Cambridge (CIE) A Level Chemistry

Your notes

Formulas

Contents

Formulae of Ionic Compounds

- **Ionic compounds** are formed from a **metal** and a **non-metal** bonded together
- Ionic compounds are electrically neutral; the positive charges equal the negative charges

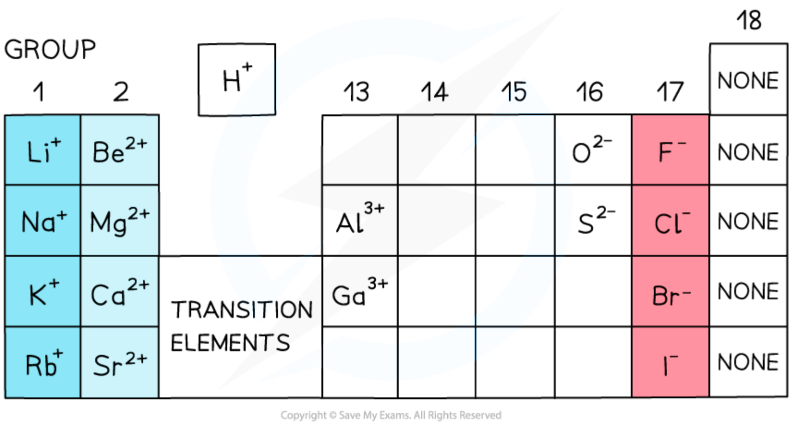
Charges on positive ions

- All metals form **positive** ions
  - There are some non-metal positive ions such as ammonium, NH_4^+, and hydrogen, H^+
- The **metals** in Group 1, Group 2 and Group 13 have a charge of 1+ and 2+ and 3+ respectively
- The charge on the ions of the **transition elements can vary** which is why **Roman numerals** are often used to indicate their charge
- **Roman numerals** are used in some compounds formed from transition elements to show the **charge** (or **oxidation state**) of metal ions
  - E.g. in copper (II) oxide, the copper ion has a charge of 2+ whereas in copper (III) nitrate, the copper has a charge of 3+

Non-metal ions

- The **non-metals** in Group 15 to 17 have a negative charge and have the suffix '**ide**'
  - E.g. nitride, chloride, bromide, iodide
- Elements in Group 17 gain 1 electron so have a 1– charge, eg. Br^-
- Elements in Group 16 gain 2 electrons so have a 2– charge, eg. O^{2-}
- Elements in Group 15 gain 3 electrons so have a 3– charge, eg. N^{3-}
- There are also more **complex** negative ions, which are negative ions made up of more than one type of atom

Periodic table outlining common ion charges



The charges of simple ions depend on their position in the Periodic Table

Formulae of ions

- Silver(I); Ag^{+}
- Ammonium: NH_4^{+}
- Zinc(II): Zn^{2+}
- Hydroxide: OH^{-}
- Nitrate: NO_3^{-}
- Sulfate: SO_4^{2-}
- Carbonate: CO_3^{2-}
- Hydrogen carbonate: HCO_3^{-}
- Phosphate: PO_4^{3-}

Worked Example

Determine the formulae of the following ionic compounds:

- Magnesium chloride
- Iron(III) oxide
- Aluminium nitrate

Answer 1: Magnesium chloride

- Magnesium is in Group 2 so has a charge of 2+
- Chlorine is in Group 17 so has a charge of 1–
- Magnesium needs two chlorine atoms for each magnesium atom to be balanced so the formula is **$MgCl_2$**

Answer 2: Iron(III) oxide

- The Roman numeral states that iron has a charge of 3+

Your notes

- Oxygen is in Group 16 so has a charge of 2–
- The charges need to be equal so 2 iron to 3 oxygen atoms will balance electrically, so the formula is **Fe_2O_3**

Answer 3: Aluminium nitrate

- Aluminium is in Group 13 so has a charge of 3+
- Nitrate is a **compound ion** and has a charge of 1–
- The complex ion needs to be placed in a bracket if more than 1 is needed
- The formula of aluminium nitrate is **$Al(NO_3)_3$**

Examiner Tips and Tricks

Remember: **Compound ions** are ions that contain more than one type of element, such as OH^-

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Balancing Equations

- A **symbol** equation is a shorthand way of describing a chemical reaction using **chemical symbols** to show the number and type of each atom in the reactants and products
- A **word** equation is a longer way of describing a chemical reaction using only **words** to show the reactants and products

