SL IB Chemistry

Your notes

How Much? The Amount of Chemical Change

Contents

Your notes

Balancing Equations

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

Your notes

Worked example

Balance the following equation:

Magnesium + Oxygen → Magnesium oxide

Answer:

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

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

Worked example

1. Write a balanced symbol equation for the following equation

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

Answer

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

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

Your notes

Your notes

Reacting Mass Calculations

Reacting Mass Calculations

- The number of moles of a substance can be found by using the following equation:

$$\text{number of moles} = \frac{\text{mass of substance in grams}}{\text{molar mass } \left(\text{g mol}^{-1}\right)}$$

- It is important to be clear about the type of particle you are referring to when dealing with moles
  - E.g. 1 mole of CaF_2 contains one mole of CaF_2 **formula units**, but one mole of Ca^{2+} and two moles of F^- **ions**

Reacting masses

- The **masses** of reactants are useful to determine how much of the reactants **exactly** react with each other to prevent waste
- To calculate the reacting masses, the chemical equation is required
- This equation shows the ratio of moles of all the reactants and products, also called the **stoichiometry**, of the reaction
- To find the mass of products formed in a reaction the following pieces of information are needed:
  - The mass of the reactants
  - The molar mass of the reactants
  - The balanced equation

Your notes

Worked example

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

magnesium (s) + oxygen (g) → magnesium oxide (s)

Answer:

Step 1: The symbol equation is:

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

Step 2: The relative atomic masses are:

magnesium : 24.31 oxygen : 16.00

Step 3: Calculate the moles of magnesium used in reaction

$$\text{number of moles} = \frac{6.0 \text{ g}}{24.31 \text{ g mol}^{-1}} = 0.25 \text{ mol}$$

Step 4: Find the ratio of magnesium to magnesium oxide using the balanced chemical equation

	Magnesium	Magnesium oxide
Mol	2	2
Ratio	1	1
Change in mol	-0.25	+0.25

Therefore, **0.25** mol of MgO is formed

Step 5: Find the mass of magnesium oxide

mass = mol x M

mass = 0.25 mol x 40.31 g mol^{-1}

mass = 10.08 g

Therefore, **mass of magnesium oxide produced is 10 g** (2 sig figs)

Your notes

Worked example

Calculate the mass of aluminium, in tonnes, that can be produced from 51 tonnes of aluminium oxide. The equation for the reaction is:

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

Answer:

Step 1: Calculate the moles if aluminium oxide used

mass of Al_2O_3 in g = 51 x 10^6 = 51,000,000 g

moles = $\frac{51,000,000 \text{ g}}{101.96 \text{ g mol}^{-1}}$ = 500,196.16 mol

Step 2: Find the ratio of Al_2O_3 to Al using the molar ratio from the balanced equation

$2Al_2O_3$: 4Al

Ratio is thus 1 : 2

So 500,196.16 mol moles of Al_2O_3 produces 100,0392.31 moles of Al

Step 3: Calculate mass of Al

mass = Moles x M_r

mass = 1,000,392.31 mol x 26.98 g mol^{-1} = 26,990,584.54 g

Step 4: Convert mass from grams to tonnes

$\frac{26,990,584.54 \text{ g}}{10^6}$ = 26.99 tonnes

Examiner Tip

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.

Your notes

Avogadro's Law & Molar Volume of Gas

Avogadro's Law & Molar Volume of Gas

Volumes of gases

- In 1811 the Italian scientist Amedeo **Avogadro** developed a theory about the volume of gases
- **Avogadro's law** (also called **Avogadro's hypothesis)** enables the mole ratio of reacting gases to be determined from volumes of the gases
- **Avogadro** deduced that equal volumes of gases must contain the same number of molecules
- At standard temperature and pressure(**STP**) **one mole** of any gas has a volume of **22.7 dm^3**
- The units are normally written as **dm^3 mol^{-1}**(since it is 'per mole')
- The conditions of **STP** are
  - a temperature of **0° C** (**273 K**)
  - a pressure of **100 kPa**

