Edexcel GCSE Chemistry

Your notes

Calculations Involving Masses

Contents

Relative Formula Mass

Your notes

Relative Formula Mass

- We have seen previously that the symbol for the relative atomic mass is A_r
- This is calculated from the **mass number** and **relative abundances** of all the **isotopes** of a particular element
- The symbol for the **relative formula mass** is M_r and it refers to the **total mass** of the substance
- If the substance is molecular you can use the term **relative molecular mass**, but this term should not be used for ionic compounds such as sodium chloride
- 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	M_r
Hydrogen (H_2)	$2 \times H$	$(2 \times 1) = 2$
Water (H_2O)	$(2 \times H) + (1 \times O)$	$(2 \times 1) + 16 = 18$
Potassium Carbonate (K_2CO_3)	$(2 \times K) + (1 \times C) + (3 \times O)$	$(2 \times 39) + 12 + (3 \times 16) = 138$
Calcium Hydroxide ($Ca(OH)_2$)	$(1 \times Ca) + (2 \times O) + (2 \times H)$	$40 + (2 \times 16) + (2 \times 1) = 74$
Ammonium Sulfate ($(NH_4)_2SO_4$)	$(2 \times N) + (8 \times H) + (1 \times S) + (4 \times O)$	$(2 \times 14) + (8 \times 1) + 32 + (4 \times 16) = 132$

- In accordance with the Law of Conservation of Mass, the sum of the relative formula masses of the reactants will be the same as the sum of the relative formula masses of the products

Examiner Tips and Tricks

If you are in any doubt about whether to use the term relative molecular mass or relative formula mass, use the latter because it applies to all compounds whether they are ionic or covalent.

Your notes

Empirical Formulae

Your notes

Empirical Formulae

Empirical Formula from Reacting Masses

- An **empirical formula** gives the **simplest whole number ratio** of atoms of each element in the compound
- It is calculated from knowledge of the ratio of masses of each element in the compound
- Suppose a compound contains 10 g of hydrogen and 80 g of oxygen. We can calculate the empirical formula by
  1. Dividing the reacting masses by the **relative atomic mass** of each element (this gives the moles)
  2. Divide each result by the lowest number obtained to give the simplest ratio
- This can be shown by the following calculations:
  - Amount of hydrogen atoms = Mass in grams ÷ A_r of hydrogen = (10 ÷ 1) = **10 moles**
  - Amount of oxygen atoms = Mass in grams ÷ A_r of oxygen = (80 ÷ 16) = **5 moles**
- **The ratio of moles of hydrogen atoms to moles of oxygen atoms:**

	HYDROGEN	OXYGEN
MOLES	10	5
RATIO	2	1

- Since equal numbers of moles of atoms contain the same number of atoms, the ratio of hydrogen atoms to oxygen atoms is 2:1
- Hence the **empirical formula** is H_2O

Empirical Formula from Molecular Formula

- By inspection you simply reduce the molecular formula to the simplest ratio and you have the empirical formula

- Sometimes the empirical formula is the same as the molecular formula, as in the example of methane
- The formula of ionic compounds is always the empirical formula

Relationship between Empirical and Molecular Formula

NAME OF COMPOUND	EMPIRICAL FORMULA	MOLECULAR FORMULA
METHANE	CH_4	CH_4
ETHANE	CH_3	C_2H_6
ETHENE	CH_2	C_2H_4
BENZENE	CH	C_6H_6

Molecular Formula from Empirical Formula

- To find the molecular formula from the empirical formula you need to know the **relative formula mass** of the substance
- The steps involved are:
    1. Find the empirical formula mass
    2. Divide the relative formula mass by the empirical formula mass to obtain the multiple
    3. Multiple this number by the empirical formula to obtain the molecular formula

Worked Example

The empirical formula of X is $C_4H_{10}S_1$ and the relative formula mass of X is 180. What is the molecular formula of X?

