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**Year 10 – Knowledge Booklet TRILOGY**

Key Stage 4 Science:

**Chemical Changes**

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**Ensure that your booklet is returned to your class book box at the end of the lesson.**

**Lesson Breakdown**

Lesson 1: 4.4.1.1 Metal oxides (practical – pattern seeking enquiry)

Lesson 2: 4.4.1.2 The reactivity series (practical – pattern seeking enquiry)(water and acids)

4.4.2.1 Reactions of acids with metals

Lesson 3: 4.4.1.2 Displacement reactions

Lesson 4: 4.4.1.3 Extraction of metals and reduction

Lesson 5: 4.4.2.4 The pH scale and neutralisation & 4.4.2.6 Strong and weak acids (HT only)

Lesson 6: 4.4.2.2 Neutralisation of acids and salt production

Lesson 7: 4.4.2.3 Soluble salts & **Required practical:** Preparing a salt (practical – reinforcing theory)

Lesson 8: 4.4.3.1 The process of electrolysis & 4.4.3.2 Electrolysis of molten ionic compounds

Lesson 9: 4.4.3.3 Using electrolysis to extract metals

Lesson 10: 4.4.3.4 Electrolysis of aqueous solutions

Lesson 11: **Required practical 3**: Electrolysis of aqueous solutions

***Embedded in several lessons: 4.4.1.4 Oxidation and reduction in terms of electrons (HT only)***

***Embedded in lesson 8 – 11: 4.4.3.5 Representation of reactions at electrodes as half equations (HT only)***

**Keystone words**

Oxidise

Reduce

Atom

Element

Ion

Neutralisation

Solution

**Lesson 1: 4.4.1.1 Metal oxides**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24 | In chemistry, substances react when:   1. They interact with other substances and form new products. For example, hydrogen reacts with oxygen to make water. Hydrogen and oxygen are the reactants and water is the product. 2. A molecule breaks up into smaller molecules; we call this decomposition. For example, calcium carbonate decomposes to form calcium oxide and carbon dioxide.   In all cases, the atoms rearrange to form new substances. Electrons from the outer shell of atoms may be given, taken or shared differently after the reaction has occurred (see Structures and Bonding).  Some substances react more readily than others; they are more reactive. For example, potassium is more reactive than lithium metal. Potassium has a higher reactivity.  Substances that have a higher reactivity also have a greater tendency to form ions. For example, group 1 metals form positive ions. However, potassium has a greater tendency to form ions because it is more reactive. This is because the attraction between the nucleus and outer shell electrons is weaker in more reactive substances.  b) In this lesson you will learn about oxidation and reduction reactions. You will also use examples of these to put three metals in order of reactivity.  Oxidation reactions happen when oxygen is added to a substance. For example, when magnesium reacts with oxygen, magnesium oxide is produced. |  |

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| 1  2 | One way of representing this is shown in the Connect task. We can also use word and symbol equations and also show the particles:  Diagram  Description automatically generated |  |
| 1  2  3  4 | Reduction reactions are the opposite of oxidation reactions. Oxygen is removed from a substance.  There is another definition of oxidation and reduction that you will come across in this topic.  2 - Finding the formula of Water - IGCSE Chemistry |  |

**Lesson 2: 4.4.1.2 The reactivity series**

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| 1  2  3  4  5  6  7  8  9  10  12  13  14  15  16  17  18  19  20  21  22  23  24  25  26 | **Connect**  Writing and drawing electronic configuration  Take sodium as an example. Sodium has 11 electrons. It has:   * 2 electrons in the first shell * 8 electrons in the second shell * 1 electron in the third shell   You can write sodium’s electronic configuration with numbers and commas: 2,8,1  The electronic configuration of sodium can be shown in a diagram. The dots represent electrons. Electrons may be shown using dots or crosses.    Elements “aim” to have a full outer shell. They can do this in a number of ways. One method is for atoms to gain or lose electrons.  Because atoms cannot change the number of protons that are in the nucleus this means they would have a different number of protons to electrons. If protons and electrons are not balanced then there must be a charge.  They can gain a full outer shell by losing electrons and becoming positive or gaining electrons and becoming negative. We call these charged particles ions.  Figure 1 shows how we draw the electrons on an ion. Again we use the dot and cross system. Use one symbol for the electrons that the element usually has, and another symbol for any extra electrons. (If the element has lost electrons, don’t worry about the second symbol!)  At GCSE we will only ever have to draw dot and cross diagrams for the first 20 elements    **Figure 1** |  |
| 1  2  3  4  5  6  7  8  9  10  11  12 | When metals react with water there are several indications that a chemical reaction is happening. A gas is produced, heat is generated (K and Na melt into a ball) and ‘smoke’ appears. The pH of the water in the trough also changes. The high pH tells us that an alkali is being made.  Group 1 metals with water to make metal hydroxides and hydrogen gas.  Other metals tend to form metal oxides and hydrogen gas.  Metal oxides and hydroxides contain positive metal ions; the metal has been oxidised because it has lost electrons.  With metal oxides, we can also say the metal is oxidised because the metal gains oxygen. |  |
| 1  2 | Metals react with acids more readily than they do with water.  When metals react with acids they make a salt and hydrogen gas. |  |

