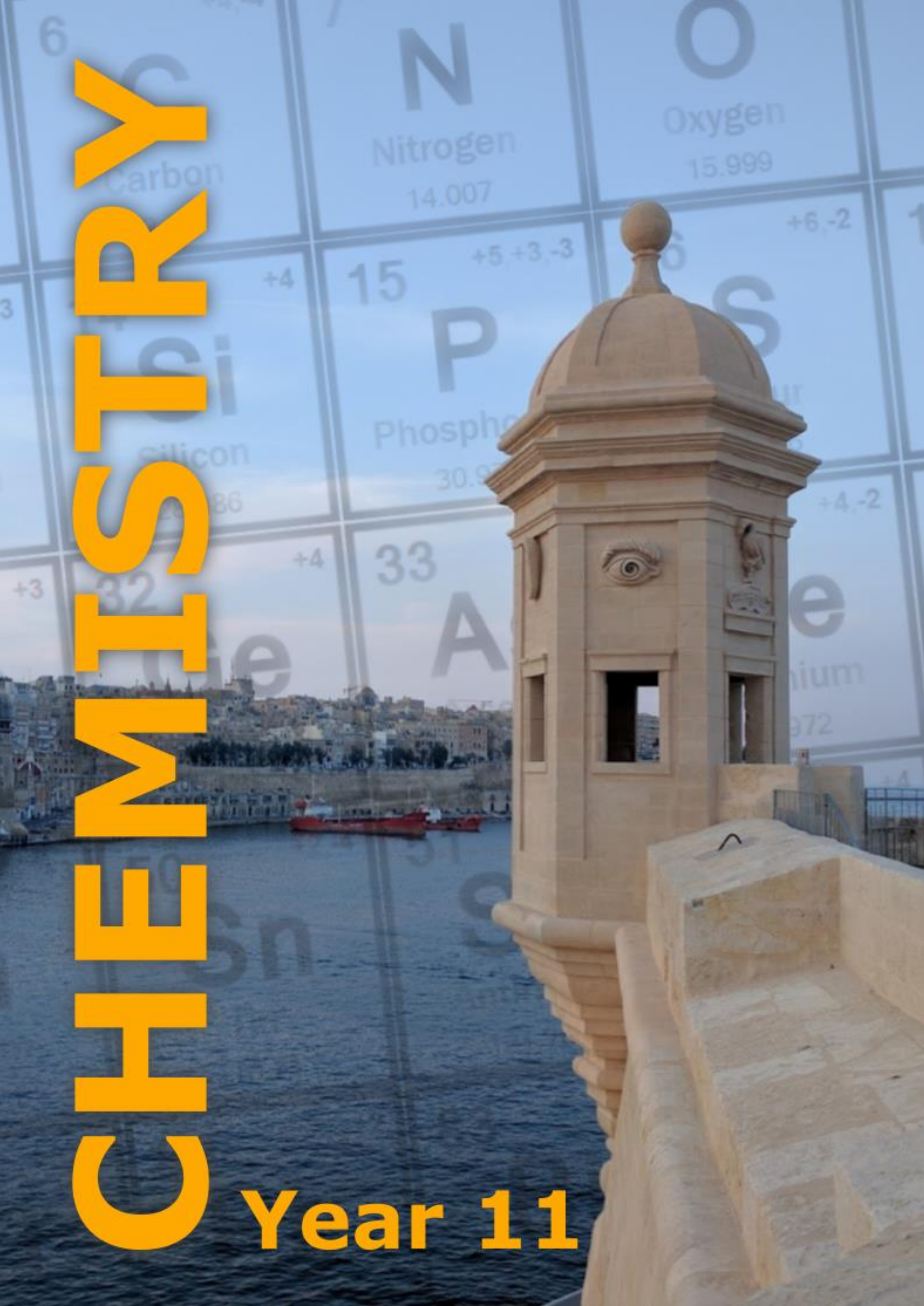
CHEMISTRY

Year 11



This textbook will be available in digital format (PDF) only. It is intended to be used by secondary school students and teachers in Malta and Gozo. It leads students to sit for the SEC06 Chemistry examination from 2025 onwards.

Version history

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HOW TO USE THIS E-BOOK.

This e-book is closely linked to the SEC 06 Chemistry syllabi from 2025 onwards which can be downloaded from https://www.um.edu.mt/matsec/syllabi/.

Each chapter represents the work expected in one learning outcome. Each chapter, except for chapter one, starts with a graphic that indicates the subject focus, which the learning outcome is part of.

Chapters are split into sections that deal with groups of related assessment criteria. These assessment criteria are declared so that students can focus on what needs to be learned. Each section is accompanied by a series of questions that are related specifically to the content in that section.

Hyperlinks link to content that is online and serve as further reading. Online content is not examinable however students are encouraged to follow these links to further their understanding of the role of chemistry in the real world.

An internet connection is required to access video clips, animations, and online content. When using the digital version of this e-book, QR codes can be clicked to access the online resources.

When using this e-book on a computer use the navigation feature provided by your PDF viewer.

Throughout this e-book you will find one of five types of textboxes shown below:

This textbox shows **exercises** related to content that students can work on their own or with their teachers.

This textbox shows **snippets** of information that are not meant to be studied for your examination, but which give extra information related to discoverers, discoveries, and further relevant content.

This textbox shows a QR code which links to an **online video clip**. Click on the QR code when using the digital version to access the resource.

This textbox shows a QR code which links to an **animation**. Click on the QR code when using the digital version to access the resource.

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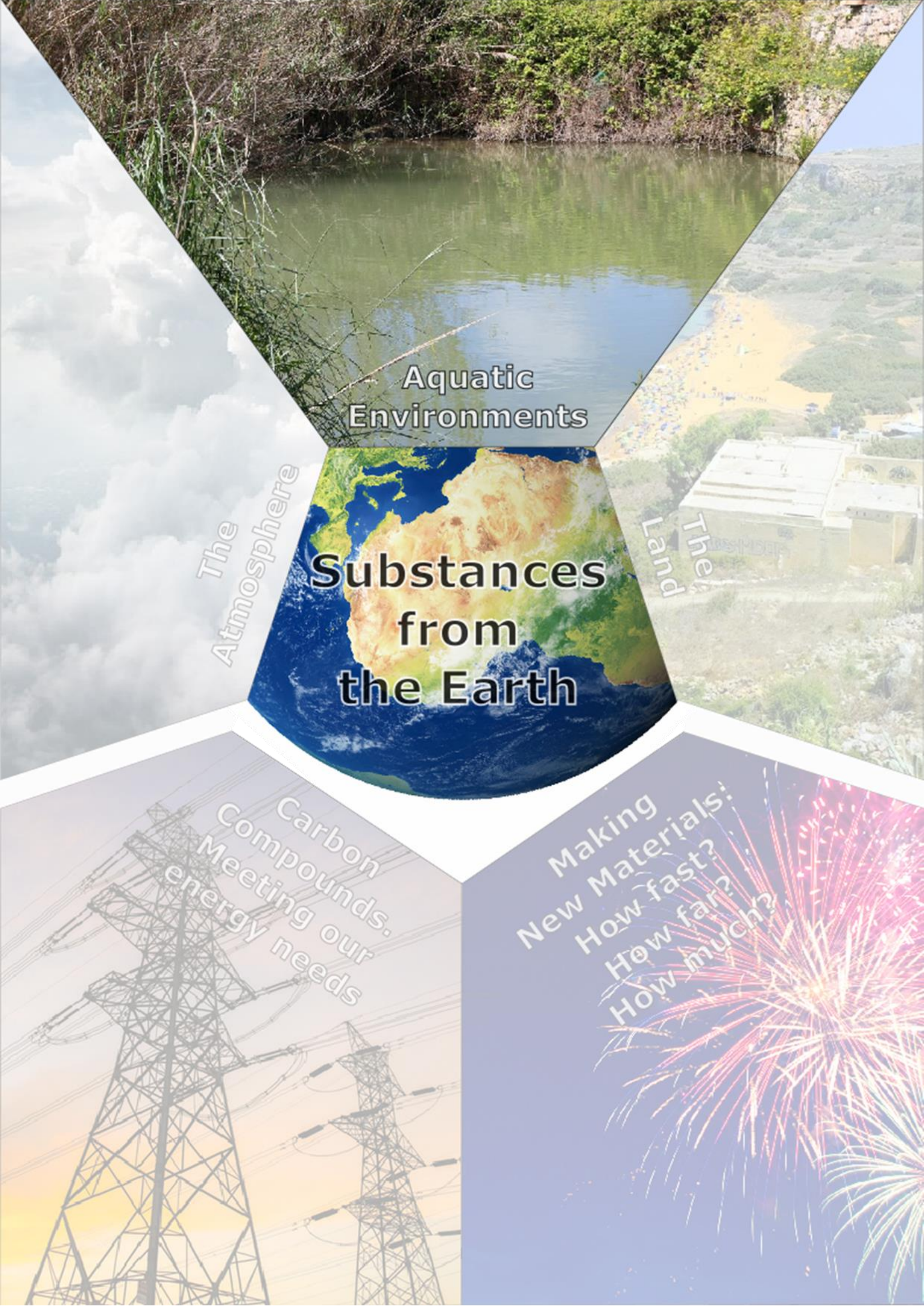
Contents

Substances from the Earth



1 Identification of soluble substances and measurement of their concentration – LO 7

1.1 Chromatography

By the end of this section, you should be able to:

- Use paper chromatography to identify the components of a coloured mixture. *(Solvent limited to water.)* (7.1a)
- Perform paper chromatography. *(Solvents limited to water and ethanol.)* (7.2a)
- Interpret chromatograms. (7.3a)

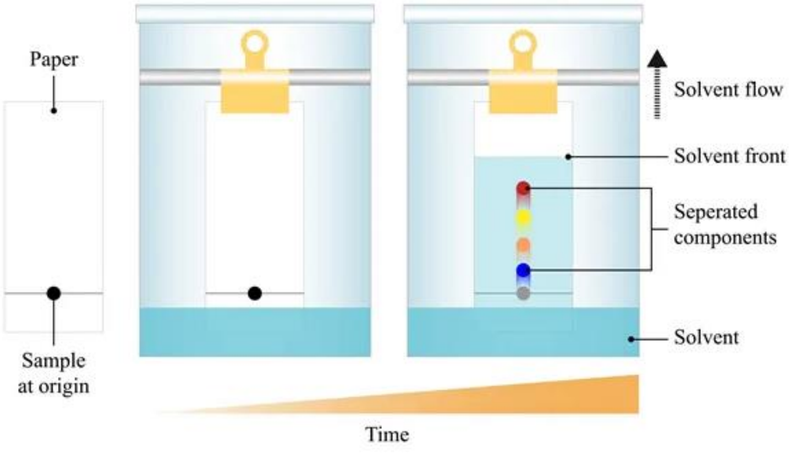
Chromatography is an analytical technique commonly used for separating a mixture of chemical substances into its individual components. There are various types of chromatography techniques, but all employ the same basic principles. At this level only paper chromatography will be dealt with.

Chromatography gets its name from the application of a technique first used in the late 19th century to separate pigments in a mixture.

1.1.1 Performing paper chromatography

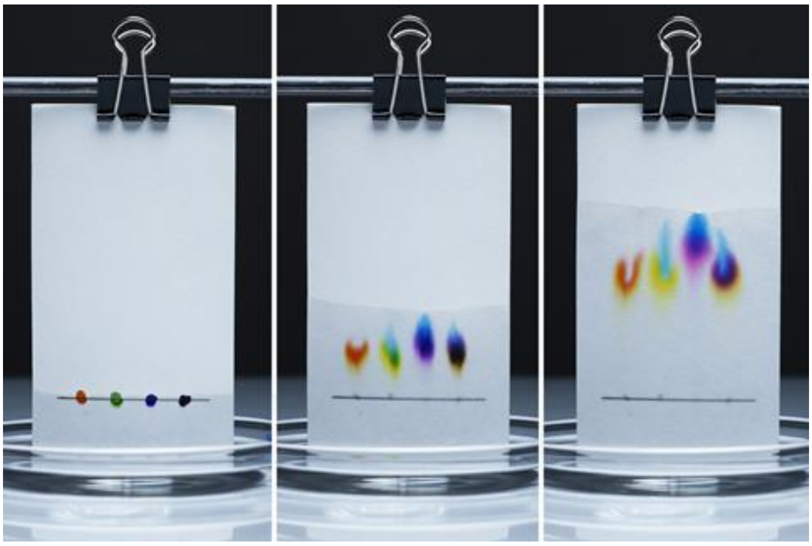
The following steps are carried out to perform paper chromatography for example on a sample of black ink:

1. Draw a baseline on a chromatography paper using a pencil. Since the graphite in the pencil is insoluble, it will not be affected by the solvent and so remains stationary.
2. Place a small spot of the black ink on the baseline. Using a small amount of ink helps to distinguish between the components of the ink (if any) at the end of experiment.
3. Pour a suitable solvent (water or ethanol) into a beaker. Hang the chromatography paper over the solvent in such a way that the base of the paper touches the solvent without the baseline being submerged.
4. Cover the beaker to avoid excessive evaporation of the solvent from the chromatography paper.
5. As the solvent slowly travels upwards through the chromatography paper, it carries with it the soluble components from the sample mixture (black ink) producing different spots along the paper.
6. Before the solvent reaches the top, remove the paper from the beaker and set aside to dry. At this point the chromatography paper including the spots can now be called a **chromatogram**.
7. Analyse the chromatogram to identify the different components of the sample of black ink where each spot represents a different component of the mixture.



Source: https://bit.ly/4216Fi5

The above chromatography experiment can be performed on different samples of ink at the same time. Below is a photograph of the same chromatogram obtained for different samples of ink at different points in time.



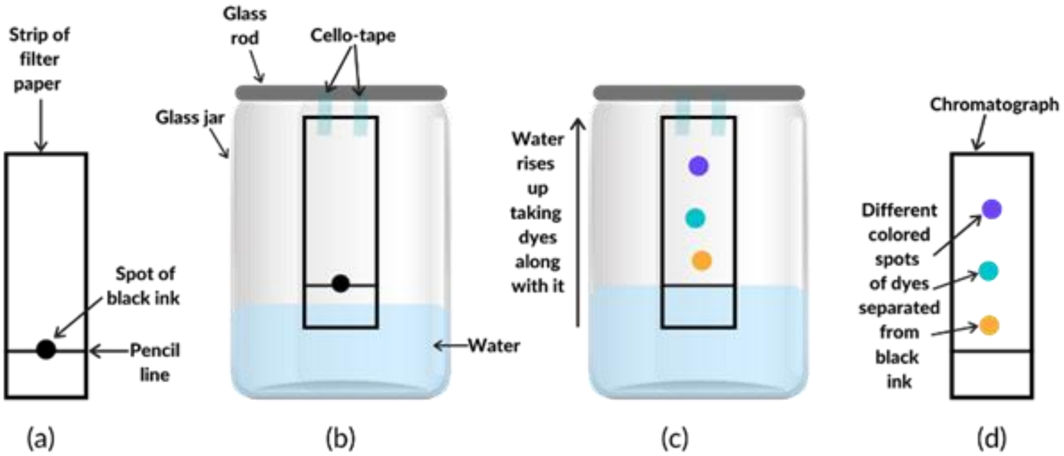
Source: edu.rsc.org

1.1.2 Interpreting chromatograms

By referring to the above photograph of a chromatogram, as the absorbent paper absorbs the solvent, the mixture of solvent and the dissolved pigments move up the paper, but the components of the pigment will not travel upwards at the same rate.

This happens because the largest molecules of the mixture travel more slowly (they're heavier) while the smallest ones move faster, causing the chromatography paper (also known as the stationary phase) to develop separate bands of colour corresponding to each component of the mixture. This gives the technique the name "chromatography" or "colour writing".

What this means is that the more soluble and/or lighter (less heavy) dye components travel quickly and further away from the original spot of pigment. The less soluble and/or heavier ones, take longer and travel slower. So you end up with something like this:



Source: https://www.shalom-education.com/courses/gcsechemistry/lessons/chemical-analysis/topic/chromatography/

1. Mystery at the Crime Scene! Police suspect a thief used a permanent marker to write graffiti. They have a black marker found at the scene and a black marker belonging to the suspect. How could chromatography help them figure out if the markers are the same? Explain your answer, considering the different components in a marker ink.

2. Food Detectives! You suspect your favourite candy contains food colouring. How could you use chromatography to separate the different colours and identify them? What materials would you need?

3. Plant Power! Scientists are studying the pigments (colourful molecules) in different leaves. How could chromatography be used to identify and compare the pigments in spinach leaves and rose petals?

Watch this video on how to do paper chromatography.

Uses of chromatography

1. Creating vaccinations
2. Food testing
3. Beverage testing
4. Drug and forensic testing

Different chromatography techniques

Different chromatography techniques are used in advanced applications and involve different technologies, e.g.:

- Gas chromatography
- Liquid chromatography
- Thin-layer chromatography
- Ion exchange chromatography (Electrophoresis, which is used to separate genetic material uses this technique.)

1.2 Identifying gases, cations, and anions

1.2.1 Identifying gases

By the end of this section, you should be able to:

- Perform chemical tests to identify gases. *(Limited to water vapour, oxygen, hydrogen, carbon dioxide, chlorine, and ammonia.)* (7.1b)
- Describe the test of gases and the expected observations. *(Limited to water vapour, oxygen, hydrogen, carbon dioxide, chlorine, and ammonia.)* (7.2b)
- Identify gases from descriptions of chemical tests. *(Limited to water vapour, oxygen, hydrogen, carbon dioxide, chlorine, and ammonia.)* (7.3b)

Many gases can be produced because of a chemical reaction. You've looked at many ways to make some gases, now let's look at ways to confirm the identity of the gases you've made.

Gas	Properties	Test	Equation	Link
Water vapour, H_2O	Colourless, odourless, tasteless, neutral	Turns white anhydrous copper(II) sulfate blue or Turns blue anhydrous cobalt(II) chloride pink	$CuSO_4(s) + 5H_2O \rightleftharpoons CuSO_4.5H_2O$ $CoCl_2 + 6H_2O \rightleftharpoons CoCl_2.6H_2O$	
Oxygen O_2	Colourless, odourless, tasteless, neutral	Relights a glowing splint	n/a	
Hydrogen H_2	Colourless, odourless, tasteless, neutral	Burns with a 'pop'	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$	
Carbon dioxide CO_2	Colourless, odour- less, tasteless, weakly acidic	Turns lime water milky	$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$	
Ammonia NH_3	Colourless, pungent, alkaline	Forms white fumes (smoke) with hydrogen chloride	$HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$	
Chlorine Cl_2	Yellow-green, pungent, acidic	Turns **damp** litmus paper red and then bleaches it white	$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HOCl(aq)$	

1.2.2 Identifying cations

By the end of this section, you should be able to:

- Perform flame tests. *(Limited to identification of* Li^+*,* Na^+*,* K^+*, and* Ca^{2+} *ions)* (7.1c)
- Identify cations present in salts/solutions using flame tests. *(Limited to identification of* Li^+*,* Na^+*,* K^+*, and* Ca^{2+} *ions)* (7.2c)
- Identify cations present in solutions. *(Limited to identification of* Mg^{2+}*,* Ca^{2+}*,* NH_4^+*,* Cu^{2+}*,* Fe^{2+}*, and* Fe^{3+} *with sodium hydroxide solution.)* (7.2d)
- Identify cations present in solutions. *(Limited to identification of:* Al^{3+}*,* Pb^{2+} *with sodium hydroxide solution,* Pb^{2+} *with KI solution.)* (7.3d)
- Represent reactions for cations using chemical equations. (7.2f part)
- Represent reactions for cations using net ionic equations. *(Except the test for nitrate ions.)* (7.3f part)
- Represent the precipitation of an insoluble salt using net ionic equations. (4.3e)

The identification of cations involves several tests. In order to streamline the process, it is a good idea to have a plan on the order of execution of these tests. Normally, the process starts with a flame test, then a sodium hydroxide test followed by other confirmatory tests if necessary.

1.2.2.1 Flame test

The flame test is a common and simple test carried out to indicate or confirm the presence of certain cations. Here is one way to carry out a flame test.

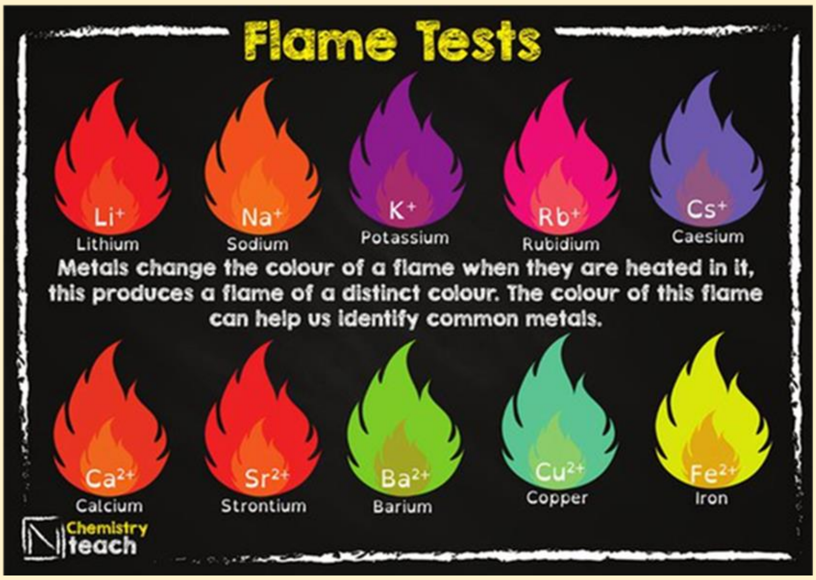
1. Dip a nichrome or platinum wire in concentrated hydrochloric acid.
2. Put the wire in a blue flame. If a colour is observed, go back to step 1, and repeat until the wire imparts no colour to the flame.
3. Dip the nichrome in clean concentrated hydrochloric acid and dip into the solid sample of the compound.
4. Place the wire in the top third of a blue Bunsen flame for a short period of time.
5. Note any colour changes.

Alternatively, the salt to be tested can be dissolved in distilled water and sprayed into a flame. The flame will change colour depending on the kind of metal ion it contains.

Colour of Flame	Cation Probably Present
Lilac	K^+
Golden Yellow	Na^+
Orange Red	Ca^{2+}
Red	Li^+

Watch a video on how to do flame tests.

Other metals also give a flame colour when exposed to a flame as shown in the picture below. These colours are used in fireworks displays.



Flame Tests

Metals change the colour of a flame when they are heated in it, this produces a flame of a distinct colour. The colour of this flame can help us identify common metals.

1.2.2.2 Sodium hydroxide test

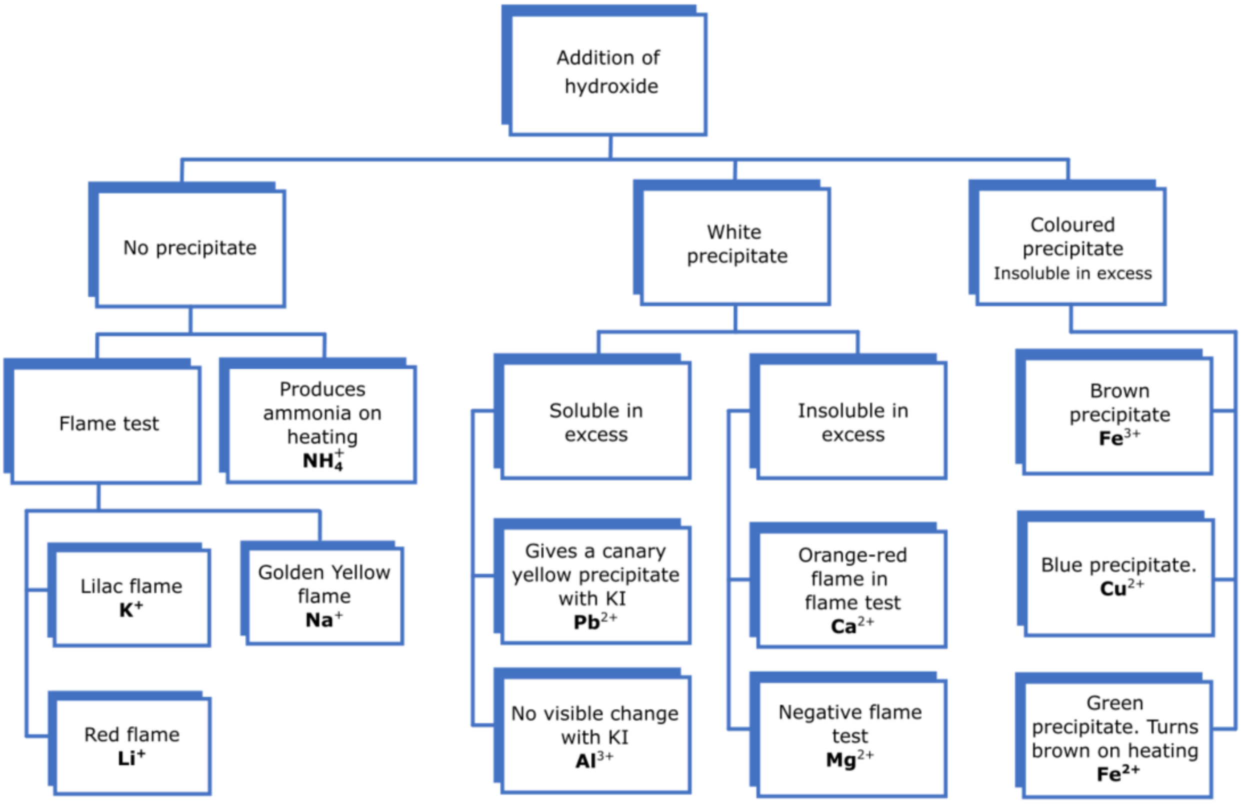
The addition of hydroxide ions, usually in the form of sodium hydroxide solution, is a very general test for cations. It is not always conclusive and where possible it should be accompanied by other tests such as flame tests for confirmation or addition of KI solution. The steps are as follows:

1. Dissolve a small amount of the substance in water. If the substance is insoluble, you can try to add an acid such as hydrochloric or nitric acid to make the substance soluble in water.
2. Add a few (three or four) drops of hydroxide solution and note any observations.
3. Add an excess (one full pipette) of hydroxide solution and note any observations.
4. If necessary, boil the mixture and note any changes. Test any gas produced with damp red litmus paper.

