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GCSE Chemistry

AQA Topic – Chemical changes

Overview

Chemical changes

Reactivity of metals

- Metal oxides
- The reactivity series
- Extraction of metals by reduction
- Oxidation and reduction in terms of electrons (HT)

Reactions of acids

- Metals and acids
- Strong and weak acids (HT)
- pH scale and Neutralisation
- Salts
- *Titration (chemistry only)*

Electrolysis

- Electrolysis of molten ionic compounds
- Electrolysis of an aqueous solution
- Using electrolysis to extract metals
- Representation at electrodes as half equations (HT)

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Reactivity of metals

- Metals oxides
- The reactivity series
- Extracting metals by reduction
- Oxidation and reduction in terms of electrons (HT)

Metal Oxides

A **metal compound** within a **rock** is an **ore.** Ores are **mined** and then **purified**.

Whether it is worth extracting a particular metal depends on:

- **How easy it is to extract it from its ore**
- **How much metal the ore contains**
- **The changing demands for a particular metal**

Most **metals** in ores are **chemically bonded** to **other elements** in compounds. Many of these metals have been **oxidised** (have oxygen added) by oxygen in the air to form their oxides.

Iron + oxygen → iron (III) oxide

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

To extract metals from their oxides, the metal oxides must be **reduced** (have oxygen removed).

Reactivity series

Metals can be arranged in order of reactivity in a **reactivity series**.

Order of reactivity	Reaction with water	Reaction with acid
Potassium	Fizz, giving off hydrogen gas and leaving an alkaline solution of metal hydroxide	Reacts violently and explodes
Sodium		
Lithium		
Calcium		Fizz, giving off hydrogen gas and forming a salt
Magnesium	Very slow reaction	
Aluminium		
Zinc		
Iron		
Tin	No reaction with water at room temperature	React slowly with warm acid
Lead		
Copper	No reaction	No reaction
Silver		
Gold		

Reactivity series

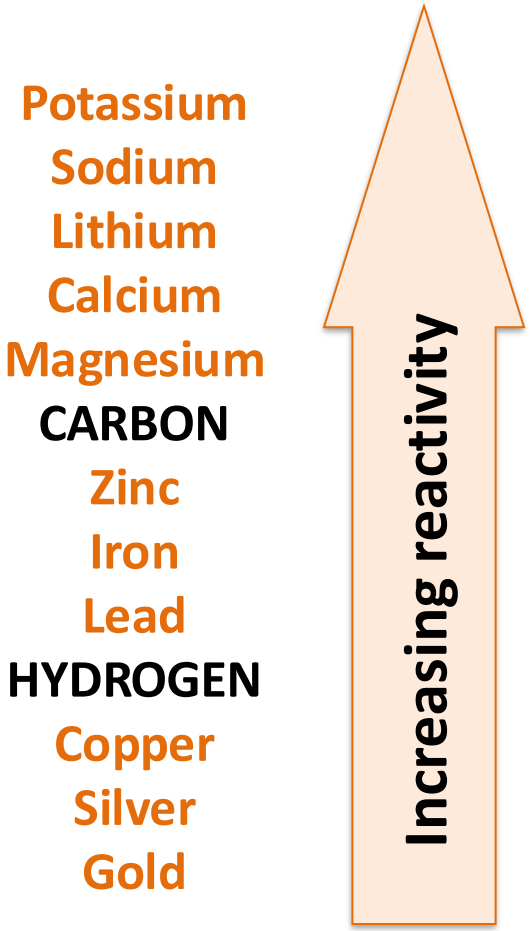
Metals can be arranged in order of reactivity in a **reactivity series**.

When metals react with other substances the metal atoms form **positive ions**.

The reactivity of a metal is linked to its tendency to form positive ions.

The **non-metals hydrogen** and **carbon** are often included in the series as they can be used to extract less reactive metals.

Metal + acid → salt + hydrogen



Extracting metals

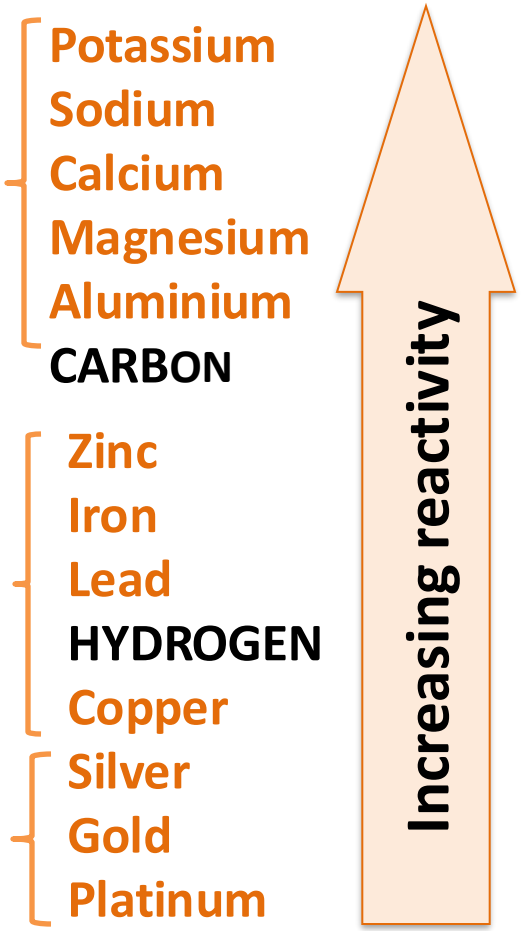
The **reactivity** of a metal determines the **method** of **extraction.**

Metals **above** carbon must be extracted from their ores by using **electrolysis.**

Metals **below** carbon can be extracted from their ores by **reduction** using **carbon.**
REDUCTION involves the loss of oxygen.

metal oxide + carbon → metal + carbon dioxide

Gold and **silver** do not need to be extracted.
They occur **native** (naturally).



