IGCSE CHEMISTRY 0620 · CHAPTER 3 NOTES · ~19 MIN READ

Stoichiometry — formulae & moles

Formulae, balanced equations and mole calculations, one step at a time — including limiting reactants, titration, empirical formulae and percentage yield.

Formulae | Balancing equations | Mr | The mole | Reacting masses | Gas volumes | Concentration | Empirical formula

RS by Riyaz Sir
mentorriyaz.in

No tag = **Core** (everyone) EXTENDED Extended only highlight = comes up again and again

Test yourself at the end

IN 30 SECONDS

- A **formula** shows which atoms are in a substance. An **equation** shows what changes in a reaction.
- **Balanced** = the same number of each atom on both sides. Change the **numbers in front**, never the small numbers.
- **Mr** = add up the Ar values from the Periodic Table.
- EXT **Moles:** n = mass ÷ Mr for solids, n = V ÷ 24 for gases (at r.t.p.), n = c × V for solutions.

BY THE END YOU CAN

- Write formulae and balanced equations with state symbols.
- Calculate relative formula mass and reacting masses.
- Use moles to link mass, gas volume and concentration. EXT
- Find the limiting reactant, empirical and molecular formulae, percentage yield and purity. EXT

1 Formulae

What a formula tells you

- A formula uses **chemical symbols** and small numbers (**subscripts**). The subscript shows how many atoms of the element **before** it: H_2O has 2 hydrogen atoms and 1 oxygen atom; CO_2 has 1 carbon and 2 oxygen.
- If a group of atoms is repeated, it goes in **brackets**: $Ca(OH)_2$ has 1 Ca, 2 O and 2 H. $(NH_4)_2SO_4$ has 2 N, 8 H, 1 S and 4 O.
- A **molecular formula** shows the actual number of each atom in a molecule: glucose $C_6H_{12}O_6$.

Substance	Formula	Substance	Formula
water	H_2O	sodium chloride	NaCl
carbon dioxide	CO_2	magnesium oxide	MgO
ammonia	NH_3	calcium carbonate	$CaCO_3$
methane	CH_4	sodium hydroxide	NaOH
hydrogen chloride	HCl	sulfuric acid	H_2SO_4
oxygen	O_2	nitric acid	HNO_3

- Name endings: **-ide** = two elements only (sodium chloride, magnesium oxide). **-ate** = contains oxygen too (sulfate, nitrate, carbonate).

Ionic formulae from charges

EXTENDED

- An ionic compound must have **zero** total charge. Balance the positive and negative charges.

Cations	Anions
Na^+ K^+ Li^+ Ag^+ NH_4^+	Cl^- Br^- I^- OH^- NO_3^-
Mg^{2+} Ca^{2+} Zn^{2+} Cu^{2+} Fe^{2+}	O^{2-} S^{2-} SO_4^{2-} CO_3^{2-}
Al^{3+} Fe^{3+}	–

- **Cross-over trick:** write the charges, swap the numbers (ignore the signs) and make them the subscripts.

WORKED EXAMPLE

EXT **Write the formula of aluminium sulfate.**

1. Ions: Al^{3+} and SO_4^{2-}
2. Swap the numbers: Al gets 2, SO_4 gets 3
3. Put the sulfate in brackets because there are 3 of them: **$Al_2(SO_4)_3$**
4. **Check the charge: 2 × (+3) = +6 and 3 × (−2) = −6 → total 0**

WORKED EXAMPLE

EXT **Write the formula of calcium hydroxide.**

1. Ca^{2+} and OH^-
2. One Ca^{2+} needs two OH^- to balance
3. **OH is a group, so use brackets: $Ca(OH)_2$**

COMMON MISTAKE

Putting brackets round a single ion ($CaCl_2$ is correct, not $Ca(Cl)_2$) — and writing $Ca(OH)_2$ as $CaOH_2$.