Observation	Ion/s probably present	Equations	Follow-up test	Link
No apparent change	Na^+, K^+, Li^+, H^+	n/a	Flame test	n/a
No change in colour. Release of a gas with a pungent smell that turns moist red litmus paper blue. May require heating.	NH_4^+	$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + H_2O(l)$	n/a	

Observation	Ion/s probably present	Equations	Follow-up test	Link
White precipitate. Insoluble in excess OH^-	Ca^{2+} Mg^{2+}	$Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$ $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$	Flame test	
White precipitate. Soluble in excess OH^-	Pb^{2+} Al^{3+}	$Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$ $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$	Addition of iodide ions.	
Blue precipitate. Insoluble in excess.	Cu^{2+}	$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$	n/a	
Green precipitate. Turns brown on heating / standing in air.	Fe^{2+}	$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$	n/a	
Brown precipitate. Insoluble in excess.	Fe^{3+}	$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$	n/a	

The following diagram shows a summary of the results of the sodium hydroxide test.



1.2.2.3 Confirmatory tests for cations

After having carried out the sodium hydroxide test, a flame test can be used to distinguish between sodium and potassium ions and between magnesium and calcium ions.

If one suspects the presence of an acid (hydrogen ions, H^+), any of the reactions of acids can be used for confirmation together with a pH test. Acids release hydrogen when reacted with metals and carbon dioxide when reacted with carbonates.

Aluminium and lead ions are both amphoteric and as such dissolve in excess hydroxide ions. The presence of lead(II) can be confirmed by adding an aqueous iodide, usually aqueous potassium iodide. A canary yellow precipitate forms since lead(II) iodide is insoluble in water.

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

Aluminium, on the other hand, gives no visible change in this reaction.

After having confirmed the cation in an ionic compound, the next step is to identify the anion.

Watch a video on how lead(II) ions react with iodide ions in aqueous solution to produce a yellow precipitate of lead(II) iodide.

1.2.3 Identifying anions

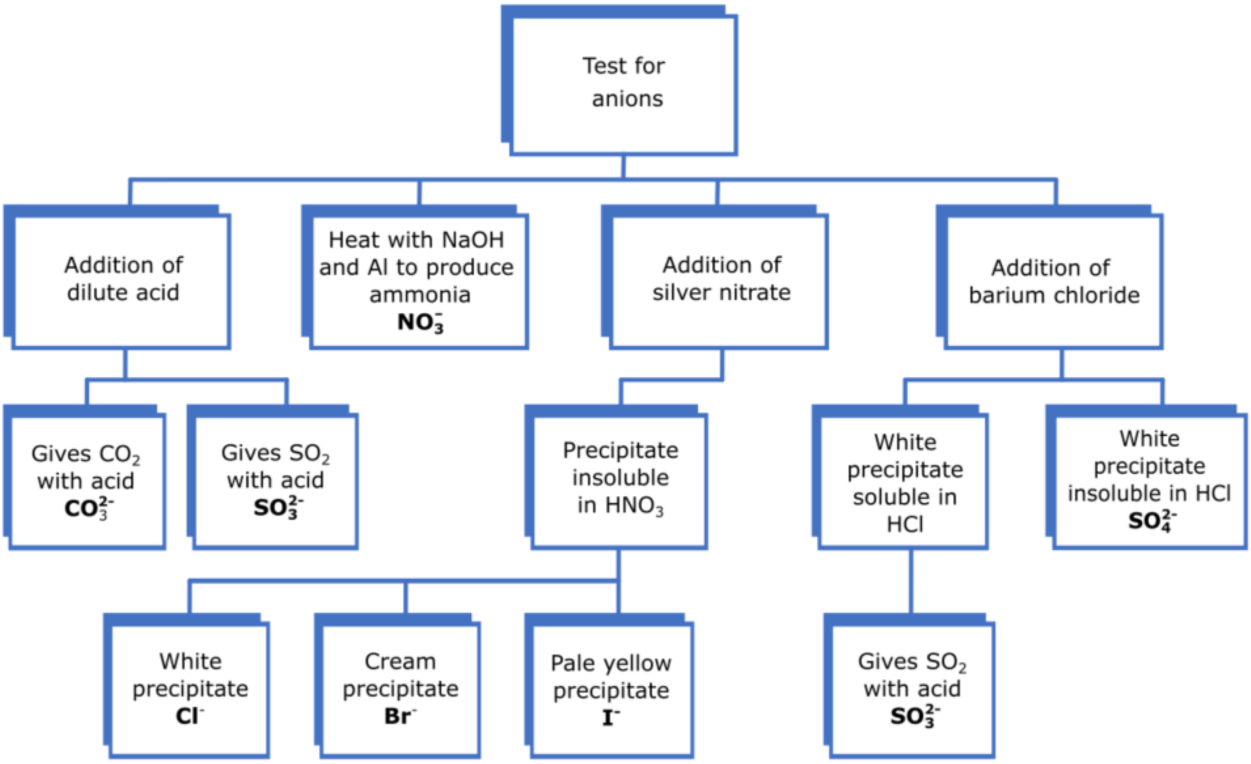
By the end of this section, you should be able to:

- Identify anions present in solutions. *(Limited to identification of: Cl^-, Br^-, I^- with acidified $AgNO_3$ solution, CO_3^{2-} with dilute acid and identifying CO_2)* (7.2e)
- Identify anions present in solutions. *(Limited to identification of: SO_3^{2-} and SO_4^{2-} with acidified $BaCl_2$ solution, NO_3^- by reduction with aluminium and alkali.)* (7.3e)
- Represent reactions for anions using chemical equations. (7.2f part)
- Represent reactions for anions using net ionic equations. *(Except the test for nitrate ions.)* (7.3f part)
- Represent the precipitation of an insoluble salt using net ionic equations. (4.3e)

The following table summarises the tests that must be performed on ionic substances to determine the anion part of an ionic compound. Certain tests require the solution to be acidified in order to eliminate or confirm the possible presence of other anions that could give the same result.

Anion	Test	Observation	Equation
CO_3^{2-}	Addition of dilute acid.	A gas which turns lime water milky.	$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

Anion	Test	Observation	Equation
SO_3^{2-}	Addition of barium chloride solution followed by dilute hydrochloric acid.	White precipitate which reacts with hydrochloric acid to produce a clear solution and a pungent gas of sulfur dioxide.	$SO_3^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_3(s)$ $BaSO_3(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + H_2O(l) + SO_2(g)$
SO_4^{2-}	Addition of barium chloride solution followed by dilute hydrochloric acid.	White precipitate that does not react with hydrochloric acid.	$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$
Cl^-	Addition of dilute nitric acid followed by silver nitrate.	White precipitate.	$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$
Br^-	Addition of dilute nitric acid followed by silver nitrate.	Cream precipitate.	$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s)$
I^-	Addition of dilute nitric acid followed by silver nitrate.	Pale yellow precipitate.	$I^-(aq) + Ag^+(aq) \rightarrow AgI(s)$
NO_3^-	To the sample add NaOH solution and aluminium turnings. Then heat gently.	A pungent gas that turns red litmus blue is produced (ammonia).	n/a



1. The following salts were analysed. Determine the cation and anion present in each salt.

Salt	Test and observations	cation	anion
A	A lilac flame is produced during a flame test. When acidified barium chloride is added to a solution of A, a white precipitate forms.		
B	When sodium hydroxide solution is added to a solution of **B**, a green gelatinous precipitate is formed which is insoluble in excess alkali. When nitric acid followed by silver nitrate solution is added to a solution of **B** a cream precipitate is produced.		
C	A white precipitate is produced when a solution of C is added to sodium hydroxide solution. This precipitate is insoluble in excess sodium hydroxide solution. No colour is produced when a flame test is carried out on C. When barium chloride is added to a solution of **C** a white precipitate is produced. This precipitate dissolves in acid to give a pungent smelling gas.		
D	A white precipitate is produced when sodium hydroxide is added to a solution of **D**. The white precipitate dissolves in excess sodium hydroxide solution. No precipitate is produced when solution D is added to potassium iodide. Sodium hydroxide solution and aluminium powder are added to a solution of **D** and heated. A pungent smelling gas which turns damp red litmus paper blue is evolved.		

2. Amy carried out tests to identify two unknown white solids **A, B** and **C.**
A flame test is carried out on a sample of solid **A** and it burns with a red flame. Solution A turns phenolphthalein pink and when an equal volume of hydrochloric acid is added to A the indicator turns colourless.

Sodium hydroxide solution is added to a solution of compound **B** and a white precipitate is formed which is insoluble in excess sodium hydroxide solution. An orange red flame is produced when a flame test is carried out. A pale-yellow precipitate forms when acidified silver nitrate solution is added to a solution of B.

Compound C is insoluble water. When C is added to nitric acid a colourless gas is evolved which turns limewater milky and a colourless solution D is produced. Sodium hydroxide solution is added to solution D forming a white precipitate which is soluble in excess sodium hydroxide. A canary yellow precipitate is formed when potassium iodide solution is added to solution D.

a) Identify compounds **A, B** and **C**.
b) Describe how a flame test is carried out.
c) Write a balanced chemical equation, including state symbols, for the reaction between:
   i) solution **A** with hydrochloric acid.
   ii) solution **B** and sodium hydroxide solution.
   iii) compound C and nitric acid.

1.3 Titrations and calculations involving concentrations

1.3.1 Calculations involving moles and molar concentrations

By the end of this section, you should be able to:

- Perform calculations involving moles and molar concentrations. *(Do not use the formula:* $\frac{MaVa}{mole\ ratio\ (a)} = \frac{MbVb}{mole\ ratio\ (b)}$ *)*(7.3g)

Volumetric analysis is a process used to determine the amount of the desired constituent by its **volume**. Hence it is considered as a branch of **quantitative analysis**.

By means of volumetric analysis we can determine the **concentration** of a solution. The procedure used is known as a **titration**. Titrations are commonly used when carrying out **acid-base reactions**.

First, one must understand what **concentration** of a solution means.

1.3.1.1 Concentration

Apart from knowing what substances are present in a solution, it is sometimes also important to know **how much** of a solute there is in a solution. You can calculate the **amount of a substance** (in **moles** or in **grams**) in a solution if you know the **volume** and **concentration**.

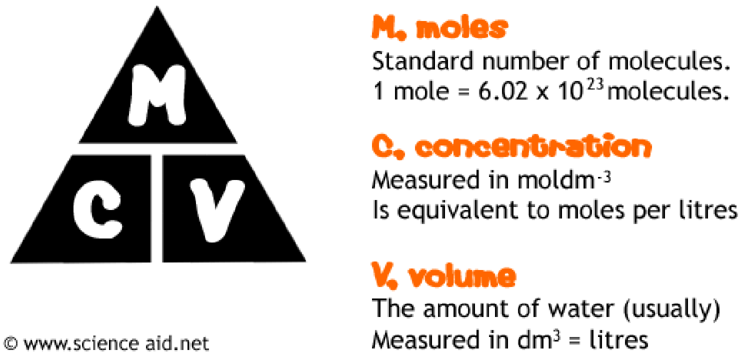
Concentration is the number of moles of substance dissolved in one decimetre cubed (dm^3) of solvent.

It is measured in **mol dm^{-3}**.

Concentration measured in **mol dm^{-3}** can also be expressed in **g dm^{-3}** if the mass of solute dissolved in solution is taken in consideration. The following equation is used to calculate concentration

$$\text{Concentration} = \frac{\text{Moles of Solute}}{\text{Volume of Solvent}}$$

Recall that **1 dm^3** is equal to **1 litre**, which is equivalent to **1000 cm^3** or **1000 ml**.



M, moles

Standard number of molecules.

1 mole = 6.02×10^{23} molecules.

C, concentration

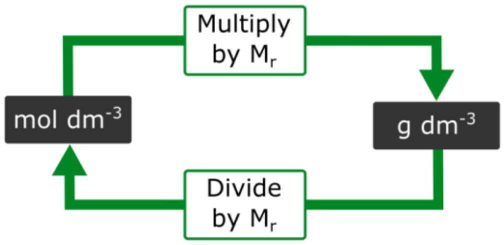
Measured in moldm^{-3}

Is equivalent to moles per litres

V, volume

The amount of water (usually)

Measured in dm^3 = litres



Concentration in **mol dm^{-3}** is sometimes referred to as **molarity (M)**.

You may find the latter unit used in chemistry literature.

Watch this video clip about concentration formula and calculations.

Worked examples

1. Find the concentration (in mol dm^{-3}) of a sample of hydrochloric acid if 2 moles of hydrogen chloride gas were dissolved in 4 dm^3 of water.

   Recall: Concentration = moles of substance dissolved in 1 dm^3

   In 4 dm^3 there are 2 moles HCl gas dissolved

   In 1 dm^3 there are ? (2 x 1) / 4 = **0.5 mol dm^{-3}** HCl(aq)

2. How many moles of substance are present in 50 dm^3 of sulfuric acid of concentration 0.5 mol dm^{-3}?

   Since concentration is 0.5 mol dm^{-3}

   So, in 1 dm^3 there are 0.5 moles H_2SO_4

   in 50 dm^3 there are ? (0.5 x 50) / 1 = **25.0 moles** H_2SO_4

3. 0.03 moles of sodium hydroxide (NaOH) are needed for an experiment. What volume of a 0.1 mol dm^{-3} solution of sodium hydroxide must be used?

   Since concentration is 0.1 mol dm^{-3}

   So, 0.1 moles NaOH are dissolved in 1 dm^3

   0.03 moles are dissolved in ? (1 x 0.03) / 0.1 = **0.3 dm^3**

4. 212 g of sodium carbonate (Na_2CO_3) were dissolved in 2000 cm^3 of water. Find the concentration (in mol dm^{-3}) of the solution.

   Relative Formula Mass (RFM) for Na_2CO_3 = (23 x 2) + (12 x 1) + (16 x 3) = 106

   Therefore moles of Na_2CO_3 in 212 g = 212 / 106 = 2 moles

   To calculate concentration: 2 moles Na_2CO_3 are dissolved in 2000 cm^3 (or 2 dm^3)

   ? moles Na_2CO_3 are dissolved in 1000 cm^3 (or 1 dm^3)

   (2 x 1000) / 2000 = **1.0 mol dm^{-3}**

5. In an experiment, 18.625 g of potassium chloride (KCl) need to be prepared by reacting 0.1 mol dm^{-3} potassium hydroxide (KOH) with 0.2 mol dm^{-3} hydrochloric acid. Calculate:

   a. the amount in moles of potassium chloride that must be prepared,
   b. the amount in moles of potassium hydroxide and hydrochloric acid needed for the preparation, and
   c. the volume of potassium hydroxide and hydrochloric acid needed.

   Balanced equation for reaction: $KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

   d. RFM for KCl = 39 + 35.5 = 74.5

      Moles of KCl in 18.625g which must be prepared = 18.625 / 74.5 = **0.25 moles**

   e. Using mole ratios from balanced equation:

      KCl : KOH = 1 : 1 So, **0.25 moles of KOH** needed.

      KCl : HCl = 1 : 1 So, **0.25 moles of HCl** needed.

f. To calculate volume of KOH needed:

Since concentration is 0.1 mol dm^{-3}

So, 0.1 moles KOH are dissolved in 1 dm^3

0.25 moles are dissolved in ? (1 x 0.25) / 0.1 = **2.5 dm^3**

To calculate volume of HCl needed:

Since concentration is 0.2 mol dm^{-3}

So, 0.2 moles HCl are dissolved in 1 dm^3

0.25 moles are dissolved in ? (1 x 0.25) / 0.2 = **1.25 dm^3**

1. Find the concentration (in mol dm^{-3}) of a salt solution if 0.75 moles of sodium chloride were dissolved in 1500 cm^3 of water.
2. 100 g of copper(II) sulfate pentahydrate ($CuSO_4.5H_2O$) were dissolved in 0.4 dm^3 of water. Calculate the concentration of the copper(II) sulfate solution in mol dm^{-3}.
3. 500 cm^3 of potassium carbonate (K_2CO_3) solution of concentration 0.8 mol dm^{-3} were evaporated to dryness. What mass of potassium carbonate should be obtained?

1.3.2 Standard solutions

By the end of this section, you should be able to:

- Prepare a standard solution using step by step instructions. *(Limited to sodium carbonate.)* (7.2h)
- Prepare a standard solution. *(Limited to sodium carbonate.)* (7.3h)

Volumetric analysis involves ways of calculating the concentration of a solution.

The procedure used is commonly known as a **titration**. Titrations are commonly used when carrying out **acid-base reactions**.

The 1st stage of volumetric analysis is the preparation of a **standard solution**.

A **standard solution** is a solution containing
a precisely known concentration of a substance.

In the 2nd stage of this analytical process, a measured volume of the standard solution is reacted or **titrated** with a solution of unknown concentration in the presence of an **indicator**.

Indicators give a colour change at the **end-point** of a reaction. This end-point indicates that the reaction is complete.

A standard solution is generally prepared by dissolving an accurately known mass of solute in accurately known volume of water.

An ideal substance commonly used for standard solutions is **anhydrous sodium carbonate** because it is a pure, stable, soluble solid that can be accurately weighed.

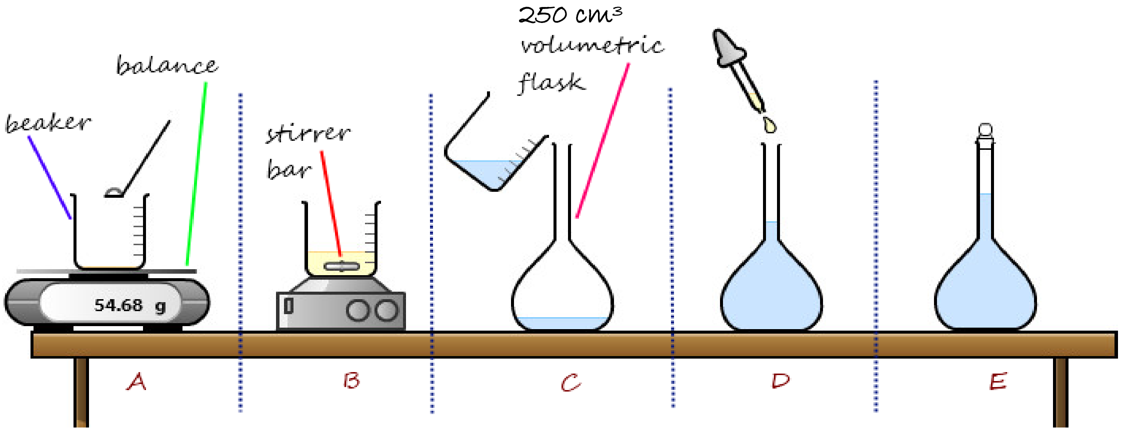
Choice of solute or standard solution

Not all compounds are suitable to prepare standard solutions for various reasons.

- Some compounds absorb water vapour from air, so their mass varies: e.g. solid sodium hydroxide and potassium hydroxide are both deliquescent; concentrated sulfuric acid is hygroscopic.
- Some compounds react with CO_2 in air: e.g. NaOH and KOH.
- Some compounds evaporate easily: e.g. concentrated hydrochloric acid and ammonia.
- Standard solutions for such compounds cannot be prepared with the procedure that will be outlined hereunder, however they can be prepared using slightly different procedures.

1.3.2.1 Preparing a standard solution

How to prepare a 250 cm^3 standard solution of 0.1 mol dm^{-3} anhydrous sodium carbonate.



Stage A: The first stage in the procedure actually involves 3 steps.

1. **Calculate the mass** of anhydrous sodium carbonate needed:
   RFM for Na_2CO_3: (23 x 2) + (12 x 1) + (16 x 3) = 106

   Mass required for 1 dm^3 of 0.1 mol dm^{-3} standard solution:
   1 mole of Na_2CO_3 has a mass of 106 g
   0.1 moles of Na_2CO_3 has a mass of: (0.1 x 106) / 1 = **10.6 g**

   Mass required for 250 cm^3 of 0.1 mol dm^{-3} standard solution:
   1000 cm^3 solution contains 10.6 g of Na_2CO_3
   250 cm^3 solution contains: (10.6 x 250) / 1000 = **2.65 g**
2. Make sure the salt is completely **anhydrous**, i.e. it does not contain any water of crystallization. This is done by heating the salt to constant mass. (N.B.: sodium carbonate is thermally stable so there is no risk of decomposition when heating it.

Weighing boat

3. **Accurately weigh** the calculated mass (2.65 g) on a **weighing boat** or **watch glass** using an **electronic mass balance.**

Stage B: A beaker is washed thoroughly with distilled water. The weighed solid is transferred into the clean beaker. The weighing boat is rinsed with distilled water to ensure that all the solid is transferred. The transferred solid is completely dissolved in approximately 100 cm^3 of distilled water mixing with a stirring rod.

Stage C: The solution is transferred to a **250 cm^3 volumetric flask** using a clean **glass funnel**. Any equipment used to handle the compound such as funnel, beaker and stirring rod, are washed thoroughly with distilled water and this is also transferred to the volumetric flask. This ensures **quantitative transfer** (exact known amount).

Stage D: The funnel is removed, and the volumetric flask is **filled up to the mark** with distilled water.

A **teat pipette** should be used to make up the last few drops until the bottom of the meniscus is on the graduation mark.

Stage E: Finally, the stopper is fitted, and the filled volumetric flask is thoroughly shaken to ensure the solution is **homogenous**.


A volumetric flask filled exactly to the mark.


Teat pipette

Watch this video about preparing a standard solution of sodium carbonate.

Worked example

What mass of potassium hydrogen phthalate, $C_8H_5KO_4$, in grams, must be weighed out in order to make 200 ml standard solution of 0.250 mol dm^{-3} $C_8H_5KO_4$?

Molar mass for $C_8H_5KO_4$: (12 x 8) + (1 x 5) + (39 x 1) + (16 x 4) = 204

Mass required for 1 dm^3 of 0.25 mol dm^{-3} standard solution: 0.25 x 204 = 51.0 g

Mass required for 200 cm^3 of 0.25 mol dm^{-3} standard solution: (51.0 x 200) / 1000 = **10.20 g**

Further reading on the uses of standard solutions.

A virtual environment dealing with solution concentrations.

1.3.3 Titrations

By the end of this section, you should be able to:

- Conduct an acid/base titration using step by step instructions. *(Limited to hydrochloric acid, sulfuric acid, with sodium hydroxide, potassium hydroxide and sodium carbonate.)* (7.2i)
- Conduct an acid/base titration to determine the concentration of a given solution. *(Example hydrochloric acid, sulfuric acid, nitric acid, ethanoic acid with sodium hydroxide, potassium hydroxide and sodium carbonate.)* (7.3i)
- Calculate the concentration/volume of a solution taking part in a reaction. (7.3j)

1.3.3.1 Acid-base titrations

Titrations are commonly used when carrying out acid-base neutralisation reactions. There are two main purposes of an acid-base titration:

1. Preparation of **soluble salts**. (Note that this was discussed in Year 9 syllabus.)
2. As a means of **volumetric analysis** in order to **determine the unknown concentration of a base or an acid**.

In volumetric analysis, the 1st stage consists of the preparation of a **standard solution**.

In the 2nd stage of this analytical process, a measured volume of the standard solution is reacted or **titrated** with a solution of **unknown concentration** in the presence of an **indicator**.

Indicators give a colour change at the **end-point** of a reaction. This end-point indicates that the reaction is complete, that is **complete neutralisation** has been achieved.

Choice of Indicator

Not every indicator is necessarily suitable for all types of acid-base reactions. The table below indicates the two most common indicators used and for which type of acid-base titrations they are better suited.

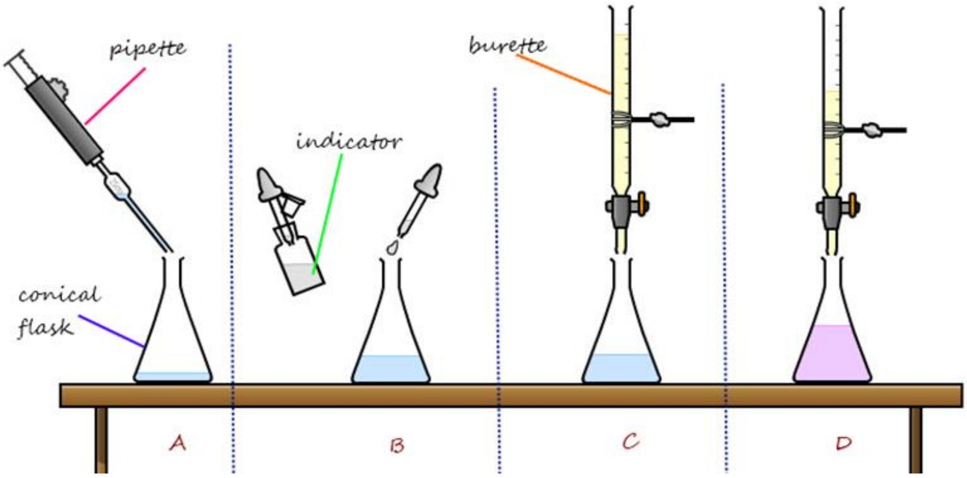
Indicator	Colour of indicator in:			Type of acid – base titration
	acidic solution	**neutral** solution	**alkaline** solution	
Methyl orange	**Red**	**Orange**	**Yellow**	strong acid – strong alkali strong acid – weak alkali
Phenolphthalein	**Colourless**	**Colourless**	**Pink**	strong acid – strong alkali weak acid – strong alkali

Note that **universal indicator** is not a good indicator to use in titrations because it does not have a single end-point since it is a mixture of indicators.

