Science (Chemistry) 5088/03 Paper 3, Oct/Nov 2026

Model answers aligned to **Discover Chemistry for GCE 'O' Level Science and Normal (A) Level Science, Textbook, 3rd Edition (Marc Chang, Alistair Chew, John Sadler, Tan Yin Toon, Wong Heng-Vee, Woo Chang Hong; Marshall Cavendish Education)** · Printed page numbers, which run continuously through the book (Ch 3 starts p. 37, Ch 19 starts p. 289). pdf_page is the page index in the full-book PDF 'Discover Chemistry GCE 'O' & NA Level (3rd ed).pdf' (= printed page + 11); in the separate chapter PDFs the page is printed page minus (chapter start page - 1).

> Teacher-grade model answers for Questions 1 to 5 (33 marks) of the real Oct/Nov 2026 Science (Chemistry) Paper 3. Cambridge has not released an official mark scheme, so every mark point and accept/reject note is the author's own, written in Cambridge style. Each mark point is worded from the Discover Chemistry text wherever the book has the wording, and each part carries the full passages it is built from, quoted verbatim with printed page numbers; a status says whether the book teaches the idea (in this section all 24 parts are taught in the book).

Concepts not (fully) covered by the textbook

Parts by status: In the textbook **44** · Partly in the textbook **3** · NOT in the textbook **0** · Data skill (answered from the question) **2**

Part	Concept	Status	What to do
8(a) T	Deducing that the black solid left after a metal dissolves in acid, and which burns in air to give CO_2, is carbon	Partly in the textbook	Teach it as a deduction habit: metal dissolves (iron); black solid that does not react with dilute acid and burns to CO_2 = carbon (steel = iron + carbon, Table 5.3). Add one line to the notes: carbon is a black non-metal that does not react with dilute acids and burns in air to form carbon dioxide.
9(c)(i)	"Include all the electrons" in a dot-and-cross diagram, i.e. the first shell as well, and the term "lone pair"	Partly in the textbook	The book draws outer electrons only. Tell students that "all the electrons" means the inner shell too (N: 2 in the first shell + 5 in the outermost; in NH_3 the outer shell shows 3 bonding pairs + 1 lone pair). Practise NH_3, H_2O and CH_4 with inner shells drawn, and teach the word "lone pair".
10(b)(ii)	Method for finding a pH with Universal Indicator: add it, then compare the colour with the pH chart	Partly in the textbook	Add the explicit step "compare the colour with the pH colour chart" to the Universal Indicator notes (Figures 8.11 and 8.13 show the chart but no sentence says to use it). Remind students that a pH sensor or data logger is not an "other" method when a pH meter is already named.

Question 1 — Choosing compounds from a list (formulae, photosynthesis, organic classes) [5]

1(a) **[1]**

Identify a compound in the list that is a molecule with only five atoms.

- CH_4 (methane): the symbols show carbon and hydrogen, and the subscript shows four hydrogen atoms, so one molecule has 1 carbon + 4 hydrogen = five atoms [1]

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Count the atoms in each formula (a subscript gives the number of atoms of that element; no subscript means 1): CO = 2; CO₂ = 3; CH₄ = 1 + 4 = 5; CH₃OH = 1 + 4 + 1 = 6; CH₃CH₂OH = 2 + 6 + 1 = 9; CH₃CH₂CO₂H = 3 + 6 + 2 = 11. Only CH₄ has five atoms.
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CH_4 or methane. Reject CH_3OH (6 atoms: the hydrogen in the –OH group is easy to forget) and CO_2 (3 atoms).

Ch 6 Chemical Formulae and Equations › 6.1 What Are Chemical Formulae? › p. 91 (PDF p. 102)

"A chemical formula is like a recipe that shows the ratio of elements that are combined in one unit of a substance. It is made up of: • chemical symbol(s) that indicate the element(s) present; and • subscript(s) indicating the number of atoms of each element present."

Ch 6 Chemical Formulae and Equations › 6.1 What Are Chemical Formulae? › p. 91 (PDF p. 102)

"These subscripts indicate that: • there are two oxygen atoms in one molecule of carbon dioxide; and • there are two hydrogen atoms in one molecule of water."

Ch 6 Chemical Formulae and Equations › 6.1 What Are Chemical Formulae? › p. 92 (PDF p. 103)

"Name of Substance Chemical Formula Atoms Present ammonia NH_3 1 nitrogen and 3 hydrogen copper(II) sulfate $CuSO_4$ 1 copper, 1 sulfur and 4 oxygen sodium hydroxide NaOH 1 sodium, 1 oxygen and 1 hydrogen hydrochloric acid HCl 1 hydrogen and 1 chlorine ethanoic acid CH_3COOH 4 hydrogen, 2 carbon and 2 oxygen"

Ch 16 Hydrocarbons › 16.2 What Are Alkanes? › p. 242 (PDF p. 253)

"When there is one carbon atom in the alkane molecule (n = 1): • The number of hydrogen atoms will be 2(1) + 2 = 4. • Hence, the formula of the alkane is CH_4 (Figure 16.7)."

In the textbook

LO: 4.1(b)

1(b) **[1]**

Identify a compound that reacts with water to make glucose during photosynthesis.

- CO_2 (carbon dioxide): plants take in carbon dioxide and water to make glucose and oxygen in the presence of sunlight [1]

Photosynthesis: $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$. The compound that reacts with water is carbon dioxide; oxygen is the other product, not a reactant, and it is not in the list.

CO_2 or carbon dioxide. Reject CO (carbon monoxide).

Ch 19 Maintaining Air Quality › 19.3 What Is the Carbon Cycle? › p. 296 (PDF p. 307)

"Photosynthesis Plants take in carbon dioxide and water to make glucose and oxygen in the presence of sunlight. $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$ Glucose is used in other processes that sustain plant life."

Ch 19 Maintaining Air Quality › 19.3 What Is the Carbon Cycle? › p. 296 (PDF p. 307)

"The element carbon is essential for life. All organic compounds are made of carbon skeletons that originate from carbon dioxide and water during photosynthesis."

In the textbook

LO: 12(g)

1(c) **[1]**

Identify a compound that is a carboxylic acid.

- $CH_3CH_2CO_2H$ (propanoic acid): carboxylic acids are organic compounds with the carboxyl (–COOH, written $–CO_2H$ in the paper) functional group [1]

Only $CH_3CH_2CO_2H$ contains the carboxyl group; CH_3OH and CH_3CH_2OH are alcohols (–OH), CH_4 is an alkane, CO and CO_2 are inorganic oxides.

$CH_3CH_2CO_2H$, CH_3CH_2COOH, C_2H_5COOH or propanoic acid. Reject CH_3CH_2OH (an alcohol). The paper and the syllabus write the group as $–CO_2H$; the book writes –COOH.

Ch 17 Alcohols and Carboxylic Acids › 17.2 What Are Carboxylic Acids? › p. 265 (PDF p. 276)

"Carboxylic acids are organic compounds with the carboxyl (–COOH) functional group. All carboxylic acids have this functional group."

Ch 17 Alcohols and Carboxylic Acids › 17.2 What Are Carboxylic Acids? › p. 265 (PDF p. 276)

"The formulae of the first three members of the carboxylic acid homologous series are shown in Table 17.3."

Ch 17 Alcohols and Carboxylic Acids › 17.2 What Are Carboxylic Acids? › p. 265 (PDF p. 276)

"propanoic acid When n = 2, C_2H_5COOH $C_3H_6O_2$ CH_3CH_2COOH"

In the textbook

LO: 11.3(e)

1(d) **[1]**

Identify a compound that is a hydrocarbon.

- CH_4 (methane): a hydrocarbon is a compound made up of only carbon and hydrogen atoms [1]

`CO and CO₂ contain no hydrogen; CH₃OH, CH₃CH₂OH and CH₃CH₂CO₂H contain oxygen. CH₄ is the only compound made of C and H only (it is the simplest alkane).`

CH_4 or methane. Reject every other formula in the list.

Ch 16 Hydrocarbons › Chapter opener › p. 239 (PDF p. 250)

"When a compound is made up of only covalently-bonded carbon and hydrogen atoms, we call it a hydrocarbon."

Ch 15 Fuels and Crude Oil › 15.1 Why Are Natural Gas and Crude Oil Important in Our Lives? › p. 227 (PDF p. 238)

"Fossil fuels are mainly organic compounds that contain only carbon and hydrogen, known as hydrocarbons. When fossil fuels are burnt in the presence of oxygen, the main products formed are carbon dioxide and water."

Ch 15 Fuels and Crude Oil › 15.1 Why Are Natural Gas and Crude Oil Important in Our Lives? › p. 227 (PDF p. 238)

"Natural gas is a gaseous fossil fuel consisting mainly of methane."

Ch 16 Hydrocarbons › 16.2 What Are Alkanes? › p. 242 (PDF p. 253)

"Alkanes are hydrocarbons that contain only carbon–carbon single bonds and carbon–hydrogen bonds."

In the textbook

LO: 11.2(b)

1(e) **[1]**

Identify a compound that is formed by the fermentation of aqueous glucose.

- CH_3CH_2OH (ethanol): in alcohol fermentation, yeast acts on glucose to produce ethanol and carbon dioxide [1] (the expected answer; CO_2 is the other fermentation product, see accept and notes)

`Fermentation (yeast, absence of oxygen): C₆H₁₂O₆(aq) → 2C₂H₅OH(aq) + 2CO₂(g). Two compounds in the list are formed: ethanol and carbon dioxide.`

Full accepted set: CH_3CH_2OH / C_2H_5OH / ethanol (the examiners' most likely key, because the syllabus outcome is 'formation of ethanol by fermentation of glucose', and the book's own practice question calls carbon dioxide the 'other product') and CO_2 / carbon dioxide (chemically also formed, it is in the list, and the book names it as a product; a careful scheme should credit it). Reject CH_3OH (methanol), $CH_3CH_2CO_2H$ and CO. A candidate who writes both ethanol and carbon dioxide has given two correct compounds and should score. Teach students to answer ethanol.

Ch 17 Alcohols and Carboxylic Acids › 17.1 What Are Alcohols? (Formation of Ethanol: Fermentation of Glucose) › p. 263 (PDF p. 274)

"In alcohol fermentation, yeast acts on carbohydrates such as glucose to produce ethanol and carbon dioxide. This process takes place in the absence of oxygen."

Ch 17 Alcohols and Carboxylic Acids › 17.1 What Are Alcohols? (Formation of Ethanol: Fermentation of Glucose) › p. 263 (PDF p. 274)

"The fermentation of glucose can be represented by the following equation: glucose → ethanol + carbon dioxide $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$"

Ch 17 Alcohols and Carboxylic Acids › 17.1 What Are Alcohols? (Formation of Ethanol: Fermentation of Glucose) › p. 263 (PDF p. 274)

"Ethanol and carbon dioxide gas are the products of the fermentation reaction."

Ch 17 Alcohols and Carboxylic Acids › 17.1 What Are Alcohols? (Let's Practise 17.1) › p. 264 (PDF p. 275)

"2 Ethanol can be obtained from the fermentation of glucose. (a) What are the conditions required in this reaction? (b) What other product is formed in the fermentation of glucose?"

In the textbook

LO: 11.3(d)

Question 2 — Reactivity series: Figure 2.1 and metal E [5]

2(a) **[1]**

Name a metal not shown in Figure 2.1 that is more reactive than sodium.

- potassium: in the reactivity series potassium is above sodium (most reactive at the top); in Group 1 the reactivity of the alkali metals increases down the group [1]

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Figure 2.1 lists sodium, calcium, magnesium, E, iron, lead, copper, silver. The textbook's
reactivity series (Fig. 12.2) puts potassium above sodium, and the Group 1 order is lithium <
sodium < potassium < rubidium < caesium < francium.
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potassium / K. Also chemically correct and credited by a fair marker: rubidium, caesium, francium (they are below sodium in Group 1). Reject lithium (less reactive than sodium), and any metal already in Figure 2.1 (calcium, magnesium, ...), which the question excludes.

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? › p. 184 (PDF p. 195)

"Most reactive potassium (K) sodium (Na) calcium (Ca) magnesium (Mg) carbon (C) zinc (Zn) iron (Fe) lead (Pb) hydrogen (H) copper (Cu) silver (Ag) Least reactive Figure 12.2 The reactivity series"

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? › p. 184 (PDF p. 195)

"In the reactivity series, metals are arranged from the most reactive to the least reactive."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 1 Elements - Alkali Metals) › p. 175 (PDF p. 186)

"From Table 11.4, we can see that the reactivity of alkali metals increases down the group (Figure 11.10). These reactions also explain why they are called alkali metals. All the metals react with water to form soluble bases (alkalis). Order of reactivity: lithium < sodium < potassium < rubidium < caesium < francium"

In the textbook

LO: 8.4(a)

2(b) **[1]**

Name the metal in Figure 2.1 that is the most difficult to obtain from its ore.

- sodium: it is the most reactive metal in Figure 2.1; the more reactive the metal is, the harder it is to extract the metal from its ore (its compounds are very difficult to break down and can only be extracted by electrolysis) [1]

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The arrow in Figure 2.1 points up to 'most reactive', so sodium (top) is the most reactive
metal shown. The metals lower in the series are extracted by reducing their oxides with carbon.
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sodium / Na. Reject silver (the least reactive, easiest to obtain), and potassium (not in Figure 2.1).

Ch 12 The Reactivity Series › 12.3 How Are Metals Extracted From Their Ores? › p. 190 (PDF p. 201)

"There are two main methods for extracting metals from their ores: • reducing the metal compound (ore) to the metal using carbon • electrolysis — using electricity to decompose the molten metal compound (ore) to the metal The method to extract the metals from their ores is determined by the position of the metal in the reactivity series. The more reactive the metal is, the harder it is to extract the metal from its ore. The oxides of metals lower in the reactivity series are more easily reduced by carbon than the oxides of more reactive metals."

Ch 12 The Reactivity Series › 12.3 How Are Metals Extracted From Their Ores? › p. 191 (PDF p. 202)

"Metals high up in the reactivity series are reactive. Their compounds are very difficult to break down. These metals can only be extracted by electrolysis."

