| Learning Intention | Vocab | Concept | Retrieval | Success Criteria | Hinge Questions for this lesson | Red Zone |
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| **Week 9 Lesson 1 Inset Day** | | | | | | |
| **Week 9 Lesson 2**<br>What are the properties of ionic substances?<br>Ionic Properties EMI.pptx | High melting point<br>Soluble<br>Conductivity<br>Solid<br>Aqueous<br>Brittle<br>Lattice structure | Bonding and properties | Are ionic compounds usually solid or liquid at room temperature? Why? and draw the particle arrangement of solid liquid and gas | Explain the properties of ionic compounds limited to:<br>a) high melting points and boiling points, in terms of forces between ions<br>b) whether or not they conduct electricity as solids, when molten and in aqueous solution. | Why do ionic compounds conduct electricity when molten but not when solid?<br>A. Electrons move freely in the solid<br>B. Ions are free to move when molten<br>C. The ions disappear when solid<br>D. The solid contains no charged particles<br>Misconception targeted: Students think electrons conduct in ionic compounds. | Explain why ionic compounds have high melting and boiling points.<br><br>Describe the role of strong electrostatic forces between oppositely charged ions.<br>Explain why a lot of energy is needed to overcome these forces during a change of state.<br><br>Describe the electrical conductivity of ionic compounds in different states:<br><br>Why do they not conduct electricity as solids?<br>Why do they conduct when molten or dissolved in water?<br>Use particle movement to support your explanation.<br><br>List two examples of ionic compounds (e.g., sodium chloride, calcium carbonate).<br>For each, describe one property and explain how it relates to its real-world use (e.g., table salt, antacid, road grit). |

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| **Week 9 Lesson 3**<br>Ionic bonding practise. | Ionic bond<br>Electrostatic attraction<br>Metal<br>Non-metal<br>Compound<br>Formula<br>Giant ionic lattice | Bonding and properties | What type of elements usually form positive ions? and What type of elements usually form negative ions? | | What is the correct formula for calcium nitrate?<br>A. $CaNO_3$<br>B. $Ca(NO_3)_2$<br>C. $Ca_2NO_3$<br>D. $Ca_2(NO_3)$<br>Misconception targeted: Students forget that nitrate is a polyatomic ion with charge –1. | |
| **Week 9 Lesson 4**<br>How do non metals bond?<br>Covalent-Bonding EMI.pptx | Covalent bond<br>Shared electrons<br>Molecule<br>Non-metal<br>Simple molecular structure<br>Bond pair | Bonding and properties | What is a molecule?<br>What happens to electrons in a covalent bond? | 1. Explain how a covalent bond is formed when a pair of electrons is shared between two atoms.<br>2. Recall that covalent bonding results in the formation of molecules.<br>3. Recall the typical size (order of magnitude) of atoms and small molecules.<br>4. Explain the formation of simple molecular, covalent substances, using dot and cross diagrams, including:<br>a) hydrogen<br>b) hydrogen chloride<br>c) water<br>d) methane<br>e) oxygen<br>f) carbon dioxide | What happens in a covalent bond?<br>A. Electrons are transferred from one atom to another<br>B. A pair of electrons is shared between atoms<br>C. Protons are shared between atoms<br>D. Neutrons are transferred between atoms<br>Misconception targeted: Students confuse ionic transfer with covalent sharing. | In your own words, explain how a covalent bond forms.<br><br>What happens to the electrons? Why does this make atoms more stable?<br><br>Draw dot and cross diagrams to show covalent bonding in the following molecules:<br><br>Hydrogen ($H_2$)<br>Water ($H_2O$)<br>Methane ($CH_4$)<br>Oxygen ($O_2$)<br>Carbon dioxide ($CO_2$)<br>Use different symbols (dots and crosses) for electrons from different atoms. |