Let us consider the following titration experiment:

- Acid of **unknown** concentration: **Ethanoic acid**, CH_3COOH – a **weak acid**
- Alkaline **standard solution**: **Sodium hydroxide**, $NaOH$ – a **strong alkali**
- Suitable **indicator**: **Phenolphthalein**

Example: Titration of Ethanoic acid solution against 0.2 mol dm^{-3} sodium hydroxide standard solution.



Titration procedure

Precautions

Before starting titration, it is important to:

- ✓ Wash the volumetric pipette with distilled water **followed by** the solution it is going to contain (vinegar in this case).
- ✓ Wash the conical flask with distilled water **only**.

Stage A

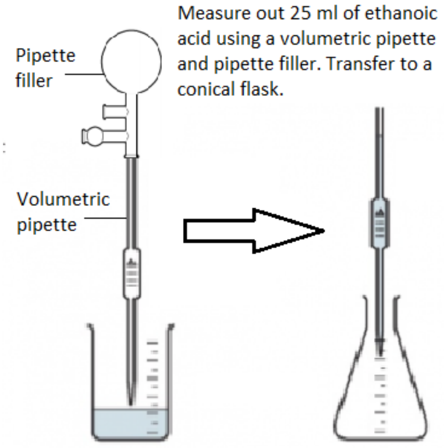
25 cm^3 of white vinegar (ethanoic acid solution) are transferred to a clean conical flask using a 25 cm^3 **volumetric pipette** fitted with a **pipette filler**.

Stage B

Using a teat pipette, two drops of phenolphthalein **indicator** are added to the vinegar in the conical flask.

Stage C

A clean **burette** is filled with the standard solution (0.2 mol dm^{-3} sodium hydroxide) using a glass funnel. The **initial volume** of solution in the burette is recorded.


Measure out 25 ml of ethanoic acid using a volumetric pipette and pipette filler. Transfer to a conical flask.

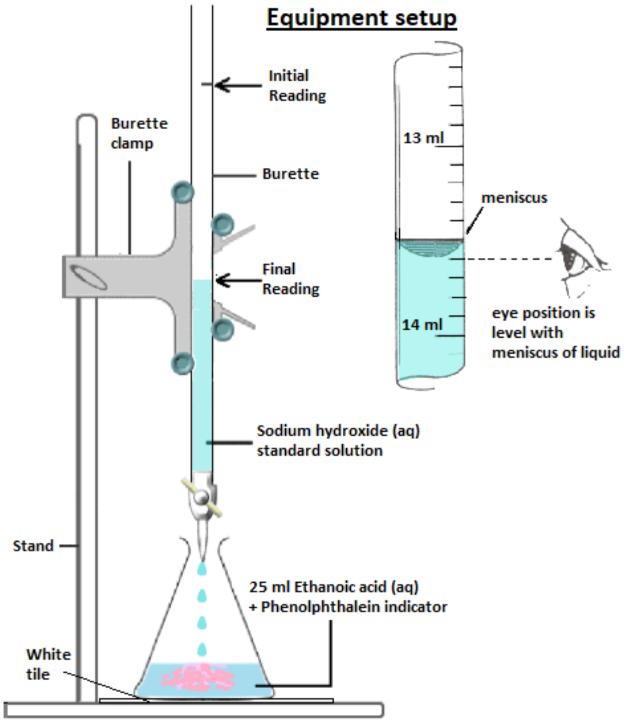
Precautions for stage C

It is important to:

- ✓ Wash burette with distilled water **followed by** the solution it is going to contain (sodium hydroxide in this case) before this step is carried out.
- ✓ Remove glass funnel after burette is filled and before reading volume.
- ✓ Read volume at eye level from the bottom of the meniscus.
- ✓ Ensure that there are no air bubbles in the burette.
- ✓ Ensure that the burette is vertical.

Stage D

1. The burette valve is opened so that the NaOH solution is added to the conical flask, while the flask is swirled regularly, until the indicator changes to a **light pink colour**, at which point the burette valve is closed. This final volume is read and is recorded as a **rough titre value**. It gives an estimate of the volume of standard NaOH solution needed to fully neutralise the 25 cm^3 acid in the conical flask.
2. This procedure is repeated, this time adding NaOH **dropwise** to accurately determine the exact volume needed.
3. The experiment is repeated multiple times until **at least two concordant**


Equipment setup

titre values that do not vary more than **±0.1 cm³** from each other are obtained. All titre values are recorded in a table.

4. The **average titre value** is calculated. The rough titre value is **not** included in this calculation.
5. Calculations are then carried out to determine the concentration of the ethanoic acid in the vinegar.

Precautions for stage D

It is important to:

- ✓ Always carry out a rough titration first.
- ✓ Add burette solution dropwise to the conical flask especially when getting closer to the expected titre value.
- ✓ Conical flask must be swirled regularly to ensure a homogenous mixture.
- ✓ Use a white tile underneath the conical flask.
- ✓ Read all burette volumes at eye level from the bottom of the meniscus.

Sample results

	Rough titration	1st Titration	2nd Titration
Initial reading (cm^3)	0.00	0.00	13.65
Final reading (cm^3)	14.40	13.65	27.20
Volume NaOH used (cm^3)	14.40	**13.65**	**13.55**
	Average titre value (cm^3)	**13.60**	

Calculations to determine concentration of ethanoic acid in the vinegar solution used.

Step 1: Calculate the **number of moles of standard solution** used.

Standard solution used (average titre): 13.60 cm^3

In 1000 cm^3 → 0.2 moles NaOH

In 13.60 cm^3 → (0.2 x 13.60) / 1000 = **0.00272 moles** of NaOH solution used

Step 2: Write down a **balanced** chemical equation for the reaction.

$NaOH(aq) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$

Step 3: Determine the **mole ratio** of reactants from the **balanced** chemical equation.

According to balanced equation, 1 mole NaOH neutralises 1 mole CH_3COOH. Therefore:

$NaOH : CH_3COOH = 1 : 1$

Step 4: Calculate the **number of moles** of the solution of **unknown concentration** (in conical flask). In this titration, 25 cm^3 ethanoic acid (vinegar) were in the conical flask.

Since mole ratio **NaOH : CH_3COOH** is **1 : 1** this means that:

0.00272 moles NaOH react with **0.00272 moles CH_3COOH.**

These are present in 25 cm^3 of solution in conical flask.

<u>Recall</u>: Concentration is the number of moles dissolved in 1 dm^3 (or 1000 cm^3) of solution.

In 25 cm^3 → 0.00272 moles CH_3COOH

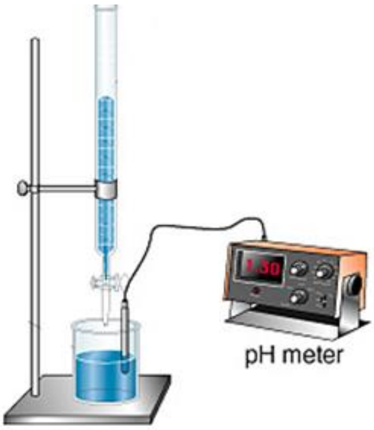
In 1000 cm^3 → (0.00272 x 1000) / 25 = **0.1088 mol dm^{-3}**

Therefore, the concentration of ethanoic acid in the vinegar in the conical flask is **0.1088 mol dm^{-3}**

Watch this video on how to perform a titration of ethanoic acid on sodium hydroxide including the necessary calculations.

1.3.3.2 Digital pH meters

Accurate digital pH meters can be used instead of indicators to determine the end-point. These are particularly useful when the solution in the conical flask is dark in colour (e.g. red vinegar or red wine). Indicators are useless in such cases as any colour change would not be visible.



Uses of titration in industry

- In the pharmaceutical sector, titrations can be used to analyse the purity of medicines, for example by titrating ephedrine found in cough syrup against perchloric acid.
- It is used to check the purity of raw materials such as binding substances used in tablets.
- When manufacturing biodiesel, recycled vegetable oils are often used. Their composition and acidity vary and can be balanced through titration.
- Potable and wastewater water are constantly tested for pH and chemical content to be able to treat it. It also ensures contaminants from factories are well-accounted for and do not contaminate groundwater.
- In the dairy industry, titration is used to determine the right amount of microbes needed to ferment milk into other products. This is due to the lactic acid produced by the bacteria used to ferment the milk.

Perform a virtual titration.

Further reading on the uses of titrations.

1. Give reasons why you think the following PRECAUTIONS taken during a typical titration are important and explain how they can affect the results obtained.
   a. **Volumetric pipette** is washed with the with the solution that they shall be filled with, after initially rinsing with distilled water.
   b. **Burette** is washed with the with the solution that they shall be filled with, after initially rinsing with distilled water.
   c. The **conical flask** is washed with distilled water only.
   d. The glass funnel is removed after filling the burette and before reading initial volume.
   e. A rough titration is carried out at first before repeating the procedure more accurately.
   f. Burette solution is added to conical flask contents dropwise when close to end-point.
   g. The conical flask is swirled regularly while burette solution is being added.

2. A student is asked to find the acidity of a red wine. He decides to carry out a titration. He puts the red wine in the conical flask and the alkali standard solution in the burette. He is undecided about which indicator to use and whether using a white tile would help him be more accurate.
   a. Which indicator would you use? Explain.
   b. Do you think using a white tile would help in this particular titration? Explain.

3. 50 cm^3 of ammonia solution was neutralised by 30 cm^3 of 2 mol dm^{-3} of sulfuric acid.
   a) Calculate the concentration of the ammonia solution in mol dm^{-3}.
   b) Calculate the mass of ammonium sulfate produced from this reaction.

4. Work out the volume of 0.5 mol dm^{-3}sodium carbonate solution that would be required to neutralise 25 cm^3 of 0.75 mol dm^{-3}hydrochloric acid?

5. 25 cm^3 portions of an aqueous solution of 0.15 mol dm^{-3} of potassium hydroxide were titrated against an aqueous solution of dilute phosphoric acid. Three titrations were performed, and the amount of acid used in each titration was recorded below.

	1st titration /cm^3	2nd titration / cm^3	3rd titration / cm^3
Final burette reading	27.60	31.25	15.10
Initial burette reading	13.60	16.05	00.00
Volume of acid used			

   a) Complete the missing values in the table above.
   b) Name the piece of glassware used to measure 25cm^3 samples of potassium hydroxide.
   c) Give the colour change that occurs at end point when methyl orange is used as an indicator.
   d) Work out the average titre value.
   e) Write a balanced chemical equation of the reaction.
   f) Calculate the concentration of phosphoric acid.

Substances from the Earth

Aquatic Environments

The Atmosphere

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

2 Energetics – LO 15

2.1 Energy changes of chemical reactions

By the end of this section, you should be able to:

- Identify chemical reactions that are exothermic or endothermic. (15.1a)
- Associate an exothermic reaction with a negative value of ΔH and an endothermic reaction with a positive value of ΔH. (15.2a)
- Identify exothermic and endothermic reactions from given energy level diagrams. (15.1b)
- Draw energy level diagrams to represent exothermic and endothermic reactions including activation energy. (15.2b)
- Explain energy level diagrams in terms of bond energies. *(Calculations are not required.)* (15.3b)
- Define heat of combustion. (15.2d)
- Define heat of neutralisation. (15.2e)

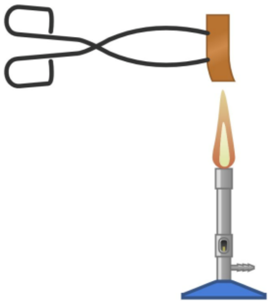
2.1.1 The law of conservation of energy

Different forms of matter possess energy in one form or another. There are different forms of energy which include light, sound, heat, movement, electrical, and stored chemical.

The law of conservation of energy states that:

Energy can neither be created nor destroyed.
Energy can only change from one form to another.

For instance, when some magnesium ribbon is burnt, the chemical energy of the fuel being used to produce the flame is changed into heat and light energy as the magnesium reacts with the oxygen in the air to form magnesium oxide.



Chemical Energy → Heat and Light Energy

The Law of Conservation of Energy

The ideas related to the law of conservation of energy were discussed even in ancient philosophy (c. 550BC), however the concept took quite long to be resolved and understood, as it focused on forces and momentum.

The turning point came after the work of Émilie du Châtelet (after 1722), a French philosopher and mathematician, who proved that energy (not momentum) is totally conserved and can be transformed from one form to another.

Energy & Chemistry: Crash Course Chemistry #17

2.1.2 Bond energy changes in chemical reactions

There are two main energy sources stored in chemicals. These are: kinetic and bond energy.

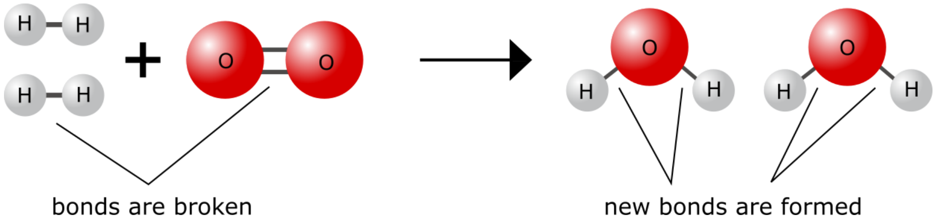
- **Kinetic Energy:** Atoms and molecules within chemicals possess kinetic energy because they are in a state of constant motion. The higher the temperature, the faster is the motion of these particles.
- **Bond Energy:** Bond energy is the energy stored in the attractive forces holding together atoms, ions, or molecules.

During a chemical reaction, the more significant of these two types of energy sources is bond energy. Energetics deals with this type of energy.

When chemical reactions occur there is **bond breaking** between the atoms of the reactants and **bond forming** between the atoms of the products.

For example, in the reaction between hydrogen molecules and oxygen molecules to form water:

$$2H_2 + O_2 \rightarrow 2H_2O$$



there is:

- bond breaking between the atoms of the hydrogen and oxygen molecules.
- bond forming between the atoms of hydrogen and oxygen to form water.

For bonds to break, energy needs to be supplied. (**Endothermic**)

When bonds form, energy is liberated. (**Exothermic**)

The stronger the bond the more energy is required to break it and the more energy it liberates when it forms. Conversely, the weaker the bond, the less energy is required to break it and less energy is liberated when it forms.

The units for bond energy are the **kJ mol^{-1}** or **kJ/mol**.

Bond energy is defined as:
The energy needed to break 1 mole of a particular chemical bond.
OR
The energy released to make 1 mole of a particular chemical bond.

2.1.3 Exothermic and endothermic changes

The term exothermic means that thermal (heat) energy is lost to the external surroundings. Since bond formation liberates heat energy, bond forming is an exothermic change.

The term endothermic means that thermal energy needs to be supplied. Since bond breaking requires heat energy to occur, bond breaking is an endothermic change.

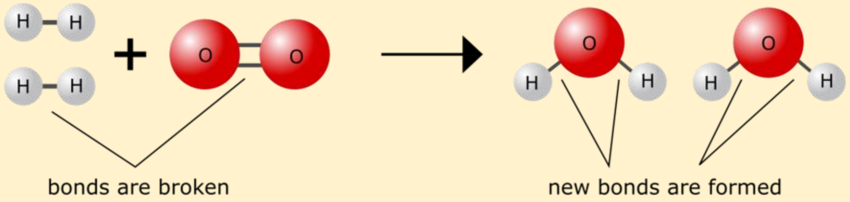
An **exothermic reaction** is one in which the overall reaction releases heat. This shows that the energy released upon bond formation when the products were formed was greater than the energy required for bond breaking of the reactants. Combustion reactions are examples of exothermic reactions.

An **endothermic reaction** is one in which the overall reaction requires heat. This shows that the energy required for bond breaking of the reactants was greater than the energy released upon bond formation when products were made. The thermal decomposition of calcium carbonate is an example of an endothermic reaction.

Worked example involving bond energies for the reaction of hydrogen with oxygen to produce water.

hydrogen + oxygen → water

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$



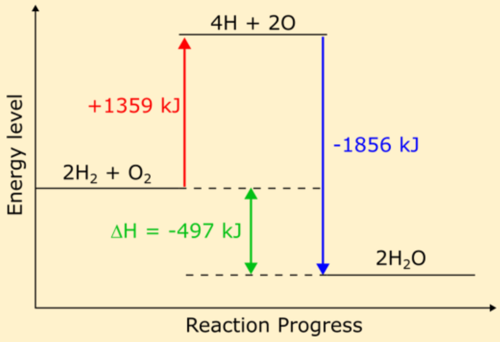
2(H-H) + 1(O=O) → 4(O-H)

Note that to make or break a bond the same amount of energy is needed. For:

- H-H: 432 kJ/mol
- O=O: 495 kJ/mol
- O-H: 464 kJ/mol

Therefore:

Energy absorbed during bond breaking in reactants.	Energy released during bond formation in products.
(2 x 432) + 495 = 1359 kJ	4 x 464 = 1856 kJ



One notes that in this reaction, the difference between the energy levels of the reactants and products (ΔH) has a value of - 497 kJ. This result comes about since the overall exothermic (bond making) processes outweigh the overall endothermic (bond breaking) processes.

Further reading on exothermic and endothermic changes.

2.1.4 Enthalpy change of a reaction and the ∆H convention

The overall heat change that occurs during a chemical reaction is the sum of energy needed to break the bonds of the reactants and the energy liberated when bonds are formed within the products. This overall heat change is given the symbol **∆H**.

Change in enthalpy (∆H)
is the difference in energy content of the reactants and the products
in a chemical change.

The enthalpy change associated with a chemical reaction is measured in **kilojoules per mole (kJ mol^{-1}** or **kJ/mol)**.

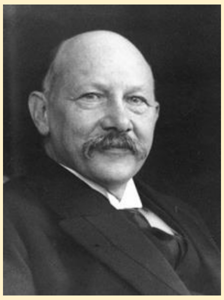
An **exothermic reaction** is one in which the overall reaction liberates heat energy. The ∆H value for an exothermic reaction has a **negative** value.

An **endothermic reaction** is one in which the overall reaction requires heat energy to be absorbed. The ∆H value for an endothermic reaction has a **positive** value.

Enthalpy of a reaction

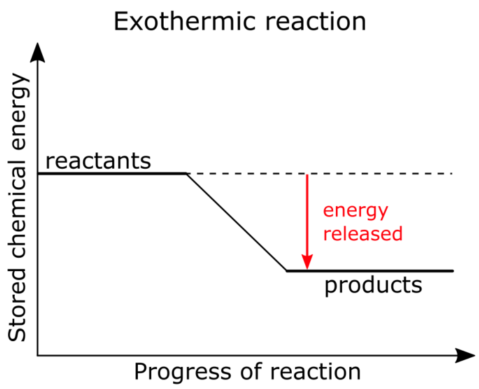
The term enthalpy, coming from Greek words that mean 'warmth' and 'heat', was introduced in thermodynamics by Heike Kamerlingh Onnes (Dutch physicist) in 1909. It refers more specifically to the heat content at constant pressure, but this is very difficult to measure accurately.

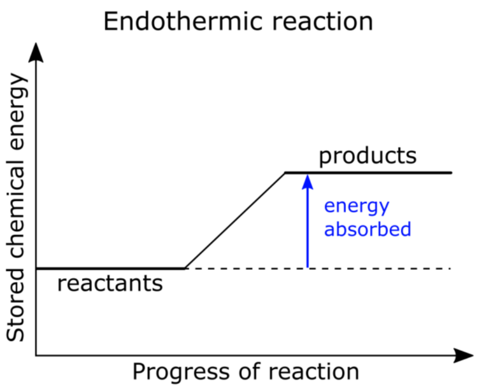
The symbol H was previously used to represent heat in general, but in 1922 it was suggested that the use of this symbol should be strictly limited to enthalpy.



2.1.5 Energy level diagrams

Energy level diagrams depict the changes in energy during the process of a reaction. These diagrams portray the levels of energy of the components of a chemical reaction as this progresses from reactants to products. The energy level diagrams for exothermic and endothermic reactions are illustrated and explained below.

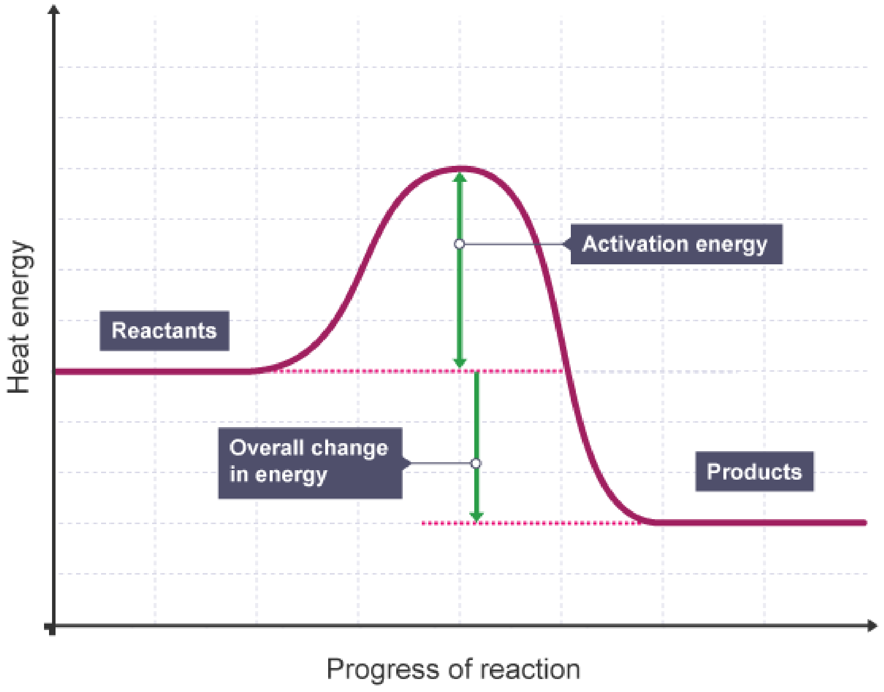

Exothermic reaction


Endothermic reaction

ENDOTHERMIC reaction	EXOTHERMIC reaction
• **Reactants** have **less** energy than the **products.** • The **difference in energy** is **taken in** (absorbed). **ΔH is Positive** • **Temperature decreases** with progression of reaction.	• **Reactants** have **more** energy than the **products.** • The **difference in energy** is **given off** (released). **ΔH is Negative** • **Temperature increases** with progression of reaction.

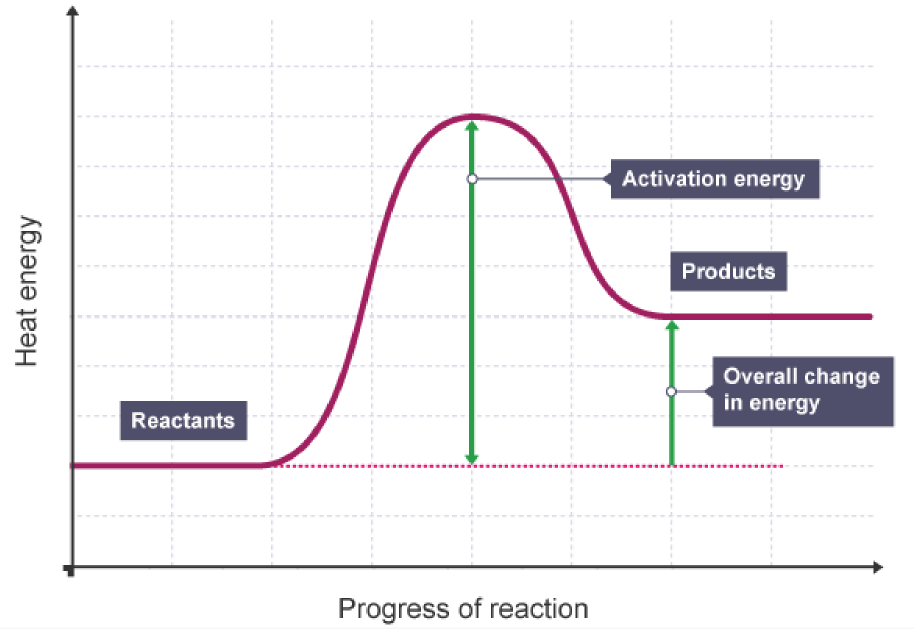
Energy is required to initiate a reaction. This energy is called **activation energy (E_A)**. Activation energy is necessary to start the bond breaking process of the stable reactant molecules. Activation energy is the minimum energy needed to break enough bonds to get a reaction started. All reactions need energy to start even exothermic ones!

An exothermic reaction is one in which heat energy is liberated to the surroundings. The overall enthalpy change of an exothermic reaction has a negative value. This means that, as the reaction progresses between reactants and products, energy is lost to the environment. So, the energy level of the reactants is greater than the energy level of the products. We can also say that the energy released during bond formation of the products is larger than the energy absorbed to break the bonds of the reactants. The difference in energy levels between reactants and products is the change in enthalpy (ΔH).



Source: https://www.bbc.co.uk/bitesize

An endothermic reaction is one in which heat energy is required for the reaction to occur. The overall enthalpy change (ΔH) of an endothermic reaction has a positive value. This means that, as the reaction progresses between reactants and products, energy is gained. Thus, the energy level of the reactants is lower than the energy level of the products. We can also say that the energy released to form the new bonds of the products is smaller than the energy needed to break the bonds of the reactants. The difference in energy levels between reactants and products is the change in enthalpy (ΔH).