Ch 12 The Reactivity Series › 12.3 How Are Metals Extracted From Their Ores? › p. 191 (PDF p. 202)

"Metals which are less reactive, i.e. from the middle to the bottom of the reactivity series can be extracted by reducing their oxides with carbon."

Ch 12 The Reactivity Series › 12.3 How Are Metals Extracted From Their Ores? › p. 191 (PDF p. 202)

"Metals above carbon in the reactivity series can be extracted from their ores using electrolysis. Metals below carbon can be extracted by reduction with carbon."

In the textbook

LO: 8.4(e)

2(c) **[1]**

Predict the products of the reaction between metal E and steam.

- metal oxide and hydrogen (metal + steam → metal oxide + hydrogen; for zinc, zinc oxide and hydrogen), both needed [1]

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E sits between magnesium and iron. In the textbook's series (Fig. 12.2) the only metal in that
gap is zinc, so E behaves like zinc: Zn(s) + H₂O(g) → ZnO(s) + H₂(g). E need not be named; the
mark is for the two products.
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E oxide / metal oxide / zinc oxide AND hydrogen (H_2). If a candidate guesses aluminium, aluminium oxide + hydrogen is the same type of answer. Reject metal hydroxide + hydrogen (that is the reaction with cold water, and metals in this region do not react with cold water), hydrogen only, or an oxide with water.

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? › p. 184 (PDF p. 195)

"Some metals such as zinc and iron do not react with cold water but react with steam. More reactive metals react more violently with steam. metal + steam → metal oxide + hydrogen"

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? › p. 185 (PDF p. 196)

"magnesium Reacts very slowly with cold water to form magnesium hydroxide and hydrogen gas. A test tube of hydrogen gas is produced only after a few days. $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$ Hot magnesium reacts violently with steam to form magnesium oxide (a white solid) and hydrogen gas. A bright white glow is produced during the reaction. $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$ zinc No reaction occurs. Hot zinc reacts readily with steam to produce zinc oxide and hydrogen. Zinc oxide is yellow when hot and white when cold. $Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$ iron No reaction occurs. Red-hot iron reacts slowly with steam to form iron oxide. The iron must be constantly heated for the reaction to progress. $3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$"*

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? › p. 187 (PDF p. 198)

"zinc (Zn) no reaction with cold water reacts readily with steam reacts moderately fast with dilute hydrochloric acid"

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? (Let's Practise 12.1) › p. 187 (PDF p. 198)

"1 Name a metal that reacts with steam but not with cold water to give an oxide and hydrogen."

In the textbook

LO: 8.4(a)

2(d) **[2]**

State the two essential conditions needed for the corrosion of iron.

- oxygen (in air) [1]
- water [1] (the presence of both oxygen (in air) and water are necessary for rusting to occur)

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Rusting needs both oxygen and water together: iron + oxygen + water → hydrated iron(III) oxide (rust).
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oxygen / air (the book writes 'oxygen (in air)'; a strict marker may want 'oxygen'); water / moisture / damp / water vapour. Reject salt, acid rain, carbon dioxide, sulfur dioxide and 'high temperature': the book says sodium chloride and acidic substances speed rusting up, but they are not essential. Each condition must be stated; 'air and water' is normally accepted.

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? › p. 192 (PDF p. 203)

"The presence of both oxygen (in air) and water are necessary for rusting to occur."

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? › p. 192 (PDF p. 203)

"Metals may react with oxygen in the air and water and corrode. When an object made of iron is left in moist air for some time, a red-brown substance slowly forms on the surface of the metal (Figure 12.11). This substance is called rust and has the chemical name hydrated iron(III) oxide."

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? › p. 192 (PDF p. 203)

"Rusting or the corrosion of iron is the process that produces rust. The equation for rusting is: iron + oxygen + water → hydrated iron(III) oxide (rust)"

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? › p. 193 (PDF p. 204)

"From the results of Let's Investigate 12A, we can conclude that: • both oxygen and water are needed for rusting to occur; and • the presence of sodium chloride increases the rate of rusting. Other than sodium chloride, acidic substances such as sulfur dioxide and carbon dioxide also speed up the rusting process."

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? › p. 193 (PDF p. 204)

"Table 12.5 Results of Let's Investigate 12A Test Tube Conditions Observation Is Water Present? Is Oxygen Present? A yes no no rusting B no yes no rusting C yes yes nails rusted D yes yes nails rusted heavily"

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? (Barrier Methods) › p. 194 (PDF p. 205)

"The barrier methods prevent rusting by keeping iron and steel away from oxygen (in air) and water (Figure 12.14)."

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? (Barrier Methods) › p. 194 (PDF p. 205)

"The paint protects the iron gate from moisture and oxygen in the air. It prevents the gate from rusting."

Ch 12 The Reactivity Series › 12.4 What Are the Conditions for Rusting? (Rusting as a Special Case of Corrosion) › p. 195 (PDF p. 206)

"Rusting only occurs in iron and steel. It involves both water and oxygen."

In the textbook

LO: 8.4(f)

Question 3 — Copper(II) carbonate: relative formula mass, moles and reaction with acid [7]

3(a)(i) **[1]**

Explain the term relative formula mass.

- the relative formula mass (Mr) of an ionic compound is the average mass of one unit of that compound relative to 1/12 the mass of an atom of carbon-12; it is found by adding together the relative atomic masses (Ar) of all the atoms in its chemical formula [1]

Either wording scores: 'adding up the relative atomic masses of all the atoms in the formula' (the syllabus wording: the sum of relative atomic masses), or 'average mass of one formula unit compared with 1/12 of a carbon-12 atom'. Reject 'mass of one mole' (that is molar mass, in g/mol), 'total mass of the compound' with no mention of relative atomic masses, and sums of proton or nucleon numbers. 'Formula' (not 'molecular') is the term for an ionic compound such as $CuCO_3$.

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Formula Mass) › p. 107 (PDF p. 118)

"The relative formula mass (Mr) of an ionic compound is the average mass of one unit of that compound relative to 1/12 the mass of an atom of carbon-12."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Formula Mass) › p. 107 (PDF p. 118)

"Relative formula mass is calculated the same way as relative molecular mass and also has the symbol Mr (Table 7.2)."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Formula Mass) › p. 107 (PDF p. 118)

"Ionic compounds do not exist as molecules. We use the term relative formula mass to refer to the relative mass of an ionic compound."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Molecular Mass) › p. 107 (PDF p. 118)

"To find the relative molecular mass (Mr) of a substance, we add together the relative atomic masses (Ar) of all the atoms in its chemical formula (Table 7.1)."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Molecular Mass) › p. 107 (PDF p. 118)

"The relative molecular mass (Mr) of a molecular substance is the average mass of one molecule of that substance relative to 1/12 the mass of an atom of carbon-12."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? › p. 106 (PDF p. 117)

"Define relative molecular mass, Mr, and calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? › p. 106 (PDF p. 117)

"The relative atomic mass (Ar) of an element is the average mass of one atom of that element relative to 1/12 the mass of an atom of carbon-12."

In the textbook

LO: 4.2(b)

3(a)(ii) **[1]**

Show, using calculations, that the Mr of $CuCO_3$ is 124.

- Mr = 64 + 12 + (3 × 16) = 124: add together the relative atomic masses (Ar) of all the atoms in the chemical formula [1]

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Ar values from the Periodic Table on page 20 of the paper: Cu 64, C 12, O 16. 64 + 12 + 48 =
124. The working must be shown because the answer 124 is given in the question.
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Any layout that shows 64 (Cu), 12 (C) and 3 × 16 (O) added to give 124. Reject the bare number 124 with no working.

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Molecular Mass) › p. 107 (PDF p. 118)

"To find the relative molecular mass (Mr) of a substance, we add together the relative atomic masses (Ar) of all the atoms in its chemical formula (Table 7.1)."

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Formula Mass) › p. 107 (PDF p. 118)

"Relative formula mass is calculated the same way as relative molecular mass and also has the symbol Mr (Table 7.2)."

Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Worked Example 7A) › p. 110 (PDF p. 121)

"Calculate the mass, in grams, of 4.5 moles of lead(II) sulfate ($PbSO_4$). Answer Molar mass of $PbSO_4$ in g/mol = 207 + 32 + 4(16) = 303 g/mol Mass of $PbSO_4$ in g = number of moles of $PbSO_4$ × molar mass of $PbSO_4$ in g/mol = 4.5 × 303 = 1360 g"

Ch 7 Mole Concept and Stoichiometry › 7.1 What Are Relative Atomic Mass and Relative Molecular Mass? (Relative Formula Mass) › p. 107 (PDF p. 118)

"Since the relative atomic mass (Ar), relative molecular mass and relative formula mass (Mr) are comparisons of mass, we do not assign them with a unit."

In the textbook

LO: 4.2(b)

3(b)(i) **[1]**

Calculate the amount, in moles, of copper(II) carbonate in 1.24 g.

- number of moles = mass of the substance in g / molar mass of the substance in g/mol = 1.24 / 124 = 0.0100 mol [1]

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Molar mass of CuCO₃ = Mr = 124 g/mol. Number of moles = 1.24 ÷ 124 = 0.0100 mol.
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0.0100, 0.01 or 1.0 × 10^{-2} mol. Unit mol is printed on the answer line.

Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Moles and Molar Mass) › p. 109 (PDF p. 120)

"One mole of any substance has a mass equal to its relative atomic mass (Ar), relative molecular mass or relative formula mass (Mr) in grams. This mass is the molar mass of the substance. The unit for molar mass is g/mol (Table 7.4)."

Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Moles and Molar Mass) › p. 109 (PDF p. 120)

"Number of moles of a substance = mass of the substance in g / molar mass of the substance in g/mol"

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Calculations from Chemical Equations) › p. 114 (PDF p. 125)

"Step 1 Calculate the number of moles of the reactant or product using the given information. In this case, we find the number of moles of $MgCO_3$."

In the textbook

LO: 4.2(f)

3(b)(ii) **[1]**

Calculate the mass of copper(II) oxide, CuO, formed when all the $CuCO_3$ decomposes.

- from the balanced equation 1 mol of $CuCO_3$ forms 1 mol of CuO, so 0.0100 mol of CuO; mass = number of moles × molar mass = 0.0100 × 80 = 0.80 g [1]

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CuCO₃(s) → CuO(s) + CO₂(g): mole ratio 1 : 1, n(CuO) = 0.0100 mol. Mr(CuO) = 64 + 16 = 80. Mass = 0.0100 × 80 = 0.80 g. Check by conservation of mass: CO₂ lost = 0.0100 × 44 = 0.44 g; 1.24 − 0.44 = 0.80 g. Alternative one-step route: 1.24 × 80/124 = 0.80 g.
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0.80 g, 0.8 g or 0.800 g. Error carried forward: 80 × the amount found in (b)(i). Unit g is printed on the answer line.

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Calculations from Chemical Equations) › p. 114 (PDF p. 125)

"Step 2 Find the number of moles of the substance we are solving for using the mole ratio from the balanced equation and the result from step 1. In this case, we need to find the number of moles of CO_2. From the equation, 1 mol of $MgCO_3$ will form 1 mol of CO_2. Number of moles of CO_2 = number of moles of $MgCO_3$ = 2 mol"

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Calculations from Chemical Equations) › p. 114 (PDF p. 125)

"A balanced chemical equation shows the mole ratio of the reactants and products in a chemical reaction. We can use these ratios to calculate certain measurements such as mass and volume."

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Calculations from Chemical Equations) › p. 114 (PDF p. 125)

"This ratio between the quantities of the different substances in the reaction is the mole ratio. Combining this ratio with the mole concept allows us to find the masses, volumes and concentrations of substances in a chemical reaction."

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Worked Example 7E) › p. 115 (PDF p. 126)

"Iron(III) oxide, $Fe_2O_3(s)$ reacts with carbon monoxide gas to produce iron and carbon dioxide gas. Calculate the mass of iron produced when 2000 g of iron(III) oxide is reacted completely. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$"

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Worked Example 7E) › p. 115 (PDF p. 126)

"From the equation, 1 mol of Fe_2O_3 produces 2 mol of Fe. Number of moles of Fe = 2 × 12.5 = 25.0 mol Mass of Fe = number of moles of Fe × Ar of Fe = 25.0 × 56 = 1400 g"

Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Worked Example 7A) › p. 110 (PDF p. 121)

"Mass of $PbSO_4$ in g = number of moles of $PbSO_4$ × molar mass of $PbSO_4$ in g/mol"

In the textbook

LO: 4.2(f), 4.1(d)

3(c) **[3]**

Name the three products of the reaction between copper(II) carbonate and dilute hydrochloric acid.

- copper(II) chloride (the salt; chlorides are formed from hydrochloric acid) [1]
- water [1]
- carbon dioxide [1] (acids react with carbonates to form a salt, water and carbon dioxide gas)

carbonate + acid → salt + water + carbon dioxide: $CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l) + CO_2(g)$.

Any order. 'copper chloride' is accepted; correct formulae ($CuCl_2$, H_2O, CO_2) are accepted. Reject copper(I) chloride / CuCl, hydrogen, chlorine, carbonic acid and 'hydrochloric acid' (a reactant). Adding a wrong extra product to a correct three can cost a mark.

Ch 8 Acids and Bases › 8.1 What Is an Acid? (Properties of Acids) › p. 126 (PDF p. 137)

"6 Acids react with carbonates to form a salt, water and carbon dioxide gas. The general equation for the reaction of carbonates with acids is: carbonate + acid → salt + water + carbon dioxide For example, when hydrochloric acid is added to calcium carbonate, the following reaction occurs: calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ We can test for carbon dioxide gas by bubbling the gas through limewater."

Ch 8 Acids and Bases › 8.1 What Is an Acid? (Properties of Acids) › p. 125 (PDF p. 136)

"A salt is an ionic compound that can be obtained from acid reactions. Different salts are formed when various acids are used in the reactions: • Sulfates are formed from sulfuric acid. • Nitrates are formed from nitric acid. • Chlorides are formed from hydrochloric acid."

In the textbook

LO: 5.1(d)

Question 4 — Bromine: isotopes, sub-atomic particles, the bromide ion and Group 17 trends [8]

4(a) **[1]**

Define the term isotopes.

