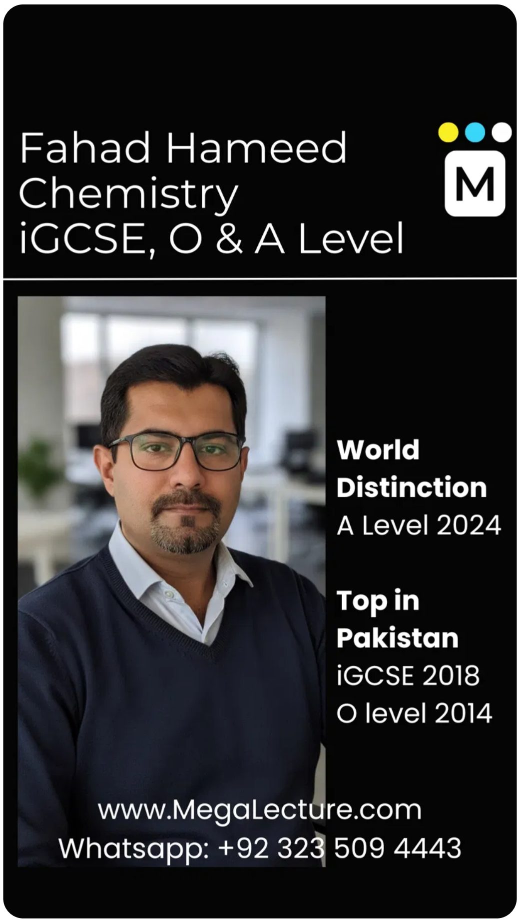
MEGA LECTURE · A LEVEL CHEMISTRY

Chapter 4: Acids and Redox

OCR A Level Chemistry A (H432) — Structured Practice Worksheet



SPECIFICATION	TOTAL MARKS	SUGGESTED TIME
2.1.4 Acids · 2.1.5 Redox · PAG 2 Titration	100	2 hours

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Instructions to candidates

- Answer **all** questions in the spaces provided. Show all working in calculations.
- Include state symbols where equations are requested with them.
- Marks are shown in brackets [ ] at the end of each part.
- You may use a scientific calculator and the periodic table on the OCR Data Sheet.

Question 1 — Acids, bases and alkalis 2.1.4

a. Give the formulae of sulfuric acid and nitric acid. **[2]**

b. State what is released when an acid is dissolved in water. **[1]**

c. Explain the difference between a *strong* acid and a *weak* acid, giving one example of each. **[3]**

d. Define the terms *base* and *alkali*, making the difference between them clear. **[2]**

Question 1 total: 8 marks

Question 2 — Neutralisation 2.1.4

a. Write the full equation for the reaction of sulfuric acid with sodium hydroxide. **[2]**

b. Write the full equation, with state symbols, for the reaction of calcium carbonate with hydrochloric acid. **[2]**

c.

Write the full equation for the reaction of magnesium oxide with nitric acid. **[2]**

d. Write the equation for the reaction of ammonia with hydrochloric acid, and name the product. **[2]**

e. Write the ionic equation, with state symbols, for the reaction of any carbonate ion with any acid. **[2]**

Question 2 total: 10 marks

Question 3 — Acid–base titration

PAG 2

A student titrates 25.0 cm^3 portions of aqueous sodium carbonate of unknown concentration with 0.150 mol dm^{-3} hydrochloric acid. The mean titre is 23.40 cm^3.
$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

a. Describe the procedure for carrying out one accurate titration, from filling the burette to reading the endpoint. **[4]**

b. Name a suitable indicator for this titration and state its colour change at the endpoint. **[2]**

c. Explain why a conical flask, rather than a beaker, is used for the alkali. **[1]**

d. Calculate the concentration of the sodium carbonate solution in mol dm^{-3}. **[3]**

e. Convert your answer to a concentration in g dm^{-3}. **[2]**

Question 3 total: 12 marks

Question 4 — Oxidation numbers

2.1.5

a. Deduce the oxidation number of the stated element in each species. **[6]**

Species	Element	Oxidation number
SO_4^{2-}	S	
NO_3^-	N	
$Cr_2O_7^{2-}$	Cr	
NaClO	Cl	
MnO_4^-	Mn	
Fe_2O_3	Fe	

b. Oxygen usually has oxidation number −2 and hydrogen +1. Give one example of a compound in which each of these rules is *not* followed, stating the oxidation number in your example. **[2]**

c. Use oxidation numbers to explain the name *potassium chlorate(V)* and write its formula. **[2]**

Question 4 total: 10 marks

Question 5 — Identifying redox reactions

2.1.5

Zinc metal is added to aqueous copper(II) sulfate: $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

a. Define *oxidation* and *reduction* in terms of electron transfer. **[2]**

b. Use oxidation numbers to identify which species is oxidised and which is reduced in this reaction. Justify your answer. **[3]**

c. Write the two half-equations for this reaction. **[2]**

d.

