4.4. Revision Checklist: Chemical Changes

Reactivity of metals

When metals react with other substances the metal atoms form positive ions. The reactivity of a metal is related to its tendency to form positive ions.

Metals react with oxygen to produce metal oxides. The reactions are **oxidation** reactions because the metals **gain oxygen**.

calcium + oxygen → calcium oxide

$2\,Ca + O_2 \rightarrow 2CaO$

$4K + O_2 \rightarrow 2\,K_2O$

Reactive metals such as magnesium will burn with a flame in oxygen. Less reactive metals like iron will tarnish and change colour without a flame.

Metals can be arranged in order of their reactivity in a reactivity series.

The metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids.

The non-metals hydrogen and carbon are often included in the reactivity series.

Reactive metals such as potassium, sodium and calcium will react with cold water producing bubbles of hydrogen gas
All metals above hydrogen in the reactivity series will react with acids producing bubbles of hydrogen gas.
The more reactive the metal the faster/more vigorous the reaction will be producing hydrogen gas quicker.
Metals below hydrogen in the reactivity series will not react with acids.

Reactivity series of Metals

most reactive

Potassium
Sodium
Calcium
Magnesium
Aluminium
Carbon
Zinc
Iron
Tin
Lead
Hydrogen
Copper
Silver
Gold

least reactive

The most important metals in the reactivity series can be learnt by using mnemonics like: **P**lease **S**top **C**alling **M**y **A**unty **Z**ebra **I**n **T**he **C**lass- **p**otassium, **s**odium, **c**alcium, **m**agnesium, **a**luminium, **z**inc, **i**ron, **t**in, **c**opper

Displacement Reactions

A more reactive metal can displace a less reactive metal from a compound

e.g. copper sulfate + magnesium → copper + magnesium sulfate

$CuSO_4 + Mg \rightarrow Cu + MgSO_4$

Magnesium is more reactive than copper so there is a displacement reaction.

e.g. zinc sulfate + copper → no reaction

Copper is less reactive than zinc so there is no reaction.

Extracting metals

Unreactive metals such as gold are found in the Earth as the metal itself but most metals are found as compounds that require chemical reactions to extract the metal.

Reduction is defined as **loss of oxygen**

Metals that are **less reactive than carbon** can be extracted from their oxides **by reduction** with carbon

iron oxide is reduced in the blast furnace by reacting with carbon or carbon monoxide to make iron

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

carbon monoxide + iron(III) oxide → carbon dioxide + iron
*(in this equation the iron oxide is being **reduced** and the carbon monoxide is causing the reduction)*

Metals above carbon in the reactivity series are often extracted by **electrolysis** or by displacement reactions with more reactive metals

Oxidation and Reduction

Ionic equations

In reactions **oxidation** can be defined as **loss of electrons** and **reduction** as **gain of electrons**

In displacement reactions, it is only the metal ions that are reacting in terms of electrons being transferred. The non metal ions e.g. sulfate (SO_4^{2-}) ions are 'spectating'.

$Zn\ (s) + CuSO_4\ (aq) \rightarrow Cu\ (s) + ZnSO_4\ (aq)$

We can write the equation above as an **ionic** equation which does not include the spectator ion.

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

In this reaction zinc has lost electrons (to become more positive) and so has **oxidised**

This can be written as a **half equation** $Zn \rightarrow Zn^{2+} + 2e^-$

In this reaction copper has gained electrons (to become neutral) and so has **reduced**

This can be written as a **half equation** $Cu^{2+} + 2e^- \rightarrow Cu$

Reactions of acid with metals
Acids react with some metals to produce salts and hydrogen
Metal + acid → salt + hydrogen

$Mg + 2HCl \rightarrow MgCl_2 + H_2$
$Zn + 2HCl \rightarrow ZnCl_2 + H_2$
$Fe + 2HCl \rightarrow FeCl_2 + H_2$

$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$
$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

All metals above hydrogen in the reactivity series will react with acids producing bubbles of hydrogen gas.
Metals below hydrogen in the reactivity series will **not** react with acids.

