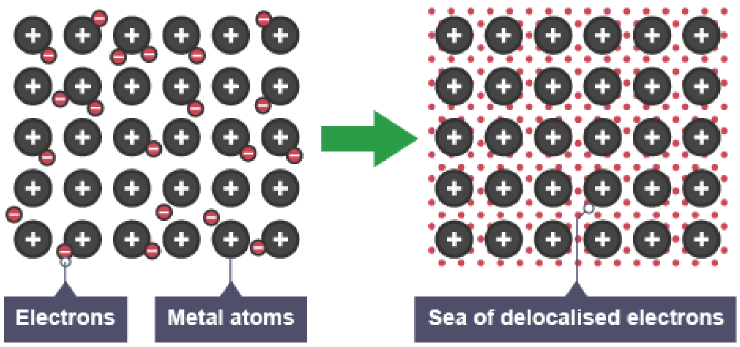
National 5 Chemistry

Unit 3 Chemistry in Society

Summary Notes

Topic 1

Metals



Learning Outcomes

Metallic bonding

- Metallic bonding is the electrostatic force of attraction between positively charged ions and delocalised electrons.
- Metallic elements are conductors of electricity because they contain delocalised electrons.

Reactions of metals

- Equations, involving formulae, can be written to show the reaction of metals with oxygen, water, and dilute acids:
    - metal + oxygen $\longrightarrow$ metal oxide
    - metal + water $\longrightarrow$ metal hydroxide + hydrogen
    - metal + dilute acid $\longrightarrow$ salt + hydrogen
- Metals can be arranged in order of reactivity by comparing the rates at which they react.
- Metals can be used to produce soluble salts - Excess metal is added to the appropriate acid, the mixture is filtered and the filtrate evaporated to dryness.

Redox

- Reduction is a gain of electrons by a reactant in any reaction.
- Oxidation is a loss of electrons by a reactant in any reaction.
- In a redox reaction, reduction and oxidation take place at the same time.
- Ion-electron equations can be written for reduction and oxidation reactions.
- Ion-electron equations can be combined to produce redox equations.

Extraction of metals

- During the extraction of metals, metal ions are reduced forming metal atoms.
- The method used to extract a metal from its ore depends on the position of the metal in the reactivity series.
- Equations can be written to show the extraction of metals.
- Methods used are:
    - heat alone (for extraction of Ag, Au and Hg)
    - heating with carbon or carbon monoxide (for extraction of Cu, Pb, Sn, Fe and Zn)
    - electrolysis (for extraction of more reactive metals including aluminium)
- Electrolysis is the decomposition of an ionic compound into its elements using electricity.
- A d.c. supply must be used if the products of electrolysis are to be identified.
- Positive ions gain electrons at the negative electrode and negative ions lose electrons at the positive electrode.

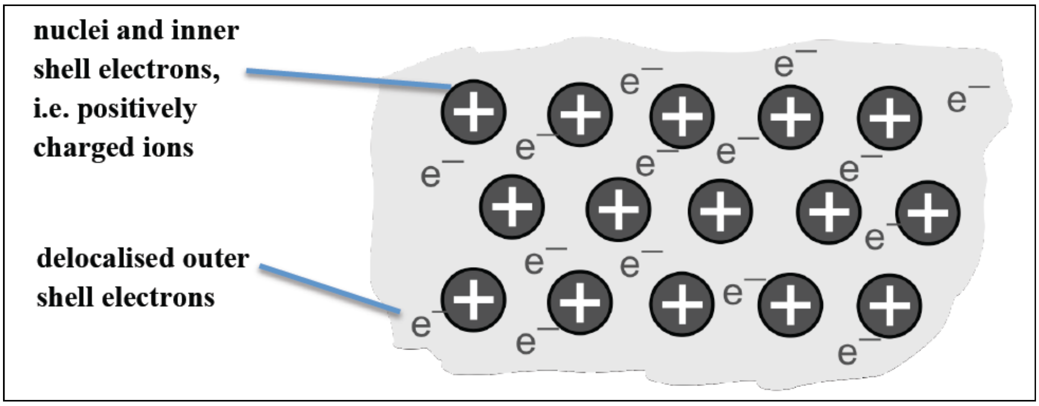
Electrochemical Cells

- Electrically conducting solutions containing ions are known as electrolytes.
- A simple cell can be made by placing two metals in an electrolyte.
- Another type of cell can be made using two half-cells (metals in solutions of their own ions).
- An 'ion bridge' (salt bridge) can be used to link the half-cells.
- Ions can move across the bridge to complete an electrical circuit.
- Electricity can be produced in cells where at least one of the half-cells does not involve metal atoms/ions. A graphite rod can be used as the electrode in such half-cells.
- Different pairs of metals produce different voltages. These voltages can be used to arrange the elements into an electrochemical series.
- The further apart elements are in the electrochemical series, the greater the voltage produced when they are used to make an electrochemical cell.
- Electrons flow in the external circuit from the species higher in the electrochemical series to the one lower in the electrochemical series.
- For an electrochemical cell, including those involving non-metals, ion-electron equations can be written for:
    - the oxidation reaction
    - the reduction reaction
    - the overall redox reactions
- The direction of electron flow can be deduced for electrochemical cells including those involving non-metal electrodes.

Properties of Metals

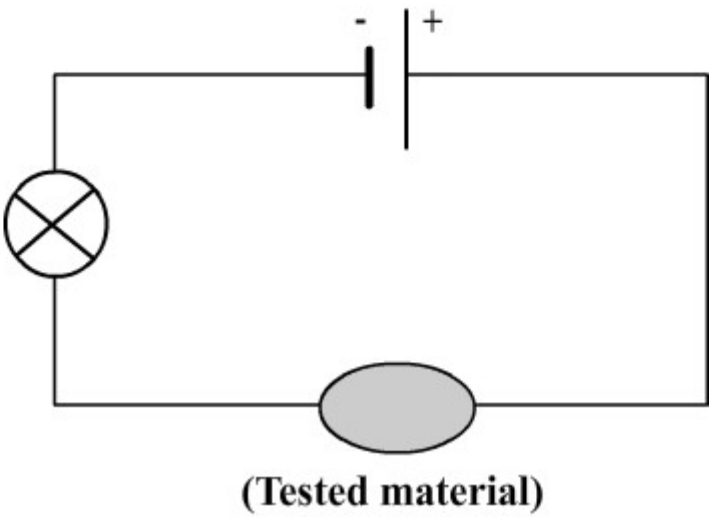
Metallic Bonding

The electrons in the outer electron shell of metal atoms are not tightly held in position. Chemists consider the electrons in metal atoms to be **delocalised** which helps to explain why all metals are able to conduct electricity. Overall the structure of metals (metallic bonding) is regarded as a **regular array of positively charged ions surrounded by a sea of delocalised electrons.** Each ion is attracted to the pool of electrons and vice versa. Since the electrons in the pool do not belong to any particular ion they are free to move from one ion to another - this explains electrical conduction.



Metallic Bonding - an array of positively charged ions surrounded by a sea of delocalised electrons.

Metals conduct electricity in any state because of the delocalised electrons (the electrons are free to move). To test the conductivity of a metal (or any substance) a simple circuit can be set up as shown below:



Reactions of Metals

Metals and Oxygen

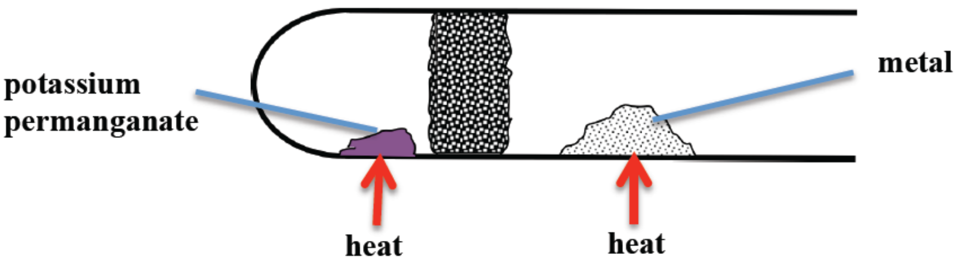
Most metals react with oxygen. When this happens, a metal oxide is formed:

In General:

Metal	**+**	**Oxygen**	$\longrightarrow$	**Metal Oxide**

For example magnesium and oxygen will react to produce magnesium oxide.

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



Metals and Water

Some metals will react with water to form a metal hydroxide solution and release hydrogen gas.

In General:

Metal	**+**	**Water**	$\longrightarrow$	**Metal Hydroxide**	**+**	**Hydrogen**

Some metals are so reactive in water that they must be stored under oil to prevent any moisture getting in contact with the surface of the metal. Potassium will react with water to produced potassium hydroxide and hydrogen:

$$2K(s) + \quad 2H_2O(l) \longrightarrow 2KOH(aq) \quad + \quad H_2(g)$$

Metals and Dilute Acids

Some metals will react with dilute acids to form a salt solution and release hydrogen gas.

In General:

Metal	**+**	**Acid**	$\longrightarrow$	**Salt**	**+**	**Hydrogen**

The balanced ionic equation for magnesium and hydrochloric acid is

$$Mg(s)+ \quad 2HCl(aq) \longrightarrow MgCl(aq) \quad + \quad H_2(g)$$

Reactions of Metals

The Reactivity Series

The reactivity series lists metals in order of reactivity with the most reactive listed first and the least reactivity at the bottom. The following table summarises the reactions of metals with oxygen, water and dilute acid.

<table>
<tr><th rowspan="2">Metal</th><th rowspan="2">Reactivity</th><th colspan="3">Reaction with</th></tr>
<tr><th>Oxygen</th><th>Water</th><th>Dilute Acid</th></tr>
<tr><td>Potassium</td><td rowspan="15">Most Reactive
↓
Least Reactive</td><td rowspan="11">Metals which will react with oxygen</td><td rowspan="4">Metals which will react with water</td><td rowspan="9">Metals which will react with dilute acids</td></tr>
<tr><td>Sodium</td></tr>
<tr><td>Lithium</td></tr>
<tr><td>Calcium</td></tr>
<tr><td>Magnesium</td><td rowspan="11">No Reaction</td></tr>
<tr><td>Aluminium</td></tr>
<tr><td>Zinc</td></tr>
<tr><td>Iron</td></tr>
<tr><td>Lead</td></tr>
<tr><td>Hydrogen</td><td rowspan="6">No Reaction</td></tr>
<tr><td>Copper</td></tr>
<tr><td>Mercury</td><td rowspan="4">No Reaction</td></tr>
<tr><td>Silver</td></tr>
<tr><td>Gold</td></tr>
<tr><td>Platinum</td></tr>
</table>

Reactions of Metals

Oxidation and Reduction

Some metals are more reactive than others. When metals take part in chemical reactions they lose electrons. For example magnesium will react with oxygen forming magnesium oxide.

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

If we consider what happens to the magnesium atoms:

Magnesium atoms have electron arrangement 2, 8, 2. Mg

Magnesium ions have stable electron arrangement 2, 8 Mg^{2+}

Overall $Mg \longrightarrow Mg^{2+} + 2e^-$

A chemical reaction where a species losses electrons is called **oxidation**. Equations which show the movement of electrons are called **ion-electron equations**. A list of reduction equations is shown on page 10 of the data book.

If we consider what happens to the oxygen atoms:

Oxygen atoms have electron arrangement 2, 6. O

Oxygen ions have electron arrangement 2, 8 O^{2-}

Overall $O + 2e^- \longrightarrow O^{2-}$

A chemical reaction where a species gains electrons is called **reduction**.