**Lesson 3: 4.4.1.2 Displacement reactions**

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| 1  2 | **Connect**  See Periodic Table at end of booklet. |  |
| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16 | In the previous two lessons we used experimental observations to put metals in order of reactivity. A more complete Reactivity Series exists that contains more metals. It also contains carbon and hydrogen; these non-metals are included because scientists often use them in displacement reactions.  Displacement reactions are reactions in which a more reactive element displaces a less reactive element from a compound. For example, potassium can displace sodium from sodium chloride.  **Potassium + sodium chloride 🡪 potassium chloride + sodium**  It is always the case that the more reactive element displaces the less reactive one.  It is also always the case that the bigger the difference in reactivity, the more vigorous the reaction. So, using potassium to displace copper from copper chloride would be a much more vigorous reaction than the previous one.  **The Reactivity Series**  Reactivity Series of Metals | Secondary Science 4 All |  |

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21 | **(HT only)** Displacement reactions can be represented by ionic equations. These are similar to symbol equations. However, they only show changes to ions and atoms.  There are several steps needed if you are to successfully write ionic equations. The first is to write a balanced symbol equation. The atoms ions associated with the reactants and products are then written out separately.  We only want to see changes to atoms and ions, so we can cancel out any ions that do not change. This leaves us with the final ionic equation.  **Writing ionic equations:**  **Step 1:**Write out the full balanced equation:  2KI (aq) +  Cl2 (aq) → 2KCl (aq) + I2 (aq)  **Step 2:**Identify the ionic substances and write down the ions separately  2K+(aq) + 2I- (aq) +  Cl2 (aq) → 2K+(aq) + 2Cl-(aq) + I2 (aq)  **Step 3:**Rewrite the equation eliminating the ions which appear on both sides of the equation (spectator ions ) which in this case are the K+ ions:   2I- (aq) +  Cl2 (aq) → 2Cl-(aq) + I2 (aq)  **Taken from** [**https://www.savemyexams.co.uk/gcse/chemistry/ocr-gateway/18/revision-notes/3-chemical-reactions/3-1introducing-chemical-reactions/3-1-4-writing-balanced-ionic-equations/**](https://www.savemyexams.co.uk/gcse/chemistry/ocr-gateway/18/revision-notes/3-chemical-reactions/3-1introducing-chemical-reactions/3-1-4-writing-balanced-ionic-equations/) |  |
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**Lesson 4: 4.4.1.3 Extraction of metals and reduction**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15 | One of the most important industrial uses of displacement reactions is the extraction of metals from ores. Ores are rocks that contain enough metal, or a compound of a metal, to make extracting the metal economically viable; in other words, somebody can make a profit from it.  Some metals are found as pure metals in the ground. For example, gold and silver. This is because they are unreactive and so do not react with water, oxygen etc. Physical methods are used to extract them from the ground. For example, sieving and washing.  Metals that are below carbon in the reactivity series can be extracted by reacting them with carbon. For example, iron ore contains iron oxide. When iron oxide is reacted with carbon, the carbon displaces the iron.  Metals that are above carbon in the reactivity series cannot be extracted by reacting them with carbon. This is because the carbon is not reactive enough to displace the metal. These metals are extracted using electrolysis. We learn about this at the end of the topic. |  |
| 1  2  3  4  5  6  7  8 | Several metals have an extraction process that uses displacement with carbon; the metal is reduced because oxygen is removed (and the ions gain electrons). These include extracting iron, tungsten and titanium. Often the extraction procedure has multiple steps and only some of them use carbon.  Typically, the displacement reaction produces carbon dioxide:  Metal oxide + carbon  metal + carbon dioxide  MO + C 🡪 M + CO2 |  |