- isotopes are atoms of the same element that have the same proton (atomic) number but different nucleon (mass) numbers, i.e. different numbers of neutrons [1]

'same proton (atomic) number but different nucleon (mass) number' scores. Both ideas are needed: same element/same protons AND different neutrons/nucleon number. Reject 'elements with different numbers of neutrons' (they are atoms of the same element), 'same atomic mass', and 'different numbers of electrons'.

> **Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Isotopes) › p. 42 (PDF p. 53)**
> *"Isotopes are atoms of the same element that have the same proton (atomic) number but different nucleon (mass) numbers. This means that they have different numbers of neutrons."*

> **Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Isotopes) › p. 42 (PDF p. 53)**
> *"We have learnt that: • protons and neutrons are found clustered together in an atom's nucleus; and • the number of protons is fixed for all atoms of a given element. However, the number of neutrons in the nucleus can vary. Neutrons do not affect the electrical charge of the atom, although they affect the total mass."*

> **Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Isotopes) › p. 43 (PDF p. 54)**
> *"Isotopes of the same element have the same number of protons. Hence, they also have the same number of electrons. This means that they undergo the same types of chemical reactions even though they do not have the same mass. However, isotopes may have different physical properties such as density, melting point and boiling point."*

> **Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? › p. 40 (PDF p. 51)**
> *"This number is also called the atomic number. It tells us what element the atom comes from so that we know what type of atom it is. All atoms of the same element have the same proton number."*

In the textbook

LO: 2.2(e)

4(b) **[3]**

Complete Table 4.1 (electrons, neutrons, protons) for the isotope $^{79}_{35}Br$.

- number of electrons = 35 (the atom is electrically neutral, so the number of electrons is the same as the number of protons) [1]
- number of neutrons = 44 (nucleon number − proton number = 79 − 35) [1]
- number of protons = 35 (the proton number) [1]

```
Proton number 35 = number of protons = 35. Neutrons = nucleon number – proton number = 79 – 35
= 44. The isotope is a neutral atom, so electrons = protons = 35.
```

One mark per row. Reject 36 electrons (that is Br^-) and 79 neutrons (that is the nucleon number).

> **Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? › p. 40 (PDF p. 51)**
> *"The proton number (Z) of an atom is the number of protons in its nucleus."*

> **Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? › p. 40 (PDF p. 51)**
> *"The nucleon number (A) is the total number of protons and neutrons in the nucleus. Nucleon number (A) = proton number (Z) + number of neutrons"*

Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? › p. 40 (PDF p. 51)

"We can use proton and nucleon numbers to calculate the number of neutrons in an atom. Number of neutrons = Nucleon number (A) – Proton number (Z)"

Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? › p. 40 (PDF p. 51)

"There are 6 protons and 6 neutrons in the nucleus. Hence, carbon has a proton number of 6 and a nucleon number of 12 (6 protons + 6 neutrons). Since a carbon atom is neutral, it will have the same number of electrons as protons (6 electrons)."

Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Nuclide Notation) › p. 41 (PDF p. 52)

"We can represent the proton and nucleon numbers of an atom using nuclide notation (Figure 3.8)."

Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Worked Example 3A) › p. 41 (PDF p. 52)

"Thought Process • This atom has a nucleon number of 184. There must be a total of 184 protons and neutrons. • The proton number of 74 tells us that it has 74 protons. • We can subtract the proton number from the nucleon number to find the number of neutrons. • The atom is electrically neutral, so the number of electrons is the same as the number of protons. Answer Number of neutrons in tungsten = 184 – 74 = 110 An atom of tungsten has 74 protons, 110 neutrons and 74 electrons."

Ch 3 Atomic Structure › 3.1 What Is an Atom Made Up Of? › p. 38 (PDF p. 49)

"Atoms are electrically neutral. The number of protons in an atom is always the same as the number of electrons. Number of protons in an atom = number of electrons in that atom"

In the textbook

LO: 2.2(f), 2.2(d)

4(c) **[2]**

Explain why the bromide ion, Br^-, is negative, using ideas about the particles in Br^- and their charges.

- Br^- has 35 protons (each +1) and 36 electrons (each –1): the ion has more electrons than protons [1]
- so it has more negative charges than positive charges (net charge 35 − 36 = –1): the ion formed is negatively charged because it has more electrons than protons [1]

```
The bromine atom has 35 protons and 35 electrons (neutral). It gains one electron to become
Br⁻: 35 protons (+35) and 36 electrons (−36), net charge = +35 − 36 = −1. Neutrons carry no
charge and do not change the overall charge. The book's chloride example (17 protons, 18
electrons, charge −1) is exactly the same argument.
```

'gains one electron so it has one more electron than protons' scores the first mark; the second needs the comparison of negative and positive charge (or 'electrons are negative, protons are positive, so there is an excess of negative charge'). 'It has gained an electron' alone describes the process and does not use the numbers or the charges.

Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Negative Ions) › p. 56 (PDF p. 67)

"A negative ion or anion is formed when an atom or combination of atoms gains one or more electrons. The ion formed is negatively charged, because it has more electrons than protons. Most non-metal atoms will gain electrons and form negative ions."

Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Negative Ions) › p. 56 (PDF p. 67)

"After gaining 1 electron, the chlorine atom is no longer electrically neutral. It now has 17 protons and 18 electrons, so it has a charge of –1. The negative ion formed is known as a chloride ion. It has the symbol Cl^-."

Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Negative Ions) › p. 56 (PDF p. 67)

"Negative ions (anions) have a net negative charge and usually have a noble gas electronic configuration."

Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Sub-atomic Particles in Ions) › p. 42 (PDF p. 53)

"Recall that electrons have a charge of –1. Hence, when an atom gains electrons, it will have more negative charges, –1 for each electron gained."

Ch 3 Atomic Structure › 3.2 How Many Sub-atomic Particles Does an Atom Have? (Sub-atomic Particles in Ions) › p. 42 (PDF p. 53)

"An oxide ion has the symbol O^{2-}. This means that it has 8 protons and 8 neutrons (just like an oxygen atom), but with a net charge of –2. Thus, it has a total of 2 more electrons than protons. It has 8 + 2 = 10 electrons. Therefore, the charge numbers on an ion indicates • the number of electrons that have been lost in a positive ion; and • the number of electrons that have been gained in a negatively charged ion."

Ch 3 Atomic Structure › 3.1 What Is an Atom Made Up Of? › p. 38 (PDF p. 49)

"Protons • Carry a positive charge (+1) Neutrons • Neutral particles with no charge (0) • Found in the nucleus between protons Electrons • Extremely tiny particles that move very quickly • Carry a negative charge (–1)"

In the textbook

LO: 2.2(f), 2.2(a), 3.1(a)

4(d) **[2]**

Describe two trends in the properties of the Group 17 elements from chlorine to iodine.

- Any TWO of the following trends, 1 mark each [2]:
- going down the group the colours become darker (colour intensities increase): chlorine yellow-green, bromine red-brown, iodine purple-black [1]
- state at r.t.p. changes from gas (chlorine) to liquid (bromine) to solid (iodine) [1]
- the melting and boiling points of the halogens increase [1]
- the reactivity of halogens decreases down the group [1]

Table 11.5 of the textbook: chlorine yellow-green gas (m.p. −101 °C, b.p. −34 °C); bromine red-brown liquid (−7 °C, 59 °C); iodine purple-black solid (114 °C, 184 °C). Reactivity: Cl_2 displaces Br^- and I^-, Br_2 displaces only I^-, I_2 displaces neither.

Each trend needs its direction (darker, increases, decreases). Reject 'colour changes' with no direction, 'gets lighter', 'reactivity increases'. 'State changes from gas to solid' and 'melting/boiling points increase' are one physical trend seen two ways, so a cautious marker may credit only one of them: the safest two come from different families (colour, state or melting/boiling point, reactivity). Other true trends (density rises, atoms get bigger) are not stated as trends in the book's halogen section; do not rely on them.

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Halogens) › p. 176 (PDF p. 187)

"Halogen Melting Point / °C Boiling Point / °C Appearance at r.t.p. chlorine –101 –34 yellow-green gas bromine –7 59 red-brown liquid iodine 114 184 purple-black solid"

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Halogens) › p. 176 (PDF p. 187)

"From Table 11.5, we can see that going down the group, • the melting and boiling points of the halogens increase; and • the colours become darker (colour intensities increase)."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Halogens) › p. 176 (PDF p. 187)

"Figure 11.12 The melting points, boiling points and colour intensities of the halogens increase down the group."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Chemical Properties of Halogens) › p. 177 (PDF p. 188)

"Unlike the alkali metals, the reactivity of halogens decreases down the group."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Chemical Properties of Halogens) › p. 177 (PDF p. 188)

"Order of reactivity: fluorine > chlorine > bromine > iodine > astatine > tennessine Halogens undergo displacement reactions with halide solutions."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Chemical Properties of Halogens) › p. 177 (PDF p. 188)

"In a displacement reaction, one element takes the place of another element in the compound. A more reactive halogen will displace a less reactive halogen from its halide solution."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Chemical Properties of Halogens) › p. 177 (PDF p. 188)

"Chlorine displaces bromine from a bromide solution. Chlorine displaces iodine from an iodide solution. Bromine displaces iodine from an iodide solution."

Ch 11 The Periodic Table › 11.2 What Group Trends Are There? (Group 17 Elements - Chemical Properties of Halogens) › p. 177 (PDF p. 188)

"Figure 11.13 Going down Group 17, the halogens become less reactive."

In the textbook

LO: 8.2(b), 8.1(g)

Question 5 — Alkenes: general formula, hydrogenation, combustion, addition polymer and recycling [8]

5(a) **[1]**

State the general formula for the homologous series of alkenes.

- CnH2n (n = 2, 3 and so on), where n is the number of carbon atoms in each molecule [1]

CnH2n (n = 2, 3, ...). Reject CnH2n+2 (alkanes), CnH2n−2 and a single formula such as C_2H_4.

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? (General Formula of Alkenes) › p. 247 (PDF p. 258)

"The general molecular formula of alkenes with one carbon−carbon double covalent bond is CnH2n, where n is the number of carbon atoms in each molecule (n = 2, 3 and so on). When there are two carbon atoms in each molecule (n = 2): • The number of hydrogen atoms will be 2(2) = 4. • Hence, the formula of the alkene is C_2H_4."

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? › p. 246 (PDF p. 257)

"Alkenes are hydrocarbons that contain carbon–carbon double (C=C) bonds."

In the textbook

LO: 11.2(f)

5(b) **[1]**

Draw the full structural formula of the product of ethene and hydrogen, showing all the bonds.

- ethane drawn with every bond shown: alkenes react with hydrogen to form alkanes; ethane has only carbon–carbon single bonds and carbon–hydrogen bonds: two C atoms joined by a single C–C bond, each C bonded to three H atoms (six C–H bonds, seven bonds in all), with no C=C [1]

$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$ (hydrogenation: about 150 °C, nickel catalyst). The C=C double bond opens and each carbon gains one hydrogen. Drawn as H–C(H)(H)–C(H)(H)–H with each C–H bond shown as a line.

Any orientation of the displayed ethane molecule, as long as every C–H line touches both atoms. Reject CH_3CH_3 or C_2H_6 (not all bonds shown), a drawing with a missing bond or a floating H, and a structure that still contains C=C.

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? (Chemical Properties of Alkenes: Addition Reactions) › p. 249 (PDF p. 260)

"Addition of Hydrogen (Hydrogenation) Alkenes react with hydrogen to form alkanes under the following conditions: • a temperature of 150 °C • in the presence of a catalyst such as nickel This reaction is known as hydrogenation. For example, ethene reacts with hydrogen as follows: ethene + hydrogen → ethane $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$"

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? (Chemical Properties of Alkenes: Addition Reactions) › p. 249 (PDF p. 260)

"Addition Reactions The carbon–carbon double bonds in alkenes are reactive. Hence, alkenes will readily undergo addition reactions. The general equation is as follows:"

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? (Chemical Properties of Alkenes: Addition Reactions) › p. 249 (PDF p. 260)

"• Other atoms can add to the carbon atoms in the C=C bond. • The carbon–carbon double bond breaks to form new single bonds. Thus, each carbon atom which was originally part of the C=C bond forms four single covalent bonds."

Ch 16 Hydrocarbons › 16.2 What Are Alkanes? › p. 242 (PDF p. 253)

"A carbon–carbon single bond is a single covalent bond between two carbon atoms. Figure 16.6 shows the structure of ethane."

Ch 16 Hydrocarbons › 16.2 What Are Alkanes? › p. 242 (PDF p. 253)

"Full structural formula of ethane, an example of an alkane."

In the textbook

LO: 11.2(j), 11.2(c)

5(c) **[2]**

Name the two products of the complete combustion of ethene.

- carbon dioxide [1]
- water [1] (alkenes burn in excess oxygen to produce carbon dioxide and water)

$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$.

CO_2 and H_2O (formulae) are accepted for a 'name' question; steam / water vapour accepted for water. Reject carbon monoxide, carbon or soot (products of incomplete combustion) and hydrogen.

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? (Chemical Properties of Alkenes: Combustion) › p. 248 (PDF p. 259)

"Combustion Like alkanes, alkenes burn in excess oxygen to produce carbon dioxide and water. For example, ethene burns in excess oxygen as follows: ethene + oxygen → carbon dioxide + water $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$"

Ch 16 Hydrocarbons › 16.3 What Are Alkenes? (Chemical Properties of Alkenes: Combustion) › p. 248 (PDF p. 259)

"Compared to alkanes, the percentage by mass of carbon in alkene molecules is higher. Alkenes are more likely to undergo incomplete combustion. Incomplete combustion produces carbon monoxide and carbon. Hence, alkenes usually burn with a sootier flame than alkanes (Figure 16.15)."

In the textbook

LO: 11.2(j)

5(d) **[2]**

But-2-ene is polymerised to make poly(but-2-ene). Draw the structure of poly(but-2-ene).