Balancing equations

- During chemical reactions, atoms cannot be **created** or **destroyed**
- The number of each atom on each side of the reaction must therefore be the **same**
    - E.g. the reaction needs to be **balanced**
- When balancing equations remember:
    - Not to change any of the formulae
    - To put the numbers used to balance the equation **in front** of the formulae
    - To balance firstly the carbon, then the hydrogen and finally the oxygen in **combustion reactions** of organic compounds
- When balancing equations follow the following the steps:
    - Write the formulae of the reactants and products
    - Count the numbers of atoms in each reactant and product
    - Balance the atoms one at a time until all the atoms are balanced
    - Use appropriate state symbols in the equation
- The **physical state** of reactants and products in a chemical reaction is specified by using **state symbols**
    - **(s)** solid
    - **(l)** liquid
    - **(g)** gas
    - **(aq)** aqueous

Ionic equations

- In aqueous solutions, ionic compounds **dissociate** into their ions
- Many chemical reactions in aqueous solutions involve ionic compounds, however, only some of the ions in solution take part in the reactions
- The ions that do **not** take part in the reaction are called **spectator ions**

- An **ionic equation** shows **only** the ions or other particles taking part in a reaction, without showing the spectator ions

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Worked Example

Balance the following equation:

magnesium + oxygen → magnesium oxide

Answer:

- **Step 1:** Write out the symbol equation showing reactants and products

$$\mathbf{Mg + O_2 \rightarrow MgO}$$

- **Step 2:** Count the number of atoms in each reactant and product

	Mg	O
Reactants	1	2
Products	1	1

- **Step 3:** Balance the atoms one at a time until all the atoms are balanced

$$\mathbf{2Mg + O_2 \rightarrow 2MgO}$$

This is now showing that 2 moles of magnesium react with 1 mole of oxygen to form 2 moles of magnesium oxide

- **Step 4:** Use appropriate **state symbols** in the fully balanced equation

$$\mathbf{2Mg\,(s) + O_2\,(g) \rightarrow 2MgO\,(s)}$$

Worked Example

1. Balance the following equation:

zinc + copper(II) sulfate → zinc(II) sulfate + copper

2. Write the ionic equation for the above reaction.

Answer 1:

- **Step 1:** To balance the equation, write out the symbol equation showing reactants and products

$$\mathbf{Zn + CuSO_4 \rightarrow ZnSO_4 + Cu}$$

- **Step 2:** Count the number of atoms in each reactant and product. The equation is already balanced

	Zn	Cu	S	O
Reactants	1	1	1	4

Products	1	1	1	4

- **Step 3:** Use appropriate **state symbols** in the equation

$$Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

Answer 2:

- **Step 1:** The full chemical equation for the reaction is

$$Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

- **Step 2:** Break down reactants into their respective ions

$$Zn\,(s) + Cu^{2+}\,SO_4{}^{2-}\,(aq) \rightarrow Zn^{2+}SO_4{}^{2-}\,(aq) + Cu\,(s)$$

- **Step 3:** Cancel the spectator ions on both sides to give the ionic equation

$$Zn\,(s) + Cu^{2+}SO_4{}^{2-}\,(aq) \rightarrow Zn^{2+}SO_4{}^{2-}\,(aq) + Cu\,(s)$$

$$Zn\,(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}\,(aq) + Cu\,(s)$$

Your notes

Your notes

Empirical & Molecular Formulae

- The **molecular formula** is the formula that shows the **number** and **type** of each atom in a molecule
  - Eg. the molecular formula of ethanoic acid is $C_2H_4O_2$
- The **empirical formula** is the simplest whole number ratio of the elements present in one molecule or formula unit of the compound
  - Eg. the empirical formula of ethanoic acid is CH_2O
- **Organic molecules** often have **different** empirical and molecular formulae
- **Simple inorganic molecules** however have **often similar** empirical and molecular formulae
- **Ionic compounds always** have **similar** empirical and molecular formulae

Empirical & Molecular Formulae Calculations

Empirical formula

- The **empirical formula** is the **simplest whole-number ratio** of the elements present in one molecule or formula unit of the compound
- It is calculated from knowledge of the ratio of masses of each element in the compound
- The empirical formula can be found by determining the **mass** of each element present in a sample of the compound
- It can also be deduced from data that gives the **percentage compositions by mass** of the elements in a compound

Worked Example

Calculating empirical formula from mass

Determine the empirical formula of a compound that contains 10 g of hydrogen and 80 g of oxygen.

