C4 – Chemical Changes

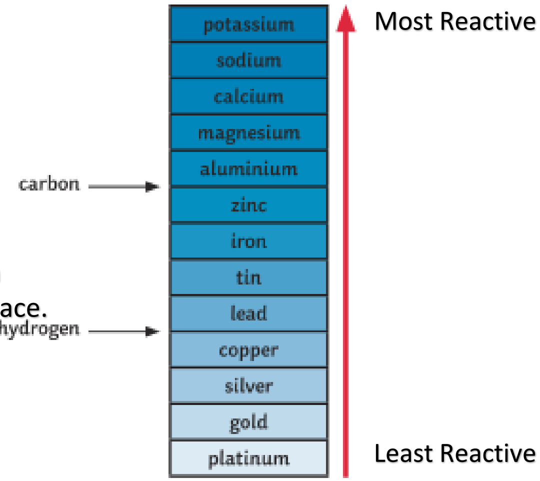
The Reactivity Series

- A more reactive metal will replace a less reactive metal in a compound (**displacement**)

- e.g. potassium + magnesium chloride → potassium chloride + magnesium

Potassium is more reactive than magnesium

Potassium **displaces** magnesium from the compound and takes it's place.



Reactions of acids with metals

- Metal + acid → salt + hydrogen

E.g. iron + sulfuric acid → iron sulfate + hydrogen

salt

To name salt :
1st name Metal
2nd name Acid used

Naming Salts

Acid used	Salt produced
Hydrochloric	Chloride
Sulfuric	Sulfate
Nitric	Nitrate

Extraction of Metals

- Extraction = remove metal from an ore or a compound.

Ore = a rock containing enough metal to make extracting metal worthwhile.

How to extract metals:

Less reactive than carbon – reduction with carbon

Reduction = loss of oxygen

E.g. iron oxide + carbon → iron + carbon dioxide

Oxygen has been removed to extract iron.

Carbon and the oxygen removed from the iron react to make carbon dioxide

More reactive than carbon – electrolysis is used.

- Some metals are found in **native** form (not reacted, so in element form) – usually platinum and gold as **very unreactive**.

Reaction of metals with oxygen

- Metal + oxygen → metal oxide

e.g magnesium + oxygen → magnesium oxide

$2Mg + O_2 \rightarrow 2MgO$

Oxidation reaction as metal gained oxygen

- Oxidation = gaining oxygen
- Reduction = losing oxygen

Reaction of metals with water

- Most metals don't react well with water
- Group 1 and group 2 react to form alkalis

- Metal + water → metal hydroxide + hydrogen

e.g lithium + water → lithium hydroxide + hydrogen

$2Li + 2H_2O \rightarrow 2LiOH + H_2O$

Metal hydroxides are alkaline

Reactions of acids with alkalis

- Acid + alkali → salt + water

neutralisation

Hydrochloric acid + sodium hydroxide → sodium chloride + water

salt

$\mathbf{HCl + NaOH \rightarrow NaCl + H_2O}$

Reactions of acids with carbonates

- Acid + carbonate → salt + water + carbon dioxide

sulfuric acid + calcium carbonate → calcium sulfate + water + carbon dioxide

salt

$\mathbf{H_2SO_4 + CaCO_3 \rightarrow CaSO_4 + H_2O + CO_2}$

C4 – Chemical Changes

1. What is meant by displacement?
2. Name a very reactive metal
3. Name two metals which are less reactive than hydrogen.

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1. Define extraction.
2. What is an ore?
3. How do you extract a metal less reactive than carbon?
4. What is meant by reduction?
5. What is meant by a 'native metal'?
6. Give an example of a metal found in native form.

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1. State the general equation for the reaction of metal with oxygen.
2. Write a word equation for the reaction of iron with oxygen.

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1. State the general equation for the reaction of metal with water.
2. Are hydroxides acid/alkaline?

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1. State the general equation for the reaction of metal with acid.
2. State the salts produced from hydrochloric acid, sulfuric acid and nitric acid.

