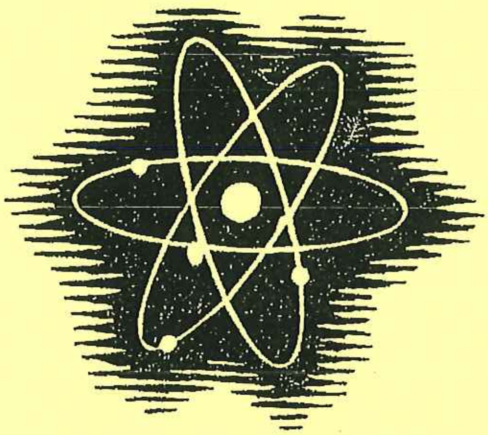
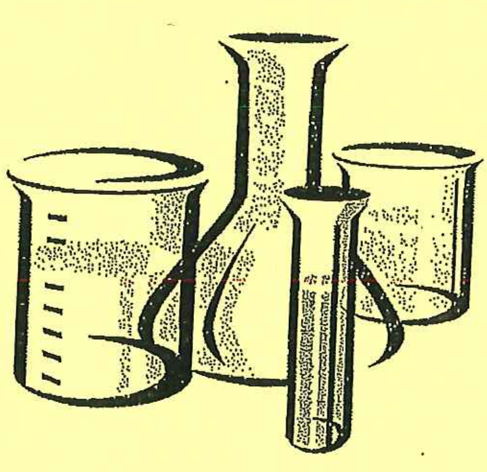
NAME:



N4/N5
CHEMISTRY IN SOCIETY

METALS



WHAT I SHOULD KNOW : NATIONAL FOUR

• Certain metals react with **water, oxygen** & dilute **acid** • Some metals react with water & dilute acid to produce **hydrogen gas** • Explain how to **test** for hydrogen gas • Metals can be placed into an **order of reactivity** by comparing their reactions with water, oxygen and dilute acid	**The properties of metals and alloys** **Determination of the reactivity series using reactions of metals**
• Use the reactivity series to investigate other metal reactions *e.g. to show how one reactive metal atom can displace a less reactive metal ion solution* • A metal can be extracted from its ore by **heating with carbon** • Show by experiment how to extract a metal from its ore using **electrolysis** • The **method** used to extract a metal depends on its **position** in the reactivity series	**Displacement reactions.** **Extraction of metals related to their reactivity.**
• Corrosion is the result of a chemical reaction on **surface** of a metal • Identify **factors** that affect corrosion from experimental data • The **rate** of corrosion of metals depends on their position in the reactivity series • **Rust** is the corrosion of iron • Ferroxyl indicator turns **blue** if rust is present • Examples of **physical protection** are paint, oil, grease, plastic, layer of another metal • Rusting is prevented when iron is connected to another **more reactive** metal	**Corrosion** **Physical and chemical protection of metals.**
• Construct a **simple cell** using pairs of different metals • Investigate the effect of different **pairs of metals** on the voltage of a cell • Compare voltages produced when different pairs of metals are connected in a cell and construct an **electrochemical series** • Use the electrochemical series to **predict direction** of the flow of current when different pairs of metals are connected in a cell • Demonstrate **electroplating** of one metal by another • Find and present information on about a variety of batteries • Describe how electroplating can be used to **protect** metals from corrosion	**Electrochemical series and electrochemical cells.** **Voltage and electroplating.**
• An alloy is a **mixture** of metals, or metals and non-metals • Analyse data about alloys and give some **examples** of alloys and their **uses** • State alloys have **different** physical properties compared to their constituent elements	**Composition, uses and physical properties of alloys.**

WHAT I SHOULD KNOW : NATIONAL FIVE

• Explain how **metallic bonding** allows metals to conduct electricity • Construct **balanced ionic equations** when metals react with water, oxygen & acids • oxidation is **loss** of electrons while reduction is the **gain** of electrons • A **redox** reaction is when oxidation and reduction occur together • identify ion-electron equations as either reduction or oxidation reactions • construct an **electro-chemical cell** using different metals or using a non-metal electrode and explain why an electrolyte is essential • use the electrochemical series to predict the direction of current when different metals are connected in a cell • construct **ion-electron equations** based on electro-chemical cells • describe different types of batteries and show by experiment that a lead-acid battery can be recharged • research technological advances in the construction of batteries e.g. fuel cells, and find out how redox reactions are utilised in them • explain why the extraction of a metal from its ore is termed a reduction reaction • identify the **reducing agent** in the extraction of metals from balanced ionic equations • show by calculation how to work out the **percentage by mass** of a metal in an ore	Metals Metallic bonding & resulting electrical conductivity. Balanced ionic equations for metal reactions, metal extractions & reduction reactions. Electrochemical cells including a non -metal electrode. Reactions of metals — electrons flow, redox reaction, oxidation, reduction. Fuel cells and rechargeable batteries.

