFINITION

Molar mass tells you the mass of one mole. It is the bridge between the mass you can weigh and the amount in moles.

1. Find M by adding the A_r contributions in the formula. Use g mol^{-1} for molar mass.

2. For grams to moles, divide by M: n = m/M.

3. For moles to grams, multiply by M: m = nM.

Check the units before substituting. When M is in g mol^{-1}, mass must be in grams.

A quick check is to compare with one mole. If M = 100.1 g mol^{-1}, then 0.350 mol must have a mass below 100.1 g.

$$n = m/M;\ m = nM;\ M = m/n$$

M molar mass of the specified substance, in g mol^{-1}; **m** mass of the specified sample or substance, in g; **n** amount of the specified substance or particles, in mol.

WORKED EXAMPLE

Mass to amount

Find the amount in moles of H_2SO_4 in 9.81 g. M(H_2SO_4) =98.1 g mol^{-1}.

Worked answer

n = m/M
n =9.81/98.1

0.100 mol H_2SO_4.

WORKED EXAMPLE

Amount to mass

Find the mass of 0.125 mol Na_2CO_3. M =106 g mol^{-1}.

Worked answer

m = n × M
m =0.125 ×106 =13.25 g

13.3 g Na_2CO_3 to three significant figures.

EXAM TRAP M must describe the entire chemical formula, not one element inside it. A mass about one-tenth of the molar mass corresponds to about 0.1 mol.

WHEN DENSITY IS GIVEN

For a pure liquid, mass = density × volume. Match the volume unit to the density unit: a density in g cm^{-3} multiplied by a volume in cm^3 gives mass in g. Then use n = m/M.

M molar mass of the specified substance, in g mol^{-1}; **m** mass of the specified sample or substance, in g; **n** amount of the specified substance or particles, in mol.

TEST IT

Try question 6 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Use n = m/M in both directions with mass in grams and molar mass in g mol^{-1}.

7. Amount and particle number

SPEC 2.1.3(a)(iii)

CRACK IT

The Avogadro constant tells you how many specified entities are in one mole. Use it to move between moles and particle number.

- **Moles to particles:** multiply by N_A, using $N = nN_A$.
- **Particles to moles:** divide by N_A, using $n = N/N_A$.

N is a count, so it has no unit. Label the answer with the entity, such as molecules or chloride ions. Use the supplied value of N_A; these questions normally use 6.02 × 10^{23} mol^{-1}.

If the target is an atom or ion within a formula, include its formula multiplier as well.

$$N = nN_A; n = N/N_A$$

N number of the specified particles; a count with no unit; N_A Avogadro constant, 6.02 × 10^{23} mol^{-1}; **n** amount of the specified substance or particles, in mol.

WORKED EXAMPLE

Moles to formula units

How many $MgCl_2$ formula units are represented by 0.0450 mol?

Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

$N = nN_A$
$N = 0.0450 \times 6.02 \times 10^{23} = 2.709 \times 10^{22}$

2.71×10^{22} $MgCl_2$ formula units.

WORKED EXAMPLE

Atoms to moles

Find the amount in moles represented by 1.204×10^{23} Ne atoms.

Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

$n = (1.204 \times 10^{23})/(6.02 \times 10^{23})$

0.200 mol Ne.

EXAM TRAP N is a count with no unit, but still needs the label atoms, molecules, formula units, ions or electrons.

TEST IT

Try question 7 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Convert between amount and a named particle count using the Avogadro constant.

8. Count atoms, ions and electrons

SPEC 2.1.3(a)

CRACK IT

The amount of a compound is not automatically the amount of every particle inside it. Use the formula to count how many of the requested particles occur in one molecule or formula unit.

1. Find the amount in moles of the stated substance.

2. Multiply by the number of target atoms or ions in each formula.

3. Use $N = nN_A$ if the question asks for the actual particle number.

For electron counts, use atomic number and ionic charge first. Al^{3+} has 13 - 3 = 10 electrons; Cl^- has 17 + 1 = 18. Multiply the amount of those ions by their electron count, then by N_A.

Write the entity beside intermediate amounts. This makes it easier to spot an incorrect multiplier.

WORKED EXAMPLE

Oxygen atoms in molecules

Find the number of O atoms in 0.250 mol CO_2.

Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

Moles of oxygen atoms = moles of CO_2 × oxygen atoms per molecule
n(O atoms) $= 0.250 \times 2 = 0.500$ mol

$N(\text{O atoms}) = nN_A$
$N(\text{O atoms}) = 0.500 \times 6.02 \times 10^{23}$

3.01×10^{23} O atoms.

WORKED EXAMPLE

Total ions

Find the total number of ions represented by 0.0300 mol $Al_2(SO_4)_3$.

Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

Each formula unit represents 2 +3 =5 ions.

Total moles of ions = moles of $Al_2(SO_4)_3$ × total ions per formula unit
n(ions) $= 0.0300 \times 5 = 0.150$ mol

$N(\text{ions}) = nN_A$
$N(\text{ions}) = 0.150 \times 6.02 \times 10^{23}$

9.03×10^{22} ions.

EXAM TRAP SO_4^{2-} contains five atoms but is one polyatomic ion. Do not split it when counting ions.

TEST IT

Try questions 8, 9, 10 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Apply formula multipliers once and keep polyatomic ions together.

9. The mass–mole–particle route

SPEC 2.1.3(a), 2.1.3(e)(i)

CRACK IT

For a mass-to-particles question, make two clear moves: **mass to moles, then moles to particles**. The mole stage keeps the units and the target entity visible.

1. Divide the mass by the molar mass to find moles of the substance.

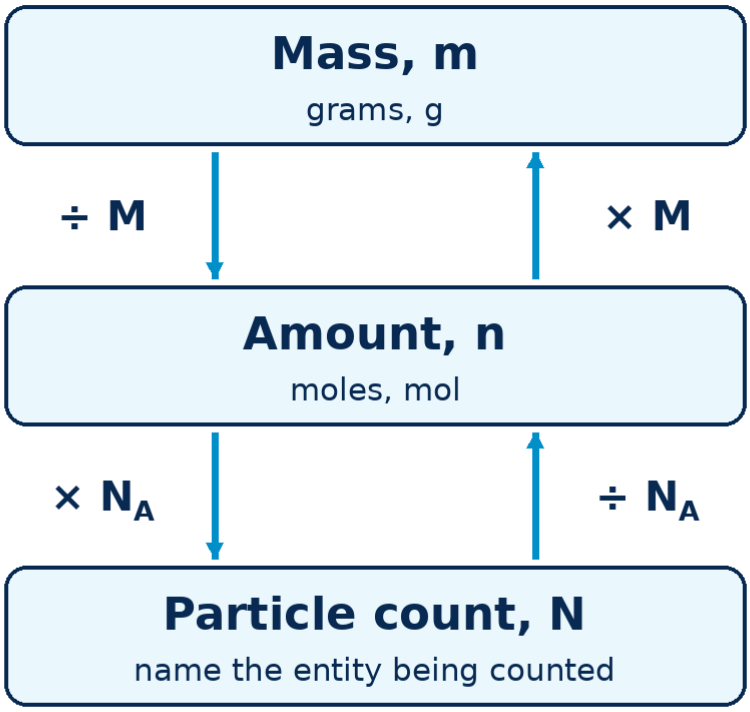
2. If needed, use the formula to find moles of the particular atom or ion being counted.

3. Multiply the appropriate mole amount by N_A.

The combined expression $N = (m/M)N_A$ is valid when N counts the same basic entity described by M. If you need a component particle, remember its formula multiplier. Writing both stages can make that easier to see.

Mass → divide by M → amount → formula multiplier → multiply by N_A → particle count

Use moles as the bridge



Use molar mass M in g mol^{-1}.

$N_A = 6.02 \times 10^{23}$ mol^{-1}

For a component atom or ion, include its formula multiplier.

WORKED EXAMPLE

Mass to chloride ions

Find the number of chloride ions represented by 5.55 g $CaCl_2$. $M(CaCl_2)$ =111 g mol^{-1}.

Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

$n(CaCl_2) = m/M$
$n(CaCl_2)$ =5.55/111 =0.0500 mol

Moles of chloride ions = moles of $CaCl_2$ × chloride ions per formula unit
$n(Cl^-)$ =0.0500 ×2 =0.100 mol

$N(Cl^-) = nN_A$
$N(Cl^-)$ =0.100 ×6.02 $\times 10^{23}$

6.02 $\times 10^{22}$ chloride ions.

EXAM TRAP The first particle count from mass/M ×N_A counts basic entities. Apply a further multiplier only when the question asks for component atoms, ions or electrons.

TEST IT

Try questions 11, 12 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Connect measured mass to the exact particle requested.