Relative Formula Masses: carbon : 12 hydrogen : 1 sulfur : 32

Answer:

- **Step 1 –** Calculate the relative empirical formula mass

$(C \times 4) + (H \times 10) + (S \times 1) = (12 \times 4) + (1 \times 10) + (32 \times 1) = 90$

Your notes

Your notes

- **Step 2 –** Divide relative formula mass of X by the relative empirical mass

180 / 90 = 2

- **Step 3 –** Multiply each number of elements by 2

$(C_{4 \times 2}) + (H_{10 \times 2}) + (S_{1 \times 2}) = (C_8) + (H_{20}) + (S_2)$

- **Molecular formula of X =** $C_8H_{20}S_2$

Examiner Tips and Tricks

Sometimes when you are finding the empirical formula from the reacting masses of two elements you do not get an exact whole number in step 2 after dividing by the relative atomic masses. However, it should be close to a whole number, so just round up or down to get the answer.

Your notes

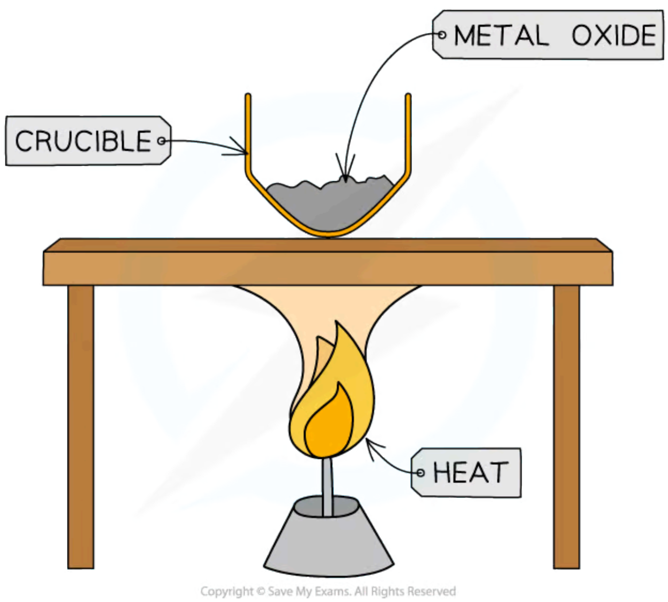
Determine an Empirical Formula

Determine an Empirical Formula

Aim:

To determine the empirical formula of magnesium oxide by combustion of magnesium

Diagram:



Finding the empirical formula of magnesium oxide involves heating magnesium ribbon very strongly in a crucible. A lid is used to trap any smoke (not shown)

Method:

- Measure mass of crucible with lid
- Add sample of magnesium into crucible and measure mass with lid (calculate the mass of the metal by subtracting the mass of empty crucible)
- Strongly heat the crucible over a Bunsen burner for several minutes
- Lift the lid frequently to allow sufficient air into the crucible for the magnesium to fully oxidise without letting magnesium oxide smoke escape
- Continue heating until the mass of crucible remains constant (maximum mass), indicating that the reaction is complete

- Measure the mass of crucible and contents (calculate the mass of metal oxide by subtracting the mass of empty crucible)

Working out the empirical formula:

Mass of metal:

- Subtract mass of crucible from magnesium and the mass of the empty crucible

Mass of oxygen:

- Subtract mass of the magnesium used from the mass of magnesium oxide

Step 1 – Divide each of the two masses by the relative atomic masses of the elements

