Meden School Curriculum Planning							
Subject	Chemistry	**Year Group**	10	**Sequence No.**	16	**Topic**	Chemical changes C4

Retrieval	Core Knowledge	Student Thinking
What do teachers need to **retrieve** from students before they start teaching **new content**?	What **specific ambitious knowledge** do teachers need teach students in this sequence of learning?	What real life examples can be applied to this sequence of learning to **development of our students thinking, encouraging them to see the inequalities around them** and 'do something about them!'
KS3 Year 7: Acids and alkalis topic. Students were introduced to the pH scale. The numbers it runs to and from and the colours it produces with universal indicator. **KS3 Year 7: Acids and alkalis topic.** Students were introduced to the general equations for a neutralisation reaction and the reaction between a metal carbonate and acids. Naming salts was also covered. **KS2:** Students learnt about filtering and evaporation. **KS3 Year 7: Separation techniques.** Students were re-introduced to the process of evaporation and filtration in this topic.	**L1: Acids/alkalis.** The pH scale is a scale which measures the **acidity or alkalinity** of a substance. It runs from the numbers 0-14 with 0-6 being acids, 7 neutral and 8-14 alkaline. The lower the number, the more acidic the substance. Neutral substances are neither acidic or alkaline. To test for pH, you can use an indicator or a pH probe/sensor. A **pH probe** is used to give a numerical value and are more **accurate** than an indicator. The pH scale is based on the colours produced from universal indicator. A wide range indicator is a dye that gradually changes colour over a broad range of pHs. These are useful for estimating the pH of a solution. All acids, when dissolved in water, form **H^+ ions** and alkalis the **OH^- ion**. **Bases** are substances that neutralise acids. Examples include copper oxide and zinc oxide. Alkalis are bases which dissolve in water, examples include sodium hydroxide and potassium hydroxide. **L2: Reactions of acids.** When a metal oxide or hydroxide reacts with an acid, a **neutralisation reaction** occurs where a salt and water are produced. To name this salt, the first part of the name comes from the metal and the second part from the acid. If hydrochloric acid is used, a **chloride** salt is formed, if nitric acid is used then a **nitrate** salt is formed and if sulfuric acid is used then a **sulfate** salt is formed. For example, if hydrochloric acid was reacted with copper oxide then the salt formed would be copper chloride. Another example of a neutralisation reaction is when a metal carbonate is reacted with an acid. The products of this reaction are a salt, carbon dioxide and water. **L3/L4: Making soluble salts (required practical).** When an **insoluble** reagent such as a metal oxide or metal carbonate is added to an acid, a soluble salt is produced. The acid needs to be warm in order to speed up the reaction but not boiling so that the acid does not spit out of the beaker. When the insoluble reagent is added, we say it is added **'in excess'**. It is added in small quantities at a time and stirred until no more will dissolve. At this point the acid has been neutralised and the reaction has finished. The excess insoluble reagent is then filtered to obtain the salt solution. To obtain the crystals, **evaporation**	L1: Acids and alkalis are found in our daily lives for things such as cleaning and cooking. We also eat and drink acidic/alkaline substances. In the ocean, acids and bases are even more critical. Molluscs in the ocean rely on certain chemicals to build their shells. Sharks rely on a specific pH in water for their hypersensitive noses. As humans produce more carbon dioxide from fossil fuels, some of it ends up in the ocean (up to 30% of the total carbon dioxide released) — where it acidifies the water. The UN has a sustainable development goal to minimise and address the impacts of ocean acidification. Many careers use the knowledge of pH such as an environmentalist, manufacturing, pool maintenance, doctors, nurses, food scientists and technologists.

KS4 Year 10: C1 Atomic structure and the periodic table. Students were reminded about the process of evaporation and introduced to crystallisation in this topic.	and **crystallisation** are used. The salt solution is gently heated until some of the water is evaporated, the solution becomes more concentrated and it is then left to cool. The resulting crystals are then dried.	
KS3 Year 7: Acids and alkalis topic. Students were introduced to some general properties of acids and alkalis including the terms strong and weak.	**L5: Strong and weak acids (higher only).** Strong acids such as sulfuric acid, hydrochloric acid and nitric acid completely ionise in water. This means that all the acid particles release H^+ ions. To write an equation, an arrow must be used to show this **dissociation**. For example, nitric acid dissociates into hydrogen ions and nitrate ions: $HNO_{3(l)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$. Weak acids such as carboxylic acids, carbonic acids and citric acid only **partially ionise** in water. An example is the dissociation of ethanoic acid: $CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-{}_{(aq)}$. Acid strength will have an effect on pH and reactivity, a strong acid will be more reactive as there will be a greater number of H^+ ions, this also leads to having a lower value of pH as strong acids are more acidic. For every decrease of 1 on the pH scale, the concentration of H^+ ions increases by 10. The concentration of a solution tells us how many of the dissolved acid molecules there are in a certain volume of water. You can have a dilute and concentrated solutions.	
KS3 Year 7: Acids and alkalis topic. Students were introduced to the general equation for a reaction between a metal and acid. **KS3 Year 8: Reactivity of metals topic.** Students recapped naming salts, the products from when an acid reacts with a metal and were introduced to the practical to investigate the reactivity of metals.	**L6: Reactivity series and reaction of metals.** The reactivity series is a list of metals with the most reactive being at the top and the least at the bottom. It also includes **carbon and hydrogen** as they are used to compare where metals are in relation to them which is then later used to determine the method of extraction. The reactivity of a metal is defined as how easily it will **lose electrons and form positive ions** rather than how it will react with water, acid or oxygen. When metals react with acids, a salt and hydrogen are formed. To determine the order of reactivity of metals from experiments, the rate of hydrogen production can be monitored. The faster the rate, the more reactive the metal. Some metals will react explosively while others will be less violent. Copper won't react with cold, dilute acids. When metals react with water, both hydrogen gas and a metal hydroxide are produced.	
KS3 Year 8: Reactivity of metals topic. Students were introduced to how displacement reactions work. **KS4 Year 10: C1 and C8 topic.** Students were re-introduced to displacement reactions for the group 7 elements.	**L7: Displacement reactions.** (Practical) When a more reactive metal is placed in the solution of a dissolved metal compound containing a less reactive metal ion, the more reactive metal will **displace** the less reactive metal from its compound. For example, iron will displace copper from a solution of copper sulfate to produce iron sulfate and copper. The reactivity series can be used to predict if a reaction will occur. **L8: Extraction of metals**. Metals can be extracted from metal ores (rocks which contain enough metal to make it profitable to extract the metal from it) using reduction. The metal is usually found as a metal oxide, **reduction** is defined as the loss of oxygen from a compound and **oxidation** is the gain of oxygen by	L7: Some applications of displacement reactions are thermite welding, steel making, extraction of metals, and relief from acid indigestion.

