



### 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 – the gain of oxygen

Reduction – 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 and the carbon has been oxidised.

### 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<sup>+</sup> and OH<sup>-</sup> 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<sup>-</sup>, Br<sup>-</sup>, I<sup>-</sup>**, they will give up their electrons to become molecules of **Cl<sub>2</sub>, Br<sub>2</sub> and I<sub>2</sub>**. If no halogen is present, **OH<sup>-</sup>** 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

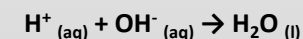
The name of the salt made depends on the acid and metal used:

- Hydrochloric acid (HCl) makes chlorides
- Sulfuric acid (H<sub>2</sub>SO<sub>4</sub>) makes sulfates
- Nitric acid (HNO<sub>3</sub>) makes nitrates

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

### 11) pH Scale

- Acids produces H<sup>+</sup> ions in solution.
- Alkaline solutions contain OH<sup>-</sup> ions
- In neutralisation reactions hydrogen ions from the acid react with hydroxide ions to produce water as shown in the equation below.



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.

**12) Strong and weak acids (HT ONLY)**

- A strong acid completely ionises in aqueous solution.  
e.g.  $\text{HCl} \rightarrow \text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$
- Examples of strong acids are hydrochloric, sulfuric & nitric.
- Weak acids are only partially ionised in aqueous solution.  
e.g.  $\text{CH}_3\text{COOH}_{(\text{aq})} \rightleftharpoons \text{H}^+_{(\text{aq})} + \text{CH}_3\text{COO}^-_{(\text{aq})}$
- 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  $\text{H}^+$  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**

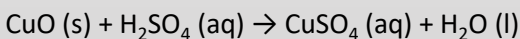
Preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution.

**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 crystallisation point
- Leave for rest of water to evaporate 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  $\text{CuO}$  is filtered and the water can be evaporated off to obtain pure crystals of copper sulfate.

**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 would form at the negative electrode and oxygen would 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.