An easy way to remember this is **OILRIG;**

O xidation
I s
L oss
R eduction
I s
G ain

When you have an oxidation and reduction reaction occurring **together** the overall reaction is called **redox.** When metals react with oxygen, water and dilute acids a redox reaction occurs. The metal is always undergoing oxidation.

Reactions of Metals

Redox Equations

The reactions of metals with oxygen, water and dilute acids are all example of redox since oxidation and reduction take place. Redox equations are written by combining oxidation and reduction equations.

Example 1: Calcium and oxygen

Oxidation $Ca(s) \longrightarrow Ca^{2+}(s) + 2e^-$

Reduction $\frac{1}{2}O_2(g) + 2e^- \longrightarrow O^{2-}(s)$

A redox equation can be written by combining the oxidation and reduction equations but it must **never contain electrons**. The oxidation and reduction equations may have to multiplied so that the **electrons cancel**. In this example the 2 electrons in each equation will cancel anyway.

Redox equation $Ca(s) + \frac{1}{2}O_2(g) \longrightarrow Ca^{2+}O^{2-}(s)$

Example 2: Sodium and water

Oxidation $Na(s) \longrightarrow Na^+(aq) + e^-$

Reduction $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

In this example the oxidation equation must be **multiplied by two** and allow the electrons to cancel.

Redox equation $2Na(s) + 2H^+(aq) \longrightarrow 2Na^+(aq) + H_2(g)$

Example 3: Magnesium and hydrochloric acid

Oxidation $Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$

Reduction $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

In this example the electrons are balanced and will cancel out.

Redox equation $Mg(s) + 2H^+(aq) \longrightarrow 2Mg^{2+}(aq) + H_2(g)$

Example 4: Aluminium and hydrochloric acid

Oxidation equation $Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$

Reduction equation $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

To cancel out the electrons, the **oxidation must be multiplied by 2**, and the **reduction by 3.**

Oxidation $2Al(s) \longrightarrow 2Al^{3+}(aq) + \mathbf{6e^-}$

Reduction $6H^+(aq) + \mathbf{6e^-} \longrightarrow 3H_2(g)$

Redox $2Al(s) + 6H^+(aq) \longrightarrow 2Al^{3+}(aq) + 3H_2(g)$

Extraction of Metals

Metal Ores

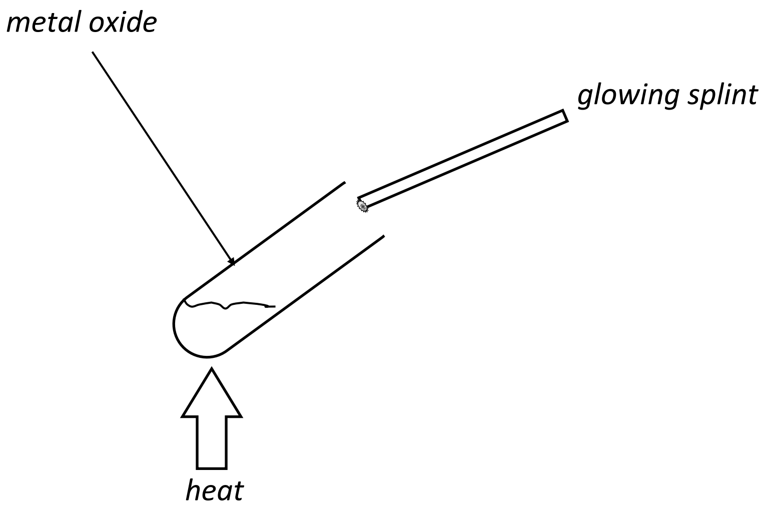
Metals such as gold and silver are unreactive and are found uncombined in the Earth's crust as the metallic element.

Most metals are found in the ground combined with other elements in the form of ionic compounds called ores. An **ore** is defined as a naturally occurring metal compound. Examples of ores are listed below.

Name of ore	Metal compound found in the ore
haematite	iron oxide
bauxite	aluminium oxide
iron pyrite (fool's gold)	iron sulfide

A metal can be extracted from its ore. Extraction involves separating the metal from the other elements in the ore. The method of extraction depends on the position of the metal in the reactivity series. In general the less reactive the metal the easier it is to extract the metal from its ore. **Extraction** of metals from ores is an example of a **reduction reaction** since the metal ions gain electrons to produce metal atoms.

Extraction by Heat Alone



Some metals can be extracted by heat alone such as mercury and silver, from their corresponding oxides. Most metals are too reactive to be extracted by heat alone.

In General:

$$\text{Metal Oxide} \xrightarrow{\text{Heat}} \text{Metal} + \text{Oxygen}$$

Extraction of Metals

Extraction by Heating with Carbon

Some metals can be extracted by heating with carbon or carbon monoxide such as zinc, tin, lead, copper and iron from their corresponding oxides. Metals above zinc in the reactivity series are too reactive to be extracted by heating with carbon or carbon monoxide gas. In this extraction process the metal ions are reduced and the carbon is oxidised.

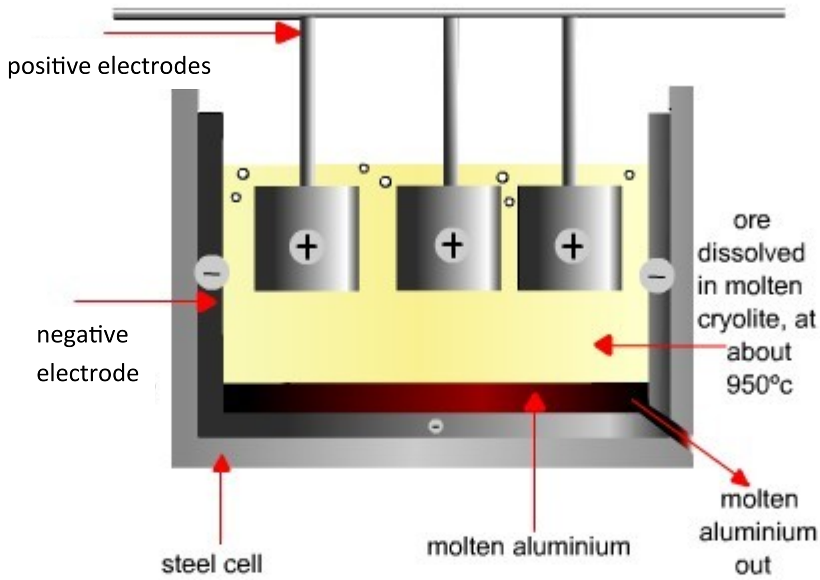
In General:

Metal Oxide	**+**	**Carbon (monoxide)**	$\xrightarrow{\textbf{Heat}}$	**Metal**	**+**	**Carbon Dioxide**

Electrolysis

Very reactive metals are strongly bonded in their ores and cannot be extracted using carbon/carbon monoxide. More energy is needed and this is provided by electrolysis. Electrolysis is a process which uses electricity to break down an ionic substance.

Metals above zinc in the reactivity series are extracted using electrolysis such as aluminium. Aluminium is extracted from its molten ore, bauxite (Al_2O_3).



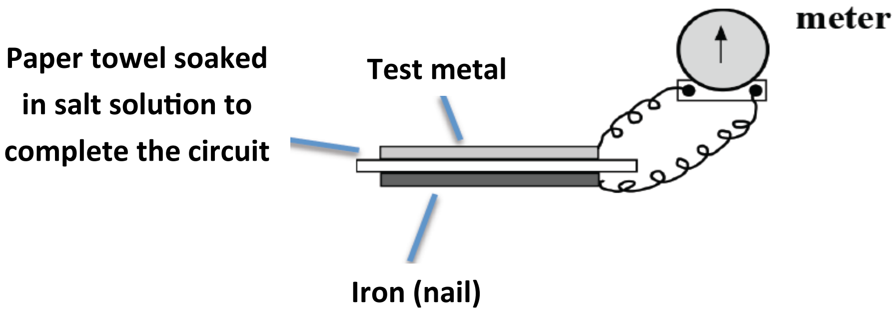
When electricity is passed through the molten ore, the positive ions are attracted to the negative electrode where they gain electrons producing metal atoms. The aluminium flows out as a liquid.

Reaction at the negative electrode: $Al^{3+}(l) + 3e^- \longrightarrow Al(l)$ (reduction)

Reaction at the positive electrode: $2O^{2-}(aq) \longrightarrow O_2(g) + 4e^-$ (oxidation)

Electrochemical Cells

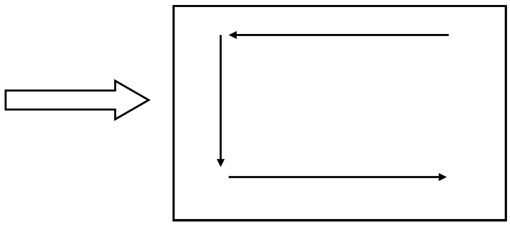
If two different metals are linked together in a certain way a flow of electrons (electricity) can be produced. A simple electrochemical cell will consist of two different metals connected through an electrolyte and external wires. Electrons flow through the wires between the two metals and ions flow through the electrolyte completing the circuit. The direction of electron flow depends on the metals used.



The Electrochemical Series

Electrons flow from the metal higher in the electrochemical series to the metal below. The size of the voltage produced depends on the relative positions of the two metals. The bigger the gap between the two metals in the electrochemical series the larger the voltage. (See data booklet page 10.)

Direction of electron flow on the electrochemical series.



Redox Equations for Cells

The electrochemical series places metals in order of their ability to lose electrons to form ions. Metals higher in the electrochemical series are more likely to lose electrons than metals lower down, e.g. magnesium is more likely to lose electrons than copper.

$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$ **More likely**

$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$ **Less likely**

When these two metals are attached in an electrochemical cell electrons will flow from magnesium to copper. This means the magnesium atoms are oxidised and the copper ions reduced, in other words a redox reaction occurs. Electrons must be cancelled in a redox equation.

Oxidation: $Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$

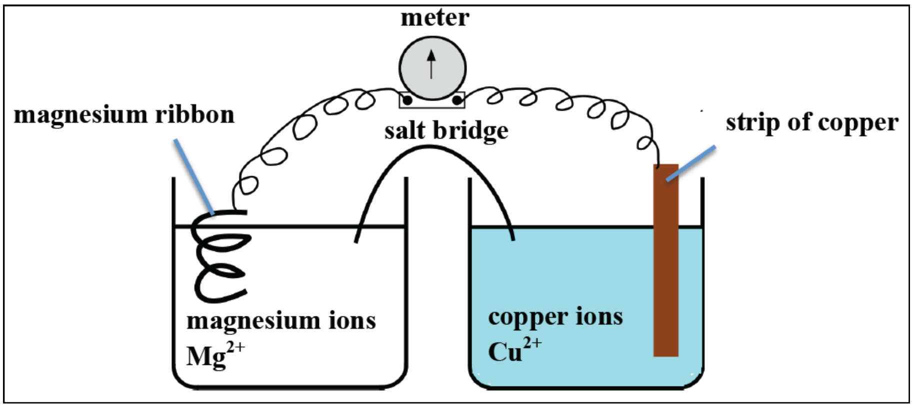
Reduction: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Redox Eq: $Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$

Electrochemical Cells

Half Cells

Metals ionise more efficiently when they are dipped into a solution already containing their own ions. A half cell can be produced where the two metals are placed in a solution of their own ions and linked with a salt bridge (ion bridge) to complete the circuit. The two metals are attached through wires, electrons flow through the wires and ions through the salt bridge.