10. Reverse and unfamiliar calculations

SPEC 2.1.3(a), 2.1.3(e)(i)

CRACK IT

A reverse calculation uses the same relationships in the opposite direction. Start by writing down the target and the information you have.

1. For a particle count to moles, divide by N_A.

2. Undo any formula multiplier if you need moles of the whole compound rather than its component particles.

3. For moles to mass, use m = nM.

4. If molar mass is the unknown, rearrange to M = m/n.

Label the substance and entity at every step. The worked examples show how the route changes when the question starts with a particle count or asks for an unfamiliar molar mass.

WORKED EXAMPLE

Particles to mass

Find the mass of 1.204×10^{23} CO_2 molecules. M =44.0 g mol^{-1}.

Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

$n = (1.204 \times 10^{23}) / (6.02 \times 10^{23}) = 0.200$ mol

m = n × M
m =0.200 ×44.0

8.80 g CO_2.

WORKED EXAMPLE

Identify an element from chloride-ion count

A 9.53 g sample of XCl_2 represents 1.204×10^{23} chloride ions. Use $A_r(Cl) = 35.5$.

Calculate the relative atomic mass of X and identify the element. Use $N_A = 6.02 \times 10^{23}$ mol^{-1}.

Worked answer

$n(Cl^-) = 0.200$ mol; Moles of XCl_2 = moles of chloride ions ÷ chloride ions per formula unit
$n(XCl_2) = 0.200/2 = 0.100$ mol

$M(XCl_2) = m/n$
$M(XCl_2) = 9.53/0.100 = 95.3$ g mol^{-1}

$A_r(X) = 95.3 - 2(35.5) = 24.3$

X is magnesium, Mg ($A_r = 24.3$).

EXAM TRAP Reverse the formula multiplier before finding the mass of the whole compound. Keep unrounded intermediate amounts.

TEST IT

Try questions 13, 14, 15 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Work from a particle count to mass or an unknown molar mass.

11. Precision and final checks

SPEC 1.1.3(c); 2.1.3(a), (e)(i)

CRACK IT

A good method can still lose marks through units, powers of ten or early rounding. Make a final check part of every calculation.

• Keep the calculator's full value during intermediate steps. Round only the final answer.

• Use sensible significant figures from the measured or supplied data. Exact subscripts and equation coefficients do not limit them.

• Check the unit and the entity: mol, g, atoms and molecules describe different answers.

• Check the direction. Dividing by a large molar mass should give fewer moles for a fixed mass.

In standard form, one non-zero digit comes before the decimal point. For example, 0.003045 mol = 3.045×10^{-3} mol. At three significant figures, 7.849×10^{22} becomes 7.85×10^{22}.

Check	What good working shows
Units	m in g; n in mol; M in g mol^{-1}; N has no unit.
Standard form	One non-zero digit before the decimal point.
Significant figures	Round the final result using the data precision or the instruction.
Entity	Name atoms, molecules, formula units, ions or electrons.
Sense check	Less than 1 mol of specified entities contains fewer of those entities than 1 mol.^{23} basic entities.

EXAM TRAP Do not round an intermediate amount such as 0.04996 mol to 0.050 mol and then reuse it. Store the full value.

TEST IT

Try question 16 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Finish with correct units, standard form, significant figures and a named entity.

Equations & reacting masses

12. Balanced equations and state symbols

SPEC 2.1.2(b)

CRACK IT

A reaction rearranges atoms; it does not create or destroy them. A balanced equation must therefore have the same number of atoms of each element on both sides.

1. Write the correct formulae before balancing.

2. Count the atoms on each side. Start with an element appearing in one reactant and one product if possible.

3. Change the coefficients in front of whole formulae. A coefficient multiplies every atom in that formula.

4. Recount every element, use the smallest whole-number coefficients and check the state symbols.

Never change a subscript to make the numbers balance. Changing H_2O to H_2O_2 would change the substance, not balance the original reaction.

State symbol	Meaning
(s)	Solid
(l)	Liquid
(g)	Gas
(aq)	Dissolved in water

WORKED EXAMPLE

Aluminium reacts with oxygen

Balance Al + O_2 → Al_2O_3.

Worked answer

1. Use 2 Al_2O_3 to make 6 oxygen atoms on the right.
2. Use 3 O_2 to supply 6 oxygen atoms.
3. Use 4 Al to supply 4 aluminium atoms.

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$.

EXAM TRAP A coefficient multiplies every atom in the following formula. An equation ratio is a mole ratio, not a mass ratio.

TEST IT

Try question 17 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Balance atoms using coefficients and show the physical states.

13. Ionic equations and spectator ions

SPEC 2.1.2(b)

CRACK IT

KEY DEFINITION

WRITE IONIC FORMULAE FIRST

An ionic formula has zero total charge. Choose the smallest whole-number ratio of positive and negative ions; use brackets if more than one polyatomic ion is needed. For example, Zn^{2+} and CO_3^{2-} combine in a 1:1 ratio to give $ZnCO_3$. Recall the following ions; other ion charges are supplied when required.

Ion	Formula and charge
Nitrate	NO_3^-
Carbonate	CO_3^{2-}
Sulfate	SO_4^{2-}
Hydroxide	OH^-
Ammonium	NH_4^+
Zinc	Zn^{2+}
Silver	Ag^+

An ionic equation removes the ions that are only watching the reaction. These **spectator ions** appear unchanged on both sides.

1. Start with the balanced full equation, including state symbols.

2. Separate appropriate aqueous ionic substances, strong acids and strong alkalis into their ions. Keep the coefficients.

3. Keep solids, liquids, gases and undissociated molecular substances intact. An (aq) label alone does not mean every substance should be split.

4. Cancel spectator ions that are identical on both sides.

5. Check both atom numbers and total charge.

For a strong acid reacting with a strong alkali, the reacting particles are $H^+(aq)$ and $OH^-(aq)$, which form water. The remaining ions do not need to appear in the net equation.

WORKED EXAMPLE

A silver chloride precipitate

Mix aqueous silver nitrate and sodium chloride.

Write the balanced full equation and the ionic equation, including state symbols. Identify the spectator ions.

Worked answer

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

Cancel Na^+ and NO_3^-, the spectator ions.

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

EXAM TRAP Do not split the precipitate into ions. Check the total electrical charge as well as each element.

TEST IT

Try question 18 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Cancel unchanged aqueous ions while conserving both atoms and charge.

14. Reacting masses and reverse calculations

SPEC 2.1.3(e)(i), (g)

CRACK IT

A reacting-mass calculation has one important bridge: the balanced equation connects **moles of one substance to moles of another**. It does not connect their gram values directly.

1. Write the balanced equation above your grid.
2. Put in the known mass and the relevant molar masses.
3. Use n = m/M to reach the moles row.
4. Move across that row using the equation's coefficient ratio.
5. Use m = nM to find the required mass.

The grid works in either direction. If the product mass is known, start there and follow the mole ratio backwards. Use g mol^{-1} for M, which is numerically equal to the corresponding relative mass.

WORKED EXAMPLE

Mass of aluminium oxide

5.40 g Al reacts with excess oxygen. Find the mass of Al_2O_3. M(Al) = 27.0; M(Al_2O_3) = 102.0 g mol^{-1}.

Worked answer

$4Al + 3O_2 \rightarrow 2Al_2O_3$

Quantity	Al	Al_2O_3
Mass / g	5.40	m = n × M 0.100 × 102.0
M / g mol^{-1}	27.0	102.0
Amount / mol	n = m/M 5.40/27.0 = 0.200	n = n(known) × equation ratio 0.200 × 2/4 = 0.100

10.2 g Al_2O_3.

EXAM TRAP Do not multiply a mass by a coefficient ratio. Convert to moles first.

TEST IT

Try questions 19, 20, 21 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Use a balanced mole ratio between two substances.

Limiting reactants

15. Limiting reactants and excess left over

SPEC 2.1.3(e)(i), (g)

CRACK IT

KEY DEFINITION

The limiting reactant is the one that runs out first. Once it is used up, no more product can form, even if some of the other reactant remains.

A **coefficient** is a big number written in front of an element or compound in a balanced equation. It multiplies the whole formula. If no number is written, the coefficient is 1. For example, in **$2Al + 3Cl_2 \rightarrow 2AlCl_3$**, the coefficients are 2, 3 and 2.