State symbols

Symbol	Meaning
(s)	solid
(l)	liquid
(g)	gas
(aq)	aqueous: dissolved in water

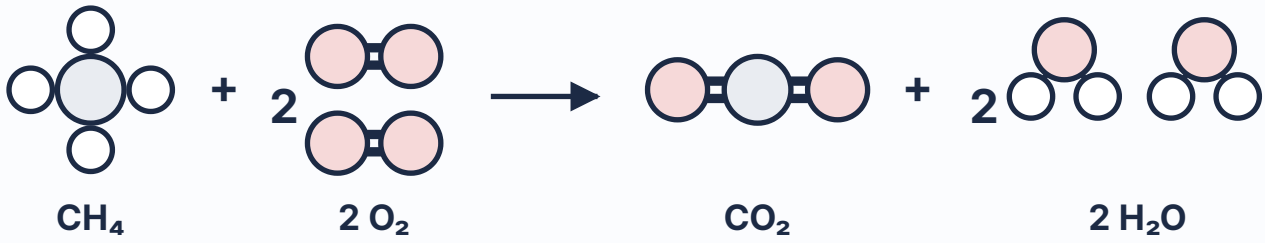
- **aqueous** means dissolved in water. Pure water itself is (l).

2 Equations

Word equations and symbol equations

- A **word equation** shows the names. A **symbol equation** uses the formulae and shows how many of each particle react.
- Example: magnesium + oxygen → magnesium oxide. In symbols: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Balancing: the method



CH_4 $2\ O_2$ CO_2 $2\ H_2O$

count each atom on both sides
C: 1 = 1 · H: 4 = 4 · O: 4 = 4 ✓ balanced

change only the big numbers in front, never the small numbers inside a formula

Burning methane: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. Atoms are never created or destroyed, so both sides must have the same number of each atom.

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- **Step 1:** write the correct formulae for everything. They can never be changed afterwards.
- **Step 2:** count the atoms of each element on each side.
- **Step 3:** add **big numbers (coefficients) in front** of formulae to balance. Do one element at a time. Leave the atoms that appear in several substances (usually O and H) until last.
- **Step 4:** recount both sides. Add the state symbols.

COMMON MISTAKE

Changing a small number to make it balance. $H_2O \rightarrow H_2O_2$ is a different substance (hydrogen peroxide). Change only the **large numbers in front**.

WORKED EXAMPLE

Balance: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

1. C: 1 on each side — balanced
2. H: 4 on the left, 2 on the right → put 2 in front of H_2O
3. O: now 2 + 2 = 4 on the right, so put 2 in front of O_2
4. **$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$**

WORKED EXAMPLE

Balance: $Fe_2O_3 + CO \rightarrow Fe + CO_2$

1. Fe: 2 on the left → put 2 in front of Fe
2. Fe_2O_3 gives away 3 oxygen atoms, and each CO can take only 1 → use 3 CO, which makes 3 CO_2
3. check C: 3 = 3
4. check O: 3 + 3 = 6 on the left, and 3 × 2 = 6 on the right
5. **$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$**

WORKED EXAMPLE

Balance: $Al + HCl \rightarrow AlCl_3 + H_2$

1. $AlCl_3$ has 3 Cl, so start with 3 HCl
2. that gives 3 H, but H_2 comes in pairs — double everything so far: 2 $AlCl_3$, 6 HCl, 3 H_2
3. 2 $AlCl_3$ contains 2 Al, so use 2 Al on the left
4. check Al: 2 = 2, Cl: 6 = 6, H: 6 = 6
5. **$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$**

WORKED EXAMPLE

Balance: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

1. C: 3 on the left → 3 CO_2
2. H: 8 on the left → 4 H_2O
3. O on the right: 3 × 2 + 4 = 10, so 5 O_2 on the left
4. **$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$**

Ionic equations

EXTENDED

- In an **ionic equation** you show only the ions that actually change. The ions that are the same on both sides are **spectator ions** and are left out.

WORKED EXAMPLE

EXT **Write the ionic equation for $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$.**

1. Write everything that is dissolved as ions: $Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$
2. Na^+ and NO_3^- are the same on both sides → they are the spectator ions
3. **$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$**

- The solid (precipitate) stays as a formula — it is **not** split into ions.
- Neutralisation: **$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$** is the ionic equation for any acid and alkali reacting.

EXAM TIP

Check **charges** balance as well as atoms in an ionic equation: $Ag^+ + Cl^-$ gives 0 on the left, and AgCl is neutral on the right.