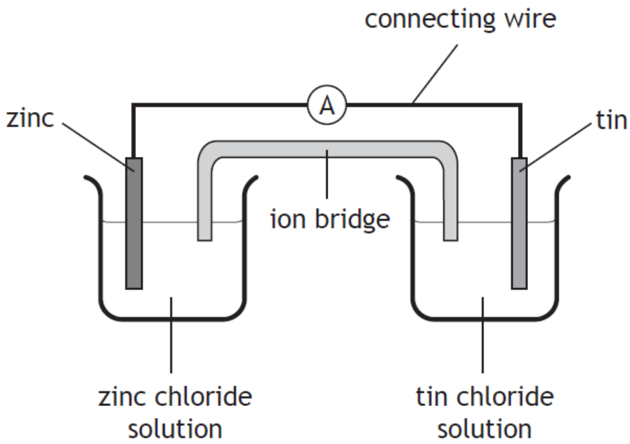
Magnesium-Copper Half Cell

Oxidation: $Mg(s) \longrightarrow Mg^{2}(aq) + 2e^{-}$

Reduction: $Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$

Redox Eq: $Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2}(aq) + Cu(s)$

Zin-Tin Half Cell



Oxidation: $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$

Reduction: $Sn^{2+}(aq) + 2e^{-} \longrightarrow Sn(s)$

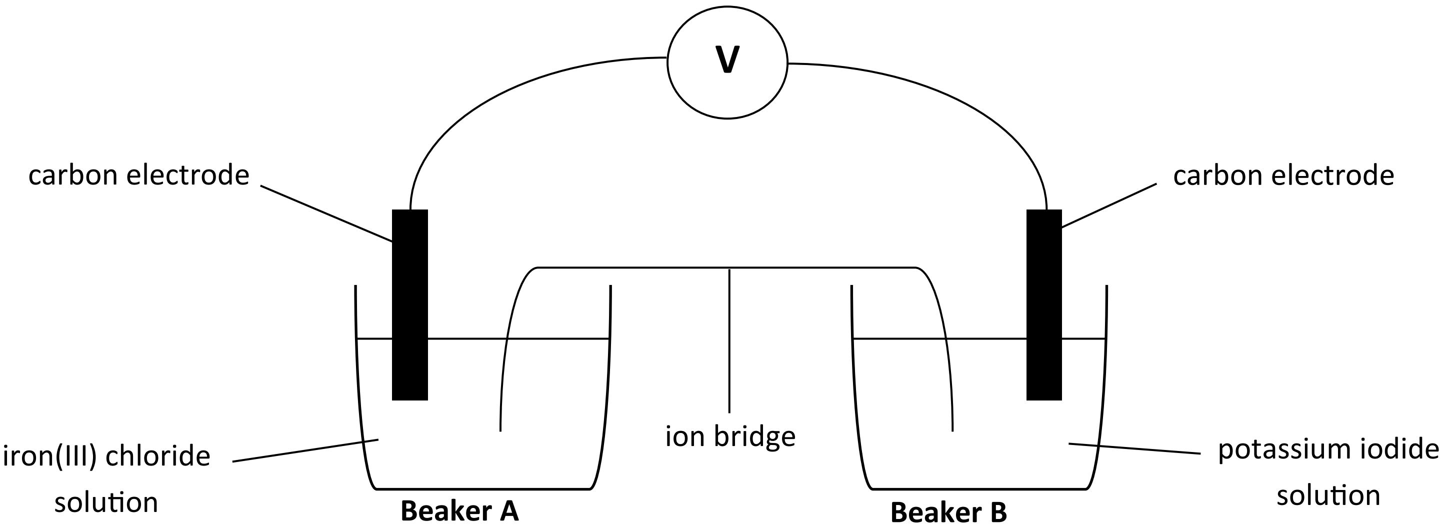
Redox Eq: $Zn(s) + Sn^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Sn(s)$

Electrochemical Cells

Non-Metal Cells

Electrochemical cells can be produced using non-metals where carbon in the form of graphite is used as the electrodes. Electrons can be removed from one half cell (oxidation) and donated to the other half cell where reduction occurs.

The Iron(III)-Iodide Cell



The Iron(III)-Iodide Cell Equations:

Oxidation:	$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$	
Reduction:	$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	**(multiplied by 2)**
Redox:	$2I^-(aq) + 2Fe^{3+}(aq) \longrightarrow I_2(aq) + 2Fe^{2+}(aq)$	

Topic Questions

Exercise 1

1. Write a word equation for the reaction of sodium and water.
2. Write a balanced chemical equation for the reaction of sodium and water.
3. What type of solution is formed when a metal reacts with water?
4. Write a balance ionic equation for the reaction of lithium and water.
5. Write a balanced ionic equation for the reaction of calcium and water.
6. Write a word equation for the reaction of magnesium and oxygen.
7. Write a balanced chemical equation for the reaction of magnesium and oxygen.
8. Write the ionic equation for the reaction of magnesium and oxygen.
9. Write the ionic equation for the reaction between aluminium and oxygen.

Exercise 2

1. Name the salt formed when the following metals and acids react.
   a) Zinc and hydrochloric acid
   b) Iron and nitric acid
   c) Magnesium and sulfuric acid
   d) Aluminium and nitric acid
   e) Calcium and hydrochloric acid
   f) Magnesium and nitric acid
   g) Lithium and hydrochloric acid
   h) Sodium and sulfuric acid
2. Name 4 metals which will not react with dilute acid.
3. Write a word equation for the reaction of magnesium and hydrochloric acid.
4. Write a balanced chemical equation for the reaction between magnesium and hydrochloric acid.
5. Write a word equation for the reaction between sodium and hydrochloric acid.
6. Write a balanced chemical equation for the reaction between sodium and hydrochloric acid.

Exercise 3

1. Name two metals which are found uncombined in the Earth's crust.
2. Explain why these two metals are found uncombined.
3. What is meant by the term ore?
4. What is meant by the term extraction?
5. What type of chemical reaction occurs when a metal is extracted from its ore?
6. Give an example of a metal which would be easily extracted from its ore.

7. Give an example of a metal which would be difficult to extract from its ore.
8. What are the 3 methods of extracting metals from their ores?
9. Name two metals which can be extracted by heat alone.
10. Name three metals which can be extracted using carbon.
11. Name a reducing agent other than carbon which can be used to extract the metals you gave in Q10.
12. Name 3 metals which must be extracted using electrolysis.
13. Explain why metals like aluminium and sodium were not extracted until the 19th century when other metals like gold have been known since ancient times.
14. Give the reduction equation for the extraction of aluminium from aluminium ore.

Exercise 4

1. Calculate the percentage by mass of iron in haematite, Fe_2O_3.
2. Calculate the percentage by mass of sodium in sodium carbonate, Na_2CO_3.
3. Calculate the percentage by mass of copper in malachite, $Cu_2CO_3(OH)_2$.
4. Calculate the mass of aluminium which could be obtained from 500 g of aluminium oxide, Al_2O_3.
5. Calculate the mass of lead which could be obtained by the electrolysis of 50 kg of molten lead bromide, $PbBr_2$.

Exercise 5

1. What is meant by the term reduction?
2. What is meant by the term oxidation?
3. For each of the following equations state if they represent an oxidation or reduction reaction.
   a) $K^+(aq) + e^- \longrightarrow K(s)$
   b) $Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$
   c) $Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$
   d) $Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$
   e) $Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$
4. Write ion-electron equations for the following reactions.
   a) Oxidation of nickel
   b) Reduction of sodium
   c) Reduction of zinc
   d) Oxidation of bromide ions
   e) Oxidation of gold
5. What is meant by the term redox?

Exercise 6

1. Aluminium will react with hydrochloric acid.

    a) Write the ion-electron equation for the oxidation reaction when aluminium reacts with hydrochloric acid.

    b) Write the ion-electron equation for the reduction reaction when aluminium reacts with hydrochloric acid.

    c) Write a balanced ionic equation for this reaction.

2. Explain in terms of the position of gold relative to hydrogen in the electrochemical series on page 10 of the data booklet why gold will not react with acidic solutions.

Exercise 7

1. Predict the direction of electron flow through a wire containing the following pairs of metals;

    a) zinc/magnesium

    b) copper/tin

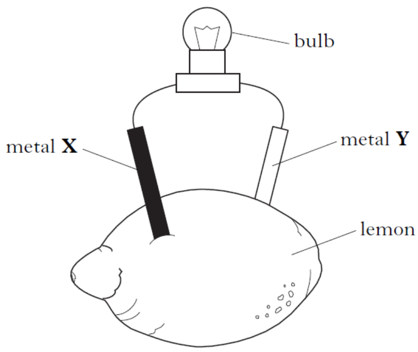
    c) aluminium/iron

    d) silver/tin

    e) nickel/iron

2. Explain why an electrochemical cell with magnesium and silver as the two metals will produce a greater voltage than a cell made with magnesium and copper.

3. A pupil set up the following lemon cell.



Which pair of metals in the table would give the brightest bulb?

	Metal X	Metal Y
A	magnesium	copper
B	copper	copper
C	zinc	copper
D	iron	copper

4. For each of the following cells state the oxidation and reduction equations and then combine them to give the overall redox equation. (Remember electrons must cancel.)

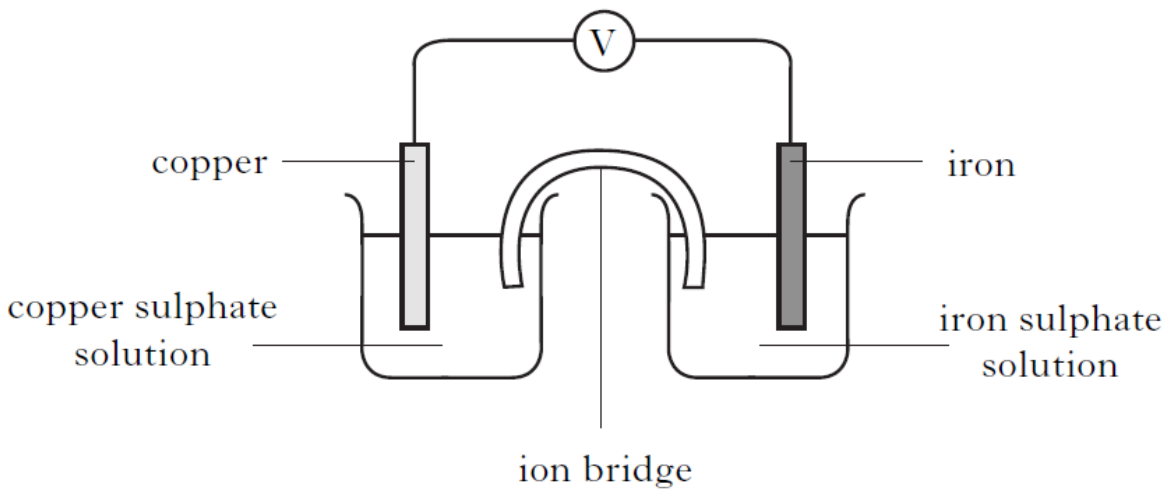
a)	Magnesium/zinc cell	b)	Iron/nickel cell
c)	Copper/tin cell	d)	Silver/magnesium cell
e)	Aluminium/copper cell	f)	Iron/aluminium cell

5. For each of the following, state which cell would give the greatest voltage.

a)	Ni/Pb and Fe/Cu	b)	Fe/Ag and Sn/Ag
c)	Al/Ag and Mg/Au	d)	Zn/Pb and Al/Hg

Exercise 8

1. Explain the direction of the electron flow in a magnesium-copper half cell.
2. What is the purpose of the ion (salt) bridge?
3. What passes through the ion bridge?
4. Why was the ion bridge soaked in salt solution yet sugar solution would be unsuitable?
5. A student set up the following electrochemical cell.



a) Copy the diagram above and indicate on the wires the direction of electron flow.

b) Give the oxidation and reduction equations for this cell.

c) Combine the oxidation and reduction equations to produce a redox equation.

d) Suggest a voltage for this cell based on the voltage you recorded in activity 5.

e) Suggest why sodium carbonate solution would not be suitable for use in the ion bridge. *(You may wish to refer to page 8 in the data booklet.)*

Exercise 9

1. Which of the following metals can be obtained from its ore by heating with carbon monoxide?

   A Aluminium B Calcium C Magnesium D Nickel

2. Which of the following metals is produced in a blast furnace?

   A Calcium B Copper C Iron D Sodium

3. Some metals can be obtained from their metal oxides by heat alone. Which of the following oxides would produce a metal when heated?

