

Oxidation Is Loss (of electrons) Reduction Is Gain (of electrons)

Ionic half equations (HT only)

For displacement reactions	<i>Ionic half equations show what happens to each of the reactants during reactions</i>	For example: The ionic equation for the reaction between iron and copper (II) ions is: $\text{Fe} + \text{Cu}^{2+} \rightarrow \text{Fe}^{2+} + \text{Cu}$ The half-equation for iron (II) is: $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$ The half-equation for copper (II) ions is: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
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Acid name	Salt name
Hydrochloric acid	Chloride
Sulfuric acid	Sulfate
Nitric acid	Nitrate

Oxidation and reduction in terms of electrons (HT ONLY)

Neutralisation of acids and salt production

sodium hydroxide + hydrochloric acid $\rightarrow$ sodium chloride + water
calcium carbonate + sulfuric acid $\rightarrow$ calcium sulfate, + carbon dioxide + water

Neutralisation	<i>Acids can be neutralised by alkalis and bases</i>	An alkali is a soluble base e.g. metal hydroxide. A base is a substance that neutralises an acid e.g. a soluble metal hydroxide or a metal oxide.
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Metals and oxygen	<i>Metals react with oxygen to form metal oxides</i>	magnesium + oxygen $\rightarrow$ magnesium oxide $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
Reduction	<i>This is when oxygen is removed from a compound during a reaction</i>	e.g. metal oxides reacting with hydrogen, extracting low reactivity metals
Oxidation	<i>This is when oxygen is gained by a compound during a reaction</i>	e.g. metals reacting with oxygen, rusting of iron

HT ONLY: Reactions between metals and acids are redox reactions as the metal donates electrons to the hydrogen ions. This displaces hydrogen as a gas while the metal ions are left in the solution.

Reactions with acids	<i>metal + acid $\rightarrow$ metal salt + hydrogen</i>	magnesium + hydrochloric acid $\rightarrow$ magnesium chloride + hydrogen zinc + sulfuric acid $\rightarrow$ zinc sulfate + hydrogen
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Acids react with some metals to produce salts and hydrogen.

Reactions of acids and metals

Reactions of acids

Extraction using carbon

Metals less reactive than carbon can be extracted from their oxides by reduction.

For example:
zinc oxide + carbon $\rightarrow$ zinc + carbon dioxide

Unreactive metals, such as gold, are found in the Earth as the metal itself. They can be mined from the ground.

Extraction of metals and reduction

AQA Trilogy C4: Chemical Changes

Reactivity of metals

The reactivity series

	Reactions with water	Reactions with acid
Group 1 metals	<i>Reactions get more vigorous as you go down the group</i>	<i>Reactions get more vigorous as you go down the group</i>
Group 2 metals	<i>Do not react with water</i>	<i>Observable reactions include fizzing and temperature increases</i>
Zinc, iron and copper	<i>Do not react with water</i>	<i>Zinc and iron react slowly with acid. Copper does not react with acid.</i>

Metal oxides

Metals form positive ions when they react	<i>The reactivity of a metal is related to its tendency to form positive ions</i>	The reactivity series arranges metals in order of their reactivity (their tendency to form positive ions).
Carbon and hydrogen	<i>Carbon and hydrogen are non-metals but are included in the reactivity series</i>	These two non-metals are included in the reactivity series as they can be used to extract some metals from their ores, depending on their reactivity.
Displacement	<i>A more reactive metal can displace a less reactive metal from a compound.</i>	Silver nitrate + Sodium chloride $\rightarrow$ Sodium nitrate + Silver chloride

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

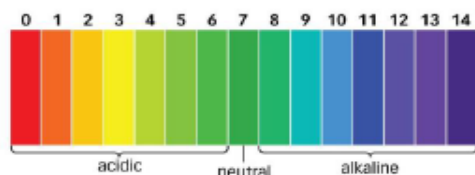
The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.

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.
At the positive electrode	Oxygen is formed at positive electrode. If you have a halide ion (Cl^- , I^- , Br^-) then you will get chlorine, bromine or iodine formed at that electrode.

Electrolysis of aqueous solutions

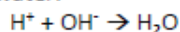
Strong acids	Completely ionised in aqueous solutions e.g. hydrochloric, nitric and sulfuric acids.
Weak acids	Only partially ionised in aqueous solutions e.g. ethanoic acid, citric acid.
Hydrogen ion concentration	As the pH decreases by one unit (becoming a stronger acid), the hydrogen ion concentration increases by a factor of 10.

Soluble salts	Soluble salts can be made from reacting acids with solid insoluble substances (e.g. metals, metal oxides, hydroxides and carbonates).
Production of soluble salts	Add the solid to the acid until no more dissolves. Filter off excess solid and then crystallise to produce solid salts.



You can use universal indicator or a pH probe to measure the acidity or alkalinity of a solution against the pH scale.

In neutralisation reactions, hydrogen ions react with hydroxide ions to produce water:



Acids	Acids produce hydrogen ions (H^+) in aqueous solutions.
Alkalis	Aqueous solutions of alkalis contain hydroxide ions (OH^-).

Process of electrolysis	Splitting up using electricity	When an ionic compound is melted or dissolved in water, the ions are free to move. These are then able to conduct electricity and are called electrolytes. Passing an electric current through electrolytes causes the ions to move to the electrodes.
Electrode	Anode Cathode	The positive electrode is called the anode. The negative electrode is called the cathode.
Where do the ions go?	Cations Anions	Cations are positive ions and they move to the negative cathode. Anions are negative ions and they move to the positive anode.

Electrolysis

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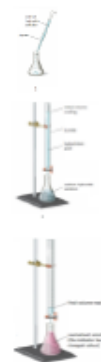
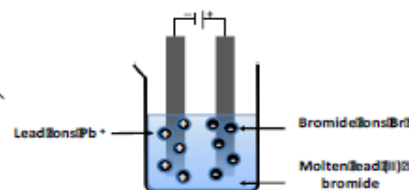
C4: Chemical Changes

Reactions of acids

Strong and weak acids (HT ONLY)

Soluble salts

The pH scale and neutralisation



Extracting metals using electrolysis	Metals can be extracted from molten compounds using electrolysis.
	This process is used when the metal is too reactive to be extracted by reduction with carbon.
	The process is expensive due to large amounts of energy needed to produce the electrical current. Example: aluminium is extracted in this way.

Higher tier: You can display what is happening at each electrode using half-equations:

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

At the anode: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$