1. Convert each supplied reactant quantity into moles.

2. Divide each mole amount by its own coefficient in the balanced equation.

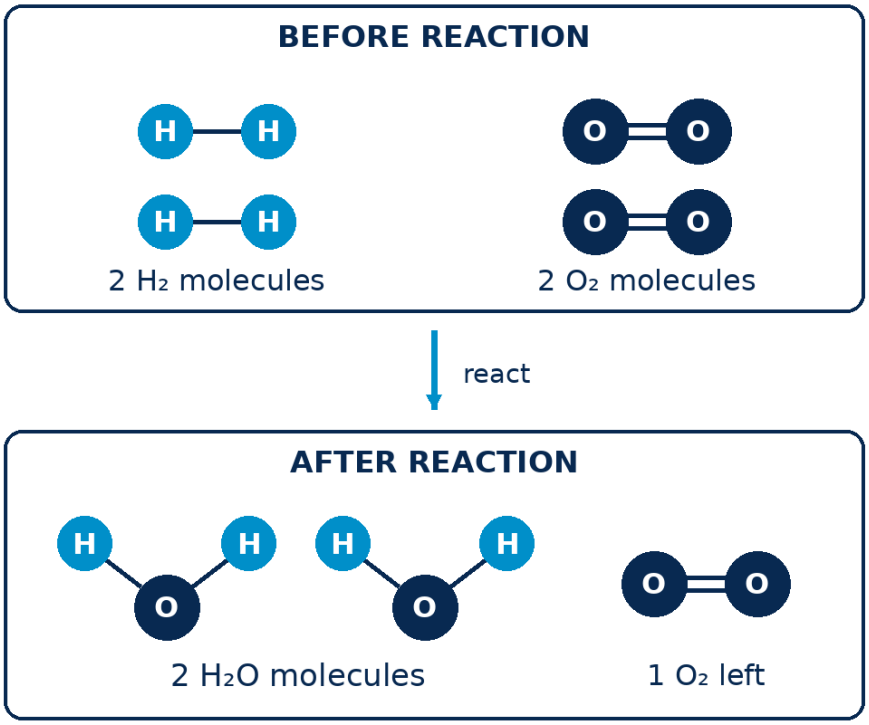
3. The smaller value identifies the limiting reactant and sets the maximum product amount.

4. To find excess left over, use the limiting amount to calculate how much of the other reactant reacts. Subtract that from its starting amount, then convert to mass if needed.

Compare n/coefficient, not the original masses or the mole amounts alone. Different coefficients mean different reaction requirements.

Which reactant runs out?

$2H_2 + O_2 \rightarrow 2H_2O$



H_2 is limiting: it is all used up.
O_2 is in excess: some remains.

Molecule model; atom sizes are not to scale.

WORKED EXAMPLE

Magnesium and oxygen

2.40 g Mg reacts with 0.960 g O_2. Use M(Mg) = 24.3, M(O_2) = 32.0 and M(MgO) = 40.3 g mol^{-1}.

Write the balanced equation, identify the limiting reactant and calculate the maximum mass of magnesium oxide formed.

Worked answer

$2Mg + O_2 \rightarrow 2MgO$

Quantity	Mg available	O_2 available	MgO formed
Mass / g	2.40	0.960	m = n × M 0.0600 × 40.3
M / g mol^{-1}	24.3	32.0	40.3
Amount / mol	n = m/M n = 2.40/24.3 = 0.098765…	n = m/M n = 0.960/32.0 = 0.0300	n = n(known) × equation ratio 0.0300 × 2/1 = 0.0600
n ÷ coefficient	Comparison value = n/coefficient 0.0493827…	0.0300 (limiting)	Not compared

Oxygen is limiting; mass of MgO = 2.42 g.

EXAM TRAP The smallest mass or the smallest number of moles is not automatically limiting. Include the balanced coefficients.

TEST IT

Try questions 22, 23, 24 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Compare the available amounts with their balanced coefficients.

Empirical & molecular formulae

16. Empirical formulae from composition

SPEC 2.1.3(b), (c)

CRACK IT

An empirical formula gives the **simplest whole-number ratio of atoms**. A mass ratio is not an atom ratio because different atoms have different masses.

1. Find the mass of each element. For percentages, assume a 100 g sample so each percentage becomes grams.

2. Divide each mass by its A_r to find moles.

3. Divide every mole amount by the smallest one.

4. If a simple fraction remains, multiply the entire ratio to obtain whole numbers.

5. Use those whole numbers as the formula subscripts.

Do not round 1 : 1.5 to 1 : 2. Multiply both values by 2 to get **2 : 3**. For a ratio of **1 : 1.33…**, multiply both by 3 to get **3 : 4**. This may look like 1 : 1.3 if rounded too early, so keep the full calculator values when choosing the multiplier.

Ratio after division	Next step
1 : 1.5	Multiply both by 2 → 2 : 3
1 : 1.33…	Multiply both by 3 → 3 : 4
1 : 1.25	Multiply both by 4 → 4 : 5

WORKED EXAMPLE

An iron oxide

1.116 g Fe forms 1.596 g of an oxide. Use A_r(Fe) = 55.8 and A_r(O) = 16.0.

Determine the empirical formula of the iron oxide.

Worked answer

Mass of oxygen = 1.596 − 1.116 = 0.480 g

Quantity	Fe	O
Mass / g	1.116	0.480
A_r	55.8	16.0
Amount / mol	n = m/M n = 1.116/55.8 = 0.0200	n = m/M n = 0.480/16.0 = 0.0300
÷ smallest	Ratio = amount ÷ smallest amount 0.0200/0.0200 = 1	Ratio = amount ÷ smallest amount 0.0300/0.0200 = 1.5
Whole-number ratio	2	3

Fe_2O_3.

EXAM TRAP Do not round 1.5 to 2. Multiply the whole ratio by 2. Find a gained or lost element mass by subtraction before calculating its amount.

TEST IT

Try questions 25, 26 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Convert element masses into the simplest whole-number atom ratio.

17. Molecular formulae and percentage composition

SPEC 2.1.3(b), (c)

CRACK IT

KEY DEFINITION

The empirical formula gives the simplest atom ratio. The molecular formula gives the **actual numbers of atoms in one molecule**, so it is a whole-number multiple of the empirical formula.

1. Find the empirical formula first. If composition is given as percentages, begin with a 100 g sample.

2. Add the A_r contributions to find the empirical-formula mass.

3. Divide the supplied M_r by the empirical-formula mass.

4. Multiply every empirical-formula subscript by that whole-number multiplier.

Check that the final molecular formula adds up to the supplied M_r. The multiplier applies to every element, not just the last one.

Multiplier = M_r / empirical formula mass

M_r relative molecular or formula mass; no unit.

WORKED EXAMPLE

From percentages to a molecular formula

A compound contains 54.5% C, 9.1% H and 36.4% O and has M_r = 88.0. Use C 12.0, H 1.0 and O 16.0.

These are percentages by mass. Determine the empirical formula and then the molecular formula.

Worked answer

Quantity	C	H	O
Mass / g	54.5	9.1	36.4
A_r	12.0	1.0	16.0
Amount / mol	n = m/M n = 54.5/12.0 = 4.54167...	n = m/M n = 9.1/1.0 = 9.10	n = m/M n = 36.4/16.0 = 2.275
÷ smallest	1.996... ≈ 2	Ratio = amount ÷ smallest amount 9.10/2.275 = 4	Ratio = amount ÷ smallest amount 2.275/2.275 = 1
Whole-number ratio	2	4	1

Multiplier = M_r/empirical-formula mass

Empirical formula mass of C_2H_4O = 44.0; multiplier = 88.0/44.0 = 2

Molecular formula $C_4H_8O_2$.

EXAM TRAP Multiply every empirical subscript by the integer multiplier. Empirical formula and molecular formula can be identical when the multiplier is 1.

TEST IT

Try question 27 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Find the empirical formula, then use M_r to obtain the molecular formula.

18. Formulae from combustion analysis

SPEC 2.1.3(c), (g)

CRACK IT

Complete combustion lets you count carbon and hydrogen through the products they form. For a compound containing only C, H and O, all its carbon enters CO_2 and all its hydrogen enters H_2O.

1. Find moles of CO_2. Each mole contains one mole of C atoms, so $n(C) = n(CO_2)$.

2. Find moles of H_2O. Each mole contains two moles of H atoms, so $n(H) = 2n(H_2O)$.

3. Convert the C and H amounts into masses. Subtract them from the original sample mass to find the oxygen originally in the compound.

4. Convert that oxygen mass to moles, then find the empirical atom ratio.

Oxygen in the combustion products also comes from the oxygen gas used in the reaction. That is why you cannot simply count it all as oxygen from the sample. If M_r is supplied, use it to find the molecular formula.

WORKED EXAMPLE

A compound containing C, H and O

0.460 g of compound forms 0.880 g CO_2 and 0.540 g H_2O. Use C 12.0, H 1.0 and O 16.0.

The compound contains carbon, hydrogen and oxygen only and is burned completely. Determine its empirical formula.

