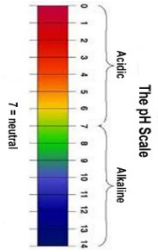
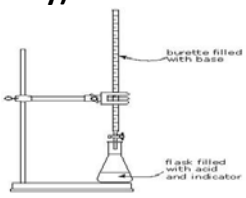
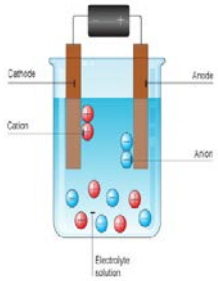
Chemical changes – Knowledge organiser

	Topic 1 - Reactivity of metals
4.4.1.1 Metal oxides	Metals react with oxygen to produce metal oxides. The reactions are **oxidation reactions** because the **metals gain oxygen**. **OILRIG – oxidation is loss (of electrons) Reduction is gain (of electrons**).
4.4.1.2 The reactivity series	• When metals react with other substances the **metal atoms form positive ions**. The reactivity of a metal relates to its tendency to 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 their **reactions with water and dilute acids**. The non-metals hydrogen and carbon are often included in the reactivity series. • Displacement: A **more reactive metal can displace a less reactive metal** from a compound. Potassium is the most reactive; it floats on water and fizzes (producing gas) when added to water. Sodium lithium, calcium, magnesium are comparatively less, while zinc and iron react slowly with cold water (it rusts) along with copper.
4.4.1.3 Extraction of metals and reduction	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** e.g. iron oxide + carbon → iron + carbon dioxide **Reduction involves the loss of oxygen. Oxidation is the gain of oxygen**
4.4.1.4 Oxidation and reduction in terms of electrons (HT only)	Oxidation is the loss of electrons and reduction is the gain of electrons (**OILRIG)**. This can be shown as **symbol equations**: **copper sulphate + iron → iron sulphate + copper** **$CuSO_4$ + Fe → $FeSO_4$ + Cu** Oxidation and reduction can also be shown as **half ionic equations** where **electrons** are shown as **e^-.** **Cu^{2+} + $2e^-$ → Cu** (copper has gained electrons, so is reduced) **Fe - $2e^-$ → Fe^{2+}** Iron has lost electrons, so is oxidised.
	Topic 2 - Reactions of acids
4.4.2.1 Reactions of acids with metals	Acids react with some metals to produce salts and hydrogen. Metal + **sulfuric** acid → metal **sulfate** + hydrogen Metal + **hydrochloric** acid → metal **chloride** + hydrogen (HT only) The metal loses electrons, so is oxidised.
4.4.2.2 Neutralisation of acids and salt production	**Acids** are **neutralised** by **alkalis** (eg soluble metal hydroxides) and bases (eg insoluble metal hydroxides and metal oxides) to produce salts and water, and by metal carbonates to produce salts, water and carbon dioxide. The particular salt produced in any reaction between an acid and a base or alkali depends on: • the acid used (hydrochloric acid produces chlorides, **nitric acid** produces **nitrates**, sulfuric acid produces sulfates) • the positive ions in the base, alkali or carbonate.

4.4.2.3 **Soluble salts**	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. **Required practical 1:** preparation of a pure, dry sample of a soluble salt from an insoluble oxide
4.4.2.4 the pH scale and neutralisation 	**Acids** produce **hydrogen ions (H+)** in aqueous solutions. Aqueous solutions of **alkalis** contain **hydroxide ions (OH).** 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 **< 7** and aqueous solutions of **alkalis** have pH values **> 7**. In **neutralisation** reactions between an acid and an alkali, hydrogen ions react with hydroxide ions to produce water. **H^+ (aq) + OH^- (aq) → H_2O (l)**
4.4.2.5 **Titrations (chemistry only)** 	The volumes of acid and alkali solutions that react with each other can be measured by **titration** using a suitable indicator. **Required practical 2:** Describe how to carry out titrations using strong acids and strong alkalis to find the reacting volumes accurately • (HT Only) calculate the chemical quantities in titrations involving concentrations in mol/dm^3 and in g/dm^3. **Concentration (moles / dm^3) = number of moles / volume in dm^3**
4.4.2.6 **Strong and weak acids (HT only)**	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. For a given concentration of aqueous solutions, the **stronger an acid**, the **lower the pH.** As the pH decreases by one unit (order of magnitude), the hydrogen ion concentration of the solution increases by a factor of 10.
4.4.3	**Topic 3 - Electrolysis**
4.4.3.1 **The process of 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**. 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**). **Opposites charges attract**. Ions are discharged at the electrodes producing elements. This process is called **electrolysis.** (HT only) half equations: Eg Cu^{2+} (ion) + 2e → Cu (atom)
4.4.3.2 **Electrolysis of molten ionic compounds**	When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes, the metal (lead) is produced at the cathode and the non-metal (bromine) is produced at the anode. **Metals** for **positive ions** which are attracted to the **negative electrode (cathode)**.
4.4.3.3	Metals can be extracted from molten compounds using electrolysis.

Using electrolysis to extract metals	**Electrolysis** is used if the metal is **too reactive** to be extracted by reduction with carbon or if the metal reacts with carbon. **Large amounts of energy** are used in the extraction process to **melt** the compounds and to **produce the electrical current**. **Aluminium** is manufactured by the electrolysis of a molten mixture of **aluminium oxide** and **cryolite** (helps lower the melting point, to save energy) using carbon as the positive electrode (anode). Oxygen is formed at the positive electrode, and reacts with it to produce **carbon dioxide**. The electrode wears away and needs to be replaced.
4.4.3.4 Electrolysis of aqueous solutions Putting it together The table shows some common ionic compounds, and the elements released when their solutions are electrolysed. **ionic substance in solution** / **element at the negative electrode** / **element at the positive electrode** copper chloride, $CuCl_2$ / copper / chlorine copper sulfate, $CuSO_4$ / copper / oxygen sodium chloride, NaCl / hydrogen / chlorine hydrochloric acid, HCl / hydrogen / chlorine sulfuric acid, H_2SO_4 / hydrogen / oxygen	The ions discharged (given out) when an aqueous solution is electrolysed using inert (unreactive) electrodes depend on the reactivity of the elements involved. At the **negative electrode** (cathode), **hydrogen** is produced if the **metal is more reactive** than hydrogen. E.g is the metal is sodium. At the **positive electrode** (anode), **oxygen** is produced unless the **solution contains halide ions** when the halogen is produced. This happens because in the aqueous solution **water** molecules **break down** producing **hydrogen ions** and **hydroxide ions** that are discharged. **Required practical 3**
4.4.3.5 representation of reactions at electrodes as half equations (HT only)	During electrolysis, at the cathode (negative electrode), positively charged ions gain electrons and so the reactions are reductions. At the anode (positive electrode), negatively charged ions lose electrons and so the reactions are oxidations. Reactions at electrodes can be represented by half equations, for example: $2H^+ + 2e^- \rightarrow H_2$ and $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$