NAME DATE STARTED TARGET END DATE

GCE A LEVEL — CHEMISTRY UNIT 3 QUESTION PACK

1095-01 (Legacy CH5) · New spec Unit 3 Topic 1 · A2 unit, first sat 2017, 80 marks, 1h 45min paper

REVISE.wales

CHEMISTRY — UNIT 3 · Redox, Electrode Potentials & Titrations

Topics 3.1 & 3.2 — Redox half-equations, standard electrode potentials and redox titrations

Balancing redox half-equations, building electrochemical cells from standard electrode potentials, predicting feasibility from E° values, and carrying out manganate(VII), dichromate(VI) and iodine–thiosulfate titrations.

Legacy 2008 specification

Estimated time for entire question pack: ~2 h 29 min

Derived from the legacy CH5 paper's pace of ~1.3 min/mark, padded for long-prose and calculation answers (93 marks over 5 questions).

*You are advised to **not** attempt to complete all of this in one sitting.*

ABOUT THIS QUESTION PACK

This is a **comprehensive practice question pack**, not a single mock paper. It contains every question from the legacy WJEC CH5 papers (2008 modular spec, Jun 2010 – Jun 2016) that maps onto the new-spec A2 Unit 3 Topic 3.1 & 3.2.

Questions are ordered by source paper date.

INSTRUCTIONS

Use black ink or black ball-point pen. Show all working — quality of written communication will affect marks. A calculator is allowed. You will need the WJEC Periodic Table / Data Booklet.

For Examiner's use only

Q	Source	Max	Mark
1	Jun 10 Q3	15	
2	Jun 10 Q5	20	
3	Jun 13 Q4	20	
4	Jun 14 Q3	18	
5	Jun 15 Q4	20	
Total		**93**	

Redox, Electrode Potentials & Titrations — what the new spec asks

WJEC GCE A Level Chemistry (from 2015) · Unit 3: Physical & Inorganic Chemistry · Topic 3.1 & 3.2.

Oxidation states

- Element: 0. Monatomic ion: charge.
- Group 1: +1; Group 2: +2; Al: +3.
- H: +1 (except metal hydrides −1); O: −2 (except peroxides −1, F_2O +2).
- Sum of oxidation states = total charge.

Half-equations & balancing

- Write oxidation and reduction halves separately.
- Balance atoms (other than O, H), then O with H_2O, H with H^+, charge with e^-.
- Multiply through so electrons cancel; add halves.
- Identify oxidising agent (accepts e^-) vs reducing agent (donates).

Standard electrode potentials E°

- Measured against standard hydrogen electrode (SHE = 0.00 V).
- Standard conditions: 298 K, 1 mol dm^{-3}, 1 atm.
- $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ (right minus left).
- Positive E°_{cell} ⇒ reaction feasible (thermodynamically).

Manganate(VII) titrations

- $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$.
- Acidic with H_2SO_4 (HCl reacts with MnO_4^-).
- Self-indicating: deep purple → colourless; first persistent pink = end point.
- Titrate against Fe^{2+}, ethanedioate, H_2O_2.

Iodine/thiosulfate titrations

- $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$.
- Pale-yellow I_2; add starch near end point (deep blue-black).
- End point: blue colour disappears.
- Use to determine Cu^{2+}, ClO^-, IO_3^-.

Dichromate(VI) titrations

- $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$.
- Orange → green colour change (not self-indicating, use sodium diphenylamine sulfonate).
- Can use HCl as acid (unlike MnO_4^-).

Redox, Electrode Potentials & Titrations in one page

Quick-reference notes — revisit before each question.

Half-eqn balancing

Atoms (non-O,H) → O with H_2O → H with H^+ → charge with e^-. Cancel electrons across both halves before adding.

MnO_4^-/Fe^{2+}

$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$. Self-indicating: first persistent pink = end point.

E°_{cell}

$E^\circ_{cell} = E^\circ_{RHS} - E^\circ_{LHS}$. Positive ⇒ feasible.
$\Delta G^\circ = -nFE^\circ_{cell}$.

$I_2/S_2O_3^{2-}$

$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$. Starch only near end point (would otherwise bind). Blue → colourless.