**Lesson 5: 4.4.2.4 The pH scale and neutralisation & 4.4.2.6 Strong and weak acids (HT only)**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24  25 | **Connect**  **Universal indicator and the pH scale**  Universal indicator is supplied as a solution or as universal indicator paper. It is a mixture of several different indicators. Unlike litmus, universal indicator can show us how strongly acidic or alkaline a solution is, not just that the solution is acidic or alkaline.  This is measured using the pH scale, which is a continuous coloured number scale measuring acids from below zero to alkali's above fourteen.  Universal indicator has many different colour changes, from red for strongly acidic solutions to dark purple for strongly alkaline solutions. In the middle, neutral pH 7 is indicated by green.  A coloured pH scale, ranging from dark red at pH0 and green at pH7, to dark purple at pH14.  Colour chart of universal indicator colours at different pH values  When you use universal indicator paper, you get more accurate results if you only put a small spot of the test solution on the paper, and then leave the colour to develop for about 30 seconds before comparing it with the colour chart.  These are the important points about the pH scale:  neutral solutions are pH 7 exactly  acidic solutions have pH values less than 7  alkaline solutions have pH values more than 7  the closer to pH 0 you go, the more strongly acidic a solution is  the closer to pH 14 you go, the more strongly alkaline a solution is |  |
| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24  25  26  27  28  29  30  31  32  33  34  35  36  37  38 | You learnt about the pH scale at KS3. You used it alongside Universal Indicator to see if different household substances were acidic, alkali or neutral.  The pH scale is used with aqueous solutions; acids and alkalis that are dissolved in water. pH 7 I neutral. Number slower than this are acidic; the lower the number, the stronger the acid (for a given concentration). Numbers above 7 are alkaline. The higher the number the stronger the alkali (for a given concentration).  All acids have one thing in common; they release hydrogen ions. The more hydrogen ions they release, the more acidic they are. Alkalis release hydroxide ions. The more hydroxide ions they release the stronger the alkali.  The pH number tells you the relative number of hydrogen ions compared to another pH. For example, pH3 has ten times the concentration of hydrogen ions as pH4. pH3 has 100 times (10 x 10) the concentration of hydrogen ions as pH5.  Good opportunity to use Show Me followed by Cold Call to see if students understand that 1 pH unit represents an order of magnitude difference in concentration.  The reaction between acids and alkalis are a type of neutralisation reaction. In these reactions the hydrogen ions from the acid react with the hydroxide ions from the alkali. This reaction forms water:  **H+(aq) + OH—(aq) 🡪 H2O(l)**  The pH changes because hydrogen ions and hydroxide ions are removed from the solution.  d) Acids can be classified as strong or weak acids. The strength of an acid is determined by the extent to which they ionise. Strong acids fully ionise. This means that every acid molecule releases hydrogen ions. For example:  **HCl(aq) 🡪 H+(aq) + Cl-(aq).**  Weak acids only partially ionise. This means that only a tiny percentage of acid molecules release hydrogen ions at any one time.  Acids can also be concentrated or dilute. This depends upon how much water the acid is dissolved in. If there is a lot of water, the acid is dilute (low concentration). The acid is more concentrated if it is dissolved in less water.  Sulphuric, nitric and hydrochloric acid are strong acids. Ethanoic, citric and carbonic acids are weak acids. |  |