- correct repeat unit: the two carbon atoms of the C=C joined by a single C–C bond, each carbon bonded to one H and one CH_3 (an alkyl group on each carbon: a $–CH(CH_3)–CH(CH_3)–$ unit); the carbon–carbon double bond has broken open, no C=C left [1]

- drawn as a polymer: the repeat unit in square brackets with continuation (extension) bonds through the brackets at both ends and n written at the bottom right (or at least two repeat units with continuation bonds at each end) [1]

Monomer $CH_3CH=CHCH_3$, written in the book's pattern with a group on each of the four bonds of C=C (here H, CH_3, H, CH_3). The C=C opens; the two carbons become the 2-carbon backbone and the four groups hang off it: $-[CH(CH_3)-CH(CH_3)]n-$. The two CH_3 groups are side groups (alkyl groups), not part of the chain, so the backbone is two carbons per repeat unit, not four.

H and CH_3 may be drawn above or below the backbone in either arrangement (cis/trans arrangement of the side groups is not tested). Reject a retained C=C, a four-carbon backbone such as $–CH_2–CH=CH–CH_2–$, and brackets with no n and no continuation bonds (second mark lost). A full-structural (displayed) formula is the safe layout, as in the textbook.

Ch 18 Polymers › 18.2 How Are Addition Polymers Formed and Used? (Deducing the Structure of a Polymer from its Monomer) › p. 277 (PDF p. 288)

"It is easier to deduce the structure of an addition polymer if the structural formula of its alkene monomer is written in the following way: [diagram: C=C with a box on each of its four bonds] (where [box] = hydrogen, alkyl group, halogen etc.) For example, the formula of propene can be written as [diagram]. The polymer formed, which is poly(propene), can then be represented as [diagram]."

Ch 18 Polymers › 18.2 How Are Addition Polymers Formed and Used? (Addition Polymerisation) › p. 275 (PDF p. 286)

"Addition Polymerisation When monomers join together without losing any molecules or atoms, the polymer formed is called an addition polymer. This process is called addition polymerisation. Addition polymerisation occurs when unsaturated monomers join together without losing any molecules or atoms to form addition polymers. The monomer used for addition polymerisation is usually unsaturated. This means that it contains either a carbon–carbon double bond or a carbon–carbon triple bond. Thus, alkenes can undergo addition polymerisation. At high pressure and temperature, thousands of alkene molecules join together to form polymers."

Ch 18 Polymers › 18.2 How Are Addition Polymers Formed and Used? (Addition Polymerisation) › p. 275 (PDF p. 286)

"Ethene is unsaturated as it contains a carbon–carbon double bond. Poly(ethene) is saturated as it contains only carbon–carbon single bonds."

Ch 18 Polymers › 18.2 How Are Addition Polymers Formed and Used? (Addition Polymerisation) › p. 276 (PDF p. 287)

"A polymer is usually made up of similar units joined together. These units are known as repeat units."

Ch 18 Polymers › 18.2 How Are Addition Polymers Formed and Used? (Addition Polymerisation) › p. 276 (PDF p. 287)

"The formula of a polymer can be written using its repeat unit. The letter n represents a large number. Many ethene molecules are required to form each poly(ethene) molecule."

Ch 18 Polymers › 18.2 How Are Addition Polymers Formed and Used? (Let's Practise 18.2) › p. 279 (PDF p. 290)

"2 Figure 18.14 represents part of a polymer chain. Draw the structural formula of its monomer."

In the textbook

LO: 11.4(d), 11.4(b)

5(e)(i) **[1]**

Describe one physical method of recycling polymers.

- after pre-treatment, small pieces of plastic (such as poly(ethene)) are melted, cooled, pulled into long, thin strands and cut into pellets; the pellets can then be made into new products [1]

melting the plastic and remoulding it / making pellets for new products; 'mechanical recycling'. Reject burning or incineration, 'reuse' or 'reduce' (not recycling), sorting or washing alone (that is pre-treatment, not the method itself), and cracking (the chemical method).

Ch 18 Polymers › 18.4 How Are Plastics Recycled and What Are the Issues Related to Recycling Plastics? (Recycling Plastics: Physical Method) › p. 282 (PDF p. 293)

"Physical Method The physical method of recycling plastics is also known as mechanical recycling. Using this method, the chemical composition of the recycled plastic remains unchanged. This means that if waste poly(ethene) bottles are collected, the recycled plastic products will still be made of poly(ethene). After pre-treatment, small pieces of plastics such as poly(ethene) are melted, cooled, pulled into long, thin strands, and cut into pellets (Figure 18.18). These pellets can then be made into new products."

Ch 18 Polymers › 18.4 How Are Plastics Recycled and What Are the Issues Related to Recycling Plastics? (Recycling Plastics: Physical Method) › p. 282 (PDF p. 293)

"Recycling of plastics can be carried out using the physical method or the chemical method. Before recycling is carried out, the plastic waste needs to undergo pre-treatment (Figure 18.17). This is because not all plastics can be recycled. The recycling method for different types of plastic can be different."

In the textbook

LO: 11.4(h)

5(e)(ii) **[1]**

Describe one chemical method of recycling polymers.

- plastic waste, such as poly(ethene) waste, undergoes cracking (high temperature and a catalyst) to form short chains of alkanes and alkenes; the short-chain alkanes can be used as fuel, while the short-chain alkenes can be used to make other useful chemicals [1]

cracking of plastic waste into fuel / smaller molecules. Breaking the polymer back into smaller molecules or monomers (depolymerisation, pyrolysis) is chemically sound but is beyond the Science (Chemistry) syllabus wording, which names cracking. Reject melting and remoulding (physical), burning to release energy, and incineration.

Ch 18 Polymers › 18.4 How Are Plastics Recycled and What Are the Issues Related to Recycling Plastics? (Recycling Plastics: Chemical Method) › p. 282 (PDF p. 293)

"Chemical Method The chemical method of recycling plastics converts plastic waste into different raw materials through chemical reactions such as cracking. Plastic waste, such as poly(ethene) waste, can undergo cracking to form short chains of alkanes and alkenes. The short-chain alkanes can be used as fuel, while the short-chain alkenes can be used to make other useful chemicals. Cracking requires a high temperature and the presence of a catalyst."

Ch 18 Polymers › 18.4 How Are Plastics Recycled and What Are the Issues Related to Recycling Plastics? (Recycling Plastics: Chemical Method) › p. 282 (PDF p. 293)

"Recall from Chapter 16: Cracking can be used to break down long-chain hydrocarbons into smaller hydrocarbon molecules."

Ch 16 Hydrocarbons › 16.4 How Do Saturated and Unsaturated Hydrocarbons Differ? / Producing Alkanes and Alkenes by Cracking › p. 250 (PDF p. 261)

"Cracking is a process in which larger hydrocarbon molecules, usually alkanes, are broken down into smaller hydrocarbon molecules."

In the textbook

LO: 11.4(h)

Question 6 — Particles in ice and water vapour; salt and the melting and boiling points of water [6]

6(a) **[3]**

Compare the arrangement, the distance between and the motion of the molecules in solid water (ice) and gaseous water (water vapour).

- Arrangement: in ice the molecules are arranged in an orderly manner; in water vapour they are arranged in a disorderly manner [1]
- Distance: in ice the molecules are very closely packed; in water vapour they are very far apart [1]

- Motion: in ice the molecules vibrate and rotate about fixed positions; in water vapour they move quickly and randomly in any direction [1]

Each mark needs BOTH states compared (ice ... vapour ...). Accept for ice: closely packed, regular pattern / lattice, vibrate in fixed positions. Accept for vapour: irregular, randomly arranged, move freely / at high speed in all directions. Reject: ice molecules do not move at all; gas molecules are bigger or expand; arrangement given as "close together" (that is the distance mark). Forces between molecules are not asked for. The book has no separate "distance" row: "very closely packed" versus "very far apart" in its Particle Arrangement row carries the distance idea, and its "orderly" versus "disorderly" carries the arrangement idea.

Ch 2 Kinetic Particle Theory › 2.1 How Are Solids, Liquids and Gases Different? (Table 2.1; columns Solids | Liquids | Gases) - row Particle Arrangement › p. 25 (PDF p. 36)

"Particle Arrangement | very closely packed in an orderly manner | closely packed in a disorderly manner | very far apart in a disorderly manner"

Ch 2 Kinetic Particle Theory › 2.1 How Are Solids, Liquids and Gases Different? (Table 2.1; columns Solids | Liquids | Gases) - row Particle Movement › p. 25 (PDF p. 36)

"Particle Movement | vibrate and rotate about fixed positions | slide past each other freely throughout the liquid | move quickly and randomly in any direction"

Ch 2 Kinetic Particle Theory › 2.4 What Happens During Solid–Gas Transitions? (Sublimation) › p. 32 (PDF p. 43)

"During sublimation, the particles gain sufficient thermal energy from the surroundings to overcome the strong forces of attraction. Particles in the solid change from vibrating about fixed positions to moving freely and randomly in the gas."

Ch 2 Kinetic Particle Theory › 2.1 How Are Solids, Liquids and Gases Different? (Disciplinary Idea) › p. 25 (PDF p. 36)

"The three physical states of matter (i.e. solid, liquid and gas) could be differentiated based on the arrangement and movement of the particles."

In the textbook

LO: 2.1(a)

6(b)(i) **[1]**

Describe the relationship between the melting point of the mixture and the mass of sodium chloride added (Table 6.1).

- As the mass of sodium chloride added increases, the melting point decreases (adding salt lowers the melting / freezing point); it falls by 1.5 °C for every 1.0 g added [1]

```
Table 6.1: 1.0 g → -1.5 °C; 2.0 g → -3.0 °C; 3.0 g → -4.5 °C; 4.0 g → -6.0 °C; 5.0 g → -7.5 °C.
Each extra 1.0 g lowers the melting point by the same 1.5 °C, so the fall is steady (linear).
```

The direction (decreases) is the mark; "steady / by 1.5 °C per gram" is a bonus. Accept "negative correlation". Reject "melting point increases" and a description with no direction. "Inversely proportional" is technically wrong (the fall is linear, not 1/mass) and is risky; the DECREASE in melting point is directly proportional to the mass added.

Ch 1 Experimental Chemistry › 1.4 How Can the Purity of Substances Be Determined? › p. 20 (PDF p. 31)

"For example, pure water boils at 100 °C. Adding salt to water raises its boiling point. Similarly, pure water freezes at 0 °C, but adding salt to water lowers its freezing point."

Ch 1 Experimental Chemistry › 1.4 How Can the Purity of Substances Be Determined? › p. 19 (PDF p. 30)

"Mixtures containing two or more substances will melt or boil over a range of temperatures. The greater the amount of impurities, the larger the change in melting or boiling points, and the wider the range."

Ch 1 Experimental Chemistry › 1.4 How Can the Purity of Substances Be Determined? › p. 19 (PDF p. 30)

"A pure substance has a specific melting and/or boiling point under fixed conditions."

Data skill (answered from the question)

Gap: The book gives only the qualitative idea (salt lowers the freezing point, p. 20). The pattern, 1.5 °C per gram, has to be read from Table 6.1.
LO: 1.2(e)

6(b)(ii) **[2]**

Use Table 6.1 to estimate the difference between the melting point and the boiling point when 6.0 g of sodium chloride are added. Show your working.

- Extrapolating the constant steps in Table 6.1 to 6.0 g: melting point = -7.5 - 1.5 = -9.0 °C AND boiling point = 110 + 2 = 112 °C [1]
- difference = 112 - (-9.0) = 121 °C [1]

```
melting point falls 1.5 °C per 1.0 g:  6.0 × (-1.5) = -9.0 °C   (or -7.5 - 1.5)
boiling point rises 2 °C per 1.0 g:  100 + 6.0 × 2 = 112 °C   (or 110 + 2)
difference = boiling point - melting point = 112 - (-9.0) = 121 °C
check by another route: at 5.0 g the difference is 110 - (-7.5) = 117.5 °C; each extra gram
adds 2 + 1.5 = 3.5 °C, so 117.5 + 3.5 = 121 °C
```

Accept 121 (°C given). Error carried forward: if one extrapolated value is wrong the first mark is lost, but the second is still awarded for subtracting their own two values correctly. Accept any consistent method (e.g. 117.5 + 3.5). Commonest error: 112 - 9.0 = 103 (the minus sign of the melting point is dropped), which loses the second mark. Reject a negative difference.

> **Ch 1 Experimental Chemistry › 1.4 How Can the Purity of Substances Be Determined? › p. 20 (PDF p. 31)**
> *"For example, pure water boils at 100 °C. Adding salt to water raises its boiling point."*

> **Ch 1 Experimental Chemistry › 1.4 How Can the Purity of Substances Be Determined? › p. 19 (PDF p. 30)**
> *"The greater the amount of impurities, the larger the change in melting or boiling points, and the wider the range."*

Data skill (answered from the question)

Gap: The book only says that salt raises the boiling point and lowers the freezing point; the extrapolation to 6.0 g is a pure data-handling step.
LO: 1.2(e)

Question 7 — Brass, properties of metals, elements and compounds [6]

7(a) **[2]**

Figure 7.1 shows the structure of solid brass. Give two reasons why it is a structure of brass and not of pure zinc or pure copper.

- The diagram shows atoms of different sizes (a pure metal has metal atoms of the same size), i.e. two different metals are mixed [1]
- The regular arrangement of atoms is disrupted: the arrangement of atoms is irregular (pure metals have a regular structure) [1]

```
Figure 7.1 (counted on the scan): 23 smaller circles and 4 larger circles, packed irregularly
with distorted rows. A pure metal would show every circle the same size in straight, regular
layers (the book's Figure 5.19).
```

Both reasons must use evidence from the diagram: (1) a second kind / size of atom; (2) irregular / disrupted layers. Accept "contains atoms of two elements" for (1) and "atoms are not in regular rows / layers" for (2). Reject "brass is an alloy" or "it is a mixture" on their own (restates the stem), and reasons about properties (harder, stronger) that cannot be seen in the diagram.

> **Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Figure 5.17) › p. 82 (PDF p. 93)**
> *"Figure 5.17 shows the arrangement of atoms in pure copper and in bronze. Bronze is an alloy that is about 90 % copper and 10 % tin. It contains atoms of different sizes, as each tin atom is about 1.6 times the size of a copper atom. Therefore, the regular arrangement of copper atoms is disrupted."*

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Figure 5.18) › p. 82 (PDF p. 93)

"Alloys are harder and stronger than pure metals. The irregular arrangement of atoms in alloys leads to the difference in properties between alloys and pure metals (Figure 5.18)."

