Cambridge (CIE) IGCSE Chemistry

Your notes

Formulae & Relative Masses

Contents

- Formulae
- Empirical Formulae & Formulae of Ionic Compounds
- Writing Equations
- Ar & Mr

Your notes

Molecular formulae

- Elements are often represented using their **chemical symbol** from the Periodic Table

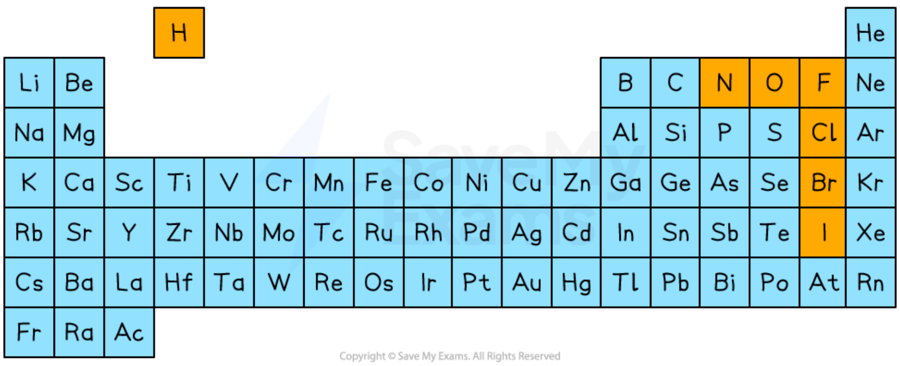
Chemical symbols of elements table

Element	Symbol	Element	Symbol
Lithium	Li	Calcium	Ca
Chromium	Cr	Gold	Au
Aluminium	Al	Sulfur	S

- Seven elements exist as diatomic **molecules**, which means that they are molecules made of two atoms:

1. Hydrogen, H_2
2. Nitrogen, N_2
3. Oxygen, O_2
4. Fluorine, F_2
5. Chlorine, Cl_2
6. Bromine, Br_2
7. Iodine, I_2

Periodic Table identifying the 7 diatomic elements



The highlighted elements all exist as diatomic molecules with the formula X_2

- These 7 elements are also classed as **simple molecules**

- Atoms combine together in **fixed ratios** that will give them full outer shells of electrons
    - When this happens, a molecule is formed
- The molecular formula of a molecule shows:
    - The type of atoms involved, given by the chemical symbol
      AND
    - The number of atoms, given by the subscript (little) number after a chemical symbol

Examples of molecular formulae

Substance	Molecular formula	Made from
Hydrogen	H_2	2 hydrogen atoms
Chlorine	Cl_2	2 chlorine atoms
Water	H_2O	2 hydrogen atoms 1 oxygen atom
Methane	CH_4	1 carbon atom 4 hydrogen atoms
Ammonia	NH_3	1 nitrogen atom 3 hydrogen atoms
Sulfuric acid	H_2SO_4	2 hydrogen atoms 1 sulfur atom 4 oxygen atoms

- The table also shows that the molecular formula can be deduced from the relative number of atoms present
    - E.g. Ammonia contains 3 atoms of hydrogen and 1 atom of nitrogen, which means its molecular formula is NH_3

Diagrammatic representation of chemicals

- Diagrams or models can be used to represent and/or deduce the molecular formula of elements and simple compounds:

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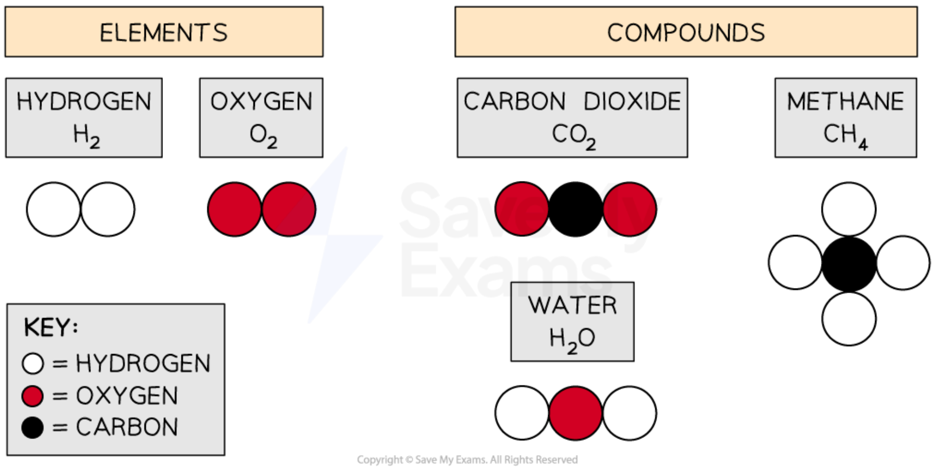
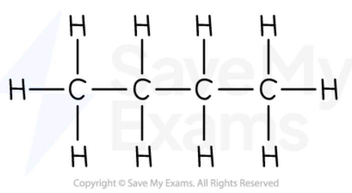