Metals

Over _____ % of the elements in the Periodic Table are metals.

Physical Properties of Metals

These vary depending on the metal. The physical properties of a metal decide the use to which it is put.

In general, metals are:

- Denser than non-metals
- Good conductors of heat and electricity in both the solid and liquid state
- Malleable – can be beaten into different shapes
- Strong
- Ductile – can be drawn out into wires

Physical Property	Uses	Example
strength		
hardness		
lightness and strength		
electrical conduction		
thermal conduction (heat)		
malleability		

Metals as Finite Resources

We get all our metals from the earth's crust. Just like fossil fuels (oil, coal and gas) our supply of metals are ____________.

There will come a time when they will run out.

Metal	Estimated Reserves (millions of tonnes)
Copper	300
Lead	125
Zinc	110
Tin	15

Metal	Used per Year (millions of tonnes)
Copper	15
Lead	8
Zinc	10
Tin	5

Plot this information as a bar graph.

Tackling the Problem

The rate at which we use up reserves of metals could be greatly reduced by **recycling** i.e. the metals we throw away could be re-used.

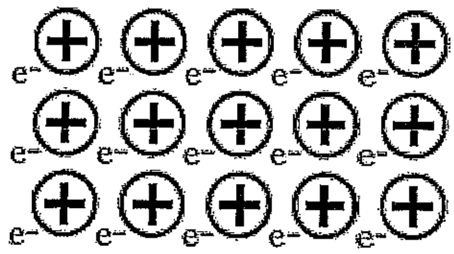
Recycling has two advantages:

1. ____

2. ____

Metallic Bonding National 5

In a metal atom the outer electrons are delocalised (free to move). Metallic bonding is the electrostatic force of attraction between the positively charged metal ions (cores), and the negatively charged delocalised electrons.



Alloys

An alloy is a mixture of metals or of metals with non-metals

Alloys are made by melting the elements involved, mixing them together and allowing to cool.

Properties of Alloys

The physical properties of alloys are different to those of the element from which it was made.

Examples of alloys and their properties:-

Alloy	Elements it is made from	Properties	Use
Steel		less brittle than iron	Construction industry, cars, ships bridges
Stainless steel		resists corrosion	Knives, forks, spoons, machinery for food industry
Solder		melts easily - good conductor	Making electrical connections in circuit boards
Brass		soft and easily melted	Ornaments, door handles, letter boxes
Bronze		harder than copper or tin	Statues, cutting edges

The properties of a particular alloy vary with the percentage of each element present in it. E.g. the percentage of carbon in steel alters the steels strength.

Reactions of Metals with Oxygen

Aim To investigate the reaction of metals with oxygen

Method

Results

Metal	What changes did you see?	Order of reactivity
		Most reactive
		Least reactive

GENERAL EQUATION:

Metal + Oxygen → Metal oxide

$Ca + O_2 \rightarrow CaO$

$4Ca + O_2 \rightarrow 2CaO$

Write the word and (balanced) chemical equations for the reactions of magnesium and iron with oxygen (assume that iron (II) oxide is formed).

Reactions of Metals with Water

Aim To investigate the reaction of metals with water

Method **Teacher Demo First**

Results

Metal	What changes did you see?	Order of reactivity
		Most reactive
		Least reactive

GENERAL EQUATION:

Metal + Water → Metal hydroxide + Hydrogen

$K + H_2O \rightarrow KOH + H_2$

$2K + 2H_2O \rightarrow 2KOH + H_2$

Write the word and balanced chemical equations for the reactions of calcium and sodium with water.

Reactions of Metals with Dilute Acid

Aim To investigate the reaction of metals with dilute acid

Method

Results

Metal	What changes did you see?	Order of reactivity
		Most reactive Least reactive

GENERAL EQUATION:

Metal + Acid → A Salt + Hydrogen

$Al + (H^+)_3PO_4^{3-} \rightarrow AlPO_4 + H_2$

$2Al + 2(H^+)_3PO_4^{3-} \rightarrow 2AlPO_4 + 3H_2$

Write the word and balanced chemical equations for the reactions of zinc and iron with hydrochloric acid (HCl). Assume zinc(II)chloride and iron(II) chloride are formed.