Worked answer

$n(C) = m(CO_2)/M(CO_2)$
$n(C) = 0.880/44.0 = 0.0200$ mol; $m(C) = 0.240$ g

$n(H) = 2m(H_2O)/M(H_2O)$
$n(H) = 2(0.540/18.0) = 0.0600$ mol; $m(H) = 0.0600$ g

$m(O \text{ in sample}) = 0.460 - 0.240 - 0.0600 = 0.160$ g

Quantity	C	H	O
Mass / g	0.240	0.0600	0.160
A_r	12.0	1.0	16.0
Amount / mol	n = m/M n = 0.240/12.0 = 0.0200	n = m/M n = 0.0600/1.0 = 0.0600	n = m/M n = 0.160/16.0 = 0.0100
÷ smallest	Ratio = amount ÷ smallest amount 0.0200/0.0100 = 2	Ratio = amount ÷ smallest amount 0.0600/0.0100 = 6	Ratio = amount ÷ smallest amount 0.0100/0.0100 = 1
Whole-number ratio	2	6	1

Empirical formula C_2H_6O.

EXAM TRAP Oxygen in combustion products also comes from the oxygen supply. Use sample mass minus carbon mass minus hydrogen mass for oxygen originally in the compound.

TEST IT

Try question 28 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Trace carbon from CO_2 and hydrogen from H_2O, then find oxygen by difference.

Hydrated salts

19. Hydrated salts and water of crystallisation

SPEC 2.1.3(d), (g)

CRACK IT

A hydrated salt has water molecules in a fixed ratio within its crystal structure. This is **water of crystallisation**; it is not simply a salt dissolved in water.

DEFINITIONS & KEY TERMS

Hydrated salt

A salt containing water molecules as part of its crystal structure.

Water of crystallisation

Water molecules present in a fixed ratio within the crystal structure of a salt.

Anhydrous salt

A salt containing no water of crystallisation.

In salt·xH_2O, x tells you how many water molecules occur per formula unit of salt. Heating can remove that water and leave an anhydrous salt.

1. Find the mass of anhydrous salt left after heating.

2. Find the mass of water lost by subtracting the final sample mass from the initial sample mass.

3. Convert salt and water masses separately into moles.

4. Divide moles of water by moles of salt to find x.

Check the stated chemistry: this method assumes that the loss is water and that the residue is the intended anhydrous salt.

$$x = n(H_2O) / n(\text{anhydrous salt})$$

n amount of the specified substance or particles, in mol; **x** number of water molecules per formula unit of the anhydrous salt; a whole-number ratio with no unit.

WORKED EXAMPLE

A hydrate from water percentage

$MgSO_4 \cdot xH_2O$ contains 51.2% water by mass. $M(MgSO_4) = 120.4$; $M(H_2O) = 18.0$ g mol^{-1}.

Determine x and write the formula of the hydrated salt.

Worked answer

Quantity	$MgSO_4$	H_2O
Mass / g	48.8	51.2
M / g mol^{-1}	120.4	18.0
Amount / mol	n = m/M n = 48.8/120.4 = 0.405316…	n = m/M n = 51.2/18.0 = 2.84444…
Salt : water	1	7.01785… ≈ 7

$MgSO_4 \cdot 7H_2O$.

EXAM TRAP The water percentage is not x. Convert both water and anhydrous salt masses to moles before comparing them.

TEST IT

Try questions 29, 30 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Find the whole-number ratio of water to anhydrous salt.

20. Hydrates by heating to constant mass

SPEC 2.1.3(d), (i); 1.1.3(a), (c)

CRACK IT

The experiment measures a mass loss, so accurate weighing and controlled heating are essential.

1. Weigh the dry empty crucible, then weigh it with the hydrated salt.

2. Heat gently at first so the salt does not spit out of the crucible.

3. Cool safely and weigh. Repeat heating, cooling and weighing until the mass is constant.

4. Subtract the crucible mass to find the anhydrous salt mass.

5. Subtract the final sample mass from the initial sample mass to find water lost, then convert both masses to moles.

Repeat heating, cooling and weighing until consecutive readings agree. This constant mass is evidence that all the water has been removed. Heat enough to remove the water without

decomposing the salt, and cool before each weighing. Losing solid as well as water would spoil the calculated mole ratio.

WORKED EXAMPLE

Hydrated copper(II) sulfate

Crucible: 25.000 g. Crucible + hydrate: 27.496 g. After successive heatings: 26.622, 26.596, 26.596 g. M($CuSO_4$) = 159.6; M(H_2O) = 18.0 g mol^{-1}.

Each mass after heating includes the crucible. Use the final constant mass to determine x in $CuSO_4 \cdot xH_2O$.

Worked answer

Quantity	$CuSO_4$	H_2O lost
Mass / g	Mass of salt = mass with salt − mass of empty container 26.596 − 25.000 = 1.596	Mass of water = mass before heating − mass after heating 27.496 − 26.596 = 0.900
M / g mol^{-1}	159.6	18.0
Amount / mol	0.0100	0.0500
Salt : water	1	5

$CuSO_4 \cdot 5H_2O$.

EXAM TRAP Use the final constant mass, not the first heated mass. The crucible is not part of the salt sample.

Allow the crucible to cool before each weighing: a hot object creates air currents and buoyancy effects that can make the balance reading unreliable.

TEST IT

Try question 31 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Use mass differences to calculate water lost and anhydrous salt left.

21. Evaluate hydrate experiments

SPEC 2.1.3(d), (i); 1.1.4(c)

CRACK IT

To explain an experimental error, follow its effect on **both** quantities in x = n(water lost)/n(salt left). Saying 'x is inaccurate' does not explain the direction.

For example, incomplete removal of water makes the measured water loss too small and leaves the residue too heavy. Use the ratio to explain what that does to x.

Heat, cool and weigh the sample, then repeat until consecutive masses agree. This is evidence that no more water is being lost. Also check that the salt has not decomposed: a steady mass alone cannot identify the solid left behind.

Error	Recorded anhydrous mass	Recorded water loss	Calculated x
Incomplete water removal	Too high	Too low	Too low
Solid spits out during heating	Too low	Too high	Too high
Salt reabsorbs water before weighing	Too high	Too low	Too low

WORKED EXAMPLE

Heating is stopped too soon

A hydrate still contains some water at the final weighing.

Explain how this affects the calculated value of x in salt·xH_2O when the residue is assumed to be anhydrous.

Worked answer

The residue is recorded as anhydrous salt, so its calculated amount is too large. Too little mass is recorded as water lost, so the calculated water amount is too small.

The calculated value of x is too low.

EXAM TRAP Do not round a seriously non-integer ratio to a convenient hydrate without considering the measurements and procedure.

TEST IT

Try question 32 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Predict how a practical error changes the calculated water ratio.

Concentration & dilution

22. Amount and mass concentration

SPEC 2.1.3(e)(iii), (g)

CRACK IT

Concentration tells you how much solute is present in a given volume of **solution**. Use the final solution volume, not the volume of water added.

For concentration in mol dm^{-3}, use c = n/V with V in dm^3. Divide cm^3 by 1000 to convert to dm^3.

Mass concentration tells you grams of solute per dm^3. Calculate mass divided by solution volume, or multiply molar concentration by molar mass: mass concentration = cM.

Write the concentration unit before choosing the calculation. mol dm^{-3} and g dm^{-3} describe different quantities.

c = n/V; n = cV; V = n/c

V solution volume, in dm^3 when concentration is in mol dm^{-3}; **c** concentration of the specified solute, in mol dm^{-3}; **n** amount of the specified substance or particles, in mol.

WORKED EXAMPLE

Sodium carbonate solution

5.30 g Na_2CO_3 is made up to 250 cm^3 of solution. M = 106.0 g mol^{-1}.

Calculate the solution concentration in mol dm^{-3} and in g dm^{-3}.

Worked answer

Quantity	Na_2CO_3
Mass / g	5.30
M / g mol^{-1}	106.0
Amount / mol	n = m/M 5.30/106.0 = 0.0500

V(dm^3) = V(cm^3)/1000
V = 250/1000 = 0.250 dm^3; c = n/V
c = 0.0500/0.250 = 0.200 mol dm^{-3}

Mass concentration = mass of solute ÷ solution volume
Mass concentration = 5.30/0.250

0.200 mol dm^{-3} and 21.2 g dm^{-3}.

EXAM TRAP Use the final solution volume. Adding solute to 250 cm^3 of water is not the same as making a solution up to 250 cm^3.

TEST IT

Try question 33 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Use concentration with the volume of the solution in the correct units.

23. Dilution and choosing a stock volume

SPEC 2.1.3(e)(iii), (g)

CRACK IT

Adding water spreads the same amount of solute through a larger volume. The concentration falls, but the solute's mole amount stays the same, assuming no reaction or loss.