When simple molecules are represented using coloured atoms, a key is used to show the colours for each type of atom

- Using the water examples, it is important to know that these representations also show the arrangement of the atoms in the molecule
    - Water, H_2O
        - The hydrogen atoms are on either side of the oxygen atom
        - It does not have two hydrogen atoms and an oxygen atom all joined together in a row

Worked Example

What is the molecular formula of the following compound?



Answer:

- The molecule contains:
    - 4 carbon atoms
    - 10 hydrogen atoms
- Therefore, the molecular formula is C_4H_{10}

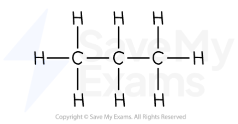
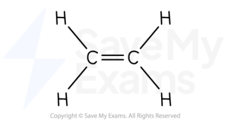
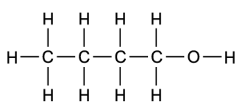
Empirical formulae

Extended tier only

- The **empirical formula** is the simplest whole number ratio of the atoms of each element present in one molecule or formula unit of the compound
- The empirical formula of an organic molecule is often different to its molecular / chemical formula
  - For example, ethanoic acid has the chemical formula CH_3COOH or $C_2H_4O_2$ but its empirical formula is CH_2O
- The molecular / chemical formula of an ionic compound is always its empirical formula
  - For example, sodium chloride has the chemical formula NaCl, which is also its empirical formula

Worked Example

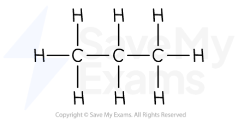
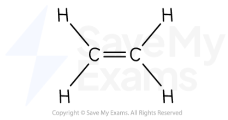
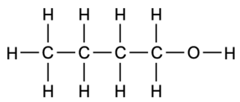
Complete the table to give the molecular and empirical formulae of the given compounds.

Chemical	Molecular formula	Empirical formula
		
		
		

Answers:

The completed table is:

Chemical	Molecular formula	Empirical formula

	C_3H_8	C_3H_8
	C_2H_4	CH_2
	$C_4H_{10}O$	$C_4H_{10}O$

- The first compound contains 3 carbon atoms and 8 hydrogen atoms
  - This 3:8 ratio of atoms cannot be simplified
  - Therefore, the molecular and empirical formula are both C_3H_8
- The second compound contains 2 carbon atoms and 4 hydrogen atoms
  - This 2:4 ratio of atoms can be simplified to 1:2
  - Therefore, the molecular formula is C_2H_4 and the empirical formula is CH_2
- The third compound contains 4 carbon atoms, 10 hydrogen atoms and 1 oxygen atom
  - This 10:4:1 ratio of atoms cannot be simplified
  - Therefore, the molecular and empirical formula are both $C_4H_{10}O$

Deducing formulae of ionic compounds

Extended tier only

- Metals and non-metals react together to form ionic compounds
  - Ionic compounds are **not** simple molecules
  - **Remember:** Simple molecules are formed when non-metals react together to form compounds
- Ionic compounds involve the metal losing electrons and the non-metal gaining electrons to form ions
- Some ions that you will be expected to be able to use, because they are stated in the exam specification, include:
  - Hydrogen ions, H^+ - sometimes referred to as protons
  - Group 1 ions, e.g. Li^+, Na^+, K^+
  - Group 7 ions, F^-, Cl^-, Br^-
  - Copper(II) ions, Cu^{2+}
  - Iron(II) ions, Fe^{2+}

