science
chemistry

GCSE C4 – Chemical Changes

1) Reactions of metals

- Metals can react with oxygen (oxidised) to make metal oxides.
- Some metals will react with water (alkali metals) and most metals will react with dilute acids to produce hydrogen gas and a salt.
- The reactivity of a metal depends on its ability to lose electrons (to form positive ions).

2) The Reactivity Series

Potassium > sodium > lithium > calcium > magnesium > **carbon** > zinc > iron > **hydrogen** > copper > silver > gold are in order of most reactive to the least reactive based on their reactions with water and dilute acids.

3) Displacement reactions

- When a more reactive metal can displace a less reactive one from a compound.

e.g. zinc + lead nitrate → zinc nitrate + lead

4) Oxidation & Reduction

Oxidation is the gain of oxygen   Reduction is the loss of oxygen
e.g. Iron oxide + carbon → iron + carbon dioxide
We can say the iron oxide has been reduced because it has lost oxygen & the carbon has been oxidised as it has gained oxygen.
When both occur, this is called a REDOX reaction.

5) Extracting metals

- Unreactive metals such as gold are found uncombined in the Earth.
- Most other metals are found as compounds in ores.

- Metals less reactive than carbon can be extracted by heating the ores with carbon (reduction see above)
- Tungsten is a special case where the ore is heated with hydrogen (this is also a reduction reaction)

- Metals more reactive than carbon are extracted by electrolysis (expensive due to large of amounts of energy needed).

6) Electrolysis

- When an ionic compound is melted (molten) or dissolved in water (aqueous), the ions are free to move. These liquids are able to conduct electricity and are called electrolytes.
- Electrolysis is the decomposition (breaking down) of an ionic substance using electricity (dc).
- The electrodes are made from inert material such as graphite or platinum.

7) Electrolysis of molten compounds (pure liquids)

- Metal ions move to the negative electrode.
- Non-metal ions move to the positive electrode.
- **Think PANIC (Positive Anode Negative Is Cathode)**

8) The electrolysis of solutions

The products of a solution will be different from a pure liquid due to the presence of the water **(contains H^+ and OH^- ions)**

During electrolysis, these ions compete with the metal and non-metal ions in solution, to gain or lose electrons.

So what forms at each?

At the cathode:
At the negative electrode, the least reactive element (between hydrogen and the metal) forms.

At the anode:
At the positive electrode, oxygen is produced unless the solution contains halide ions when the halogen is produced.

This is because if halide ions are present, **Cl^-, Br^-, I^-,** they will give up their electrons to become molecules of **Cl_2, Br_2** and **I_2**.
If no halogen is present, **OH^-** will lose electrons and oxygen forms.

9) REDOX (HT ONLY)

Oxidation & reduction can also be defined in terms of electrons
Oxidation – the loss of electrons (OIL)
Reduction – the gain of electrons (RIG)
When both occur, this is called a REDOX reaction.

10) Soluble Salts

Made by reacting acids with:

- Metals (some metals are too reactive & some don't react).
- Insoluble bases (see box 13).

Bases include: **Metal oxides**, **metal hydroxides** and metal carbonates. These will all neutralise acids.
Acid + metal → salt + hydrogen
Acid + metal oxide → salt + water
Acid + metal carbonate → salt + water + carbon dioxide

The name of the salt made depends on the acid and metal used:

- Hydrochloric acid (HCl) makes chlorides
- Sulfuric acid (H_2SO_4) makes sulfates
- Nitric acid (HNO_3) makes nitrates

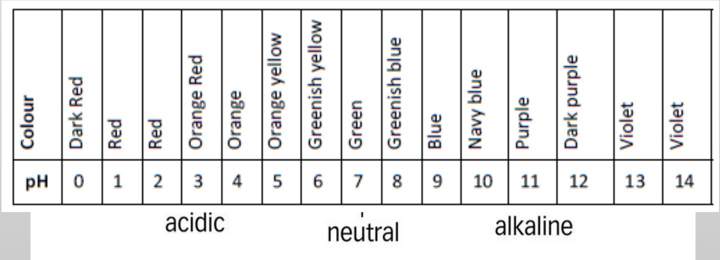
e.g. hydro**chloric** acid + zinc oxide → zinc **chloride** + water

11) pH Scale

- Acids produces H^+ ions in solution.
- Alkaline solutions contain OH^- ions

- In neutralisation reactions hydrogen ions from the acid react with hydroxide ions to produce water as shown in the equation below.

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



Colour	Dark Red	Red	Red	Orange Red	Orange	Orange yellow	Greenish yellow	Green	Greenish blue	Blue	Navy blue	Purple	Dark purple	Violet	Violet
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14

acidic neutral alkaline

- pH scale goes from 0-14
- Remember a lower pH is more acidic!
- The colours observed in the above scale are for Universal Indicator.

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12) Strong and weak acids (HT ONLY)

- **A strong acid completely ionises in aqueous solution.**

e.g. $HCl \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

- Examples of strong acids are hydrochloric, sulfuric & nitric.

- **Weak acids are only partially ionised in aqueous solution.**

e.g. $CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

- Examples of weak acids are ethanoic, citric & carbonic.

- The degree of acidity is measured by the pH scale (power of hydrogen ions).
- When the concentration of H^+ ions is decreased by a factor of 10, the pH goes up by one unit (remember pH scale goes from 0 (most acidic) to 14 (most basic).
- So pH3 would have 100x more hydrogen ion concentration than pH5.

13) Extraction of aluminium

- Manufactured by electrolysis of a molten mixture of **aluminium oxide** and **cryolite** using carbon electrodes.
- The positive electrode has to continually be replaced as the oxygen that forms there reacts with the carbon to form carbon dioxide and the electrodes slowly disappear.

- Expensive process since the aluminium oxide has to be melted to a very high temperature.
- The melting point can be lowered by adding cryolite.
- Lots of electrical energy also needed which is expensive.

14) Required Practical

Preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution.

Key steps

- Add named acid to a beaker
- Gently warm the acid
- Add named metal oxide or carbonate to the acid and stir
- Continue adding until the metal oxide is in excess (solid seen at bottom of beaker)
- Filter using a filter paper and funnel to remove excess metal oxide (or carbonate)
- Pour solution into an evaporating dish
- Heat solution until some of the water evaporates
- Leave to cool overnight and crystals appear.

e.g. making copper oxide

- Add excess copper oxide powder to a beaker of warm sulfuric acid to ensure all the sulfuric acid is used up.
- The excess CuO is filtered and the water can be evaporated off to obtain pure crystals of copper sulfate.

$CuO\ (s) + H_2SO_4\ (aq) \rightarrow CuSO_4\ (aq) + H_2O\ (l)$

15) Required Practical – Electrolysis of solutions

Investigate what happens when aqueous solutions are electrolysed using inert electrodes. This should be an investigation involving developing a hypothesis.

- Use inert electrodes used (carbon) to electrolyse sodium chloride solution and copper chloride solution.
- Look for the symbols aq (solution) and l (molten liquid)

- If you electrolyse a solution of copper sulfate, copper forms at the negative electrode and oxygen form at the positive electrode.
- If you repeated with silver carbonate, you would get silver and oxygen forming.
- You might conclude from this that you always get the metal and oxygen from electrolysis.
- But then you electrolyse sodium chloride solution and you get hydrogen at the negative electrode and chlorine forming at the positive electrode.
- If you repeat with different metals solutions of differing reactivity, you would find that….

- At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen.

- At the positive electrode (anode), oxygen is produced unless the solution contains halide ions then the halogen is produced.