Reactions of Metals

<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="14">Most reactive

Least reactive</td><td rowspan="10">Metals which react with</td><td rowspan="5">Metals which react with water</td><td rowspan="10">Metals which react with</td></tr>
<tr><td>Sodium</td></tr>
<tr><td>Lithium</td></tr>
<tr><td>Calcium</td></tr>
<tr><td>Magnesium</td></tr>
<tr><td>Aluminium</td><td rowspan="9">No reaction</td></tr>
<tr><td>Zinc</td></tr>
<tr><td>Iron</td></tr>
<tr><td>Tin</td></tr>
<tr><td>Lead</td></tr>
<tr><td>Copper</td><td rowspan="4">No reaction</td><td rowspan="4">No reaction</td></tr>
<tr><td>Mercury</td></tr>
<tr><td>Silver</td></tr>
<tr><td>Gold</td></tr>
</table>

Metal Displacement Reactions

Experiment 1

Aim To see what happens when zinc powder is added to a solution of copper (II) sulphate

Method

Result

Reason The zinc atoms have lost electrons and turned into zinc ions which go into the solution.

Zn	$\rightarrow$	Zn^{2+}	+	$2e^-$
atoms		ions		electrons

The copper (II) ions take up the electrons and turn into copper atoms (copper metal).

Cu^{2+}	+	$2e^-$	$\rightarrow$	Cu
ions		electrons		atoms

These are **ion-electron** equations.

Experiment 2

<u>Aim</u> To investigate the order of displacement of metals

<u>Method</u>

<u>Results</u>

Metal	Added to a solution of			
	$MgSO_4$	$ZnSO_4$	$FeSO_4$	$CuSO_4$
Magnesium				
Zinc				
Iron				
Copper				

<u>Conclusions</u>

The metal which is higher in the reactivity series will displace a metal lower in the series from its compounds in solution.

Example:

magnesium + copper sulphate $\longrightarrow$ magnesium sulphate + copper

$$Mg(s) + Cu^{2+}SO_4^{2-}(aq) \longrightarrow Mg^{2+}SO_4^{2-}(aq) + Cu(s)$$

The spectator ion in the above equation is the SO_4^{2-}ion. We can remove the solid copper by the technique **filtration**.

- (N4) Write the word and chemical equations for all the displacement reactions in experiment 2.
- (N5) Write the word and **ionic** equations for all the displacement reactions in experiment 2:

Extraction of Metals

Most metals are obtained from rocks in the earth's crust. These metals are in the rocks as compounds, mostly *metal oxides, metal sulphides or metal carbonates.*

Rocks that contain naturally occurring metal compounds are called **ores**

Only a few elements are found uncombined in the earth's crust, these are the **most unreactive metals**, at the bottom of the reactivity series such as *mercury, silver and gold.*

More reactive metals, i.e. metals above mercury in the reactivity series have to be **extracted** from their ores.

There various extraction methods depend on how strongly the metal is bonded in the compound. This depends on **how reactive** the metal is.

Extracting a metal from its ore is usually done in several stages:

- Crushing the ore (so that the metal compound can be separated from all the other compounds in the rocks)
- Roasting (if the metal compound is not an oxide, the ore is roasted in air to change it into an oxide)
- Extraction (of the metal from the metal oxide)

<u>Obtaining Metals from Metal Oxides</u>

1 *<u>Using Heat Alone</u>*

This only works for silver (I) oxide and mercury (II) oxide

$$\text{silver (I) oxide} \xrightarrow{\text{heat}} \text{silver} + \text{oxygen}$$

$$2Ag_2O \rightarrow 4Ag + O_2$$

Why does heat alone work?
The silver/oxygen bonds and the mercury/oxygen bonds <u>are weak</u>, since neither metal is very reactive. This means that the metal/oxygen bonds can <u>easily be broken by heat alone</u>.