Reactions of acid and metals are **redox reactions**

These equations can be written as an ionic equation
$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$
$Fe + 2H^+ \rightarrow Fe^{2+} + H_2$
$Zn + 2H^+ \rightarrow Zn^{2+} + H_2$

They can also be written as a half equations

$Mg \rightarrow Mg^{2+} + 2e^-$ The magnesium is **oxidised** as it is **losing** electrons

$2H^+ + 2e^- \rightarrow H_2$ The hydrogen is **reduced** as it is **gaining** electrons

Neutralisation Reactions

Acids are neutralised by alkalis (eg soluble metal hydroxides) and bases (eg insoluble metal hydroxides and metal oxides) to produce salts and water

base + acid → salt + water

Example:
copper oxide + sulfuric acid → copper sulfate + water
$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

zinc hydroxide + nitric acid → zinc nitrate + water
$Zn(OH)_2 + 2HNO_3 \rightarrow Zn(NO_3)_2 + 2H_2O$

Acids are also neutralised by metal carbonates to produce salts, water and carbon dioxide.
metal carbonate + acid → salt + water + carbon dioxide

Example
$CuCO_3 + 2HCl \rightarrow CuCl_2 + H_2O + CO_2$

The particular salt produced in any reaction between an acid and a base or alkali depends on:

- the acid used
  - hydrochloric (HCl) acid produces **chlorides**,
  - nitric acid (HNO_3)produces **nitrates**,
  - sulfuric acid (H_2SO_4) produces **sulfates**
- the metal in the base or alkali.

Making soluble salts

Soluble salts can be made from acids by reacting them with **solid insoluble** substances, such as metals, metal oxides, hydroxides or carbonates.

- not all metals are suitable, some are too reactive/dangerous and others are not reactive enough
- copper is too unreactive and **does not react with acid**.
- Sodium is too reactive and **reacts too violently**

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.

Required Practical Detailed method for forming a salt from a solid base (or metal)

- Measure 50 cm^3 acid in beaker.
- **Warm acid** (*to speed up reaction rate*)
- **Add spatulas of solid base** (or metal) to acid
- stir
- Keep adding solid **until no more reacts** and **base is in excess.** (*You will see solid sinking to the bottom of the beaker*)(*This step ensures all acid is reacted*)
- **Filter off excess solid base** with filter funnel and filter paper.
- Pour solution into **evaporating basin.**
- Heat using a **water bath** to crystallisation point.
- **Leave to crystallise.**
- Dry crystals on filter paper.

pH Scale and neutralisation

Acids produce **hydrogen ions (H^+)** in aqueous solutions.

Aqueous solutions of **alkalis** contain **hydroxide ions ($OH^–$)**.

Base and Alkalis
Bases neutralise acids
Metal oxides and hydroxides are bases .
Soluble hydroxide bases are called **alkalis**

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.
Aqueous solutions of alkalis have pH values greater than 7.

Neutralisation

In neutralisation reactions between an acid and an alkali, hydrogen ions react with hydroxide ions to produce water.
This reaction can be represented by the ionic equation: **$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.**

Universal indicator colours
Red pH 1-2 strong acid
Orange/yellow pH3-6 weak acid
Green pH7 neutral
blue pH8-11 weak alkali
Purple pH 12-14 strong alkali

Strong Acids
A strong acid is **completely ionised** in **aqueous solution**. Examples of strong acids are hydrochloric, nitric and sulfuric acids.
$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Weak Acids
A weak acid is only **partially ionised** in **aqueous solution**. Examples of weak acids are ethanoic, citric and carbonic acids.
$CH_3COOH(aq) \rightleftharpoons H^+ + CH_3COO^-$
Ethanoic acid ethanoate ion
Less than 1% of ethanoic acid molecules will ionise in this reversible reaction.

For a given concentration of aqueous solutions, the stronger an acid, the lower the pH.

Dilute and Concentrated Acids
A concentrated acid will have more moles of H^+ ions per unit volume than a dilute acid.
A concentrated acid is made into a dilute acid by adding water.
Dilute and concentrated does not mean the same as strong and weak.

pH and Concentration
As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.