Cell diagram

Pt | H_2(g) | H^+(aq) ‖ Cu^{2+}(aq) | Cu(s)

Left = anode (oxidation); right = cathode (reduction).

Pitfalls

- Don't add electrons to overall eq.
- Use H_2SO_4 for MnO_4^- (HCl reduces it).
- E° assumes std states; deviations shift feasibility.

Examiner only

3. *Read the passage below and then answer questions (a) to (d) in the spaces provided.*

Hydrogen Fuel Cells

Although fuel cells have been around since 1839, it took another 120 years until NASA demonstrated some of their potential applications when providing power during space flights.

A fuel cell works like an electrochemical cell (battery) but does not run down or need recharging. It will produce electricity and heat as long as fuel (hydrogen) is supplied. A fuel cell consists of two electrodes—an anode where oxidation occurs and a cathode for reduction—sandwiched around an electrolyte. Replacing the salt bridge of conventional electrochemical cells, several electrolyte systems have been tried such as phosphoric acid or a solid electrolyte based on polymeric fluorocarbons. The relevant electrode potentials are

$$2H^+ + 2e^- \rightleftharpoons H_2 \qquad E^\ominus = 0\,V$$

$$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O \qquad E^\ominus = 1.23\,V$$

Hydrogen is fed to the anode, and oxygen (air) to the cathode. Activated by a catalyst, usually involving a layer of platinum and carbon a few nanometres thick, hydrogen atoms separate into protons and electrons, which take different paths to the cathode. The electrons go through an external circuit, creating a flow of electricity. The protons migrate through the electrolyte. Fuel cells can be used to power vehicles or to provide electricity and heat to buildings.

A significant barrier to using fuel cells in vehicles is hydrogen storage. Most fuel-cell vehicles powered by hydrogen store the hydrogen as a compressed gas in pressurized tanks. Due to the low energy density of hydrogen, it is difficult to store enough hydrogen onboard to allow vehicles to travel the same distance as petrol-powered vehicles.

A potentially energy-dense water-based fuel is based on sodium tetrahydridoborate(III) (30% by mass $NaBH_4$ in water). A catalyst induces rapid hydrogen production

$$NaBH_4 + 2H_2O \rightarrow 4H_2 + NaBO_2 \qquad \Delta H^\ominus = -300\,kJ\,mol^{-1}$$

and pure humidified H_2 is delivered to the engine or fuel cell. The exothermic reaction requires no heat input and sodium borate, $NaBO_2$, can be recycled into $NaBH_4$.

– End of passage –

Examiner only

(a) State the function performed by both the salt bridge in an electrochemical cell and the electrolyte in a fuel cell. *(lines 6-7)* [2]

....

....

(b) (i) Explain why the $2H^+ + 2e^- \rightleftharpoons H_2$ electrode has an electrode potential of zero. *(line 10)* [1]

....

....

(ii) Calculate the EMF of the hydrogen fuel cell. *(lines 10-11)* [1]

....

....

(iii) Give **one** reason why the EMF calculated in *(b)*(ii) is not attained in practice, with 0.7 V being a typical value for a fuel cell. [1]

....

....

(iv) Write a balanced equation for the overall reaction which occurs in the cell. *(lines 10-11)* [1]

....

....

(v) Given that $\Delta H_f^{\ominus}$ $H_2O(l) = -285.8\,kJ\,mol^{-1}$, calculate the enthalpy change, $\Delta H^{\ominus}$, for the equation in *(b)*(iv). [1]

....

....

Examiner only

(c) (i) State **one** disadvantage, mentioned in the passage, of using hydrogen fuel cells to power vehicles. *(lines 18-21)* [1]

....

....

(ii) Give a second disadvantage, not mentioned in the passage, of using hydrogen as a fuel in vehicles. [1]

....

....

(iii) State **one** advantage of using a hydrogen fuel cell compared to the combustion of petrol. [1]

....

....

(d) When 1 kg of the water-based fuel (30% $NaBH_4$ by mass) is reacted to produce hydrogen, calculate *(lines 22-26)*

(i) the mass, and hence the number of moles, of $NaBH_4$ in 1 kg of the water-based fuel, [2]

....

....