That gives $c_1V_1 = c_2V_2$. The volumes must use matching units. If the stock volume is unknown, rearrange first to find the volume you need to measure.

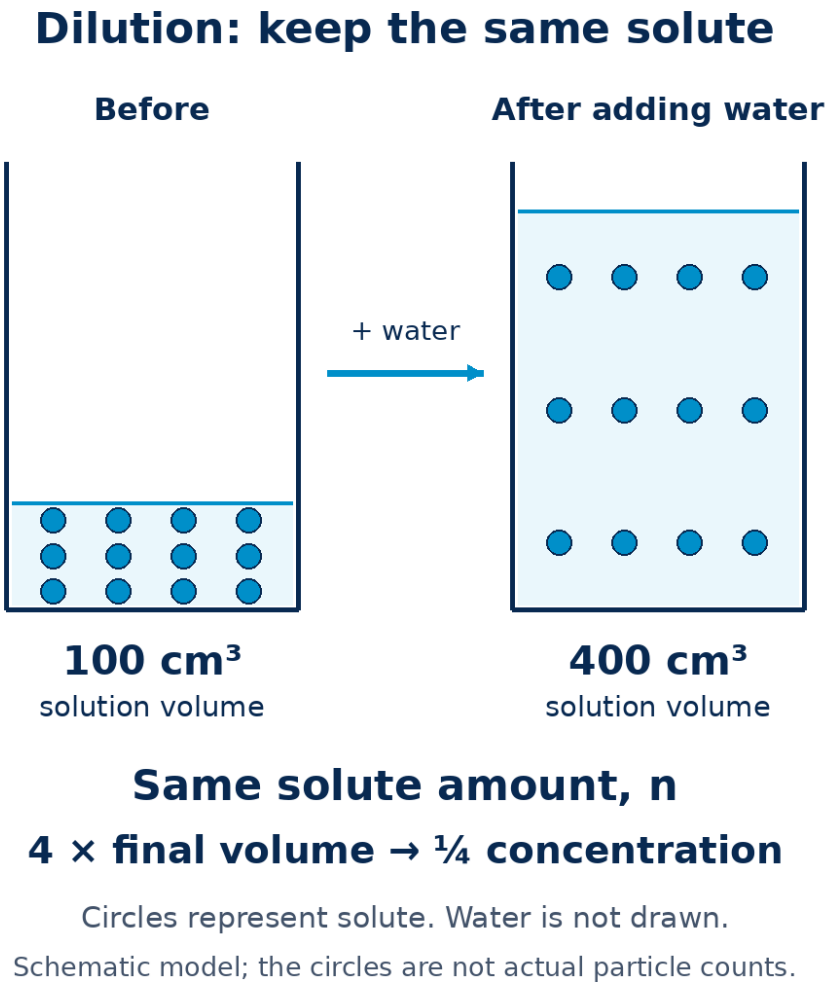
1. Measure the required stock solution with suitable volumetric apparatus.

2. Transfer it completely to a volumetric flask.

3. Add deionised water to the calibration mark, using drops near the end.

4. Stopper and invert repeatedly to mix.

Make up to the required **final volume**. Do not simply add that volume of water to the stock solution.



WORKED EXAMPLE

Dilute a stock solution

25.0 cm^3 of 0.800 mol dm^{-3} NaCl is made up to 200.0 cm^3.

Calculate the concentration of the diluted sodium chloride solution in mol dm^{-3}.

Worked answer

n(NaCl) = cV
n(NaCl) = 0.800 × 0.0250 = 0.0200 mol

$c_2 = n/V_2$
$c_2 = 0.0200/0.2000$

0.100 mol dm^{-3}.

To choose a stock volume: $V_1 = c_2V_2/c_1$

V_1 volume of stock solution used; use matching units for both volumes; **V_2** final solution volume; use matching units for both volumes; **c_1** initial stock concentration, in mol dm^{-3}; **c_2** final diluted concentration, in mol dm^{-3}.

EXAM TRAP $c_1V_1 = c_2V_2$ applies to dilution of the same solute. A reaction between different substances needs a balanced equation and its mole ratio.

TEST IT

Try question 34 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Conserve solute amount during a dilution.

24. Prepare a standard solution

SPEC 2.1.3(i); 2.1.4(d); PAG 2

CRACK IT

KEY DEFINITION

A standard solution has an accurately known concentration. To prepare one, both the mass of solute and the final solution volume must be accurate.

1. Calculate the required mass of a suitable pure, dry, stable solid. Weigh it accurately, using weighing by difference where appropriate.

2. Dissolve the solid fully in deionised water in a beaker, stirring with a glass rod.

3. Transfer the solution through a funnel into the volumetric flask.

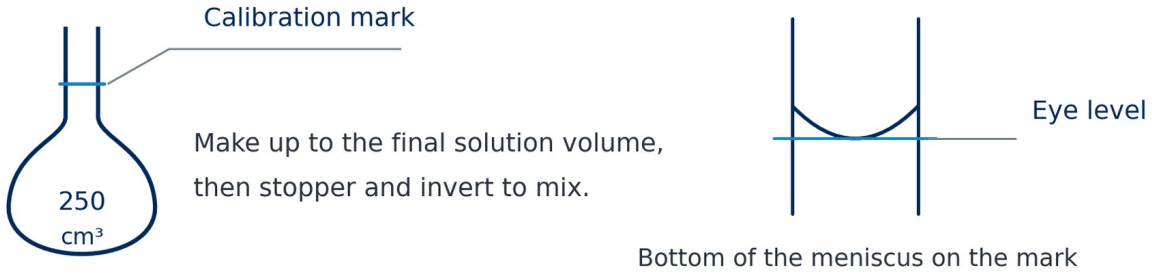
4. Rinse the beaker, rod and funnel into the flask. This transfers solute that would otherwise be left behind.

5. Let a warm solution cool to room temperature. Add water nearly to the mark, then dropwise until the bottom of the meniscus is on the mark at eye level.

6. Stopper and invert repeatedly, then label the solution.

KEY DEFINITION

Quantitative transfer means transferring all the measured solute. A volumetric flask is calibrated to contain its stated volume at the calibration temperature.



WORKED EXAMPLE

Plan a standard solution

Prepare 250.0 cm^3 of 0.0500 mol dm^{-3} Na_2CO_3. M = 106.0 g mol^{-1}.

Calculate the mass of anhydrous sodium carbonate needed.

Worked answer

n = cV
n = 0.0500 × 0.2500 = 0.01250 mol

m = n × M
m = 0.01250 × 106.0

Weigh 1.325 g Na_2CO_3 (subject to the balance precision).

EXAM TRAP Overshooting the mark makes the concentration too low. Removing some mixed solution does not restore the required concentration; prepare a fresh solution.

TEST IT

Try question 35 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Transfer a known amount quantitatively into a measured solution volume.

Acids & neutralisation

25. Acids, alkalis and strength

SPEC 2.1.4(a), (b)

CRACK IT

KEY DEFINITION

An acid releases $H^+(aq)$ in aqueous solution. An alkali is a water-soluble base that produces $OH^-(aq)$. Familiar examples are listed in the table.

KEY DEFINITION

Strength and concentration are different ideas. A strong acid ionises essentially completely in water; a weak acid ionises only partly. Concentration tells you how much acid is present per unit volume.

Common acid	Formula	Common alkali	Formula
Hydrochloric acid	HCl	Sodium hydroxide	NaOH
Sulfuric acid	H_2SO_4	Potassium hydroxide	KOH
Nitric acid	HNO_3	Ammonia	NH_3
Ethanoic acid	CH_3COOH	—	—

Diluting a strong acid makes it less concentrated, not weak. Use the ionisation equations below to see the difference between essentially complete and partial ionisation.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Term	Meaning
Strong acid	Essentially completely dissociated in aqueous solution.
Weak acid	Partially dissociated in aqueous solution.
Concentration	Amount of solute per unit solution volume.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

WORKED EXAMPLE

Equal concentrations, different strengths

Equal volumes of hydrochloric acid and ethanoic acid have the same concentration in mol dm^{-3}. Compare their initial hydrogen-ion concentrations and the amounts in moles of hydroxide ions needed for complete neutralisation. Explain both comparisons.

Worked answer

HCl is essentially completely dissociated; ethanoic acid is only partially dissociated, so HCl has a higher H^+ concentration initially. Both acids are monoprotic, and the samples contain the same total amount of acid.

They need the same amount of OH^- for complete neutralisation.

EXAM TRAP Strong does not mean concentrated. A dilute strong acid and a concentrated weak acid are both possible.