Magnesium reacts with dilute hydrochloric acid. Write the full equation and use oxidation numbers to show that this is a redox reaction. **[3]**

Question 5 total: 10 marks

Question 6 — Redox equations

2.1.5

a. Acidified manganate(VII) ions oxidise iron(II) ions: **[4]**
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $Fe^{2+} \rightarrow Fe^{3+} + e^-$
Combine the two half-equations to give the overall ionic equation.

b. Chlorine reacts with water: $Cl_2 + H_2O \rightarrow HCl + HClO$. Assign oxidation numbers to chlorine in each species and explain why this reaction is described as *disproportionation*. **[4]**

c. Write the equation for the reaction of chlorine with cold dilute aqueous sodium hydroxide. **[2]**

Question 6 total: 10 marks

Question 7 — Metals and acids

2.1.4 / 2.1.5

a. State two observations when calcium is added to dilute sulfuric acid. **[2]**

b. Write the full equation, with state symbols, for this reaction. **[2]**

c. Identify, with oxidation numbers, the element oxidised and the element reduced. **[3]**

d.

Write the ionic equation for the reaction of calcium with dilute hydrochloric acid, with state symbols. **[3]**

Question 7 total: 10 marks

Question 8 — Gas tests

2.1.4

a. Describe the test for carbon dioxide gas and its positive result. **[2]**

b. Explain, with an equation, why limewater turns cloudy in a positive test. **[1]**

c. Describe the test for hydrogen gas and its positive result. **[2]**

d. Describe how to show that a solid contains ammonium ions, NH_4^+, including the observation for a positive result. **[3]**

Question 8 total: 8 marks

Question 9 — Percentage purity

2.1.3 / 2.1.4

A 1.00 g sample of impure calcium carbonate (the impurities do not react) is added to an excess of dilute hydrochloric acid. 220 cm^3 of carbon dioxide is collected at RTP.

a. Write the full equation for the reaction of calcium carbonate with hydrochloric acid. **[2]**

b. Calculate the percentage by mass of calcium carbonate in the sample. **[5]**

c.

State one assumption made in this calculation. **[1]**

Question 9 total: 8 marks

Question 10 — Planning a titration

PAG 2

Vinegar is a dilute solution of ethanoic acid, CH_3COOH. Plan an experiment to determine the concentration of ethanoic acid in a sample of vinegar by titration with a standard solution of sodium hydroxide. Include the apparatus, procedure, choice of indicator, and an outline of the calculation. **[6]**

This question is marked using a level of response.

a. Define the term *standard solution*. **[2]**

Question 10 total: 8 marks

Question 11 — Oxidation number drill

2.1.5

a. Deduce the oxidation number of the stated element in each species. **[6]**

Species	Element	Oxidation number
H_2O_2	O	
NaH	H	
NH_4^+	N	
PO_4^{3-}	P	
SO_3^{2-}	S	
CO_3^{2-}	C	

Question 11 total: 6 marks

PAPER TOTAL: 100 MARKS

MARK SCHEME — Chapter 4: Acids and Redox

Award marks independently unless stated. ECF applies to calculations.

Q1 (a) [2]

- H_2SO_4 (1); HNO_3 (1)

Q1 (b) [1]

- H^+ ions (protons) are released. (1)

Q1 (c) [3]

- Strong acid: fully/completely dissociates in aqueous solution, e.g. HCl / H_2SO_4 / HNO_3. (1 + 1 for example)
- Weak acid: only partially dissociates, e.g. CH_3COOH. (1 including example)

Q1 (d) [2]

- Base: a substance that neutralises an acid / accepts H^+ (e.g. metal oxide, hydroxide, carbonate). (1)
- Alkali: a base that dissolves in water, releasing OH^- ions. (1)

Q2 (a) [2]

- $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$ (1 species, 1 balancing)

Q2 (b) [2]

- $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ (1 balanced, 1 state symbols)

Q2 (c) [2]

- $MgO + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$ (1 species, 1 balancing)

Q2 (d) [2]

- $NH_3 + HCl \rightarrow NH_4Cl$ (1)
- Ammonium chloride. (1)

Q2 (e) [2]

- $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$ (1 species, 1 state symbols)

Q3 (a) [4]

- Rinse and fill the burette with acid, removing the air bubble from the jet; record the initial reading at eye level. (1)
- Pipette 25.0 cm^3 of the alkali into a conical flask, add a few drops of indicator. (1)
- Add acid with swirling; near the endpoint add dropwise. (1)
- Stop at the first permanent colour change; record the final reading; repeat until concordant. (1)

Q3 (b) [2]

- Methyl orange. (1)
- Yellow → red (accept orange at endpoint). (1)
- *Phenolphthalein is unsuitable: it changes before the second equivalence point of carbonate.*

Q3 (c) [1]

- A conical flask can be swirled without splashing/loss of solution. (1)

Q3 (d) [3]

- $n(HCl) = 0.02340 \times 0.150 = 3.51 \times 10^{-3}$ mol (1)
- $n(Na_2CO_3) = 3.51 \times 10^{-3} \div 2 = 1.755 \times 10^{-3}$ mol (1)
- $c = 1.755 \times 10^{-3} \div 0.0250 = 0.0702$ mol dm^{-3}. (1)

Q3 (e) [2]