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| **Week 10 Lesson 1 Responsive Curriculum** | | | | | | |
| **Week 10 Lesson 2** What are the properties of covalent structures? Covalent Properties EMI.pptx | Low melting point Insulator Weak intermolecular forces Simple molecular Giant covalent Insoluble | Bonding and properties | Are covalent substances usually gases, liquids, or solids at room temperature? Do covalent substances conduct electricity? Why or why not? | 1. Explain the properties of typical covalent, simple molecular compounds limited to a) low melting points and boiling points, in terms of forces between molecules (intermolecular forces) b) poor conduction of electricity 2. Describe, using poly(ethene) as the example, that simple polymers consist of large molecules containing chains of carbon atoms | Why do simple molecular substances like oxygen have low melting points? A. Their covalent bonds are weak B. Their intermolecular forces are weak C. They have no forces between molecules D. Their atoms can slide over each other Misconception targeted: Students think covalent bonds break during melting. | Explain why simple molecular covalent compounds (e.g., water, carbon dioxide) have low melting and boiling points.<br><br>Focus on the weak intermolecular forces between molecules. Compare this to the strong ionic or metallic bonds in other substances.<br><br>Why do simple molecular substances not conduct electricity?<br><br>Mention the absence of free-moving charged particles (like ions or electrons).<br><br>Describe what a polymer is and explain how poly(ethene) is formed.<br><br>Include the idea of long chains of carbon atoms. |

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| | | | | | | How does the size of polymer molecules compare to simple molecules? |
| **Week 10 Lesson 3**<br>How many ways can carbon bond?<br>CC7b-Allotropes-of-Carbon.pptx<br><br>6 marker – Marked piece 1 | Carbon<br>Diamond<br>Graphite<br>Fullerenes<br>Nanotubes<br>Allotropes | Bonding and properties | What is the name of the carbon structure used in pencils?<br>What is a fullerene? | 1. Recall that graphite and diamond are different forms of carbon and that they are examples of covalent giant molecular substances.<br>2. Describe the structures of graphite and diamond.<br>3. Explain, in terms of structure and bonding, why graphite is used to make electrodes and as a lubricant, whereas diamond is used in cutting tools.<br>4. Explain the properties of fullerenes (e.g. C60) and graphene in terms of their structures and bonding. | Why can graphite conduct electricity but diamond cannot?<br>A. Diamond has no electrons<br>B. Graphite has delocalised electrons between layers<br>C. Diamond has weak bonds<br>D. Graphite contains metal atoms<br>Misconception targeted: Students think only metals conduct electricity. | Write a short description of the structures of:<br><br>Diamond – include how many atoms each carbon is bonded to and the shape of the structure.<br>Graphite – describe the layers of hexagonal rings and the forces between them.<br><br>Why is diamond used in cutting tools?<br>Why is graphite used in lubricants and electrodes?<br>Link your answers to their structure and bonding.<br><br>What is a fullerene (e.g., $C_{60}$)? Describe its shape and bonding.<br>What is graphene and why is it considered a special material? Mention at least one use for each. |
| **Week 10 Lesson 4** | | | | | Which dot-and-cross diagram correctly represents methane ($CH_4$)?<br>A. Carbon sharing 1 | |

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| Covalent bonding practise. | | | | | electron with each H<br>B. Carbon sharing 2 electrons with each H<br>C. Carbon transferring electrons to H<br>D. Carbon forming ionic bonds with H<br>Misconception targeted: Students think hydrogen forms double bonds or ionic bonds. | |
| **Week 11 Lesson 1**<br>How do metals bond?<br>Metallic bonding lesson.pptx | Metallic bond<br>Delocalised electrons<br>Positive ions<br>Lattice<br>Conductivity<br>Malleable | Bonding and properties | What happens to electrons in metallic bonding?<br>What is meant by a "sea of delocalised electrons"?<br>Why are metals good conductors of electricity? | 1. Describe how metal ions are bonded within a metal.<br>2. Recall which ions are formed in different metals.<br>3. Compare the structure and movement of ions in pure metals and alloys. | What holds metal ions together in a metal lattice?<br>A. Attraction between ions and delocalised electrons<br>B. Sharing of electrons between two atoms<br>C. Attraction between positive ions<br>D. Repulsion between electrons<br>Misconception targeted: Students confuse metallic bonding with ionic or covalent bonding. | Describe how metal atoms form positive ions and how they are held together in a metal.<br><br>What is meant by a 'sea' of delocalised electrons?<br>Why is this important for bonding?<br>Recall the ions formed by these metals and write their symbols:<br><br>Iron<br>Copper<br>Sodium<br>(e.g., $Fe^{2+}$, $Cu^{2+}$, $Na^{+}$)<br>Structure of Metals vs Alloys (4 minutes)<br>Compare the structure of pure metals and alloys: |