Oxidation and reduction in terms of electrons (HT)

A **more reactive metal can displace a less reactive metal from its compound** in **displacement reactions**.

Iron + copper(II) sulfate → iron sulfate + copper

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

Higher:

OILRIG

Oxidation **I**s **L**oss of electrons
Reduction **I**s **G**ain of electrons

When reactions involve **oxidation and reduction**, they are known as **redox reactions**

Higher:

An **ionic equation** shows only the atoms and ions that change in a reaction:

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Half equations show what happens to each reactant:

$Fe \rightarrow Fe^{2+} + 2e^-$ *The iron atoms are **oxidised** (lose 2 electrons) to form **ions**.*

$Cu^{2+} + 2e^- \rightarrow Cu$ *The 2 **electrons** from the iron are **gained** (**reduction**) by copper ions as they become **atoms**.*

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QuestionIT!

Extracting metals

- The reactivity series
- Metals oxides
- Extracting metals by reduction

Extracting metals – QuestionIT

1. What is an ore?
2. What is produced when metals react with oxygen?
3. What is this process called and why?
4. What is reduction in terms of oxygen?
5. What type of ions do metals produce?
6. Which is more reactive potassium or iron?
7. Which two non-metals can be included in the reactivity series?

8. Why is gold found as an element in the Earth?

9. How are metals, less reactive than carbon, extracted from their ores?

10. HT: Describe oxidation in terms of electrons.

11. HT: Describe reduction in terms of electrons.

12. Write the word equation for the reaction between lithium and water.

AnswerIT!

Extracting metals

- The reactivity series
- Metals oxides
- Extracting metals by reduction

1. What is an ore?
   Metal compound in a rock.
2. What is produced when metals react with oxygen?
   Metal oxide.
3. What is this process called and why?
   Oxidation, gain of oxygen.
4. What is reduction in terms of oxygen?
   Loss of oxygen.
5. What type of ions do metals produce?
   Positive.
6. Which is more reactive potassium or iron?
   Potassium.
7. Which two non-metals can be included in the reactivity series?
   Carbon and hydrogen.

8. Why is gold found as an element in the Earth?
   Unreactive metal.
9. How are metals, less reactive than carbon, extracted from their ores?
   Reduction with carbon.
10. HT: Describe oxidation in terms of electrons.
    OIL – loss of electrons
11. HT: Describe reduction in terms of electrons.
    RIG – gain of electrons
12. Write the word equation for the reaction between lithium and water.
    lithium + water → lithium hydroxide + hydrogen

8. Write the balanced symbol equation for the reaction between lithium and water.
   $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

9. HT: Zinc can be extracted from zinc oxide by heating it with carbon in the blast furnace. Carbon monoxide is also produced. Which reactant is:
   a) Oxidised? **Carbon**
   b) Reduced? **Zinc oxide**

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Reactions of acids
PART 1

- Acids and metals
- Neutralisation of acids
- Soluble salts

Reactions of acids - PART 1

Acids react with some **metals** to produce **salts** and **hydrogen**.

Metal + acid → salt + hydrogen

Reactions between metals and acids only occur if the metal is **more reactive** than the **hydrogen** in the acid. If the metal is too reactive, the reaction with acid is **violent**.

The **salt** that is made depends on the **metal** and **acid** used.

Salts made when metals react *nitric acid* are called *nitrates*.

Zinc + Nitric acid → Zinc Nitrate + Hydrogen

Salts made when metals react with *sulfuric acids* are called *sulfates*.

Iron + Sulfuric Acid → Iron Sulfate + Hydrogen

Salts made when metals react with *hydrochloric* acid are called *chlorides*.

Magnesium + Hydrochloric acid → Magnesium Chloride + Hydrogen

Reactions of acids - PART 1 - HIGHER

In the reaction between magnesium and hydrochloric acid, the hydrogen ions are **displaced** from the solution by magnesium as the magnesium is more reactive than hydrogen.

The following **ionic equation** occurs:

$$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

The chloride ions are not included as they do not change in the reaction. These are known as **spectator ions**.

The reaction can be further represented by **half equations**, showing that the reaction between a metal and acid is a **redox reaction**.

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

The magnesium atoms lose two electrons, they have been **oxidised**.

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

The hydrogen ions have gained electrons, they have been **reduced**.

Reactions of acids - PART 1

Acids are neutralised by **alkalis** (**eg: soluble metal hydroxides**) and **bases** (**eg: insoluble metal hydroxides and metal oxides**) to produce **salts** and **water** and by **metal carbonates** to produce **salts**, **water** and **carbon dioxide**.

The salt **name** depends on the **acid** used and the **positive ions** in the **alkali, base or carbonate**.

Making Soluble Salts from acids and alkalis

Salts can be made by reacting an acid with an alkali.