2 *<u>Using Heat and Another Substance</u>*

a) Heat and Carbon

$$\text{metal oxide} + \text{carbon} \xrightarrow{\text{heat}} \text{metal} + \text{carbon dioxide}$$

b) Heat and Carbon Monoxide

$$\text{metal oxide} + \text{carbon monoxide} \xrightarrow{\text{heat}} \text{metal} + \text{carbon dioxide}$$

c) Heat and Hydrogen

$$\text{metal oxide} + \text{hydrogen} \xrightarrow{\text{heat}} \text{metal} + \text{water}$$

These methods works for all metal oxides **below**, and including **zinc** oxide.

Why Is Heat <u>AND</u> Another Substance Needed?

The metals involved (Zn, Fe, Sn, Cu) are **<u>more reactive</u>** than Hg and Ag. This means that they <u>form **stronger** bonds</u> with oxygen and so **<u>more energy</u>** <u>is required to break the bonds.</u>

3 <u>Using Electrolysis</u>

All metal oxides above zinc oxide require more energy to remove the metal.

$$\text{metal oxide (\underline{\textit{melt}})} \xrightarrow{\text{electric current (d.c.)}} \begin{cases} \text{metal (-ve electrode)} \\ \text{oxygen (+ve electrode)} \end{cases}$$

Why is Electrolysis Needed?

The metals involved (K, Na, Li, Ca, Mg, Al) are **<u>very reactive</u>** and the "extra" energy that they need to separate them from their ores is provided by electricity.

General Conclusion

As reactivity of metal increases, so does the strength of the metal-oxygen bond and the more energy is required to break the bond

Note

When we obtain metals from metal oxides we say that we have **<u>reduced</u>** the metal oxides.

This is because each of the methods used resulted in the metal ions in the oxide <u>gaining electrons</u> to become metal atoms, i.e. they have been <u>reduced.</u>

$$M^{2+} + 2e^- \rightarrow M$$

Corrosion N4

<u>Corrosion</u> is a chemical reaction in which the surface of a metal reacts with its surroundings and changes from an element into a compound.

Not all metals corrode at the same rate. The more reactive a metal, the quicker it corrodes.

- K, Na and Li tarnish almost immediately when exposed to air. This is a sign of corrosion. This is why these metals are stored under oil
- Ca and Mg corrode quickly, but not as quickly as the group 1 metals
- Cu takes years to corrode
- Au does not corrode at all

<u>Rusting</u> is the term applied to the corrosion of ____________ ONLY.

Chemically, rust is _____________ (____________)

Requirements for rusting

1.

2.

Rusting N4

In the **first stage of rusting** the iron atoms at the surface react to form iron ions

$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$$

This is an **oxidation** reaction

We can show that Fe^{2+} ions are present by using **ferroxyl indicator** (sometimes called rust indicator).

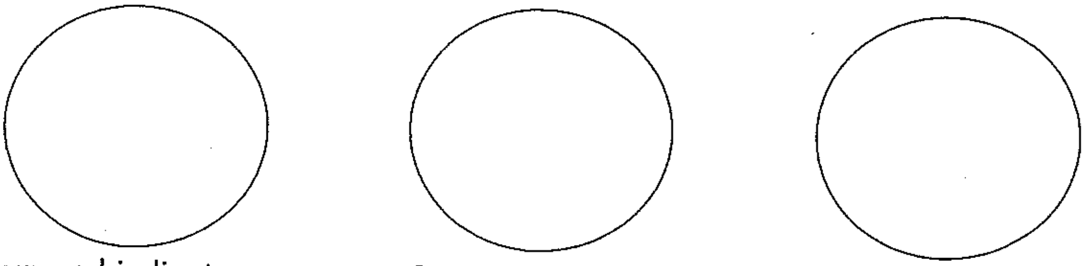
The indicator changes from ____________ to ________________ when Fe^{2+} is present.

Speed of Rusting N4

The rusting of iron may be affected by the presence of some solutions.

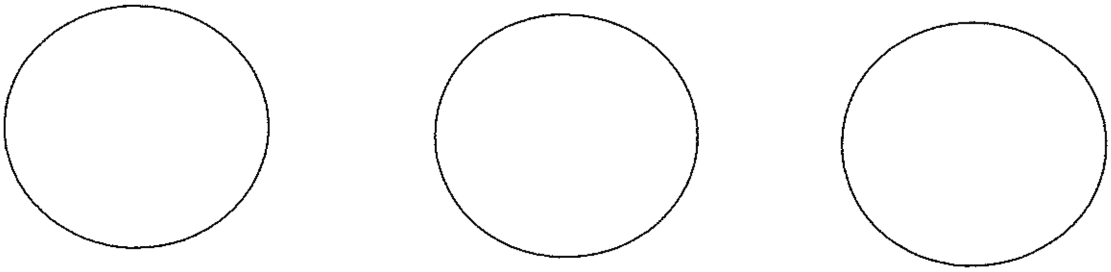
EXPERIMENT

The following plates were set up and left for several days.