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- Iron(III) ions, Fe^{3+}
- There are some polyatomic (containing more than one atom) ions stated in the exam specification:
  - Carbonate ions, CO_3^{2-}
  - Sulfate ions, SO_4^{2-}
  - Hydroxide ions, OH^-
  - Nitrate ions, NO_3^-
  - Ammonium ions, NH_4^+

How to determine the formulae of ionic compounds

- Ionic compounds typically have **no overall charge**
  - This means that the size of any positively charged ion is cancelled by the size of any negatively charged ion
  - **Careful:** This should not be confused with an atom having no overall charge

Direct comparison

- The formula of an ionic compound can be determined by directly comparing the charges of the ions:
  - For example, iron(II) sulfate
    - The iron(II) ion is Fe^{2+}, which means that it has a 2+ or +2 charge
    - The sulfate ion is SO_4^{2-}, which means that it has a 2– or –2 charge
    - The charges cancel each other out

      Mathematically, $(+2) + (-2) = 0$
    - This means that one SO_4^{2-} ion is needed to cancel the +2 charge on Fe^{2+}
    - Therefore, the formula of iron(II) sulfate is $FeSO_4$

The swap-and-drop method

- When the ions in the ionic compound have **different charges**, it can be easier to use the **swap-and-drop** method
  - **Careful:** If you use this method with ions that have the same charge, then you must give the simplest whole number ratio to get the correct answer
- For example, copper(II) chloride:
  - The copper(II) ion is Cu^{2+}, which means that it has a 2+ or +2 charge
  - The chloride ion is Cl^-, which means that it has a 1– or –1 charge
  - The size of the charge on the copper(II) ion indicates the number of chloride ions needed, and the size of the charge on the chloride ion indicates the number of

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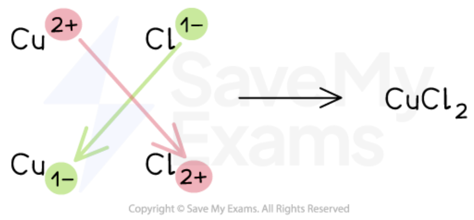
copper(II) ions needed

Determining the formula of copper(II) chloride



The charges swap from element to element and drop down. The positive and negative signs are removed and there is no need for the number 1.

- This gives the overall formula of copper(II) chloride as $CuCl_2$

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Worked Example

The compound produced in the reaction between iron wool and chlorine contains the ions Fe^{3+} and Cl^{-}.

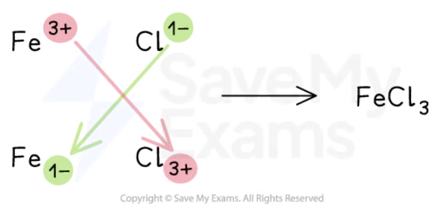
a) Give the formula of this compound.

b) State the name of this compound.

Answers:

Part a)

- **Direct comparison method:**
  - The iron ion is Fe^{3+}, which means that it has a 3+ or +3 charge
  - The chloride ion is Cl^{-}, which means that it has a 1– or –1 charge
  - The charges do not cancel each other out
    - Mathematically, $(+3) + (-1) \neq 0$
  - Three Cl^{-} ions are needed to cancel the +3 charge on Fe^{3+}
  - Therefore, the formula is $\mathbf{FeCl_3}$
- **Swap-and-drop method**



- The formula is $\mathbf{FeCl_3}$

Part b)

- The metal is iron and the chlorine will change to chloride
- Therefore the name is **iron chloride**

Examiner Tips and Tricks

Take your time determining the chemical formula of ionic compounds with

- Different charges on the ions
- Polyatomic ions

Your notes

Writing word equations & symbol equations

Word equations

- Word equations show the **reactants** and **products** of a chemical reaction using their full chemical names

reactants → products

- The reactants are the substances on the **left-hand side** of the arrow
  - They can be thought of as the chemical **ingredients** of the reaction
- They react with each other to form new substances, which are the products
- The products are on the **right-hand side** of the arrow
- The arrow (which is spoken as "*to form*" or "*produces*") implies the **conversion** of reactants into products
  - Reaction **conditions** or the name of a **catalyst** (a substance added to make a reaction go faster) can be written above the arrow
- An example is the reaction of sodium hydroxide (a base) and hydrochloric acid to produce sodium chloride (common table salt) and water:

sodium hydroxide + hydrochloric acid ⟶ sodium chloride + water

Worked Example

1. Ammonia reacts with nitric acid to form the fertiliser ammonium nitrate. Write a word equation for the reaction taking place.
2. Iron(II) hydroxide and sodium sulfate are formed when iron(II) sulfate solution and sodium hydroxide react together. Write a word equation for the reaction taking place.
3. Carbon is the main element found in coal and burns in air to produce carbon dioxide. Write a word equation for the reaction taking place.

Answers:

1. Ammonia + nitric acid → ammonium nitrate
   - This question has all the information in the correct order
   - *Ammonia reacts with nitric acid*
     - This becomes ammonia + nitric acid
   - *to form*
     - This is the arrow in the equation

- *to form the fertiliser ammonium nitrate*
    - This tells you that the product is ammonium nitrate

2. Iron(II) sulfate + sodium hydroxide → iron(II) hydroxide + sodium sulfate
    - **Careful:** This question has all the required information but the products are written first
    - *Iron(II) hydroxide and sodium sulfate are formed*
        - This becomes → iron(II) hydroxide + sodium sulfate
    - *when iron(II) sulfate solution and sodium hydroxide react together*
        - This becomes Iron(II) sulfate + sodium hydroxide →

3. Carbon + oxygen → carbon dioxide
    - **Careful:** Not all of the required information is given in the question
    - You are expected to know that burning in air means that the chemical is reacting with oxygen
    - *Carbon... ...burns in air*
        - This becomes carbon + oxygen
    - *to produce*
        - This is the arrow in the equation
    - *to produce carbon dioxide*
        - This tells you that the product is carbon dioxide

Symbol equations

- A **symbol equation** uses the formulae of the reactants and products to show what happens in a chemical reaction
- When writing symbol equations, you should:
    - Ensure reactants are on the left of the equation and products are on the right
    - Write the following non-metals as molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2
    - Include state symbols
        - Solid = (s)
        - Liquid = (l)
        - Gas = (g)
        - Aqueous = (aq)
- Sometimes it can be hard to know what the correct state symbol is and we have to look for clues in the identity of substances in a reaction
- Generally, unless they are in a solution:
    - Metal compounds will always be solid, although there are a few exceptions
    - Ionic compounds will usually be solids

Your notes

Your notes

- Non-metal compounds can be solids, liquids or gases
  - So, it depends on information given
- Precipitates formed in solution count as solids
- A symbol equation must be balanced to give the correct ratio of reactants and products:
  - For example, the combustion of sulfur:

$$S\ (s) + O_2\ (g) \rightarrow SO_2\ (g)$$

- This equation shows that one atom of solid sulfur, S, reacts with one gaseous molecule of oxygen, O_2, to make one gaseous molecule of sulfur dioxide, SO_2

Examiner Tips and Tricks

- In exams, you will not need to include them in all equations unless you are specifically asked to .
  - However, it is good practice to include state symbols in your equations so that you don't miss any marks.
- Be careful when writing the state symbol of solutions of liquids.
  - For example, ethanol, or common alcohol, is a liquid at room temperature, so if it is pure alcohol then you would be using (l) as the state symbol.
  - However, most of the time alcohol is used as a solution in water so the state symbol should be (aq).

Balancing symbol equations

- When balancing equations, there must be the **same number of atoms** of each element on either side of the equation following the law of conservation of mass
- To balance an equation you work across the equation from left to right, checking one element after another
  - If there is a group of atoms such as a nitrate group (NO_3^-) that has not changed from one side to the other, then count the whole group as one entity rather than counting the individual atoms
- Examples of balanced symbol / chemical equations include:
  - Acid-base neutralisation reaction:

$$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$$

- Redox reaction:

$$2Fe_2O_3\ (aq) + 3C\ (s) \rightarrow 4Fe\ (s) + 3CO_2\ (g)$$

- In each equation, there are equal numbers of each atom on either side of the reaction arrow so the equations are balanced

- The best approach is to practice lot of examples of balancing equations
- This can be by trial and error - changing the coefficients (numbers) in front of the formulae one by one and checking the result on the other side
- Balance elements that appear on their own, last in the process

Your notes

Worked Example

When magnesium oxide, MgO, reacts with nitric acid, HNO_3, it forms magnesium nitrate, $Mg(NO_3)_2$, and water.

magnesium oxide + nitric acid ⟶ magnesium nitrate + water

Write the balanced **symbol** equation for this reaction.