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| | | | | | | Draw a simple diagram of each (label the regular vs irregular arrangement of atoms).<br>Explain how this affects malleability and strength.<br><br>Explain why metals conduct electricity.<br><br>What role do delocalised electrons play?<br>Why is this useful in real-world applications?<br><br>In 2–3 sentences, explain why alloys are often harder than pure metals.<br>Give one example of a common alloy and its use (e.g., steel, bronze) |
| **Week 11 Lesson 2**<br>What are the properties of metals and alloys?<br>CC7c-Properties-of-Metals.pptx | Alloy<br>Malleable<br>Ductile<br>Conductor<br>Strong<br>Layers | Bonding and properties | What is an alloy?<br>Why are alloys usually stronger than pure metals?<br>Name two properties of metals. | 1. Explain the properties of metals, including malleability and the ability to conduct electricity.<br>2. Describe most metals as shiny solids which have high melting points, high density and are good conductors of electricity whereas most non-metals have low boiling | Why are alloys harder than pure metals?<br>A. Their ions are arranged in perfect layers<br>B. Different-sized atoms distort layers and prevent sliding<br>C. They contain more delocalised electrons<br>D. They have weaker | Write a short paragraph explaining the typical properties of metals.<br>Include the following terms:<br><br>Malleable<br>High melting point<br>High density<br>Electrical conductivity<br>Shiny appearance |

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| | | | | points and are poor conductors. | metallic bonds<br>Misconception targeted: Students think alloys are softer because they are "mixed." | Write a paragraph comparing non-metals to metals.<br><br>Mention their boiling points, electrical conductivity, and physical state at room temperature.<br>Give two examples of non-metals and their uses. |
| **Week 11 Lesson 3**<br>How do the different types of bonding compare?<br>CC7d-Bonding-Models.pptx | Ionic<br>Covalent<br>Metallic<br>Structure<br>Properties<br>Conductivity<br>Melting point | Bonding and properties | Which type of bonding involves the transfer of electrons?<br>Which type of bonding involves shared electrons?<br>Which type of bonding allows substances to conduct electricity when solid? | 1. Explain why elements and compounds can be classified as: ionic, simple molecular, giant covalent and metallic.<br>2. Explain how the structure and bonding of these types of substances results in different physical properties.<br>3. Describe the limitations of particular representations and models to include dot and cross, ball and stick models and two- and three-dimensional representations | Which substance is correctly matched to its bonding and structure?<br>A. Sodium chloride — simple molecular<br>B. Oxygen — giant covalent<br>C. Diamond — giant covalent<br>D. Copper — ionic lattice<br>Misconception targeted: Students mix up bonding types and structures. | Classify each of the following substances as ionic, simple molecular, giant covalent, or metallic:<br><br>Sodium chloride<br>Water<br>Diamond<br>Copper<br><br>For each substance, explain how its structure and bonding affect its:<br><br>Melting point<br>Hardness<br>Electrical conductivity<br>Use 1–2 sentences per substance.<br><br>Choose one type of bonding (e.g., covalent, ionic, or metallic).<br>Describe the strengths and |

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| | | | | | | limitations of using the following models to represent it:<br><br>Dot and cross diagrams<br>Ball and stick models<br>3D representations<br>Use examples to support your explanation (e.g., water for covalent, NaCl for ionic). |
| **Week 11 Lesson 4**<br>What are acids and alkalis/bases?<br>Acids, alkalis & indicators EMI.pptx | Acid<br>Alkali<br>Base<br>pH scale<br>Indicator<br>Neutral<br>Corrosive | Chemical reactions | What colour does litmus paper turn in an acid?<br>What is the pH range of an acid?<br>What is the pH of a neutral substance? | 1. Describe the use of hazard symbols on containers<br>2. Recall that acids in solution are sources of hydrogen ions and alkalis in solution are sources of hydroxide ions.<br>3. Recall that a neutral solution has a pH of 7 and that acidic solutions have lower pH values and alkaline solutions higher pH values.<br>4. Recall the effect of acids and alkalis on indicators, including litmus, methyl orange and phenolphthalein.<br>5. Recall that the higher the concentration of hydrogen | What ion is responsible for acidity in aqueous solutions?<br>A. $OH^-$<br>B. $H^+$<br>C. $O^{2-}$<br>D. $H_2$<br>Misconception targeted: Students confuse hydrogen atoms with hydrogen ions. | Look at the four common hazard symbols below .<br>For each, write:<br>(toxic, corrosive, danger to environment put pictures on the slide)<br>The name of the symbol<br>What it warns you about (e.g., corrosive, toxic)<br>One example of a substance that might carry this symbol<br>Acids and Alkalis (3 minutes)<br><br>What ion do acids release in solution?<br>What ion do alkalis release in solution?<br>Write the pH range for acids, alkalis, and neutral substances.<br><br>Explain how the concentration of |