Stoichiometric relationships

- The stoichiometry of a reaction and **Avogadro's Law** can be used to deduce the **exact volumes** of gaseous reactants and products
  - Eg. in the **combustion** of 50 cm^3 of propane, the volume of oxygen needed is (5 x 50) 250 cm^3, and (3 x 50) 150 cm^3 of carbon dioxide is formed, using the ratio of propane: oxygen: carbon dioxide, which is 1: 5: 3 respectively, as seen in the equation

$$C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(l)$$

- Remember that if the gas volumes are not in the same ratio as the coefficients then the amount of product is determined by the limiting reactant so it is essential to identify it first

Your notes

Worked example

What is the total volume of gases remaining when 70 cm^3 of ammonia is combusted completely with 50 cm^3 of oxygen according to the equation shown?

$$4NH_3\ (g) + 5O_2\ (g) \rightarrow 4NO\ (g) + 6H_2O\ (l)$$

Answer:

Step 1: From the equation deduce the molar ratio of the gases, which is $NH_3 : O_2 : NO$ or 4:5:4 (water is not included as it is in the liquid state)

Step 2: We can see that oxygen will run out first (the **limiting reactant**) and so 50 cm^3 of O_2 requires 4/5 x 50 cm^3 of NH_3 to react = 40 cm^3

Step 3: Using Avogadro's Law, we can say 40 cm^3 of NO will be produced

Step 4: There will be 70–40 = 30 cm^3 of NH_3 left over

Therefore **the total remaining volume will be 40 + 30 = 70 cm^3 of gases**

Examiner Tip

Since gas volumes work in the same way as moles, we can use the '**lowest is limiting**' technique in limiting reactant problems involving gas volumes. This can be handy if you are unable to spot which gas reactant is going to run out first. Divide the volumes of the gases by the cofficients and whichever gives the lowest number is the **limiting reactant**

- E.g. in the previous problem we can see that
  - For NH_3 70/4 gives 17.5
  - For O_2 50/5 gives 10, so **oxygen is limiting**

Molar Gas Volume

- The **molar gas volume** of 22.7 $dm^3\ mol^{-1}$ can be used to find:
  - The volume of a given number of moles of gas:

    volume of gas (dm^3) = amount of gas (mol) x 22.7 $dm^3\ mol^{-1}$

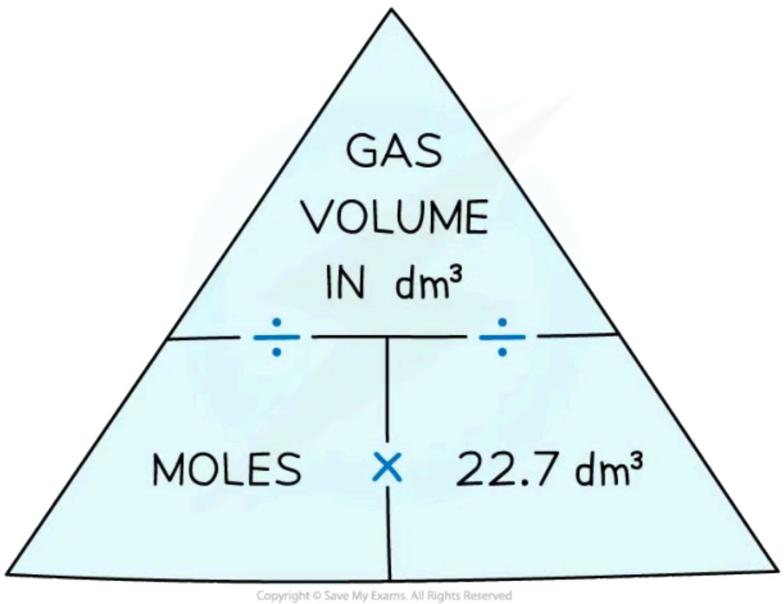
  - The number of moles of a given volume of gas:

$$\text{amount of gas (moles)} = \frac{\text{volumes of gas in dm}^3}{22.7\ \text{dm}^3\ \text{mol}^{-1}}$$

- The relationships can be expressed using a formula triangle

Gas formula triangle

Your notes



To use the gas formula triangle cover the one you want to find out about with your finger and follow the instructions

Worked example

What is the volume occupied by 3.0 moles of hydrogen at stp?