Answer:

- The balanced symbol equation is:
  $MgO\ (s) + \mathbf{2}HNO_3\ (aq) \longrightarrow Mg(NO_3)_2\ (aq) + H_2O\ (l)$
- **Step 1** - writing the unbalanced equation
  - Magnesium oxide, MgO, reacts with nitric acid, HNO_3, it forms magnesium nitrate, $Mg(NO_3)_2$, and water
    - $MgO + HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$
  - The Mg and O atoms (not including the O in the NO_3 group appear to be balanced), so we should focus on the H atoms and NO_3 groups
- **Step 2** - balancing hydrogen atoms
  - There are 2 hydrogen atoms on the product side, so 2 hydrogen atoms are needed on the reactant side
  - This means that $2HNO_3$ will be needed as we cannot change the chemical formula
    - $MgO + \mathbf{2}HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$
  - This also balances the nitrate, NO_3, groups
- **Step 3** - checking the equation
  - The equation appears balanced so we need to check that it is:
  - Reactant side:
    - Mg atom
    - 1 O atom - not including those in the NO_3 group
    - 2 H atoms
    - 2 NO_3 groups - remember to keep groups as a single entity if they are unchanged on both sides of the equation
  - Product side:
    - 1 Mg atom
    - 2 NO_3 groups - remember to keep groups as a single entity if they are unchanged on both sides of the equation
    - 2 H atoms

- 1 O atom - not including those in the NO_3 group
- The equation is now balanced

Worked Example

Aluminium reacts with copper(II) oxide to produce aluminium oxide and copper.
Balance the symbol equation for the reaction taking place.

$$\_Al\,(s) + \_CuO\,(s) \longrightarrow \_Al_2O_3\,(s) + \_Cu\,(s)$$

Answer:

- The balanced symbol equation is:
  $$\mathbf{2}Al\,(s) + \mathbf{3}CuO\,(s) \longrightarrow Al_2O_3\,(s) + \mathbf{3}Cu\,(s)$$
- Step 1 - balancing aluminium atoms
  - There are 2 aluminium atoms on the product side, so 2 aluminium atoms are needed on the reactant side
    - $\mathbf{2}Al + CuO \longrightarrow Al_2O_3 + \_Cu$
- Step 2 - balancing oxygen atoms
  - There are 3 oxygen atoms on the product side, so 3 oxygen atoms are needed on the reactant side
  - This means that 3 CuO will be needed as we cannot change the chemical formula
    - $\mathbf{2}Al + \mathbf{3}CuO \longrightarrow Al_2O_3 + Cu$
- Step 3 - balancing copper atoms
  - There are 3 copper atoms on the reactant side, so 3 copper atoms are needed on the product side
    - $\mathbf{2}Al + \mathbf{3}CuO \longrightarrow \_Al_2O_3 + \mathbf{3}Cu$
- The equation is now balanced

Deducing symbol equations

Extended tier only

- For some reactions, you will not be given the unbalanced equation
  - You will be expected to know or deduce the formula of compounds and then balance the equations

Worked Example

Aluminium burns in chlorine to form the white solid, aluminium chloride.

Your notes

Write the balanced symbol equation, including state symbols, for the reaction.