Ferroxyl indicator ONLY | ferroxyl indicator + ACID | ferroxyl indicator + SALT solution

RESULTS



CONCLUSION

Rusting takes place **FASTER** in solutions that contain **IONIC** bonding. Salt solution and acid are both ionic solutions and so cause rusting to happen faster.

This means that in areas with lots of acid rain pollution or where the road is made safer in winter by 'gritting' the speed of rusting will be faster for any iron structures, eg cars, bridges, lamp posts etc.

Protection Against Corrosion N4

Two basic methods

-Barrier (physical) methods
-Electron flow methods

Barrier Methods of Preventing Corrosion

The metal can be:

1.
2.
3.
4.
5.
6.

This prevents corrosion by preventing oxygen and water from contacting the metal.

When the zinc coating is scratched, because zinc is above iron in the electrochemical series, the electron flow is from zinc to iron, giving sacrificial protection.

In contrast, if the tin coating is scratched, since iron is above tin in the electrochemical series, the electron flow is **from iron** to tin. This promotes the change

$$Fe(s) \qquad\qquad Fe^{2+}(aq) \ \ +2e^-$$

So the rate of rusting increases.

Electron Flow Protection (2 types) N4

1. Sacrificial Protection

Result

Reason

Magnesium is higher in the electrochemical series than iron
Therefore the magnesium gives up its electrons much more readily than iron i.e. the magnesium will oxidise and its electrons will flow to the iron. This prevents the iron corroding and is known as **sacrificial protection**

Conclusion

The corrosion rate of any metal is prevented by attaching it to a metal higher up the electrochemical series

2. Cathodic Protection

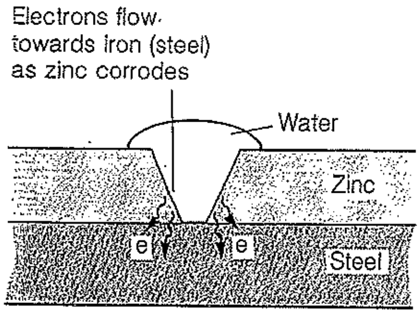
Result and Reason

The nail attached to the **negative terminal** (cathode) of the battery had electrons supplied to it and this **prevented corrosion.**

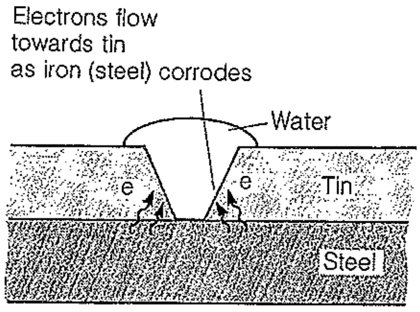
The nail attached to the positive terminal (anode) had electrons pulled away from it, this encouraged oxidation and therefore increased corrosion.

Comparison of Tin plating and Galvanising N4

When the coatings are intact, both give physical protection. But things are different once the coatings are scratched.



Scratched zinc plating



Scratched tin plating

When the zinc coating is scratched, because zinc is above iron in the electrochemical series, the electron flow is from zinc to iron, giving sacrificial protection.

In contrast, if the tin coating is scratched, since iron is above tin in the electrochemical series, the electron flow is **from iron** to tin. This promotes the change

$$Fe(s) \longrightarrow Fe^{2+}(aq) \quad +2e^-$$

So the rate of rusting increases.

Showing the Redox nature of Rusting

Rusting is an example of oxidation. For example, when iron rusts, the iron atoms lose electrons to form ions.