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Structural Properties, Figure 5.19) › p. 82 (PDF p. 93)

"Pure metals have a regular structure, as shown in Figure 5.19. … metal atoms of the same size"

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Structure of Metals and Alloys) › p. 81 (PDF p. 92)

"Metals can exist as pure elements. They can also be mixed with other elements to produce alloys. … An alloy is a mixture of a metal with one or more other elements."

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Table 5.3; columns Alloy | Constituent Elements) - row Brass › p. 83 (PDF p. 94)

"Brass | copper and zinc"

In the textbook

LO: 3.4(c), 3.4(b)

7(b) **[2]**

Give two physical properties typical of metals.

- One property, in the book's words: metals usually have high melting and boiling points; or metals and alloys are good conductors of heat; or metals and alloys are good electrical conductors (metal atoms have valence electrons that are free to move throughout the structure); or pure metals are malleable and ductile (can be bent or flattened into a thin sheet, and pulled into a wire without breaking) [1]
- A second, different property from the same list [1]

"Conducts heat" and "conducts electricity" count as two separate properties if each is named. Probably also credited although the book does not stress them: shiny, high density, sonorous. Not credited: chemical properties (forms positive ions, reacts with acids, forms basic oxides); "conducts" with no heat or electricity named; properties of alloys only ("harder than the pure metal"). "Strong" or "hard" are risky (Group 1 metals are soft) and are not the syllabus properties.

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? › p. 81 (PDF p. 92)

"Describe the general physical properties of metals as solids having high melting and boiling points, malleable, and good conductors of heat and electricity."

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Melting Point and Boiling Point) › p. 82 (PDF p. 93)

"Metals usually have high melting and boiling points. With the exception of mercury, metals are solids at room temperature. … Both metals and alloys are good conductors of heat. … Metals and alloys usually have high melting and boiling points. They are good conductors of heat."

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Structural Properties) › p. 82 (PDF p. 93)

"Pure metals have a regular structure, as shown in Figure 5.19. If enough force is applied, the layers of atoms can slide over one another easily. Thus, when hammered, pure metals can be bent or flattened into a thin sheet (malleable). They can also be pulled into a wire without breaking (ductile). … Pure metals are malleable and ductile."

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Electrical Conductivity) › p. 83 (PDF p. 94)

"Metal atoms have valence electrons that are free to move throughout the structure. Hence, most metals and alloys can conduct electricity. … Metals and alloys are good electrical conductors."

In the textbook

LO: 3.3(b)

7(c) **[2]**

Describe the difference between an element and a compound.

- An element is a pure substance that cannot be broken down into two or more simpler substances by chemical methods (it is made of only one element, i.e. one type of atom) [1]
- A compound is a pure substance containing two or more elements that are chemically combined in a fixed ratio [1]

The contrast must be clear for each. Accept "a compound is made of two or more elements chemically bonded together" and, as an extra, "a compound can be separated into its elements by chemical methods". Reject: "a compound is a mixture of elements" or elements "physically mixed"; "an element is one atom" (it must be one TYPE of atom); "elements are atoms, compounds are molecules" (O_2 is an element made of molecules, and ionic compounds have no molecules).

Ch 5 Structure and Properties of Materials › 5.1 How Are Elements, Compounds and Mixtures Different? (Difference Between Elements and Compounds) › p. 73 (PDF p. 84)

"An element is a pure substance that cannot be broken down into two or more simpler substances by chemical methods. Elements can appear as: atoms, if they are not chemically bonded to each other (e.g. helium); or molecules, if they are covalently bonded (e.g. oxygen molecules, O_2)."

Ch 5 Structure and Properties of Materials › 5.1 How Are Elements, Compounds and Mixtures Different? (Difference Between Elements and Compounds) › p. 73 (PDF p. 84)

"A compound is a pure substance containing two or more elements that are chemically combined in a fixed ratio. The elements are chemically bonded together to form either a covalent compound or an ionic compound."

Ch 5 Structure and Properties of Materials › 5.1 How Are Elements, Compounds and Mixtures Different? (Table 5.1; columns Element | Compound | Mixture) - row What Is It Made Of? › p. 72 (PDF p. 83)

"What Is It Made Of? | only one element | two or more elements that are chemically combined | two or more elements and/or compounds that are not chemically combined"

Ch 5 Structure and Properties of Materials › 5.1 How Are Elements, Compounds and Mixtures Different? (Separation of Compounds Into Their Constituents) › p. 73 (PDF p. 84)

"Compounds can be separated into their constituent elements using chemical methods. The separation can be achieved by: exposing the compound to strong heat (thermal decomposition); or passing an electric current through the compound (electrolysis)."

In the textbook

LO: 3.4(a)

Question 8 — Flow chart: dilute acid S acting on a mixture of two elements (iron and carbon) [10]

8(a) S and Y **[2]**

Identify acid S and solid Y.

- S = sulfuric acid, H_2SO_4: sulfates are formed from sulfuric acid, and solution U gives a white precipitate with acidified aqueous barium nitrate (the test for sulfate ions) [1]
- Y = barium sulfate, $BaSO_4$: a white precipitate of barium sulfate is formed [1]

Name or correct formula accepted ("identify"); "dilute" is not needed. Reject hydrochloric or nitric acid for S (neither gives a white precipitate with barium nitrate; barium chloride is soluble). Reject "barium nitrate" for Y (that is the reagent added).

Ch 9 Qualitative Analysis › 9.2 How Do We Test for Anions? (Table 9.2; columns Anion | Test | Observations) - row sulfate ion › p. 145 (PDF p. 156)

"Add dilute nitric acid, then add aqueous barium nitrate. | A white precipitate of barium sulfate, $BaSO_4$, is formed."

Ch 9 Qualitative Analysis › 9.2 How Do We Test for Anions? › p. 145 (PDF p. 156)

"From Table 9.2, we can see that dilute nitric acid was added in the test to identify chloride and sulfate ions. This is to remove the presence of any carbonate and hydroxide ions, ensuring that there is no error in the analysis obtained."

Ch 8 Acids and Bases › 8.1 What Is an Acid? (Properties of Acids), point 4 › p. 125 (PDF p. 136)

"Different salts are formed when various acids are used in the reactions: Sulfates are formed from sulfuric acid. Nitrates are formed from nitric acid. Chlorides are formed from hydrochloric acid."

In the textbook

LO: 6(b)

8(a) T **[1]**

Identify black solid T, which is left after the acid acts on R and which burns in air to give gas W.

- T = carbon (the black solid that does not react with the acid and burns in air to form carbon dioxide, W) [1]

```
Reasoning: R is a mixture of two elements. One of them reacts with the acid to give hydrogen
(V) and a green solution (U): iron. The other does not react with the acid, is black and burns
in air to give a gas that turns limewater milky: carbon. So R is iron + carbon, as in steel (R
itself is not asked). Black copper(II) oxide is not an element and would dissolve in the acid
to give a blue solution.
```

Accept carbon, C, graphite (charcoal or soot probably also). Reject iron (it dissolves), copper, coal, carbon dioxide or carbon monoxide.

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Table 5.3; columns Alloy | Constituent Elements | Purpose | Example) - row Steel › p. 83 (PDF p. 94)

"Steel | iron and carbon | to increase hardness | warehouse shelves"

Ch 5 Structure and Properties of Materials › 5.4 What Are the Properties of Metals and Alloys? (Figure 5.16) › p. 81 (PDF p. 92)

"Stainless steel is an alloy of iron with small amounts of other elements such as chromium and carbon."

Partly in the textbook

Gap: The book never says that carbon is a black solid that does not react with dilute acid and burns in air to form carbon dioxide. Students must deduce T from the clues (black solid + burns to give CO_2 + the other element is iron). The only nearby teaching is that steel is made of iron and carbon (Table 5.3) and the limewater test for CO_2.

LO: 6(c)

8(a) U and X **[2]**

Identify solution U and solid X.

- U = iron(II) sulfate, $FeSO_4$: iron reacts with sulfuric acid to form the salt iron(II) sulfate, a green solution [1]
- X = iron(II) hydroxide, $Fe(OH)_2$: the green precipitate of iron(II) hydroxide formed when aqueous sodium hydroxide is added to Fe^{2+} [1]

```
Data Sheet: iron(II) hydroxide is green, iron(III) hydroxide is red-brown. A green precipitate
therefore shows Fe²⁺ in U, and U is the sulfate because the acid S is sulfuric acid.
```

Name or formula accepted; the oxidation state must be correct (names need the (II)). "Fe^{2+}(aq)" alone for U is borderline; give the compound. Reject iron(III) sulfate / iron(III) hydroxide (red-brown precipitate) and "green precipitate" as a name.

Ch 9 Qualitative Analysis › 9.1 How Do We Test for Cations? (Table 9.1; columns Cation | NaOH(aq) a few drops | NaOH(aq) excess) - row iron(II) ion › p. 143 (PDF p. 154)

"Green precipitate of iron(II) hydroxide, $Fe(OH)_2$ | Precipitate is insoluble in excess. The precipitate turns brown on standing."

Ch 9 Qualitative Analysis › 9.1 How Do We Test for Cations? › p. 144 (PDF p. 155)

"Precipitation is the formation of insoluble solids from two solutions. The precipitate formed in each of the reactions in Table 9.1 is the hydroxide of the metal ion."

Ch 9 Qualitative Analysis › 9.1 How Do We Test for Cations? › p. 142 (PDF p. 153)

"The reagents used to identify cations are aqueous sodium hydroxide, NaOH(aq) and aqueous ammonia, NH_3(aq). All cations (except Na^+, K^+ and NH_4^+) give precipitates with these alkalis."

Ch 12 The Reactivity Series › 12.2 How Can We Deduce the Order of Reactivity of Metals? (Figure 12.7) › p. 188 (PDF p. 199)

"For example, when iron filings are added to a solution of copper(II) sulfate (Figure 12.7), copper metal is displaced out of the solution as a pink solid (when freshly formed) which turns into a red-brown solid; and the solution turns green. … green aqueous iron(II) sulfate"

In the textbook

LO: 6(a)

8(a) V **[1]**

Identify gas V, the colourless gas that "pops" with a lighted splint.

- V = hydrogen, H_2: acids react with reactive metals to form a salt and hydrogen gas, and hydrogen extinguishes a burning splint with a "pop" sound [1]

Accept hydrogen or H_2. Reject hydrogen ions / H^+, and oxygen (relights a glowing splint).

Ch 8 Acids and Bases › 8.1 What Is an Acid? (Properties of Acids), point 4 › p. 125 (PDF p. 136)

"Acids react with reactive metals to form a salt and hydrogen gas. … We can test for hydrogen gas by placing a burning splint at the mouth of the test tube. Hydrogen gas extinguishes the burning splint with a "pop" sound (Figure 8.3)."

Ch 9 Qualitative Analysis › 9.3 How Do We Test for Gases? (Table 9.3; columns Gas | Colour and Odour | Test | Observations) - row hydrogen › p. 148 (PDF p. 159)

"colourless and odourless | Place a burning splint at the mouth of the test tube. | The burning splint is extinguished with a "pop" sound."

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? (Table 12.2, row iron) › p. 186 (PDF p. 197)

"Metals above hydrogen in the reactivity series react with hydrochloric acid to produce hydrogen gas. … iron | Reacts slowly to give hydrogen gas."

In the textbook

LO: 6(c), 5.1(d)

8(a) W **[1]**

Identify gas W, the colourless gas that forms a white precipitate in limewater.

- W = carbon dioxide, CO_2: carbon dioxide gas forms a white precipitate with limewater [1]

Accept carbon dioxide or CO_2. Reject carbon monoxide and calcium carbonate (that is the precipitate, not the gas).

Ch 8 Acids and Bases › 8.1 What Is an Acid? (Properties of Acids), point 6 › p. 126 (PDF p. 137)

"We can test for carbon dioxide gas by bubbling the gas through limewater. Carbon dioxide gas forms a white precipitate with limewater (Figure 8.6)."

Ch 9 Qualitative Analysis › 9.3 How Do We Test for Gases? (Table 9.4, row carbon dioxide) › p. 148 (PDF p. 159)

"Bubble the gas through limewater. | A white precipitate is formed. The precipitate dissolves upon further bubbling."

Ch 9 Qualitative Analysis › 9.2 How Do We Test for Anions? (Table 9.2, row carbonate ion) › p. 145 (PDF p. 156)

"The gas given off forms a white precipitate of calcium carbonate, $CaCO_3$, with limewater. Carbon dioxide gas is produced."

In the textbook

LO: 6(c)

8(b) **[3]**

Write a balanced chemical equation, including state symbols, for any one of the reactions shown in Figure 8.1.

- Model answer (acid + metal gives salt + hydrogen): $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$
- Correct formulae of the reactants and products of ONE reaction shown in the flow chart [1]
- Balanced: every atom on the left-hand side of the arrow also appears on the right-hand side (here 1 Fe, 1 S, 4 O and 2 H on each side) [1]
- State symbols correct for every species: solid (s), liquid (l), gas (g), aqueous (aq) [1]

Other acceptable equations for reactions shown in Figure 8.1 (each checked balanced in atoms and charge):

$C(s) + O_2(g) \rightarrow CO_2(g)$ [carbon burns in air, T giving W]

$FeSO_4(aq) + 2NaOH(aq) \rightarrow Fe(OH)_2(s) + Na_2SO_4(aq)$ [U + aqueous sodium hydroxide giving X]

$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$ [ionic form of the same reaction]

$FeSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + Fe(NO_3)_2(aq)$ [U + acidified barium nitrate giving Y]

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ [ionic form of the same reaction]

$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$ [W + limewater giving a white precipitate]

$Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$ [ionic form of iron + acid]

Any ONE reaction in the figure. Wrong formulae are the usual reason an equation fails: they lose the formulae mark and normally the balancing mark too. The state-symbol mark needs every symbol right ($Fe(OH)_2$ is (s), $BaSO_4$ is (s), CO_2 is (g), aqueous salts are (aq)). An ionic equation should be credited (syllabus 4.1(e) includes ionic equations), but a full equation is the safer choice. The equation must belong to the flow chart: iron with oxygen, or magnesium with acid, scores nothing. WARNING: the answer to the book's Worked Example 6C (p. 101) writes barium sulfate as $BaSO_4(aq)$; that is a printing error (barium sulfate is the insoluble white precipitate, (s)) and students who copy it lose the state-symbol mark.

