

2.3 Formulas

Name: _____ Class: _____ Date: _____

Total: 22 marks

KEY WORDS empirical • molecular • relative molecular mass
 • equations • balanced

Objective

Build the skills to answer exam questions on **formulas** – writing formulas of ionic compounds, balancing **equations** (including **ionic equations**), and finding **empirical** and **molecular** 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 Key knowledge and 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**):
 $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$.

■ Empirical formula

Divide each element's mass (or %) by its $A_r \rightarrow$ moles; divide by the smallest; round to whole numbers. **Example.** 40.0% C, 6.7% H, 53.3% O: moles $3.33 : 6.7 : 3.33 \rightarrow 1 : 2 : 1 \rightarrow \text{CH}_2\text{O}$. 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 $\text{Fe}_2(\text{SO}_4)_3$
☐ D $\text{Fe}_3(\text{SO}_4)_2$

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: $\text{Mg} + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. [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

Each answer shows the steps, then the **result**.

2.1

(a)

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

(b)

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

2.2

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

3.1 C

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

3.2 C

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

3.3

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

3.4

(a)

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

(b)

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

3.5

(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

(b)

- $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{H}_2\text{O}(\text{g})$

(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

(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