Acid + Alkali → Salt + Water

Making Soluble Salts from acids and bases

Salts can be made by reacting an acid with a insoluble base.

Acid + Bases → Salt + Water

Making Soluble Salts from acids and metal carbonates

Salts can be made by reacting an acid with a metal carbonate.

Acid + Metal carbonate → Salt + Water + Carbon dioxide

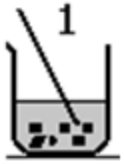
Salts are made of positive metal ions (or ammonia ions - NH_4^+) and a negative ion from the acid. Like all ionic compounds, salts have **no overall charge**, so once you know the charges on the ions, you can work out the **formula**.

Example: **magnesium sulfate** is **$MgSO_4$**

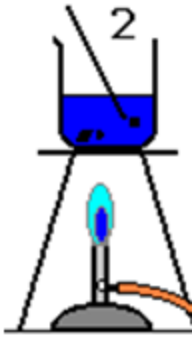
ion	formula	ion	formula
Group 1	Li^+ Na^+ K^+	Transition metals	Cu^{2+} Fe^{3+}
Group 2	Mg^{2+} Ca^{2+}	Group 7	F^- Cl^- Br^-
Aluminium	Al^{3+}	Nitrate	NO_3^-
Ammonium	NH_4^+	Sulphate	SO_4^{2-}

Reactions of acids - PART 1

Soluble salts can be made from **acids** by reacting them with **solid insoluble substances**, such as **metals, metal oxides, hydroxides or carbonates**. The solid is added to the acid until no more reacts and the excess solid is **filtered** off to produce a solution of the salt. Salt solutions can be **crystallised** to produce **solid** salts. **You will complete this as a required practical.**

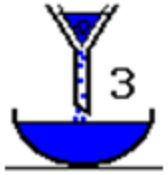


1. Measure the required volume of acid with a measuring cylinder and add the weighed solid (insoluble metal, oxide, hydroxide or carbonate) in small portions with stirring.



2. Safety goggles required - the mixture may be heated to speed up the reaction. When no more of the solid dissolves it means ALL the acid is neutralised and there should be a little excess solid. You should see a residue of the solid (oxide, hydroxide, carbonate) left at the bottom of the beaker.

3. Filter the solution to remove the excess solid metal/oxide/carbonate, into an evaporating dish. On filtration, only a solution of the salt is left.



4. Then hot concentrated solution is left to cool and crystallise.
After **crystallisation**, you collect and dry the crystals with a filter paper.
If the solution is heated, the solvent will evaporate faster. Heating a solution until all the solvent has evaporated is known as **heating to dryness**.

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QuestionIT!

Reactions of acids
PART 1

- Acids and metals
- Neutralisation of acids
- Soluble salts

Reactions of acids – QuestionIT

1. What is produced when acids react with metals?

2. HT: What is a redox reaction?

3. What is produced when an acid reacts with a carbonate?

4. What salt is produced by the following acids?
   a) Hydrochloric acid
   b) Sulfuric acid
   c) Nitric acid

5. How are soluble salts made from acids and insoluble substances?

6. Name the process of producing solid salts from salt solution.

7. Complete the following equations:
   a) magnesium + sulfuric acid →
   b) sodium hydroxide + hydrochloric acid →
   c) lithium carbonate + nitric acid →

8. Write a balanced symbol equation, with state symbols, for the reaction between zinc and hydrochloric acid.

9. **Higher**:
   Write an ionic equation, with state symbols, to show magnesium reacting with hydrochloric acid.

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AnswerIT!

Reactions of acids PART 1

- Acids and metals
- Neutralisation of acids
- Soluble salts

1. What is produced when acids react with metals?
   Salt + water.
2. HT: What is a redox reaction?
   Oxidation and reduction.
3. What is produced when an acid reacts with a carbonate?
   Salt + water + carbon dioxide.
4. What salt is produced by the following acids?
   a) Hydrochloric acid **Chloride**
   b) Sulfuric acid **Sulfate**
   c) Nitric acid **Nitrate**
5. How are soluble salts made from acids and insoluble substances?
   Solid added to acid until no more reacts; excess solid filtered off.
6. Name the process of producing solid salts from salt solution.
   Crystallisation.

7. Complete the following equations:
   a) magnesium + sulfuric acid → **magnesium sulfate + hydrogen**
   b) sodium hydroxide + hydrochloric acid → **sodium chloride + water**
   c) lithium carbonate + nitric acid → **lithium nitrate + water + carbon dioxide**

8. Write a balanced symbol equation, with state symbols, for the reaction between zinc and hydrochloric acid.
   $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

9. **Higher**:
   Write an ionic equation, with state symbols, to show magnesium reacting with hydrochloric acid.
   $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

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Reactions of acids
PART 2

- pH and neutralisation
- Titration (chemistry only)
- Strong and weak acids (HT)

Reactions of acids - PART 2

Indicators are substances which change **colour** when you add them to acids and alkali.

Litmus goes red in acid and blue in alkali.

Universal indicator, made from many dyes is used to tell you **pH**. The scale runs from 0 (most acidic) to 14 (most alkaline). Aqueous solutions of **acids** have a pH value **less than 7**, and for **alkalis greater than 7** and anything in the middle is **neutral** (pH 7). You can use a pH meter to record the change of a pH over time.