Answer:

1. Work out the formulae and state symbols of the reactants and products:
    - Aluminium is a solid metal, like other pure metals, it is an element so its formula is the same as its chemical symbol: **Al (s)**
    - From your knowledge of Group VII elements, you should know that chlorine is a gas that exists as a diatomic molecule: **Cl_2 (g)**
    - Aluminum chloride is a solid - this information is given in the question as you would not be expected to know this.
        - Its formula is deduced from the charges on the ions present:
        - Aluminium has a 3+ charge
        - Chloride ions have a 1– charge
        - Therefore, for the compound to be neutral, 3 chloride ions are needed for every 1 aluminium ion: **$AlCl_3$ (s)**
2. Construct an unbalanced symbol equation:
    - The unbalanced symbol equation is:
        - Al (s) + Cl_2 (g) → $AlCl_3$ (s)
3. Balance the equation:
    - Make the number of Cl atoms on the right-hand side an even number by adding a 2 in front of $AlCl_3$:
        - Al (s) + Cl_2 (g) → **2**$AlCl_3$ (s)
    - This gives 6 Cl atoms on the right-hand side
    - So, now balance the number of Cl atoms, on the left-hand side, by adding a 3 in front of Cl_2:
        - Al (s) + **3**Cl_2 (g) → **2**$AlCl_3$ (s)
    - This gives 2 Al atoms on the right-hand side
    - So, balance the number of Al atoms, on the left-hand side, by adding a 2 in front of the Al:
        - **2**Al (s) + **3**Cl_2 (g) → **2**$AlCl_3$ (s)

Examiner Tips and Tricks

When balancing equations you **cannot** change any of the formulae, only the amount of each atom or molecule. This is done by changing the numbers that go in **front** of each chemical species.

Balancing Ionic Equations

- In aqueous solutions, ionic compounds **dissociate** into their ions
    - This means that they separate into their component ions

Your notes

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- For example, hydrochloric acid and potassium hydroxide dissociate as follows:

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-(aq)$$

$$KOH\ (aq) \rightarrow K^+\ (aq) + OH^-\ (aq)$$

- It is important that you can recognise common ionic compounds and their constituent ions
  - These include:
  - Acids such as HCl and H_2SO_4
  - Group I and Group II hydroxides e.g. sodium hydroxide
  - Soluble salts e.g. potassium sulfate, sodium chloride
- The steps to writing an **ionic equation** are:
  1. Write the full, balanced symbol equation
  2. Replace the ionic compounds in the balanced symbol equation with the component ions
  3. Remove any ions that appear on both sides of the equation

Worked Example

Write the ionic equation for the displacement reaction of aqueous chlorine and aqueous potassium iodide.

Answer:

1. Write out the full balanced equation:
   - $2KI\ (aq) + Cl_2\ (aq) \rightarrow 2KCl\ (aq) + I_2\ (aq)$
2. Replace the ionic compounds in the balanced symbol equation with the component ions
   - $2K^+ (aq) + 2I^- (aq) + Cl_2\ (aq) \rightarrow 2K^+ (aq) + 2Cl^- (aq) + I_2\ (aq)$
3. Remove any ions that appear on both sides of the equation:
   - $2I^- (aq) + Cl_2\ (aq) \rightarrow 2Cl^- (aq) + I_2\ (aq)$

Examiner Tips and Tricks

Ionic equations should **always** have state symbols included.

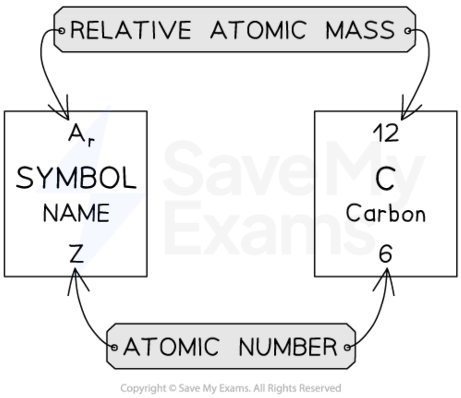
Your notes

Relative masses

Relative atomic mass

- The symbol for relative atomic mass is A_r
- The relative atomic mass for each element can be found in the Periodic Table, along with the atomic number
  - Relative atomic mass is shown on the atomic symbol
  - It is always larger than the atomic number (except for hydrogen, where they are the same)
  - Use the key on the Periodic Table to correctly identify the mass number

Key for chemical symbols on the Periodic Table



This key is given on the Periodic Table in exams and identifies the number that is the relative atomic mass.