3 Relative masses

Ar and Mr

- The **relative atomic mass (A_r)** is the mass of an atom compared with carbon-12. You can read it from the Periodic Table: H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, S = 32, Cl = 35.5, Ca = 40, Fe = 56, Cu = 64.
- EXT Formal definition: the average mass of the naturally occurring atoms of an element, on a scale where an atom of carbon-12 has a mass of exactly 12.

Relative formula mass (M_r) — the sum of the relative atomic masses of all the atoms in the formula. It has **no units**.

WORKED EXAMPLE

Calculate the Mr of calcium carbonate, $CaCO_3$.

1. Ca = 40, C = 12, O = 16 × 3 = 48
2. **40 + 12 + 48 = 100**

WORKED EXAMPLE

Calculate the Mr of $(NH_4)_2SO_4$.

1. N: 2 × 14 = 28
2. H: 8 × 1 = 8
3. S: 32
4. O: 4 × 16 = 64
5. **28 + 8 + 32 + 64 = 132**

Reacting masses in simple proportions

- A balanced equation shows the ratio of particles. Convert to masses using Ar or Mr, then use **ratios**.

WORKED EXAMPLE

What mass of magnesium oxide forms when 6.0 g of magnesium burns completely in oxygen?

1. $2Mg + O_2 \rightarrow 2MgO$
2. Masses: 2 × 24 = 48 g of Mg makes 2 × 40 = 80 g of MgO
3. 6.0 g is 6.0 ÷ 48 = 0.125 of the amount
4. **0.125 × 80 = 10 g**

EXAM TIP

Set out the three lines: **equation → mass ratio → scale to the amount in the question**. Even if you get stuck, you will pick up method marks.

4 The mole EXT

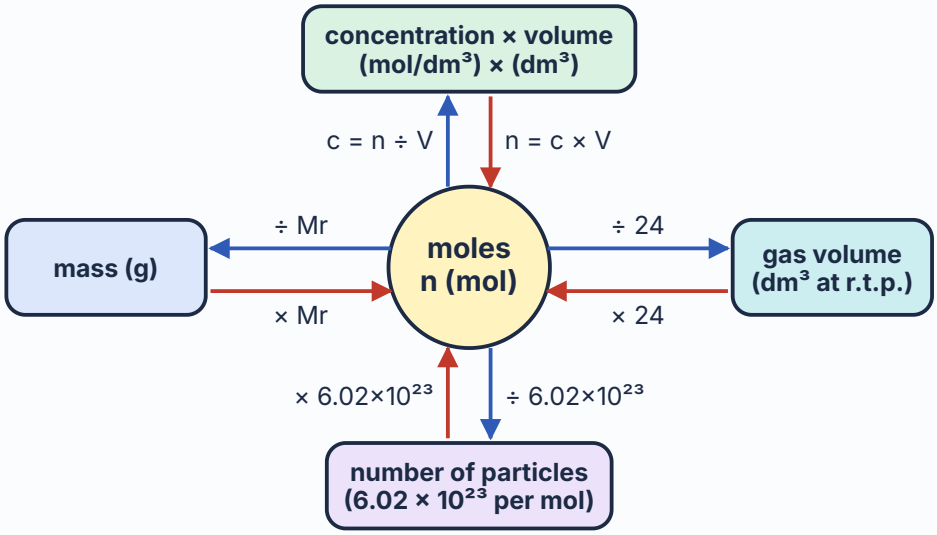
Counting by weighing

Mole — the amount of substance that contains 6.02×10^{23} particles (atoms, molecules or ions). This number is the **Avogadro constant**.

PICTURE IT

A **dozen** is 12 of anything. A **mole** is 6.02×10^{23} of anything. You cannot count atoms, but you can weigh them: **one mole of any substance weighs its Mr in grams.**

- 1 mol of carbon (Ar = 12) weighs **12 g**. 1 mol of water (Mr = 18) weighs **18 g**. 1 mol of CO_2 (Mr = 44) weighs **44 g**.



The mole is the central quantity. From moles you can reach mass, gas volume, number of particles and solution concentration.