Answer:

volume of gas (dm^3) = amount of gas (mol) x 22.7 dm^3 mol^{-1}

3.0 mol x 22.7 dm^3 mol^{-1}= **68 dm^3**

Your notes

Worked example

How many moles are in the following volumes of gases?

1. 7.2 dm^3 of carbon monoxide
2. 960 cm^3 of sulfur dioxide

Answer 1:

Step 1: Use the formula

$$\text{amount of gas (moles)} = \frac{\text{volumes of gas in dm}^3}{22.7\ \text{dm}^3\ \text{mol}^{-1}}$$

$$\text{amount of gas (moles)} = \frac{7.2\ \text{dm}^3}{22.7\ \text{dm}^3\ \text{mol}^{-1}} = 0.32\ \text{mol}$$

Answer 2

Step 1: Convert the volume from cm^3 to dm^3

$$\frac{960}{1000} = 0.960\ \text{dm}^3$$

Step 2: Use the formula

$$\text{amount of gas (moles)} = \frac{0.960\ \text{dm}^3}{22.7\ \text{dm}^3\ \text{mol}^{-1}} = 4.22 \times 10^{-2}\ \text{mol}$$

Concentration Calculations

Your notes

Concentration Calculations

Titrations

- **Volumetric analysis** is a process that uses the volume and concentration of one chemical reactant (**a standard solution**) to determine the concentration of another unknown solution
- The technique most commonly used is a **titration**
- The volumes are measured using two precise pieces of equipment, a **volumetric** or **graduated pipette** and a **burette**
- The steps in a titration are:
  - Measuring a known volume (usually 20 or 25 cm^3) of one of the solutions with a **volumetric** or **graduated pipette** and placing it into a **conical flask**
  - The other solution is placed in the **burette**
  - A few drops of the **indicator** are added
  - The tap on the **burette** is carefully opened and the solution added, portion by portion, to the **conical flask** until the **indicator** just changes colour
  - Multiple trials are carried out until **concordant** results are obtained

Calculating concentration

- Concentration calculations involve bringing together the skills and knowledge you have acquired in molar concentration and applying them to problem solving
- You should be able to easily convert between moles, mass, concentrations and volumes (of solutions and gases)
- The four steps involved in problem solving are:
  - write the balanced equation for the reaction
  - determine the mass/ moles/ concentration/ volume of the of the substance(s) you know about
  - use the balanced equation to deduce the mole ratios of the substances present
  - calculate the mass/ moles/ concentration/ volume of the of the unknown substance(s)

Your notes

Worked example

25.0 cm^3 of 0.050 mol dm^{-3} sodium carbonate was completely neutralised by 20.0 cm^3 of dilute hydrochloric acid. Calculate the concentration in mol dm^{-3} of the hydrochloric acid.

Answer:

Step 1: Write the balanced equation for the reaction

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

Step 2: Determine the moles of the known substance, in this case sodium carbonate. Don't forget to divide the volume by 1000 to convert cm^3 to dm^3

moles = volume x concentration

amount (Na_2CO_3) = 0.0250 dm^3 x 0.050 mol dm^{-3} = 0.00125 mol

Step 3: Use the balanced equation to deduce the mole ratio of sodium carbonate to hydrochloric acid:

1 mol of Na_2CO_3 reacts with 2 mol of HCl, so the mole ratio is 1 : 2

Therefore 0.00125 moles of Na_2CO_3 react with 0.00250 moles of HCl

Step 4: Calculate the concentration of the unknown substance, hydrochloric acid

$$\text{concentration} = \frac{\text{moles}}{\text{volume}}$$

$$\text{concentration (HCl)} = \frac{0.00250 \text{ mol}}{0.0200 \text{ dm}^3} = 0.125 \text{ mol dm}^{-3}$$

Your notes

Worked example

Calculate the volume of hydrochloric acid of concentration 1.0 mol dm^{-3} that is required to react completely with 2.5 g of calcium carbonate.