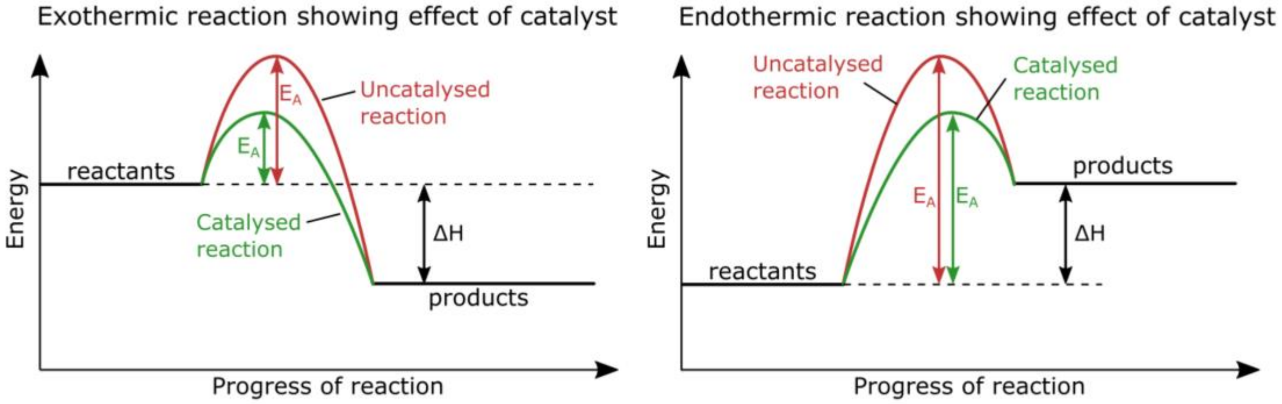


Source: https://www.bbc.co.uk/bitesize

2.1.6 Catalysts and energy level diagrams

A catalyst is a chemical which, when introduced in a chemical process, will increase / alter the rate at which the reaction occurs. Catalysts are specific in nature that is they are only functional for one or a few chemical reactions. Catalysts that increase the rate of a chemical reaction offer an alternative pathway which has a lower activation energy.

If one had to illustrate the use of the catalysts on the previous energy level diagrams, the effect of the catalyst is depicted as shown below.



GCSE Chemistry – Exothermic and Endothermic Reactions #43

Further reading on energy level diagrams.

2.1.7 Energy changes

Energy changes can be classified according to the type of reaction to which they refer. Reference is going to be made to two such energy changes.

2.1.7.1 Heat of Combustion

The heat of combustion (ΔH_C) of a substance is defined as:

the heat change which takes place
when one mole of substance is completely burnt in oxygen.

- ΔH_C is always exothermic.
- E.g. burning of fuels (e.g. butane):

$$2C_4H_{10}(g) + 8O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)\ \Delta H_C = -2877\ \text{kJ mol}^{-1}$$

- The value for the change in heat of combustion for different fuels has different values.

2.1.7.2 Heat of Neutralisation

Heat of neutralisation (ΔH_N) is defined as:

the energy released upon the formation of one mole of water
when an acid reacts with an alkali.

- ΔH_N is always exothermic.
- For any strong acid **and** any strong alkali, the energy change is the same since it occurs for the common reaction:

$$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)\ \Delta H_N = -\ 57\ \text{kJ mol}^{-1}$$

- This value is less if a weak acid and/or alkali is used. This is so because some of the energy released is used to produce more soluble acid / alkali.

2.2 Determining the heat change of reactions

By the end of this section, you should be able to:

- Carry out experiments to compare energy released by different food samples. (15.2c)
- Determine the heat of combustion of different food samples (in kJ g^{-1}). (15.3c)
- Calculate the heat of combustion of a fuel (in kJ mol^{-1}). (15.3d)
- Calculate the heat of neutralisation (in kJ mol^{-1}). (15.3e)
- Carry out experiments to determine the change in heat (in kJ mol^{-1}). *(Limited to combustion of safe liquid fuels and neutralisation of an acid with an alkali.)* (15.3f)

To determine the heat change of a reaction, one must be able to measure the quantity of heat absorbed or produced by the surroundings in relation to the quantity of substance that is reacting.

The way an experiment is set up depends on the kind of reaction, for example whether it is a combustion or a neutralisation reaction. This can be done for other types of reaction, however here we shall be dealing with combustion and neutralisation only.

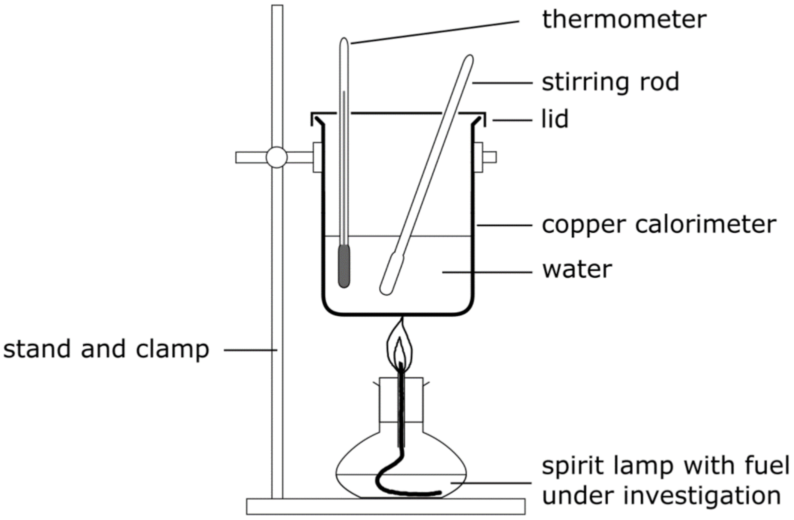
2.2.1 Experimental determination of the heat of combustion

The heat of combustion can be measured for any substance that burns. We shall be dealing with the determination of the heat of combustion of ethanol, which is a common liquid fuel, as well as other foodstuffs.

2.2.1.1 To determine the heat of combustion of a liquid fuel

Method

1. A measured amount of water is placed in a metal container (calorimeter).
2. The initial temperature of the water is recorded.
3. A spirit lamp filled with fuel is accurately weighed, after which it is placed under the calorimeter and lit.
4. The water is stirred frequently to allow even distribution of heat.
5. When about 10 °C rise in temperature is recorded, the spirit lamp's flame is put out.
6. The highest temperature reached is recorded.
7. The spirit lamp is weighed again, and the mass of fuel burned is calculated.



Precautions

1. The metal calorimeter must be thin-walled so that heat absorbed by it is minimal.
2. The spirit lamp is placed as close as possible to the bottom of calorimeter to facilitate heat transfer to the liquid.
3. Enough air circulation must be allowed around the flame to avoid incomplete combustion of the fuel.
4. The water is stirred frequently to allow even distribution of heat in the water.
5. The thermometer must not touch the bottom of the calorimeter as this would be hotter and false temperature readings are recorded.
6. When the spirit lamp is switched off, some time is allowed before reading the final temperature as this will momentarily continue to rise. The highest temperature reached must be recorded.
7. The calorimeter must be surrounded with an insulation barrier to reduce heat losses as much as possible. This will also act as a wind shield. This barrier is not shown in the diagram.
8. A lid is used to minimise heat loss by evaporation of water.
9. A copper calorimeter is used as it transfers heat to the water easily.

Sample results

Mass of spirit lamp before experiment: 110.85 g

Mass of spirit lamp after experiment: 110.62 g

Mass of ethanol burnt: 110.85 – 110.62 = 0.23 g

Initial temperature: 22 °C

Final temperature: 37 °C

Change in temperature: 37 – 22 = 15 °C

Mass of water in calorimeter: 100 g

Calculations

Given data: Specific heat capacity of water = 4.2 J $°C^{-1}$ g^{-1}

1. Calculate heat absorbed by water:
   $H = m\ c\ \Delta\theta = 100 \times 4.2 \times 15 = 6300$ J = 6.3 kJ

2. Calculate RMM for ethanol, C_2H_5OH:
   $(12 \times 2) + (1 \times 6) + (16 \times 1) = 46$

3. Calculate the heat of combustion:
   Recall that ΔH_C is the heat evolved when 1 mole of fuel is burned.

Assuming NO HEAT LOSSES, i.e. ALL heat generated by the burned fuel is absorbed by the water in the calorimeter.

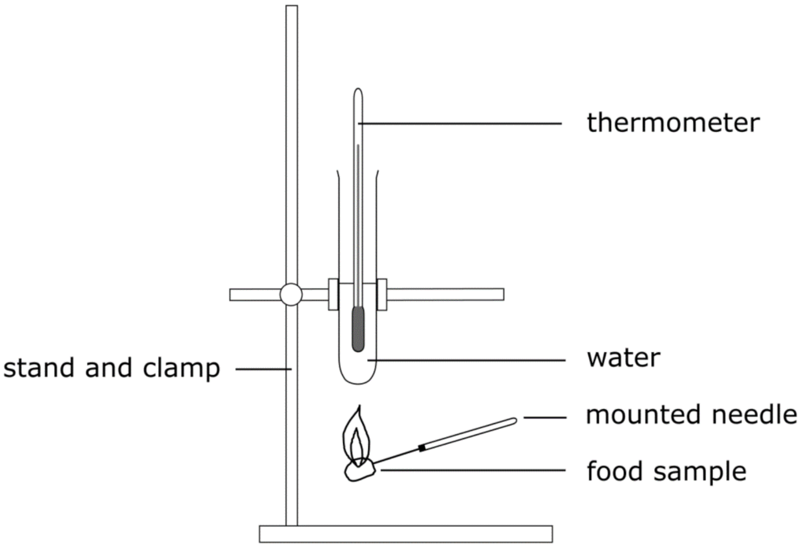
0.23 g ethanol liberated 6.3 kJ of heat

46 g (mass of 1 mole) would liberate (6.3 x 46) / 0.23 = **1260 kJ**

Since the reaction is **exothermic** (as temperature rises): $\mathbf{\Delta H_C = -1260\ kJ\ mol^{-1}}$

2.2.1.2 To determine the heat of combustion of a food stuff

To perform this experiment the apparatus is set up as shown in the diagram below:



Method

1. A known volume of water is measured in a measuring cylinder and added to a boiling tube.
2. A small quantity of the food stuff (pasta) to be investigated is weighed and its mass noted.
3. The initial temperature of the water is recorded.
4. The food is lit and allowed to burn while being held immediately underneath the boiling tube.
5. The water in the boiling tube is heated directly until the food stuff burns completely and extinguishes itself.
6. The final, maximum temperature of the water in the boiling tube is measured.
7. The heat of combustion of the food is measured by using the equation $H = m\,c\,\Delta\theta$ where m is the mass of water heated, c is the specific heat capacity of the water and $\Delta\theta$ is the rise in temperature.

Sample results

- Mass of pasta: 1.14 g
- Mass of water: 100 g
- Initial temperature of water: 20 °C
- Final temperature of water: 36 °C
- Temperature change = 36 -20 = 16 °C

Calculations

To calculate the change in heat of combustion of the food stuff, follow the instruction given in the previous section but instead of using moles, use grams as the unit of measurement for mass.

Step 1: Calculate Heat Absorbed by Water

Using the formula for heat absorbed: $q = m \times C \times \Delta T$

where:

- q = heat absorbed (in joules)
- m = mass of water (100 g)
- C = specific heat capacity of water (4.2 J g^{-1} $°C^{-1}$)
- ΔT = change in temperature (16 °C)

Substituting the values:

$$q = 100\,\text{g} \times 4.2\ \text{J g}^{-1}\ °\text{C}^{-1} \times 16\ °\text{C} = 6720\ \text{J} = 6.72\ \text{kJ}$$

Step 2: Calculate Heat Released per Gram of Food

Next, calculate the heat released per gram of the foodstuff:

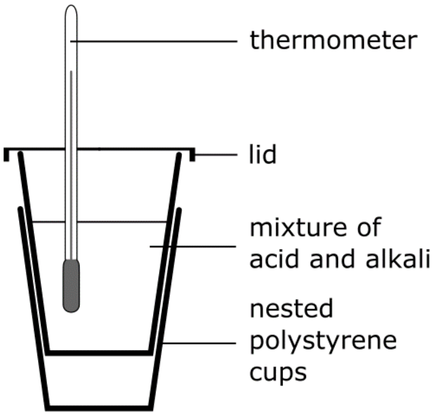
Heat released per gram = q / mass of food = 6720 J/1.14 g = 5894.74 J/g = 5.89 kJ/g

This experiment can be repeated for different food stuffs to compare the change in heat of combustion per gram of different food stuffs.

Watch this video clip showing how to determine the change in heat of combustion.

2.2.2 To determine the heat of neutralisation

To calculate the heat of neutralisation of a strong acid reacting with a strong alkali, the following apparatus is set up. The following example deals with the reaction of 50 cm^3 of 1.0 mol dm^{-3} sodium hydroxide solution with 25 cm^3 of 2.0 mol dm^{-3} hydrochloric acid.



Method

1. Exactly 50 cm^3 of sodium hydroxide solution were placed in a plastic cup or beaker nested in a Styrofoam cup for better insulation.
2. The initial temperature is recorded.
3. Exactly 25 cm^3 of hydrochloric acid solution, at the same temperature, are quickly added to the alkali and the cup covered with a lid.
4. The mixture is stirred thoroughly, and the highest steady temperature reached is recorded.

Precautions

- Calorimeter must be made of poor conducting material in order to keep heat generated by the reaction inside it thus reducing heat losses.
- Mixture is stirred so that even distribution of heat occurs.
- A cardboard cover lid reduces heat losses.
- A digital thermometer can be used for more accurate temperature readings.

Sample results

Volume of NaOH(aq) used:	50 cm^3
Volume of HCl(aq):	25 cm^3
Total volume of mixture:	75 cm^3
Mass of NaOH(aq):	50 g
Mass of HCl(aq):	25 g
Total mass of mixture: 50 + 25 =	75 g
Initial temperature:	22.0 °C
Final temperature:	31.1 °C
Change in temperature: 31.1 – 22.0 =	9.1 °C

Assumption: Density of NaOH(aq) and HCl(aq) are similar to that of pure water, i.e. 1 g cm^{-3}.

Specific heat capacity of solution is similar to that of water that is 4.2 J $°C^{-1}$ g^{-1}.

(This assumption can be made if solutions used are dilute.)

Calculations

Given data: Specific heat capacity of water = 4.2 J $°C^{-1}$ g^{-1}

1. Calculate heat evolved by reaction mixture:
   $H = m\ c\ \Delta\theta = 75 \times 4.2 \times 9.1 = 2866.5$ J $= 2.8665$ kJ

2. Write a balanced equation for reaction:
   $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

3. Calculate moles of NaOH and HCl used:
   $0.05\ dm^3 \times 1.0\ mol\ dm^{-3} = 0.05$ moles NaOH
   $0.025\ dm^3 \times 2.0\ mol\ dm^{-3} = 0.05$ moles HCl

4. Calculate heat of neutralisation:
   Recall ΔH_N is the heat evolved when 1 mole of H_2O is produced.
   Mole ratio from balanced equation: NaOH : HCl : H_2O = 1 : 1 : 1
   So, 0.05 moles NaOH react with 0.05 moles HCl to produce 0.05 moles of H_2O
   Assuming NO HEAT LOSSES.
   Production of 0.05 moles H_2O liberates: 2.8665 kJ of heat
   Producing 1 mole H_2O would liberate: 2.8665 / 0.05 = 57.33 kJ
   Since reaction is **exothermic** (as temperature rises): **$\Delta H_N = -57.33$ kJ mol^{-1}**

Perform a virtual experiment to determine the heat of neutralisation.

Substances from the Earth

Aquatic Environments

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The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

3 Dynamic equilibria – LO 12

3.1 Different kinds of chemical reactions

By the end of this section, you should be able to:

- Classify reactions as acid-base, combustion, thermal decomposition, and precipitation. (12.2a)
- Classify reactions as displacement and/or redox. (12.3a)
- Describe changes of state as an example of a reversible change. (12.1b)
- Describe reversible changes such as hydration of copper(II) sulfate and thermal dissociation of ammonium chloride. (12.2b)

3.1.1 Classification of chemical reactions

Most chemical reactions can be classified into one or more of five basic types:

1. Acid-base (Neutralisation)
2. Combustion
3. Thermal decomposition
4. Precipitation
5. Displacement

Further reading about classification of chemical reactions.

Classification of chemical reactions

Combustion, fermentation, and the reduction of ores to metals are some of the chemical reactions that have been used since ancient times. Greek philosophers made their own theories about the transformation of materials. During the Middle Ages the transformation of substances was studied by alchemists, who had a particular interest in finding a way that could convert lead to gold. We now know that this is unachievable.

3.1.1.1 Acid Base reactions

To be able to recognize this type of reaction you must first be able to identify the base and the acid. An acid is a substance like hydrochloric acid that dissolves in water to produce H^+ ions, and a base is a substance like sodium hydroxide that dissolves in water to produce hydroxide (OH^-) ions or like copper(II) oxide that has oxide ions (O^{2-}). So, in acid base reactions the H^+ of the acid combines with the OH^- or O^{2-} of the base to form water. The anion of the acid joins the cation of the base to form what we call a salt. Both the salt and water have a pH of 7 and are therefore neutral. Therefore, this type of reaction is also referred to as **neutralisation**.

Acid + Base → Salt + Water

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

$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

Acid base reactions were dealt with in detail in Year 9.

3.1.1.2 Combustion reactions

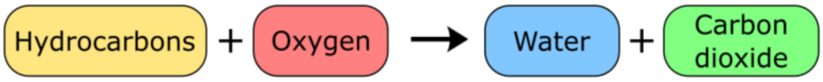
Combustion reactions involve the reaction of an element or compound with oxygen to form other substances and heat energy.

For example, magnesium burns in air (oxygen) to form magnesium oxide, heat, and light energy.

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

Another example, hydrocarbons combust in an abundant amount of oxygen to form carbon dioxide, water, heat, and light energy.

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



3.1.1.3 Thermal decomposition reactions

Thermal decomposition happens when one reactant breaks into two or more products. Such reactions involve the use of heat. Below are some examples of this kind of reaction:

$$Ca(HCO_3)_2 \rightarrow CaCO_3 + CO_2 + H_2O$$

$$CaCO_3 \rightarrow CaO + CO_2$$

3.1.1.4 Precipitation reactions

Precipitation reactions are recognised when two soluble compounds react producing two new substances, one of which is an insoluble compound.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

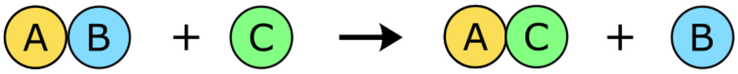
$$Ca(OH)_2(aq) + Na_2SO_4(aq) \rightarrow CaSO_4(s) + 2NaOH(aq)$$

3.1.1.5 Displacement reactions

Displacement reactions happen when a more reactive element replaces a less reactive element from its compound. The type of reactants for a displacement reaction are an element and a compound. Displacement reactions can occur both in metals and non-metals. Refer to the chapter on the reactivity series in Year 10 to be able to determine whether displacement is possible or not.

$$CuSO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + Cu(s)$$

$$2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(aq)$$

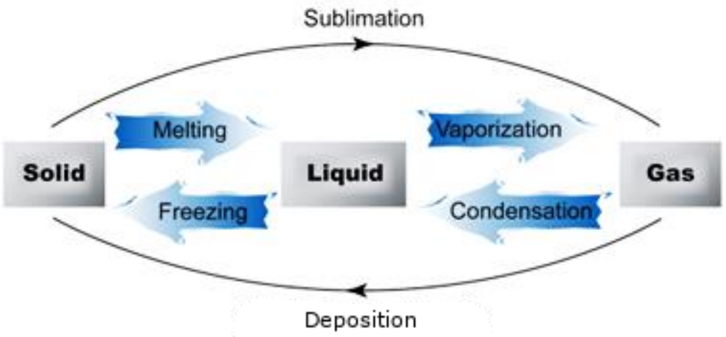


Watch a lesson by Mr Bugeja on the classification of chemical reactions.

3.1.2 Reversible changes

3.1.2.1 Reversible physical changes

A physical reversible change is a change where **no new materials** are created, and the original material can be recovered easily. Examples include freezing water to make ice or melting chocolate.



Adapted from: https://www.coloridoprinting.com/news/what-is-sublimation-printing

When boiling water to turn it into water vapour, it certainly looks like the water has become something else entirely. However, if you zoom in closer, you can see that the water molecules remained the same!

3.1.2.2 Reversible chemical changes

Chemical reactions are represented by chemical equations. These equations typically have a unidirectional arrow (→) to represent irreversible reactions. Other chemical equations may have a two-way arrow (⇌) that represent reversible reactions.

Examples of reversible chemical reactions are:

Hydration of copper(II) sulfate

White anhydrous copper(II) sulfate changes to blue hydrated copper(II) sulfate with the addition of water. However blue hydrated copper(II) sulfate changes to white anhydrous copper(II) sulfate with the application of heat. Since hydrated copper(II) sulfate can be converted to anhydrous copper(II) sulfate and vice versa, this is called a **reversible reaction** and is shown by the ⇌ sign.

	Anhydrous copper(II) sulfate	**Hydrated copper(II) sulfate**
Formula	$CuSO_4$	$CuSO_4.5H_2O$
Colour	White	Blue
What happens on exposure to water vapour?	Turns blue	No change
What happens on the application of heat?	No change	Turns white

The bond that exists between the salt and the water of crystallisation in hydrated salts is a chemical bond. Therefore reversible chemical changes can be classified as such because the bond between the water molecules and the cations and anions in the hydrated salt need to be broken and reformed when converting from the hydrated to the anhydrous form and vice versa.

Thermal dissociation of ammonium chloride

Ammonium chloride is a solid that undergoes thermal dissociation on heating. It dissociates (splits) into a mixture of two colourless gases, ammonia and hydrogen chloride. On cooling ammonia and hydrogen chloride recombine to form ammonium chloride.

$$NH_4Cl \rightleftharpoons NH_3 + HCl$$

Another example follows:

$$N_2O_4 \rightleftharpoons 2NO_2$$

Note that there is an important difference between thermal decomposition and thermal dissociation. Both reactions involve the breakdown of the chemical substance by heating. However, while **thermal decomposition is irreversible**, **thermal dissociation is reversible** because the products easily recombine on cooling to reform the original substance.

It was believed that all chemical reactions were irreversible until 1803, when French chemist Claude Louis Berthollet introduced the concept of reversible reactions. Initially he observed that sodium carbonate and calcium chloride react to yield calcium carbonate and sodium chloride; however, after observing sodium carbonate formation around the edges of salt lakes, he realized that large amount of salts in the evaporating water reacted with calcium carbonate to form sodium carbonate, indicating that the reverse reaction was occurring.

Further reading on reversible changes.

3.2 Dynamic equilibria

By the end of this section, you should be able to:

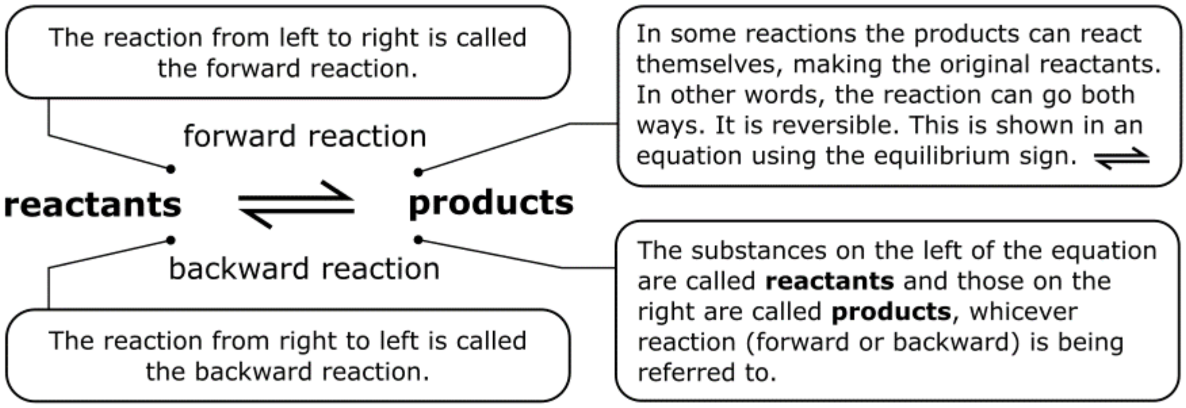
- Use the appropriate symbol to represent a reversible change. (12.1c)
- Explain how some chemical reactions in closed conditions do not go to completion but reach dynamic equilibrium. (12.3c)
- Explain how changing temperature or pressure affects the position of equilibrium in a reversible reaction. (12.3d)

To be able to understand the concept of dynamic equilibrium we are going to look at the reaction which takes place in the Haber process.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

1. At the start there is no ammonia.
2. As the nitrogen and hydrogen react, ammonia starts to form. (Forward reaction)

3. As the concentration of ammonia builds up inside the **closed container**, the dissociation of ammonia starts to take place. (Backward reaction)
4. There is a **forward reaction** and a **backward reaction** taking place simultaneously.
5. Eventually a state is reached in which the **rate of combination** of hydrogen and nitrogen **is equal** to the **rate of dissociation** of ammonia.



The reaction from left to right is called the forward reaction.

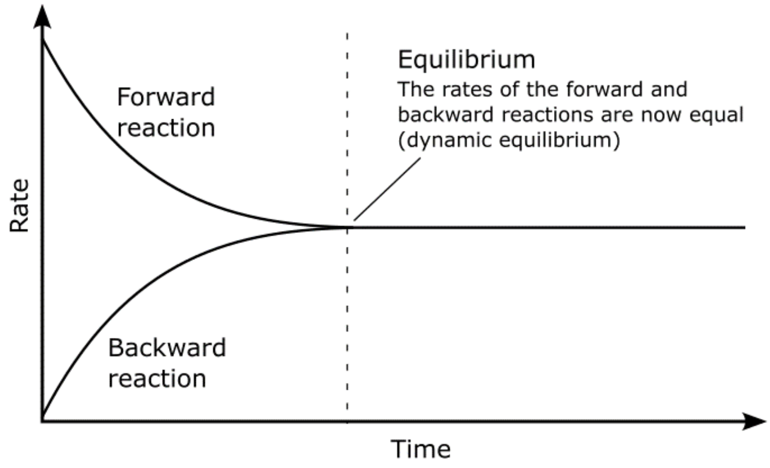
In some reactions the products can react themselves, making the original reactants. In other words, the reaction can go both ways. It is reversible. This is shown in an equation using the equilibrium sign. $\rightleftharpoons$

The reaction from right to left is called the backward reaction.

The substances on the left of the equation are called **reactants** and those on the right are called **products**, whicever reaction (forward or backward) is being referred to.