If an acid is **diluted 10 times** its pH will increase by **one unit** e.g. if a solution of hydrochloric acid has a concentration of 0.01 mol/dm^3 it will have a pH of 2. If it is then diluted a 100 times it will have a pH of 4

Concentration of $H^+(aq)$ ions in mol/dm^3	pH
1.0	0.0
1.0×10^{-1}	1.0
1.0×10^{-2}	2.0
1.0×10^{-3}	3.0
1.0×10^{-4}	4.0
1.0×10^{-5}	5.0
1.0×10^{-6}	6.0

Required practical :Titrations

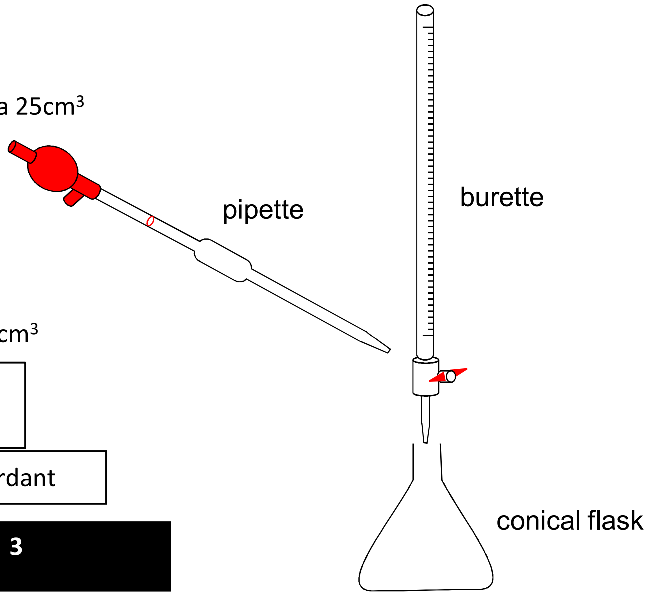
Chemistry only

The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator.
If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated.

A pipette measures one fixed volume accurately.
A burette measures variable volume

Know the basic method for doing a titration

- alkali in **burette**
- acid in **conical flask** measured out with a 25cm^3 **pipette**
- few drops of **indicator**
- add alkali to acid until colour changes
- swirl conical flask
- add alkali dropwise towards the end
- note final burette reading
- repeat until two readings are within 0.1cm^3



Common indicator: Phenolphthalein
Colour in acid: colourless Colour in alkali: pink

Results within 0.10cm^3 of each other are called concordant

Titration number	1	2	3
Initial burette reading (cm^3)	0.50	2.50	1.55
Final burette reading (cm^3)	24.50	27.00	25.95
Titre (cm^3)	24.00	24.50	24.40

Working out average titre results
Only make an average of the concordant titre results

Average titre = (24.50+ 24.40)/2 = 24.45

Titration calculations

General method
Step 1: calculate the number of moles of the substance for which the volume and concentration has been given.
Using **number of moles = concentration x vol (in dm^3)**
Step 2 : use the balanced equation to work out the moles of the other substance
e.g. NaOH and HCl react on a 1:1 ratio ($NaOH + HCl \rightarrow NaCl + H_2O$)
Step 3: calculate the concentration of the second substance
Using **concentration = moles / volume**

Example 1
25.0 cm^3 of HCl is reacted with 22.4cm^3 of 2 mol/dm^3 NaOH. Calculate the concentration of the HCl.

step 1: work out the number of moles of NaOH
Number of moles = conc x vol = 2 x 22.4/1000 = 0.0448 mol

step2 :use the balanced equation to work out the moles of HCl
$NaOH + HCl \rightarrow NaCl + H_2O$
1 mole NaOH reacts with 1 mole HCl **1:1 ratio**
So 0.0448 moles NaOH reacts with 0.0448 moles HCl

Step 3: work out the conc of HCl
concentration = moles / volume
= 0.0448/ (25/1000)
= 1.79 mol/dm^3

Example 2
What volume of 0.1 mol/dm^3 HCl is needed to neutralise 25.0cm^3 of 0.14 mol/dm^3 NaOH?

step 1: work out the number of moles of NaOH
Number of moles = conc x vol = 0.14 x 25.0/1000 = 0.0035 mol

step2 : use the balanced equation to work out the moles of HCl
$NaOH + HCl \rightarrow NaCl + H_2O$
1 mole NaOH reacts with 1 mole HCl **1:1 ratio**
So 0.0035 moles NaOH reacts with 0.0035 moles HCl

Step 3: work out the volume of HCl
volume= moles / concentration
= 0.0035/ 0.1
= 0.035 dm^3 or 35 cm^3

Electrolysis

When an ionic compound is melted or dissolved in water, the **ions are free to move** about within the liquid or solution. These liquids and solutions are able to conduct electricity and are called **electrolytes**.