- Atoms are too small to accurately weigh but scientists needed a way to compare the masses of atoms
  - Carbon-12 is used as the standard atom and has a fixed mass of 12 units
  - The mass of all other atoms are compared against carbon-12
- The relative atomic mass of carbon is 12
  - The relative atomic mass of magnesium is 24 which means that magnesium is twice as heavy as carbon
  - The relative atomic mass of hydrogen is 1 which means it has one-twelfth the mass of one carbon-12 atom

Relative molecular (formula) mass

Your notes

- The symbol for the relative molecular mass is M_r
- Relative molecular mass is the sum of the relative atomic masses of all the atoms in a molecule
  - The term **relative formula mass** is used when referring to the total mass of an **ionic** compound
- To calculate the M_r of a substance, you have to add up the relative atomic masses of all the atoms present in the formula

Relative Formula Mass Calculations Table

Substance	Atoms present	Calculation	M_r
Hydrogen **H_2**	2 x H	(2 x 1)	2
Water **H_2O**	(2 x H) + (1 x O)	(2 x 1) + (1 x 16)	18
Potassium carbonate **K_2CO_3**	(2 x K) + (1 x C) + (3 x O)	(2 x 39) + (1 x 12) + (3 x 16)	138
Calcium hydroxide **$Ca(OH)_2$**	(1 x Ca) + (2 x O) + (2 x H)	(1 x 40) + (2 x 16) + (2 x 1)	74
Ammonium sulfate **$(NH_4)_2SO_4$**	(2 x N) + (8 x H) + (1 x S) + (4 x O)	(2 x 14) + (8 x 1) + (1 x 32) + (4 x 16)	132

Worked Example

Calculate the relative formula mass of:

1. Sodium chloride, NaCl
2. Copper oxide, CuO
3. Magnesium nitrate, $Mg(NO_3)_2$

Answers:

1. Sodium chloride
   - NaCl = 23 + 35.5 = 58.5
2. Copper oxide
   - CuO = 63.5 + 16 = 79.5
3. Magnesium nitrate

- $Mg(NO_3)_2 = 24 + (14 \times 1 \times 2) + (16 \times 3 \times 2) = 148$

Reacting masses

- The Law of Conservation of mass tells us that mass cannot be created or destroyed
  - In a chemical reaction, the total mass of reactants equals the total mass of the products
- We can use this, along with relative atomic / formula masses to perform calculations to identify the quantities of reactants or products involved in a chemical reaction
- **Example**:

$$\mathbf{2Ca + O_2 \rightarrow 2CaO}$$

- Relative atomic masses: Ca = 40; O = 16
- Using the balanced symbol equation:
  - Reactants:
  - 2 x 40 = 80 units of mass of calcium
  - 2 x 16 = 32 units of mass of oxygen (O_2 molecule, 16 + 16 = 32)
  - Products:
  - 2 x (40 + 16) = 112 units of mass of CaO

$$\mathbf{2Ca + O_2 \rightarrow 2CaO}$$

$$\mathbf{80 + 32 \ = \ 112}$$

- The ratio of the mass of calcium and oxygen reacting will always be the same, regardless of the units
  - 80 g of calcium will react with 32 g of oxygen to form 112 g of calcium oxide
  - 80 tonnes of calcium will react with 32 tonnes of oxygen to form 112 tonnes of calcium oxide
  - So, 40 kg of calcium will react with 16 kg of oxygen to form 56 kg of calcium oxide

Worked Example

Calculate the mass of carbon dioxide produced when 32 g of methane, CH_4, reacts completely in excess oxygen:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Relative atomic masses, A_r: H = 1; C = 12; O = 16

Answer:

- Using the balanced symbol equation:

Your notes

- Reactants:
- $12 + (4 \times 1) = 16$ units of mass of methane
- $2 \times 16 = 32$ units of mass of oxygen (O_2 molecule, $16 + 16 = 32$)
- Products:
- $12 + (2 \times 16) = 44$ units of mass of carbon dioxide
- $2 \times ((2 \times 1) + 16) = 36$ units of mass of water

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

$$16 + 64 \rightarrow 44 + 36$$

- So, 16 g of methane would react in excess oxygen to form 44 g of carbon dioxide
- 32 g of methane is double the amount of methane from the balanced symbol equation
- So, this would produce double the amount of caron dioxide
  - $2 \times 44 =$ **88 g** of carbon dioxide

Your notes