Answer:

Step 1: Write the balanced equation for the reaction

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

Step 2: Determine the moles of the known substance, calcium carbonate

Amount of $CaCO_3$ = $\frac{2.5 \text{ g}}{100.09 \text{ g mol}^{-1}} = 0.025 \text{ mol}$

Step 3: Use the balanced equation to deduce the mole ratio of calcium carbonate to hydrochloric acid:

1 mol of $CaCO_3$ requires 2 mol of HCl

So 0.025 mol of $CaCO_3$ requires 0.050 mol of HCl

Step 4: Calculate the volume of HCl required

Volume of HCl = $\frac{\text{moles}}{\text{concentration}} = \frac{0.050 \text{ mol}}{1.0 \text{ mol dm}^{-3}} = 0.050 \text{ dm}^3$

Examiner Tip

When performing titration calculations using **monoprotic** acids (meaning one H^+) such as HCl, the number of moles of the acid and alkali will be the same. This allows you to use the relationship

$$C_1V_1 = C_2V_2$$

where C_1 and V_1 are the concentration and volume of the acid and C_2 and V_2 are the concentration and volume of the alkali. There is no need to convert the units of volume to dm^3 as this is a ratio.Simply re-arrange the formula to solve for the unknown quantity.

Your notes

Worked example

A 0.675 g sample of a solid acid, HA, was dissolved in distilled water and made up to 100.0 cm^3 in a volumetric flask. 25.0 cm^3 of this solution was titrated against 0.100 mol dm^{-3} NaOH solution and 12.1 cm^3 were required for complete reaction. Determine the molar mass of the acid.

Answer:

Step 1: Write the equation for the reaction

$$HA\ (aq) + NaOH\ (aq) \rightarrow NaA\ (aq) + H_2O\ (l)$$

Step 2: Calculate the number of moles of the NaOH

$$n(\text{NaOH})_{\text{sample}} = \left(\frac{12.1\ \text{cm}^3}{1000}\right)\text{dm}^3 \times 0.100\ \text{mol dm}^{-3} = 1.21 \times 10^{-3}\ \text{mol}$$

Step 3: Deduce the number of moles of the acid

Since the acid is monoprotic the number of moles of HA is also 1.21 x 10^{-3} mol

This is present in 25.0 cm^3 of the solution

Step 4: Scale up to find the amount in the original solution

$$\text{n(NaOH)}_{\text{original}} = \frac{1.21 \times 10^{-3}\ \text{mol} \times 100.0\ \text{cm}^3}{25.0\ \text{cm}^3} = 4.84 \times 10^{-3}\ \text{mol}$$

Step 5: Calculate the molar mass

$$\text{moles} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{molar mass} = \frac{\text{mass}}{\text{moles}} = \frac{0.675\ \text{g}}{4.84 \times 10^{-3}} = 139\ \text{g mol}^{-1}$$

Back titration

- A **back titration** is a common technique used to find the **concentration** or amount of an unknown substance indirectly
- The principle is to carry out a reaction with the unknown substance and an **excess** of a further reactant such as an acid or an alkali
- The **excess** reactant, after reaction, is then analysed by **titration** and the **mole ratios** are used to deduce the **moles** or **concentration** of the original substance being analysed

Your notes

Worked example

The percentage by mass of calcium carbonate, $CaCO_3$, in a sample of marble was determined by adding excess hydrochloric acid to ensure that all the calcium carbonate had reacted. The excess acid left was then titrated with aqueous sodium hydroxide. A student added 27.20 cm^3 of 0.200 mol dm^{-3} HCl to 0.188 g of marble. The excess acid required 23.80 cm^3 of 0.100 mol dm^{-3} NaOH for neutralisation. Calculate the percentage of calcium carbonate in the marble.