   A Calcium oxide B Copper oxide C Zinc oxide D Silver oxide

4. The ion-electron equation for the oxidation and reduction steps in the reaction between magnesium and silver(I) ions are:

$$Mg \rightarrow Mg^{2+} + 2e^-$$
$$Ag^+ + e^- \rightarrow Ag$$

   The overall redox equation is

   A $Mg + 2Ag^+ \longrightarrow Mg^{2+} + 2Ag$

   B $Mg + Ag^+ \longrightarrow Mg^{2+} + Ag$

   C $Mg + Ag^+ + e^- \longrightarrow Mg^{2+} + Ag + 2e^-$

   D $Mg + 2Ag^+ \longrightarrow Mg^{2+} + 2Ag^+$.

5. Aluminium can be extracted from aluminium oxide by

   A heating alone B heating with carbon

   C heating with carbon monoxide D electrolysis.

6. Which of the following metals does not react with dilute acid?

   A Magnesium B Calcium C Copper D Zinc

7. A student adds a powder to dilute hydrochloric acid. A gas which burns with a pop is produced. The powder could be

   A carbon B calcium oxide C sodium carbonate D zinc

Exercise 10

1. Metal can be extracted from their ores by different methods.

   a) Copy and complete the following table to list the methods correctly.

   reacting with carbon **electrolysis** **heat alone**

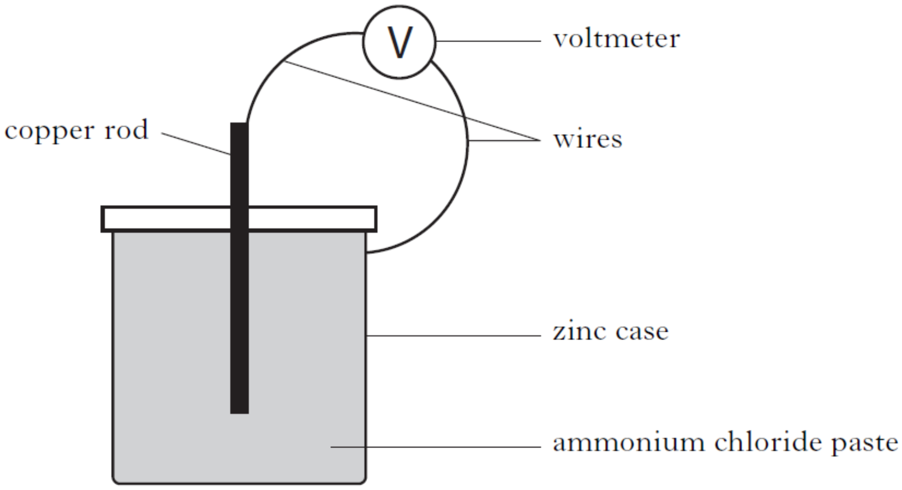
Metal	Method
mercury	
iron	
magnesium	

   b) i) Mercury can be extracted from the ore cinnabar, HgS.
   Calculate the percentage by mass of mercury in cinnabar.

   ii) Write the formula for the mercury ion in cinnabar.

   iii) Write the reduction equation forv the reaction of the mercury ion to form mercury.

2. A student set up the following cell.



   a) In this cell the purpose of the ammonium chloride paste is to complete the circuit.

   i) What name is used to describe an ionic compound which is used for this purpose?

   ii) What type of charged particle will move through the ammonium chloride paste?

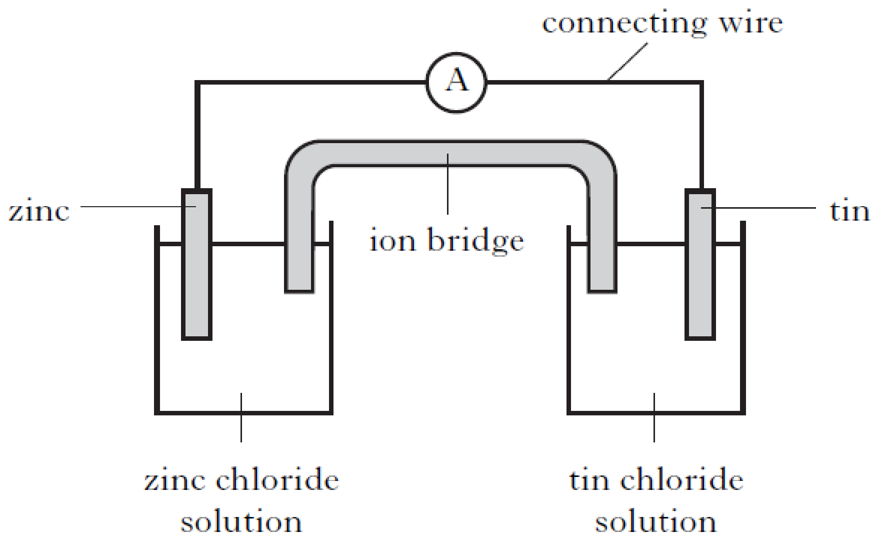
   iii) Why will no voltage be produced if the ammonium chloride paste dries out?

   b) What way do the electrons flow in this cell through the wires?

   c) Write the ion-electron equations for each of the oxidation and reduction reactions.

   d) Combine the oxidation and reduction equations to write an overall redox equation for the cell.

3. The following cell was set up.



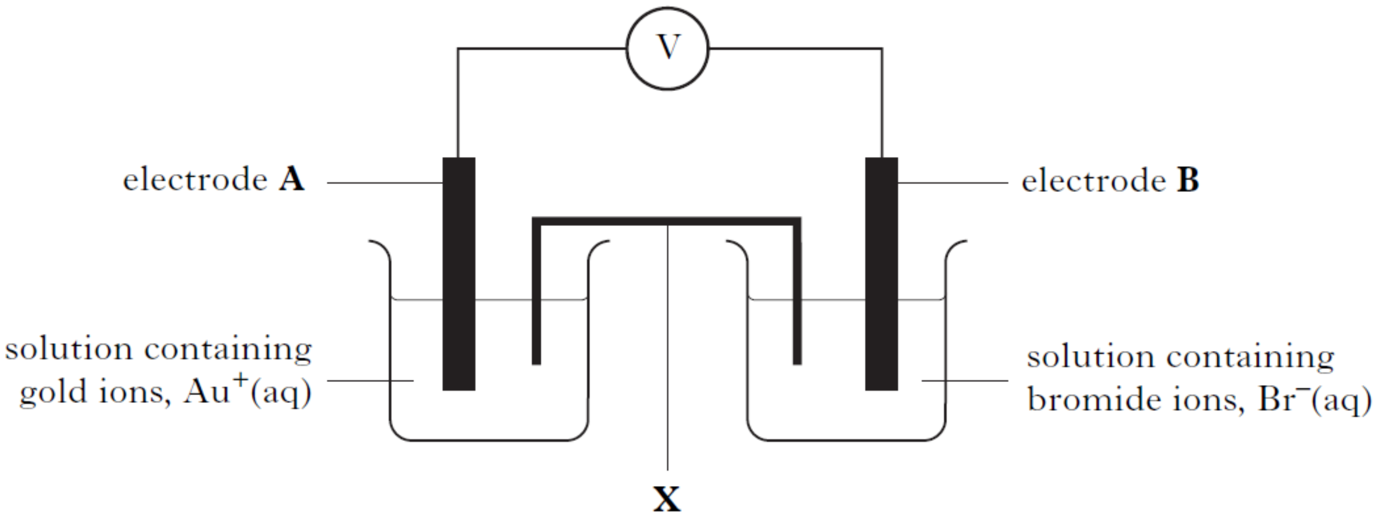
a) What is the purpose of the ion bridge?

b) What direction will electrons flow through the connecting wire?

c) Give both the oxidation and reduction equations for this cell.

d) Combine the oxidation and reduction equations to give the overall redox equation for the cell.

4. A technician set up the following non-metal cell.



a) Oxidation occurs at electrode **A** while reduction occurs at electrode **B**. In what direction will electrons flow through the wires?

b) State the ion-electron equations for both the oxidation and reduction reactions.

c) Write the redox equation for this cell.

d) Name the piece of apparatus labelled **X**.

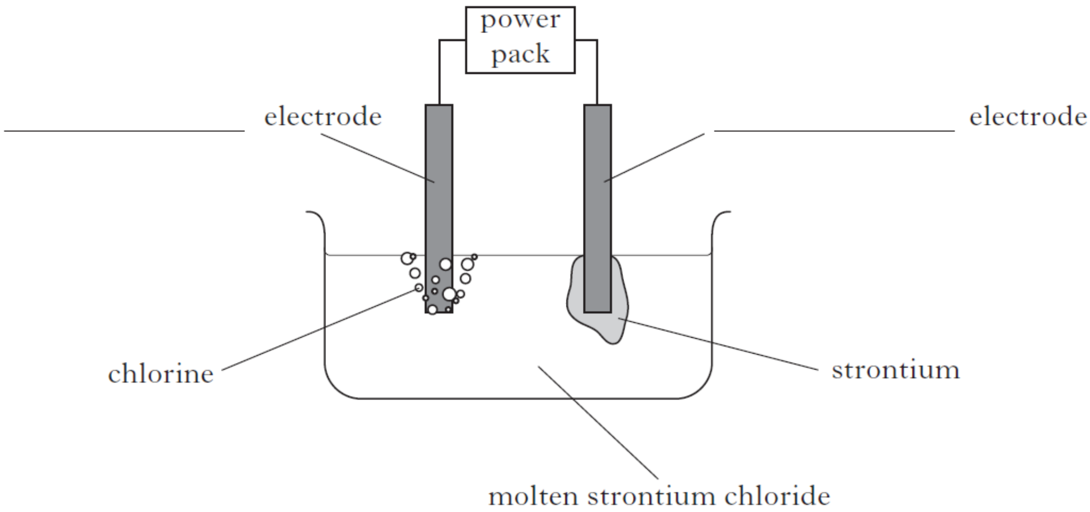
e) Gradually a colour appears in beaker with electrode **B**. Suggest why a colour appears.

Exercise 11

1. a) Strontium compounds have many uses such as in fireworks and toothpaste. Strontium chloride hexahydrate can be used in toothpaste for sensitive teeth as it plugs the holes in the tooth enamel. This is possible because strontium has similar chemical properties to calcium. Why does strontium have similar chemical properties to calcium?

   b) Strontium can be extracted from the compound strontium chloride, $SrCl_2$, using electrolysis.