Acids produce **hydrogen ions (H^+)** in aqueous solutions and **alkalis** produce **hydroxide ions (OH^-).** In **neutralisation** reactions between an acid and alkali, hydrogen ions react with hydroxide ions to produce **water**.

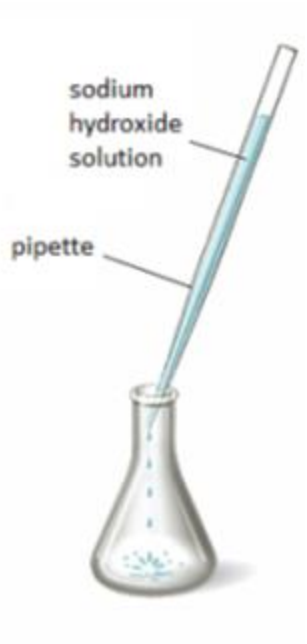
Neutralisation symbol equation:

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

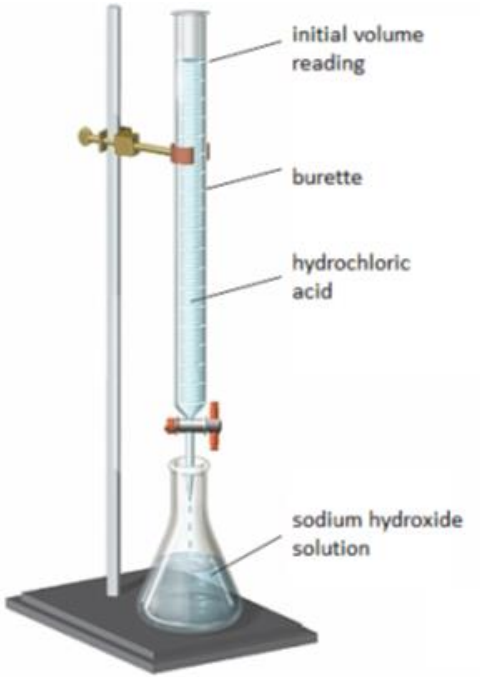
pH	Examples of solutions
0	Battery acid, strong hydrofluoric acid
1	Hydrochloric acid secreted by stomach lining
2	Lemon juice, gastric acid, vinegar
3	Grapefruit juice, orange juice, soda
4	Tomato juice, acid rain
5	Soft drinking water, black coffee
6	Urine, saliva
7	"Pure" water
8	Sea water
9	Baking soda
10	Great Salt Lake, milk of magnesia
11	Ammonia solution
12	Soapy water
13	Bleach, oven cleaner
14	Liquid drain cleaner

Reactions of acids - PART 2 - CHEMISTRY ONLY

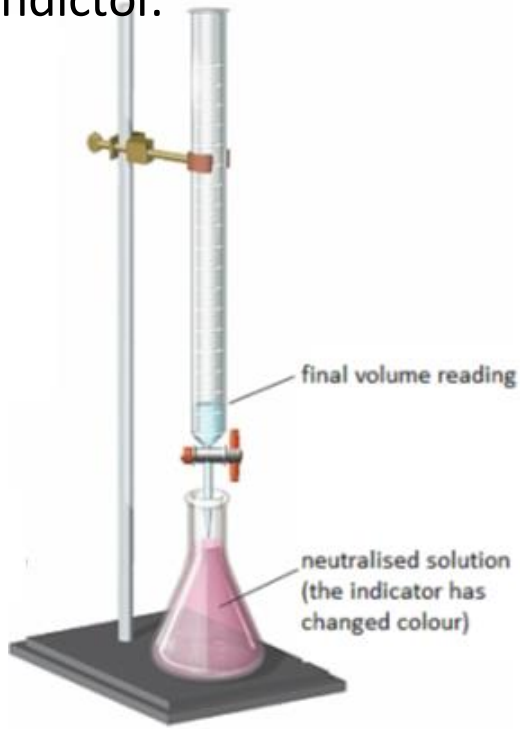
The **volumes** of acid and alkali solutions that react with each other can be measured by **titration** using a suitable indictor.



1. Use the pipette to add 25 cm^3 of alkali to a conical flask and add a few drops of indicator.



2. Fill the burette with acid and note the starting volume. Slowly add the acid from the burette to the alkali in the conical flask, swirling to mix.



3. Stop adding the acid when the end-point is reached (the appropriate colour change in the indicator happens). Note the final volume reading. Repeat steps 1 to 3 until you get consistent readings.

Reactions of acids – PART 2 – CHEMISTRY ONLY
Higher

The **concentration** of a **solution** is the amount of **solute per volume of solution**. Chemists measure concentration in moles per cubic decimetre **(mol/dm³)**.

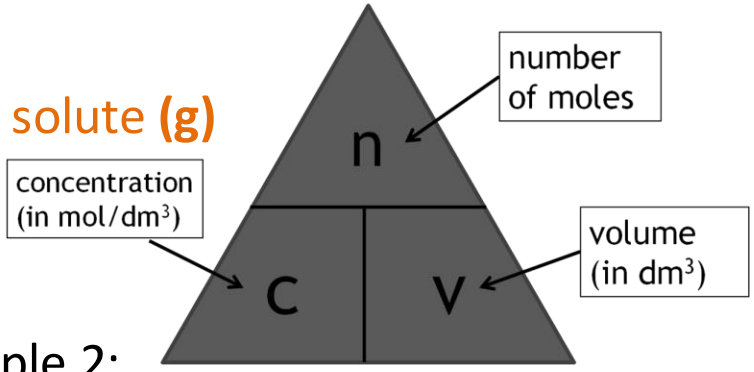
$$n = v\ c$$

Where:

n is the number of moles **(mol)** or the mass of the solute **(g)**

c is the concentration **(mol/dm³** or **g/dm³)**

v is the volume **(dm³)**



Example 1:

What is the concentration of a solution that has 35.0g of solute in 0.5dm³of solution?