Ch 6 Chemical Formulae and Equations › 6.3 What Are Chemical Equations and How Do We Balance Them? (Balancing Chemical Equations) › p. 100 (PDF p. 111)

"A balanced chemical equation shows the expected ratio of reactants to products. Every atom on the left-hand side of the arrow must also appear on the right-hand side of the arrow. However, it may be bonded to different atoms."

Ch 6 Chemical Formulae and Equations › 6.3 (Table 6.10, Steps 1, 3 and 4) › p. 100 (PDF p. 111)

"Write the chemical formulae of the reactants and products in an unbalanced chemical equation. … Add molecule(s) of the unbalanced species to balance the equation. Do not change the chemical formulae or subtract atoms. … Add the state symbols into the balanced chemical equation."

Ch 6 Chemical Formulae and Equations › 6.3 (Chemical Equations) › p. 99 (PDF p. 110)

"Common states are: solid (s) liquid (l) gas (g) aqueous (aq)"

Ch 8 Acids and Bases › 8.1 What Is an Acid? (Properties of Acids), point 4 › p. 125 (PDF p. 136)

"When magnesium ribbon is added to dilute sulfuric acid, the following reaction occurs: magnesium + dilute sulfuric acid → magnesium sulfate + hydrogen $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$"

Ch 12 The Reactivity Series › 12.1 What Is the Order of Reactivity of Metals? (Table 12.2, row iron) › p. 186 (PDF p. 197)

"iron | Reacts slowly to give hydrogen gas. | $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$"

Ch 9 Qualitative Analysis › 9.1 How Do We Test for Cations? (Formation of Precipitates Upon Addition of Reagents) › p. 144 (PDF p. 155)

"iron(III) ion + hydroxide ion → iron(III) hydroxide $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$"

In the textbook

LO: 4.1(e), 4.1(d)

Question 9 — Phosphorus, calcium phosphide and ammonia (Section B, option 1) [10]

9(a) **[1]**

The electronic configuration of phosphorus is 2.8.5. Explain why nitrogen and phosphorus are in the same group of the Periodic Table.

- Elements in the same group have the same number of valence electrons: nitrogen (2.5) and phosphorus (2.8.5) both have 5 valence electrons (the number of valence electrons is the group number minus 10, so Group 15) [1]

```
Phosphorus 2.8.5 and nitrogen 2.5 (proton number 7, from the Periodic Table) both have 5 outer electrons; Group 15 means 15 - 10 = 5 valence electrons.
```

Accept "same number of outer-shell / valence electrons (5)". Reject "same number of electron shells" (nitrogen has 2 shells, phosphorus 3), "same number of electrons", and "similar properties" (restates the question).

> **Ch 11 The Periodic Table › 11.1 How Are Elements Arranged in the Periodic Table? (Electronic Configurations Down a Group) › p. 172 (PDF p. 183)**
> *"Elements in the same group have the same number of valence electrons. Elements in Groups 1 and 2 (e.g. lithium and beryllium) have 1 and 2 valence electrons respectively. Elements in Groups 13 to 18 (except helium) have 3 to 8 valence electrons respectively. The number of valence electrons is the group number minus 10 (e.g. chlorine has 17 – 10 = 7 valence electrons)."*

> **Ch 11 The Periodic Table › 11.1 How Are Elements Arranged in the Periodic Table? (Electronic Configurations Down a Group) › p. 172 (PDF p. 183)**
> *"Elements with the same number of valence electrons have similar chemical properties. We can thus deduce that elements in the same group have similar chemical properties (See Section 11.2)."*

> **Ch 11 The Periodic Table › 11.1 How Are Elements Arranged in the Periodic Table? (Proton Number and Electronic Configuration; Table 11.1(a)) › p. 172 (PDF p. 183)**
> *"Let us look at the element with a proton number of 7 in Table 11.1(a). We can see that the element would be in Period 2, Group 15 (i.e. nitrogen)."*

In the textbook

LO: 8.1(b), 8.1(d)

9(b)(i) **[1]**

Calcium phosphide contains Ca^{2+} and P^{3-} ions. Explain how a phosphorus atom forms a phosphide ion.

- A negative ion is formed when an atom gains one or more electrons: a phosphorus atom gains three electrons, giving the noble gas electronic configuration 2.8.8 and the ion P^{3-} [1]

```
P (2.8.5) + 3e⁻ → P³⁻ (2.8.8): 15 protons and 18 electrons give a charge of +15 - 18 = -3. (The book's worked examples are Cl⁻ and O²⁻; P³⁻ follows the same rule with three electrons.)
```

Accept "gains 3 electrons from calcium atoms". Reject: loses electrons; shares electrons; gains 3 protons; "gains a charge of 3-" with no mention of electrons.

> **Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Negative Ions) › p. 56 (PDF p. 67)**
> *"A negative ion or anion is formed when an atom or combination of atoms gains one or more electrons. The ion formed is negatively charged, because it has more electrons than protons. Most non-metal atoms will gain electrons and form negative ions."*

> **Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Negative Ions) › p. 56 (PDF p. 67)**
> *"Negative ions (anions) have a net negative charge and usually have a noble gas electronic configuration."*

> **Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Learning Outcomes) › p. 54 (PDF p. 65)**
> *"Describe the formation of (a) positive ions by losing electron(s); (b) negative ions by gaining electron(s); and that these ions usually have the electronic configuration of a noble gas."*

Ch 4 Chemical Bonding › 4.2 What Is Ionic Bonding? (Common Ions) › p. 57 (PDF p. 68)

"Positive ions seldom have a net charge higher than +3 because it gets increasingly harder for atoms to lose more electrons. Negative ions seldom have a net charge lower than –3 because it is difficult for atoms to gain and keep a large number of electrons."

In the textbook

LO: 3.1(a)

9(b)(ii) **[2]**

Predict two physical properties of calcium phosphide (an ionic compound with a giant lattice).

- Ionic compounds have high melting and boiling points (strong electrostatic forces of attraction between oppositely charged ions need a large amount of energy to overcome); they are solids at room temperature [1]
- A second, different property: ionic compounds conduct electricity when molten or in aqueous solution, but not in the solid state; or are hard but brittle [1]

Accept also "insoluble in organic solvents". "High melting point" and "high boiling point" are best treated as ONE property. Reject: "conducts electricity" with no condition (solid ionic compounds do not), low melting point, gas at room temperature, malleable or ductile. CAUTION: "soluble in water" (the book's generic ionic property, p. 77) would probably be credited by a generic scheme, but it is chemically wrong here: calcium phosphide, Ca_3P_2, reacts with water to give phosphine. The two chemically safest answers are high melting point and conduction when molten.

Ch 5 Structure and Properties of Materials › 5.2 What Are the Properties of Ionic Substances? (Melting Point and Boiling Point) › p. 76 (PDF p. 87)

"Ions of opposite charges have strong electrostatic forces of attraction between them. A large amount of energy is needed to overcome these forces. Hence, ionic compounds tend to have very high melting and boiling points. They are solids at room temperature. … Ionic compounds have high melting and boiling points."

Ch 5 Structure and Properties of Materials › 5.2 What Are the Properties of Ionic Substances? (Electrical Conductivity) › p. 77 (PDF p. 88)

"When an ionic compound is in the solid state, the ions in the lattice are only able to vibrate about their fixed positions. Hence, the ions are not mobile. However, when the compound is molten (in the liquid state) or in the aqueous state, the ions are mobile and can conduct electricity. … Ionic compounds conduct electricity when molten or in aqueous solution, but not in the solid state."

Ch 5 Structure and Properties of Materials › 5.2 What Are the Properties of Ionic Substances? (Hardness) › p. 76 (PDF p. 87)

"The strong forces of attraction between oppositely charged ions make ionic compounds resistant to deforming. Ionic compounds tend to be hard, as these forces cause the ions to resist motion. … Hence, ionic compounds are hard but brittle. … Ionic compounds are physically hard but brittle."

Ch 5 Structure and Properties of Materials › 5.2 What Are the Properties of Ionic Substances? (Solubility) › p. 77 (PDF p. 88)

"Most ionic compounds are soluble in water and insoluble in organic solvents."

Ch 5 Structure and Properties of Materials › 5.2 What Are the Properties of Ionic Substances? (Structural Properties) › p. 76 (PDF p. 87)

"Ionic compounds have giant ionic crystal lattice structures (Figure 5.6). The lattice consists of alternating positive and negative ions. These ions are held together by electrostatic forces of attraction."

In the textbook

LO: 3.1(d)

9(c)(i) **[2]**

Draw a 'dot-and-cross' diagram for a molecule of ammonia, NH_3. Include ALL the electrons.

- Three bonding pairs: each hydrogen atom shares 1 electron with the nitrogen atom (nitrogen has 5 electrons in its outermost shell and shares 3), each bonding pair shown as one dot and one cross; each hydrogen has a duplet [1]
- Nitrogen has an octet: its outermost shell holds 8 electrons (the 3 bonding pairs and one non-bonded pair), and (because the question says "all the electrons") its first shell of 2 electrons is shown as well [1]

```
Electron inventory (10 in all): N atom = 7 (2 in the first shell + 5 in the outermost shell);
each H atom = 1.
Draw: an N circle with its first shell holding 2 electrons, and an outermost shell holding 8 =
one non-bonded pair (2 N electrons) + three bonding pairs; each bonding pair = 1 N electron
(dot) + 1 H electron (cross) in the overlap of the N circle and that H circle. Each H circle
holds only its bonding pair.
Key: dot = electron from N, cross = electron from H.
Check: first shell 2 + non-bonded pair 2 + 3 bonding pairs × 2 = 10 electrons.
```

Accept any consistent dots and crosses, shells drawn as overlapping circles (as the book does) or as concentric circles; the two inner electrons may be a pair of dots on a small inner circle. The non-bonded pair is usually called a lone pair. Reject: nitrogen with no non-bonded pair (only 6 outer electrons); four bonding pairs (NH_4); an H atom with two of its own electrons; a bond drawn as a line only. Unofficial reading of the scheme: one mark for the three N-H bonding pairs, one for the lone pair / complete outer shell. Drawing the inner shell as well is the safe route: it cannot lose a mark if the scheme wants outer electrons only, and it gains one if the scheme wants all electrons.

> **Ch 4 Chemical Bonding › 4.3 What Is Covalent Bonding? (Covalent Molecules) › p. 65 (PDF p. 76)**
> *"Generally, we can determine how atoms will bond covalently by comparing the number of electrons each atom shares. Suppose one atom shares x electrons to achieve a noble gas electronic configuration. This atom will form a covalent compound with one or more other atoms that can "match" this sharing. For example, ammonia is made up of hydrogen and nitrogen atoms. Let us compare the number of electrons each hydrogen atom and nitrogen atom shares."*

> **Ch 4 Chemical Bonding › 4.3 What Is Covalent Bonding? (Covalent Molecules, Figure 4.20) › p. 65 (PDF p. 76)**
> *"Hydrogen atom: It has 1 electron in its outermost shell. It shares 1 electron to have a duplet electronic configuration. Nitrogen atom: It has 5 electrons in its outermost shell. It shares 3 electrons to have an octet electronic configuration."*

> **Ch 4 Chemical Bonding › 4.3 What Is Covalent Bonding? (Covalent Molecules, Figure 4.20) › p. 65 (PDF p. 76)**
> *"Therefore, three hydrogen atoms will share electrons with one nitrogen atom (Figure 4.20)."*

> **Ch 4 Chemical Bonding › 4.3 What Is Covalent Bonding? (The Covalent Bond, Figure 4.15) › p. 62 (PDF p. 73)**
> *"The overlap shows that this pair of electrons is shared between the two atoms. These electrons are called a bonding pair. A single covalent bond has been formed."*

> **Ch 3 Atomic Structure › 3.3 How Are Sub-atomic Particles Distributed in an Atom? (Electron Shells and Electrons) › p. 44 (PDF p. 55)**
> *"Each electron shell can only hold a certain number of electrons. The first electron shell can hold a maximum of 2 electrons. The second and third shells can usually have a maximum of 8 electrons each."*

> **Ch 11 The Periodic Table › 11.1 How Are Elements Arranged in the Periodic Table? (Proton Number and Electronic Configuration; Table 11.1(a)) › p. 172 (PDF p. 183)**
> *"Let us look at the element with a proton number of 7 in Table 11.1(a). We can see that the element would be in Period 2, Group 15 (i.e. nitrogen)."*

> **Ch 4 Chemical Bonding › 4.3 What Is Covalent Bonding? (Worked Example 4B) › p. 67 (PDF p. 78)**
> *"Methane is the simplest covalent compound formed between carbon and hydrogen. Draw a dot-and-cross diagram for methane. (Only the outermost electrons need to be shown.)"*

Partly in the textbook

Gap: The book teaches NH_3 (Figure 4.20, p. 65) and the first-shell capacity of 2 (p. 44), but its dot-and-cross diagrams show only the outermost electrons, and its worked examples say "Only the outermost electrons need to be shown" (p. 67). It never draws the inner shell in a dot-and-cross diagram and never uses the term "lone pair" (Figure 4.20 shows the non-bonded pair on N but does not name it). The instruction "Include all the electrons" is therefore new to students.

LO: 3.2(b), 3.2(c)

9(c)(ii) **[1]**

The reaction $N_2 + 3H_2 \rightarrow 2NH_3$ is exothermic. State the meaning of the term exothermic.

- Exothermic changes release energy to the surroundings (and increase the temperature of the surroundings) [1]

Accept "gives out heat (energy) to the surroundings" and "temperature of the surroundings increases". "Heat is given out" on its own is the commonest borderline answer: add "to the surroundings". Reject "takes in energy", "enthalpy change is positive", "energy is produced or created".

