C4 Learning Journey

|  |  |
| --- | --- |
| 1 | **Metals** react with **oxygen** to produce **metal oxides**. The reactions are **oxidation** reactions because the metals gain oxygen. If a compound **loses oxygen** then that reaction is **reduction**. A general equation for oxidation of metals is  Metal + oxygen metal oxide |
| 2 | When **metals react**, the **metal atoms lose electrons** and form **positive ions**. The **reactivity** of a metal is related to its **readiness** to lose electrons and **form positive ions**. 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 observations of their reactions with water and dilute acids. Copper is unreactive with water and acid; iron and zinc are not reactive with water but slightly reactive with acid; magnesium, calcium, lithium, sodium and potassium are increasingly reactive, and are all more reactive with acid than water. |
| 3 | The **reactivity series** is a list of metals arranged in **order of reactivity** and is useful when making predictions of reactions. **Displacement** reactions happen if a metal higher in the reactivity series pushes a less reactive metal out of its compound and takes its place. The non-metals **carbon** and **hydrogen** are also included in reactivity series for reference. |
| 4 | **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. Metals **less reactive** than **carbon** can be extracted from their oxides by **reduction** with carbon. Reduction involves the heating the metal oxide with a source of carbon- the **carbon removes the oxygen** from the metal oxide and making **carbon dioxide**. **More reactive** metals need **electrolysis** to extract the metal from its oxide. |
| 5 | **Oxidation** and **reduction** can refer to the **loss and gain of electrons** as well as the loss and gain of oxygen. **Oxidation** is the **loss** of electrons, **reduction** is the **gain** of electrons. Oxidation and reduction happen at the same time. The acronym **OIL** **RIG** (oxidation is loss, reduction is gain) can be used to help remember the movement of electrons in oxidation and reduction. |
| 6 | **Acids** react with some **metals** to produce **salt** and **hydrogen**- the general word equation for this reaction is acid + metal salt + hydrogen. Magnesium, zinc and iron react with **hydrochloric acid** to make magnesium **chloride**, zinc chloride and iron chloride salts. Magnesium, zinc and iron react with **sulfuric acid** to make magnesium **sulphate**, zinc sulphate and iron sulphate salts. |
| 7 | **Acids** are **neutralised** by **alkalis** (eg **soluble metal hydroxides**) and **bases** (eg insoluble metal hydroxides and metal oxides) to produce **salts** and **water**. This reaction is called **neutralisation**. **Salt names** **have 2 parts**- the **firs**t part of the name comes from the **meta**l in the alkali or base. The **second** part of the name comes from the **acid** used- so hydrochloric acid makes chloride salts, sulphuric acid makes sulphate salts and nitric acid makes nitrate salts. |
| 8 | **Acids** are **neutralised** by metal **carbonates** to produce **salt**, **water** and **carbon** **dioxide**. The general word equation for this reaction is acid + carbonate salt + water + carbon dioxide. The name of the salt comes from the metal in the carbonate and the acid used. |
| 9 | Required practical - Salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates. Copper oxide is added to the warmed sulphuric acid until it no more reacts and the **excess** copper oxide is **filtered** off. The **salt solutions** is then **crystallised** to produce solid copper sulphate salt. |
| 10 | **Acids** produce **hydrogen ions (H+)** in aqueous solutions. **Alkalis** contain **hydroxide** **ions (OH–)**. 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 equation:  H+ (aq) + OH– (aq) → H2O (l). 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**. **Acid**s have pH values **0 (very strong) – 6 (weak**); substances with a **pH 7** are neutral**; alkalis** have **pH values 8 (weak) – 14 (very strong)** |
| 11 T | Required practical- The **volumes** of **acid** and **alkali** solutions that react with each other can be measured by **titration** using a suitable indicator. |
| 12 | **Acids dissociate/ ionise** in solution to release **H⁺ ions** and **alkalis dissociate/ionise** in solution to release **OH⁻ ions**. **Strong acids** (hydrochloric, sulphuric and nitric acids) and strong alkalis are almost **fully ionised in solution**; **weak acids** (citric, ethanoic and carbonic acids) and weak alkalis are **only slightly ionised** in solution. **pH** is a measure of the amount of **H⁺ ions** in solution. Each decrease of 1 on the pH scale is an increase in the amount of H⁺ ions by a factor of 10 |
| 13 | **Electrolysi**s is the splitting of an **ionic** compound using **electricity**. The compound must be **molten** or **dissolved** so that the **ions are free to move**. When the current passes through the **electrolyte** (the ionic compound in a molten/ dissolved state) the **positive ions (cations)** move to the **negative electrode (cathode)** and the **negative ions (anions)** move to the **positive electrode (anode)**. When the ions reach the electrodes, the ions become elements. Electrolysis cannot take place using solid ionic compounds because the ions are not free to move and carry the charge. |
| 14 | **Half equations** show the reactions at the **anode** and the **cathode** in terms of movement of **electrons.** For example, at the cathode X⁺ + e⁻ X and at the anode Y⁻ Y + e⁻ |
| 15 | **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**. Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current. |
| 16 | The **ions discharged** when an aqueous solution is **electrolysed** depends on the **relative reactivity** of the elements involved. At the **negative electrode** (cathode), **hydrogen** is produced if the **metal is more reactive** than hydrogen. If the metal is copper, then a thin coating of pinkish copper will be seen forming on the cathode. At the **positive electrode** (the anode), At the positive electrode (anode), **oxygen** is produced **unles**s the solution contains **halide ions** (F⁻, Cl⁻ or Br⁻) when the halogen is produced. This happens because in the aqueous solution water molecules also break down producing H⁺ ions and OH⁻ ions can be discharged. |
| 17 | Required practical- make predictions about the products of electrolysis and test them out. |
| 18 | During **electrolysis**, at the **cathode** (negative electrode), positively charged ions **gain electrons** and so the reactions are **reduction** reactions. The **number** of **electrons** gained is **equal** to the size of the **positive charge** on the cation. At the **anode** (positive electrode), negatively charged ions **lose electrons** and so the reactions are **oxidation** reactions. The **number** of electrons **lost** is **equal** to the size of the **negative charge** on the anion.  Reactions at electrodes can be represented by half equations, for example:  2H⁺ + 2e⁻ → H₂  and  4OH⁻ → O₂ + 2H₂O + 4e⁻  or  4OH⁻ – 4e⁻ → O₂ + 2H₂O |