Answer:

Step 1: Write the equation for the titration reaction:

$$\textbf{HCl (aq) + NaOH (aq)} \rightarrow \textbf{NaCl (aq) + H}_2\textbf{O (l)}$$

Step 2: Calculate the number of moles of the NaOH

n(NaOH) = 0.02380 dm^3 x 0.100 mol dm^{-3} = 2.380 x 10^{-3} mol

Step 3: Deduce the number of moles of the excess acid

Since the reacting ratio is 1:1 the number of moles of HCl is also 2.380 x 10^{-3} mol

Step 4: Find the amount of HCl in the original solution and then the amount reacted

$n(HCl)_{original}$ = 0.02720 dm^3 x 0.200 mol dm^{-3} = 5.440 x 10^{-3} mol

$n(HCl)_{reacted}$ = 5.440 x 10^{-3} mol – 2.380 x 10^{-3} mol = 3.060 x 10^{-3} mol

Step 5: Write the equation for the reaction with the calcium carbonate

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

Step 6: Deduce the number of moles of the calcium carbonate that reacted

Since the reacting ratio is 2:1 the number of moles of $CaCO_3$ is (3.060 x 10^{-3} mol) ÷ 2

$n(CaCO_3)$ = 1.530 x 10^{-3} mol

Step 7: Calculate the mass of calcium carbonate in the sample of marble

mass = moles x molar mass = 1.530 x 10^{-3} mol x 100.09 g mol^{-1} = 0.1531g

Step 8: Calculate the percentage of calcium carbonate in the marble

Percentage of $CaCO_3$ in marble = $\frac{0.1531 \times 100}{0.188} = 81.5\%$

Your notes

Examiner Tip

Rounding off when you take averages. When you have an average of burette readings that comes to three decimal places, e.g.(23.20 cm^3 + 23.25 cm^3) ÷ 2 = 23.225 cm^3

You CANNOT show more than two decimal places because that would make the average more precise than the readings.

To manage this situation you need to follow a simple rule. If the last digit is between a 5 and 9 then you round up; if the digit is between 0 and 4 you round down. So in this case the value recorded would be 23.23 cm^3

Your notes

Limiting & Excess Reactants

Limiting & Excess Reactants

Excess & limiting reactants

- Sometimes, there is an **excess** of one or more of the reactants (**excess reactant**)
- The reactant which is not in excess is called the **limiting reactant**
- To determine which reactant is limiting:
  - The number of moles of the reactants should be calculated
  - The ratio of the reactants shown in the equation should be taken into account eg:

$$C + 2H_2 \rightarrow CH_4$$

- **What is limiting when 10 mol of carbon are reacted with 3 mol of hydrogen?**
  - Hydrogen is the **limiting reactant** and since the ratio of C : H_2 is 1:2 only 1.5 mol of C will react with 3 mol of H_2

Worked example

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

Answer:

Step 1: Calculate the moles of each reactant

$$\text{number of moles (Na)} = \frac{9.2\ \text{g}}{22.99\ \text{g mol}^{-1}} = 0.40\ \text{mol}$$

$$\text{number of moles (S)} = \frac{8.0\ \text{g}}{32.07\ \text{g mol}^{-1}} = 0.25\ \text{mol}$$

Step 2: Write the balanced equation and determine the coefficients

$$2Na + S \rightarrow Na_2S$$

Step 3: Divide the moles by the coefficient and determine the limiting reagent

- divide 0.40 moles of Na by 2, giving 0.20 - **lowest**
- divide 0.25 moles of S by 1, giving 0.25

Therefore, **sodium is limiting** and **sulfur is in excess**

Examiner Tip

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 10 moles of C by 1, giving 10
  - divide 3 moles of H by 2, giving 1.5, so hydrogen is limiting

Your notes

Your notes

Percentage Yield Calculations

Percentage Yield Calculations

Percentage yield

- In a lot of reactions, not all reactants react to form products which can be due to several factors:
  - Other reactions take place simultaneously
  - The reaction does not go to **completion**
  - Products are **lost** during separation and purification

- The **percentage yield** shows how much of a particular product you get from the reactants compared to the maximum theoretical amount that you can get:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

  - The **actual yield** is the number of moles or mass of product obtained **experimentally**
  - The **theoretical yield** is the number of moles or mass obtained by a reacting mass calculation

Your notes

Worked example

In an experiment to displace copper from copper(II) sulfate, 6.5 g of zinc was added to an excess of copper(II) sulfate solution. The resulting copper was filtered off, washed and dried. The mass of copper obtained was 4.8 g.