Step 2 – Simplify the ratio

	magnesium	**oxygen**
Mass	a	b
Moles	a / A_r	b / A_r
	= x	= y
Ratio	x :	y

Step 3 – Represent the ratio into the form 'M_xO_y' E.g, MgO

Your notes

Your notes

Conservation of Mass

Conservation of Mass

- The Law of Conservation of Mass states that no matter is **lost** or **gained** during a chemical reaction.
- Mass is always **conserved**, therefore the total mass of the **reactants** is **equal** to the total mass of the **products**, which is why all chemical equations must be **balanced**
- The sum of the relative atomic/molecular masses of the reactants will be the same as the sum of the relative atomic/molecular masses of the products
- A precipitation reaction is one in which two solutions react to form an **insoluble solid** called a precipitate
- If the reaction flask is closed and no other substance can enter or leave the system, then the total mass of the reaction flask will remain **constant**
- For example, the reaction between calcium chloride and sodium sulfate produces a precipitate of calcium sulfate.
- If carried out in a **closed system** then the mass before and after the reaction will be the same
- The balanced equation is:

$$CaCl_2\ (aq) + Na_2SO_4\ (aq) \longrightarrow CaSO_4\ (s) + 2NaCl\ (aq)$$

Your notes

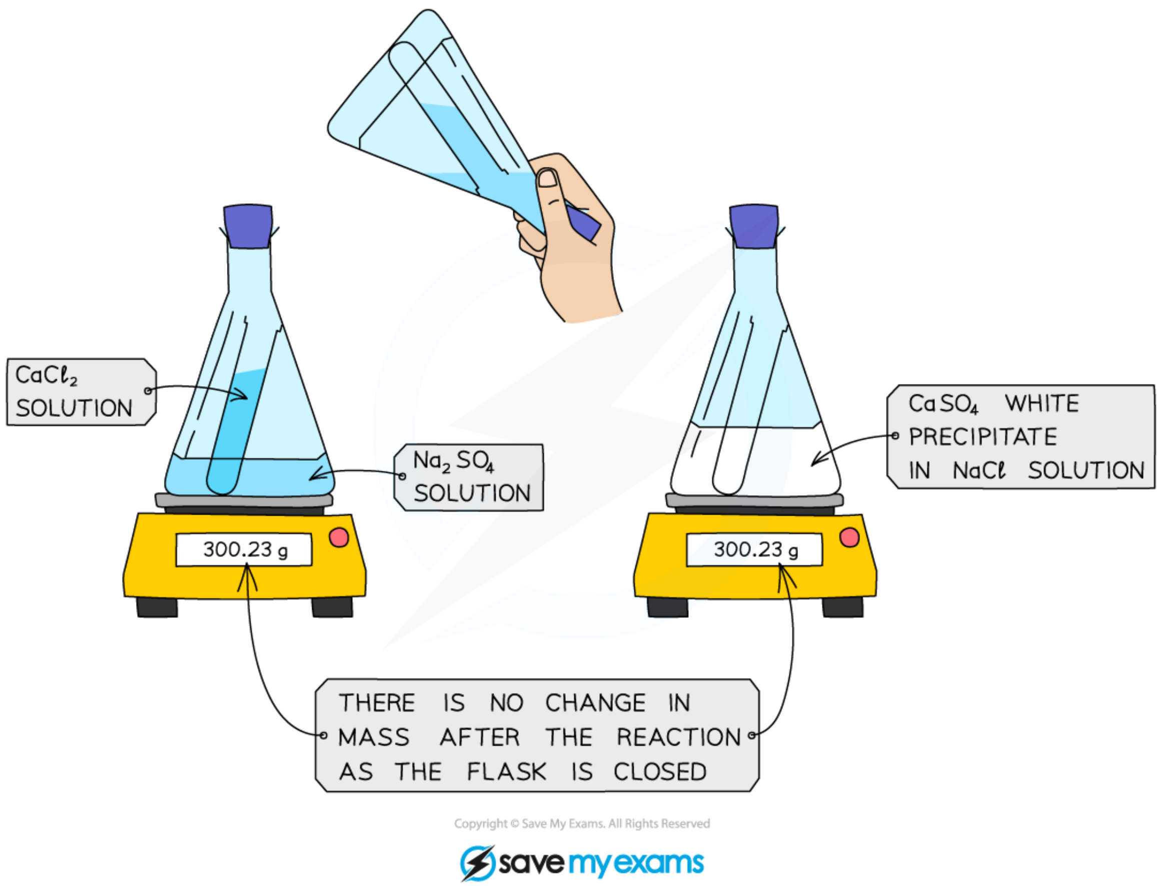


Diagram showing the conservation of mass in a precipitation reaction

- If the reaction flask is open and a **gaseous product** is allowed to escape, then the total mass of the reaction flask will change as product mass is lost when the gas leaves the system
- For example, the reaction between hydrochloric acid and calcium carbonate produces carbon dioxide gas:

$$2HCl\,(aq) + CaCO_3\,(s) \rightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2\,(g)$$

- Mass will be lost from the reaction flask unless it is closed
- If the mass of a reaction flask is found to **increase** then it may be due to one of the reactants being a gas found in the air and all of the products are either solids or liquids

Your notes

Examiner Tips and Tricks

Matter cannot be created or destroyed, so the total amount of matter before and after a reaction is the same. What changes is the chemical and physical properties of the reactants as they transform into products.

Reacting Masses

Your notes

Reacting Masses

- Chemical equations can be used to calculate the **reacting masses** of reactants and products
- The **mass ratio** between the substances is identified using the balanced chemical equation
- The steps are:
  1. Write down the balanced equation for the reaction
  2. Write the relative formula masses under the substances and add the units in the question
  3. Multiply the relative formula masses by the coefficients in the equation
  4. Find the mass of product for 1 g of reactant
  5. Scale up for the mass given in the question

Worked Example

Example 1

Calculate the mass of magnesium oxide that can be made by completely burning 6.0 g of magnesium in oxygen.

Answer:

- **Step 1:** Write the balanced equation

$$2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$$

- **Step 2:** Add RFMs and units

24 g	40 g

- **Step 3:** Multiply by coefficients

2 x 24 = 48 g	2 x 40 = 80 g

- **Step 4:** Cross multiply for 1 g

1 g	80 / 48 = 1.66 g

- **Step 5:** Scale up to mass in question

6 g	**6 x 1.66 = 10 g**

Your notes

Worked Example

Example 2

Calculate the mass of aluminium, in tonnes, that can be produced from 51 tonnes of aluminium oxide.

Answer:

- **Step 1:** Write the balanced equation

$$2Al_2O_3 \longrightarrow 4Al + 3O_2$$

- **Step 2:** Add the RFMs and units

102 tonnes	27 tonnes

- **Step 3:** Multiply by coefficients

2 x 102 = 204 tonnes	4 x 27 = 108 tonnes

- **Step 4:** Cross multiply for 1 tonne

1 tonne	108 / 204 = 0.53 tonne

- **Step 5:** Scale up to mass in question

51 tonnes	**51 x 0.53 = 27 tonnes of Al**

Examiner Tips and Tricks

As long as you are consistent it doesn't matter whether you work in grams or tonnes or any other mass unit as the reacting masses will always be in proportion to the balanced equation.

Percentage Composition

- The percentage by mass of an element in a compound can be calculated using the following equation:

$$\% \text{ mass of an element} = \frac{A_r \times \text{number of atoms of the element}}{M_r \text{ of the compound}} \times 100$$

Worked Example

Calculate the percentage by mass of calcium in calcium carbonate, $CaCO_3$.

Answer:

$A_r Ca = 40$
$A_r C = 12$
$A_r O = 16$

STEP 1: WRITE DOWN THE RELATIVE ATOMIC MASSES OF EACH ELEMENT

$$\%Ca = \frac{1 \times 40}{[40 + 12 + (3 \times 16)]} \times 100$$

$\%Ca$ BY MASS $= 40\%$

STEP 2: INPUT THE VALUES INTO THE EQUATION AND SOLVE

Examiner Tips and Tricks

Don't forget to multiply your answer by 100 in order to convert it to a percentage.

