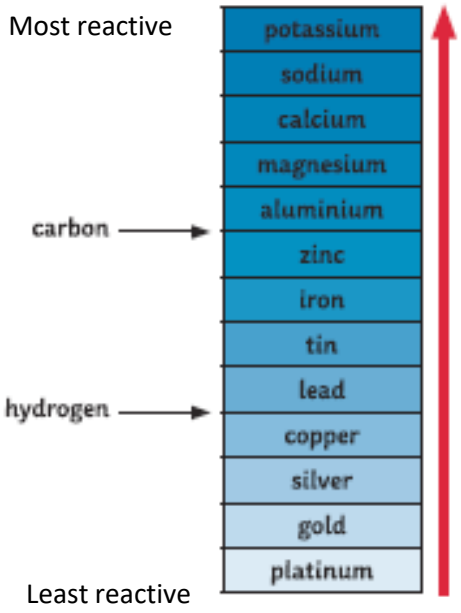
C4 Knowledge Organiser – Chemical Changes

Reactivity Series 1



Metals less reactive than carbon can be extracted from their oxides by reduction.
For example: zinc oxide + carbon → zinc + carbon dioxide.
Metals more reactive than carbon require electrolysis for extraction.
Unreactive metals, such as gold, are found in the Earth as the metal itself. They can be mined from the ground.

Reactions of metals with water: 2

Metal + Water → Metal Hydroxide + Hydrogen

Reactions of metals with dilute acid:
Metal + Acid → Salt + Hydrogen

Ores 3

Rock containing enough mineral or metal for extraction.
There must be enough mineral or metal to make a profit. Metals are usually in the form of compounds within the ore (typically metal oxides).

More Important Reactions: 4

Metal + Oxygen → Metal oxide

Thermal decomposition: breakdown of compounds using heat:
$CaCO_3 \rightarrow CaO + CO_2$

Recycling Metals 5

Recycling is used to conserves ores, reduce energy required for extraction and minimise pollution. It also reduces the use of landfill and the destruction of habitats. Some metals cannot be reused because of damage, need for paint removal, rusting/corrosion, metal fatigue.

Displacement reactions 6

A less reactive metal is displaced from its compound by a more reactive metal.
e.g.
Tin oxide + Sodium → Sodium oxide + Tin

Extracting metals using electrolysis 7

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.

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Redox Reactions and Ionic Half Equations (H Tier) 1

Oxidation **I**s L**oss** (of electrons)

Reduction I**s** **G**ain (of electrons)

The ionic equation for the reaction between iron and copper (II) ions is:

$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$

The half-equation for iron (II) is:

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

The half-equation for copper (II) ions is:

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

More Acid Reactions 2

Neutralisation reactions: Acid + Alkali → Salt + Water

Acid + BASE → Salt + Water

Acid + Carbonate → Salt + Water + Carbon Dioxide

Base = insoluble alkalis e.g. insoluble metal oxides and hydroxides

pH Scale 3

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

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:

$H^+ + OH^- \rightarrow H_2O$

Strong and weak acids (H Tier) 4

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.*

Naming salts 5

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

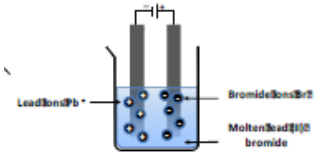
First part comes from the metal in the compound, second from which acid it reacted with.

Half equations (H tier) 6

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

At the cathode: $Pb2+ + 2e- \rightarrow Pb$

At the anode: $2Br- \rightarrow Br2+ 2e-$



Basic electrolysis 7

Positive
Anode
Negative
Is
Cathode

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 though 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.

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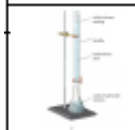
Making Soluble Salts RP. 1

1. Make a saturated solution by stirring copper oxide into the sulfuric acid until no more will dissolve.
2. Filter the solution to remove the excess copper oxide solid.
3. Half fill a beaker with water and set this over a Bunsen burner to heat the water. Place an evaporating dish on top of the beaker.
4. Add some of the solution to the evaporating basin and heat until crystals begin to form.
5. Once cooled, pour the remaining liquid into a crystallising dish and leave to cool for 24 hours.
6. Remove the crystals with a spatula and pat dry between paper towels

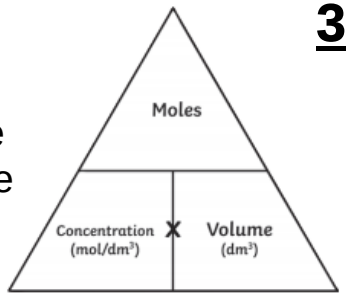
Titration RP (Chem only). 2

Titrations are used to work out the precise volumes of acid and alkali solutions that react with each other.

	1. Use the pipette to add 25 cm³ of alkali to a conical flask and add a few drops of indicator. **Indicator: methyl orange. Turns from orange to red upon neutralisation**
	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.

Titration RP (Chem only). 3

Using the results from a titration experiment, it is possible to calculate the concentration of a solution or the volume of solution required to neutralise an acid or alkali.



Calculating the chemical quantities in titrations involving concentrations in mol/dm³ and in g/dm³ (HT ONLY):

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

It takes 12.20cm³ of sulfuric acid to neutralise 24.00cm³ of sodium hydroxide solution, which has a concentration of 0.50mol/dm³.

Calculate the concentration of the sulfuric acid in g/dm³

0.5 mol/dm³ x (24/1000) dm³ = 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 12.20cm³ of sulfuric acid is (0.012/2) = 0.006 mol of sulfuric acid

Calculate the concentration of sulfuric acid in mol/ dm³

0.006 mol x (1000/12.2) dm³ =0.49mol/dm³

Calculate the concentration of sulfuric acid in g/ dm³

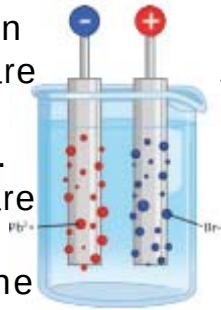
H_2SO_4 = (2x1) + 32 + (4x16) = 98g

0.49 x 98g = 48.2g/dm³

Electrolysis of aqueous solutions and molten ionic compounds. 4

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.

Lead bromide is an ionic compound. Ionic compounds, when solid, are not able to conduct electricity. When molten or in solution, the ions are free to move and are able to carry a charge. The positive lead ions are attracted toward the negative cathode at the same time as the negative bromide ions are attracted toward the positive anode.



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

Aluminium is manufactured by electrolysis from aluminium oxide which has a very high melting point. It takes large amount of energy and money to turn it molten. Therefore CRYOLITE is added to aluminium oxide to lower the melting point and reduce cost.

OIL RIG (Higher Tier Only).

We represent what is happening at the electrodes by using half equations.

Lead ions reduced (gain e-):

$Pb^2 + 2e \rightarrow Pb$

Bromide ions reduced (lose e-):

$2Br\ Br_2 + 2e$

Oxidation Is Loss (OIL)

Reduction Is Gain (RIG)