When the forward and the backward reaction reach the same rate, the system is said to be in **equilibrium**. Chemical equilibria are **dynamic** because the particles are constantly interchanging, while the individual quantities of the reactants and products remain constant. Note that at equilibrium, the quantity of reactants and products may not be equal. This depends on the nature of the chemical reaction.

Dynamic equilibrium is reached when the rate of the forward reaction equals the rate of the backward reaction.
It can only be established if the reaction happens in a closed system, where no substances can enter or leave.



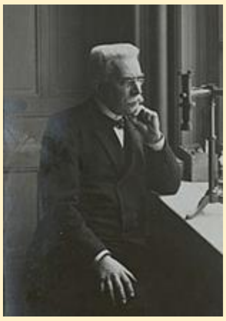
The rates of the forward and backward reactions are now equal (dynamic equilibrium)

This can be explained by a situation which you may be familiar with, that is of a tap filling a container which is losing water. If the rate at which the container is being filled with water is equal to the rate at which the container is losing water, then the level of water in the container does not change. Yet, the water in the tank is being replaced continuously!

Watch this video from Fuse School on dynamic equilibria.

Watch a video by Mr Bugeja on reversible changes.

Dynamic equilibria



Henry Louis Le Chatelier (1850-1936) was a French chemist and engineer, who is best known for developing Le Chatelier's principle (published in 1884). This principle explains how a system in chemical equilibrium can be altered when a change in a condition is applied. Growing up Le Chatelier was greatly influenced by his father, who was an established engineer and chemist, and who had worked on the industrial process to produce aluminium from bauxite. As a career, Le Chatelier chose to teach chemistry, in Paris, while continuing his research and publications related to various industrial processes.

3.2.1 Shifting a reaction in equilibrium

When a reversible reaction is in dynamic equilibrium and conditions such as pressure or temperature are altered (changed), the equilibrium will shift to reach a **new equilibrium position**. As the conditions change, the rate of the forward and backward reactions are affected in such a way that they do not remain equal. Subsequently, a new equilibrium position is reached as the rates of the forward and backward reactions become equal again with a different value.

Depending on the change, the equilibrium can be **shifted towards the right** if the rate of the forward reaction is greater than the rate of the backward reaction. This results in the formation of more products.

Alternatively, the equilibrium can be **shifted towards the left** if the rate of the backward reaction is greater than the rate of forward reaction. This results in the formation of more reactants.

In reactions where the products are needed, it is important to shift the equilibrium to the right so that more products are formed. This is the case in various industrial processes.

It is important to note that the use of catalysts in these processes shortens the time needed for the equilibrium to be reached, however catalysts do not alter the equilibrium position. This means that with or without the catalyst given enough time the reaction will reach the same equilibrium position containing the same concentrations of reactants and products.

3.2.1.1 Changing pressure

One can predict what happens to the amount of product in an equilibrium mixture of **gases** if the **pressure** is changed. All you have to do is count the molecules in the equation!

If the **pressure is increased** the equilibrium will shift in the direction of the fewest moles of gas.

Alternatively, if the **pressure is decreased** the equilibrium will shift in the direction of the highest moles of gas.

Consider the equation below:

$N_2O_4(g)$	$\rightleftharpoons$	$2NO_2(g)$
1 mole of gas		2 moles of gas
22.4 dm^3 at STP		44.8dm^3 at STP
No colour		Brown
Low pressure side		High pressure side

In the reaction above, increasing the pressure will cause the equilibrium to shift to the left (the side of less molecules). The amount of N_2O_4 will increase while the amount of NO_2 will decrease. This is observed as the brown colour of the NO_2 becomes lighter.

If on the other hand the pressure is decreased. The equilibrium will try to shift to the right (the side of more molecules). The amount of NO_2 will increase and the amount of N_2O_4 will decrease. This is observed as the brown colour of NO_2 becomes darker.

Let us consider another example:

$2SO_2(g) + O_2(g)$	$\rightleftharpoons$	$2SO_3(g)$
3 moles of gas		2 moles of gas
High pressure side		Low pressure side

Increasing the pressure will shift the equilibrium to the right (the side of less molecules). The amount of SO_3 will increase. Lowering the pressure will cause the equilibrium to shift to the left (the side of more molecules). So, more SO_2 and O_2 will be produced.

Note that in reactions in which the **number of reactant molecules** is equal to the **number of product molecules,** change in total pressure will have no effect on the position of equilibrium.

Further reading on dynamic equilibria and the factors that shift a reaction in equilibrium.

Watch a lesson by Mr Bugeja on shifting a reaction in equilibrium.

3.2.1.2 Changing temperature

To be able to understand the effect of temperature we are going to look at the Haber process. This is the industrial process used to produce ammonia on a large scale. The equation for the reaction is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$$

The value of ΔH given in reversible reactions ALWAYS refers to the forward reaction. Therefore a negative ΔH value is associated with an exothermic forward reaction.

In this case the forward reaction is exothermic (it gives off 92 kJ of heat energy), so the backward reaction is endothermic (takes in 92 kJ of heat energy).

For a reversible chemical reaction whose ΔH value is positive, the reverse argument applies.

The following table explains what happens when temperature is changed during the Haber process.

	Effect of temperature	Equilibrium
What happens if the temperature of the system increases?	The kinetic energy of both the reactants and products increases. This increases the rate of both the forward and backward reactions.	The equilibrium shifts to the left where a greater proportion of the reactants (hydrogen and nitrogen) is produced.
	However, the rate of the endothermic process increases more than that of the exothermic process. Therefore the backward reaction is favoured.	
What happens if the temperature of the system decreases?	The kinetic energy of both reactants and products decreases. This decreases the rate of both the forward and backward reactions.	The equilibrium shifts to the right where a greater proportion of products (ammonia) is produced.
	However, the rate of the endothermic process decreases more than that of the exothermic process. Therefore the forward reaction is favoured.	

Note that changing temperature has a greater effect on the endothermic processes.

Therefore, it is evident that in order to favour the production of ammonia a lower temperature is needed.

In general, **increasing temperature** of a system in equilibrium will favour the endothermic reaction.

On the other hand, **decreasing temperature** of a system in equilibrium will favour the exothermic reaction.

3.3 Chemical products and related environmental issues

By the end of this section, you should be able to:

- Explain how in the Haber process the best yield of ammonia is obtained by applying compromised conditions with respect to temperature and pressure and the use of a catalyst. *(Values for pressure (200 atm.) and temperature (450 °C) will be given.)* (12.3e)
- Identify needs for chemical products such as ammonia and substances produced from it. *(Limited to fertilizers.)* (12.2f)
- Discuss the environmental issues related to the use and misuse of chemical products such as ammonia and substances produced from it. *(Limited to fertilizers and explosives.)* (12.3f)

3.3.1 The Haber process (synthesis of ammonia)

The reaction between nitrogen and hydrogen gases is used to produce ammonia on an industrial scale. The process is known as the Haber process. The balanced chemical equation follows:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$$

This is a homogeneous reaction (all substances are in the same physical state)

Specific conditions are used in the process so that the equilibrium is shifted to favour the production of ammonia.

3.3.1.1 Effect of pressure

Since the number of moles of gas on the right-hand side of the equation (2 moles of ammonia) is less than the number of moles of gases on the left-hand side (1 mole of nitrogen + 3 moles of hydrogen), the pressure must be increased so that the equilibrium shifts to the right to produce more ammonia. So, a **high pressure of 200 atmospheres** is used.

3.3.1.2 Effect of temperature

The formation of ammonia is exothermic. If the temperature of the system is lowered, the equilibrium will shift to the right because the backward endothermic process is disadvantaged when compared to the forward reaction. So a **low temperature favours the production of ammonia.**

However, if the temperature is too low the rate of production of ammonia would be too slow to be commercially viable. Therefore a **compromise temperature of about 450 °C** is used. This temperature ensures an acceptable yield of ammonia at a reasonable rate.

3.3.1.3 Use of a catalyst

In order to aid the process further, a catalyst is used.

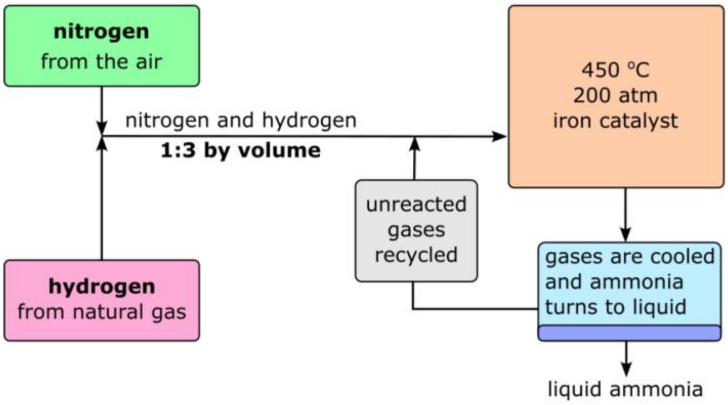
In reversible reactions catalysts increase the rate of both the forward and backward reactions to the same extent. While this does not change the position where equilibrium is reached, it decrease the time it takes for the equilibrium to be reached.

In the Haber process, the catalyst used is **fine powdered iron**.

3.3.1.4 Conditions used in practice

To conclude the following conditions are used in the Haber process:

- very high pressure - 200 atmospheres.
- compromise temperature - 450 °C.
- catalyst - fine powdered iron.



Adapted from: https://www.chemguide.co.uk/physical/equilibria/haber.html

Watch a lesson by Mr Bugeja on the Haber Process.

Take control over a Haber-Bosch Plant.

The Haber process

The Haber Process (also known as The Haber-Bosch Process) was developed at the beginning of the 1900s, by two German chemists Fritz Haber and Carl Bosch. This was done after the demand for fertilizers and feedstocks was continuously increasing and not efficiently supplied. The main challenge that was met by chemists was related to the use of nitrogen in the production of ammonia. Nitrogen although very abundant in air, is also extremely stable under normal conditions.

Haber and Bosch, managed to develop a high-pressure devise (among the other conditions needed) that could be used industrially to convert nitrogen successfully into ammonia, and for this they were both awarded Nobel prizes.

3.3.2 Uses of ammonia and compounds made from ammonia

About 80% of the ammonia produced by industry is used in agriculture to make fertilizer. Plants need nutrients to grow which they absorb from the soil via the plant's root system. Fertilizers provide the major nutrients (nitrogen, phosphorus, potassium and important secondary elements) that plants need. Unless the nutrients are replenished, the soil's productive capacity declines with every harvest.

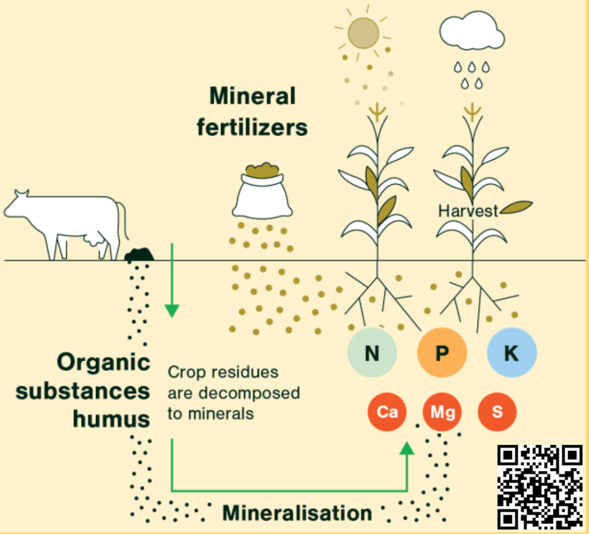
The most common nitrogen fertilizers are:

- ammonium nitrate
- calcium ammonium nitrate
- urea

Watch a lesson by Mr Bugeja on the uses of ammonia.

Ammonia is a basic building block for the compounds in the list above. Before the discovery of the Haber process humans produced less crops and populations suffered hunger. Fritz Haber was the scientist who discovered the Haber process. His work was recognised by the scientific community and in 1918 he received a Nobel prize.

The diagram shows why fertilizers are important to produce crop.



3.3.3 Environmental impact of ammonia and its derivatives

Ammonia serves as a starting material for the production of many commercially important nitrogen compounds.

3.3.3.1 Environmental Impact of Ammonia

Ammonia is produced worldwide, mostly in large-scale manufacturing plants. While it is primarily used as a fertilizer, ammonia also finds application in the production of plastics, fibers, and other chemicals.

Ammonia spills

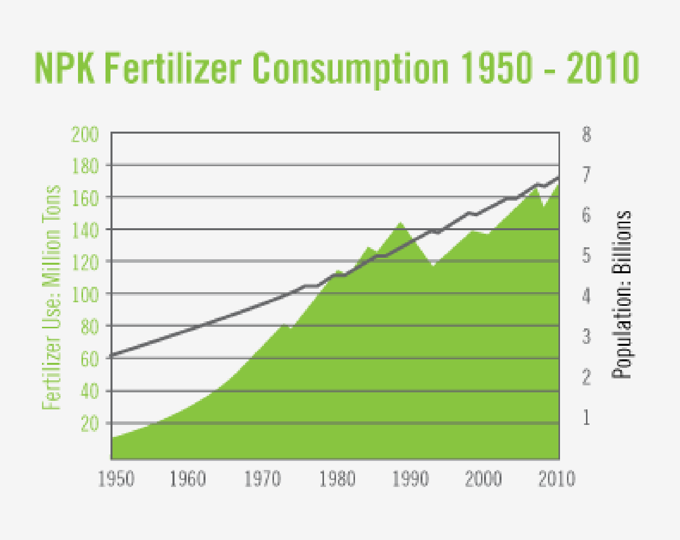
Ammonia spills can have significant environmental impacts. A study found that estuaries, mangroves, and wetlands are particularly sensitive to potential ammonia fuel spills. Fish are most sensitive to ammonia spills, with birds and mammals affected to a lesser degree. An ammonia spill has a relatively smaller dispersion distance and lower persistence within the environment when compared to both heavy fuel oil and maritime gas oil.

Derivatives of ammonia and their impact

Ammonia is used to produce several important chemicals, including urea, ammonium phosphate, ammonium nitrate, and other nitrates. Other chemicals produced from ammonia include nitric acid, hydrazine, cyanides, and amino acids.

3.3.3.2 Fertilizers

The major use of ammonia is as a fertilizer. It is usually applied directly to the soil from tanks containing the liquefied gas. The ammonia can also be in the form of ammonium salts, such as ammonium nitrate, NH_4NO_3, ammonium sulfate, $(NH_4)_2SO_4$, and various ammonium phosphates. Urea, $(H_2N)_2C{=}O$, is the most commonly used source of nitrogen for fertilizer worldwide.



Source: https://earthrestorationfoundation.org/nitrogen-overuse/

The application of nitrogen fertilizers to various crops has been continuously increasing globally. Although nitrogen fertilizer contributes to plant growth, but excessive use of this fertilizer has posed serious threats to environment and human health. Excessive use of nitrogen fertilizers application has a close relationship with nitrate ion accumulation in surrounding environments, groundwater, as well as the sea. Consumption of diets having high nitrate contents has contributed to thyroid conditions, various kinds of human cancers, neural tube defects (during foetus development), and diabetes. Carcinogenic activity and other adverse effects of nitrogen fertilizers compounds might be prevented by consuming vitamin C and antioxidants containing fruits and vegetables.

Environmental impact

The production and use of ammonia-based fertilizers contribute to environmental issues such as eutrophication and acidification. Moreover, ammonia production depends on plentiful supplies of energy, making it one of the most energy-intensive processes.

This picture is showing an algal bloom because of eutrophication. This happens when farmers use excess fertilizers which end up in rivers or the sea causing algae to grow at an abnormal rate.

Source: NOAA Great Lakes Environmental Research Laboratory, CC BY-SA 2.0 via Wikimedia Commons

Ammonium nitrate

Ammonium nitrate (NH_4NO_3) is one of the substances produced from ammonia. It is used as a fertilizer and also in the manufacture of commercial explosives.

Watch a lesson by Mr Bugeja on the environmental issues related to fertilizers.

3.3.3.3 Explosives

Explosives containing ammonium nitrate include Amatex (ammonium nitrate, TNT, and RDX), Amatol (ammonium nitrate and TNT), Ammonal (ammonium nitrate and aluminium powder), ANFO (ammonium nitrate and fuel oil), and others. These explosives have been used in several terrorist acts, including the Oklahoma City bombing in 1995.

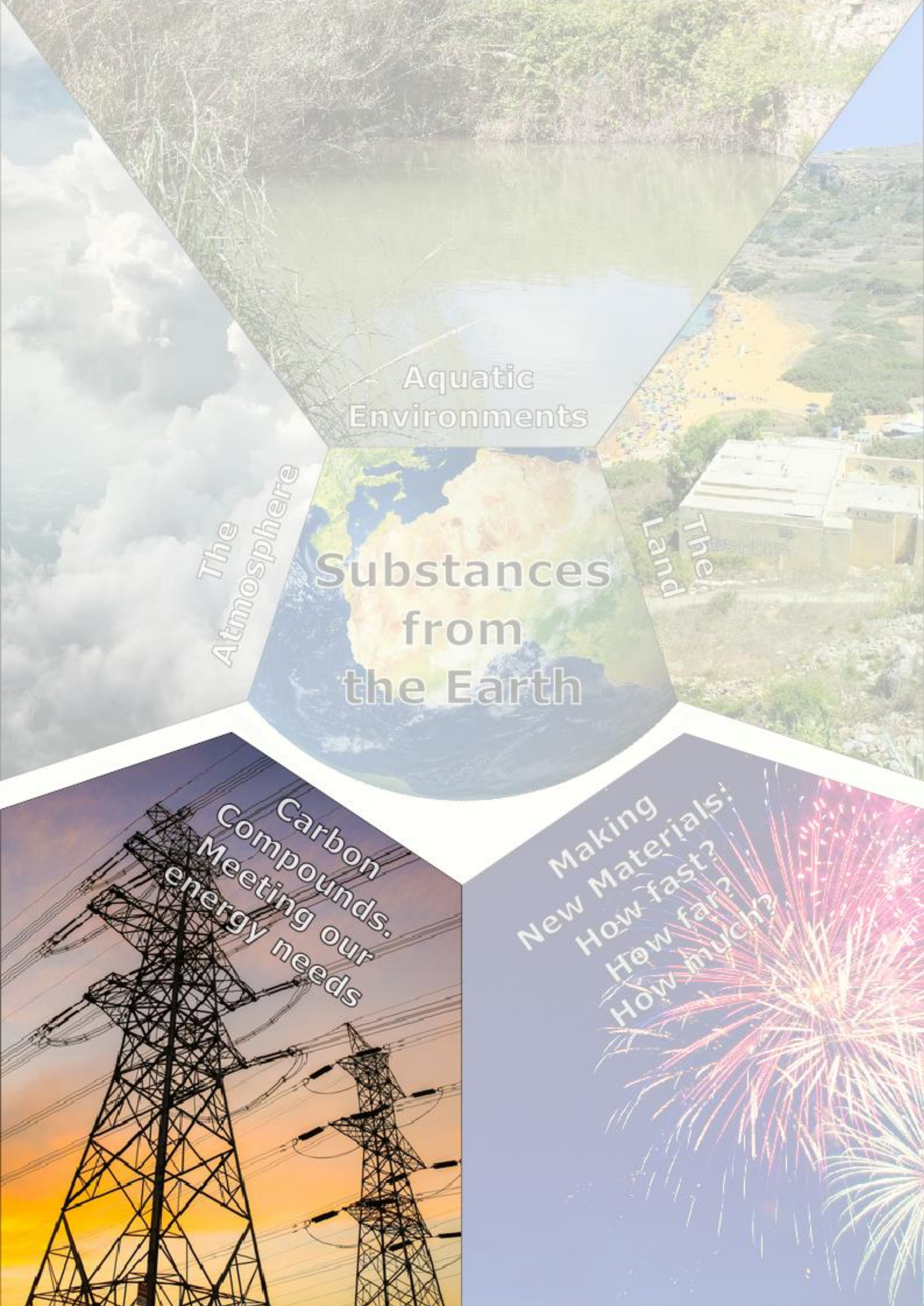
Environmental impact

The misuse of ammonium nitrate in explosives can lead to devastating consequences, as seen in the Beirut explosion in 2020. Moreover, the production and use of ammonium nitrate contribute to environmental issues such as eutrophication and acidification.

Watch a video clip of the Beirut explosion caused by the improper storage of ammonium nitrate.

While ammonia and its derivatives play a crucial role in various industries, their misuse can lead to significant environmental and health issues. Therefore, it is essential to handle these substances responsibly and understand their impact on our environment.

Watch a lesson by Mr Bugeja on ammonium nitrate.



Substances from the Earth

Aquatic Environments

The Atmosphere

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

4 Fossil fuels – LO 13

4.1 The importance of fossil fuel including its risks and benefits

By the end of this section, you should be able to:

- Identify coal, crude oil, and natural gas as fossil fuels. (13.1a)
- Describe the importance of fossil fuels as a source of energy for transport and production of electricity as well as the use of crude oil as feedstock for chemical production. (13.2a)
- Evaluate the importance of fossil fuels as a source of energy for transport and production of electricity as well as the use of crude oil as feedstock for chemical production. (13.3a)
- Present an argument demonstrating that fossil fuels are crucial raw materials and that their control in the world is a possible source of conflict. (13.3b)
- Describe the risks and benefits of the transport of fossil fuels to and storage on an island and the use of crude oil as a finite fuel. (13.2c)
- Evaluate the risks and benefits of the transport of fossil fuels to and storage on an island and the use of crude oil as a finite fuel. (13.3c)

4.1.1 Fossil fuels

Fossil fuels are naturally occurring fuels that are produced over millions of years from fossilized plants and animals.

A fuel is a substance that combusts (reacts with oxygen in air) to release significant quantities of energy (exothermic process) which can either be used as heat energy or converted into other forms of energy *e.g.* electrical and kinetic energy.

There are 3 types of fossil fuels: coal, crude oil, and natural gas. All 3 fuels are extracted from the Earth’s crust.

- **Coal** is a solid material found in deposits of sedimentary rock. Coal is a combustible black or brownish-black sedimentary rock with a high amount of carbon and hydrocarbons. Coal contains the energy stored by plants that lived hundreds of millions of years ago in swampy forests.

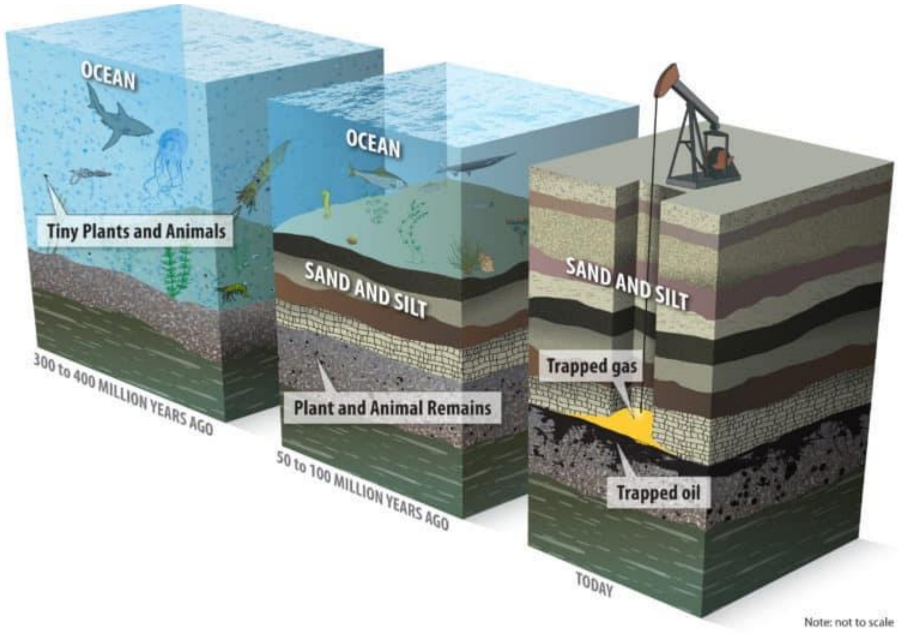
- **Crude oil** consists mainly of hydrocarbons, is often referred to as petroleum.

- **Natural gas**, which mainly consists of methane, is found in gaseous pockets within sedimentary rock.

  Oil and gas are formed from organic material mainly deposited as sediments on the seabed and then broken down and transformed over millions of years.

Sedimentary rock contains dead plant and animal material that has been pressurized over millions of years, due to the weight of other rock and organic material on top of it. This results in the formation of fossil fuels which are considered to be a finite resource.

A **finite (non-renewable) resource** is one which is used at a rate which far exceeds the rate at which it is produced.



Formation of crude oil and natural gas
Source of image: https://stem.guide/wp-content/uploads/2018/09/oilNaturalGasFormation_Primary-768x527.jpg

4.1.2 The importance of fossil fuels

Fossil fuels have been fundamental to the development and growth of modern civilization. They have a wide range of applications, including in transport, electricity production, and as feedstock for chemical production.

Transportation: Fossil fuels, particularly petroleum, are very important for the transportation industry. They power cars, trucks, ships, planes, and trains. The energy density of fossil fuels, their ease of storage and transport, and the existing infrastructure built around them make them a convenient and efficient fuel source for transportation.

Electricity Production: Fossil fuels are also a major source of electricity. Fossil fuels generate a significant portion of the world’s electricity. Power plants burn fossil fuels to heat water, producing steam that drives turbines connected to generators.

Chemical Feedstock: Crude oil is not just a fuel source; it’s also a vital feedstock for the chemical industry. It’s used to produce a wide range of products, including plastics, fertilizers, synthetic fibres, and pharmaceuticals. The complex mix of hydrocarbons in crude oil can be broken down and reassembled into a variety of configurations, making it a versatile raw material.

However, the use of fossil fuels also has significant environmental impacts, including air pollution and climate change. The challenge for the future will be to transition to more sustainable energy sources while still meeting the world’s energy needs.

Watch this video about the far-reaching uses of crude oil & petroleum products.

4.1.3 Fossil fuels as a source of conflict

Being non-renewable and incredibly sought for by several industries worldwide, fossil fuels often lead to conflicts of greed and power between many nations.