Concentration

Your notes

Calculating Concentration

- A solid substance that dissolves in a liquid is called a **solute**, the liquid is called a **solvent** and the two when mixed together form a **solution**
- Most chemical reactions occur between solutes which are dissolved in solvents, such as water or an organic solvent
- **Concentration** simply refers to the amount of solute there is in a specific volume of the solvent
- The greater the **amount** of solute in a given volume then the **greater** the concentration
- A general formula for concentration is thus:

$$\textit{concentration (g dm}^{-3}\textit{)} = \frac{mass\ of\ solute\ (g)}{volume\ of\ solution\ (dm^3)}$$

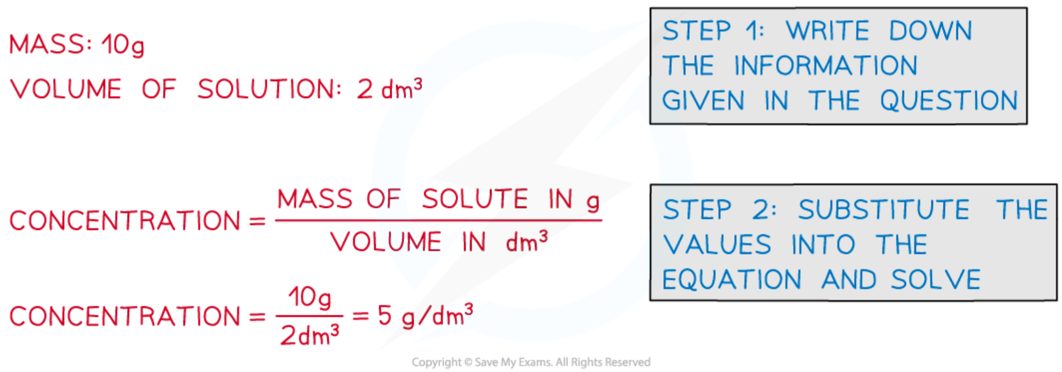
- Concentration can be measured in grams per cubic decimetre
- 1 decimetre cubed (dm^3) = 1000 cm^3
  - 1 decimetre cubed (dm^3) is the same as 1 litre
- You may be given data in a question which needs to be converted from cm^3 to dm^3 or the other way around
  - To go from cm^3 to dm^3 :
    - Divide by 1000
  - To go from dm^3 to cm^3 :
    - Multiply by 1000

Worked Example

A student dissolved 10 g of sodium hydroxide, NaOH, in 2 dm^3 of distilled water. Calculate the concentration of the solution.

Answer:

Your notes



Examiner Tips and Tricks

Be careful when doing volume unit conversions as it is easy to multiply instead of dividing by 1000 and vice-versa. Always ask yourself – is the result going to be a bigger or smaller number than I started with? Do I get more or fewer cubic decimetres when I convert from cubic centimetres?