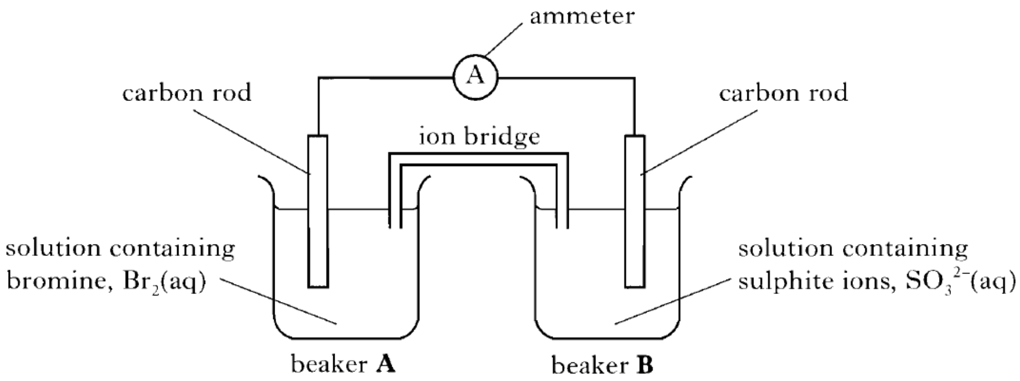
   Copy and complete the diagram to show the charge on each electrode.



   c) i) Write the ion-electron equation for the reaction which occurs at each electrode.

   ii) Combine the equations in c) i) to form a redox equation.

2.. The following cell can be used to produce a current.



The reaction in beaker B is $SO_3^{2-}(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

   a) What way will electrons flow between the carbon rods?

   b) i) What term is used to describe the type of reaction taking place in beaker **B**?

   ii) Suggest what would happen to the pH in beaker **B.**

   c) Write the ion-electron equation for the chemical reaction taking place in beaker **A**.

   d) Write the redox equation for this electrochemical cell.

Topic 2

Plastics

& Fertilisers

Learning Outcomes

Addition polymerisation

- Plastics are examples of materials known as polymers.
- Polymers are long chain molecules formed by joining together a large number of small molecules called monomers.
- Addition polymerisation is the name given to a chemical reaction in which unsaturated monomers are joined, forming a polymer.
- The names of addition polymers are derived from the name of the monomer used. Note: brackets can be used in polymer names to aid identification of the monomer unit.

Representation of the structure of monomers and polymers

- A repeating unit is the shortest section of polymer chain which, if repeated, would yield the complete polymer chain (except for the end-groups).
- The structure of a polymer can be drawn given either the structure of the monomer or the repeating unit.
- From the structure of a polymer, the monomer or repeating unit can be drawn.

Commercial production of fertilisers

- Growing plants require nutrients, including compounds containing nitrogen, phosphorus or potassium.
- Fertilisers are substances which restore elements, essential for healthy plant growth, to the soil.
- Ammonia and nitric acid are important compounds used to produce soluble, nitrogen-containing salts that can be used as fertilisers.
- Ammonia is a pungent, clear, colourless gas which dissolves in water to produce an alkaline solution.
- Ammonia solutions react with acids to form soluble salts.
- ammonia solution + an acid $\longrightarrow$ an ammonium salt + water

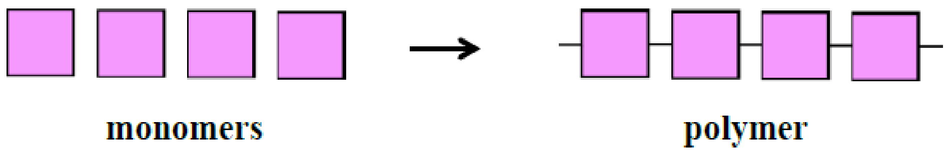
Haber and Otswald process

- The Haber process is used to produce the ammonia required for fertiliser production.
- $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3\,(g)$
- At low temperatures the forward reaction is too slow to be economical. If the temperature is increased, the rate of reaction increases but, as the temperature increases, the backward reaction becomes more dominant.
- An iron catalyst is used to increase reaction rate.
- Ammonia is the starting material for the commercial production of nitric acid.
- The Ostwald process uses ammonia, oxygen and water to produce nitric acid.
- A platinum catalyst is used in this process.

Plastics

Making Plastics

The molecules in plastics are made by joining up lots of small molecules called **monomers** (‘mono’ meaning one or single unit) to make very long chains, often consisting of thousands of atoms linked together. The large molecules which are formed from monomers are called **polymers**. Plastics are examples of polymers. When polymers join together this is called a **polymerisation** reaction.

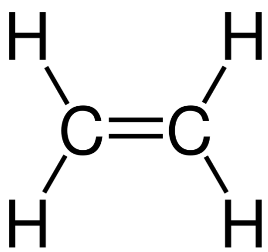


Monomers

Monomers are defined as small single units which join together to create large chains called polymers.

Remember many monomers make a polymer!!

The simplest monomer is ethene:



A polymer made of many ethenes joined together would be called **poly**ethene (meaning *many* ethenes’).

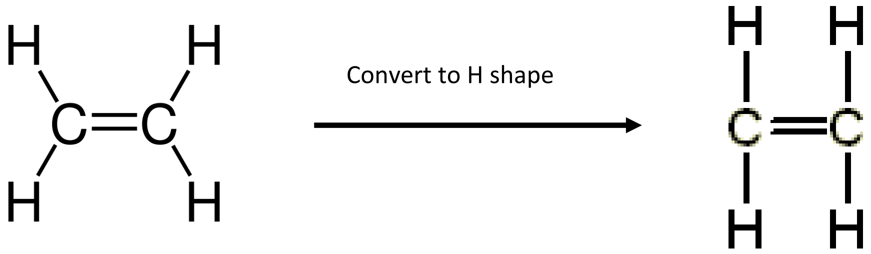
Addition Polymerisation

Addition polymerisation occurs between unsaturated monomers (monomers which contain at least one carbon-to-carbon double bond eg. ethene)

The shape of the monomer must first be changed into a ‘H’ shape and then the monomers are joined together by breaking the carbon-to-carbon double bond and creating new single bonds between the monomer units.

Changing the shape of the monomer unit

Ethene

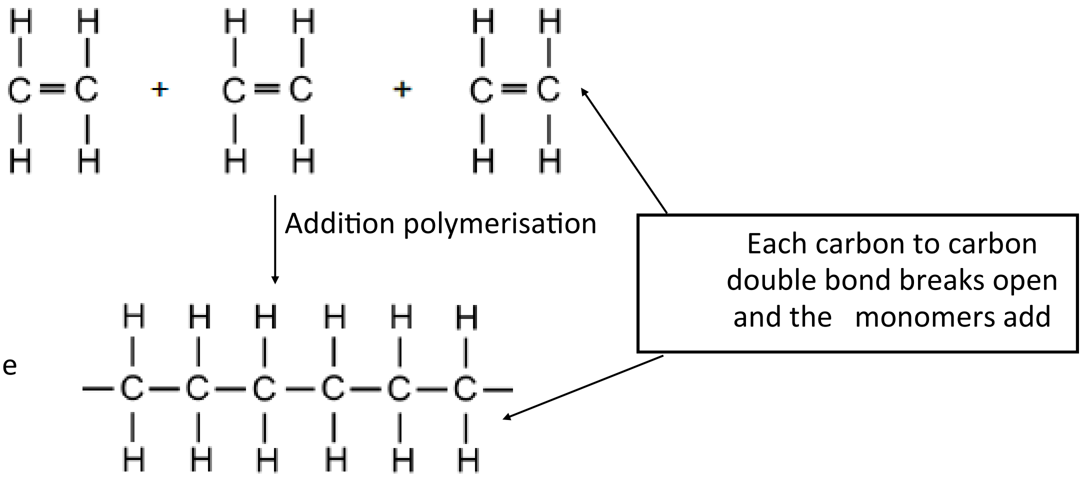


Example 1: Poly(ethene)

Poly(ethene) is formed from ethene monomers. When drawing part of an addition polymer structure we show how three monomer molecules join to produce part of the polymer chain as shown below.

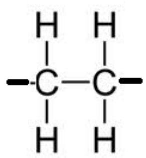
Ethene monomers

Poly(ethene) molecule



Addition polymers have a two carbon **repeating unit**. This is based on the two carbons originally formed by the double bond.

The repeating unit in poly(ethene) is:

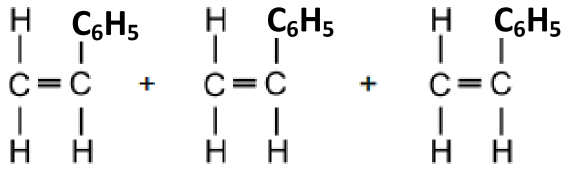


Addition Polymerisation

Example 2: Poly(phenylethene) or Poly(styrene)

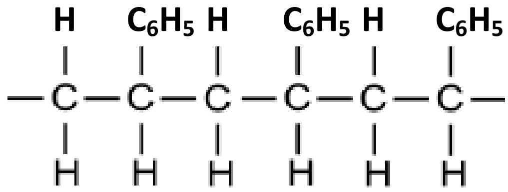
Poly(styrene) is a common addition polymer formed when styrene (or phenylethene) monomers join together as the double carbon to carbon bonds break.

Styrene monomers

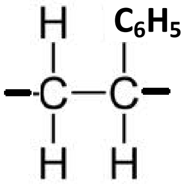


Addition polymerisation

Poly(styrene) molecule



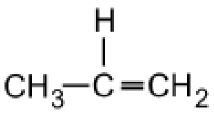
The repeating unit in poly(styrene) is:



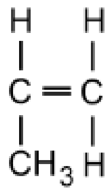
Example 3: Poly(propene)

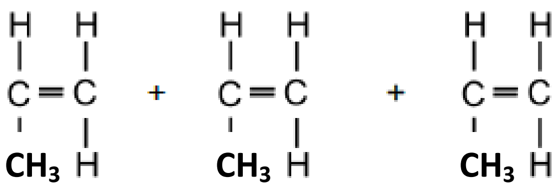
Poly(propene) is a thermoplastic often used in food containers and ropes. In drawing the polymer structure for propene it is easier to show the monomer in the rugby goalpost style.

Propene molecule

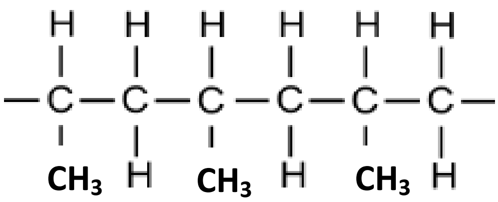


or alternatively as a monomer

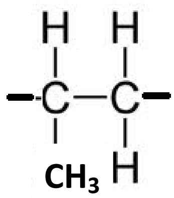




Addition polymerisation



repeating unit:



Addition Polymerisation

Points to Remember when drawing polymers:

- When drawing larger groups 'C_2H_5' or 'CH_3' make sure your bonds goes to the correct element in the group.
- Polymer chains must have open bonds at the end of the molecule to show that they can continue on.
- Repeating units must also show open bonds as they are a representation of a section within the polymer chain.
- Monomers do not have open bonds but do have a carbon-to-carbon double bond in the middle of the unit!

Naming Addition Polymers

Many of the monomers are small unsaturated molecules, i.e. small molecules with a carbon to carbon double covalent bond. The monomer molecules join together as the carbon to carbon double bonds open and to give a long chain of carbon to carbon single covalent bonds. **Polymers formed in this way are called addition polymers and the process is called addition polymerisation**.