Gozo Oil Exploration – late 1990s.
Picture taken by Mr David Bugeja.

4.1.4 The transport and storage of fossil fuels

We must remember that fossil fuels are finite. This means that they are non-renewable. Prices of fossil fuels continue to increase due to the rapid exhaustion of these natural sources of energy. It takes millions of years for fossil fuels to form, but it only takes a relatively short period of time to combust and consume them entirely.

Fossil Fuels 101

Storage

Fuels must be safely stored in appropriate storage tanks. Gaseous fuels are stored in liquefied form. In this form the vaporisation of fuel is minimized, and leakages are avoided. Many fuels are volatile and have very low flashpoints. The flashpoint of a substance is the lowest temperature at which it ignites if a spark is applied (ignition source). Common fuels have very low flashpoints that range from -180 to 65 °C. This means that they easily combust.

Often, large quantities of fuel are safely stored in clearly marked restricted areas so as to present minimal risk to human life.

In Malta fuels are imported by means of specialised ships in liquid form. The possibility for the construction of an undersea pipeline to carry fuel from neighbouring countries is currently being explored.

The **LNG** tanker at Delimara stores **liquefied natural gas** which is then supplied to the Delimara power plant which produces electricity for the whole nation.

Picture taken by Michael Mercieca showing the LNG storage tanker at Delimara

Transportation

		Risks	Benefits
Modes of transportation	**Pipeline**	Many fractions of crude oil are volatile i.e. vaporise to form a gas very easily. Fuels in the gaseous state are extremely flammable. Fuel pipeline distributors must ensure that there is no build-up of excess gas, especially during the hot seasons of spring and summer. Otherwise, this would lead to very fatal incidents.	Fast and continuous delivery of large quantities of fuel to destination.
	Ship	Oil spills can happen during shipment of fuel overseas. Oil spills are detrimental to aquatic life.	Slow and periodic delivery of large quantities of fuel to port of a given destination.
	Lorry/Train	Motor accidents can lead to fatal events in the case of sudden build-up and agitation of fuel in the gaseous state, in a given container.	Fuel is collected from port and delivered to destination.

Malta's national grid is powered by a mix of energy sources:

Natural Gas: As of 2017, most of the electricity generated in Malta was from natural gas, with oil as a backup. Natural gas has only been used for generation on Malta since CCGT systems were installed at Delimara Power Station in 2015.

Oil: Oil has been the primary fuel for electricity generation for many decades before 2015.

Renewable Energies: Renewable energies have a small but growing share of the electricity generation mix. Malta supports mainly PV for electricity production and solar water heaters for heating and cooling.

Electricity Imports: Since 2015, the Malta–Sicily interconnector allows Malta to be connected to the European power grid and import a significant share of its electricity.

A lesson on the importance of fossil fuel including its risks and benefits recorded by David Bugeja, chemistry teacher at Verdala Secondary School.

4.2 Crude oil and its fractions

By the end of this section, you should be able to:
- Define hydrocarbons. (13.1d)
- State that crude oil consists of a mixture of hydrocarbons. (13.1e)
- Describe the uses of fractions obtained from crude oil. *(Students should be able to list the following fractions in this order: refinery gases, gasoline/petrol, naphtha, kerosene, diesel oil, fuel oil and residue. Details of carbon chain length and fraction temperatures are not required.)* (13.2e)
- Describe how crude oil is separated by fractional distillation. (13.3e)
- Distinguish between miscible and immiscible liquids. (13.1f)
- Separate immiscible liquids using a separating funnel. (13.2f)

4.2.1 Crude oil

The word ‘crude’ indicates that the oil is impure and unrefined. In fact, oil as it is first extracted from the Earth’s crust, primarily contains a mixture of various **hydrocarbons**. In crude oil there are also compounds containing nitrogen, oxygen, sulfur, and metals.

Hydrocarbons are covalent compounds that are made up of carbon and hydrogen. These molecules contain carbon-to-carbon and carbon-to-hydrogen covalent bonds.

Crude oil

The uses of the different fractions of crude oil have been documented since ancient history. For example, the walls and towers of Babylon (ancient city which was situated in western Asia, and which existed more than four thousand years ago), were built with a mixture containing natural asphalt.

In other documents of the same era, it was written that in an ancient civilization in eastern Europe there was a well which contained a very thick oil (bitumen) and this was used amongst other things, to waterproof building and constructions.

On the other hand, in the fourth century BC, petroleum was already being used as fuel, in China. Some centuries later, oil wells were discovered, and oil was extracted. The drilling was done with long bamboo poles to which drill bits were attached.

Crude oil is not suitable for burning (as a fuel). As a result, it needs to go through a refining process.

Refining is the process of separating crude oil into fractions. These fractions are then treated to obtain other products.

The refining of crude oil involves 2 processes:

1. **Fractionation** (or fractional distillation).
2. **Cracking**. This will be discussed in more detail in section 5.4.

4.2.2 Miscible and immiscible liquids

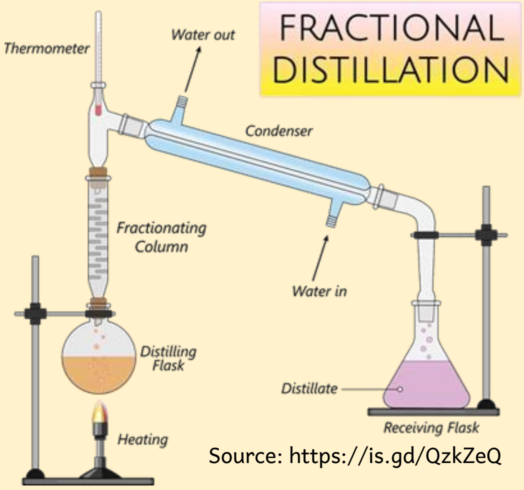
Crude oil is an example of a mixture that contains miscible liquids.

Miscible liquids mix to form a homogeneous mixture.

This type of mixture can be separated into its different components by means of fractional distillation. These components are separated based on their varying boiling points.

This technique can also be used on a small scale in the laboratory to separate miscible liquid mixtures such as ethanol and water.

As the two miscible liquids are heated, both evaporate and flow through the fractionating column which condenses the liquid with the highest boiling point and allows the vapour of the lower boiling point through which is then condensed and collected as the distillate.



Source: https://is.gd/QzkZeQ

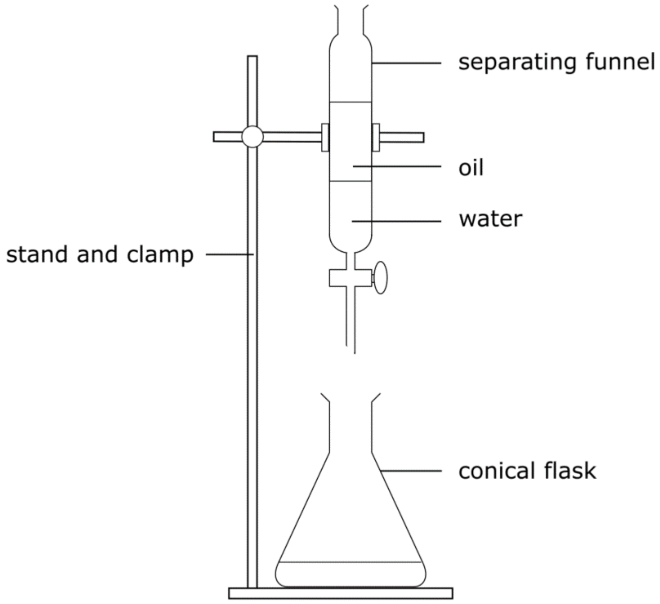
Immiscible liquids do not mix together and form a heterogenous mixture.

Immiscible liquids, on the other hand, do not mix. These liquids form distinct layers. Oil and water are a very common example of a mixture containing immiscible liquids.

Immiscible liquids can be separated by means of a laboratory technique that uses a separating funnel. The mixture is poured into a separating funnel and the liquids should be given enough time to separate. Make sure to leave the separating funnel open.

The following steps need to be carried out to separate the two immiscible liquids:

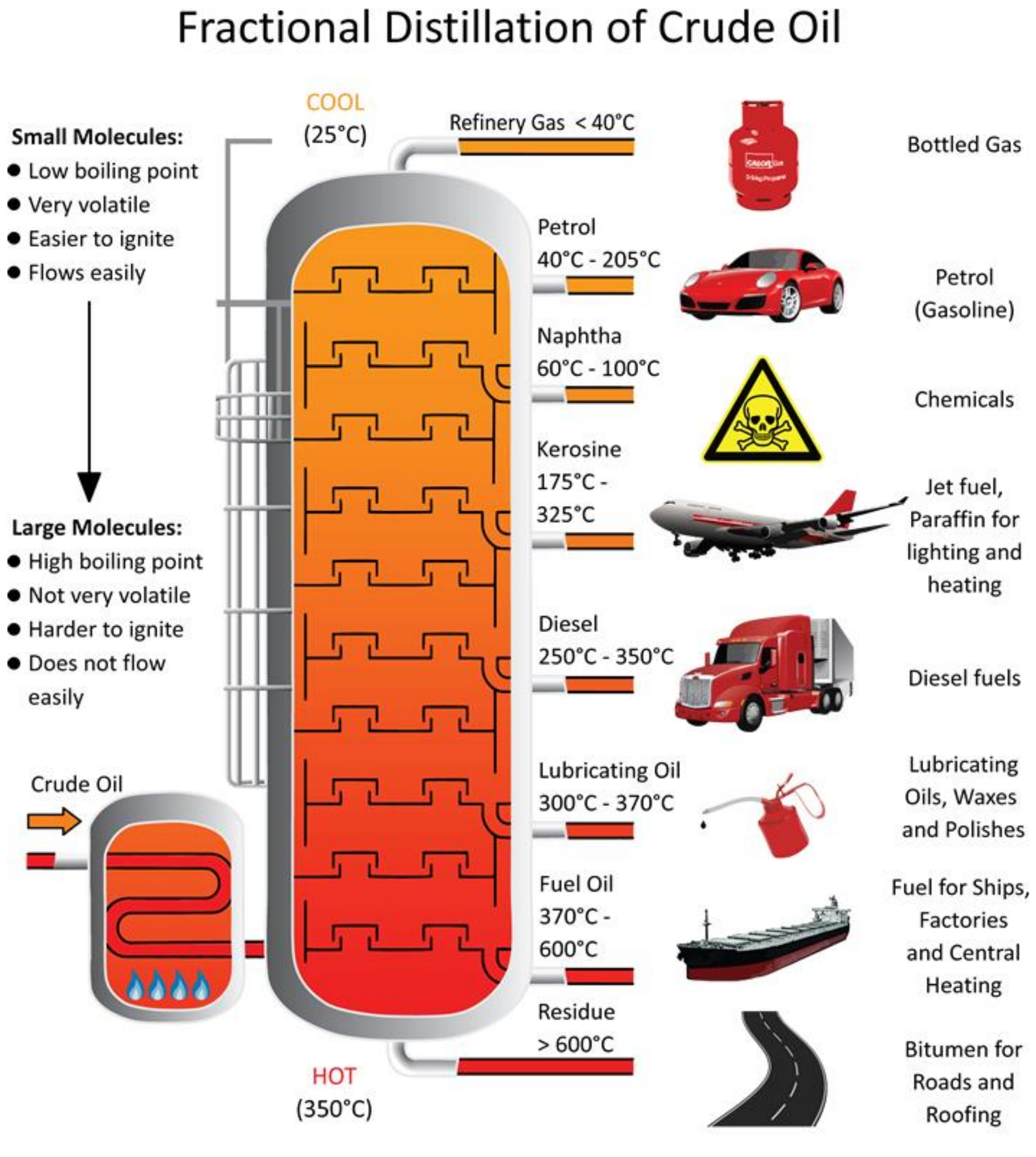
1. Open the separating funnel tap to allow the liquid which settles at the bottom to flow into the conical flask. This liquid is the denser liquid, in this case water.
2. To ensure that the liquid collected is free from contamination from the less dense liquid in this case oil, the tap should be closed when the last quantity of denser liquid reaches the funnel's tap.
3. The middle layer (where both liquids meet) should be discarded.
4. The top layer (less dense layer) can then be collected in a different container.



4.2.3 Fractional distillation of crude oil

The following steps describe how fractional distillation is carried out in industry.

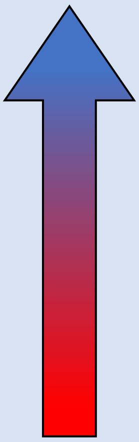
1. The crude oil is preheated to 400 °C until most of its components vaporise entirely.
2. The fractionating column is divided into compartments. The top compartments are cooler than the ones below. The lowest compartment is the hottest, while the highest compartment is the coolest.
3. In the fractionating column, the hydrocarbon components of crude oil are separated by virtue of their varying boiling points.
4. The most volatile components (with low boiling points) are collected from the top compartments. These are usually simple, short-chained hydrocarbons *e.g.* petroleum gases.
5. The least volatile components (with high boiling and/or melting points) are collected from the bottom compartments. These are usually bulkier, carbon-heavy molecules *e.g.* heavy fuel oil used for ships, and bitumen.



Fractionating column

Copyright © 2009 science-resources.co.uk

The main fractions of crude oil and their uses are:

Temperature (°C)	Fraction	Uses of Fraction
COOL (25 °C) 	Refinery Gases / Bottled Gas	Liquefied Petroleum Gas (LPG) in gas cylinders / bottles.
	Petrol / Gasoline	Petrol (Gasoline) for fuel in cars.
	Naphtha	Naphtha for synthesis of industrial materials e.g., solvents, cleaning fuels, painting oils.
	Kerosene (paraffin oil)	Kerosene for jet fuel and paraffin for candle wax and crayons.
	Diesel oil	Diesel for cars, buses, and lorries.
	Fuel Oil	Fuel for ships, power stations, factories, and central heating.
HOT (350 - 400 °C)	Residue	Bitumen for road surfaces and roofing.

Further reading on crude oil and refining.

A lesson on crude oil and its fractions recorded by David Bugeja, chemistry teacher at Verdala Secondary School.

4.3 Pollution related to fossil fuels

By the end of this section, you should be able to:

- Describe the problems of high sulfur content in fossil fuels. (13.1g)
- Discuss the importance of desulfurisation of fuels. (13.2g)
- Describe how the use of fossil fuels contributes to pollution. *(By liberating particulates, carbon monoxide and carbon dioxide during combustion, and oil spills.)* (13.2h)
- Explain how the use of fossil fuels contributes to pollution. *(By liberating particulates, carbon monoxide, carbon dioxide, nitrogen oxides, and sulfur during combustion , and oil spills, etc.)* (13.3h)
- Interpret data on the use of fossil fuels and the gases generated. (13.3i)

There are two kinds of pollutant gases in the atmosphere:

- Those that are present in excessively high concentrations but are original components of air *e.g.* CO_2.
- Those that are not present in clean unpolluted air *e.g.* NO_x

Pollutants can be caused by natural causes such as volcanic eruptions or by humans due to heavy industry and transportation.

The gases of concern are listed in the following table:

Pollutant	Source	Effect	Remedy
Excess carbon dioxide	• Complete combustion of fossil fuels.	• Enhanced greenhouse effect which leads to global warming.	• Alternative fuels *e.g.* hydrogen
Carbon monoxide	• Incomplete combustion of fossil fuels.	• Toxic	• Efficient engines that burn fuels optimally.
Nitrogen monoxide & Nitrogen dioxide	• Combustion at very high temperatures (in vehicles and in power stations).	• NO_2 is toxic. • Causes respiratory problems. • Causes acid rain.	• Catalytic converters* (convert nitrogen oxides into harmless nitrogen).
Sulfur dioxide	• Combustion of fossil fuels.	• Causes acid rain. • Causes lung disease.	• Desulfurisation (removal of sulfur) from fuels.
Particulate matter (PM) e.g. soot from car exhaust and dust created due to tyre wear.	• Vehicles. • Industry. • Power stations.	• Causes respiratory and other health problems.	• Engines that use less fuel. • Responsible driving. • Regular maintenance of vehicles and roads.

* The catalytic converter is a device that is fitted to cars to prevent the release of toxic gases into the atmosphere. It helps convert toxic gases such as CO and NO_x into less harmful ones.

Pollution can also be caused by oil spills. Oil spills can happen during the transport of crude oil when there is an oil tanker accident at sea or by spilling of fuel and other oils into the environment. Oil spills generate harmful volatile gases (VOCs) as well as contaminate water bodies to the detriment of the living things that live in these ecosystems.


Source: https://is.gd/M8Drrn

Volatile organic compounds, or **VOCs** for short, are basically organic chemicals that easily turn into a gas at room temperature. Imagine you have a bottle of nail polish remover. The strong smell you perceive comes from the VOCs in the remover evaporating into the air. VOCs are all around us, in things like paints, cleaning products, and even air fresheners. While some VOCs are harmless, others can irritate our eyes and noses, or even cause headaches and dizziness at high levels. In chemistry, we're interested in VOCs because they can react with sunlight and air pollution to create smog, which can make it harder to breathe, especially for people with asthma.

4.3.1 Combustion of sulfur deposits in fossil fuels

Fossil fuels contain large quantities of sulfur in its elemental form (S_8 commonly written as S) and in compounds *e.g.* hydrogen sulfide (H_2S).

When fossil fuels undergo combustion, sulfur and compounds of sulfur react with oxygen in air to form sulfur dioxide gas (SO_2) as shown below:

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

$$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(g)$$

The sulfur dioxide produced causes respiratory problems as well as acid rain. These were discussed in learning outcome 2 (Year 9 textbook).

The picture shows the old Marsa power station which used to be powered by burning coal. It used to produce huge amounts of pollution related to the production of sulfur dioxide, nitrogen oxides and soot. In 1995, the Marsa power station was converted to run with other fossil fuels in a bid to reduce the amount of pollutants that it used to generate until it was decommissioned in 2015.

Marsa Power Station – late 1990s.
Picture taken by Mr David Bugeja

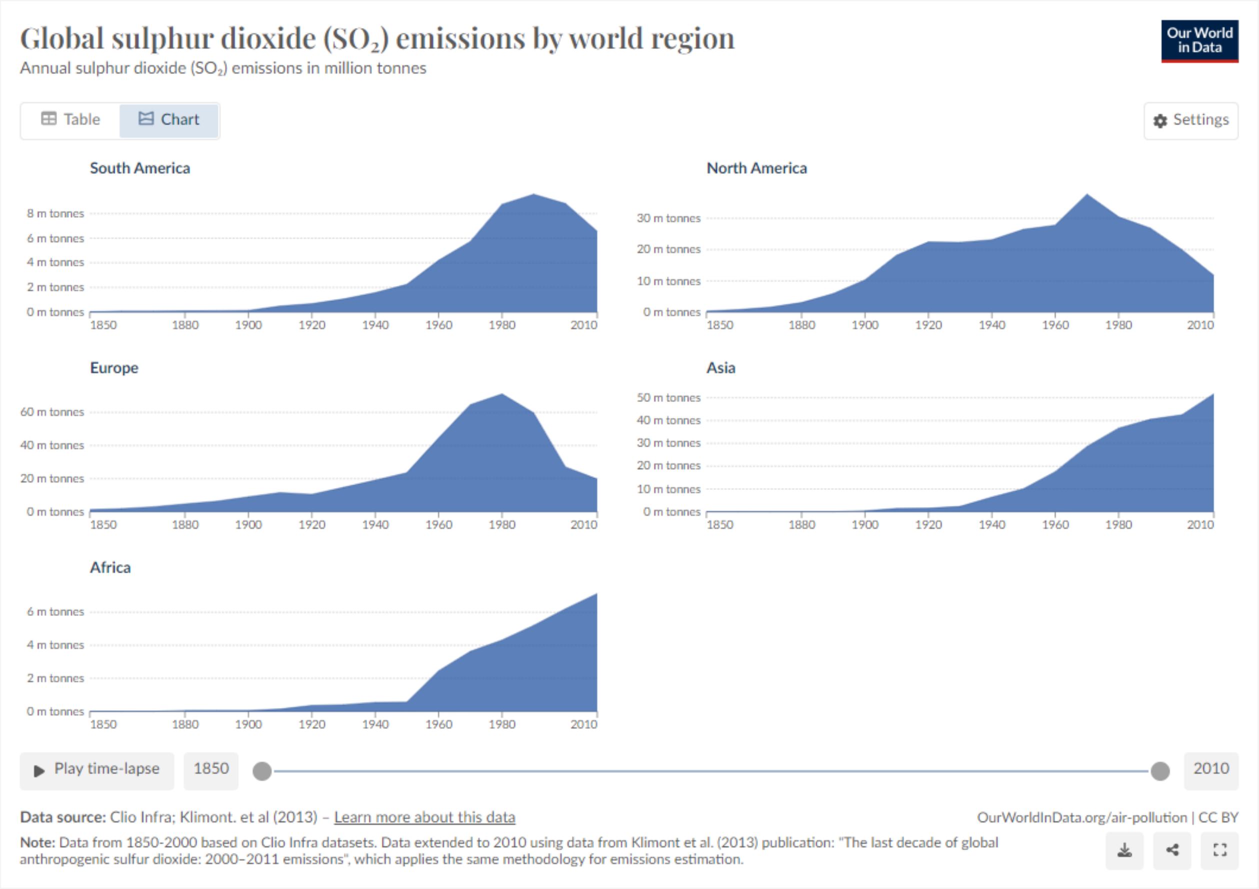
4.3.2 Desulfurisation of fossil fuels

Desulfurization of fossil fuels refers to the process of removing sulfur compounds from coal, crude oil, and natural gas to reduce sulfur emissions during combustion. The primary goal of desulfurization is to mitigate environmental pollution and health hazards associated with sulfur dioxide emissions, which contribute to acid rain, smog formation, and respiratory problems.

In accordance with the EU Industrial Emissions Directive of 2010, countries in the European Union are obliged to regulate pollutant emissions from industrial settings.

To significantly lower the release of sulfur dioxide (SO_2) into the atmosphere, the fossil fuel industry has adopted processes of 'desulfurisation' i.e. the removal of sulfur and its compounds from such fuels.

The following chart shows the global sulfur dioxide production for different areas in the world. It is evident which parts of the world adopted the use of desulfurized fuel. Other factors may be at play for example the increase or decrease of the level of industry in a country.



Global sulphur dioxide (SO_2) emissions by world region

Annual sulphur dioxide (SO_2) emissions in million tonnes

Data source: Clio Infra; Klimont. et al (2013) – Learn more about this data

OurWorldInData.org/air-pollution | CC BY

Note: Data from 1850-2000 based on Clio Infra datasets. Data extended to 2010 using data from Klimont et al. (2013) publication: "The last decade of global anthropogenic sulfur dioxide: 2000–2011 emissions", which applies the same methodology for emissions estimation.

Sulfur dioxide emissions across Europe

The data on the following website ([Air pollutant emissions data viewer | Tableau Public]) portrays the SO_2 emissions of European countries over a period of years until 2015. Use your mouse to toggle over the map, click a country, and look at SO_2 emissions for that particular country.

The Industrial Emissions Directive - Environment - European Commission (europa.eu)

Industrial Emissions Directive - FuelsEurope

Flue Gas Desulfurization | Yokogawa Europe

Substances from the Earth

Aquatic Environments

The Atmosphere

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

5 Organic compounds – LO 14

5.1 Classifying organic molecules

By the end of this section, you should be able to:

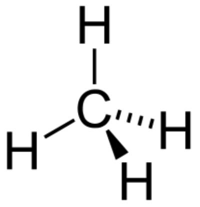
- Explain why carbon is a special element that can form many different compounds that are natural and/or synthetic. (14.2a)
- Define homologous series. (14.2b part)
- Use the terms homologous series, empirical formula, molecular formula, structural formula, displayed formula, general formula, and functional group. *(For homologous series: alkanes, alkenes, alkynes, alcohols, carboxylic acids.)* (14.3b)
- Identify the homologous series of given simple organic molecules from their names and/or displayed formulae. *(Limited to the first 5 straight chain members of alkanes, alkenes, alkynes, alcohols, and carboxylic acids.)* (14.2c)
- Draw structures of simple organic molecules from their names and vice-versa. *(Limited to the first 5 straight chain members of alkanes, alkenes, alkynes, alcohols, and carboxylic acids where the functional group (if applicable) is on the first carbon atom.)* (14.3c)

There are three main branches in the study of chemistry:

- **Physical chemistry**: concerned with the physical aspect of chemistry, e.g. energetics, rates of reactions, etc.
- **Inorganic chemistry**: concerned with the properties and behaviour of inorganic compounds.
- **Organic chemistry**: concerned with the study of compounds of carbon except for oxides, metal carbonates, and hydrogen carbonates. Many of these compounds of carbon can be manufactured by living cells.

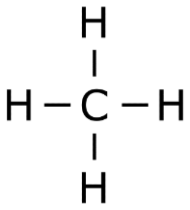
In this topic we will look at an overview of the major areas of organic chemistry. All organic compounds contain carbon. Carbon is found in group 4 of the Periodic Table. Since it has four outer electrons, it can form up to four single covalent bonds with other elements.

These four bonds have particular positions in space. They are not all in one plane; rather they tend to form a tetrahedral structure such as that found in methane (CH_4). Methane is the simplest organic molecule. However, at this level organic molecules will be displayed in a two-dimensional representation. So methane can be shown as:



3-dimensional representation of methane

or



2-dimensional representation of methane

Carbon also has a particular property in that its atoms can join up with each other to form chains or closed rings of atoms. Such compounds are known as **organic**. Organic compounds may contain other elements.

At this level we shall only be dealing with organic compounds whose structure is based on a carbon chain.

Organic compounds can be natural or synthetic.

Natural organic compounds include **keratin** which is found in nails and hair. Other organic compounds are **carbohydrates** which are a main source of energy for living things, **proteins** which are essential for building and repairing tissues in the body and **lipids** (fats, oils and waxes) that store energy and provide insulation.

