Chemistry Knowledge Organiser
Chemical changes (Trilogy Science)

Redox - oxygen

Oxidation – when a substance gains oxygen

Reduction – when a substance loses oxygen

Reactivity series

When metals react they form positive ions, reactivity is the tendency to form these positive ions.

Use observations from reactions of metals with water and dilute acids to put in order of reactivity.

Metal	Observation with cold water	Observation with acid
Potassium	Rapid effervesce	Rapid effervesce
Sodium		
Lithium		
Calcium		
Magnesium	Very slow	
Carbon	-	-
Zinc	No reaction	Slow effervesce
Iron	No reaction	Very slow effervesce
Hydrogen	-	-
Copper	No reaction	No reaction

Displacement reactions

A more reactive metal will displace a less reactive metal from its compounds.

For example iron will displace copper from iron sulfate:

Balanced symbol eq: $Fe + CuSO_4 \rightarrow Cu + FeSO_4$

An ionic equation only includes the ions which have changed in some way.

Ionic equation: $Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Fe^{2+}_{(aq)}$

Extraction of metals

Metals are extracted from metal oxides by loss of oxygen – reduction

increasing reactivity

potassium
sodium
calcium
magnesium
aluminium
(carbon)
zinc
iron
lead
(hydrogen)
copper
silver
gold
platinum

Metals above carbon in the reactivity series must be extracted using **electrolysis**. Electrolysis can also be used to purify copper.

Metals less reactive than carbon can be extracted from their ores by **reduction** using carbon, coke or charcoal.

Platinum, gold, silver and copper can occur **native** and do not need to be extracted.

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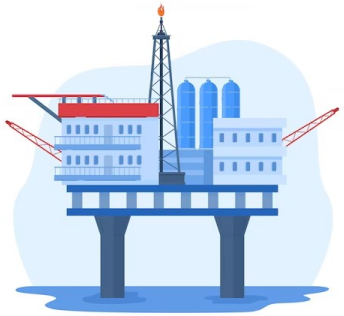
Redox - electrons

Oxidation
Is
Loss

Reduction
Is
Gain



Neutralisation

Base – substance that neutralises an acid

Alkali – soluble base – release OH^- ions

Acid – release H^+ ions in solution

In neutralisation reactions between an acid and an alkali, hydrogen ions react with hydroxide ions to form water.

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

General equations (LEARN):

acid + base → salt + water

acid + carbonate → salt + water + carbon dioxide

(acid + metal → salt + hydrogen) - not neutralisation

Naming salts

Hydrochloric acid (HCl) – **chloride**

Sulfuric acid (H_2SO_4) – **sulfate**

Nitric acid (HNO_3) - **nitrate**

The first part of the name of the salt is a positive ion (often a metal) - comes from the base (or metal) used, e.g. potassium sulfate

If ammonia is used and ammonium salt (NH_4^+) is formed

Making soluble salts (required practical)

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.

Method:

- use appropriate acid and insoluble base – name them!
- add insoluble base to the (hydrochloric) acid in a beaker
- stir the mixture and warm – do not heat
- continue adding until the base is in excess
- shown by excess solid and no more effervescence (if carbonate)
- filter the reaction mixture
- to remove the excess base
- heat the solution
- using a water bath or electric heater
- to crystallisation point
- leave the solution to crystallise as the solution cools down
- pat crystals dry with filter paper

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pH scale

The pH scale, from 0 to 14, is a measure of the acidity or alkalinity of a solution, and can be measured using universal indicator or a pH probe.

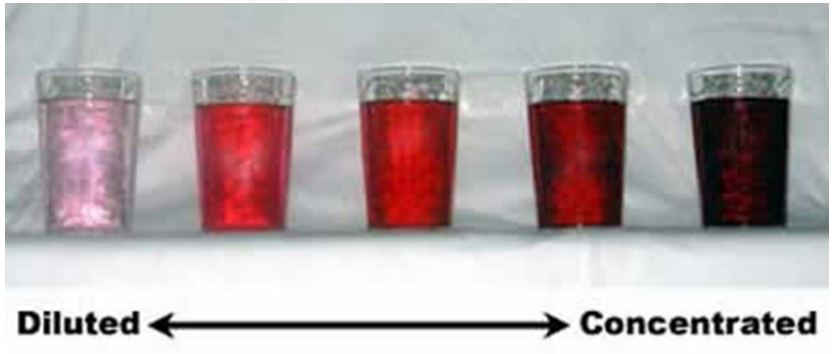
A solution with pH 7 is neutral. Aqueous solutions of acids have pH values of less than 7 and aqueous solutions of alkalis have pH values greater than 7.

The pH scale is a measure of the concentration of $H^+_{(aq)}$ ions in solution.

As the concentration of $H^+_{(aq)}$ ions increases by a factor of 10 the pH decreases by 1.

Concentrated vs dilute acids

A concentrated solution has more acid particles per unit volume than a dilute acid.



As concentration of an acid increases the pH will decrease.

Strong vs Weak acid

A strong acid is completely ionised in aqueous solution. Examples of strong acids are hydrochloric, nitric and sulfuric acids.

A weak acid is only partially ionised in aqueous solution. Examples of weak acids are ethanoic, citric and carbonic acids.