Your notes

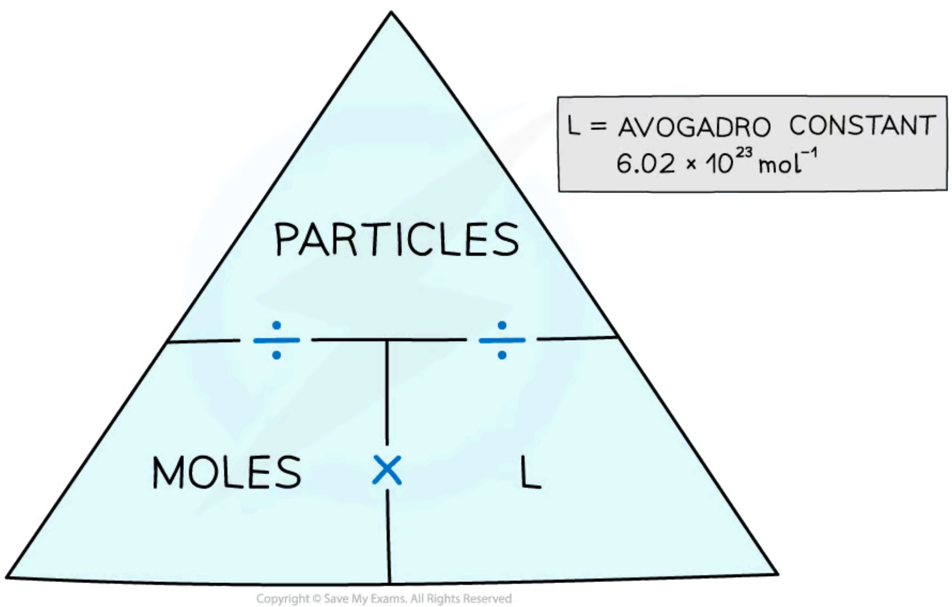
The Mole

The Mole

- Chemical amounts are measured in moles
- The symbol for the unit mole is **mol**
- One mole of a substance contains the same number of the stated particles, atoms, molecules, or ions as one mole of any other substance
- The number of atoms, molecules or ions in a mole (1 mol) of a given substance is the Avogadro constant. The value of the Avogadro constant is 6.02×10^{23} per mole

For example:

- One mole of sodium (Na) contains 6.02×10^{23} **atoms** of sodium
- One mole of hydrogen (H_2) contains 6.02×10^{23} **molecules** of hydrogen
- One mole of sodium chloride (Na*Cl*) contains 6.02×10^{23} **formula units** of sodium chloride



The formula triangle showing the relationship between moles, particles and the Avogadro constant

Your notes

Worked Example

Particles from Moles: How many hydrogen atoms are in 0.010 moles of CH_3CHO?

Answer:

- There are 4 H atoms in 1 molecule of CH_3CHO
- So, there are 0.040 moles of H atoms in 0.010 moles of CH_3CHO
- The number of H atoms is the **amount in moles x L**
- This comes to 0.040 x (6.02 x 10^{23}) = **2.4 x 10^{22} atoms**

Worked Example

Moles from Particles: How many moles of hydrogen atoms are in 3.612 x 10^{23} molecules of H_2O_2?

Answer:

- In 3.612 x 10^{23} molecules of H_2O_2 there are 2 x (3.612 x 10^{23}) atoms of H
- So, there are 7.224 x 10^{23} atoms of H
- The number of moles of H atoms is the **number of particles ÷ *L***
- This comes to 7.224 x 10^{23} ÷ (6.02 x 10^{23}) = **1.20 moles of H atoms**

Linking the Mole and Relative Atomic Mass

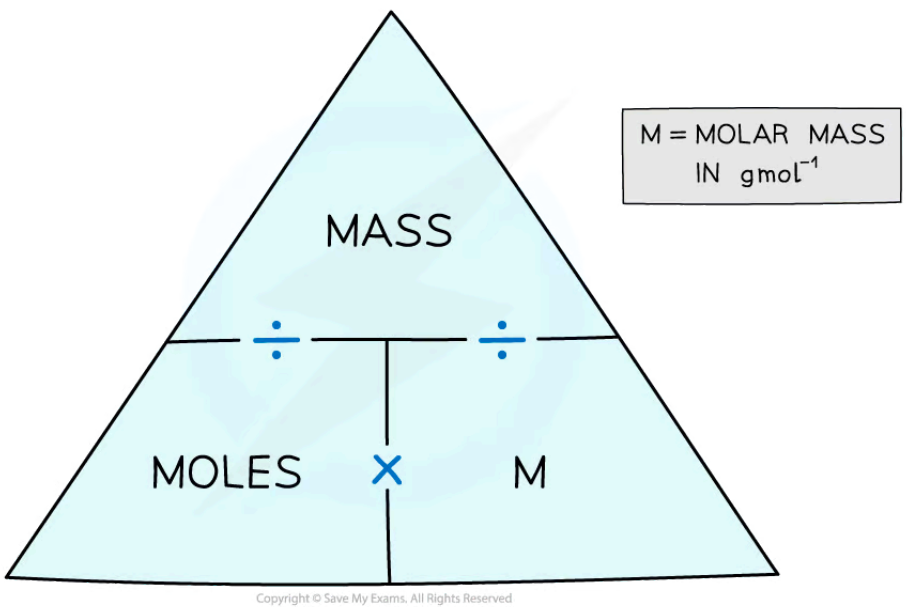
Linking the Mole and Relative Atomic Mass