- $M(Na_2CO_3) = 106$ g mol^{-1}; 0.0702×106 (1) $= 7.44$ g dm^{-3}. (1)

Q4 (a) [6] — 1 mark each

- S in $SO_4^{2-} = +6$ · N in $NO_3^- = +5$ · Cr in $Cr_2O_7^{2-} = +6$
- Cl in NaClO = +1 · Mn in $MnO_4^- = +7$ · Fe in $Fe_2O_3 = +3$

Q4 (b) [2]

- O = −1 in peroxides, e.g. H_2O_2. (1)
- H = −1 in metal hydrides, e.g. NaH. (1)

Q4 (c) [2]

- The (V) shows chlorine has oxidation number +5 in the chlorate ion. (1)
- $KClO_3$. (1)

Q5 (a) [2]

- Oxidation: loss of electrons. (1)
- Reduction: gain of electrons. (1)

Q5 (b) [3]

- Zn: 0 → +2, so zinc is oxidised. (1)
- Cu: +2 → 0, so copper(II) is reduced. (1)
- Correct justification linking increase/decrease of oxidation number to oxidation/reduction. (1)

Q5 (c) [2]

- $Zn \rightarrow Zn^{2+} + 2e^-$ (1)
- $Cu^{2+} + 2e^- \rightarrow Cu$ (1)

Q5 (d) [3]

- $Mg + 2HCl \rightarrow MgCl_2 + H_2$ (1)
- Mg: 0 → +2 (oxidised); H: +1 → 0 (reduced). (1)
- Both oxidation and reduction occur together, so the reaction is redox. (Cl stays −1, a spectator.) (1)

Q6 (a) [4]

- Multiply the iron half-equation by 5 to balance the electrons. (1)
- $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$ (2 for fully correct; 1 if one error)
- No electrons remaining in the overall equation; charges balance (+17 each side). (1)

Q6 (b) [4]

- Cl in Cl_2 = 0. (1)
- Cl in HCl = −1; Cl in HClO = +1. (1)
- The same element (chlorine) is simultaneously oxidised (0 → +1)... (1)
- ...and reduced (0 → −1), which is disproportionation. (1)

Q6 (c) [2]

- $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$ (1 species, 1 balancing)

Q7 (a) [2]

- Effervescence / bubbles of gas. (1)
- The metal dissolves / solution warms up. (Accept: white solid $CaSO_4$ may form.) (1)

Q7 (b) [2]

- $Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2(g)$ (1 balanced, 1 state symbols)

Q7 (c) [3]

- Ca: 0 → +2, oxidised. (1)
- H: +1 → 0, reduced. (1)
- S and O unchanged (+6, −2) — spectators. (1)

Q7 (d) [3]

- $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$ (1 species, 1 balancing/charges, 1 state symbols)

Q8 (a) [2]

- Bubble the gas through limewater. (1)
- Positive: limewater turns cloudy/milky. (1)

Q8 (b) [1]

- $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$ — insoluble calcium carbonate forms. (1)

Q8 (c) [2]

- Hold a lighted splint at the mouth of the tube. (1)
- Positive: the gas ignites with a squeaky pop. (1)

Q8 (d) [3]

- Warm the solid with aqueous sodium hydroxide. (1)
- Test the gas released with damp red litmus paper. (1)
- Positive: litmus turns blue (ammonia gas is released). (1)

Q9 (a) [2]

- $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ (1 species, 1 balancing)

Q9 (b) [5]

- $n(CO_2) = 220 \div 24\,000 = 9.17 \times 10^{-3}$ mol (1)
- 1 : 1 ratio → $n(CaCO_3) = 9.17 \times 10^{-3}$ mol (1)
- $m(CaCO_3) = 9.17 \times 10^{-3} \times 100.1$ (1) = 0.918 g (1)
- % purity = 91.8% (accept 91.7–91.8). (1)

Q9 (c) [1]

- The impurities do not react with the acid / do not release gas (or: all CO_2 is collected). (1)

Q10 (a) [6] — Level of response

- **Level 3 (5–6):** Complete, workable plan with correct apparatus, procedure, indicator and calculation outline.
- **Level 2 (3–4):** Workable plan with minor omissions (e.g. no dilution consideration or vague endpoint description).
- **Level 1 (1–2):** Fragmentary plan; major steps missing.

Indicative content: *pipette 25.0 cm^3 of vinegar (diluted accurately if concentrated) into a conical flask; add phenolphthalein — suitable for a weak acid–strong base titration (colourless → pink); titrate with standard NaOH from a burette to the first permanent pale pink; repeat to concordance. Calculation: $n(NaOH) = c \times V$; 1:1 ratio → $n(CH_3COOH)$; $c = n \div 0.0250$ (× dilution factor if diluted).*

Q10 (b) [2]

- A solution of accurately known concentration... (1)
- ...prepared by dissolving a known mass in a known volume (volumetric flask). (1)

Q11 (a) [6] — 1 mark each

- O in H_2O_2 = −1 · H in NaH = −1 · N in NH_4^+ = −3
- P in PO_4^{3-} = +5 · S in SO_3^{2-} = +4 · C in CO_3^{2-} = +4

END OF MARK SCHEME · 100 marks