Four equations

$$n = \frac{m}{M_r}$$

moles (mol) = mass (g) ÷ molar mass (g/mol)

$$V = n \times 24$$

gas volume (dm^3) = moles × 24 dm^3/mol (at room temperature and pressure, r.t.p.)

$$c = \frac{n}{V}$$

concentration (mol/dm^3) = moles ÷ volume (dm^3)

$$N = n \times 6.02 \times 10^{23}$$

number of particles = moles × Avogadro constant

- **Units:** 1 dm^3 = 1000 cm^3. To change cm^3 to dm^3 divide by 1000.
- Concentration in **g/dm^3** = mass of solute (g) ÷ volume (dm^3). To convert between the two: **g/dm^3 = mol/dm^3 × Mr**.

WORKED EXAMPLE

EXT **How many moles and molecules are in 11 g of carbon dioxide? (Mr = 44)**

1. n = 11 ÷ 44 = **0.25 mol**
2. molecules = 0.25 × 6.02 × 10^{23}
3. **= 1.5 × 10^{23} (1.505 × 10^{23})**

WORKED EXAMPLE

EXT **What volume of hydrogen at r.t.p. forms when 4.8 g of magnesium reacts with excess hydrochloric acid?**

1. $Mg + 2HCl \rightarrow MgCl_2 + H_2$
2. n(Mg) = 4.8 ÷ 24 = 0.20 mol
3. mole ratio Mg : H_2 = 1 : 1 → $n(H_2)$ = 0.20 mol
4. **V = 0.20 × 24 = 4.8 dm^3**

WORKED EXAMPLE

EXT **4.0 g of sodium hydroxide is dissolved to make 500 cm^3 of solution. Find the concentration in mol/dm^3 and g/dm^3.**

1. Mr(NaOH) = 23 + 16 + 1 = 40
2. n = 4.0 ÷ 40 = 0.10 mol
3. V = 500 ÷ 1000 = 0.500 dm^3
4. c = 0.10 ÷ 0.500 = **0.20 mol/dm^3**
5. **g/dm^3 = 4.0 ÷ 0.500 = 8.0 g/dm^3 (check: 0.20 × 40 = 8.0 ✓)**

The standard route

- **Step 1:** write the balanced equation.
- **Step 2:** convert what you are given into **moles** (n = m ÷ Mr, or n = V ÷ 24, or n = c × V).
- **Step 3:** use the **mole ratio** from the equation.
- **Step 4:** convert moles into what is asked for (mass, volume, concentration).

COMMON MISTAKE

Using cm^3 instead of dm^3. Always divide by 1000 first.

COMMON MISTAKE

Forgetting the mole ratio. In $2H_2 + O_2 \rightarrow 2H_2O$, 1 mol of O_2 makes **2** mol of water, not 1.

Titration calculations

- In a titration you find the unknown concentration of one solution by reacting a **known volume** of it with a solution of **known concentration**. (The practical method is in experimental techniques and acids, bases & salts.)

WORKED EXAMPLE

EXT **25.0 cm^3 of sodium hydroxide is neutralised by 20.0 cm^3 of 0.100 mol/dm^3 hydrochloric acid. Find the concentration of the sodium hydroxide.**

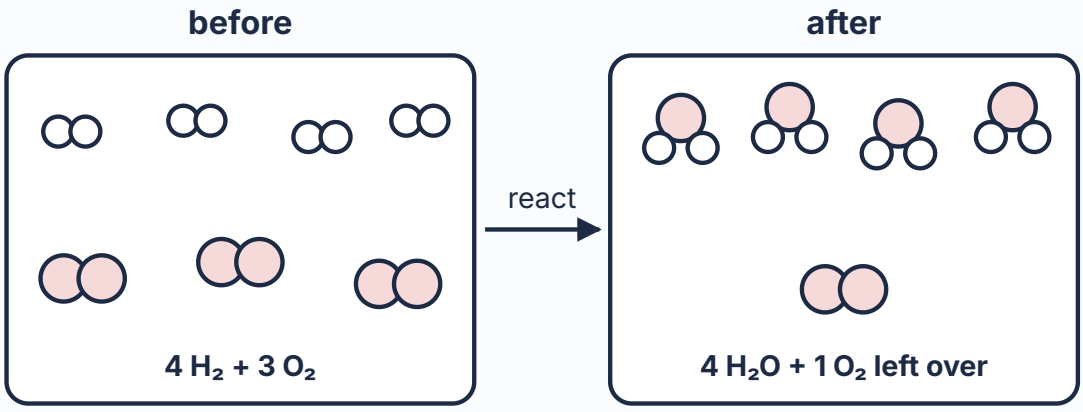
1. HCl + NaOH → NaCl + H_2O (ratio 1 : 1)
2. n(HCl) = 0.100 × 20.0 ÷ 1000 = 0.00200 mol
3. n(NaOH) = 0.00200 mol
4. **c(NaOH) = 0.00200 ÷ 0.0250 = 0.0800 mol/dm^3**