- One mole of any element is equal to the **relative atomic mass** of that element in **grams** or for a compound the **relative formula mass** in grams
- This is called the **molar mass**
- If you had 6.02 x 10^{23} atoms of carbon in your hand, that number of carbon atoms would have a mass of 12 g (because the A_r of carbon is 12)
- So one mole of helium atoms would have a mass of 4 g (A_r of He is 4), one mole of lithium would have a mass of 7 g (A_r of Li is 7) and so on
- To find the mass of one mole of a compound, we add up the relative atomic masses
  - So one mole of water would have a mass of (2 x 1) + 16 = 18 g
  - So one carbon atom has the same mass as 12 hydrogen atoms

Moles and Molar Mass

Your notes

- Although elements and chemicals react with each other in **molar ratios**, in the laboratory we use digital balances and grams to measure quantities of chemicals as it is impractical to try and measure out moles
- Therefore we have to be able to convert between moles and grams
- We can use the following formula to convert between moles, mass in grams and the molar mass:



Formula triangle for moles, mass and molar mass

Worked Example

Mass from Moles: What is the mass of 0.250 moles of zinc?

Answer:

- From the periodic table the relative atomic mass of Zn is 65.38
- So, the molar mass is 65.38 g mol^{-1}
- The mass is calculated by **moles x molar mass**
- This comes to 0.250 mol x 65.38 g mol^{-1} = **16.3 g**

Your notes

Worked Example

Moles from Mass: How many moles are in 2.64 g of sucrose, $C_{12}H_{22}O_{11}$ (M_r = 342.3)?

Answer:

- The molar mass of sucrose is 342.3 g mol^{-1}
- The number of moles is found by **mass ÷ molar mass**
- This comes to 2.64 g ÷ 342.3 g mol^{-1} = **7.71×10^{-3} mol**

Examiner Tips and Tricks

Always show your workings in calculations as its easier to check for errors and you may pick up credit if you get the final answer wrong.

Your notes

Limiting Reagents

Limiting Reagents

- A chemical reaction stops when one of the reagents is used up
- The reagent that is used up first is the **limiting reagent**, as it limits the duration and hence the amount of product that a reaction can produce
- The amount of product is therefore **directly proportional** to the amount of the limiting reagent added at the beginning of a reaction
- The limiting reagent is the reactant which is **not present in excess** in a reaction
- In order to determine which reactant is the limiting reagent in a reaction, we have to consider the ratios of each reactant in the balanced equation
- When performing reacting mass calculations, the limiting reagent is always the number that should be used as it indicates the maximum possible amount of product
- The steps are:
    1. Write the balanced equation for the reaction
    2. Calculate the moles of each reactant
    3. Compare the moles & deduce the limiting reactant

Worked Example

9.2 g of sodium is reacted with 8.0 g of sulfur to produce sodium sulfide, Na_2S. Which reactant is in excess and which is the limiting reactant?