35/0.5 = 70 g/dm³

Example 2:

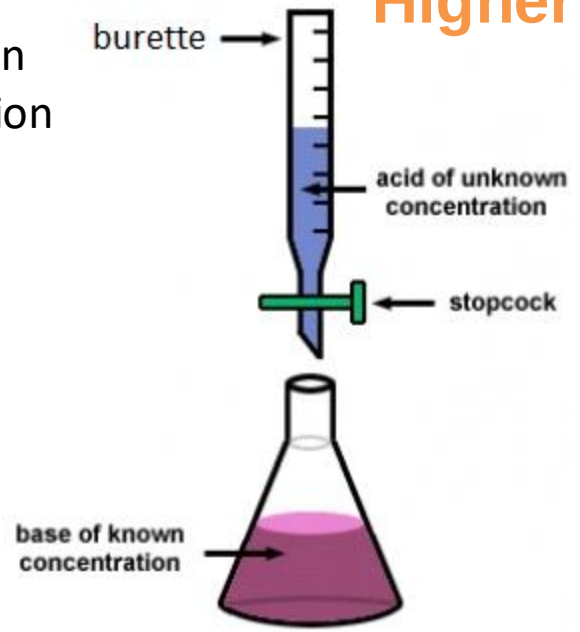
How many moles of magnesium nitrate are there in 0.50 dm³ of a 2 mol/dm³ solution?

2 x 0.50 = 1 mol

Reactions of acids – PART 2 - CHEMISTRY ONLY

Higher

If the volumes of two solutions that react completely are known and the concentrations of one solution is known, the concentration of the other solution can be calculated.



Example:

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

It takes 12.20cm^3 of sulfuric acid to neutralise 24.00cm^3 of sodium hydroxide solution, which has a concentration of 0.50mol/dm^3.

Calculate the concentration of the sulfuric acid in g/dm^3

0.5 mol/dm^3 x (24/1000) dm^3 = 0.012 mol of NaOH

The equation shows that 2 mol of NaOH reacts with 1 mol of H_2SO_4, so the number of moles in 12.20cm^3 of sulfuric acid is **(0.012/2) = 0.006 mol of sulfuric acid**

Calculate the concentration of sulfuric acid in mol/ dm^3

0.006 mol x (1000/12.2) dm^3 =0.49mol/dm^3

Calculate the concentration of sulfuric acid in g/ dm^3

H_2SO_4 = (2x1) + 32 + (4x16) = 98g

0.49 x 98g = 48.2g/dm^3

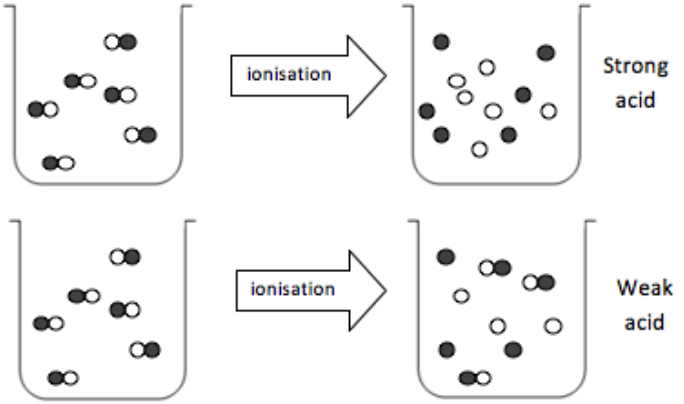
Reactions of acids - PART 2 - HIGHER

Acids must **dissolve** in water to show their acidic properties.
A **concentrated acid** has a relatively **large amount of solute** dissolved in the solvent.
A **dilute acid** has a relatively **smaller amount of solute** dissolved in the solvent

The molecules split up to form **hydrogen** ions.

A **strong** acid is **completely ionised** in aqueous solution. E.g. Hydrochloric, nitric and sulfuric acid.

A **weak** acid is only **partially ionised** in aqueous solution. E.g. Ethanoic, citric and carbonic.



A **weak acid** (aq) has a **lower pH** than a **strong acid** (aq) of the **same concentration**.

This is because a **weak acid** has a **lower concentration of hydrogen ions.**

As the pH decrease by one unit, the hydrogen ion concentration of the solution increase by a factor of 10.

Concentration of hydrogen ions in mol/dm^3	pH
0.10	1.0
0.010	2.0
0.0010	3.0
0.00010	4.0

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QuestionIT!

Reactions of acids
PART 2

- pH and neutralisation
- Titration (chemistry only)
- Strong and weak acids (HT)

Reactions of acids – QuestionIT

1. What ions do aqueous acids contain?
2. What ions do aqueous alkalis contain?
3. What is the pH scale?
4. How can pH be measured?
5. What pH is a neutral solution?
6. What pH do aqueous acid solutions have?
7. What pH do aqueous alkali solutions have?