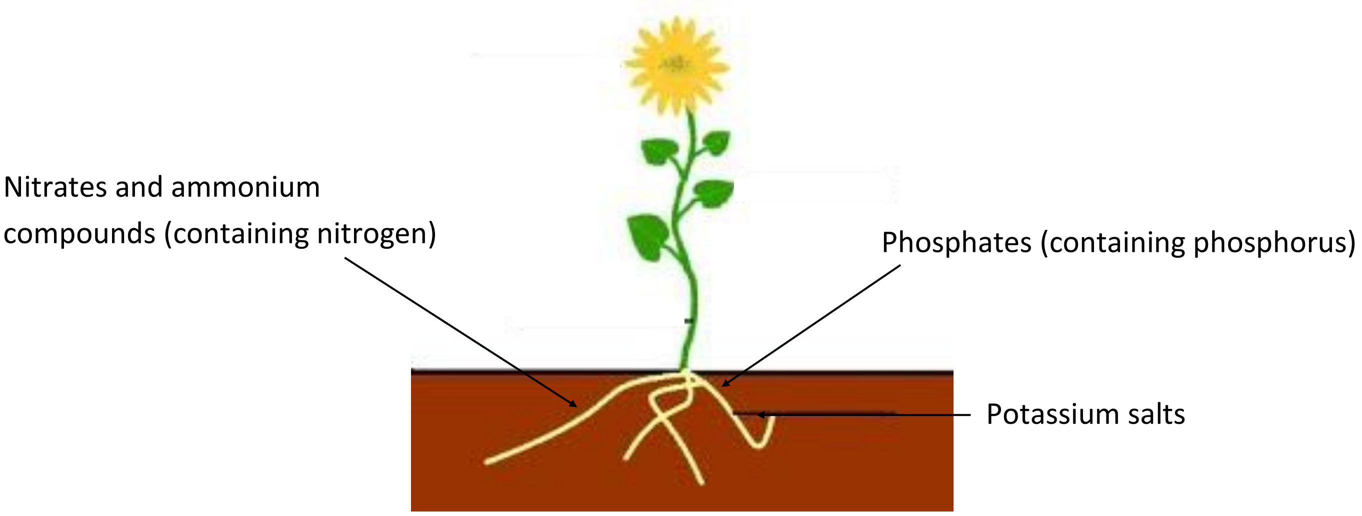
Addition polymers are named according to the monomer from which they are made. Some examples are given in the table.

Monomer name	Polymer name
ethene	poly(ethene)
propene	poly(propene)
chloroethene	poly(chloroethene)
tetrafluoroethene	poly(tetrafluoroethene)
phenylethene	poly(phenylethene)

Fertilisers and Nutrients

Essential Elements for Plant Growth

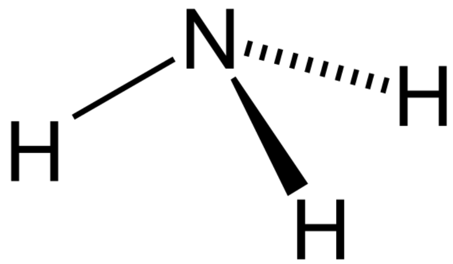
Healthy crops provide more food and growing plants require nutrients, chemicals essential for healthy plant growth. Elements required by plants include hydrogen, carbon and oxygen, most importantly plants need **nitrogen, phosphorous** and **potassium** (NPK). Plants take these elements in from the soil through their roots. The elements are taken as compounds which are dissolved in water from the soil.



When crops grow, they remove some of the essential elements from the soil since crops are harvested to feed people. The nutrients removed by the crops need to be added back to the soil and this is achieved by the addition of a fertiliser.

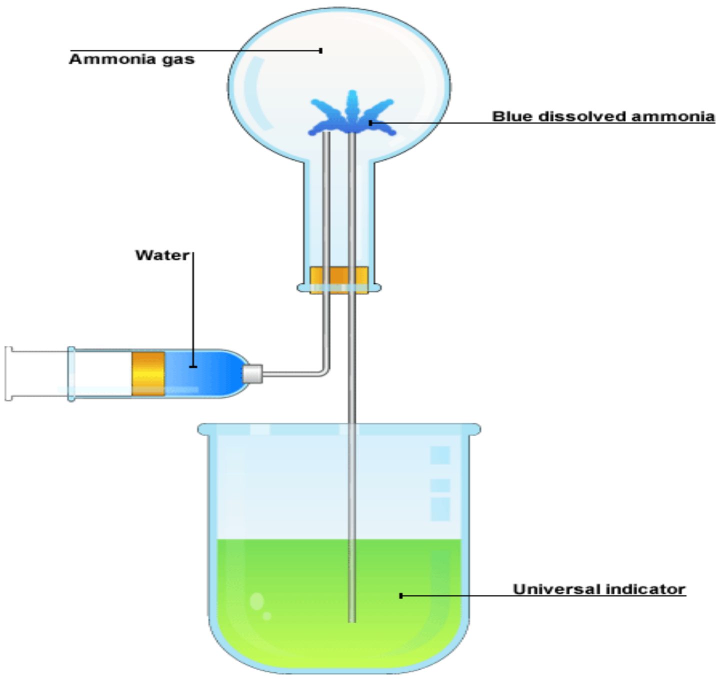
Ammonia (NH_3) and Nitric Acid (HNO_3) are important compounds used to produce soluble nitrogen containing salts which can be used as fertilisers.

Ammonia



Ammonia is a pungent, clear, colourless gas which dissolves in water to produce an alkaline solution. It has the formula NH_3.

The Ammonia Fountain



As the water with indicator is added the gas dissolves since it is **highly soluble** turning the indicator blue/purple (indicating an alkaline solution is being formed).

As an alkaline solution is produced here. Ammonia solution can react with acids in a neutralisation reaction to produce soluble salts.

ammonia solution + an acid $\longrightarrow$ an ammonium salt + water

Eg. **ammonia + hydrochloric acid $\longrightarrow$ ammonium chloride + water**

ammonia + nitric acid $\longrightarrow$ ammonium nitrate + water

Haber Process

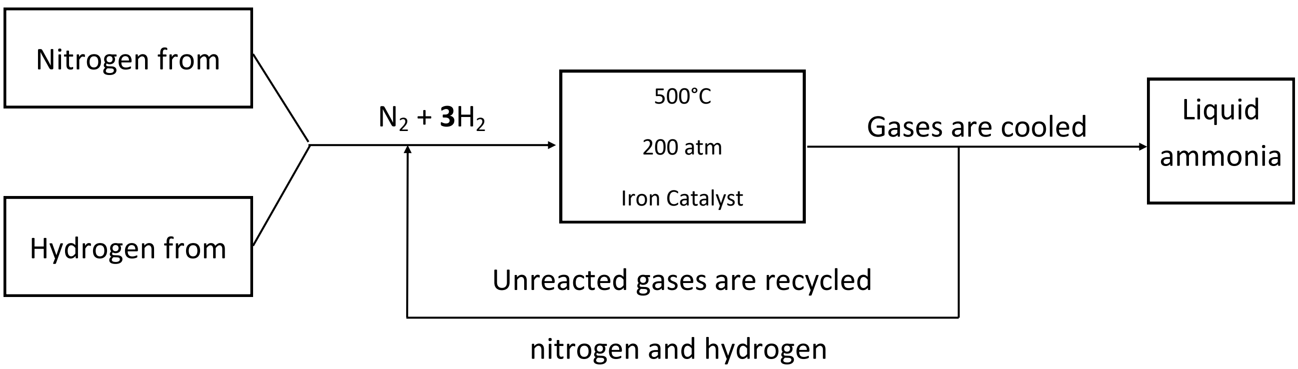
Making Ammonia - The Haber Process.

Chemists began to look at ways of producing **artificial fertilisers containing nitrogen compounds.**

Although nitrogen was the obvious starting point, they had huge difficulty getting it to react.

The breakthrough came in 1904 when a German Chemist, FRITZ HABER, found a way of turning nitrogen and hydrogen into AMMONIA.

Ammonia is manufactured commercially by the **Haber Process.**



The reaction between nitrogen and hydrogen to make ammonia is reversible. This means that as ammonia is produced, it begins to break back down into nitrogen and hydrogen again.

This means it is not possible to convert all of the hydrogen and nitrogen into ammonia.

The use of a double arrow indicates a reversible reaction.

$$N_2(g) \quad + \quad 3H_2(g) \quad \rightleftharpoons \quad 2NH_3(g)$$

In order to improve the yield of ammonia a **high pressure** and **moderately high temperature** is used. At low temperatures the forward reaction is too slow to be economical. If the temperature is increased, the rate of reaction increases but, as the temperature increases, the backward reaction becomes more dominant.

An iron catalyst is used to increase reaction rate.

The Otswald Process

Nitric Acid

Nitric acid is used to make nitrate fertilisers, like potassium nitrate and ammonium nitrate. It can be made by nitrogen dioxide produced in lightning storms and in car engines due to the spark plug which ignites the gases present in the air. To make nitrogen dioxide, nitrogen molecules need to react with oxygen molecules. This requires large quantities of energy due to the very strong triple covalent bond between the nitrogen atoms.

Word Equation: nitrogen + oxygen $\longrightarrow$ nitrogen dioxide

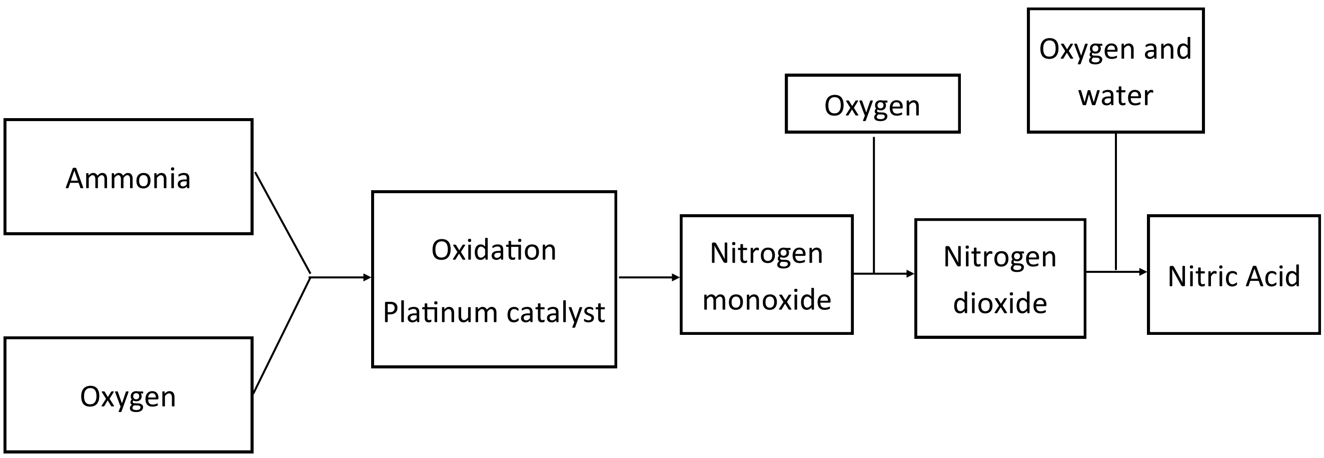
Nitrogen dioxide dissolves in water in the presence of air to make nitric acid.

Word Equation: nitrogen dioxide + oxygen + water $\longrightarrow$ nitric acid

This method for making nitric acid is very expensive due to the large amounts of electricity required to make nitrogen and oxygen react.

Making Nitric Acid - The Ostwald Process

The Ostwald Process is the industrial process for making nitric acid using ammonia and oxygen.



Stage 1. Ammonia reacts with oxygen to form nitrogen monoxide.