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1. State the general equation for the reaction of acid with an alkali.

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1. State the general equation for the reaction of acid with carbonates.

C4 – Chemical Changes

Redox Reactions (HT only)

- Redox = reduction and oxidation takes place at same time in a reaction.

- Metal + acid = redox reaction

Example

$H_2SO_4 + Ca \rightarrow CaSO_4 + H_2$

Ionic equation: $2H^+ + Ca \rightarrow Ca^{2+} + H_2$

Half equation 1: $Ca \rightarrow Ca^{2+} + 2e^-$ ← Lost 2 electrons (oxidation)

Half equation 2: $2H^+ + 2e^- \rightarrow H_2$ ← Gained 2 electrons (reduction)

Strong/Weak Acids (HT only)

Strong acid = completely dissociates in a solution

e.g. $HCl \rightarrow H^+ + Cl^-$

Examples = nitric acid and sulfuric acid

Weak acid = partially dissociates in solution.

e.g. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$\rightleftharpoons$ = reversible reaction ← Hasn't fully turned into ions – only partially

Concentration = how much is dissolved in every cm^3

Strong/weak = how well it ionises

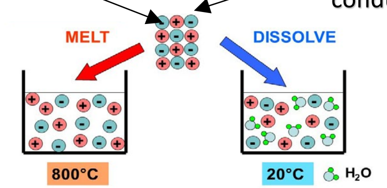
As **pH** decreases by 1 unit, **hydrogen ion concentration** of solution increases by factor of 10

Electrolysis

- **Splitting** up a **compound** using electricity.
- Used to extract metals from compounds**,** purify metals (eg copper)

Ionic compound

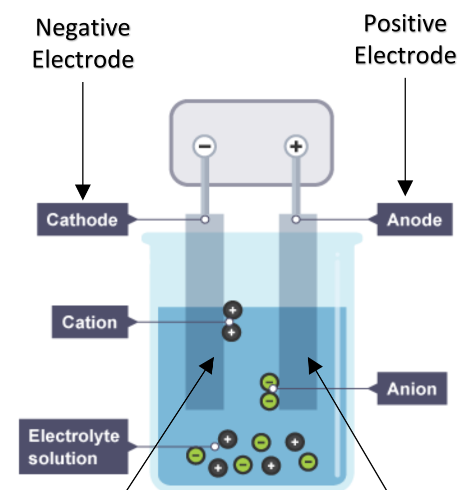
Solid – ions can't move to conduct electricity



- Must be **molten** or **aqueous** (dissolved in water) to allow **ions** to **move** to the electrodes

The Process of Electrolysis

Two **electrodes –** made of **inert** material (doesn't react)

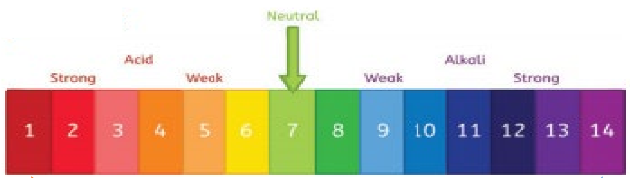


Positive ions move to the negative electrode (opposites attract)

Negative ions move to the positive electrode (opposites attract)

pH Scale

- Shows how acidic or alkaline solution is.
- pH 1-6 = acid
- pH 7 = neutral
- pH 8-14 = alkali



Lots of H^+ | Less H^+ | Less OH^- | Lots of OH^-

In aqueous solutions:

Acids – produce H^+ ions

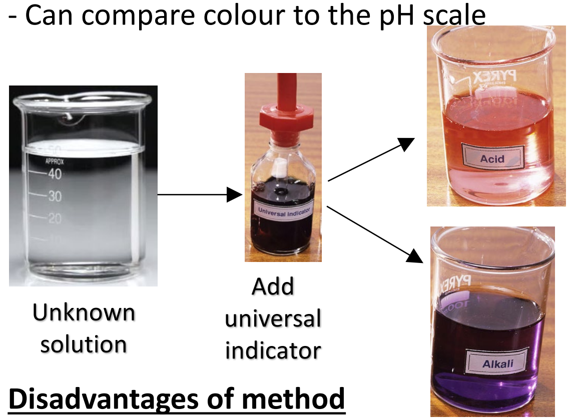
Alkalis – produce OH^- ions

In neutralisation reactions:

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Measuring pH of a solution

- Can use **universal indicator**
- Gives the solution a colour
- Can compare colour to the pH scale



Changes colour to give pH value

Disadvantages of method

- Colour is **subjective** – different people may see different colours
- Doesn't give an exact pH number (could use **pH probe** to make more **accurate**).

C4 – Chemical Changes

1. What is a redox reaction?
2. In terms of electrons, what does oxidation mean?
3. In terms of electrons, what does reduction mean?

1. What is the pH range for an acid?
2. What is the pH range for an alkali?
3. If a substance has a pH of 7, what type of substance is it?
4. What ions do acids produce in solution?
5. What ions do alkalis produce in a solution?
6. State the ionic equation for neutralisation reactions.

1. Define a strong acid.
2. Give an example of a strong acid.
3. Define a weak acid.
4. What happens to H^+ concentration as the pH value decreases by 1?

1. Describe a simple method to test the pH of an unknown solution.
2. State 2 disadvantages of using universal indicator.
3. How can pH be measured more accurately?

1. What is meant by the term electrolysis?
2. What is electrolysis used for?
3. What must the compound be for electrolysis to take place?
4. Why can solid ionic compounds not conduct electricity?
5. What does inert mean?
6. Name the positive electrode.
7. Name the negative electrode.
8. Why do positive ions move to the negative electrode?

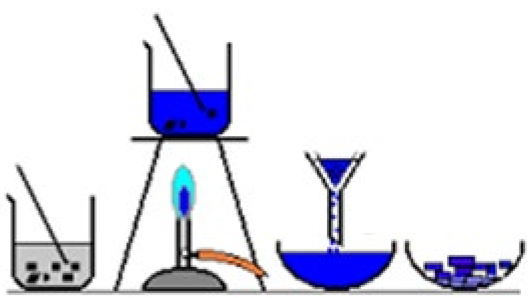
C4 – Chemical Changes – Required Practical – Preparation of soluble salts

Aim
Prepare a pure, dry sample of a soluble salt from an insoluble **oxide or carbonate.**

Equipment
- Beaker
- Measuring cylinder
- Bunsen burner and safety mat
- Filter funnel and filter paper
- Named acid (e.g. hydrochloric acid)
- Metal oxide or carbonate.
- Spatula
- Glass stirring rod

Change method depending on reactants in the question.

Method (example copper oxide and sulfuric acid to make copper sulfate)
1. Using measuring cylinder – 20cm^3 **sulfuric acid** → beaker
2. Warm the acid gently (not boiling)
3. Using spatula add **copper oxide** to the acid and stir
4. Keep adding until no more oxide will dissolve (excess).
5. Using a filter funnel and filter paper – filter excess copper oxide.
6. Evaporate some of the filtrate using a water bath.
7. Pour remaining filtrate into an evaporating basin – leave overnight to evaporate water
8. Pat the crystals dry.



Common questions

Q1) Why do you heat the acid before adding the oxide?

A1) To speed up the reaction (particles have more energy to react).

Q2) Why is the oxide added in excess?

A2) To make sure that all the acid has been neutralised.

Q3) Why is the solution filtered?

A3) Remove any unreacted, excess solid.

Q4) Why is the solution left overnight in a warm, dry place?

A4) To evaporate excess water, to form crystals (crystallise).

Q5) Name 2 safety precautions you should take during this practical.

A5) Safety goggles and allow equipment to cool before putting away

C4 – Chemical Changes – Required Practical – Preparation of soluble salts

1. Write a method to prepare a pure, **dry** sample of copper sulfate crystals (6 marks).

Q2) Why do you heat the acid before adding the oxide?