Answer:

Elements	Hydrogen	Oxygen
Mass of each element **(g)**	10	80

Atomic mass	1.0	16.0
Moles = mass / A_r	$\frac{10}{1.0} = 10$	$\frac{80}{16.0} = 5$
Ratio (divide by smallest value)	$\frac{10}{5} = 2$	$\frac{5}{5} = 1$

- So, the empirical formula of the compound is H_2O

Worked Example

Calculating empirical formula from %

Determine the empirical formula of a compound that contains 85.7% carbon and 14.3% hydrogen.

Answer:

Elements	Carbon	Hydrogen
Mass of each element **(%)**	85.7	14.3
Atomic mass	12.0	1.0
Moles = mass / A_r	$\frac{85.7}{12.0} = 7.14$	$\frac{14.3}{1.0} = 14.3$
Ratio (divide by smallest value)	$\frac{7.14}{7.14} = 1$	$\frac{14.3}{7.14} = 2.00$

- So, the empirical formula of the compound is CH_2

Molecular formula

- The **molecular formula** gives the exact number of atoms of each element present in the formula of the compound
- The molecular formula can be found by dividing the **relative formula mass** of the **molecular formula** by the **relative formula mass** of the **empirical formula**
- **Multiply** the number of each element present in the empirical formula by this number to find the molecular formula

Worked Example

Calculating molecular formula

Your notes

The empirical formula of **X** is $C_4H_{10}S$ and the relative formula mass of **X** is 180.

What is the molecular formula of **X**?

Relative atomic mass Carbon: 12 Hydrogen:1 Sulfur:32

Answer:

- **Step 1:** Calculate the relative formula mass of the empirical formula
  - Relative formula mass = (C x 4) + (H x 10) + (S x 1)
  - Relative formula mass = (12 x 4) + (1 x 10) + (32 x 1)
  - Relative formula mass = 90
- **Step 2:** Divide the relative formula mass of **X** by the relative formula mass of the empirical formula
  - Ratio between M_r of **X** and the M_r of the empirical formula = 180/90
  - Ratio between M_r of **X** and the M_r of the empirical formula = 2
- **Step 3:** Multiply each number of elements by 2
  - $(C_{4 \times 2}) + (H_{10 \times 2}) + (S_{1 \times 2}) = (C_8) + (H_{20}) + (S_2)$
  - The molecular formula of **X** is $C_8H_{20}S_2$

Your notes

Water of Crystallisation

- **Water of crystallisation** is when some compounds can form **crystals** which have **water** as part of their structure
- A compound that contains water of crystallisation is called a **hydrated compound**
- The water of crystallisation is separated from the main formula by a **dot** when writing the chemical formula of hydrated compounds
  - E.g. hydrated copper(II) sulfate is $CuSO_4•5H_2O$
- A compound which doesn't contain water of crystallisation is called an **anhydrous compound**
  - E.g. anhydrous copper(II) sulfate is $CuSO_4$
- A compound can be hydrated to **different degrees**
  - E.g. cobalt(II) chloride can be hydrated by **six** or **two** water molecules
  - $CoCl_2•6H_2O$ or $CoCl_2•2H_2O$
- The conversion of hydrated compounds to anhydrous compounds is achieved by heating the hydrated salt
- This process is reversed by adding water, which reforms the hydrated compound:

Hydrated: $CuSO_4•5H_2O \rightleftharpoons CuSO_4 + 5H_2O$:Anhydrous

- The degree of hydration can be calculated from experimental results:
  - The mass of the hydrated salt must be measured before heating
  - The salt is then heated until it reaches a constant mass
  - The two mass values can be used to calculate the number of moles of water in the hydrated salt - known as the water of crystallisation

Worked Example

Calculating water of crystallisation

10.0 g of hydrated copper sulfate are heated to a constant mass of 5.59 g.

Determine the formula of the original hydrated copper sulfate, $CuSO_4•xH_2O$.

(M_r data: $CuSO_4$ = 159.6, H_2O = 18.0)

Answer:

Compounds	$CuSO_4$	H_2O
Mass of each compound **(g)**	5.59	4.41
Formula mass	159.6	18.0
Moles = mass / M_r	$\frac{5.59}{159.6}$ = 0.035	$\frac{4.41}{18.0}$ = 0.245
Ratio (divide by smallest value)	$\frac{0.035}{0.035}$ = 1	$\frac{0.245}{0.035}$ = 7

- So, the value of x is 7
- Therefore, the formula of the hydrated salt is $CuSO_4 \bullet 7H_2O$

Your notes

Examiner Tips and Tricks

A water of crystallisation calculation can be completed in a similar fashion to an empirical formula calculation

- Instead of elements, you start with the salt and water
- Instead of dividing by atomic masses, you divide by molecular / formula masses
- The rest of the calculation works the same way as the empirical formula calculation