**Lesson 6: 4.4.2.2 Neutralisation of acids and salt production**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24 | Neutralisation reactions occur when an acid is reacted with a base or alkali. Bases are molecules that can accept hydrogen ions and so neutralise acids. Alkalis are bases that dissolve in water.  The pH of a solution increases when a neutralisation reaction happens because the hydrogen ion concentration in the solution decreases.  Bases, including alkalis, react with acids to make a salt and water. A salt is a compound containing a metal ion and a non-metal ion. They also make water.  **Metal oxide + acid 🡪 salt + water**  **Metal hydroxide + acid 🡪 salt + water**  Metal carbonates are also bases. They react with acids to produce a salt, water and carbon dioxide.  Metal carbonate + acid 🡪 salt + water + carbon dioxide  There are many different salts. The salt made in a neutralisation reaction depends upon the acid used and the metal in the base.  The names of bases have two parts. The first part of the name comes from the metal in the base. The second part is determined by the acid used.   * Nitric acid always produces a metal nitrate. * Sulphuric acid produces a metal sulphate. * Hydrochloric acid produces a metal chloride.   For example, if copper oxide neutralises sulphuric acid, the salt made is copper sulphate. Water is also made.  **Copper oxide + sulphuric acid 🡪 copper sulphate + water.** |  |
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**Lesson 7: 4.4.2.3 Soluble salts & required practical: Preparing a salt**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24  25  26 | **Method:**  1. Measure 40 cm3 sulfuric acid and put it into the 100 cm3 beaker.  2. Set up the Bunsen burner, tripod, gauze and heatproof mat. Put the beaker on the gauze and heat the acid gently until it is almost boiling. Turn off the Bunsen burner.  3. Remove the glass beaker from the tripod. Use the spatula to add a small amount of copper (II) oxide powder to the hot acid. Stir with the glass rod. The copper (II) oxide will disappear and the solution will turn clear blue.  4. Add some more copper (II) oxide and stir again.  5. Keep adding the copper (II) oxide until some of it remains after stirring.  6. Allow the apparatus to cool completely.  7. Set up the filter funnel and paper over the conical flask. Filter the contents of the beaker.  8. Pour the filtrate from the conical flask into the evaporating basin.  9. Set up a water bath using the 250 cm3 beaker on the tripod and gauze.  10. Evaporate the filtrate gently using the water bath.  11. When crystals start to form, stop heating the water bath.  12. Pour the remaining solution into the crystallising dish.  13. Leave the crystallising dish in a cool place for at least 24 hours.  14. Remove the crystals from the concentrated solution with a spatula.  Gently pat the crystals dry between two pieces of filter paper.  **Taken from** [**https://www.bbc.co.uk/bitesize/guides/zf73382/revision/1**](https://www.bbc.co.uk/bitesize/guides/zf73382/revision/1) |  |
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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23 | **Method**  1. Use the pipette and pipette filler to put exactly 25 cm3 sodium hydroxide solution into the conical flask.  2. Put the flask on a white tile.  3. Clamp the burette vertically in the clamp stand. There should be just enough room underneath for the conical flask and tile.  4. Close the burette tap.  5. Use the small funnel to carefully fill the burette with dilute sulfuric acid. Before it completely fills put a small beaker underneath the tap, gently open it to allow acid to fill the tap, before closing again and filling the burette to the 0.00 cm3 line. Remove the funnel.  6. Put 5–10 drops of phenolpthalein indicator into the conical flask. Swirl the flask to mix and put under the burette on top of the tile. The contents of the flask will go pink.  7. Carefully open the burette tap so that 10 cm3 sulfuric acid slowly flows into the flask. Constantly swirl the flask when adding the acid. Then add the acid drop by drop until you see a permanent colour change from pink to colourless in the flask.  You need to be able to shut the tap immediately after a single drop of acid causes the colour to become permanently colourless.  8. Read the burette scale carefully and record the volume of acid you added to 2dp.  9. Repeat steps 1‒8 twice more and record the results in the table. |  |
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**Lesson 8: 4.4.3.1 The process of electrolysis & 4.4.3.2 Electrolysis of molten ionic compounds**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24  25  26  27  28  29  30  31  32  1  2  3  4  5  6 | In a previous lesson we learnt about how we extract metals from their ores. The method used depended on how reactive the metal was. Unreactive metals are found as pure metals in the ground; physical methods are used to extract the metal.  Metals that are less reactive than carbon can be displaced from their ore by reacting them with carbon.  Reactive metals cannot be extracted by reacting them with carbon; carbon is not reactive enough to displace the metal. Electrolysis is used when a reactive metal is being extracted.  In the next three lessons you will learn about:  a. Electrolysis of single substances.  b. Using electrolysis to extract metals.  c. Electrolysis of aqueous solutions.  The theory is very similar for all of these situations.  Electrolysis uses electricity to separate ions in ionic substances and then oxidise or reduce them so that they become electrically neutral.  For this to happen, the ions need to be able to move. This can be achieved in two ways:  a. Melt the ionic substance  b. Dissolve the ionic substance  Two rods (electrodes) are usually suspended in the liquid. These are made of a conductor. They are attached to a power supply so that one electrode becomes positively charged (anode) and one becomes negatively charged (cathode). The ions will migrate towards the electrode with the opposite charge.  At the positive electrode, the negative ions give electrons to the electrode; the ions are oxidised to form atoms.  At the negative electrode, the positive ions gain electrons; the ions are reduced to form atoms.  When a molten compound is used it is easy to predict the products of electrolysis. At the negative electrode the positive ion is reduced to give atoms / molecules of the element; the positive ion gains electrons.  At the positive electrode the negative ions are oxidised; they donate electrons to the electrode.  The reactions that occur at the electrodes can be represented using half equations. These equations show ions gaining or losing electrons. These equations need to be balanced just like other symbol equations. However, the charges also need to be balanced. |  |
|  | Examples of apparatus used in electrolysis:  Required Practical Review      C6 - Extraction of Aluminium | Teaching Resources |  |