Q3) Why is the oxide added in excess?

Q4) Why is the solution filtered?

Q5) Why is the solution left overnight in a warm, dry place?

Q6) Name 2 safety precautions you should take during this practical.

C4 – Chemical Changes

Electrolysis of Molten Ionic Compounds

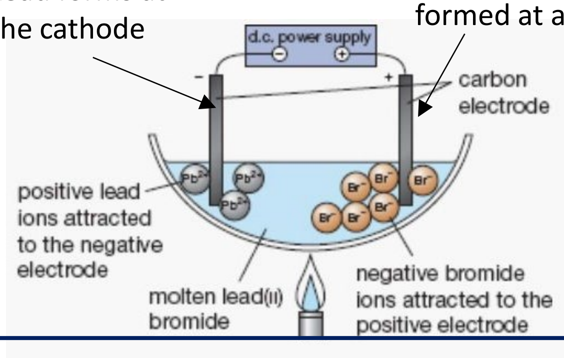
Molten = melted **so ions** can move.

- Metal = produced at **anode**
- Non-metal = produced at **cathode**

Example: Lead Bromide - $PbBr_2$

Lead forms at the cathode

Bromine gas is formed at anode



Using Electrolysis to Extract Metals

- Used if metal is **too reactive** to be extracted by reduction with carbon.
- Requires **large amount of energy** to melt the compound and produce electrical current. **(expensive)**

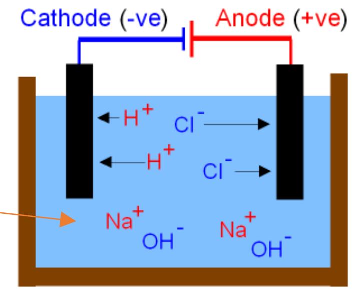
Example: Aluminium Oxide

- **Cryolite** is added – reduces the melting point (less energy needed – less expensive)
- **Carbon** used as positive electrode – needs to be replaced constantly as **oxygen** will react with it to produce CO_2 – it will degrade.

Electrolysis of Aqueous Solutions

- Compound is dissolved in water so ions can move.

When aqueous – H^+ and OH^- (from H_2O) are also present along with the two ions from the compound.



- Only **one** ion is discharged at each electrode.

Anode – Non-metal or oxygen

Cathode – Metal or hydrogen

Rules

+ ANODE Attracts – ions ('Anions')	**- CATHODE** Attracts + ions ('Cations')
If – ions are group 7 i.e. **chloride Cl^-** **bromide Br^-** **iodide I^-** Then the groups 7 element is produced as a gas	If + ions (metals) are MORE REACTIVE than hydrogen **K, Na, Ca, Mg, Zn, Fe** Then **HYDROGEN** is produced
If – ions are NOT Group 7 Eg sulphate SO_4^{2-} nitrate NO_3^- carbonate CO_3^{2-} OXYGEN is produced.	If + ions (metals) are LESS REACTIVE than hydrogen **Cu, Ag, Au** Then the METAL is produced

Examples

Solution	Product at cathode	Product at anode
Potassium chloride	Hydrogen – because K is more reactive than H	Chlorine – as it is a halogen
Copper sulfate	Copper – as copper is less reactive than H	Oxygen – as there is no halogen

Half-Equations at Electrodes (HT only)

During electrolysis:

Cathode – positive ions **gain** electrons (**reduction**)

Anode – negative ions **lose** electrons (**oxidation**)

- Ions become **discharged** (lose their charge) at the electrodes to form the atoms again.

- Reactions at electrodes can be represented by half equations.

Examples

Cathode - $2H^+ + 2e^- \rightarrow H_2$

Gained 2 electrons (reduction)

molecules of hydrogen gas produced

Anode – $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

molecules of oxygen produced

Lost electrons (oxidation)

Cathode – $Cu^{2+} + 2e^- \rightarrow Cu$

Gained electrons (reduction)

Copper atoms are formed at the cathode

Anode – $2Cl^- \rightarrow Cl_2 + 2e^-$

chlorine molecules are formed

Lost electrons (oxidation)

C4 – Chemical Changes

1. Why is an ionic compound melted before electrolysis takes place?
2. Metals are produced at the..
3. Non-metals are produced at the..