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| | | | | ions in an acidic solution, the lower the pH; and the higher the concentration of hydroxide ions in an alkaline solution, the higher the pH. (H) | | hydrogen ions ($H^+$) affects the pH of a solution.<br><br>What happens to the pH if the concentration of $H^+$ increases? What happens if the concentration of $OH^-$ increases? |
| **Week 12 Lesson 1 Responsive Curriculum** | | | | | | |
| **Week 12 Lesson 2 What makes an acid? (H only)** CC8b-Looking-at-acids-H.pptx | Hydrogen ion ($H^+$)<br>Proton donor<br>Aqueous<br>Strong acid<br>Weak acid<br>Dissociation | Chemical reactions | What ion do acids release in water?<br>Give an example of a common acid used in the lab.<br>What is the opposite of an acid? | 1. Recall that as hydrogen ion concentration in a solution increases by a factor of 10, the pH of the solution decreases by 1.<br>2. Explain the terms dilute and concentrated, with respect to amount of substances in solution.<br>3. Explain the terms weak and strong acids, with respect to the degree of dissociation into ions. | A solution becomes 10× more concentrated in $H^+$ ions. What happens to its pH?<br>A. pH increases by 1<br>B. pH decreases by 1<br>C. pH increases by 10<br>D. pH decreases by 10<br>Misconception targeted: Students confuse concentration changes with pH scale behaviour. | Explain how increasing the concentration of hydrogen ions ($H^+$) in a solution by a factor of 10 causes the pH to decrease by 1.<br><br>Define the terms dilute and concentrated in the context of acids.<br><br>Give an example of a dilute acid and a concentrated acid.<br>Does concentration affect the strength of an acid?<br><br>Explain the difference between strong acids and weak acids in terms of dissociation into ions.<br><br>What does it mean for an acid to |

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| | | | | | | fully dissociate?<br>How does this affect the pH of the solution?<br><br>Give two examples of strong acids (e.g., hydrochloric acid, sulfuric acid) and two examples of weak acids (e.g., ethanoic acid, citric acid).<br><br>Describe how their dissociation affects their pH and reactivity. |
| **Week 12 Lesson 3**<br>How do acids and bases react?<br>CC8c-bases-and-salts.pptx | Neutralisation<br>Salt<br>Water<br>Base<br>Exothermic | Chemical reactions | What are the products of a reaction between an acid and a base?<br>What is the name of this type of reaction?<br>What is a salt? | 1. Recall that a base is any substance that reacts with an acid to form salt and water only.<br>2. Explain the general reactions of aqueous solutions of acids with metal oxides to produce salts<br>3. Describe a neutralisation reaction as a reaction between an acid and a base.<br>4. Write word equations/ Write balanced chemical equations, including the use of the state symbols (s), (l), (g) and (aq) | What is formed when an acid reacts with a base?<br>A. Salt only<br>B. Salt and water<br>C. Water and hydrogen<br>D. Salt and carbon dioxide<br>Misconception targeted: Students confuse metal carbonate reactions with base reactions. | Define a base.<br>What is produced when an acid reacts with a base?<br>What is a neutralisation reaction?<br><br>Hydrochloric acid reacts with copper(II) oxide in aqueous solution.<br><br>What type of reaction is this?<br>What are the products of this reaction?<br>Write the word equation for this reaction.<br>Write the balanced chemical equation with state symbols.<br><br>Complete the following for each |