(ii) the energy given out (kJ) by 1 kg of the water-based fuel, [1]

....

....

(iii) the volume of hydrogen gas produced. [2]

[Assume 1 mol H_2 gas occupies a volume of $24\,dm^3$]

....

....

....

....

Total [15]

(b) Both boron nitride, BN, and carbon, C, form hexagonal graphite-type structures.
Explain why
- BN and C can both adopt the same hexagonal structure;
- both BN and C exhibit lubricating properties;
- C is an electrical conductor but BN is an insulator at room temperature. [6]

(QWC) [2]

Total [20]

5. *(a)* *Bordeaux Mixture* is one of the earliest fungicides, first used about 1885. It can be prepared by mixing copper sulfate solution with excess limewater (calcium hydroxide solution).

(i) State what you would observe when copper sulfate solution is mixed with limewater. [2]

(ii) Write an equation for the reaction that occurs. [1]

(b) A sample of *Bordeaux Mixture* was analysed to determine its copper content. Firstly, it was reacted with excess potassium iodide

$$2Cu^{2+} + 4I^{-} \rightarrow 2CuI + I_2$$

and the iodine produced was then titrated against sodium thiosulfate solution.

$$I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$$

(i) Name the indicator used for the titration and state the colour change at the end-point. [2]

(ii) If a 31.2 g sample of *Bordeaux Mixture* required 12.25 cm^3 of sodium thiosulfate solution with concentration 0.100 mol dm^{-3} $Na_2S_2O_3$ to react with the liberated iodine, calculate the mass of copper in the sample and hence the % Cu by mass in *Bordeaux Mixture*. Your answers should be given to **three** significant figures. [3]

(c) Copper can exist as Cu^{2+} or Cu^{+} compounds.

(i) Write the full electron configurations for Cu^{2+} ions **and** Cu^{+} ions. [2]

(ii) Explain why most Cu^{2+} compounds are coloured blue in the presence of water. [4]

(iii) Briefly explain why most Cu^{+} compounds are colourless or white. [1]

(d) (i) State what would be observed, and give equations for any reactions, when tetrachloromethane, CCl_4, and silicon(IV) chloride, $SiCl_4$, are separately added to water. [3]

(ii) Explain why lead forms a solid chloride $PbCl_2$, but the corresponding CCl_2 and $SiCl_2$ are too unstable to exist. [2]

Total [20]

GCE A level

1095/01-A

CHEMISTRY CH5
DATA SHEET

A.M. MONDAY, 28 June 2010

SECTION B

Answer **both** *questions in the separate answer book provided.*

4. *(a)* Electrochemical cells are used as power sources in many everyday applications. To decide what to use in a cell, it is necessary to know the standard electrode potential for electrodes. This is measured using a standard hydrogen electrode as a reference standard.

Draw a labelled diagram of the apparatus you would use to measure the standard electrode potential of an Fe^{3+}/Fe^{2+} electrode. [5]

(b) Vanadium is a transition metal that can form compounds with a variety of oxidation states. Zinc however forms compounds with an oxidation state of +2 only.

(i) Why can transition elements form compounds with a variety of oxidation states? [1]

(ii) Give the electronic structure of Zn. [1]

(iii) State why zinc forms Zn^{2+}. [1]

You will need the standard electrode potentials in the table below to answer part *(c)*.

Oxidation state of vanadium at start of reaction	Reaction	$E^\ominus$/V
+5	$VO_3^-(aq) + 4H^+(aq) + e \rightleftharpoons VO^{2+}(aq) + 2H_2O(l)$	+1.00
+4	$VO^{2+}(aq) + 2H^+(aq) + e \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
+3	$V^{3+}(aq) + e \rightleftharpoons V^{2+}(aq)$	–0.26
+2	$V^{2+}(aq) + 2e \rightleftharpoons V(s)$	–1.13
	$Zn^{2+}(aq) + 2e \rightleftharpoons Zn(s)$	–0.76
	$Cu^{2+}(aq) + 2e \rightleftharpoons Cu(s)$	+0.34

(c) Vanadium(V)(aq), as VO_3^-, is yellow and can be reduced by zinc and aqueous acid producing a series of coloured solutions until the reduction stops with the formation of a violet solution. The reducing agent involves the $Zn^{2+}(aq)/Zn(s)$ equilibrium.