This part of the process involves the use of a platinum catalyst along with a moderately high temperature. The platinum catalyst must be heated to start with, after a while the reaction, which is **exothermic**, gives out enough heat to keep the catalyst hot and no further heating is required.

Stage 2. Nitrogen monoxide reacts with more oxygen to form nitrogen dioxide.

Stage 3. Nitrogen dioxide reacts with further oxygen and water to produce nitric acid.

Topic Questions

Exercise 1

1. What is meant by the term unsaturated?

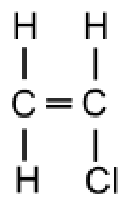
2. a) Draw a structural formula for but-2-ene.

   b) Draw three units of but-2-ene in the shape of a 'H'.

   c) Draw the product formed when these three units join together.

   d) Draw the repeating unit.

3. Vinyl chloride has the structure as shown.

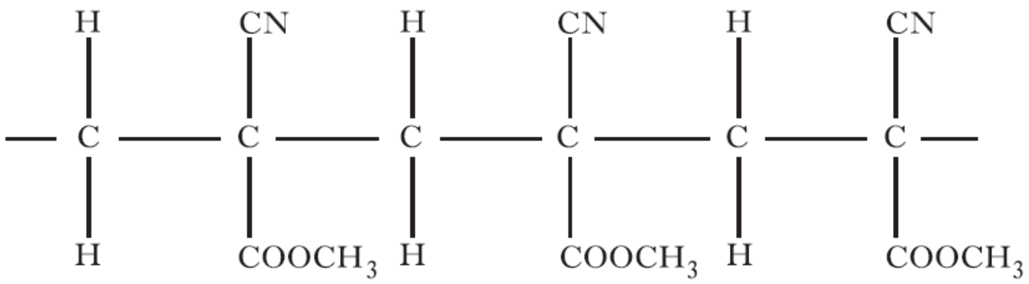


   a) Draw the structure of the polymer chain when three vinyl chloride monomer units react together.

   b) Draw the structure of the repeating unit.

   c) What would be the name of the polymer chain formed from vinyl chloride monomers.

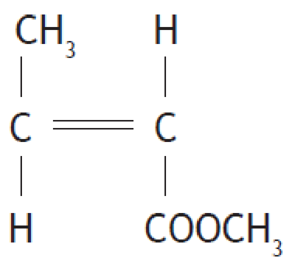
4. When superglue sets, a polymer is formed. Part of the polymer structure is shown.



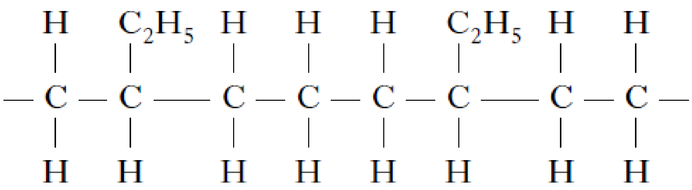
   a) Draw the repeating unit for this polymer.

   b) Draw the monomer for this polymer.

5. Given the following monomer show the part of a polymer chain formed when **three** monomers join together.



6. Part of a polymer chain is shown.



Which two alkenes were used to make this addition polymer?

A Ethene and propene

B Ethene and but-1-ene

C Propene and but-1-ene

D Ethene and but-2-ene

Exercise 2

1. Polymers are made by a polymerisation reaction.

   a) Name the type of polymerisation.

   b) What must be present in the monomer unit to have a successful polymer formed?

2. Draw the following molecules in the shape required to be a monomer ('H')

   a) propene

   b) pent-2-ene

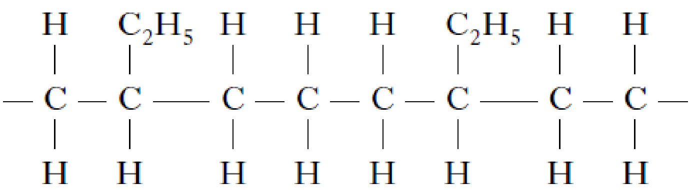
   c) 2-methylprop-1-ene

   d) hex-1-ene

   e) 1,2—dichloroethene

   f) 3—methylbut-1-ene

3. a) Draw the structure of the repeating unit found within the following polymer

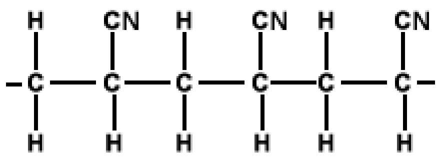


   b) Draw the structure of the two monomers used to create this polymer chain.

4. Draw a polymer chain showing 3 monomers of propene joined together.

5. Name the above polymer.

6. Given the structure of the below polymer, draw the structure of the monomer unit used to form this polymer



Exercise 3

1. Calculate the percentage by mass of potassium in potassium nitrate, KNO_3.
2. Calculate the percentage by mass of nitrogen in calcium nitrate, $Ca(NO_3)_2$.
3. Calculate the percentage by mass of nitrogen in ammonium nitrate, NH_4NO_3.
4. Calculate the percentage by mass of phosphorus in ammonium phosphate, $(NH_4)_3PO_4$.

Exercise 4

1. How is the nitrogen obtained for use in the Haber process?
2. How is the hydrogen obtained for use in the Haber process?
3. Name the catalyst used in the Haber process and state the temperature and pressure of the reaction vessel.
4. Some unreacted nitrogen and hydrogen passes out of the reactor. Explain why this is the case and state what happens to these gases.
5. a) Write a word equation for the formation of ammonia from nitrogen and hydrogen.

   b) Write a balanced chemical equation for this reaction.
6. Draw a diagram to show the shape of an ammonia molecule.

Exercise 5

1. Name the compound produced by the Ostwald process.
2. At what temperature is the Ostwald process carried out?
3. What catalyst is used in the Ostwald process?
4. Nitric acid will react with ammonia.

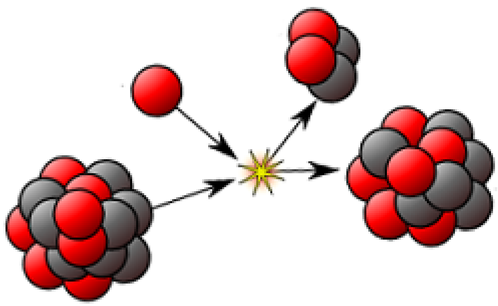
   a) What is the product formed in this reaction?

   b) Write a chemical equation for this reaction.

   c) Suggest a use for the product.

Topic 3

Nuclear Chemistry



Learning Outcomes

Radiation

- Radioactive decay involves changes in the nuclei of atoms.
- Unstable nuclei (radioisotopes) can become more stable nuclei by giving out alpha, beta or gamma radiation.
- Alpha particles (α) consist of two protons and two neutrons and carry a two positive charge. They have a range of only a few centimetres in air and are stopped by a piece of paper. Alpha particles will be attracted towards a negatively charged plate.
- Beta particles (β) are electrons ejected from the nucleus of an atom. They are able to travel over a metre in air but can be stopped by a thin sheet of aluminium. Beta particles will be attracted towards a positively charged plate.
- Gamma rays (γ) are electromagnetic waves emitted from within the nucleus of an atom. They are able to travel great distances in air. They can be stopped by barriers made of materials such as lead or concrete. Gamma rays are not deflected by an electric field.

Nuclear equations

- Balanced nuclear equations can be written using nuclide notation.
- In nuclear equations:

  - an alpha particle can be represented as ${}^{4}_{2}He$

  - a beta particle can be represented as ${}^{0}_{-1}e$

  - a proton can be represented as ${}^{1}_{1}p$

  - a neutron can be represented as ${}^{1}_{0}n$
- In the course of any nuclear reaction:
  - The sum of the atomic numbers on the left of the reaction arrow is equal to the sum of the atomic numbers on the right of the reaction arrow.
  - The sum of the mass numbers on the left of the reaction arrow is equal to the sum of the mass numbers on the right of the reaction arrow. *Candidates do not need to show electrical charges when writing balanced equations representing nuclear reactions.*

Half-life

- Half-life is the time for half of the nuclei of a particular isotope to decay.
- The half-life of an isotope is a constant, unaffected by chemical or physical conditions.
- Radioactive isotopes can be used to date materials.
- The half-life of an isotope can be determined from a graph showing a decay curve.
- Calculations can be performed using the link between the number of half-lives, time and the proportion of a radioisotope remaining.

Use of radioactive isotopes

- Radioisotopes have a range of uses in medicine and in industry. Candidates do not need to be able to name the isotope used in a particular application.
- Given information on the type of radiation emitted and/or half-lives, the suitability of an isotope for a particular application can be evaluated.

Nuclear Radiation

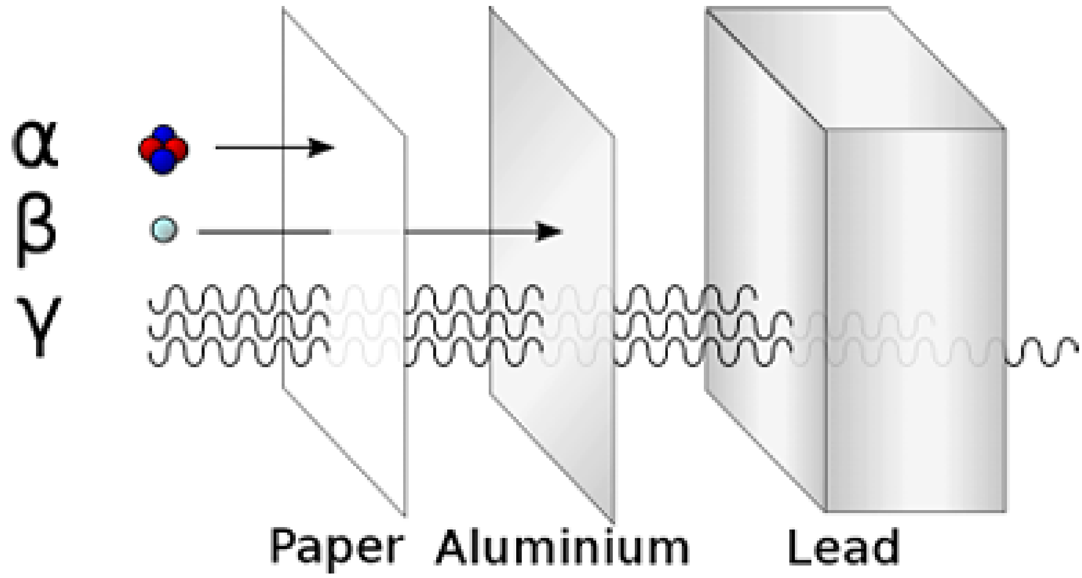
Types of Radiation

Alpha particles consist of 2 protons and 2 neutrons i.e. the nucleus of a helium atom.

Beta particles are fast moving electrons emitted from the nucleus. They are produced when a neutron in the nucleus breaks up to form a proton and electron.

Gamma radiation is a type of electromagnetic wave similar to X-rays but of higher energy.

