8464 AQA Combined Science

C4 Chemical Changes

What's the science story?

Understanding of chemical changes began when people began experimenting with chemical reactions in a systematic way and organizing their results logically. Knowing about these different chemical changes meant that scientists could begin to predict exactly what new substances would be formed and use this knowledge to develop a wide range of different materials and processes. It also helped biochemists to understand the complex reactions that take place in living organisms. The extraction of important resources from the earth makes use of the way that some elements and compounds react with each other and how easily they can be 'pulled apart'.

Previous knowledge:

KS3

Year 7 – Reactions 1 and Particles
Year 8 – Reactions 2
Year 9 – Reactions 3

KS4

C1 Atomic structure and the periodic table
C2 Bonding, structure an the properties of matter

Next steps…

KS4

C5 Energy changes
C6 The rate and extent of chemical change
C7 Organic chemistry
C8 Chemical analysis

Keywords

Reactivity	Ores	Hydrogen	pH scale	Electrolyte
Oxidised	OILRIG	Soluble	Indicator	**Concentrated**
Reduced	Ionic equation	Insoluble	Solution	**Dilute**
Reactivity	Ion	Crystallised	Cation	**Dissociate**
Displacement	Salt	Excess	Anion	**pH meter**
Electrons	Metal	Acid	Cathode	**Strong acid**
Reduction	Acid	Alkali	Anode	**Weak acid**
Oxidation	Ions	Neutral	Electrolysis	
Extract				

Working scientifically skills:

WS2 Draw/interpret diagrams
WS3 Make predictions
WS8 Following methods
WS10 Selecting and using equipment
WS11 Hazards
WS14 Drawing graphs
WS17 Making conclusions

Assessments:

Exit tickets x 2/3 (formative)

Details of each exit ticket

- **ET Making salts**
- **ET Electrolysis**

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Lesson No. and Title	Learning objectives	AQA Specification	Practical equipment
1. The reactivity series	4 – To recall the order of common metals in the reactivity series. 6 – To predict reactions of metals with oxygen, water and acid. 8 – To construct balanced symbol equations, including state symbols.	Metals react with oxygen to produce metal oxides. The reactions are oxidation reactions because the metals gain oxygen. Students should be able to explain reduction and oxidation in terms of loss or gain of oxygen.	DEMO: Alkali metals Trough, Li, Na, K, white tile, tweezers, scalpel, filter paper, UI. Gloves PRAC: Reactions ot metals with water and acid Spotting tiles, HCl 0.5M, pipettes, Mg, Zn, Fe, Cu
2. Displacement	4 – To define displacement. 6 – To explain why displacement reactions occur. 8 – To describe displacement reactions using ionic equations.	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 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. A more reactive metal can displace a less reactive metal from a compound.	PRAC: Displacement Spotting tiles, pipettes, Metals; Mg, Zn, Fe, Cu, Magnesium chloride, zinc chloride, iron chloride, copper sulphate Possible DEMO: Thermit reaction – See techs
3. Extraction of metals	4 – To describe the extraction of metals using carbon. 6 – To identify the species being reduced and oxidised in a reaction. 8 – To evaluate the extraction process to obtain a metal from its ore.	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 loss of oxygen.	PRAC: Extract copper using carbon Copper oxide, carbon, crucible, pipeclay triangle, spatula, matches, splints

4. Oxidation & Reduction HIGHER TEIR ONLY	4 – Describe oxidation and reduction in terms of electrons. 6 – Identify the species being reduced and oxidised in a reaction. 8 – Write ionic equations for displacement reactions.	Oxidation is the loss of electrons and reduction is the gain of electrons. Student should be able to: • write ionic equations for displacement reactions • identify in a given reaction, symbol equation or half equation which species are oxidised and which are reduced.	
5. Reactions of acids with metals	4 – To name salts that are formed with different acids. 6 – To describe how to make a salt using a metal and acid. 8 – To write ionic and half equations to describe reactions with acids.	Acids react with some metals to produce salts and hydrogen. (HT only) Students should be able to: • explain in terms of gain or loss of electrons, that these are redox reactions • identify which species are oxidised and which are reduced in given chemical equations. Knowledge of reactions limited to those of magnesium, zinc and iron with hydrochloric and sulfuric acids.	PRAC: Reactions or different acids and metals Test tubes, bungs, HCl 0.5M. H_2SO_4 0.5M, Mag, Zn, Fe, Cu, pipettes. Matches. Splints
6. Neutralisation of acids and salt production	4 – To describe neutralization reactions between acids and bases. 6 – To predict products from given reactants. 8 – To write balanced symbol equations for a range of reactions.	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.	PRAC: Making salts NaOH 0.4M, HCl 0.4M, H_2SO_4 0.4M, test tubes, UI, pipettes, measuring cylinder (10ml)