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| | | | | | | reaction.<br><br>A. Sulfuric acid reacts with magnesium oxide.<br>B. Nitric acid reacts with zinc oxide.<br><br>For each:<br><br>Write the word equation.<br>Write the balanced chemical equation with state symbols. |
| **Week 12 Lesson 4**<br>**Inset Day** | | | | | | |
| **Week 13 Lesson 1**<br>**Inset Day** | | | | | | |
| **Week 13 Lesson 2**<br>Core practical: Preparing Copper Sulfate (1 of 2 lessons)<br>CC8c-Core-Prac.pptx | Filtration<br>Crystallisation<br>Evaporation<br>Excess reactant<br>Soluble salt | Chemical reactions | What two substances are used to make copper sulfate crystals?<br>Why is excess copper oxide added?<br>What do you do after filtering the solution? | 1. Investigate the preparation of pure, dry hydrated copper sulfate crystals starting from copper oxide including the use of a water bath.<br>2. Explain why, if soluble salts are prepared from an acid and an insoluble reactant:<br>a) excess of the reactant is added<br>b) the excess reactant is | Why is excess copper oxide added when making copper sulfate?<br>A. To make the solution blue<br>B. To ensure all the acid reacts<br>C. To increase the temperature<br>D. To make the crystals form faster<br>Misconception targeted: Students think excess | Describe the method for preparing pure, dry hydrated copper sulfate crystals from copper oxide and sulfuric acid.<br>Explain why excess copper oxide is used in the reaction and how the excess is removed.<br>Describe why the solution after removal of excess reactant contains only copper sulfate and water, and how this ensures the purity of the crystals formed. |

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| | | | | removed<br>c) the solution remaining is only salt and water. | reactant is for colour or speed. | |
| **Week 13 Lesson 3**<br>Core practical: Preparing Copper Sulfate (2 of 2 lessons)<br>CC8c-Core-Prac.pptx<br><br>6 marker – Marked piece 2 | Filtration<br>Crystallisation<br>Evaporation<br>Excess reactant<br>Soluble salt | Chemical reactions | What two substances are used to make copper sulfate crystals?<br>Why is excess copper oxide added?<br>What do you do after filtering the solution? | 1. Investigate the preparation of pure, dry hydrated copper sulfate crystals starting from copper oxide including the use of a water bath.<br>2. Explain why, if soluble salts are prepared from an acid and an insoluble reactant:<br>a) excess of the reactant is added<br>b) the excess reactant is removed<br>c) the solution remaining is only salt and water. | Why is excess copper oxide added when making copper sulfate?<br>A. To make the solution blue<br>B. To ensure all the acid reacts<br>C. To increase the temperature<br>D. To make the crystals form faster<br>Misconception targeted: Students think excess reactant is for colour or speed. | Describe the method for preparing pure, dry hydrated copper sulfate crystals from copper oxide and sulfuric acid.<br>Explain why excess copper oxide is used in the reaction and how the excess is removed.<br>Describe why the solution after removal of excess reactant contains only copper sulfate and water, and how this ensures the purity of the crystals formed. |
| **Week 13 Lesson 4**<br>How do acids and metal hydroxides react?<br>CC8d-Acid and Metal Hydroxide.pptx | | | | 1. Explain the general reactions of aqueous solutions of acids with metal hydroxides to produce salts.<br>2. Write word equations./ Write balanced chemical equations, including the use of the state symbols (s), (l), | What type of reaction occurs when an acid reacts with a metal hydroxide?<br>A. Displacement<br>B. Neutralisation<br>C. Oxidation<br>D. Combustion<br>Misconception targeted: Students confuse | Explain how acids react with metal hydroxides in aqueous solutions to produce salts and water.<br>Write both word and balanced chemical equations for these reactions, including appropriate state symbols such as (s), (l), (g), and (aq).<br>Recall and explain that alkalis are |

