A-LEVEL CHEMISTRY • EXERCISE SHEET

2.3 Formulas

Name: ____________ Class: ________ Date: __________ **Total: 22 marks**

KEY WORDS empirical 实验式 • molecular 分子式 • relative molecular mass 相对分子质量 • equations 方程式 • balanced 配平

Objective

Build the skills to answer exam questions on **formulas** — writing formulas of ionic compounds, balancing **equations** 方程式 (including **ionic equations** 离子方程式), and finding **empirical** 实验式 and **molecular** 分子式 formulas.

You must be able to:

- write formulas of ionic compounds from ionic charges and oxidation numbers
- write and balance equations (including ionic equations) with state symbols
- calculate empirical and molecular formulas from data

1 Worked examples

Ionic formulas

$$A_r = \frac{(75 \times 35) + (25 \times 37)}{100} = 35.5$$

Formulae build from relative atomic masses

The total + charge balances the total − charge. Charges from the group: Group 1 +1, 2 +2, 13 +3, 15 −3, 16 −2, 17 −1. Learn: NO_3^-, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+. E.g. iron(III) oxide = Fe_2O_3 (two Fe^{3+} balance three O^{2-}).

Equations and ionic equations

A balanced equation has the same atoms on both sides; add **state symbols** (s, l, g, aq). An **ionic equation** shows only what changes (no **spectator ions** 旁观离子): $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

Empirical formula

Divide each element's mass (or %) by its A_r → moles; divide by the smallest; round to whole numbers. **Example.** 40.0% C, 6.7% H, 53.3% O: moles 3.33 : 6.7 : 3.33 → 1 : 2 : 1 → CH_2O. With M_r, scale up to the **molecular formula**.

2 Practice

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]

3 Exam-style questions

3.1 What is the formula of iron(III) sulfate? [1]

- **A** $FeSO_4$
- **B** Fe_2SO_4
- **C** $Fe_2(SO_4)_3$
- **D** $Fe_3(SO_4)_2$

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: $Mg + HCl \rightarrow MgCl_2 + H_2$. [2]

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]

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

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]

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

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

(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]

Solutions

2.1 (a) $Ca(NO_3)_2$; (b) $Al_2(SO_4)_3$.

2.2 $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

3.1 C. Two Fe^{3+} balance three $SO_4^{2-} \rightarrow Fe_2(SO_4)_3$.

3.2 C. Empirical mass 14; $42/14 = 3 \rightarrow C_3H_6$.

3.3 $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$.

3.4 (a) moles: C $= 2.4/12 = 0.20$, H $= 0.4/1 = 0.40$, O $= 3.2/16 = 0.20$; divide by 0.20 $\rightarrow 1:2:1$; empirical formula CH_2O.
(b) empirical mass 30; $60/30 = 2 \rightarrow C_2H_4O_2$.

3.5 (a) mass of oxygen $= 2.80 - 1.96 = 0.84$ g; $n(\text{Fe}) = \dfrac{1.96}{55.8} = 0.0351$ mol and $n(\text{O}) = \dfrac{0.84}{16.0} = 0.0525$ mol; ratio $= 1 : 1.49 = 2 : 3$, so the empirical formula is Fe_2O_3.
(b) $Fe_2O_3(s) + 3H_2(g) \rightarrow 2Fe(s) + 3H_2O(g)$.
(c) $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$; ionic: $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$ — the sulfate is a spectator.
(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.