8. Write a balanced symbol equation for the reaction between hydrogen ions and hydroxide ions.

9. CHEMISTRY ONLY: What are the units for the concentration of a solution?

10. CHEMISTRY ONLY: What is the concentration of a solution that has 40g of solute in 2dm^3 of solution?

11. HT: What is a strong acid?

12. HT: Name 3 strong acids.

13. HT: What is a weak acid?

14. HT: Name 3 weak acids.

15. HT: What is a dilute acid?

16. HT: What happens to the hydrogen ion concentration as the pH decreases by 1?

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AnswerIT!

Reactions of acids
PART 2

- pH and neutralisation
- Titration (chemistry only)
- Strong and weak acids (HT)

Reactions of acids – QuestionIT

1. What ions do aqueous acids contain?
   H^+
2. What ions do aqueous alkalis contain?
   OH^-
3. What is the pH scale?
   Measure of the acidity or alkalinity of a solution.
4. How can pH be measured?
   Universal indicator, pH probe.
5. What pH is a neutral solution?
   7
6. What pH do aqueous acid solutions have?
   Less than 7.
7. What pH do aqueous alkali solutions have?
   More than 7.

8. Write a balanced symbol equation for the reaction between hydrogen ions and hydroxide ions.

   $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

9. CHEMISTRY ONLY: What are the units for the concentration of a solution?

   mol/dm^3 or g/dm^3

10. CHEMISTRY ONLY: What is the concentration of a solution that has 40g of solute in $2dm^3$ of solution?

    Concentration = mass ÷ volume = 40 g ÷ 2 dm^3 = 20 g/dm^3

11. HT: What is a strong acid?
    Completely ionised in aqueous solution.
12. HT: Name 3 strong acids.
    Hydrochloric, nitric, sulfuric.
13. HT: What is a weak acid?
    Partially ionised in aqueous solution.
14. HT: Name 3 weak acids.
    Ethanoic, citric, carbonic
15. HT: What is a dilute acid?
    Contains less solute in the same volume
16. HT: What happens to the hydrogen ion concentration as the pH decreases by 1?
    Increases by a factor of 10

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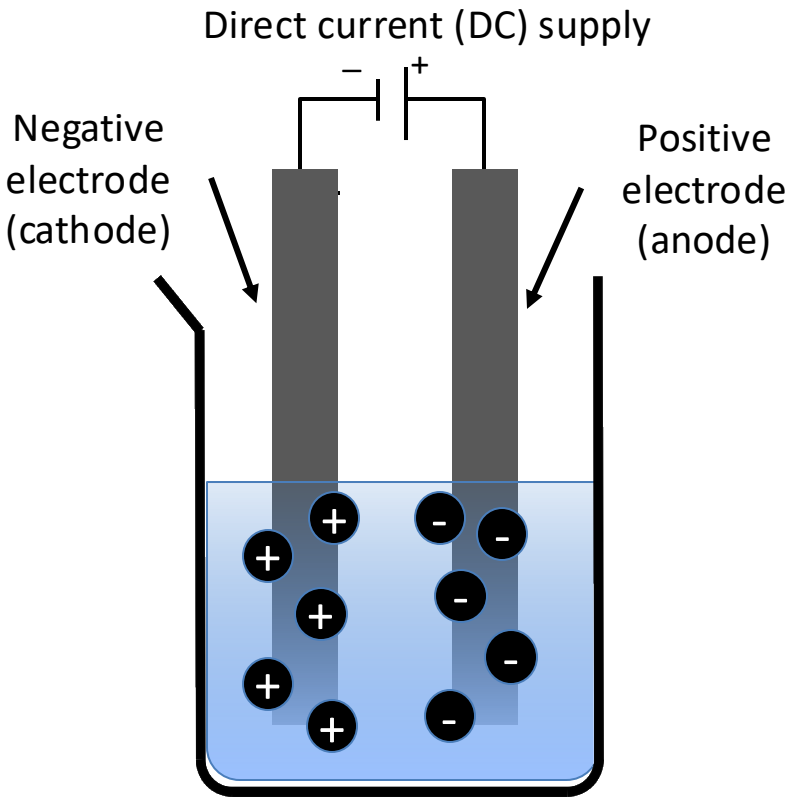
Electrolysis
PART 1

- The process of electrolysis
- Electrolysis of a molten ionic compounds
- Electrolysis of aqueous solutions

Electrolysis - PART 1

When an **ionic compound** is **melted** or **dissolved in water**, the **ions** are **free** to **move** about the liquid or solution. These liquids and solutions are able to **conduct electricity** and are called **electrolytes**. Passing an **electric current** though electrolytes causes the ions to move to the electrodes.



Positive ions go to **negative** electrode (cathode) and are **reduced** (gain of electrons).

Negative ions go to the **positive** electrode (anode) and are **oxidised** (loss of electrons).