(i) State the identity of the violet vanadium-containing solution produced in this reduction. Use standard electrode potentials to explain your answer. [3]

(ii) What is the standard potential of a cell formed from a standard $Zn^{2+}(aq)/Zn(s)$ electrode and a standard $Cu^{2+}(aq)/Cu(s)$ electrode? [1]

(iii) Write the equilibrium equation for the change occurring at the zinc electrode showing the direction in which the reaction proceeds. [1]

(iv) Use Le Chatelier's principle to predict the effect on the electrode potential of the zinc electrode of increasing the concentration of $Zn^{2+}(aq)$ in the electrode. Explain your answer. [2]

(d) Halogens can also form compounds with a variety of oxidation states. Some of these including compounds of iodate(V), IO_3^-, behave as oxidising agents.

A student was investigating the reaction that occurs when iodate(V) oxidises iodide ions to produce iodine. Two possible equations were suggested.

$$IO_3^- + 6H^+ + 5I^- \longrightarrow 3I_2 + 3H_2O \qquad \textbf{equation 1}$$

$$IO_3^- + 4H^+ + 4I^- \longrightarrow IO^- + 2H_2O + 2I_2 \qquad \textbf{equation 2}$$

He prepared a solution of potassium iodate(V) by dissolving 0.978 g of KIO_3 in 250 cm^3 of solution. He pipetted 25.0 cm^3 of this solution into a conical flask, added excess potassium iodide and titrated the iodine produced with 0.100 mol dm^{-3} sodium thiosulfate solution, $Na_2S_2O_3$. A volume of 27.40 cm^3 of this solution was needed to react with the iodate(V).

The equation for the reaction of thiosulfate with iodine is shown below.

$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$

(i) Calculate the number of moles of thiosulfate used to react with the iodine. [1]

(ii) Deduce the number of moles of iodine present in the 25.0 cm^3 sample. [1]

(iii) Calculate the number of moles of KIO_3 present in 250 cm^3 of the original solution and hence the number of moles present in 25.0 cm^3. [1]

(iv) Use your results from (ii) and (iii) to deduce which of **equation 1** and **equation 2** suggested above, correctly shows what happens when iodate(V) ions oxidise iodide ions. Show, by calculation, how you came to this conclusion. [2]

Total [20]

3. *Read the passage below and then answer the questions in the spaces provided.*

Hydrogen Peroxide

If a non-scientist knows only one chemical formula it is most likely to be H_2O for water but how much do you know about another hydrogen oxide, hydrogen peroxide? A molecule of hydrogen peroxide has the molecular formula H_2O_2.

Most chemistry students first meet hydrogen peroxide as a colourless solution that is used to prepare oxygen. Bottles of hydrogen peroxide from a pharmacist are often labelled '20 volume'. This means that one volume of solution decomposes to give 20 volumes of oxygen gas. The equation for the decomposition is:

$$\underset{1\,dm^3}{2H_2O_2(aq)} \longrightarrow 2H_2O(l) + \underset{20\,dm^3}{O_2(g)}$$

This reaction is very slow at room temperature. However the addition of a suitable catalyst increases the rate of decomposition phenomenally. Manganese(IV) oxide, potatoes and blood are all effective. Potatoes and blood both contain the enzyme catalase and one catalase molecule decomposes 50 000 molecules of H_2O_2 per second!

Is hydrogen peroxide an oxidising agent or a reducing agent?

Both in the laboratory and at home hydrogen peroxide is most commonly used as an oxidising agent (so the hydrogen peroxide itself is reduced). The half-equation is:

Reduction $$H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$$

Since some colouring matter is bleached by oxidation and the product of hydrogen peroxide's reduction is water, it is used as a safe bleaching agent particularly in hair treatment. A peroxide blonde is someone with almost white hair, usually as a result of treatment with hydrogen peroxide.