Electrolysis is the passing an **electric current** through **ionic substances** that are **molten or in solution** to **breaks them down** into elements. Ions are discharged at the electrodes producing elements.

Passing an electric current through electrolytes causes the ions to move to the electrodes. **Positively charged ions** move to **the negative electrode** (the cathode), and **negatively charged ions** move to the **positive electrode** (the anode).

Electrolysis of molten salts

When a simple ionic compound is electrolysed in the molten state using inert electrodes, the salt splits and the metal ion moves to the negative electrode and the negative ion moves to the positive electrode.
e.g. if molten lead bromide is electrolysed, the lead will form at the negative electrode and bro**mine** will form at the positive electrode

At the negative electrode (cathode)

At the negative electrode, positively charged ions **gain electrons** to become metal atoms.
This is classed as reduction. **Reduction** is gaining electrons
$Na^+ (l) + e^- \rightarrow Na (s)$ (sodium ions become sodium atoms)
$Cu^{2+} (l) + 2e^- \rightarrow Cu (s)$
$Al^{3+} (l) + 3e^- \rightarrow Al (s)$

At the positive electrode (anode)

At the positive electrode, negatively charged ions **lose electrons.**
This is classed as **oxidation**. Oxidation is losing electrons
$2Cl^- (l) \rightarrow Cl_2 (g) + 2e^-$ (chlor**ide** ions becomes chlor**ine**)
$2Br^- (l) \rightarrow Br_2 (l) + 2e^-$ (brom**ide** ions becomes brom**ine**)
$2I^- \rightarrow I_2 + 2e^-$ (Iod**ide** ions becomes Iod**ine**)
$2O^{2-} (l) \rightarrow O_2 (g) + 4e^-$ (ox**ide** ions becomes ox**ygen**)

OIL RIG can help remember that **O**xidation **is L**oss of electrons: **R**eduction **is G**ain of electrons

Extraction by electrolysis

Metals can be extracted from molten compounds using electrolysis. Electrolysis is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon.

Aluminium is manufactured by the electrolysis of a molten mixture of aluminium oxide and cryolite. The **mixture** has a lower melting point than pure aluminium oxide.

Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current. This makes aluminium an expensive metal.

Aluminium forms at the negative electrode (cathode) and oxygen at the positive electrode (anode).

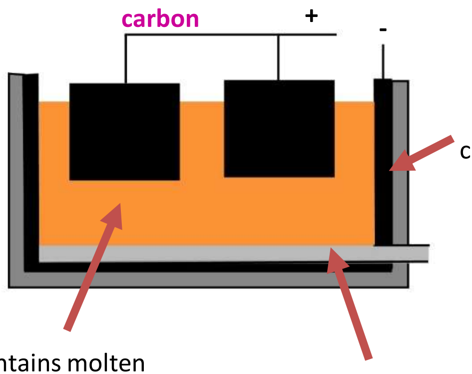
Half equation at negative electrode: $Al^{3+} (l) + 3e^- \rightarrow Al (s)$ (reduction)

Half equation at positive electrode $2O^{2-} (l) \rightarrow O_2 (g) + 4e^-$ (oxidation)

The **positive** electrode (anode) is made of **carbon**, which reacts with the **oxygen to produce carbon dioxide** and so must be continually replaced.

Cryolite is used in this process to **lower the melting point** of the **mixture** which therefore **reduces the energy costs**.

The anode is made from carbon



The cathode is made from carbon

The electrolyte contains molten aluminium oxide dissolved in molten cryolite.

The aluminium ions are attracted to the cathode, gain three electrons to form liquid aluminium.