**Lesson 9: 4.4.3.3 Using electrolysis to extract metals**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14 | The best example of using electrolysis to extract a metal is extraction of aluminium from aluminium oxide.  The melting point of aluminium oxide is extremely high. A substance called cryolite is added to the aluminium oxide; this reduces the melting point of aluminium so that a lower temperature can be used to melt aluminium oxide.  The aluminium ions are positively charged so they migrate towards the cathode. Electrons are then donated to the ions so they are reduced and form aluminium metal.  The oxygen ions migrate towards the positive electrode (anode). The oxygen ions are oxidised and form oxygen gas.  The electrodes have to be replaced regularly. This is because the electrodes are made of carbon (graphite). This is because the carbon in the electrodes reacts with the oxygen to form carbon dioxide.  C6 - Extraction of Aluminium | Teaching Resources |  |
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**Lesson 10: 4.4.3.4 Electrolysis of aqueous solutions**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17  18  19  20  21  22  23  24  25  26  27  28  29  30  31  32  33  34 | **Connect**  **Common ions:**  Al3+ Pb2+ Fe2+  Cu2+ Zn2+ Mg2+ Ca2+ Li+  Na+ K+ Ag+ H+  SO42- CO32- O2- NO3- OH- Br- Cl-  Electrolysis of molten substances is very predictable because only two ions are present in the electrolyte. It can also be use with every ionic substance; it can be used to extract even the most reactive metals.  The disadvantage is that it is very expensive. This is because it the salt needs to be melted; there are heating costs. The electricity used in electrolysis is a second cost (an inevitable one).  The alternative is to use electrolysis of aqueous solutions; the ions can move because they are dissolved in water. An advantage of this is that you don’t have to pay the costs associated with melting the substance.  A disadvantage is that there are four ions in the solution (two from water and two from the salt). This makes the chemistry more complex because two positive ions compete for electrons and two negative ions compete to lose electrons.  Fortunately, three are simple rules for predicting the outcome of electrolysis of aqueous solutions.  There are usually four ions in aqueous solutions.  1. Hydrogen ions from water (positively charged).  2. Hydroxide ions from water (negatively charged).  3. Positive ions from the salt.  4. Negative ions from the salt.  The products made at each electrode can be predicted if you follow some simple rules:  Cathode (negative electrode): Hydrogen will be produced unless the metal ion is for metal less reactive than hydrogen.  Anode (positive electrode): If a group 7 ion is present, a group 7 element will be produced. If a group 7 ion is not present, oxygen is produced.  The half equations follow the same format as for electrolysis of molten substances. The half equation that produces oxygen from hydroxide ions is more complex: |  |
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**Lesson 11: Required practical 3: Electrolysis**

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| 1  2  3  4  5  6  7  8  9  10  11  12  13  14  15  16  17 | **Method**  1. Pour approximately 50 cm3 copper (II) chloride solution into the beaker.  2. Add the petri dish lid and insert the carbon rods through the holes. The rods must not touch each other.  3. Attach crocodile leads to the rods. Connect the rods to the dc (red and black) terminals of a low voltage power supply  4. Select 4 V on the power supply and switch on.  5. Look at both electrodes and record your initial observations in the table below.  6. Use forceps to hold a piece of blue litmus paper in the solution next to the anode (positive electrode) and identify the element?  7. Rinse the electrochemical cell apparatus and collect a new set of electrodes.  Repeat steps 1‒8 using the other solution sodium chloride and complete the following tasks to show your understanding of the chemistry of electrolysis. |  |
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