	an element or compound. If the metal is below carbon in the reactivity series, it can be extracted using reduction with carbon. The metal oxide is reduced as it will lose oxygen and the carbon will be oxidised as it gains oxygen. Any metal above carbon in the reactivity series has to extracted using **electrolysis**. **L9/L10: Redox reactions (higher only).** Reduction and oxidation have another other definition in terms of electrons. Oxidation is the loss of electrons and reduction is the gain of electrons. The term OILRIG is used to remember these definitions. Redox is the term used to describe that both reduction and oxidation occur at the same time and displacement reactions are an example of redox reactions. In displacement reactions, the metal ion will gain electrons and is therefore reduced and the metal atom loses electrons and is oxidised. Redox reactions also occur between metals and acids. The metal atoms lose electrons (oxidised) and the hydrogen ions gain electrons (reduced). To show redox reactions, **ionic equations and half equations** can be written. Ionic equations only show the particles that react and the products they form. Half equations show the reduction and oxidation of individual species and the electrons that are transferred. Example: When zinc is placed in a solution of copper sulfate Ionic equation: $Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$ Half equations: $Zn \rightarrow Zn^{2+} + 2e^-$ (Oxidation) $Cu^{2+} + 2e^- \rightarrow Cu$ (Reduction)	
KS3 Year 8: Reactivity of metals topic. Students were introduced to the basics of electrolysis, that at the negative electrode a metal is produced and at the positive electrode, a non-metal will be produced. Students also learnt about metal ores and that bauxite the main ore used to extract aluminium from aluminium oxide.	**L11/L12/L13: Electrolysis (required practical).** To extract a more reactive metal than carbon, electrolysis must be used. An electrical current is passed through an aqueous or molten ionic compound, the electrical current then splits up the **ionic compound** into its separate ions. The **aqueous or molten ionic compound** is called an **electrolyte**, the free ions will move to the oppositely charged solid **electrode**s. The positive ions (called **cations**) will move towards the negative electrode (called a **cathode**) and the negative ions (**anions**) will move towards the positive electrode (**anode**). The term PANIC is used to remember the names of these electrodes. The ions will lose or gain electrons and become atoms or molecules, these are the products of electrolysis. When looking at the electrolysis of metal ores, in particular the extraction of aluminium from aluminium oxide, the **bauxite** ore is used. Aluminium oxide is dissolved in molten **cryolite** to reduce the melting point (making it cheaper and easier). Carbon electrodes are used and need to be replaced regularly as the oxygen produced will react with the carbon to produce carbon dioxide which gradually wears away the electrode. When looking at aqueous solutions, there are rules that need to be followed to predict the products formed. When an ionic compound is dissolved in water, H^+ and OH^- ions are also in the solution along with the ions from the ionic compound. At the negative electrode, there will be metal ions and H^+ ions, the least reactive element will be produced. At the positive electrode, if halide ions are present (Cl^-, Br^-, I^-) then chlorine, bromine or iodine will be produced. If no halide ions are present, the hydroxide ions (OH^-	L11: Electrolysis is used extensively in metallurgical processes such as extraction or purification of metals from ores or compounds and in deposition of metals from solutions (electroplating). In America, they have a 'hydrogen energy earthshot' initiative which is looking at more abundant, affordable and reliable clean energy. Electrolysis is used to split water into hydrogen and oxygen and it can result in zero greenhouse gas emissions. It is being looked at as an option for carbon free hydrogen production from renewable and nuclear resources (therefore fewer fossil fuels which lead to climate change).

	) will produce oxygen and water. If neither of these ions are present, then other negative ions will be discharged. In molten ionic substances, only the ions from the ionic compound will be present and so the neutral metals and neutral non-metal elements will be produced. **(Higher only).** Redox reactions occur at the electrodes. Reduction occurs at the negative electrode as the ions gain electrons and oxidation occurs at the positive electrode as the ions lose electrons. Half equations can be written to show the reactions occurring at the electrodes. **L14: Revision** **L15: EOTT** **L16: GPA**	The laws of electrolysis are attributed to the English scientist Michael Faraday. He also popularised the terms anode, cathode, electrode and ions which were actually proposed by a different scientist, William Whewell. Electrolysis was invented in 1800 by William Nicholson and Anthony Carlisle using voltaic current. They had replicated Alessandro Voltas voltaic pile.