Synthetic organic compounds include **plastics** such as polyethene, polyvinylchloride and polystyrene, **medicines** such as aspirin and paracetamol, **fibers** such as polyester and nylon, and **solvents** such as acetone and benzene.

5.1.1 Homologous series & functional groups

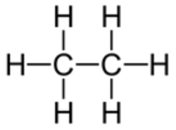
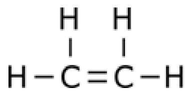
Because there are so many different organic compounds scientists had to invent a form of classification system. They called the groups they devised **homologous series**.

> A **homologous series** can be defined as a group of organic compounds, all of which have the same general formula and similar chemical properties.

In fact compounds in modern organic chemistry are all grouped in different homologous series.

Up till now, the organic molecule we have mentioned, methane contained only carbon and hydrogen atoms. There are many other compounds that are only composed of carbon and hydrogen - they are collectively known as **hydrocarbons**.

In this unit we are going to look at the three homologous series which are made of carbon and hydrogen atoms only. These are: alkanes, alkenes and alkynes.

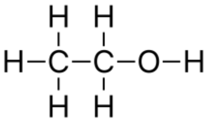
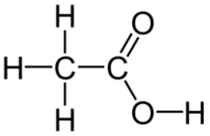
Homologous series	Contains	Example
alkane	only **C – C** single bonds	 $\mathrm{H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H}$ ethane, C_2H_6
alkene	at least one **C = C** double bond	 $\mathrm{H-\overset{H}{C}=\overset{H}{C}-H}$ ethene, C_2H_4
alkyne	at least one **C ≡ C** triple bond	$\mathrm{H-C\equiv C-H}$ ethyne, C_2H_2

But there are other organic compounds that have other kinds of atoms combined with them.

This other atom or group of atoms is called the **functional group**. It is this functional group that gives identity to the compound, i.e. it helps us group compounds according to which functional group they contain. The functional group determines the chemical properties of those compounds.

Note that hydrocarbons do not have a functional group.

Here are the homologous series that contain a functional group which we will study in this unit.

Homologous series	Functional group	Example
alcohol	**-OH** hydroxyl group	 ethanol, C_2H_5OH
alkanoic acid or **carboxylic acid**	**-COOH** carboxyl group	 ethanoic acid, CH_3COOH

All members of a homologous series have the following characteristics:

- They have the same general molecular formula, e.g. for alkanes, the general molecular formula is $\mathbf{C_nH_{2n+2}}$, where "n" is a whole number.
- They exhibit similar chemical reactions, e.g. all alkanes burn in air and give substitution reactions with chlorine, although with varying vigour.
- The physical properties change gradually as the number of carbons increases, e.g. as alkanes grow larger in molecular size, they have higher boiling points.
- They have the same general methods of preparation.

1. What is a hydrocarbon? Name an example.
2. What is a homologous series? Name a few examples.
3. What is a functional group?
4. Name the characteristic properties of members of a homologous series.

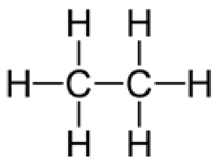
5.1.2 Different types of formulae

When describing organic molecules, different situations call for the use of different kinds of formulae to describe an organic molecule.

Empirical formula: gives the smallest ratio of particles in a substance. For example, the empirical formula for ethane whose formula is C_2H_6 is $\mathbf{CH_3}$.

Molecular formula: shows the number of atoms of different elements forming up a compound. It does not give any clue as to how these atoms are arranged, for example: $\mathbf{C_2H_6}$ (ethane).

Displayed formula: shows all the bonds connecting the elements in an organic compound as shown for ethane below:



Structural formula: shows the individual atoms in a molecule and how they are arranged in relation to each other. For example, for ethane whose molecular formula is C_2H_6, the structural formula is: **CH_3CH_3**.

5.1.3 Naming organic molecules

The modern system of naming organic molecules is briefly explained as follows:

The front part of the name usually refers to the number of carbon atoms in the longest chain in the molecule, e.g.

- meth- = 1 carbon.
- eth- = 2 carbons.
- prop- = 3 carbons.
- but- = 4 carbons.
- pent- = 5 carbons.

Names of compounds in each homologous series has a particular ending, for example the names of:

- alkanes end in -ane.
- alkenes end in -ene.
- alkynes end in -yne.
- alcohols end in -anol.
- alkanoic acids end in -anoic acid.

The table below shows how the names of the first five members of the alkanes homologous series are derived.

Number of carbon atoms	Molecular formula of straight chain alkane	Name of alkane	Prefix used in naming	Suffix used in naming
1	CH_4	methane	meth-	-ane
2	C_2H_6	ethane	eth-	-ane
3	C_3H_8	propane	prop-	-ane
4	C_4H_{10}	butane	but-	-ane
5	C_5H_{12}	pentane	pent-	-ane

A compound called ethanoic acid is a member of the organic acids homologous series and has two carbon atoms.

Name the homologous series to which the following compounds belong and state how many carbon atoms it contains:

- ethene
- propane
- butanol
- pentanoic acid

5.2 Alkanes

By the end of this section, you should be able to:

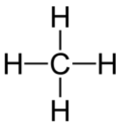
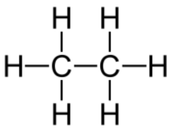
- Define alkanes. (14.2b part)
- Name common alkanes that are used as fuels. *(E.g. methane, propane, butane, etc.)* (14.2g)
- Compare the strength of intramolecular bonding (covalent) and intermolecular forces (weak forces of attraction) in alkanes and use these to explain the trends in properties of alkanes such as boiling points and melting points. (14.3h)
- Relate the production of carbon dioxide/carbon monoxide with complete/incomplete combustion of hydrocarbons. (14.2i)
- Describe the main chemical reactions of alkanes. *(Limited to cracking, combustion, and halogenation (monosubstitution).)* (14.3i)
- Define saturated and unsaturated hydrocarbons. (14.2j)
- Link the saturated nature of alkanes to their lack of reactivity. (14.3j)
- Describe how certain organic substances, other than fuels, can contribute to environmental problems. (Limited to the ongoing effect of CFCs on ozone depletion and their replacement.) (14.2m part)

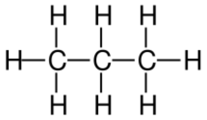
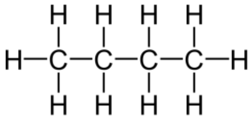
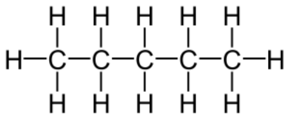
Alkanes are saturated hydrocarbons.

The word **saturated** refers to the fact that their carbon atoms are joined by **single** covalent bonds (C - C) only.

They do not contain C = C or C ≡ C bonds.

The following table shows the first five members of the alkanes homologous series.

Name	Molecular formula	Structural formula	Displayed formula	Boiling Point /°C	State at RTP
Methane	CH_4	CH_4		-161	Gas
Ethane	C_2H_6	CH_3CH_3		-89	Gas

Name	Molecular formula	Structural formula	Displayed formula	Boiling Point /°C	State at RTP
Propane	C_3H_8	$CH_3CH_2CH_3$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	-42	Gas
Butane *	C_4H_{10}	$CH_3CH_2CH_2CH_3$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	0	Gas
Pentane *	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	36	Liquid

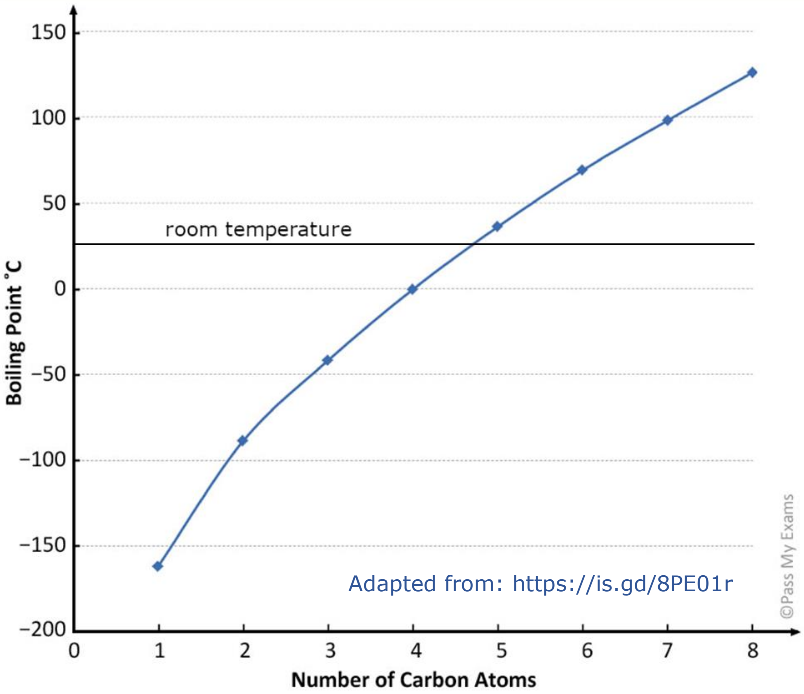
* Other isomers exist which are not examinable.

The general formula for alkanes is **C_nH_{2n+2}**. (Where **n** is the number of carbon atoms present in the compound)

Alkanes are insoluble in water but soluble in organic solvents.

The physical properties of alkanes change gradually as the molecular formula becomes larger. The first four alkanes are gases at room temperature, these are followed by liquids, such as pentane and hexane, and finally we find the solid alkanes (waxes).

Variation of Boiling Point of Straight Chained Alkanes with Number of Carbon Atoms



Adapted from: https://is.gd/8PE01r

The boiling points of alkanes increase with increasing number of carbons atoms as shown in the graph. This is because the **intermolecular forces** among other factors, although weak, become more significant as the number of atoms and electrons in the molecule increases.

The alkanes are not very reactive compounds. The only important alkane reactions are **combustion** and **substitution**.

1. Give the name, empirical formula, display formula, and structural formula of C_4H_{10}.
2. Give the empirical formula, molecular formula, display formula, and structural formula of pentane.

Alkanes

Molecules of alkanes can also be found in outer space. Scientists discovered small amounts of methane and ethane, in the atmospheres of Jupiter, Saturn, Uranus and Neptune.

On one of the moons of Saturn, Titan, there is occasional rain of liquid methane and on this same moon, seasonal lakes of methane and ethane form next to the polar regions.

5.2.1 Combustion

Alkanes such as methane and ethane (liquefied natural gas, LNG), propane and butane (liquefied petroleum gas, LPG) are commonly used as fuels. They burn in a **plentiful supply of oxygen** to produce carbon dioxide and water vapour only.

For example, methane burns completely in oxygen to form carbon dioxide and water vapour.

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

The above equation is an example of **complete combustion**. Alkanes burn with a clean blue flame during complete combustion.

When alkanes burn in a **limited supply of oxygen**, carbon monoxide gas is produced which is colourless and odourless. This gas is poisonous since it binds permanently with red blood cells rendering them useless for the respiration process. The equation below shows methane burning in a limited supply of oxygen.

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$$

The above equation is an example of **incomplete combustion**.

When burning larger organic molecules, a yellow sooty flame is noted. This happens since the amount of oxygen in the air is not enough to react completely with large hydrocarbons. The yellow flame is an indication that carbon is being produced by the reaction. The carbon produced glows and gives a yellow colour in the flame. The yellow flame is an indication that incomplete combustion is happening.

Carbon produced this way is called soot which causes environmental and health problems. This was mentioned in Year 9 in the section dealing with particulate matter.

5.2.2 Substitution reactions

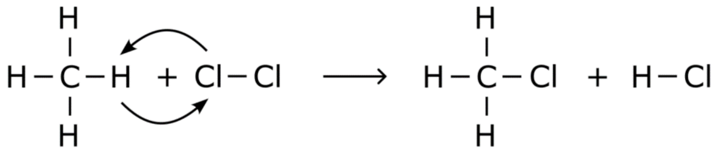
When chlorine gas reacts with an alkane one of the chlorine atoms substitutes/replaces one of the hydrogen atoms in the alkane. At the same time the substituted/replaced hydrogen atoms combines with the chlorine atom to form hydrogen chloride gas.

For example methane can react with chlorine gas to form chloromethane and hydrogen chloride gas as shown below:

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$$

In the dark, this mixture does not react. The above reaction is photo catalysed. This means that UV light can ignite a mixture of methane and chlorine to react together.

So methane undergoes a substitution reaction as shown below:



When only one hydrogen atom is allowed to be replaced, chemists refer to it as a **monosubstitution** reaction. In reality, this substitution reaction continues until all hydrogen atoms in methane are replaced one by one.

Chloromethane is an example of a halogenoalkane or haloalkane. Other examples include chlorofluorocarbons (CFCs). They were widely used as refrigerants, propellants, and solvents. However, these compounds destroy the ozone layer, which protects us from the harmful ultraviolet rays produced by the sun and so were subsequently banned from production. Safer alternatives have now been developed.

Hydrocarbon Power: Crash Course Chemistry #40

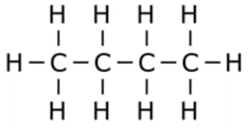
5.3 Isomerism

- Define isomerism. (14.2d)
- Identify isomers from displayed formulae of alkanes. *(Limited to alkanes with 4 and 5 carbon atoms. No naming of branched hydrocarbons is required.)* (14.2e)
- Draw isomers of alkanes from their molecular formulae. *(Limited to alkanes with 4 and 5 carbon atoms. No naming of branched hydrocarbons is required.)* (14.3e)

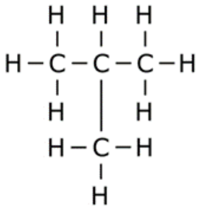
Molecules of C_4H_{10} can have different arrangement of atoms in space i.e. different structural and displayed formulae, though they both have the same molecular formula as shown below:

Such compounds are known as **isomers**.

Butane

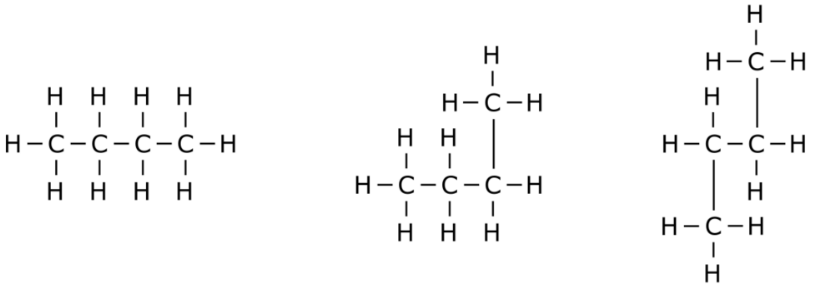


Methylpropane



Isomerism is the occurrence of two or more compounds with the same molecular formula but different structural formulae.

This excludes any different arrangements which are simply due to the molecule rotating as a whole or rotating about particular bonds. For example, the diagram below shows molecules of butane. They are **not** isomers of each other since the longest chain contains 4 carbon atoms.

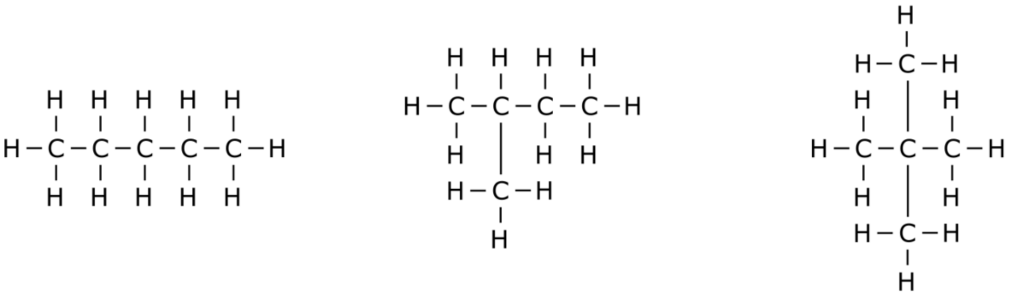


Isomers arise because of branching; in fact all hydrocarbons with four or more carbon atoms give rise to isomers. Isomers of the same molecular formula have **different** physical and, in some cases, different chemical properties because they have different structures.

There are various kinds of isomerism. The kind we will be concerned with at this level is called **structural isomerism**. In structural isomers the atoms are arranged in a different order.

Furthermore, there are various types of structural isomerism. In this unit we are going to see examples of **chain isomerism** which is concerned with the arrangement of carbon atoms in the molecule and arises due to the possibility of branches in carbon chains.

Additionally C_5H_{12} has three isomers which are shown in the diagram below.



Pentane

Methylbutane

Dimethylpropane

***** The names of the branched isomers mentioned in the above diagram are not examinable.

5.4 Cracking

By the end of this section, you should be able to:

- Describe how long chain alkanes can be converted to smaller, more useful ones. *(Limited to thermal cracking only. Specific cracking temperatures are not required.)* (14.2f)
- Identify possible alkanes and alkenes that can be obtained from thermal cracking of long chain alkanes. (14.3f)

Not all the compounds found in crude oil are useful. Hydrocarbons having larger molecules are thick liquids or solids with high boiling points. They are not good as fuels, since they are difficult to vaporise and do not burn easily.

However, we can break down large hydrocarbon molecules. This process is called **cracking**.

Cracking is a chemical process in which
less useful, long-chain hydrocarbons, are broken into
more useful, short-chain hydrocarbons.

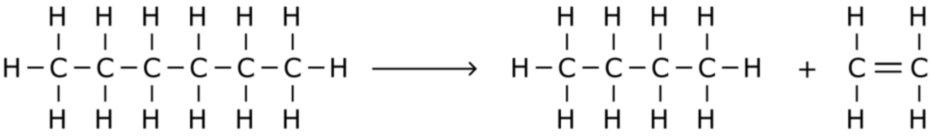
The long chain hydrocarbon is heated to a high temperature. The hydrocarbons are cracked as thermal decomposition reactions take place. The large molecules split apart to form smaller, more useful ones.

The products of cracking include alkanes and alkenes.

Example 1: hexane can be cracked to form butane and ethene:

hexane → butane + ethene

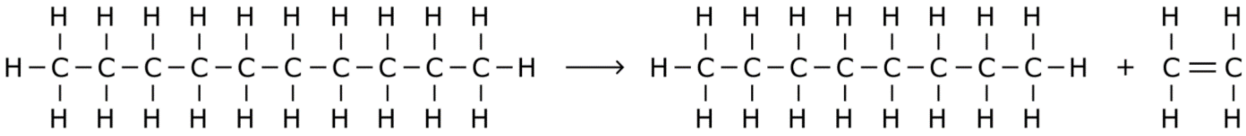
$$C_6H_{14} \rightarrow C_4H_{10} + C_2H_4$$



Example 2: decane can be cracked to form octane and ethene:

decane → octane + ethene

$$C_{10}H_{22}(g) \rightarrow C_8H_{18}(g) + C_2H_4(g)$$



Hydrocarbon cracking and why it is done | FuseSchool

A lesson on cracking recorded by David Bugeja, chemistry teacher at Verdala Secondary School.

Further reading on cracking.

Catalytic Cracking

The first commercial catalytic cracking of heavy petroleum oils was set in Pittsburgh U.S. in 1915, developed by Almer McAfee, for Gulf Refining Company. At that time the process was discontinued due to the high cost of the required catalyst. Nevertheless, over the next decades, other companies kept on improving the process and identifying catalysts which made the process more economically viable.

Later this process was efficiently used to supply the high demand of fuel needed by the air forces of Great Britain and the United States, in winning World War II.

5.5 Alkenes and their addition reactions

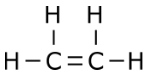
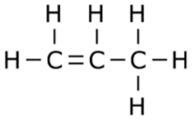
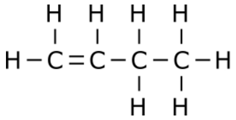
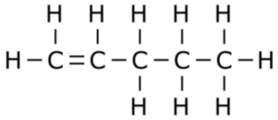
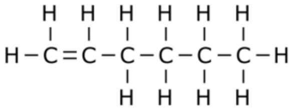
By the end of this section, you should be able to:

- Define alkenes. (14.2b part)
- Describe a test to distinguish between saturated and unsaturated hydrocarbons. (14.2k)
- Describe addition reactions of ethene. *(E.g. bromination, hydration, and hydrogenation. Details of reaction conditions are not required.)* (14.3k)
- Link the reactivity of alkenes to unsaturation. (14.3l part)

Alkenes are hydrocarbons with a **double** covalent bond between two of their carbon atoms and have the general formula, **C_nH_{2n}**. Since a double bond is present, alkenes are called **unsaturated** hydrocarbons.

Alkenes are unsaturated hydrocarbons which have at least one **double covalent bond** between two of their carbon atoms.

The following table shows alkenes up to 5 carbon atoms. Note that an alkene with one carbon atom does not exist since a double covalent bond must be present between **two** carbon atoms.

Name	Molecular formula	Structural formula	Displayed formula	Boiling Point /°C	State at RTP
Ethene	C_2H_4	CH_2CH_2	 $H-\overset{H}{C}=\overset{H}{C}-H$	-104	Gas
Propene	C_3H_6	CH_2CHCH_3	 $H-\overset{H}{C}=\overset{H}{C}-\underset{H}{\overset{H}{C}}-H$	-47	Gas
Butene *	C_4H_8	$CH_2CHCH_2CH_3$	 $H-\overset{H}{C}=\overset{H}{C}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	-6	Gas
Pentene *	C_5H_{10}	$CH_2CHCH_2CH_2CH_3$	 $H-\overset{H}{C}=\overset{H}{C}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	30	Liquid
Hexene *	C_6H_{12}	$CH_2CHCH_2CH_2CH_2CH_3$	 $H-\overset{H}{C}=\overset{H}{C}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	69	Liquid

* Other isomers exist which are not examinable.

Apart from combustion reactions alkenes undergo addition reactions.

An **addition reaction** is one where two reactants react to produce a single product.

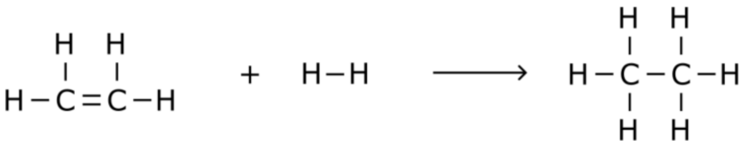
As a result, the double C=C covalent bond in alkenes is converted to a single C-C covalent bond. In the process this creates space for other atoms to combine with the alkene to form a saturated molecule as shown in the following sections.

Alkenes are more reactive than alkanes because the unsaturated bond is easy to break.

5.5.1 Hydrogenation

The addition of hydrogen gas to an alkene is called **hydrogenation**. The equation below shows the reaction of ethene with hydrogen.

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

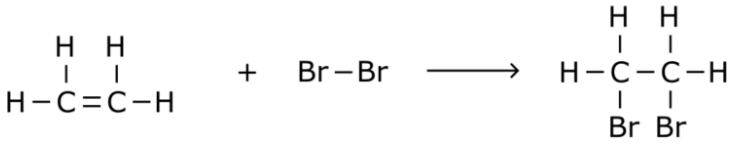


This is the reaction involved in the conversion of edible oils to margarine. Conversion of double bonds to single bonds results in margarine being a harder material than the oil it started from.

5.5.2 Bromination

Bromination is a form of halogenation whereby halogens combine with alkenes to form haloalkanes. These reactions take place at room temperature. The equation below shows the reaction of ethene with bromine to form dibromoethane. This is another example of a haloalkane.

$$C_2H_4(g) + Br_2(g) \rightarrow C_2H_4Br_2(g)$$

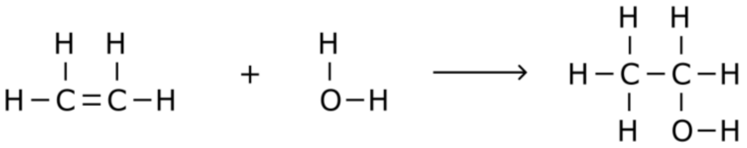


Such a reaction always produces a molecule where the bromine atoms are combined to adjacent (next to each other) carbon atoms. This reaction is used to test for the presence of unsaturated hydrocarbons as described in section 5.7.

5.5.3 Hydration

Alkenes react with water vapour to form alcohols. Ethene reacts with water vapour at high pressure to form ethanol. Synthetically made ethanol is made using this chemical pathway according to the following equation:

$$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$$



Alkene reactions

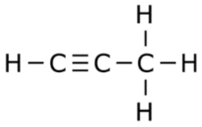
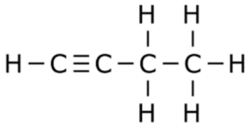
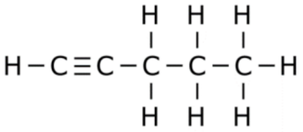
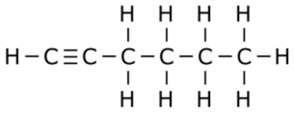
5.6 Alkynes

- Define alkynes. (14.2b part)
- Link the reactivity of alkynes to unsaturation. (14.3l part)

Alkynes, like alkenes, are unsaturated hydrocarbons. They have a **triple covalent bond** instead of a double bond. Their general formula is $\mathbf{C_nH_{2n-2}}$. Because of their unsaturation they undergo similar addition reactions like alkenes but to a greater extent since they can add more molecules than alkenes.

Alkynes are unsaturated hydrocarbons with have at least one **triple covalent bond** between two of its carbon atoms.

The following table shows alkynes having up to five carbon atoms. Like alkenes, an alkyne with one carbon atom does not exist for the same reason.