Calculate the percentage yield of copper.

Answer:

Step 1: The balanced symbol equation is:

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

Step 2: Calculate the amount of zinc reacted in moles

$$\text{number of moles} = \frac{6.5\ \text{g}}{65.4\ \text{g mol}^{-1}} = 0.10\ \text{mol}$$

Step 3: Calculate the maximum amount of copper that could be formed from the molar ratio:

Since the ratio of Zn (s) to Cu (s) is 1:1 a maximum of 0.10 moles can be produced

Step 4: Calculate the maximum mass of copper that could be formed (theoretical yield)

mass = mol x M

mass = 0.10 mol x 63.55 g mol^{-1}

mass = 6.4 g (2 sig figs)

Step 5: Calculate the percentage yield of copper

$$\text{percentage yield} = \frac{4.8\ \text{g}}{6.4\ \text{g}} \times 100 = 75\%$$

Your notes

Atom Economy

Atom Economy

- The atom economy of a reaction shows how many of the atoms used in the reaction become the desired product
  - The rest of the atoms or mass is wasted
- It is found directly from the balanced equation by calculating the M_r of the desired product

$$\text{Atom economy} = \frac{\text{molecular mass of desired product}}{\text{sum of molecular masses of all reactants}} \times 100$$

- In addition reactions, the atom economy will always be 100%, because all of the atoms are used to make the desired product
  - Whenever there is only one product, the atom economy will always be 100%
- For example, in the reaction between ethene and bromine:

$$CH_2{=}CH_2 + Br_2 \rightarrow CH_2BrCH_2Br$$

- The atom economy could also be calculated using mass, instead or M_r
- In this case, you would divide the mass of the desired product formed by the total mass of all reactants, and then multiply by 100
- Efficient processes have high atom economies and are important to sustainable development
  - They use fewer resources
  - Create less waste
- As well as atom economy and percentage yield there are other factors that can be used to gauge the efficiency of a chemical process
  - Rate
  - Quantities of reagents such as catalysts and solvents
  - Energy uses
  - Economic efficiency

Worked example

Ethanol can be produced by various reactions, such as:

Hydration of ethene: $C_2H_4 + H_2O \rightarrow C_2H_5OH$
Substitution of bromoethane: $C_2H_5Br + NaOH \rightarrow C_2H_5OH + NaBr$

Explain which reaction has a higher atom economy.

Answer:

Hydration of ethene has a higher atom economy (of 100%) because all of the reactants are converted into products, whereas the substitution of bromoethane produces NaBr as a waste product

Worked example

Quantitative atom economy

The blast furnace uses carbon monoxide to reduce iron(III) oxide to iron.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Calculate the atom economy for this reaction, assuming that iron is the desired product.

(A_r / M_r data: Fe_2O_3 = 159.6, CO = 28.0, Fe = 55.8, CO_2 = 44.0)

Answer:

Step 1: Write the equation:

$$\text{Atom economy} = \frac{\text{molecular mass of desired product}}{\text{sum of molecular masses of ALL reactants}} \times 100$$

Step 2: Substitute values and evaluate:

$$\text{Atom economy} = \frac{2 \times 55.8}{159.6 + (3 \times 28.0)} \times 100 = 45.8\%$$

Examiner Tip

Careful: Sometimes a question may ask you to show your working when calculating atom economy.

In this case, even if it is an addition reaction and it is obvious that the atom economy is 100%, you will still need to show your working.

Your notes