TEST IT

Try questions 36, 37 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Distinguish acid strength from concentration and identify aqueous ions.

26. Neutralisation and salts

SPEC 2.1.4(c)

CRACK IT

DEFINITIONS & KEY TERMS

Salt

An ionic compound formed when the replaceable hydrogen ions of an acid are replaced by metal ions or ammonium ions (NH_4^+).

KEY DEFINITION

Neutralisation is the reaction of an acid with a base. When $H^+(aq)$ reacts with $OH^-(aq)$, water forms. The full equation depends on the base: acid plus aqueous ammonia forms an ammonium salt without a water product.

The other ions form the salt. Use the acid to identify the negative ion, then balance charges with the positive ion from the base.

Different bases give different additional products. Metal oxides and hydroxides form salt and water; carbonates also produce carbon dioxide. Use the table and balanced equations to choose the correct mole ratio.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Reactants	Products
Acid + metal hydroxide	Salt + water
Acid + metal oxide	Salt + water
Acid + metal carbonate	Salt + water + carbon dioxide
Acid + ammonia	Ammonium salt

Acid	Salt family
Hydrochloric	Chloride
Sulfuric	Sulfate
Nitric	Nitrate
Ethanoic	Ethanoate

WORKED EXAMPLE

Choose the correct salt and balance

Write balanced equations, including state symbols, for (a) magnesium oxide with hydrochloric acid, (b) sodium hydroxide with sulfuric acid, (c) calcium carbonate with nitric acid and (d) aqueous ammonia with hydrochloric acid.

Worked answer

$2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

$2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

Check both the salt formula and every balanced coefficient.

EXAM TRAP Ammonia with an acid forms an ammonium salt without producing water. Do not assume every neutralisation has a 1 : 1 mole ratio.

TEST IT

Try question 38 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Write balanced equations for acid reactions and name the salt.

Titrations & practical skills

27. Titration apparatus and technique

SPEC 2.1.3(i); 2.1.4(d); PAG 2

CRACK IT

A titration finds the volume of one solution needed to react with a measured portion of another. The burette delivers the titrant; the conical flask holds the solution being analysed.

1. Rinse the burette with deionised water, then its own solution. Fill it and run solution through the tip to remove air bubbles. Remove the filling funnel.

2. Rinse the pipette with deionised water, then the solution it will measure. Use a pipette filler to transfer a fixed aliquot (a measured portion of the solution) to the conical flask.

3. Let a standard transfer pipette drain. Do not blow out the drop designed to remain in it.

4. Add a few drops of a suitable indicator and place the flask on a white tile.

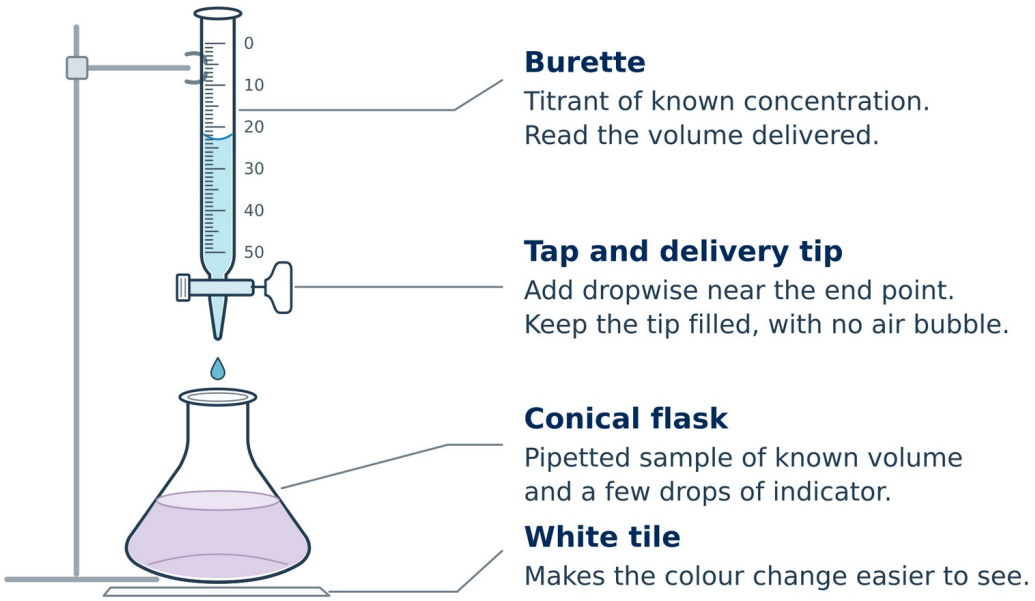
5. Read the burette meniscus at eye level and record the initial reading.

6. Do a rough titration to locate the endpoint. Repeat with fresh aliquots (measured portions), swirling and adding dropwise near the endpoint.

7. Wash splashes down the flask walls with deionised water. This changes volume but not the amount of solute being titrated.

8. Record the final reading and calculate final minus initial. Repeat until you obtain concordant titres (within 0.1 cm^3 of each other).

Record burette readings to two decimal places, normally to the nearest 0.05 cm^3 on a scale graduated every 0.10 cm^3. Use eye protection, a pipette filler and the laboratory risk assessment.



Titration	Suitable indicator
Strong acid + strong alkali	Methyl orange or phenolphthalein
Weak acid + strong alkali	Phenolphthalein
Strong acid + weak alkali	Methyl orange

WORKED EXAMPLE

A pipette was left wet with water

A pipette is left wet with distilled water before it is used to measure the analyte solution into a conical flask. Explain how this affects the number of moles transferred and the titre, compared with rinsing the pipette with the analyte solution first.

Worked answer

Residual water dilutes the solution inside the pipette. The measured aliquot contains fewer moles of analyte.

Less titrant is needed: the titre is too low.

EXAM TRAP The conical flask may be rinsed with deionised water: this changes volume but not the amount of analyte. Rinsing it with analyte would add unmeasured moles.

For the burette diagram, the delivered volume is the final scale reading minus the initial scale reading. A single scale reading is not necessarily the delivered titre.

TEST IT

Try questions 39, 40 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Obtain reliable titres using suitable apparatus, rinsing and an indicator.

28. Concordant titres and percentage uncertainty

SPEC 2.1.4(e); 1.1.3(c); 1.1.4(c)

CRACK IT

The titre is final burette reading minus initial reading. Select concordant titres (within 0.1 cm^3 of each other) before calculating a mean, and show which results you used.

Leave out the rough titre and justified non-concordant results. Use a different concordance tolerance if the question supplies one.

A titre depends on two burette readings, so add their absolute uncertainties in this treatment. Percentage uncertainty is absolute titre uncertainty divided by the titre, multiplied by 100.

For the same burette uncertainty, a larger titre gives a smaller percentage uncertainty. Use a different apparatus uncertainty if the question supplies one.

A **directional error** makes the result too high or too low. For example, consistently overshooting the end point gives too large a titre. When known acid is in the burette and unknown alkali is in the flask, this makes the calculated alkali concentration too high. Explain what changes, how the titre changes and how this changes the answer. Uncertainty alone does not tell you the direction of an error.

Titre = final reading − initial reading

WORKED EXAMPLE

Choose the readings to average

The following burette readings were recorded during a titration. All readings are in cm^3.

Reading / cm^3	Rough	1	2	3
Initial	0.00	0.10	0.20	0.00
Final	24.85	24.20	24.40	24.60

Calculate all four titres, select the concordant accurate titres and calculate their mean. Treat accurate titres that differ by no more than 0.10 cm^3 as concordant. Each burette reading has uncertainty ±0.05 cm^3. Calculate the absolute and percentage uncertainty in a titre equal to that mean.

Worked answer

Titration	Initial	Final	Titre
Rough	0.00	24.85	24.85
1	0.10	24.20	24.10
2	0.20	24.40	24.20
3	0.00	24.60	24.60

Mean of concordant titres (within 0.1 cm^3 of each other) 1 and 2 = (24.10 + 24.20)/2 = 24.15 cm^3

If each reading is ±0.05 cm^3: titre uncertainty = ±0.10 cm^3

Percentage uncertainty = (absolute uncertainty ÷ measured value) × 100
Percentage uncertainty = (0.10/24.15) × 100 = 0.414...%

Mean titre 24.15 cm^3; percentage uncertainty 0.41%.

EXAM TRAP A titre is the difference between two readings, so include both reading uncertainties. Repeating readings improves confidence but does not remove instrument resolution.

This example includes the two reading uncertainties only. Include any additional contribution specified in a question; do not assume that one uncertainty value applies to every burette or experiment.

TEST IT

Try question 41 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Select a consistent set of titres and account for both burette readings.