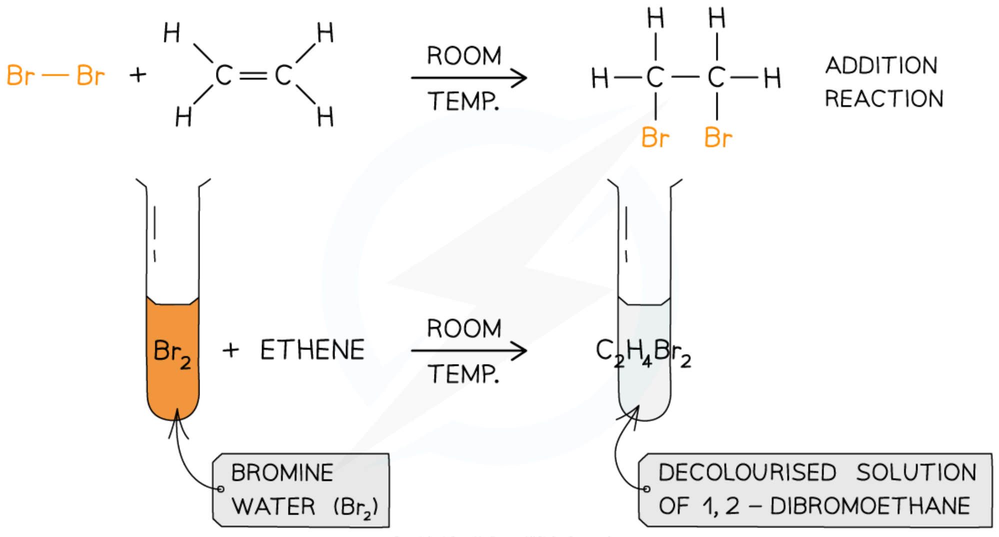
Name	Molecular formula	Structural formula	Displayed formula	Boiling Point /°C	State at RTP
Ethyne	C_2H_2	CHCH	$H-C\equiv C-H$	-84	Gas
Propyne	C_3H_4	$CHCCH_3$	 $H-C\equiv C-\underset{H}{\overset{H}{C}}-H$	-23	Gas
Butyne *	C_4H_6	$CHCCH_2CH_3$	 $H-C\equiv C-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	-1	Gas
Pentyne *	C_5H_8	$CHCCH_2CH_2CH_3$	 $H-C\equiv C-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	31	Liquid
Hexyne *	C_6H_{10}	$CHCCH_2CH_2CH_2CH_3$	 $H-C\equiv C-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	71	Liquid

***** Other isomers exist which are not examinable.

5.7 Distinguishing between saturated and unsaturated hydrocarbons

Bromine water test

The orange-red colour of bromine water is rapidly decolourised when an unsaturated hydrocarbon such as ethene or ethyne is mixed with it at room temperature. This is the test for the presence of unsaturation.



The alkanes will not rapidly decolourise bromine water since a higher temperature / UV light would be required for a substitution reaction to happen.

Watch how the bromine water test works.

Burning of hydrocarbons

If two hydrocarbons, such as ethane and ethene are burned in air, ethane burns with a blue flame while ethene burns with a sooty flame.

Alkenes are more reactive than alkanes, as a result of the presence of a carbon double bond. When an alkene burns, it therefore needs a rapid supply of oxygen. However, if the oxygen is not supplied fast enough, the alkene will burn with a limited oxygen supply and incomplete combustion will result. One product of the incomplete combustion of hydrocarbons is carbon monoxide, and the other is soot (carbon). Soot is responsible for producing a smoky flame.

This fact can be used to distinguish between alkanes and alkenes that have the same number of carbon atoms for the smaller molecules.

Note that the better way to test for the presence of saturated and unsaturated hydrocarbons is by using the bromine water test.

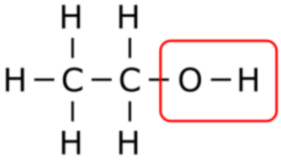
5.8 Alcohols

By the end of this section, you should be able to:

- Define alcohols and carboxylic acids. (14.2b part)
- Describe some important uses of ethanol. *(E.g. Solvent, fuel and alcoholic drinks.)* (14.1n)
- Describe how ethanol can be produced through fermentation and hydration of ethene. (14.2n)
- Evaluate the advantages and disadvantages of fermentation and hydration of ethene. (14.3n)
- Describe how ethanol can be oxidised to ethanoic acid using acidified potassium dichromate and by aerial oxidation. *(Chemical equations are not required.)* (14.3o)
- Construct the reaction between a carboxylic acid and an alcohol to form an ester. *(Limited to ethyl ethanoate.)* (14.3q)
- Identify the ester functional group in a displayed formula. (14.3r)

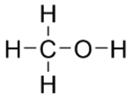
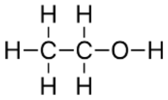
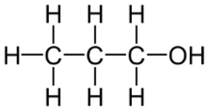
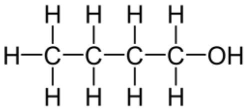
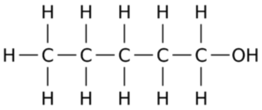
Alcohols have a general formula of **$C_nH_{2n+1}OH$**. This homologous group is also known as the **alkanols**.

The **-OH** or **hydroxyl group** is the functional group of the alcohols. Alcohols are volatile liquids at room temperature and pressure. They are colourless and have a characteristic smell. They mix well with water. Liquids that mix well together are referred to as **miscible**.



Methanol is very dangerous as it attacks the optic nerve causing blindness. Ethanol is used in alcoholic drinks (wine, beers, and spirits) and also as a solvent. Ethanol is also used as a fuel on its own or mixed with other hydrocarbons.

The table below shows the first five members of the alcohols homologous series.

Name	Molecular formula	Structural formula	Displayed formula	Boiling point /°C	State at RTP
Methanol	CH_4O	CH_3OH		65	Liquid
Ethanol	C_2H_6O	CH_3CH_2OH or C_2H_5OH		78	Liquid
Propanol *	C_3H_8O	$CH_3CH_2CH_2OH$ or C_3H_7OH		97	Liquid
Butanol *	$C_4H_{10}O$	$CH_3CH_2CH_2CH_2OH$ or C_4H_9OH		117	Liquid
Pentanol *	$C_5H_{12}O$	$CH_3CH_2CH_2CH_2CH_2OH$ or $C_5H_{11}OH$		138	Liquid

* Other isomers exist, which are not examinable.

5.8.1 Producing ethanol by fermentation

Ethanol is prepared by the **fermentation** of sugars. Enzymes produced by the yeast cells, breaks down the glucose **anaerobically** (without oxygen) to produce ethanol, carbon dioxide and energy.

In the lab, yeast is added to a sugar solution. The mixture is kept warm at about 30 °C. Carbon dioxide is produced which turns lime water milky. The solution is left to ferment for some days. The following reaction shows the fermentation of glucose:

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

Fermentation produces a dilute solution of ethanol in water. This is then distilled to produce ethanol of very high purity. This process can be used to make alcoholic drinks. In some countries that have an abundance of sugar, ethanol produced this way is used as a fuel.

Further reading on the preparation of ethanol.

Preparation of Ethanol

The preparation of alcohol, also known as fermentation, is a process that has existed since Neolithic times (c. 10000-4500 BC). In those times the product was used for beverages, medicinal purposes and for religious rituals. The first understanding of the involvement of yeast, as a living organism, in this process was first published in 1837 by various scientists. However, several chemists rejected their theory and viewed fermentation as another chemical reaction, without any contribution of living organisms.

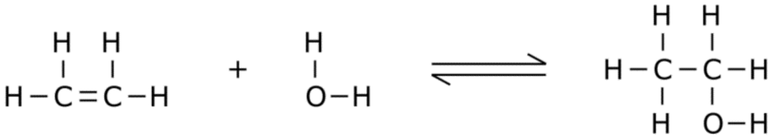
The issue was settled later in the 1850s by Louis Pasteur, a French chemist and microbiologist, who repeated earlier experiments to show that yeast could be killed if the solution was boiled, and other ones that showed that fermentation was initiated by yeast.

5.8.2 Producing ethanol by hydration of ethene

Pure ethanol can be produced by the hydration of ethene. This is an addition reaction whereby the double bond is converted to a single as water is added to ethene to produce ethanol. This reaction is done on an industrial scale as it has many uses in industry such as a solvent, as a disinfectant among many other uses.

$$\text{ethene} + \text{water vapour} \rightleftharpoons \text{ethanol}$$

$$C_2H_4 + H_2O \rightleftharpoons C_2H_5OH$$



Catalytic hydration of ethene

For the above reaction where a mixture of ethene and water vapour produces ethanol, the actual industrial conditions require that ethene and water vapour are passed over phosphoric acid as catalyst. For this reaction to take place, a temperature of 300 °C and a pressure of 60 atmospheres are required.

Making alcohols by fermentation & from ethene. | FuseSchool

Further reading on the industrial manufacture of alcohol.

5.8.3 Advantages and disadvantages of fermentation and hydration of ethene to produce ethanol

The following table shows the advantages and disadvantages of two common processes that are used to produce ethanol.

	Advantages	Disadvantages
Fermentation	**Renewable Source:** Ethanol in this process is made using sugar found in plants. This makes it a renewable source as it can be grown again. **Carbon Neutral:** The process could be seen as carbon neutral because as the plant grows, it takes in carbon dioxide from the air for photosynthesis. This therefore balances out the carbon dioxide released when it is fermented. **Economic Benefit:** In places such as Brazil, sugar cane is used in this process as it is a starchy material. In areas such as this, it is beneficial to the economy as it means that there is more employment.	**Two-Stage Process:** It involves fermentation and then distillation. **Batch Process:** It is a batch process. When the ethanol concentration reaches about 15% the yeast dies, and fermentation stops. **Impure Product:** The product obtained is not pure. The mixture must be concentrated and purified by distillation, which adds extra cost. **Ethical Questions:** It raises ethical questions as to whether we should use crops for ethanol when there are third world countries with no food.
Hydration of ethene	**Continuous process:** The process is continuous, which means that more can be made for a cheaper price in a shorter time. **No waste products:** There are no waste products in this process. **Doesn't use food crops:** It doesn't use crops that could be used as food. **Purer product:** The product made is purer than that made by fermentation.	**Non-renewable resource:** Crude oil is used, which is a non-renewable resource and will one day run out. **Future cost increase:** Although this method of obtaining ethanol is currently fairly cheap, as resources run out it will become more expensive.

In conclusion, both processes have their own advantages and disadvantages. The choice between them depends on various factors such as the scale of production, availability of resources, and the desired purity of the product. While fermentation is a natural and renewable process, it is slower and has a risk of contamination. On the other hand, the hydration of ethene is faster and produces a purer product, but it relies on non-renewable resources and requires a high amount of energy.

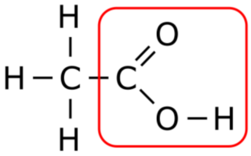
5.9 Carboxylic acids

By the end of this section, you should be able to:

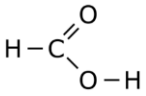
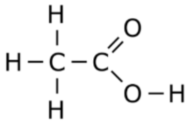
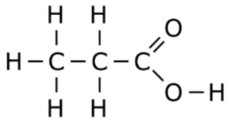
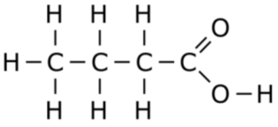
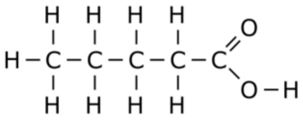
- Define carboxylic acids. (14.2b part)
- Describe how ethanol can be oxidised to ethanoic acid using acidified potassium dichromate and by aerial oxidation. *(Chemical equations are not required.)* (14.3o)

The general formula of alkanoic acids is **$C_nH_{2n+1}COOH$** where in this case **n** is the number of carbon atoms in the molecule minus 1. So in this case the general formula can be taken as **$C_{n-1}H_{2n+1}COOH$**. This is because one of the carbon atoms is already represented in the -COOH functional group. This homologous group is also known as the **carboxylic acids** group.

The functional group in carboxylic acids, is known as the carboxyl group and is determined by the **-COOH** group of atoms that is situated at the end of an organic compound. The red box in the displayed formula for ethanoic acid indicates the functional group.



The following table shows the first 5 members of the carboxylic acid group.

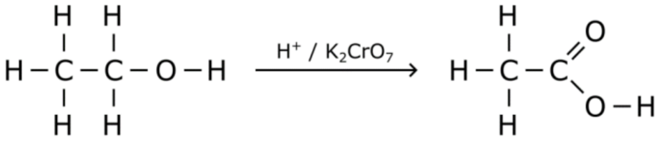
Name	Molecular formula	Structural formula	Displayed formula	Boiling point /°C	State at RTP
Methanoic acid	CH_2O_2	HCOOH		101	Liquid
Ethanoic acid	$C_2H_4O_2$	CH_3COOH		118	Liquid
Propanoic acid	$C_3H_6O_2$	CH_3CH_2COOH or C_2H_5COOH		141	Liquid
Butanoic acid *	$C_4H_8O_2$	$CH_3CH_2CH_2COOH$ or C_3H_7COOH		164	Liquid
Pentanoic acid *	$C_5H_{10}O_2$	$CH_3CH_2CH_2CH_2COOH$ or C_4H_9COOH		187	Liquid

* Other isomers exist which are not examinable.

Oxidation of ethanol to ethanoic acid

Bacteria that occurs naturally around us, turns ethanol in wine when exposed to air into ethanoic acid (vinegar) and so wine goes sour. This is known as aerial oxidation.

Ethanol may be oxidised in the lab by using a suitable oxidising agent, such as potassium dichromate. A mixture of acidified potassium dichromate solution together with some ethanol is warmed in a test tube. The mixture turns from **orange** to **green** as the dichromate ion ($Cr_2O_7^{2-}$) changes into the chromium (III) ion (Cr^{3+}). Ethanoic acid, CH_3COOH is produced.



Note that the above equation is not balanced and is not examinable.

GCSE Chemistry - Carboxylic Acids #58

Further reading on the oxidation of alcohols.

5.10 Esters

By the end of this section, you should be able to:

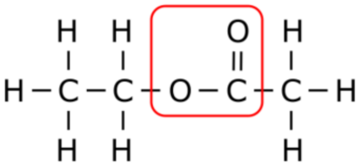
- Construct the reaction between a carboxylic acid and an alcohol to form an ester. *(Limited to ethyl ethanoate.)* (14.3q)
- Identify the ester functional group in a displayed formula. (14.3r)

Esters form when carboxylic acids react with alcohols. For example, when ethanol reacts with ethanoic acid, the ester, **ethyl ethanoate**, is produced as shown below.

$$C_2H_5OH + CH_3COOH \rightleftharpoons CH_3COOC_2H_5 + H_2O$$

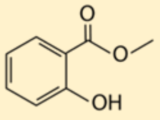
The reaction is carried out in the presence of concentrated sulfuric acid. Besides acting as a catalyst, the concentrated acid helps to absorb the water produced. Taking away the water causes the reaction to move from left to right which produces more ester.

The functional group in esters can be identified when the organic molecule has the **-COO-** sequence in its carbon chain. This also known as the ester link.

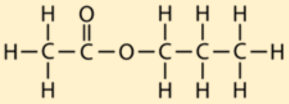


Build virtual organic molecules.

Carboxylic acids react with alcohols to form esters



methyl salicylate



$\mathrm{H{-}\overset{H}{\underset{H}{C}}{-}\overset{O}{\overset{\|}{C}}{-}O{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}H}$

propyl ethanoate

Esters are organic compounds that often give a pleasant fruity smell, and some can be found naturally in the essential oils of plants. They are commonly used in fragrances, food flavouring and synthetic lubricants. Some of the most distinct odours of esters include those of pineapple, pear, strawberry, banana, apple, nail polish remover, lemon, and cherry.

5.11 Polymers

By the end of this section, you should be able to:

- List uses of polyethene, PTFE and PVC. (14.1p)
- Discuss how applying a strategy of "reduce, reuse, recycle" can alleviate environmental problems caused by organic substances. (14.2p)
- Model the production of polymers from alkenes and other unsaturated monomers by addition polymerisation. *(Limited to polyethene, PTFE and PVC.)* (14.3p)
- Describe how certain organic substances, other than fuels, can contribute to environmental problems. *(Limited to non-biodegradable plastics.)* (14.2m part)

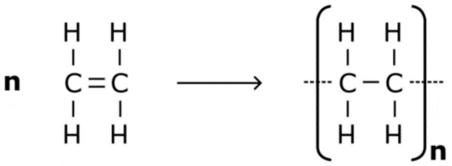
Polymers are **very large molecules** (sometimes called **macromolecules**). Poly means 'many'. Each polymer molecule is a long chain of carbon atoms.

Polymers are made from many **smaller molecules**, called **monomers**. There are two kinds of polymerisation, addition and condensation polymerisation. At this level, we shall be dealing with addition polymerisation only where the monomers (the starting materials) are unsaturated hydrocarbons such as alkenes and alkynes.

The following sections deal with only three of the many polymers that exist.

5.11.1 Polyethene

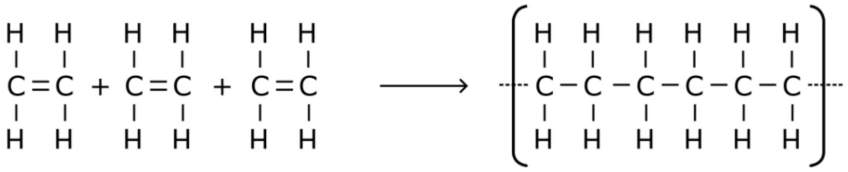
When ethene is heated at high pressure, polyethene is formed. The C=C double bond in ethene is involved in the polymerisation reaction. It allows ethene molecules to join together to form a single product, so it is an example of an **addition reaction**. Many ethene monomer molecules combine together to form polyethene which is an addition polymer.



$$\mathbf{n}\ \mathrm{\overset{H}{\underset{H}{C}}{=}\overset{H}{\underset{H}{C}}} \longrightarrow \left[\mathrm{\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}}\right]_{\mathbf{n}}$$

'**n**' can be thousands of ethene molecules.

The following diagram shows how three ethene monomer units combine together to form polyethene.

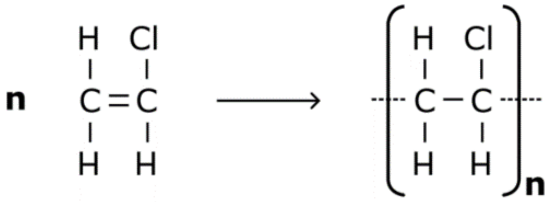


Uses of polyethene

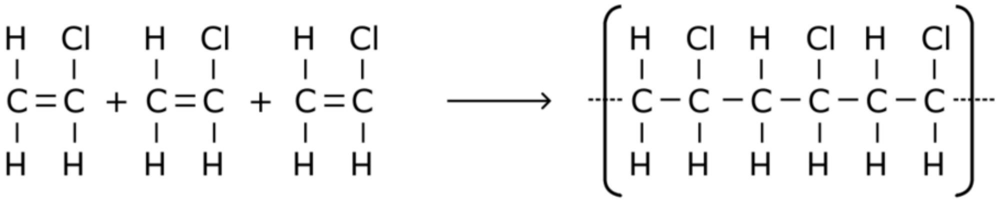
- Bags / dustbin liners
- Hoses / tubes
- Insulation
- Plastic bottles
- Food wrapping

5.11.2 Polychloroethene

When chloroethene is heated at high pressure, polychloroethene is formed. The C=C double bond in chloroethene is involved in the polymerisation reaction. It allows chloroethene molecules to join together to form a single product, so it is an example of an addition reaction. Many chloroethene monomer molecules form polychloroethene which is an addition polymer. Polychloroethene is also known as polyvinylchloride (PVC).



The following diagram shows how three chloroethene monomer units combine together to form polychloroethene.



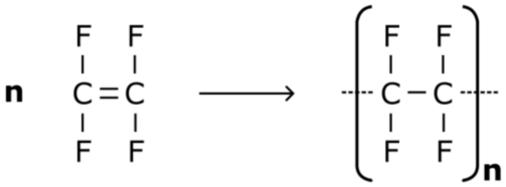
Uses of polychloroethene:

- in the building and construction industry,
- health care, electronics, automobile, and other sectors,
- in products ranging from piping to wire and cable insulation and
- windshield system components, among others.

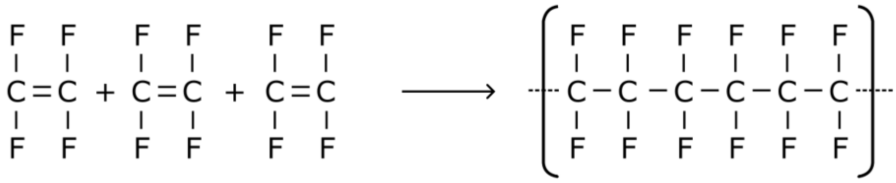
5.11.3 Polytetrafluoroethene (PTFE)

The monomer used in the production of PTFE is tetrafluoroethene. When these monomers react together, they form polytetrafluoroethene as shown below.

Many tetrafluoroethene monomer molecules form polytetrafluoroethene which is an addition polymer.



The following diagram shows how three tetrafluoroethene monomer units combine together to form Polytetrafluoroethene (PTFE).



The commercial name of PTFE is Teflon. PTFE is hydrophobic, non-reactive, high strength and flexible material. One of the uses of Teflon is in non-stick pan coatings.

5.11.4 Problems with polymers

One of the useful properties of polymers is that they are unreactive. This means they are suitable for storing food and other substances safely. Unfortunately, this property makes it difficult to dispose of polymers since they are non-biodegradable. Hence the strategy of "reduce, reuse and recycle" must be employed to use less of the Earth's resources.

Applying the strategy of reduce, reuse, and recycle can help alleviate environmental problems caused by synthetic organic substances in several ways:

1. Reduce

Decreased Production: By reducing the demand for synthetic organic substances, the overall production can be decreased, leading to reduced environmental impacts associated with their manufacturing processes. This includes energy consumption, emissions of greenhouse gases and other pollutants, and the generation of hazardous waste.

Minimized Pollution: Reducing the use of synthetic organic substances can help minimize pollution of air, water, and soil caused by their production, use, and disposal. Synthetic chemicals often persist in the environment and can have detrimental effects on ecosystems, wildlife, and human health.

2. Reuse

Extended Product Lifespan: Reusing products containing synthetic organic substances prolongs their useful life and reduces the need for new production. This reduces resource consumption, energy use, and waste generation associated with manufacturing new products.

Repurposing: Synthetic organic substances can often be repurposed for different applications before being discarded. For example, electronics and appliances containing synthetic materials can be refurbished and sold as used goods, extending their lifespan and delaying their entry into the waste stream.

3. Recycle

Closed-Loop Systems: Recycling synthetic organic substances diverts them from landfills and reintroduces them into the production cycle. This conserves resources, reduces energy consumption, and minimizes the environmental impacts associated with extracting and processing natural resources.

Material Reclamation: Many synthetic organic substances, such as plastics, can be recycled into new products. Recycling plastic, for example, reduces the demand for new plastic production, conserves fossil fuel resources, and mitigates the pollution caused by plastic waste in the environment.

Applying the reduce, reuse, recycle strategy to synthetic organic substances is crucial for mitigating environmental problems associated with their production, use, and disposal. It helps conserve natural resources, minimize pollution and waste generation, and promote a circular economy where materials are used more efficiently and continuously circulated within the system. However, effective implementation requires collaboration among governments, industries, communities, and individuals to adopt sustainable practices, invest in recycling infrastructure, and promote awareness about the environmental benefits of reducing, reusing, and recycling synthetic organic substances.

Further reading on polymers.

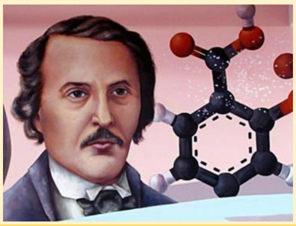
Polymers: Crash Course Chemistry #45

Further reading on homologous series.

Homologous series

The concept of homologous series can also be seen in other fields of inorganic chemistry (like phosphoric acids).

However, this classification of organic compounds was proposed initially by a French chemist, Charles Gerhardt, in 1843.



6 Periodic Table of the Elements

PERIODIC TABLE OF THE ELEMENTS

1	2											3	4	5	6	7	0
							1 **H** Hydrogen 1										4 **He** Helium 2
7 **Li** Lithium 3	9 **Be** Beryllium 4											11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10
23 **Na** Sodium 11	24 **Mg** Magnesium 12											27 **Al** Aluminium 13	28 **Si** Silicon 14	31 **P** Phosphorus 15	32 **S** Sulfur 16	35.5 **Cl** Chlorine 17	40 **Ar** Argon 18
39 **K** Potassium 19	40 **Ca** Calcium 20	45 **Sc** Scandium 21	48 **Ti** Titanium 22	51 **V** Vanadium 23	52 **Cr** Chromium 24	55 **Mn** Manganese 25	56 **Fe** Iron 26	59 **Co** Cobalt 27	59 **Ni** Nickel 28	63.5 **Cu** Copper 29	65 **Zn** Zinc 30	70 **Ga** Gallium 31	73 **Ge** Germanium 32	75 **As** Arsenic 33	79 **Se** Selenium 34	80 **Br** Bromine 35	84 **Kr** Krypton 36
85 **Rb** Rubidium 37	88 **Sr** Strontium 38	89 **Y** Yttrium 39	91 **Zr** Zirconium 40	93 **Nb** Niobium 41	96 **Mo** Molybdenum 42	99 **Tc** Technetium 43	101 **Ru** Ruthenium 44	103 **Rh** Rhodium 45	106 **Pd** Palladium 46	108 **Ag** Silver 47	112 **Cd** Cadmium 48	115 **In** Indium 49	119 **Sn** Tin 50	122 **Sb** Antimony 51	128 **Te** Tellurium 52	127 **I** Iodine 53	131 **Xe** Xenon 54
133 **Cs** Caesium 55	137 **Ba** Barium 56	139 **La** Lanthanum 57	178 **Hf** Hafnium 72	181 **Ta** Tantalum 73	184 **W** Tungsten 74	186 **Re** Rhenium 75	190 **Os** Osmium 76	192 **Ir** Iridium 77	195 **Pt** Platinum 78	197 **Au** Gold 79	201 **Hg** Mercury 80	204 **Tl** Thallium 81	207 **Pb** Lead 82	209 **Bi** Bismuth 83	210 **Po** Polonium 84	210 **At** Astatine 85	222 **Rn** Radon 86

Key:

a	*relative atomic mass*
X	***SYMBOL***
Y	***Name***
b	*atomic number*