1. When is electrolysis used to extract a metal?
2. Why is electrolysis expensive?
3. Why is cryolite added to aluminium oxide before electrolysis?
4. Why does the positive anode need constantly replacing when electrolysing aluminium oxide?

1. Why is the compound dissolved in water before electrolysing?
2. What two ions are also present in aqueous solutions (along with the compound)?
3. Which two substances can be produced at the anode?
4. Which two substances can be produced at the cathode?
5. When would a metal be produced at the cathode?
6. When would oxygen be produced at the anode?

1. In terms of electrons, what happens at the positive electrode?
2. In terms of electrons, what happens at the negative electrode?
3. Write the half equation for the production of hydrogen.
4. Write the half equation for the production of oxygen from hydroxide ions.
5. Write the half equation for the production of copper from copper ions.
6. Write the half equation for the production of chlorine from chloride ions.

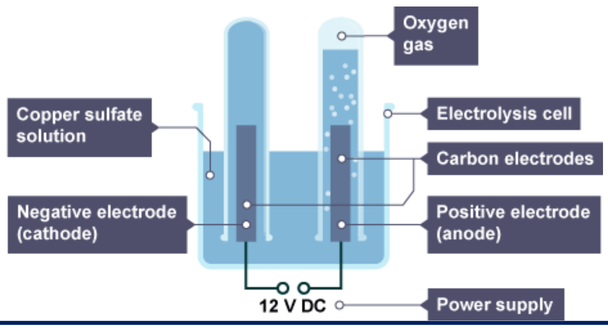
C4 – Chemical Changes – Required Practical – Electrolysis of Aqueous Solutions

Aim
To investigate the electrolysis of an aqueous solution using inert (unreactive) **electrodes**.

Equipment
- Beaker
- Two test tubes (or measuring cylinders)
- Graphite electrodes
- Two splints
- Aqueous solution
- DC powerpack

Change method depending on the question.

Method (example copper sulfate solution.)
1. Pour some copper sulfate solution into a beaker.
2. Place two graphite rods into the copper sulfate solution. Attach one electrode to the negative terminal of a dc supply, and the other electrode to the positive terminal.
3. Completely fill two small test tubes with copper sulfate solution and position a test tube over each electrode as shown in the diagram. **(use measuring cylinders if measuring volume of gas produced)**
4. Turn on the power supply and observe what happens at each electrode.
5. Test any gas produced with a glowing splint and a burning splint.
6. Record observations and the results of your tests.



Common questions

Q1) How do you test for hydrogen gas?

A1) Lit splint will make a squeaky pop.

Q2) How do you test for oxygen gas?

A2) Glowing splint – will relight.

Q3) Explain why copper is produced at the cathode.

A3) Copper ions are **positive**, so are attracted to the negative electrode (opposites attract). Copper is less reactive than hydrogen so is discharged. The copper ions **gain electrons** and are **reduced** to form **copper atoms**.

Q4) Why do hydrogen ions move to the cathode?

A4) Hydrogen ions are **positive** so move to the negative electrode as **opposites attract**.

Q5) Why are measuring cylinders better to collect the gas?

A5) Because they are more accurate when measuring the volume of gas produced.

C4 – Chemical Changes – Required Practical – Electrolysis of Aqueous Solutions

Q1. Draw a labelled diagram to show the equipment needed to electrolyse copper chloride.

Q2. Write a method for the electrolysis of aqueous copper chloride solution.

Q2) How do you test for hydrogen gas?

Q3) How do you test for oxygen gas?

Q4) Explain why copper is produced at the cathode.

Q5) Why do hydrogen ions move to the cathode?

Q6) Why are measuring cylinders better to collect the gas?