WORKED EXAMPLE

EXT **20.0 cm^3 of 0.50 mol/dm^3 sodium hydroxide neutralises 25.0 cm^3 of sulfuric acid. Find the concentration of the acid.**

1. H_2SO_4 + 2NaOH → Na_2SO_4 + $2H_2O$ (ratio 1 : 2)
2. n(NaOH) = 0.50 × 20.0 ÷ 1000 = 0.0100 mol
3. n(H_2SO_4) = 0.0100 ÷ 2 = 0.00500 mol
4. **c(H_2SO_4) = 0.00500 ÷ 0.0250 = 0.20 mol/dm^3**

Limiting reactant



H_2 is the limiting reactant: it is all used up and decides the amount of product
O_2 is in excess: some is left over

Hydrogen and oxygen. There are 4 H_2 but only 3 O_2. The H_2 runs out first, so 1 O_2 is left over.

Limiting reactant — the reactant that is used up first. It decides how much product forms. The other reactant is in **excess**.

WORKED EXAMPLE

EXT **5.6 g of iron is heated with 6.4 g of sulfur: Fe + S → FeS. Which is limiting, and what mass of FeS forms?**

1. n(Fe) = 5.6 ÷ 56 = 0.10 mol
2. n(S) = 6.4 ÷ 32 = 0.20 mol
3. the ratio is 1 : 1, so iron (0.10 mol) runs out first → **iron is limiting**
4. n(FeS) = 0.10 mol, Mr = 88 → mass = 0.10 × 88 = **8.8 g**
5. **sulfur left over = 0.10 mol × 32 = 3.2 g**

EXAM TIP

Compare **moles divided by the coefficient** — the smaller one is the limiting reactant. Always use the limiting reactant for the calculation.

PICTURE IT

You have 4 slices of bread and 3 slices of cheese and make cheese sandwiches (2 bread + 1 cheese each). Bread limits you to 2 sandwiches; cheese would allow 3. The bread runs out first and 1 slice of cheese is left over.

5 Formulae from data and yields EXT

Empirical and molecular formula

Empirical formula — the simplest whole-number ratio of atoms of each element in a compound.

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

- **Method:** (1) write the mass or % of each element; (2) divide each by its Ar to get moles; (3) divide all by the **smallest** number; (4) turn the ratio into whole numbers.

WORKED EXAMPLE

EXT **A hydrocarbon is 85.7% carbon and 14.3% hydrogen by mass. Mr = 56. Find the empirical and molecular formulae.**

1. C: 85.7 ÷ 12 = 7.14
2. H: 14.3 ÷ 1 = 14.3
3. divide by 7.14: C = 1, H = 2 → empirical formula **CH_2** (mass 14)
4. **56 ÷ 14 = 4, so multiply by 4 → molecular formula C_4H_8**

WORKED EXAMPLE

EXT **An iron oxide contains 69.9% iron and 30.1% oxygen. Find its formula.**

1. Fe: 69.9 ÷ 56 = 1.248
2. O: 30.1 ÷ 16 = 1.881
3. divide by 1.248: Fe = 1, O = 1.507
4. that is 1 : 1.5, so double both: 2 : 3
5. **Fe_2O_3**

EXAM TIP

If your ratio is 1 : 1.5 multiply by **2**; 1 : 1.33 → **3**; 1 : 1.25 → **4**. Do not round 1.5 down to 1 or up to 2.