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| | | | | (g) and (aq)<br>3. Recall that alkalis are soluble bases | neutralisation with displacement. | soluble bases and give examples of common alkalis. |
| **Week 14 Lesson 1**<br>**Responsive Curriculum** | | | | | | |
| **Week 14 Lesson 2**<br>Core Practical: Investigating Neutralisation (1 of 2 lessons)<br>CC8d-Core-Prac.pptx | pH<br>Indicator<br>Titration<br>Neutral | Chemical reactions | What equipment can be used to measure pH?<br>What happens to the pH as an acid is neutralised? | 1. Investigate the change in pH on adding powdered calcium hydroxide or calcium oxide to a fixed volume of dilute hydrochloric acid | During neutralisation, what happens to $H^+$ and $OH^-$ ions?<br>A. They form hydrogen gas<br>B. They form water<br>C. They form salt only<br>D. They disappear without reacting<br>Misconception targeted: Students think ions "cancel out" without forming water. | Outline the steps for the following experiment:<br>To investigate how the pH of dilute hydrochloric acid changes as powdered calcium hydroxide or calcium oxide is added gradually.<br><br>Include:<br>Apparatus needed (e.g., beaker, spatula, pH probe or universal indicator, stirring rod).<br>Step-by-step method (start with a fixed volume of dilute HCl, add base in small amounts, stir, and record pH after each addition).<br>How you would record your results? (e.g., in a table showing amount of base added vs. pH). |
| **Week 14 Lesson 3** | pH<br>Indicator | Chemical reactions | What equipment can be used to | 1. Investigate the change in pH on adding powdered calcium hydroxide or | During neutralisation, what happens to $H^+$ and $OH^-$ ions? | Outline the steps for the following experiment:<br>To investigate how the pH of dilute |

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| Core Practical: Investigating Neutralisation (2 of 2 lessons) CC8d-Core-Prac.pptx | Titration Neutral | | measure pH? What happens to the pH as an acid is neutralised? | calcium oxide to a fixed volume of dilute hydrochloric acid | A. They form hydrogen gas<br>B. They form water<br>C. They form salt only<br>D. They disappear without reacting<br>Misconception targeted: Students think ions "cancel out" without forming water. | hydrochloric acid changes as powdered calcium hydroxide or calcium oxide is added gradually.<br><br>Include:<br>Apparatus needed (e.g., beaker, spatula, pH probe or universal indicator, stirring rod).<br>Step-by-step method (start with a fixed volume of dilute HCl, add base in small amounts, stir, and record pH after each addition).<br>How you would record your results? (e.g., in a table showing amount of base added vs. pH). |
| **Week 14 Lesson 4** What happens to the particles during neutralisation? CC8e-Alkalis-and-Neutralisation.pptx | Hydrogen ion ($H^+$) Hydroxide ion ($OH^-$) Water molecule Ionic equation Neutralisation | Chemical reactions | What ion from the acid reacts with the hydroxide ion? What is formed when these ions react? Why is the solution neutral at the end? | 1. Explain an acid-alkali neutralisation as a reaction in which hydrogen ions (H+) from the acid react with hydroxide ions (OH–) from the alkali to form water.<br>2. Explain why, if soluble salts are prepared from an acid and a soluble reactant:<br>a) titration must be used<br>b) the acid and the soluble reactant are then mixed in the correct proportions<br>c) the solution remaining, | During neutralisation, what happens to $H^+$ and $OH^-$ ions?<br>A. They form hydrogen gas<br>B. They form water<br>C. They form salt only<br>D. They disappear without reacting<br>Misconception targeted: Students think ions "cancel out" without forming water. | Explain what happens to the particles during neutralisation, describing how hydrogen ions ($H^+$) from an acid react with hydroxide ions ($OH^-$) from an alkali to form water.<br>Explain why titration is necessary when preparing soluble salts from an acid and a soluble base, including how the acid and alkali must be mixed in the correct proportions to ensure the solution contains only salt and water after the reaction. |