Electrolysing aqueous solutions

If an aqueous solution is electrolysed, using inert electrodes, the ions discharged depend on the reactivity of the elements involved.

In aqueous solutions there is a mixture of ions: H^+ and OH^- ions are present in addition to the ions from the salt e.g. in copper chloride solution there are $\mathbf{H^+}$ **,**$\mathbf{OH^-}$ (from water) $\mathbf{Cu^{2+}}$ **,**$\mathbf{Cl^-}$ (from the salt)

The negative electrode

At the negative electrode (cathode) the **ions** of the **less reactive** element **discharges**.

In **aqueous** solutions of **reactive** metal salts where the metal is above hydrogen in the reactivity series, the metal will not therefore be evolved at the cathode. (e.g. **sodium** chloride, **calcium** fluoride) **Hydrogen gas** will be evolved at the cathode instead.

$2H^+ (aq) + 2e^- \rightarrow H_2 (g)$

In **aqueous** solutions of **unreactive** metal salts e.g. **copper** chloride or **silver** fluoride the metal will be evolved at the cathode.

$Cu^{2+} (aq) + 2e^- \rightarrow Cu (s)$

This happens because in the aqueous solution water molecules break down producing hydrogen ions and hydroxide ions that are discharged.

The positive electrode

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

$4OH^- (aq) \rightarrow O_2 (g) + 2H_2O (l) + 4e^-$ or

$4OH^- - 4e- \rightarrow O_2 + 2H_2O$

If a halide ion is present then the halogen is produced e.g.

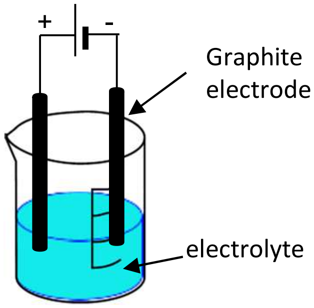
$2Cl^- (aq) \rightarrow Cl_2 (g) + 2e^-$

Oxygen is produced because **water molecules break down producing** hydrogen ions and **hydroxide ions**. The hydroxide ions are **attracted to the positive electrode** where they are **discharged to produce oxygen**

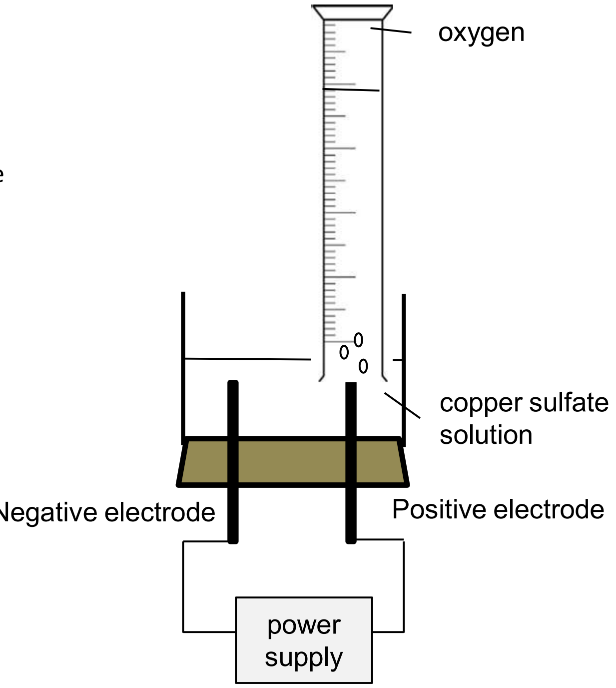
Careful with the term **discharge!** Discharge means an ion losing its charge at the electrode. Only ions can discharge. It is not correct to say elements or atoms discharge.
Do **not** say oxygen molecules are discharged at the electrode. Say **hydroxide ions are discharged** producing oxygen molecules

Required practical electrolysis

The products given off at the electrodes will depend on factors above



- If aqueous copper sulfate is used copper metal would form at the cathode (negative electrode) and the mass of the electrode would increase as copper forms on it.
- The blue colour of the electrolyte would fade as the copper ions are removed from the solution and reduced to copper atoms, so decreasing the concentration of copper ions in solution.
- Oxygen gas would form at the anode.



Volumes of gases produced can be measured using this apparatus