

Exercise walk-through

2.3

Formulas

A-Level Chemistry

Formulas and equations

empirical formula

simplest ratio – CH_3

molecular formula

actual atoms – C_2H_6

Equations are **balanced** with **state symbols** (s, l, g, aq). An **ionic equation** drops **spectator ions** : $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$.

Formulas and equations, in detail

element	% by mass	$\div A_r \rightarrow$ moles	smallest (3.33)	ratio
C	40.0	$40.0/12 = 3.33$	1.0	1
H	6.7	$6.7/1 = 6.7$	2.0	2
O	53.3	$53.3/16 = 3.33$	1.0	1

empirical formula CH_2O - if $M_r = 180$, it fits $180/30 = 6$ times: molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$

The empirical-formula recipe: divide by A_r , divide by the smallest, read off the ratio

Hydrated and anhydrous solids

Some solids hold water inside their crystals.

This water is the **water of crystallisation** .

Reacting masses and yield

Worked example: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

$4.8 \text{ g Mg} = 0.20 \text{ mol}$; ratio $1:1 \Rightarrow 0.20 \text{ mol MgO} = 0.20 \times 40 = 8.0 \text{ g}$

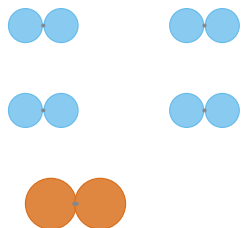
$$\text{percentage yield} = \frac{\text{actual product}}{\text{maximum possible}} \times 100\%$$

The numbers in a balanced equation give the mole ratio – the **stoichiometry** .

Limiting reagent



start with



O₂ runs out first:
limiting reagent

you get



2 H₂O made

2 H₂ left over (excess)

The **limiting reagent** runs out first and decides how much product forms.

The leftover reactant is in **excess**. Always base the calculation on the limiting reagent.

Exercises 2.1·2.2

2.1 Write the formula of (a) calcium nitrate and (b) aluminium sulfate. [2]

2.2 Write the ionic equation for the reaction of silver ions with chloride ions to form silver chloride. [1]

Exercises 2.1·2.2

2.1 Write the formula of (a) calcium nitrate and (b) aluminium sulfate. [2]

2.1

(a)

- $\text{Ca}(\text{NO}_3)_2$

(b)

- $\text{Al}_2(\text{SO}_4)_3$

2.2 Write the ionic equation for the reaction of silver ions with chloride ions to form silver chloride. [1]

2.2

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

Exercise 3.1

3.1 What is the formula of iron(III) sulfate?

[1]

☐

A FeSO_4

B Fe_2SO_4

C $\text{Fe}_2(\text{SO}_4)_3$

D $\text{Fe}_3(\text{SO}_4)_2$

Exercise 3.1

3.1 What is the formula of iron(III) sulfate?

[1]

☐

A FeSO_4

B Fe_2SO_4

C $\text{Fe}_2(\text{SO}_4)_3$

D $\text{Fe}_3(\text{SO}_4)_2$

3.1 C

- Two Fe^{3+} balance three SO_4^{2-}
- $\text{Fe}_2(\text{SO}_4)_3$

Exercises 3.2·3.3

3.2 A compound has empirical formula CH_2 and $M_r = 42$. Its molecular formula is: [1]

☐

A CH_2

B C_2H_4

C C_3H_6

D C_4H_8

3.3 Balance this equation and add state symbols: $\text{Mg} + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. [2]

Exercises 3.2-3.3

3.2 A compound has empirical formula CH_2 and $M_r = 42$. Its molecular formula is: [1]

☐

A CH_2

B C_2H_4

C C_3H_6

D C_4H_8

3.2 C

- Empirical mass 14; $42/14 = 3$
- C_3H_6

3.3 Balance this equation and add state symbols: $\text{Mg} + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. [2]

3.3

- $\text{Mg}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$

Exercise 3.4 (a)

3.4 A compound contains 2.4 g of carbon, 0.4 g of hydrogen and 3.2 g of oxygen (A_r : C 12, H 1, O 16).

(a) Calculate its empirical formula.

[3]

Exercise 3.4 (a)

3.4 A compound contains 2.4 g of carbon, 0.4 g of hydrogen and 3.2 g of oxygen (A_r : C 12, H 1, O 16).

(a) Calculate its empirical formula. [3]

(a)

- moles: C = $2.4/12 = 0.20$, H = $0.4/1 = 0.40$, O = $3.2/16 = 0.20$; divide by 0.20
- 1 : 2 : 1; empirical formula CH_2O

Exercise 3.4 (b)

3.4 A compound contains 2.4 g of carbon, 0.4 g of hydrogen and 3.2 g of oxygen (A_r : C 12, H 1, O 16).

(b) The M_r of the compound is 60. Deduce its molecular formula. [1]

Exercise 3.4 (b)

3.4 A compound contains 2.4 g of carbon, 0.4 g of hydrogen and 3.2 g of oxygen (A_r : C 12, H 1, O 16).

(b) The M_r of the compound is 60. Deduce its molecular formula. [1]

(b)

- empirical mass 30; $60/30 = 2$
- $\text{C}_2\text{H}_4\text{O}_2$

Exercise 3.5 (a)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(a) Calculate the empirical formula of the oxide, showing your working. [4]

Exercise 3.5 (a)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(a) Calculate the empirical formula of the oxide, showing your working. [4]

(a)

- mass of oxygen = $2.80 - 1.96 = 0.84$ g; $n(\text{Fe}) = \frac{1.96}{55.8} = 0.0351$ mol and $n(\text{O}) = \frac{0.84}{16.0} = 0.0525$ mol; ratio = $1 : 1.49 = 2 : 3$, so the empirical formula is Fe_2O_3

Exercise 3.5 (b)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

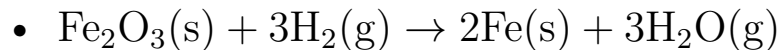
(b) Write a balanced equation, with state symbols, for the reduction of this oxide by hydrogen. [2]

Exercise 3.5 (b)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(b) Write a balanced equation, with state symbols, for the reduction of this oxide by hydrogen. [2]

(b)



Exercise 3.5 (c)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(c) The iron produced is dissolved in excess dilute sulfuric acid. Write the full equation and the ionic equation for this reaction. [3]

Exercise 3.5 (c)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(c) The iron produced is dissolved in excess dilute sulfuric acid. Write the full equation and the ionic equation for this reaction. [3]

(c)

- $\text{Fe(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{FeSO}_4(\text{aq}) + \text{H}_2(\text{g})$; ionic: $\text{Fe(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
– the sulfate is a spectator

Exercise 3.5 (d)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(d) A different oxide of iron has the empirical formula FeO. Explain why an empirical formula alone cannot give the molecular formula of a compound. [2]

Exercise 3.5 (d)

3.5 A 2.80 g sample of an oxide of iron is reduced completely by hydrogen, leaving 1.96 g of iron. (A_r : Fe 55.8, O 16.0)

(d) A different oxide of iron has the empirical formula FeO. Explain why an empirical formula alone cannot give the molecular formula of a compound. [2]

(d)

- The empirical formula gives only the **simplest whole-number ratio**; any molecule that is a whole-number multiple of that ratio has the same empirical formula, so the relative molecular mass is also needed