7. Soluble salts	4 – To describe how to make pure dry salts. 6 – To write word equations for the reactions. 8 – To write balanced symbol equations for a range of reactions.	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. Students should be able to describe how to make pure, dry samples of named soluble salts from information provided.	PRAC: Making soluble salts Calcium carbonate chips, magnesium carbonate power, sodium carbonate powder, boiling tube with bung and delivery tube, test tubes, lime water
8. REQ PRAC: Soluble salts	4 – To safely make a salt by reacting a metal carbonate with acid. 6 – To describe how to make a salt from a metal carbonate and acid. 8 – To describe neutralisation using ionic equations.	**Required practical activity 8:** preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate, using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution. AT skills covered by this practical activity: chemistry AT 2, 3, 4 and 6.	PRAC: Soluble salts 1.0 M dilute sulfuric acid, copper (II) oxide powder Spatula, glass rod, 100 cm3 beaker, filter funnel and paper, conical flask, 250 cm3 beaker, evaporating basin, crystallising dish
9. Neutralisation and pH scale	4 – To describe the pH scale. 6 – To describe how universal indicator can be used to classify a chemical. 8 – To explain how the pH of a solution changes as acid to alkali is added.	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 less than 7 and aqueous solutions of alkalis have pH values greater than 7. 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: $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$	PRAC: pH curve Burette, 250ml conical flask, 25ml measuring cylinder, pH probe, 0.1 HCl, 0.1 NaOH OR DEMO: pH probe and 25ml of each substance: 0.1 mol/dm^3 hydrochloric acid 0.01 mol/dm^3 hydrochloric acid 0.001 mol/dm^3 hydrochloric acid 0.1 mol/dm^3 nitric acid- CLEAPSS Hazcard 67 0.1 mol/dm^3 sulfuric acid- CLEAPSS Hazcard 98A 0.1 mol/dm^3 ethanoic acid- CLEAPSS Hazcard 38A 0.1 mol/dm^3 citric acid- CLEAPSS Hazcard 36C Distilled water 0.1 mol/dm^3 sodium chloride solution 0.1 mol/dm^3 sodium hydroxide

10. Strong and weak acids HIGHER TIER ONLY	6 – To describe how an acid or alkali can be concentrated or diluted. 8 – To quantitatively explain how concentration of hydrogen ions relate to the pH number.	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, the hydrogen ion concentration of the solution increases by a factor of 10. Students should be able to: • use and explain the terms dilute and concentrated (in terms of amount of substance), and weak and strong (in terms of the degree of ionisation) in relation to acids • describe neutrality and relative acidity in terms of the effect on hydrogen ion concentration and the numerical value of pH (whole numbers only).	
11. Electrolysis of molten compounds	4 – Define the process of electrolysis 6 – To predict the products at each electrode for molten compounds. 8 – Describe electrolysis using half equations.	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). Ions are discharged at the electrodes producing elements. This process is called electrolysis. (HT only) Throughout Section 4.4.3 Higher Tier students should be able to write half equations for the reactions occurring at the electrodes during electrolysis, and may be required to complete and balance supplied half equations. 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. Students should be able to predict the products of the electrolysis of binary ionic compounds in the molten state.	DEMO: Electrolysis of Zinc chloride – IN FUME CUPBOARD ONLY – see techs

12. Using electrolysis to extract metals	4 – Write the word equation for the electrolysis of aluminum oxide. 6 – Describe how aluminum is extracted. 8 – Write half equations for each electrode.	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. 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 using carbon as the positive electrode (anode). Students should be able to: • explain why a mixture is used as the electrolyte • explain why the positive electrode must be continually replaced.	
13. Electrolysis of aqueous solutions	4 – To state the products of the electrolysis of brine. 6 – To describe how to electroylse brine using the term ions. 8 – To explain the electrolysis of brine using half equations and key words.	The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved. At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen. 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.	PRAC: Electrolysis of brine Power packs set at low voltage, U-tubes or similar, electrodes, croc-plug clips, litmus paper, brine
14. REQ PRAC: Electrolysis of aqueous solutions	4 – To state the products of the electrolysis of brine. 6 – To describe how to electroylse brine using the term ions. 8 – To explain the electrolysis of brine using half equations and key words.	**Required practical activity 9:** investigate what happens when aqueous solutions are electrolysed using inert electrodes. This should be an investigation involving developing a hypothesis. AT skills covered by this practical activity: chemistry AT 3 and 7.	PRAC: Electrolysis of aqueous solutions Power packs set at low voltage, U-tubes or similar, electrodes, croc-plug clips, litmus paper, range of aqueous solutions as per required practical guidance.

15. Reactions at electrodes – half equations HIGHER TIER ONLY	6 – To explain REDOX in terms of electrons. 8 – To represent the reactions at electrodes as half equations	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 $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ or $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$	