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| | | | | after reaction, is only salt and water.<br>3. Describe how to carry out simple acid-alkali titrations, using burette, pipette and a suitable indicator, to prepare a pure, dry salt. | | Describe how to carry out a simple acid-alkali titration using a burette, pipette, and suitable indicator to prepare a pure, dry salt. |
| **Week 15 Lesson 1**<br>How do acids and metals and carbonates react?<br>CC8f-Acids-with-metals-and-carbonates.pptx | Carbonate<br>Salt<br>Hydrogen gas<br>Carbon dioxide<br>Effervescence | Chemical reactions | What gas is produced when acids react with metals?<br>What gas is produced when acids react with carbonates?<br>How can you test for carbon dioxide? | 1. Explain the general reactions of aqueous solutions of acids with metals and metal carbonates to produce salts.<br>2. Describe the chemical test for:<br>a) hydrogen<br>b) carbon dioxide (using limewater)<br>3. Write balanced ionic equations (H) | Which gas is produced when a metal carbonate reacts with an acid?<br>A. Hydrogen<br>B. Oxygen<br>C. Carbon dioxide<br>D. Chlorine<br>Misconception targeted: Students confuse carbonate reactions with metal reactions. | Explain how acids react with metals and metal carbonates in aqueous solutions to produce salts. Describe the chemical tests used to identify hydrogen gas and carbon dioxide gas.<br>Write balanced ionic equations for these reactions, showing the key ions involved. |
| **Week 15 Lesson 2**<br>Which salts are soluble?<br>CC8g-Solubility.pptx | Solubility<br>Precipitate<br>Insoluble<br>Soluble<br>Nitrate<br>Sulfate<br>Precipitation reaction | Chemical reactions | Are all nitrates soluble?<br>Are most carbonates soluble or insoluble?<br>What is a precipitate? | 1. Recall the general rules which describe the solubility of common types of substances in water.<br>2. Predict, using solubility rules, whether or not a precipitate will be formed when named solutions are | What happens when solutions of silver nitrate and sodium chloride are mixed?<br>A. No reaction<br>B. A white precipitate of silver chloride forms<br>C. A gas is produced<br>D. The solution turns acidic | Write down the general solubility rules for common ionic substances in water. Include rules for:<br><br>Nitrates<br>Chlorides<br>Sulfates<br>Carbonates |

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| | | | | mixed together, naming the precipitate if any. | Misconception targeted: Students think all nitrates or chlorides are soluble. | Hydroxides<br><br>Use your solubility rules to predict whether a precipitate will form in each of the following reactions. If a precipitate forms:<br><br>Name the precipitate<br>Explain why it is insoluble<br>A. Sodium sulfate + Barium chloride<br>B. Potassium nitrate + Silver chloride<br>C. Lead(II) nitrate + Potassium iodide |
| **Week 15 Lesson 3**<br>What is DNA?<br>DNA EMI.pptx | Chromosome<br>Double helix<br>Genetic code<br>Nucleotide | Inheritance and Evolution<br><br>Cells and Systems | Cells contain a nucleus where genetic material is stored<br>Living organisms inherit traits from their parents | 1. Recall where DNA is found in a eukaryotic cell.<br>2. Names the bases in DNA.<br>3. Recall the pairing of bases in DNA.<br>4. Describe how DNA strands are held together.<br>5. Describe the overall structure of DNA. | Which statement best describes DNA?<br>A. DNA is a type of protein that gives cells energy.<br>**B. DNA is a molecule that carries genetic instructions for how an organism develops and functions.** ☑<br>C. DNA is a cell structure that controls what enters and leaves the cell.<br>D. DNA is a chemical used by the body to digest food. | Describe and explain the structure of DNA. (6 marks) |

| Learning Intention | Vocab | Concept | Retrieval | Success Criteria | Hinge Questions for this lesson | Red Zone |
|---|---|---|---|---|---|---|
| **Week 15 Lesson 4**<br>What is Mitosis? (REVIEW)<br>Lesson 1 Mitosis EMI.pptx | Cell division<br>Diploid<br>Growth<br>Repair | Inheritance and Evolution<br><br>Cells and Systems | Cells reproduce to make new cells<br>Chromosomes carry genetic information | 1. List the names and order the stages of the cell cycle including mitosis.<br>2. Describe what happens in each stage of the cell cycle, including mitosis.<br>3. Describe why mitosis is important for an organism's (growth, repair, asexual reproduction).<br>4. Explain why organisms may rely on asexual reproduction.<br>5. Describe how mitosis produces genetically identical, diploid cells.<br>6. Describe how cancers grow. | Hinge Question: What is mitosis?<br>Which statement best describes mitosis?<br>A. The division of a cell to produce gametes with half the genetic material<br>B. The process by which a cell makes energy using oxygen<br>**C. The division of a nucleus to produce two genetically identical cells**<br>D. The movement of proteins out of the cell | A scientist observes a group of cells dividing rapidly and forming a mass. The cells show no signs of differentiation and ignore signals to stop dividing.<br><br>Your Task:<br>Write a short explanation (150–200 words) answering:<br>- What part of the cell cycle is malfunctioning?<br>- How does this relate to cancer?<br>- Why are these cells dangerous to the organism?<br>- How is this different from normal mitosis |