> **Ch 13 Chemical Energetics › 13.1 What Are Endothermic and Exothermic Changes? (Exothermic Changes) › p. 203 (PDF p. 214)**
>
> *"Exothermic changes: release energy to the surroundings; and increases the temperature of the surroundings."*

> **Ch 13 Chemical Energetics › 13.1 (Neutralisation) › p. 204 (PDF p. 215)**
>
> *"The neutralisation reaction between sodium hydroxide and hydrochloric acid is exothermic (Figure 13.10). When sodium hydroxide and hydrochloric acid react, energy is released to the surroundings and the temperature of the surroundings increases."*

> **Quick Revision Guide › Chapter 13: Chemical Energetics › p. 316 (PDF p. 327)**
>
> *"Exothermic changes involve the release of energy into the surroundings. Hence, the temperature of the surroundings increases."*

> **Ch 13 Chemical Energetics › 13.1 (Learning Outcomes) › p. 200 (PDF p. 211)**
>
> *"Describe exothermic changes as processes and chemical reactions that release energy to the surroundings."*

In the textbook

LO: 9(a)

9(c)(iii) **[3]**

Describe and explain how the rate of the reaction between nitrogen and hydrogen changes as the temperature increases. Use ideas about particles.

- Effect on rate: an increase in the temperature of reactants results in an increase in the rate of reaction (higher temperatures result in faster reaction rates) [1]
- Explanation: at higher temperatures, reactant particles have more kinetic energy; this means that they move and vibrate faster [1]
- increasing the frequency of effective collisions (the particles collide more frequently) [1]

Accept "collide more often", "more frequent (successful / effective) collisions". Alternative or extra route, in the book's terms: more reactant particles possess energy greater than or equal to the activation energy, which increases the frequency of effective collisions (effective collisions and activation energy go beyond the Science syllabus and are never required). Reject: "more collisions" with no "more frequent / per second"; "the activation energy decreases" (only a catalyst does that); "more particles" or "particles expand" (the number of particles is unchanged); "particles gain heat" with no kinetic energy or faster movement; a rate change with no direction. The reaction being exothermic does not change the answer: the rate still increases.

> **Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (Effect of Temperature) › p. 220 (PDF p. 231)**
>
> *"At higher temperatures, the rate of reaction is affected in two ways: … At higher temperatures, reactant particles have more kinetic energy. This means that they move and vibrate faster, increasing the frequency of effective collisions. … At higher temperatures, more reactant particles possess energy that is greater than or equal to activation energy. This increases the frequency of effective collisions. Hence, higher temperatures result in faster reaction rates (Figure 14.19)."*

> **Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (Effect of Temperature) › p. 220 (PDF p. 231)**
>
> *"An increase in the temperature of reactants results in an increase in the rate of reaction."*

Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (Effective Collisions) › p. 214 (PDF p. 225)

"When a chemical reaction occurs between particles, the collision is said to be an effective collision. The activation energy is the least amount of energy that colliding reactant particles must possess to react with each other. For an effective collision to occur, the reactant particles must collide with energy that is greater than or equal to the activation energy (Figure 14.10)."

Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (summary, Temperature) › p. 223 (PDF p. 234)

"As the temperature increases, reactant particles have more kinetic energy. More reactant particles possess energy that is greater than or equal to activation energy. This increases the frequency of effective collisions."

In the textbook

LO: 10(a)

Question 10 — Hydrogen peroxide: redox test, pH and a mole calculation (Section B, option 2) [10]

10(a) **[1]**

Hydrogen peroxide is a reducing agent. Describe the colour change when aqueous hydrogen peroxide is added to acidified potassium manganate(VII).

- The (dark) purple acidified potassium manganate(VII) turns colourless (a positive test for a reducing agent) [1]

Accept "decolourised", "purple to colourless". Reject any other final colour (pink, brown, green, yellow); "clear" is risky, write "colourless". The colour change must be from purple, not just "turns colourless".

Ch 10 Oxidation and Reduction › 10.2 How Can We Test for Redox Reactions? (Acidified Potassium Manganate(VII) Test For Reducing Agents) › p. 165 (PDF p. 176)

"In this test, dark purple aqueous potassium manganate(VII) that has been acidified with a little sulfuric acid is used. A few drops of the purple solution are added to a solution of the unknown substance. … In a positive test, manganate(VII) ions are reduced to manganese(II) ions by the reducing agent (Figure 10.16)."

Ch 10 Oxidation and Reduction › 10.2 How Can We Test for Redox Reactions? (Figure 10.16) › p. 165 (PDF p. 176)

"The test for reducing agents can also be applied to gases. Use a strip of filter paper with a spot of acidified potassium manganate(VII) solution. If the spot decolourises when exposed to the gas, the gas is a reducing agent. … Positive test (colourless solution)"

Ch 9 Qualitative Analysis › 9.3 How Do We Test for Gases? (Table 9.4; columns Gas | Colour and Odour | Test | Observations) - row sulfur dioxide › p. 148 (PDF p. 159)

"Place a piece of filter paper soaked with acidified potassium manganate(VII) at the mouth of the test tube. | The purple acidified potassium manganate(VII) turns colourless."

Ch 10 Oxidation and Reduction › 10.2 How Can We Test for Redox Reactions? (Oxidising and Reducing Agents) › p. 163 (PDF p. 174)

"A reducing agent is able to: reduce another substance; decrease the oxidation state of an element in another substance; and lose electrons to the other substance. Thus, a reducing agent causes reduction while being oxidised itself."

In the textbook

LO: 7.1(d)

10(b)(i) **[1]**

State the formula of the ion in aqueous hydrogen peroxide that makes it an acidic solution.

- H^+: an acid is a substance that produces hydrogen ions, H^+, in aqueous solutions [1]

Accept H_3O^+ (hydronium / oxonium). The charge must be shown: "H" alone is wrong. Reject OH^- and H_2O_2.

> **Ch 8 Acids and Bases › 8.1 What Is an Acid? › p. 124 (PDF p. 135)**
> *"From Table 8.1, we can see that all the acids produce hydrogen ions, H^+, in aqueous solutions. When acids are dissolved in water, the acid molecules ionise, or dissociate, to produce H^+ ions. This H^+ ion is mainly responsible for the behaviour of acids. … An acid is a substance that produces hydrogen ions, H^+, in aqueous solutions."*

> **Ch 8 Acids and Bases › 8.3 How Do We Compare Relative Acidity and Alkalinity? (pH and Concentration of Hydrogen Ions) › p. 131 (PDF p. 142)**
> *"pH < 7: The concentration of H^+ ions is higher than that of OH^- ions, so the solution is acidic."*

> **Ch 8 Acids and Bases › 8.3 How Do We Compare Relative Acidity and Alkalinity? (Helpful Note) › p. 131 (PDF p. 142)**
> *"A successful test for acids (e.g. blue litmus paper) indicates the presence of H^+ ions."*

In the textbook

LO: 5.1(a)

10(b)(ii) **[2]**

A pH meter is used to measure the pH of aqueous hydrogen peroxide. Describe one other method to determine the pH.

- Use Universal Indicator (a mixture of coloured dyes, in the form of a solution or paper): add a few drops of Universal Indicator solution to a sample of the hydrogen peroxide (or dip Universal Indicator paper into it) [1]
- It gives different colours depending on how acidic or alkaline the solution is: compare the colour obtained with the pH colour chart (Universal Indicator scale) to read off the pH [1]

Accept the indicator-paper version: put a drop of the solution on the paper and match its colour to the chart. Reject: litmus or another single indicator (it only shows acidic or alkaline, not a pH value); a pH sensor / probe / data logger (the question says "other" than a pH meter; the book lists the sensor separately, so a lenient examiner might credit it, but do not rely on it); a titration; adding a carbonate. A method with no colour comparison scores one mark at most.

> **Ch 8 Acids and Bases › 8.3 How Do We Compare Relative Acidity and Alkalinity? › p. 129 (PDF p. 140)**
> *"Universal Indicator is a mixture of coloured dyes. It may come in the form of a solution or paper. It gives different colours depending on how acidic or alkaline a solution is (Figure 8.11)."*

> **Ch 8 Acids and Bases › 8.3 How Do We Compare Relative Acidity and Alkalinity? (The pH Scale) › p. 130 (PDF p. 141)**
> *"A set of numbers from 0 to 14 is used to indicate whether a solution is acidic, neutral or alkaline. This is called the pH scale. The pH of a solution may be determined by using: an indicator; a pH meter; or a pH sensor connected to a data logger (Figure 8.12)."*

> **Ch 8 Acids and Bases › 8.3 How Do We Compare Relative Acidity and Alkalinity? (Helpful Note) › p. 130 (PDF p. 141)**
> *"The colours of the different pH values in Universal Indicator mimic the colours of the rainbow."*

> **Ch 8 Acids and Bases › 8.3 How Do We Compare Relative Acidity and Alkalinity? (pH and Concentration of Hydrogen Ions) › p. 130 (PDF p. 141)**
> *"The pH of a solution is related to the concentration of hydrogen ions or hydroxide ions present in a solution. An acid with a lower pH value is more acidic. It has a higher concentration of hydrogen ions."*

Partly in the textbook

Gap: The book says Universal Indicator gives different colours at different acidities (p. 129) and pictures the colour chart (Figure 8.11) and a coloured pH scale (Figure 8.13), but never states in words the step "match the colour to the pH chart to read the pH".

LO: 5.1(b), 5.1(a)

10(c)(i) **[1]**

A sample of 30.0 cm^3 of 0.125 mol/dm^3 H_2O_2 is completely decomposed. Calculate the amount, in moles, of H_2O_2 in the sample.

- Number of moles of H_2O_2 = concentration of H_2O_2 × volume of H_2O_2 = 0.125 × (30.0 ÷ 1000) = 0.00375 mol [1]

```
n = concentration × volume (in dm³) = 0.125 mol/dm³ × 0.0300 dm³ = 0.00375 mol
```

Accept 0.00375 or 3.75 × 10^{-3} (three significant figures, as the data). Reject 3.75 or 0.0375 (volume not converted to dm^3).

> **Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (worked example) › p. 115 (PDF p. 126)**
> *"Number of moles of KOH reacted = concentration of KOH × volume of KOH = 0.100 × (30.0 ÷ 1000) = 0.003 mol"*

> **Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Concentration in mol/dm^3) › p. 111 (PDF p. 122)**
> *"Concentration in mol/dm^3 = number of moles of solute [over] volume of solvent in dm^3"*

> **Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Link: Recall from Chapter 1) › p. 110 (PDF p. 121)**
> *"1 dm^3 is equivalent to 1 litre. There are 1000 cm^3 in 1 dm^3."*

In the textbook

LO: 4.2(g)

10(c)(ii) **[1]**

Calculate the amount, in moles, of oxygen produced.

- From the equation, 2 mol of H_2O_2 form 1 mol of O_2 (mole ratio 2 : 1): number of moles of O_2 = 0.00375 ÷ 2 = 0.001875 mol (0.00188 mol) [1]

```
2H₂O₂ → O₂ + 2H₂O: 2 mol H₂O₂ gives 1 mol O₂, so n(O₂) = 0.00375 ÷ 2 = 0.001875 mol
```

Accept 0.001875, 0.00188, 0.00187, 1.875 × 10^{-3}. Error carried forward: half of the answer to (i). Reject 0.0075 (ratio inverted) and 0.00375 (ratio ignored).

> **Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (worked example) › p. 115 (PDF p. 126)**
> *"From the equation, 2 mol of KOH will react with 1 mol of H_2SO_4. Number of moles of H_2SO_4 reacted = 0.003 ÷ 2 = 0.0015 mol"*

> **Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Calculations from Chemical Equations) › p. 114 (PDF p. 125)**
> *"This ratio between the quantities of the different substances in the reaction is the mole ratio. Combining this ratio with the mole concept allows us to find the masses, volumes and concentrations of substances in a chemical reaction."*

> **Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Table 7.5, Step 2) › p. 114 (PDF p. 125)**
> *"Find the number of moles of the substance we are solving for using the mole ratio from the balanced equation and the result from step 1. In this case, we need to find the number of moles of CO_2. From the equation, 1 mol of $MgCO_3$ will form 1 mol of CO_2."*

In the textbook

LO: 4.2(f)

10(c)(iii) **[1]**

Calculate the volume of oxygen, in dm^3, produced at room temperature and pressure (1 mol of any gas = 24 dm^3).

- Volume of O_2 in dm^3 = number of moles of O_2 × 24 = 0.001875 × 24 = 0.045 dm^3 [1]

```
V = n × 24 dm³/mol = 0.001875 × 24 = 0.0450 dm³ (= 45 cm³)
```

Accept 0.045, 0.0450 and 0.0451 (from the rounded 0.00188); 4.5 × 10^{-2}. Error carried forward: 24 × their answer to (ii). Reject 45 written against "dm^3" (cm^3 not converted).

Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Moles and Gaseous Volumes) › p. 110 (PDF p. 121)

"One mole of any gas occupies a volume of 24 dm^3 at room temperature and pressure (r.t.p.)."

Ch 7 Mole Concept and Stoichiometry › 7.2 What Is the Mole? (Worked Example 7B) › p. 110 (PDF p. 121)

"Calculate the volume occupied by 2 moles of oxygen gas (O_2), measured at room temperature and pressure (r.t.p.). Answer Volume of O_2 in dm^3 = number of moles of O_2× 24 dm^3 = 2 × 24 = 48 dm^3"

Ch 7 Mole Concept and Stoichiometry › 7.3 What Is Stoichiometry? (Table 7.5, Step 3) › p. 114 (PDF p. 125)

"Lastly, calculate the answer using the number of moles obtained in step 2. In this case, we need to find the volume of CO_2. Volume of CO_2 in dm^3 = number of moles of CO_2 × 24 = 2 × 24 = 48 dm^3"

In the textbook

LO: 4.2(f)

10(c)(iv) **[3]**

Describe and explain how the rate of decomposition of aqueous hydrogen peroxide changes as the concentration of hydrogen peroxide increases. Use ideas about particles.