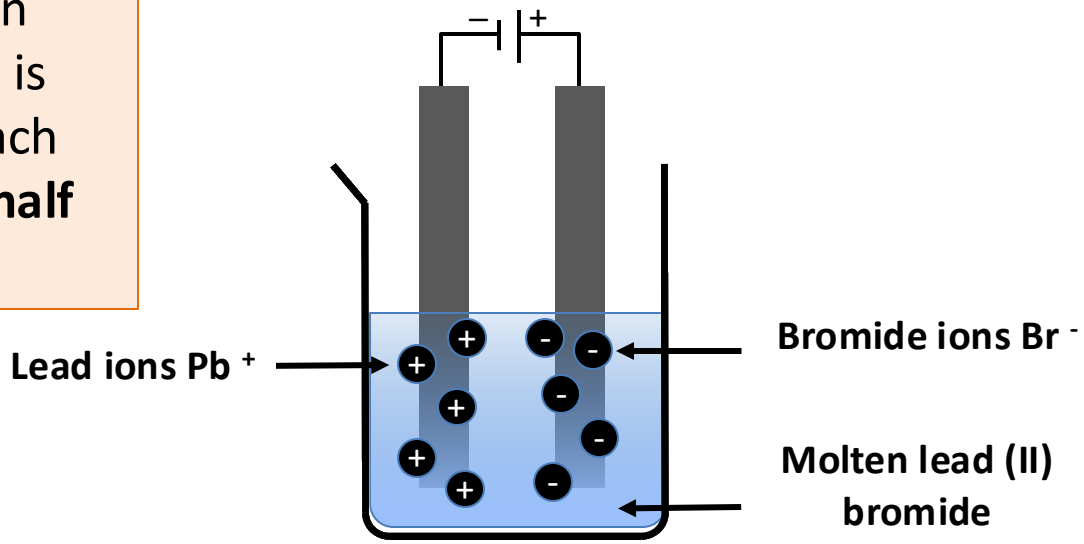
Ions are **discharged** at the electrodes producing **elements**. This is called **electrolysis**.

Electrolysis - PART 1

When an **ionic compound** is electrolysed in a **molten** state using inert electrodes, the **metal** is produced at the **cathode** and the **non-metal** is produced at the **anode**.

Higher: You can represent what is happening at each electrode using **half equations**.

lead bromide → lead + bromine



The positively charged lead ions Pb + (cations) are attracted to cathode and the negatively charged bromide ions Br - are attracted to the anode.

Higher:
At the cathode
$Pb^{2+} + 2e^- \rightarrow Pb$

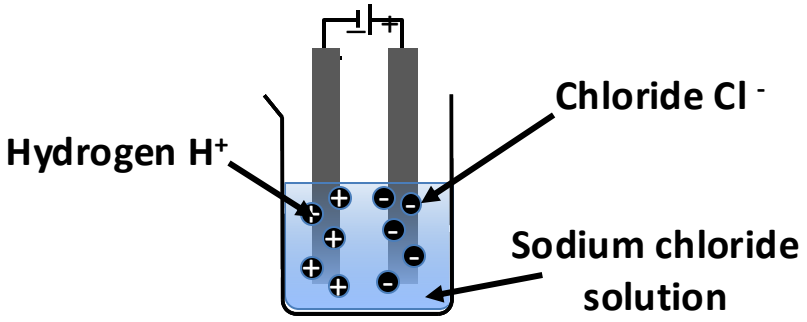
Higher:
At the anode
$2Br^- \rightarrow Br_2 + 2e^-$ or $2Br^- - 2e^- \rightarrow Br_2$

Electrolysis - PART 1

The **ions** discharged when an **aqueous solution** is electrolysed using inert electrodes depend on the relative **reactivity** of the elements involved.

At the **negative** electrode:

Metal will be produced on the electrode if it is **less** reactive than **hydrogen**.
Hydrogen will be produced if the metal is **more** reactive than hydrogen.



sodium chloride → hydrogen + chlorine + sodium hydroxide

Uses of the products:
Chlorine: Bleach and PVC
Hydrogen: Margarine
Sodium hydroxide: Bleach and soap

At the **positive** electrode:

Oxygen is formed at positive electrode.

Higher: At the anode
$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
or $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$

If you have a **halide** ion (Cl^-, I^-, Br^-) then you will get **chlorine, bromine or iodine** formed at that electrode.
This happens because in the aqueous solution, **water molecules** break down producing **hydrogen** ions and **hydroxide** ions that are discharged.
$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$

Higher:
At the cathode
$2H^+ + 2e^- \rightarrow H_2$

Higher:
At the anode
$2Cl^- \rightarrow Cl_2 + 2e^-$ or $2Cl^- - 2e^- \rightarrow Cl_2$

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QuestionIT!

Electrolysis
PART 1

- The process of electrolysis
- Electrolysis of a molten ionic compounds
- Electrolysis of aqueous solutio

1. Why can a molten or dissolved ionic compound conduct electricity?

2. What is electrolysis?

3. What is the name of the electrode that positive ions move to?

4. What is the name of the electrode that the negative ions move to?

5. What is produced at the cathode when lead bromide is electrolysed?

6. What is produced at the anode when lead bromide is electrolysed?

7. What is produced at the cathode is the metal in the solution is more reactive than hydrogen?

8. What is produced at the anode if the solution does not contain halide ions?

9. HT Only: Write half equations for the reactions that happen at the electrodes during the electrolysis of molten copper chloride.

10. Predict the products of electrolysis of copper sulfate solution

11. HT Only: Write a half equation for the reactions that happen at the electrodes during the electrolysis of copper bromide solution.

AnswerIT!