$$Fe_{(s)} \longrightarrow Fe^{2+}{}_{(aq)} + 2e^-$$

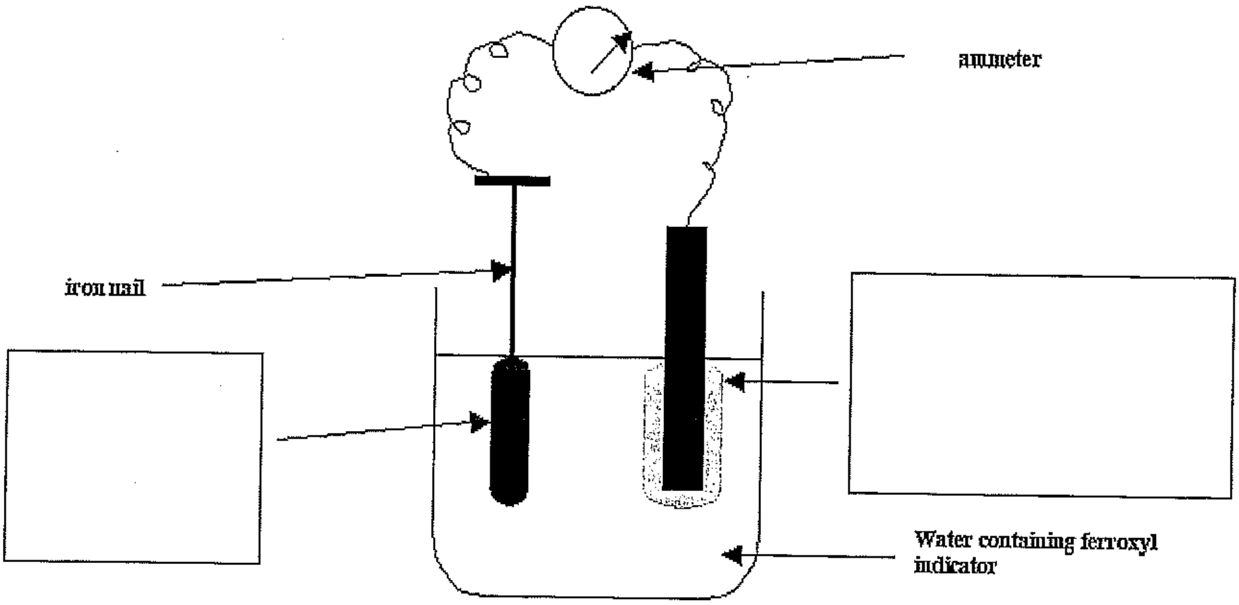
The electrons lost in this oxidation process are accepted by oxygen and water to form hydroxide ions. The corresponding reduction equation is :

$$O_{2(g)} + 2H_2O_{(l)} + 4e^- \longrightarrow 4OH^-{}_{(aq)}$$

This helps explain why oxygen and water are needed for the rusting of iron.

This Redox reaction can be illustrated by setting up the electrochemical cell shown below with ferroxyl indicator in the electrolyte.

Ferroxyl indicator also contains a second indicator which turns pink in the presence of OH^- ions.



The meter indicates that electrons are flowing from the iron to the carbon electrode. Rusting can therefore be thought of as an electrochemical cell where water, oxygen and carbon dioxide or another electrolyte are required.

<u>**Batteries and Cells**</u>

Everyday use of the word "battery" is often wrong. Most batteries should really be called cells. A true battery contains several cells joined together.

Answer the following questions:

1. Name three things that use batteries to supply electricity

   ____

2. How is electricity produced in a battery?

   ____

3. What is meant by the term a "dead battery"?

   ____

4. Give an example of a rechargeable battery

   ____

5. What are the advantages of using batteries?

   ____

   ____

6. What are the disadvantages of using batteries?

   ____

   ____

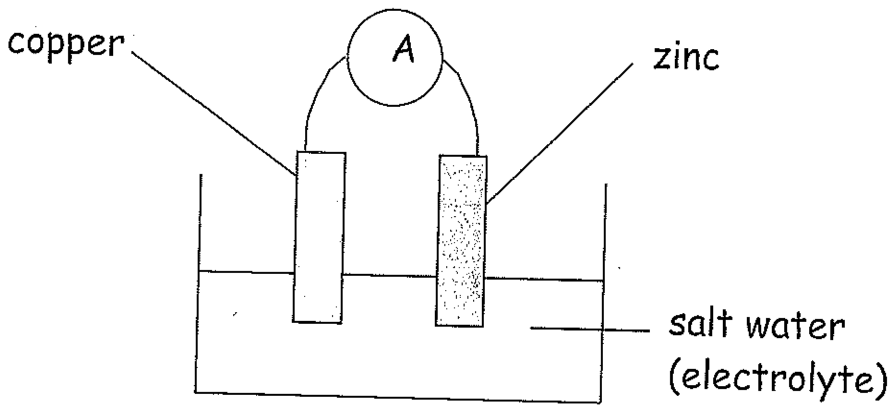
The Electrochemical Series

A **cell** is a device that contains substances that react together to provide a flow of electrons through a wire.