- Effect on rate: an increase in concentration of reactants increases the rate of reaction [1]
- Explanation: as the concentration of reactants increases, the number of reacting particles per unit volume increases [1]
- They collide more frequently and hence, the frequency of effective collisions also increases [1]

Accept "more particles in the same volume", "more frequent (successful / effective) collisions". Reject: "more particles" with no "per unit volume / in the same volume"; "more collisions" with no "more frequent / per second"; "particles move faster / have more energy" (that is the temperature answer); a rate change with no direction; any mention of a catalyst, which is not in the question.

Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (Effect of Concentration) › p. 215 (PDF p. 226)

"The concentration of a solution shows the number of particles in a given volume. As the concentration of reactants increases, the number of reacting particles per unit volume increases. They collide more frequently and hence, the frequency of effective collisions also increases (Figure 14.11)."

Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (Effect of Concentration) › p. 215 (PDF p. 226)

"An increase in concentration of reactants increases the rate of reaction."

Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (summary, Concentration) › p. 223 (PDF p. 234)

"As the concentration of reactants increases, the number of reacting particles per unit volume increases. The reacting particles collide more frequently, increasing the frequency of effective collisions."

Ch 14 Rate of Reactions › 14.2 What Determines the Rate of Reactions? (Effective Collisions) › p. 214 (PDF p. 225)

"When a chemical reaction occurs between particles, the collision is said to be an effective collision. … When more effective collisions occur between reactant particles, the rate of reaction increases."

In the textbook

LO: 10(a)

Section B choice

Q10 is the kinder option. Every part of it is taught directly in Discover Chemistry (only the colour-chart wording in 10(b)(ii) is thin), seven of its ten marks are recall of one taught fact or a standard collision-theory explanation, and the three calculation marks are a routine mole chain (n = c × V, the 2 : 1 ratio, × 24 dm^3) with every number supplied and a near-identical worked example on p. 115 of the book; the stem even tells candidates that H_2O_2 is a

reducing agent and acidic. Q9 is not hard but carries more risk: calcium phosphide and P^{3-} are unfamiliar, "predict two physical properties" of an unfamiliar compound invites chemically unsafe answers (for example "soluble in water", which is wrong for Ca_3P_2), and the 2-mark dot-and-cross diagram says "include all the electrons", which the book never practises (it draws outer electrons only), so the inner shell is easily forgotten. Q9 suits a candidate who is strong on bonding and the Periodic Table but weaker at mole calculations; for everyone else Q10 is the safer choice. Both options end with the same 3-mark collision-theory explanation (temperature in Q9, concentration in Q10).

Notes

- Transcription vs scan: Q1-Q5 were re-read from the scan (pp. 2-5 and 7-9 rendered at 1.5x; the Periodic Table on p. 20 at 2x). No disagreement was found: the stems, Figure 2.1 (order sodium, calcium, magnesium, metal E, iron, lead, copper, silver, arrow pointing up to 'most reactive'), Table 4.1 (rows electrons, neutrons, protons, blank cells), the nuclide symbol 79/35 Br, Figure 5.1 (ethene; but-2-ene drawn with the two CH3 groups on opposite corners) and every printed mark are as transcribed. The transcription's 'expected answer (unofficial)' lines for Q1-Q5 agree with this sheet; the only addition here is that 1(e) also credits carbon dioxide.
- Marks check (standalone script): Q1 = 5, Q2 = 1+1+1+2 = 5, Q3 = 1+1+1+1+3 = 7, Q4 = 1+3+2+2 = 8, Q5 = 1+1+2+2+1+1 = 8; total 33, equal to the printed [Total] of each question. Every balanced equation, atom count and calculation on the sheet was asserted in a small Python check (Mr(CuCO3) = 124, n = 0.0100 mol, m(CuO) = 0.80 g also cross-checked by conservation of mass).
- Quotation method (revised): quotations are now full passages (the complete defining sentence(s), box, worked-example lines or table rows the answer is built from), 103 textbook entries from 45 printed pages in all. Each one was machine-matched, segment by segment, against the text layer of the full-book PDF on the cited page (white space, bullet glyphs, arrows and sub/superscript digits ignored), and the printed page number was confirmed on that PDF page. Every cited page (all 45) was also rendered and read as an image to check that passages joined in one entry are really side by side on the page: where the PDF text layer and the printed layout differ (pp. 40, 38, 109, 177, 184, 249, 275) the passage is given in printed order and its pieces are verified separately. Conventions: bullets are written as a bullet sign; a stacked fraction is written with a slash (1/12 on pp. 106-107, and the mole formula on p. 109); reaction arrows are written as an arrow (a reversible arrow for the iron row of Table 12.1) and the conditions printed above or below an arrow (yeast, 150 C, Ni, high temperature and pressure) are not reproduced; square brackets inside a quotation mark a diagram or placeholder box printed at that point; subscripts and charges are Unicode; table rows are given in reading order (Tables 6.2, 11.5, 12.1, 12.4, 12.5, 17.3). Where a quotation needs a break, it is split into separate entries rather than joined, so nothing inside quotation marks is paraphrased or abridged.
- Syllabus: only the 2027 K326-327-328 text is on file (the paper is set on the 5088 syllabus). All cited outcomes are in it, as comb outcomes 2.2(a),(d),(e),(f), 4.1(b),(d), 4.2(b),(d), 5(c), 8.1(f), 8.2(b), 8.3(a),(c),(d), 11.2(b),(e),(i), 11.3(d),(e), 11.4(d),(f), 12(e). The 'lo' fields use the project's canonical (pure-numbered) codes from taxonomy.json and match the tags already in data/comb/2026 P3.json; canonical to comb: 12(g) = 12(e); 8.4(a) = 8.3(a); 8.4(e) = 8.3(c); 8.4(f) = 8.3(d); 5.1(d) = 5(c); 4.2(f) = 4.2(d); 11.2(f) = 11.2(e); 11.2(j) = 11.2(i); 11.4(h) = 11.4(f); 8.1(g) = 8.1(f); all other codes are identical. Nothing in Q1-Q5 goes beyond the syllabus: but-2-ene is not one of the alkenes the syllabus asks candidates to name or draw (ethene, propene), but the question supplies its name and structure, so 5(d) only tests 11.4(d). The syllabus asks candidates to define relative molecular mass and to calculate relative formula mass 'as the sum of relative atomic masses', so the answer to 3(a)(i) gives the book's definition together with the 'adding together' wording that matches the syllabus.
- No part of Q1-Q5 is classed partly, not_in_textbook or data_skill: every expected answer has a verbatim passage in the book, so gaps_summary is empty for this half. The two nearest cases are 4(d), where the 'state' trend (gas, liquid, solid) is only in the appearance column of Table 11.5 and not in a sentence (the colour, melting/boiling-point and reactivity trends are all stated in words), and 1(a), where the book gives the subscript rule but never asks for a total atom count.

- Rulings on the most ambiguous parts. (1) 1(e): ethanol and carbon dioxide are both formed by fermenting glucose and both are in the list; the key is most likely ethanol (the syllabus outcome is 'formation of ethanol by fermentation of glucose'), but a careful scheme must accept CO2 too, and the book states both products on p. 263. (2) 4(d): four trends can be given, but state and melting/boiling point are one physical idea, so a cautious marker may credit only one of them; two trends from different families (colour, state or melting/boiling point, reactivity) are safest. (3) 2(c): E is not named; the mark is for 'metal oxide + hydrogen', not for naming zinc. (4) 2(d): 'air' for oxygen and 'moisture' for water are normally accepted; salt, acid gases and heat are not essential conditions. (5) 5(e)(ii): the syllabus names cracking; depolymerisation to monomers is a pure-syllabus idea but is chemically correct and should not be penalised.
- E in Figure 2.1: the book's reactivity series (Fig. 12.2, p. 184) contains carbon and zinc between magnesium and iron and has no aluminium; Figure 2.1 omits carbon. The only metal on the syllabus list (8.3(a)) that fits between magnesium and iron is zinc, so E = zinc, supported by the book's own steam row (Table 12.1, p. 185).
- 5(d) is close to a repeat of a textbook exercise: Let's Practise 18.2 Q2 (Figure 18.14, p. 279) shows a chain of -CH(CH3)-CH(CH3)- units, which is poly(but-2-ene), and asks for the monomer; the paper asks the reverse. The book's 'four boxes' rule (p. 277: hydrogen, alkyl group, halogen etc. on each carbon) covers a CH3 on both carbons.
- Ar values for the calculations (Cu 64, C 12, O 16, Br 80, Zn 65) are the ones printed in the paper's own Periodic Table on p. 20. The paper writes the carboxyl group as CO2H (as the syllabus does); the book writes COOH.
- Marks checked against the scan (pp. 10-17): Q6 = 3 + 1 + 2 = 6; Q7 = 2 + 2 + 2 = 6; Q8 = 7 + 3 = 10; Q9 = 1 + 1 + 2 + 2 + 1 + 3 = 10; Q10 = 1 + 1 + 2 + 1 + 1 + 1 + 3 = 10. Every printed [Total] line agrees with the transcription (6, 6, 10, 10, 10).
- Transcription versus scan for Q6-Q10: no disagreement found. Re-read from the scan: Table 6.1 values (-1.5 to -7.5 and 102 to 110), Figure 7.1 (counted 23 smaller and 4 larger circles), Figure 8.1 (all labels and arrows, including that the Y arrow starts from solution U), the Data Sheet (p. 19: iron(II) hydroxide green, iron(III) hydroxide red-brown) and the Q10 data (30.0 cm^3, 0.125 mol/dm^3).
- Layout: Q8(a) is split into five rows (S and Y, T, U and X, V, W; 2 + 1 + 2 + 1 + 1 = 7 marks) so that each can carry its own textbook status; the paper prints it as one 7-mark part. The seven identifications score one mark each.
- Quotation conventions (Q6-Q10): every quotation is verbatim from the printed page named, copied from the page text of the full-book PDF and checked against it by script; Table 2.1, Figure 4.20, Table 5.3, Table 12.2, Worked Example 6C and the p. 125 equation were also read on the rendered page image. In a quotation, " | " separates the cells of a table row in reading order (the column heads are named in the section field), " … " joins separate passages on the same page or marks list numbers that were left out, and [over] marks a fraction written as numerator above denominator. The mark points in the answers re-use the book's own wording but are not themselves in quotation marks.
- Textbook error to warn students about: Worked Example 6C (p. 101) writes the product barium sulfate as $BaSO_4(aq)$. Barium sulfate is the insoluble white precipitate, so its state symbol is (s); this matters for Q8(b) if a candidate chooses the barium nitrate reaction.
- Q6(b): the numbers in Table 6.1 are schematic, not real. For 1.0 g of NaCl in 100 g of water the real freezing-point depression is about 0.6 °C and the real boiling-point rise about 0.2 °C, not 1.5 °C and 2 °C. The item is pure data handling (extrapolate a linear table), so answer from the table, not from chemistry.
- Q7(a): copper and zinc atoms are in reality almost the same size; Figure 7.1 is schematic and shows two sizes, so the answer comes from the diagram (two sizes + irregular layers).
- Q8: R is iron + carbon (steel-like) but R is not asked. Hydrochloric acid is excluded for S because barium chloride is soluble, so only a sulfate gives the white precipitate Y. The Data Sheet alone identifies Fe^{2+} (green iron(II) hydroxide), so U can be deduced without knowing the colour of iron(II) salts. The book adds that the green precipitate turns brown on standing (p. 143).
- Q9(b)(ii): calcium phosphide is Ca_3P_2 (3 Ca^{2+} : 2 P^{3-}) and it reacts with water to give phosphine, so "soluble in water" (the generic ionic-compound property in the book, p. 77) is chemically wrong here even though a generic scheme will probably credit it; high melting point and conduction when molten are the chemically safe answers.

- Q9(c)(i): precedent for "all the electrons" meaning the inner shell too: the topical-book key to 2025 P3 Q11(a) ("Show all the electrons in the molecule", Cl_2) draws all three shells of each atom, and the three answer sources transcribed for 2015 P3 Q4(a)(ii) (ammonia) all draw N with its inner shell (2) and an outer shell of 8 (three bonding pairs + one lone pair). Sources: References\S34 G3 Science Chemistry\Topical (16-25) g3 - Paper 3 2025 transcription.md and Old papers (2004-2015) - Paper 3 transcription.md.
- Collision wording (Q9(c)(iii), Q10(c)(iv)): the structured-question guide for this paper (References\S34 G3 Science Chemistry\How Cambridge sets G3 Science(Chem) structured questions) records that "effective collisions" and "activation energy" never appear in the examiner reports or the Science syllabus; the creditworthy core is direction + kinetic energy (temperature) or particles per unit volume (concentration) + more frequent collisions. Discover Chemistry teaches the activation-energy route as well (pp. 214, 220, 223), so students may use it but need not.
- Q10(c): carry the unrounded 0.001875 mol into (iii) (giving 0.0450 dm^3); rounding to 0.00188 first gives 0.0451 dm^3, which should be accepted. Q10(a): in practice oxygen also bubbles off; the pale pink Mn^{2+} is not visible in dilute solution, so the answer is "purple to colourless". Q10(b)(i): hydrogen peroxide is in fact only a very weak acid, but the stem asks for the ion that makes it acidic, which is H^+.
- Syllabus check (K326-K328 Science Chemistry): every part sits inside the syllabus. Points to note: activation energy in Q9(c)(iii) is outside the Science syllabus (not required); NH_3 is not in the syllabus list of dot-and-cross examples but is covered by 3.2(c) (deduce the arrangement of electrons in other covalent molecules); hydrogen peroxide is not named in the syllabus, but the stem supplies its role as a reducing agent, so only the $KMnO_4$ test (7(c)) is needed; Q8 T = carbon relies on general knowledge not stated in the syllabus.
- Book and page numbers: Discover Chemistry for GCE ‘O’ Level Science and Normal (A) Level Science, Textbook, 3rd Edition (Marshall Cavendish Education), full-book PDF C:\Users\cher_\Desktop\Discover Chemistry GCE ‘O’ & NA Level (3rd ed).pdf. "p." is the printed page (running through the whole book); "PDF p." is the index in that full-book PDF (= printed page + 11). The separate chapter PDFs carry the same printed numbers.