Percentage yield, composition and purity

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

the actual yield is what you really got; the theoretical yield is the maximum from the equation

$$\%\ \text{purity} = \frac{\text{mass of pure substance}}{\text{mass of sample}} \times 100$$

how much of the sample is the wanted substance

$$\%\ \text{of an element} = \frac{A_r \times \text{number of atoms}}{M_r} \times 100$$

percentage composition by mass

WORKED EXAMPLE

EXT **10.0 g of calcium carbonate is heated: $CaCO_3 \rightarrow CaO + CO_2$. 4.2 g of calcium oxide is obtained. Find the % yield.**

1. $n(CaCO_3) = 10.0 \div 100 = 0.100$ mol
2. $n(CaO) = 0.100$ mol, $Mr(CaO) = 56$
3. theoretical yield = $0.100 \times 56 = 5.6$ g
4. **% yield = $4.2 \div 5.6 \times 100 = 75\%$**

WORKED EXAMPLE

EXT **5.00 g of impure calcium carbonate reacts with excess acid and gives 960 cm^3 of carbon dioxide at r.t.p. Find the % purity.**

1. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$
2. $n(CO_2) = 0.960 \div 24 = 0.0400$ mol
3. $n(CaCO_3) = 0.0400$ mol → mass = $0.0400 \times 100 = 4.00$ g
4. **% purity = $4.00 \div 5.00 \times 100 = 80\%$**

WORKED EXAMPLE

EXT **Find the percentage of nitrogen in ammonium nitrate, NH_4NO_3.**

1. Mr = 14 + 4 + 14 + 48 = 80
2. two N atoms: $2 \times 14 = 28$
3. **% N = $28 \div 80 \times 100 = 35\%$**

COMMON MISTAKE

A % yield above 100%. That means the product is wet or impure, or the calculation is wrong.

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6 Test yourself

Test yourself — cover the answers first

QUICK CHECK

1. How many atoms are in one formula unit of $Ca(OH)_2$?
2. Balance: $Na + H_2O \rightarrow NaOH + H_2$
3. Find the Mr of H_2SO_4.
4. What are the four state symbols?
5. EXT How many moles are in 36 g of water?
6. EXT What volume do 0.50 mol of any gas occupy at r.t.p.?
7. EXT Write the formula of magnesium nitrate.

EXAM-STYLE

8. Calculate the mass of carbon dioxide formed when 5.0 g of calcium carbonate reacts with excess acid. [3]
9. EXT 2.4 g of magnesium reacts with 50.0 cm^3 of 2.0 mol/dm^3 hydrochloric acid: $Mg + 2HCl \rightarrow MgCl_2 + H_2$. Which is limiting? What volume of H_2 forms at r.t.p.? [4]
10. EXT A compound contains 40.0% C, 6.7% H and 53.3% O by mass, and has Mr = 180. Find the molecular formula. [4]
11. EXT 25.0 cm^3 of 0.200 mol/dm^3 NaOH is titrated with HCl. 20.0 cm^3 of acid is needed. Find the HCl concentration. [3]

Show answers

1. 1 + 2 + 2 = 5 atoms (1 Ca, 2 O, 2 H)
2. $2Na + 2H_2O \rightarrow 2NaOH + H_2$
3. 2 + 32 + 64 = 98
4. (s) solid, (l) liquid, (g) gas, (aq) aqueous
5. 36 ÷ 18 = 2.0 mol
6. 0.50 × 24 = 12 dm^3
7. Mg^{2+} and NO_3^- → $Mg(NO_3)_2$
8. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$
   $n(CaCO_3) = 5.0 \div 100 = 0.050$ mol → $n(CO_2) = 0.050$ mol
   mass = 0.050 × 44 = 2.2 g
9. n(Mg) = 2.4 ÷ 24 = 0.10 mol
   n(HCl) = 2.0 × 0.0500 = 0.100 mol
   HCl needs 2 per Mg: 0.100 mol HCl reacts with only 0.050 mol Mg → HCl is limiting
   $n(H_2) = 0.100 \div 2 = 0.050$ mol → V = 0.050 × 24 = 1.2 dm^3
10. C: 40.0 ÷ 12 = 3.33; H: 6.7 ÷ 1 = 6.7; O: 53.3 ÷ 16 = 3.33
    ratio 1 : 2 : 1 → CH_2O (mass 30)
    180 ÷ 30 = 6
    $C_6H_{12}O_6$
11. n(NaOH) = 0.200 × 0.0250 = 0.00500 mol
    1 : 1 ratio, so n(HCl) = 0.00500 mol
    c = 0.00500 ÷ 0.0200 = 0.250 mol/dm^3

Stuck on a question?

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