The flow of electrons is known as an electric current and can be detected using an ammeter or a bulb.

When 2 metals are connected by an ionic solution (electrolyte) a current will flow. In many batteries, the electrolyte is ammonium chloride.

This is an example of a cell:



Using 2 Metals to Make a Cell

Aim To observe the voltage produced when 2 different metals are connected

Method

Results

Metal A	Meter Reading in Volts

List the metals in order of voltage reading, starting with the highest

Conclusion

The order is the same as the reactivity series, this order is called the **electrochemical series.**

In a cell, the electrons flow from the metal higher in the series to the metal lower in the series.

The further apart the metals are, the higher the voltage produced.

More About Cells

Aim: To produce electricity by connecting two metals in solutions of their metal ions.

Method: The following apparatus was set up

Result:

The circuit can be completed by using one of the following ion bridges:

- A folded filter paper soaked in salt water
- A U-tube containing potassium chloride solution or gel containing potassium chloride
- String soaked in salt water

This is an alternative apparatus where the beakers are replaced by the two limbs of the U-tube and the porous plug acts as the ion bridge.

The movement of ions in an ion bridge provides ions to complete the circuit.

Oxidation and Reduction N5

The loss of electrons is called OXIDATION.

The gain of electrons is called REDUCTION.

OILRIG

Ion/Electron equations

We can represent the loss and gain of electrons by writing a chemical equation. (They are sometimes called half ion equations because each one only tells half of the story. Electrons are not lost by something unless there is something else ready to take them, so oxidation can't take place without reduction.)

Oxidation

When an atom loses electrons it will form a positive ion

$$X \rightarrow X^+$$

This means that the electron will have to go to the RHS to keep the equation balanced.

$$X \rightarrow X^+ + e^-$$

Reduction

When an atom gains electrons it becomes a negative ion.

$$X \rightarrow X^-$$

This means that the electron will have to go to the LHS to keep the equation balanced.

$$X + e^- \rightarrow X^-$$

Oxidation and Reduction N5

Oxidation is the loss of electrons
Reduction is the gain of electrons

Oxidation
Is
Loss
Reduction
Is
Gain

e.g. $Cu \rightarrow Cu^{2+} + 2e^-$ is an example of oxidation
$I_2 + 2e^- \rightarrow 2I^-$ is an example of reduction

Page 7 of the data booklet gives the ion-electron equations for a series of **reduction reactions**. To write an oxidation reaction, you need to reverse the reduction reaction equation.

Write the ion-electron equations for the following reactions and state whether they are oxidations or reductions:

Potassium ions becoming potassium atoms ____________________

Gold atoms becoming gold ions ____________________

Sulphite ions becoming sulphate ions ____________________

Chorine become chloride ions ____________________

When oxidation and reduction both occur in a reaction, the reaction is called a REDOX reaction

Oxidation and Reduction Worksheet N5

A. State which of these changes are: oxidation, reduction, both, neither

1. copper (II) oxide to copper metal ________________
2. zinc to zinc (II) chloride ________________
3. bromine atoms gaining electrons to form bromide ions ________________
4. iodine molecules forming iodide ions ________________
5. aluminium metal from aluminium oxide ________________
6. silver (I) carbonate decomposing to silver metal ________________
7. hydrogen gas being released from an acid ________________
8. zinc displacing copper (II) ions from solution to form copper and zinc (II) ions ________________
9. magnesium reacting with hydrochloric acid to form magnesium chloride and hydrogen ________________
10. copper (II) carbonate reacting with nitric acid to form copper (I) nitrate, carbon dioxide and water ________________

B. Complete the ion-electron equations below by putting the electrons in the correct place and state which are oxidation and which are reduction.

1. $Br_2 \rightarrow 2\ Br^-$ ________________
2. $Fe^{2+} \rightarrow Fe^{3+}$ ________________
3. $Mg \rightarrow Mg^{2+}$ ________________
4. $Cu^{2+} \rightarrow Cu$ ________________
5. $2I^- \rightarrow I_2$ ________________
6. $2H^+ \rightarrow H_2$ ________________

Some ion-electron equations are more complex.

For example $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$

To work out if this, or any, type of change is an oxidation or a reduction, simply look at the side of the equation that the electrons are on.