29. Titration calculations

SPEC 2.1.3(e)(iii), (g); 2.1.4(e)

CRACK IT

Use a grid with the reacting substances across the top and **c (concentration), V (volume) and n (amount in moles)** down the side. Add the coefficients from the balanced equation below the mole row.

1. Write the balanced equation and fill in the known concentrations and volumes. Convert each volume to dm^3 by dividing its cm^3 value by 1000.

2. Start in the column with both c and V known. Use n = cV to fill its mole row.

3. Work across the mole row using the equation ratio: divide by the known substance's coefficient, then multiply by the other substance's coefficient.

4. Work up the unknown substance's column: use c = n/V, dividing by that substance's **own** solution volume in dm^3.

5. Give the concentration unit and round only the final answer.

Keep substance names beside the numbers. The burette volume and pipette volume belong to different solutions and cannot be swapped.

WORKED EXAMPLE

Hydrochloric acid against carbonate

25.00 cm^3 of 0.0500 mol dm^{-3} Na_2CO_3 solution requires 23.60 cm^3 of HCl solution for complete neutralisation.

Write the balanced neutralisation equation and calculate the HCl concentration in mol dm^{-3}.

Worked answer

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$. Twice as many moles of HCl react as carbonate.

Quantity	Na_2CO_3	HCl
Concentration, c / mol dm^{-3}	0.0500	c = n/V 0.002500/0.02360 = **0.106**
Volume, V / dm^3	25.00/1000 = 0.02500	23.60/1000 = 0.02360
Amount, n / mol	n = cV 0.0500 × 0.02500 = 0.001250	0.001250 × 2 = 0.002500
Coefficient (mole ratio)	1	2

The HCl concentration is 0.106 mol dm^{-3} (3 significant figures).

EXAM TRAP Do not use the combined volume of the two solutions to find the original unknown concentration. Use the volume of that solution alone.

TEST IT

Try questions 42, 43 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Convert the known solution to moles and use the balanced reaction.

30. Aliquots and dilution factors

SPEC 2.1.3(e)(iii), (g); 2.1.4(e)

CRACK IT

KEY DEFINITION

An aliquot is a measured portion of a well-mixed solution. A titration of that portion gives the amount in the aliquot, not automatically the amount in the whole flask.

PERCENTAGE PURITY

When only the desired component reacts, calculate its amount and mass from the reaction. Percentage purity = (mass of pure component / mass of original sample) × 100. Use the whole-sample amount, not the aliquot amount.

Multiply the aliquot mole amount by whole-flask volume divided by aliquot volume to find the total. Use matching volume units in that ratio.

If the flask was made by diluting an original sample, those total solute moles came from that original sample. Divide by the original sample volume to recover its concentration. Label each stage so the dilution factor is applied only once.

WORKED EXAMPLE

Find the original acid concentration

10.00 cm^3 ethanoic acid is diluted to 250.0 cm^3. A 25.00 cm^3 aliquot of the diluted ethanoic acid solution requires 20.00 cm^3 of 0.100 mol dm^{-3} NaOH solution for complete neutralisation.

Calculate the ethanoic acid concentration in the original 10.00 cm^3 sample, in mol dm^{-3}. The reaction is $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$.

Worked answer

The aliquot is the measured 25.00 cm^3 portion of diluted acid. $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$.

Quantity	CH_3COOH (diluted)	NaOH
Concentration, c / mol dm^{-3}	c = n/V 0.00200/0.02500 = 0.0800	0.100
Volume, V / dm^3	25.00/1000 = 0.02500	20.00/1000 = 0.02000
Amount, n / mol	0.00200 (1 : 1 ratio)	n = cV 0.100 × 0.02000 = 0.00200
Coefficient (mole ratio)	1	1

Moles in the full flask = moles in aliquot × flask volume/aliquot volume = 0.00200 × 250.0/25.00 = 0.0200 mol.

Original concentration = total moles/original sample volume = 0.0200/0.01000 = **2.00 mol dm^{-3}**.

EXAM TRAP The flask scaling factor and the original sample volume do different jobs. Label each stage to avoid applying a dilution factor twice.

TEST IT

Try questions 44, 45 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Scale an aliquot amount to the whole solution before returning to the original sample.

31. Back titration and purity

SPEC 2.1.3(g); 2.1.4(e)

CRACK IT

A back titration measures what is left over. You add a known excess of reagent to the sample, then titrate the unreacted excess.

1. Calculate the amount in moles of reagent originally added.

2. Use the titration and its equation ratio to find the amount in moles of that reagent left over.

3. Subtract: amount that reacted with the sample = initial amount - amount remaining.

4. Use the sample-reaction equation to find the amount in moles of sample substance, then its mass or purity.

The two reactions may have different coefficient ratios. Apply each ratio at the stage where it belongs.

WORKED EXAMPLE

Purity of limestone

0.250 g limestone reacts with 25.00 cm^3 of 0.200 mol dm^{-3} HCl. All the hydrochloric acid remaining after the reaction is completely neutralised by 10.00 cm^3 of 0.100 mol dm^{-3} NaOH solution. M($CaCO_3$) = 100.1 g mol^{-1}.

Calculate the percentage by mass of $CaCO_3$ in the limestone. Assume the impurities do not react. The reactions are $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ and $HCl + NaOH \rightarrow NaCl + H_2O$.

Worked answer

The NaOH titration measures the remaining HCl. $HCl + NaOH \rightarrow NaCl + H_2O$.

Quantity	HCl added	NaOH titre	HCl left
Concentration, c / mol dm^{-3}	0.200	0.100	Not needed
Volume, V / dm^3	25.00/1000 = 0.02500	10.00/1000 = 0.01000	Not needed
Amount, n / mol	n = cV 0.200 × 0.02500 = 0.00500	n = cV 0.100 × 0.01000 = 0.00100	Same moles as NaOH 0.00100
Coefficient (mole ratio)	Not compared	1	1

Moles of HCl consumed = initial − remaining = 0.00500 − 0.00100 = 0.00400 mol.

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$. Moles of $CaCO_3$ = 0.00400/2 = 0.00200 mol (1 : 2 ratio).

Pure mass = n × M = 0.00200 × 100.1 = 0.2002 g. Percentage purity = pure mass/sample mass × 100 = 0.2002/0.250 × 100 = 80.08%, or **80.1%**.

EXAM TRAP The titration measures the acid left over. It does not directly give the amount in moles that reacted with the solid.

TEST IT

Try question 46 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Subtract the excess remaining from the reagent added before applying the sample ratio.

Gas calculations

32. Molar gas volume at RTP

SPEC 2.1.3(a)(v), (e)(ii), (g)

CRACK IT

DEFINITIONS & KEY TERMS

Molar gas volume, V_m

The gas volume per mole, V/n, at a stated temperature and pressure.

At room temperature and pressure, use a molar gas volume of **24.0 dm³ mol⁻¹**, unless the question supplies a different value.

To find moles, divide the gas volume by the molar gas volume. To find volume, multiply moles by it. Make the units match first: 24.0 dm³ is 24,000 cm³.

At the same temperature and pressure, equal volumes of ideal gases contain equal mole amounts. You can therefore use balanced coefficients as volume ratios when all the substances being compared are gases under those conditions.

$V = nV_m$; $n = V/V_m$; $V_m = 24.0\ dm^3\ mol^{-1} = 24000\ cm^3\ mol^{-1}$ at RTP

V gas volume; match its unit to V_m (dm³ or cm³); **V_m** molar gas volume: 24.0 dm³ mol⁻¹ or 24,000 cm³ mol⁻¹ at the stated room temperature and pressure; **n** amount of the specified substance or particles, in mol.

WORKED EXAMPLE

Hydrogen from magnesium

0.486 g Mg reacts with excess HCl at RTP. M(Mg) = 24.3 g mol^{-1}.

Calculate the volume of hydrogen formed, in dm³ and cm³. Use a molar gas volume of 24.0 dm³ mol⁻¹ at RTP.

Worked answer

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

Quantity	Mg	H_2
Mass / g	0.486	Not needed
M / g mol^{-1}	24.3	Not needed
Amount / mol	n = m/M 0.486/24.3 = 0.0200	0.0200 (1 : 1)

$V(H_2) = nV_m$

$V(H_2) = 0.0200 \times 24.0$

0.480 dm^3, or 480 cm^3, of hydrogen at RTP.

EXAM TRAP Use 24.0 only at the stated RTP conditions. For other temperatures or pressures, use the ideal gas equation when data are supplied.

TEST IT

Try questions 47, 48 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Use 24.0 dm^3 per mole at room temperature and pressure.

33. The ideal gas equation

SPEC 2.1.3(f)

CRACK IT

The ideal gas equation connects pressure, volume, temperature and amount. Write pV = nRT, then rearrange for the quantity you need.