Answer:

- **Step 1:** Write the balanced equation and determine the molar ratio
  $2Na + S \rightarrow Na_2S$ so the molar ratios is 2 : 1
- **Step 2:** Calculate the moles of each reactant
  Moles = Mass ÷ A_r

  Moles Na = 9.2/23 = 0.40

  Moles S = 8.0/32 = 0.25

- **Step 3:** Compare the moles

To react completely 0.40 moles of Na requires 0.20 moles of S and since there are 0.25 moles of S, then S is in excess. **Na is therefore the limiting reactant.**

Your notes

Examiner Tips and Tricks

An easy way to determine the limiting reactant is to find the moles of each substance and divide the moles by the coefficient in the equation. The **lowest** number resulting is the **limiting reactant**

- In the example above:
  - divide 0.40 moles of Na by 2, giving 0.20
  - divide 0.25 moles of S by 1, giving 0.25, so Na is limiting

Deducing Stoichiometry

Your notes

Writing & Balancing Equations

Nothing created – nothing destroyed

- New substances are made during chemical reactions
    - However, the same atoms are always present before and after reaction
    - They have just joined up in different ways
    - Atoms cannot be created or destroyed, so if they exist in the reactants then they absolutely must be in the products!
- Because of this the total mass of reactants is always equal to the total mass of products
- This idea is known as the **Law of Conservation of Mass**

Conservation of Mass

- The **Law of Conservation of Mass** enables us to balance chemical equations, since no atoms can be lost or created
- You should be able to:
    - Write word equations for reactions outlined in these notes
    - Write formulae and balanced chemical equations for the reactions in these notes

Word Equations

- These show the **reactants** and **products** of a chemical reaction using their full chemical names
- The reactants are those substances on the **left-hand side** of the arrow and can be thought of as the chemical **ingredients** of the reaction
- They react with each other and form new substances
- The products are the new substances which are on the **right-hand side** of the arrow
- The arrow (which is spoken as "*goes to*" 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 produces sodium chloride (common table salt) and water:

sodium hydroxide + hydrochloric acid ⟶ sodium chloride + water

Your notes

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
- During chemical reactions as 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

Worked Example

Balance the following equation:

magnesium + oxygen ⟶ magnesium oxide

Answer:

Your notes

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

$$Mg + O_2 \rightarrow MgO$$

- **Step 2:** Count the numbers 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

$$2Mg + O_2 \rightarrow 2MgO$$

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

$$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$$

Ionic Equations

Higher Only

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, and not the spectator ions

Worked Example

1. Balance the following equation

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

Your notes

2. Write down the ionic equation for the above reaction

Answer 1:

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

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- **Step 2:** Count the numbers 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

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

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

Deducing Stoichiometry

Higher Only

- Stoichiometry refers to the numbers **in front** of the reactants and products in an equation, which must be adjusted to make sure that the equation is balanced
- These numbers are called **coefficients** (or multipliers) and if we know the masses of reactants and products, the balanced chemical equation for a given reaction can be found by determining the coefficients

Your notes

- First, convert the masses of each reactant and product in to moles by dividing by the molar masses using the periodic table
- If the result yields **uneven numbers**, then multiply all of the numbers by the same number, to find the **smallest whole number** for the coefficient of each species
  - For example, if the resulting numbers initially were 1, 2 and 2.5, then you would multiply all of the numbers by 2, to give the whole numbers 2, 4 and 5
- Then, use the molar ratio to write out the balanced equation

Worked Example

64 g of methanol, CH_3OH, reacts with 96 g of oxygen gas to produce 88 g of carbon dioxide and 72 g of water. Deduce the balanced equation for the reaction.(C = 12, H = 1, O = 16).

Answer:

- Calculate the molar masses of the substances in the equation

CH_3OH = 32 g mol^{-1} O_2 = 32 g mol^{-1}

CO2 = 44 g mol^{-1} H_2O = 18 g mol^{-1}

- Divide the masses present by the molar mass to obtain the number of moles

CH_3OH = 64 g ÷ 32 g mol^{-1} = 2 mol

O_2 = 96 g ÷ 32 g mol^{-1} = 3 mol

CO_2 = 88 g ÷ 44 g mol^{-1} = 2 mol

H_2O = 72 g ÷ 18 gmol^{-1} = 4 mol

- The mole ratios are the same as the coefficients in the balanced equation

$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$

Examiner Tips and Tricks

The molar ratio of a balanced equation gives you the ratio of the amounts of each substance in the reaction.