Electrolysis
PART 1

- The process of electrolysis
- Electrolysis of a molten ion compounds
- Electrolysis of aqueous solutions

1. Why can a molten or dissolved ionic compound conduct electricity?
   Free moving ions.
2. What is electrolysis?
   When an electric current is passed through a molten or aqueous ionic solution and the salt breaks down into simpler substances.
3. What is the name of the electrode that positive ions move to?
   Cathode.
4. What is the name of the electrode that the negative ions move to?
   Anode.
5. What is produced at the cathode when lead bromide is electrolysed?
   Lead.
6. What is produced at the anode when lead bromide is electrolysed?
   Bromine.

7. What is produced at the cathode is the metal in the solution is more reactive than hydrogen?

   Hydrogen.

8. What is produced at the anode if the solution does not contain halide ions?

   Oxygen.

9. HT Only: Write half equations for the reactions that happen at the electrodes during the electrolysis of molten copper chloride.

   Negative electrode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

   Positive electrode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

10. Predict the products of electrolysis of copper sulfate solution

    Positive electrode: Oxygen gas; Negative electrode: Copper.

11. HT Only: Write a half equation for the reactions that happen at the electrodes during the electrolysis of copper bromide solution.

    Negative electrode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

    Positive electrode: $2Br^-(aq) \rightarrow Br_2(g) + 2e^-$

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LearnIT!
KnowIT!

Electrolysis
PART 2

- Using electrolysis to extract metals

Electrolysis - PART 2

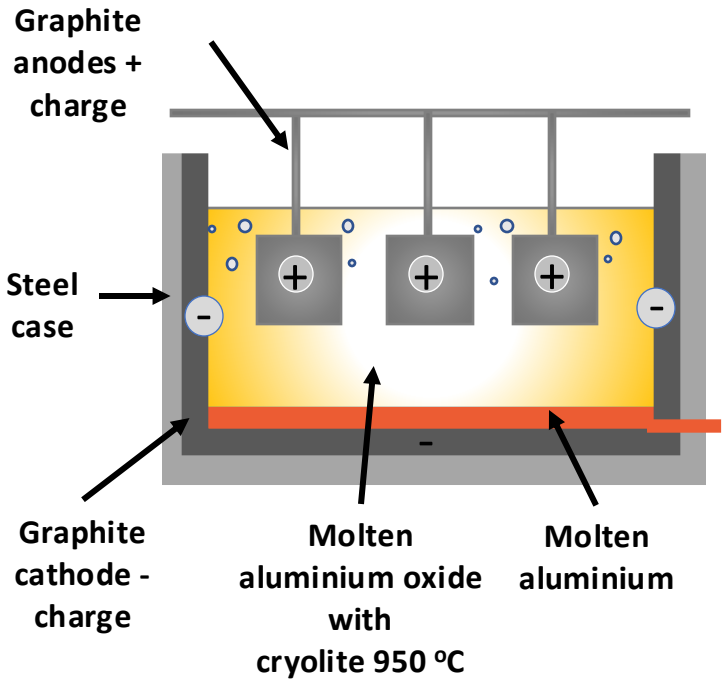
Metals can be extracted from **molten compounds** using electrolysis.

It is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon.

Large amounts of **energy** are used in the extraction process to melt the compounds and to produce the electrical current.

Aluminum is manufactured by electrolysis of molten aluminum oxide.

Aluminium oxide → aluminium + oxygen



Aluminium oxide has a very **high melting point** so is mixed with molten **cryolite** to lower the temperature required to carry out the electrolysis.

Aluminium goes to the negative electrode and sinks to bottom.

Higher: $Al^{3+} + 3e^- \rightarrow Al$

Oxygen forms at positive electrodes.

Higher: $2O^{2-} \rightarrow O_2 + 4e^-$

The oxygen reacts with the carbon electrode making carbon dioxide causing damage. The electrode needs **replacing** due to this reaction.

$C + O_2 \rightarrow CO_2$

QuestionIT!

Electrolysis
PART 2

- Using electrolysis to extract metals

1. Why is electrolysis used to extract aluminium form its ore?

2. Why is electrolysis an expensive way to extract metal from its ore?

3. Name the compound from which aluminium is extracted.

4. What is this compound dissolved in before electrolysis?

5. What is the anode made of?

6. Describe what happens at the positive electrode during the electrolysis of aluminium oxide.

7. **HT Only**: Write half equations for the reactions that occur at the positive and negative electrodes during the production of aluminium.

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AnswerIT!

Electrolysis
PART 2

- Using electrolysis to extract metals

1. Why is electrolysis used to extract aluminium form its ore?
   Aluminium is more reactive than carbon.
2. Why is electrolysis an expensive way to extract metal from its ore?
   Large amounts of energy needed.
3. Name the compound from which aluminium is extracted.
   Aluminium oxide/ bauxite.
4. What is this compound dissolved in before electrolysis?
   Cryolite.
5. What is the anode made of?
   Carbon.

6. Describe what happens at the positive electrode during the electrolysis of aluminium oxide.

   Oxide ions give up their electrons to form oxygen atoms, these join together in pairs to form oxygen gas, the oxygen reacts with the carbon electrode to make carbon dioxide gas.

7. **HT Only**: Write half equations for the reactions that occur at the positive and negative electrodes during the production of aluminium.

   Positive electrode: $2O^{2-}(aq) \rightarrow O_2(g) + 4e^-$

   Negative electrode: $Al^{3+}(aq) + 3e^- \rightarrow Al(l)$