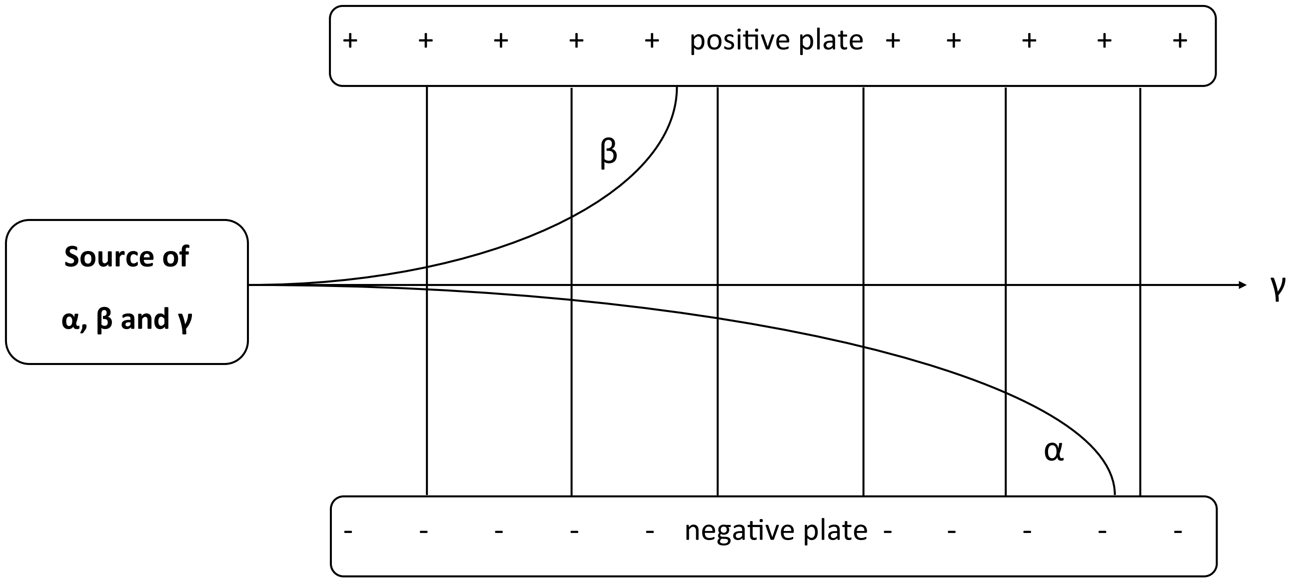
The different types of radiation have different penetrating power. Alpha is the least penetrating as it is stopped by paper or 2-3 cm of air. Beta is the next least penetrating and is stopped by thin aluminium or about 10 cm of air. Gamma radiation is the most penetrating but is stopped by 10 cm thick lead or concrete as summarised in the diagram below.



Name	Nature	Symbol	Charge
Alpha	2 protons 2 neutrons	α^{4}_{2}	2+
Beta	fast moving electron	β^{0}_{-1}	1-
Gamma	electromagnetic waves	γ	no change

Unstable Nuclei

If alpha, beta and gamma radiation are passed through an electric field there is a distinct pathway each will take based on the charges present in each. (Remember like charges repel and opposite attract.)



Nuclear Equations

Nuclear equations are used to summarise nuclear decay. Note that when writing nuclear equations the total mass number of the left hand side is equal to the total mass number on the right hand side. The total atomic number on the left hand side is also equal to the total atomic number on the right hand side.

Alpha Decay

During alpha decay the nucleus emits 2 protons and 2 neutrons (an alpha particle). This means the atomic number of the product is smaller by 2 and the mass number smaller by 4.

Example 1: An americium isotope emits an alpha particle producing a neptunium isotope.

$$^{241}_{95}\text{Am} \rightarrow {}^{237}_{93}\text{Np} + {}^{4}_{2}\alpha$$

Example 2: The alpha decay of radium to form radon.

$$^{222}_{88}\text{Ra} \rightarrow {}^{4}_{2}\text{He} + {}^{218}_{86}\text{Rn}$$

Beta Decay

During beta decay a high energy electron is emitted from the nucleus and the atomic number increases by 1. The mass number remains constant.

Example 1: Beta decay of carbon-14.

$$^{14}_{6}\text{C} \rightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e}^{-}$$

Example 2: Beta decay of iodine-131.

$$^{131}_{53}\text{I} \longrightarrow {}^{0}_{-1}\text{e} + {}^{131}_{54}\text{Xe}$$

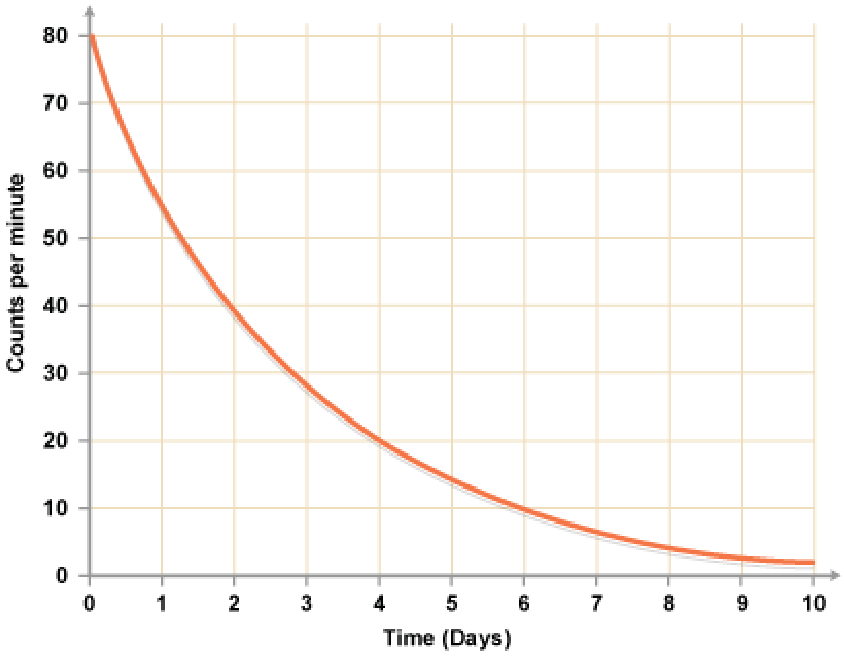
Half-life

When a sample of a radioisotope decays the decay process is random. However when the random events are averaged out the decay of a radioisotope follows a definite pattern.

It is possible to calculate how much of a radioisotope has decayed and how much remains after a given time using the half-life of the isotope.

Half-life is defined as the time it takes for half of the sample to decay or the time it takes for the activity to reduce to half its original value. Some half-life values are a matter of seconds while others are millions of years. The half-life is **not** affected by temperature or concentration unlike normal chemical reactions.

Half-life can be found from a graph of activity against time as shown below.



In this graph the initial activity is 80 counts per minute. After one half life the activity will be reduced to 40 counts per minute. It is therefore seen that the half-life is 2 days. After another 2 days the activity is seen to have reduced by half again.

Half-life can be shown as a percentage of the original sample or as a fraction of the original sample.

After			**1st**		**2nd**		**3rd**	**half-lives**
As a percentage:	100%	⟶	50%	⟶	25%	⟶	12·5%	etc
Or as fractions:	1/1	⟶	1/2	⟶	1/4	⟶	1/8	etc

Half-Life Calculations

Example 1: Lead-210 decays to 6·25% of its original radioactivity after 84 years. What is the half-life of this isotope?

After 1 half-life there would be 50% of the original sample remaining.

After 2 half-lives there would be 25% of the original sample remaining.

After 3 half-lives there would be 12·5% of the original sample remaining.

After 4 half-lives there would be 6·25% of the original sample remaining.

So 84 ÷ 4 = 21 years for each half-life.

Example 2: A radioisotope of phosphorus has a mass of 80 g and a half-life of 14 days. Calculate the mass of the isotope remaining after 56 days.

56 ÷ 14 = 4 half-lives

80 g $\longrightarrow$ 40 g $\longrightarrow$ 20 g $\longrightarrow$ 10 g $\longrightarrow$ 5 g

5 g remains.

Example 3: A radioisotope has a half-life of 7000 years. How long will it take for the 48 g of the radioisotope to decay to leave 6 g?

Find how many half-lives this requires.

48 g $\longrightarrow$ 24 g $\longrightarrow$ 12 g $\longrightarrow$ 6 g

3 half-lives are required.

Total time = 7000 x 3 = 21 000 years.

Half-life

Example 4: Carbon-14 is a naturally occurring radioisotope. It has a half-life of 5730 years and is found in the remains of living things and is therefore used to date objects.

A sample of wood was found to have 12·5% of the carbon-14 content of living wood.

Calculate the age of the wood.

100% $\longrightarrow$ 50% $\longrightarrow$ 25% $\longrightarrow$ 12·5%

0 years $\longrightarrow$ 5730 years $\longrightarrow$ 11 460 years $\longrightarrow$ 17 190 years

Uses of Radioisotopes

Radioisotopes are widely used in medicine and industry. They are often manufactured by bombarding stable isotopes with neutrons in a nuclear reactor. Since neutrons have no charge they are not repelled by the positive nucleus. Some radioisotopes are used in medical applications such as treatment of cancer cells. Radioisotopes can be injected into the body to help destroy the cancerous cells which cause a cancer tumour.

Radioisotopes are found in household smoke detectors. The radioisotope causes the smoke particles to ionise creating a small electrical current which then sets of an alarm.

The radioisotope chosen for a particular use must be determined by its half life. For example, within smoke detectors an isotope with a long half life is required. On the other hand, for medical purposes a radioisotope with a short half life would be much better suited and have less damaging effects on the human body.

Topic Questions

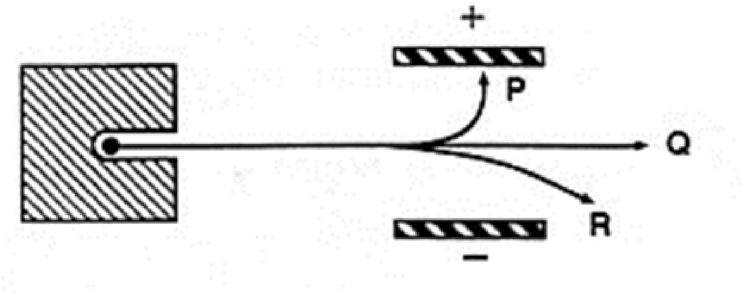
1. Write a nuclear equation to show the alpha decay of each of the following isotopes.

a) ${}^{13}_{6}C$ b) ${}^{238}_{92}U$ c) ${}^{23}_{11}Na$ d) ${}^{16}_{8}O$ e) ${}^{144}_{54}Xe$

2. Write a nuclear equation to show the beta decay of each of the following isotopes.

a) ${}^{131}_{53}I$ b) ${}^{22}_{10}Ne$ c) ${}^{31}_{15}P$ d) ${}^{27}_{13}Al$ e) ${}^{90}_{38}Sr$

3. Name each type of radiation, P , Q and R in the following diagram as each is passed through an electric field.



4. ${}^{24}Na$ is a radioactive isotope of sodium with a half life of 15 hours. A sample of ${}^{24}Na$ has a mass of 200g. What will be the remaining mass after 120 hours?
5. A sample of a radioisotope decays to 1/8th of its original activity over a period of 12 days. What is the half life of the radioisotope?
6. A radioisotope has a half-life of 14 hours. How long will it take to decay to 6·25% of its original activity?
7. A radioisotope with a half life of 9 seconds has a count rate of 64 counts per minute at the start of timing. What will the count rate be 18 seconds later?
8. A sample of a radioisotope decays to 1/32 of its original activity in 12·5 years. What is the half life of the isotope?
9. A radioisotope has a half-life of 160 days. A 1g sample of the pure radioisotope is allowed to decay for 640 days. What mass of the original isotope will remain after that time?
10. ${}^{257}Lr$ has a half life of 8 seconds. Over what period of time would the activity of a sample of this radioisotope decay from 120 to 15 counts per minute?
11. ${}^{90}Sr$ has a half life of 27 years. What fraction of the sample of this radioisotope would remain after 108 years?
12. What fraction of its original activity will a sample of a radioisotope with a half life of three minutes have after decaying for 15 minutes?