1. Use pressure in Pa, volume in m^3 and temperature in K when using R = 8.314 J mol^{-1} K^{-1}.

2. Convert before substituting: add 273 to temperature in °C for these school calculations, and use 1 dm^3 = 10^{-3} m^3 or 1 cm^3 = 10^{-6} m^3.

3. Calculate the amount in moles. If mass is supplied, use M = m/n to find molar mass.

The ideal-gas model is a better approximation at low pressure and high temperature. The particles then occupy a smaller fraction of the volume, and attractions have less effect.

pV = nRT; n = pV/(RT)

R gas constant, 8.314 J mol^{-1} K^{-1}, unless the question supplies a different value; **T** absolute temperature of the gas, in K; **V** gas volume, in m^3 in the ideal gas equation; **n** amount of the specified substance or particles, in mol; **p** gas pressure, in Pa in the ideal gas equation.

Quantity	Unit	Conversion
p	Pa	kPa × 1000
V	m^3	dm^3 ÷ 1000; cm^3 ÷ 1000000
n	mol	Use amount, not mass
T	K	°C + 273 (or +273.15 when precision requires)
R	J mol^{-1} K^{-1}	8.314

WORKED EXAMPLE

Find an amount in moles of gas

100 cm^3 of gas is measured at 101 kPa and 27 °C. Use R = 8.314 J mol^{-1} K^{-1}.

Calculate the number of moles of gas.

Worked answer

p = 101000 Pa; V = 0.000100 m^3; T = 300 K

n = pV/(RT)
n = (101000 × 0.000100)/(8.314 × 300)

n = 4.05 ×10^{-3} mol.

EXAM TRAP A temperature in °C cannot be substituted directly. A volume in dm^3 also needs conversion when R = 8.314 in SI units.

TEST IT

Try question 49 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Convert all quantities into compatible SI units before using pV = nRT.

34. Gas measurements and molar mass

SPEC 2.1.3(e)(ii), (f), (i); PAG 1

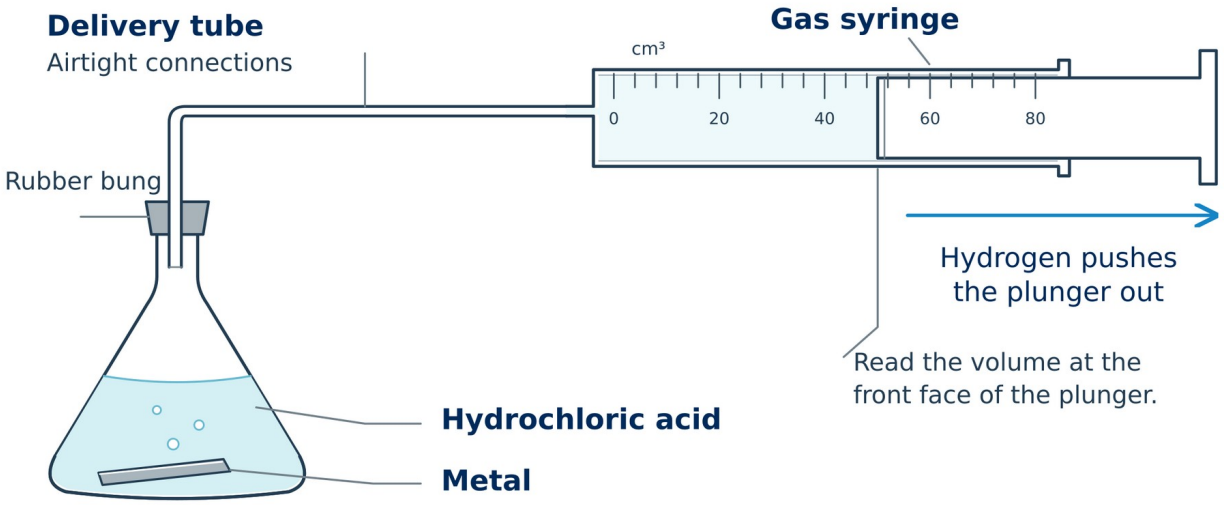
CRACK IT

A gas calculation is only as good as the volume measurement. Use a gas-tight apparatus with a syringe plunger that moves freely, and begin collection promptly so gas is not lost.

Measure temperature and pressure consistently. Collection over water is suitable only if the gas is sufficiently insoluble and does not react with the water.

To find molar mass, first calculate moles from the gas data, then use M = m/n.

If gas escapes, measured volume is too small and the calculated mole amount is too small. With the gas mass used in the calculation unchanged, dividing by that smaller n makes the calculated M too high.



WORKED EXAMPLE

Molar mass of a gas

0.184 g gas occupies 100 cm^3 at 100 kPa and 27 °C. Use R = 8.314 J mol^{-1} K^{-1}.

Calculate the molar mass of the gas in g mol^{-1}.

Worked answer

n = pV/(RT)
n = (100000 × 0.000100)/(8.314 × 300) = 0.00400946... mol

M = m/n = 0.184/0.00400946...

M = 45.9 g mol^{-1}.

EXAM TRAP Match the measured gas volume to the stated temperature and pressure. Do not substitute an RTP molar volume into a non-RTP measurement.

TEST IT

Try question 50 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Combine gas amount with mass and evaluate the apparatus.

Yield & atom economy

35. Percentage yield

SPEC 2.1.3(h)(i), (g)

CRACK IT

KEY DEFINITION

The theoretical yield is the maximum amount of desired product predicted by the balanced equation. The actual yield is what you recover from the experiment.

1. Calculate the theoretical yield from the limiting reactant.

2. Divide actual yield by theoretical yield and multiply by 100.

3. Compare matching quantities: mass with mass, amount with amount, or suitable gas volumes measured under matching conditions.

A reaction may be incomplete, competing reactions may form other products, or product may be lost during transfer and purification.

A result above 100% suggests an apparent product mass that is too high, for example because the product is wet or impure. It does not mean more pure product formed than the theoretical maximum.

Percentage yield = (actual yield/theoretical yield) × 100

WORKED EXAMPLE

Calcium oxide yield

10.0 g $CaCO_3$ is heated and 4.48 g CaO is collected. M($CaCO_3$) = 100.1; M(CaO) = 56.1 g mol^{-1}.

Write the balanced decomposition equation and calculate the percentage yield of calcium oxide.

Worked answer

$CaCO_3 \rightarrow CaO + CO_2$

Quantity	$CaCO_3$	CaO theoretical
Mass / g	10.0	5.60440...
M / g mol^{-1}	100.1	56.1
Amount / mol	n = m/M 10.0/100.1 = 0.0999001...	n = m/M n = 5.60440.../56.1 = 0.0999001...

Percentage yield = (actual yield ÷ theoretical yield) × 100
Percentage yield = (4.48/5.60440...) × 100

79.9%.

For a reverse calculation: theoretical yield = actual yield/(percentage yield/100)

EXAM TRAP Actual product mass divided by starting reactant mass is not percentage yield. Calculate the theoretical product first.

TEST IT

Try questions 51, 52 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Compare actual and theoretical amounts of the same product.

36. Atom economy and sustainable processes

SPEC 2.1.3(h)(ii), (j)

CRACK IT

Atom economy asks how much of the reactant mass is directed into the **chosen desired product** by the balanced equation. Percentage yield asks how much of that product is actually recovered.

Use the formula-mass contribution of the desired product, including its coefficient, divided by the total reactant formula-mass contributions, also including coefficients. Multiply by 100.

Conservation of mass means the denominator can also be the total formula-mass contribution of all products. Do not include catalysts, solvents or arbitrary excess quantities.

A high atom economy can reduce waste, but choosing a sustainable process also needs energy use, hazards, raw-material cost and supply, renewability, rate, yield, separation and recycling to be considered. A useful by-product may also affect the choice.

Atom economy = [coefficient × M_r of desired product / sum(coefficient × M_r of reactants)] × 100

M_r relative molecular or formula mass; no unit.

WORKED EXAMPLE

Reduction of copper(II) oxide

$2CuO + C \rightarrow 2Cu + CO_2$. Copper is the desired product. M(CuO) = 79.5, A_r(C) = 12.0 and A_r(Cu) = 63.5.

Calculate the percentage atom economy for producing copper.

Worked answer

Desired product contribution = 2 × 63.5 = 127.0

Total reactant contribution = 2 × 79.5 + 12.0 = 171.0

Atom economy = (total relative mass of desired product ÷ total relative mass of reactants) × 100
Atom economy = (127.0/171.0) × 100

74.3%.

EXAM TRAP Do not use experimental product masses in atom economy. Include coefficients and every reactant in its denominator.

TEST IT

Try question 53 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Use the balanced equation to assess useful product and waste.