For an oxidation reaction, the electrons are on the RIGHT

For a reduction reaction, the electrons are on the LEFT

So the example above shows ____________________

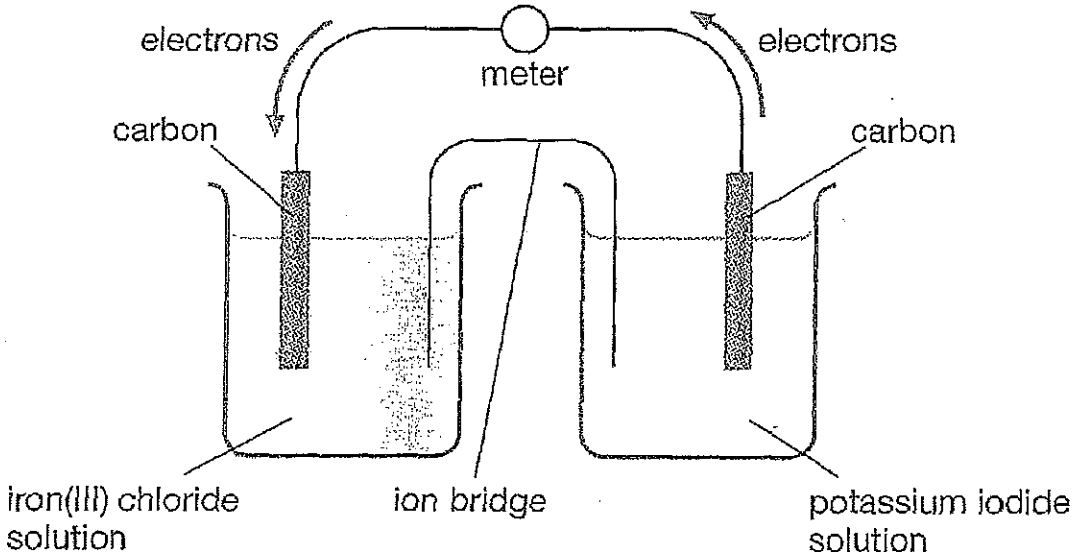
What type of change is shown in the example below?

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ ____________________

Cells involving Non Metals N5

Electricity can be produced in a cell in which ***one of the half cells does not involve a metal.***

For example, in the cell shown below, the two half cells are a solution of iodide ions and a solution of iron (III) ions.



Electrons flow from the iodide ions to the iron (III) ions. We can tell this since electrons ***are produced*** by this **oxidation** reaction.

$$2I^{-} \longrightarrow I_2 + 2e^{-}$$

The electrons pass through the wires and the meter to the iron (III) chloride side of the cell. There the iron ions ***gain electrons*** forming iron (II) ions.

$$Fe^{3+} + e^{-} \longrightarrow Fe^{2+}$$

This is the **reduction** step.

To form the overall redox reaction, the ion-electron equations for the oxidation and reduction must be combined, ensuring that number of electrons in the oxidation cancels out with the number of electrons in the reduction.

Example 1:
The reaction of magnesium with dilute sulphuric acid

oxidation	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$
reduction	$2H^+(aq) + 2e^- \rightarrow H_2(g)$
redox reaction	$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

Example 2:
The reaction of aluminium with dilute hydrochloric acid

oxidation	$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
reduction	$2H^+(aq) + 2e^- \rightarrow H_2(g)$

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

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

$$6H^+(aq) + 6e^- \rightarrow 3H_2(g)$$

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

Example 3:
The reaction of silver nitrate solution with sodium sulphite solution

oxidation	$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$
reduction	$Ag^+(aq) + e^- \rightarrow Ag(s)$

To cancel out the electrons, the reduction must be multiplied by 2.

$$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

$$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$$

$$2Ag^+(aq) + SO_3^{2-}(aq) + H_2O(l) \rightarrow 2Ag(s) + SO_4^{2-}(aq) + 2H^+(aq)$$

Fuel Cells National 5

A fuel cell is an electrochemical cell that produces electricity by combining a fuel such as hydrogen with oxygen without burning it.

The use of fuel cells is increasing. They represent a clean technology which will help industries and governments meet CO_2 emission reduction targets.

Rechargeable Batteries National 5

A flat battery is a battery in which one of the chemicals has been used up. Some batteries are **rechargeable**.

Here electrical energy is converted to chemical energy. One example of a rechargeable battery is a car battery also known as the **lead-acid battery**.