However, if hydrogen peroxide reacts with a more powerful oxidising agent such as potassium manganate(VII), the hydrogen peroxide will act as a reducing agent and will itself be oxidised. The half-equation is:

Oxidation $$H_2O_2 \longrightarrow 2H^+ + O_2 + 2e^-$$

Therefore hydrogen peroxide can act as both oxidising agent and reducing agent. In fact, it can react with itself so that alternate molecules are oxidised and reduced. The overall equation is obtained by adding the half-equations for the reduction and oxidation, giving

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

which is the standard decomposition equation!

- End of passage -

Examiner only

(a) Using outer electrons only, draw a dot and cross diagram to show the bonding in a hydrogen peroxide molecule *(line 3)*. [1]

(b) Use the equation for the decomposition of hydrogen peroxide *(line 8)* to calculate the concentration, in mol dm^{-3}, of aqueous hydrogen peroxide solution in a bottle of '20 volume hydrogen peroxide' at 25 °C. [2]

[1 mol of oxygen occupies 24 dm^3 at 25 °C]

Concentration = ... mol dm^{-3}

(c) Manganese(IV) oxide *(line 10)* and potassium manganate(VII) *(lines 20-21)* are typical transition metal compounds.

(i) Give **two** reasons why transition metal compounds can act as catalysts. [2]

....

....

....

(ii) Explain why transition metal complex ions appear coloured. [4]
QWC [1]

....

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Examiner only

(d) In an acidic solution, hydrogen peroxide is oxidised to oxygen by potassium manganate (VII) *(lines 20-23)*.

(i) Write the half-equation for the reduction of MnO_4^- to Mn^{2+} ions in acidic solution. [1]

....

(ii) Use your answer to (i) and the half-equation given in *line 23* to deduce the overall equation for this reaction. [2]

....

....

....

(iii) 20.0 cm^3 of an acidified solution of hydrogen peroxide required 14.80 cm^3 of a 0.020 mol dm^{-3} solution of potassium manganate(VII) for complete reaction.
Calculate the concentration, in mol dm^{-3}, of the hydrogen peroxide solution. [3]

Concentration = ... mol dm^{-3}

(e) Explain, using oxidation states, why the decomposition of hydrogen peroxide *(line 27)* can be classified as a redox reaction. [2]

....

....

....

Total [18]

18

Total Section A [40]

SECTION B

Answer **both** *questions in the separate answer book provided.*

4. The leaves of the rhubarb plant are rich in ethanedioic acid (oxalic acid) which is a poisonous compound. A solution containing ethanedioate ions can be formed by boiling rhubarb leaves with water. It can be separated and samples titrated against acidified potassium manganate(VII) to find the concentration of the ethanedioate solution.

(a) Suggest how the ethanedioate solution could be separated from the rhubarb leaves. [1]

(b) Write an ion-electron half-equation for the reduction of acidified manganate(VII) ions, MnO_4^-. [1]

(c) The ion-electron half-equation for the oxidation of ethanedioate ions is given below.

$$C_2O_4^{2-}(aq) \longrightarrow 2CO_2(g) + 2e^-$$

(i) Give the oxidation states for carbon at the start and end of this reaction. [1]

(ii) Write an equation for the reaction of acidified manganate(VII) ions with ethanedioate ions. [1]

(d) Give a reason why an indicator is not needed in this titration. [1]

(e) Four samples of 25.00 cm^3 of the ethanedioate solution were titrated against acidified potassium manganate(VII) solution of concentration 0.0200 mol dm^{-3}. The volumes of potassium manganate(VII) solution required for complete reaction are listed below.

	1	2	3	4
Volume of $KMnO_4$(aq) / cm^3	28.80	27.95	28.00	27.80

Use the information given to calculate the concentration of the ethanedioate solution. [4]

(f) Heating ethanedioic acid in glycerol produces methanoic acid, HCOOH.

(i) Write the expression for the acid dissociation constant, K_a, for methanoic acid. [1]

(ii) The value of K_a for methanoic acid is 1.8×10^{-4} mol dm^{-3}.
Calculate the pH of a solution of methanoic acid of concentration 0.2 mol dm^{-3}. [3]

(iii) A mixture of methanoic acid and sodium methanoate can be used as a buffer solution. State what is meant by a *buffer solution* and explain how a mixture of methanoic acid and sodium methanoate acts as a buffer. [3]
QWC [1]

(g) Acidified potassium dichromate, $K_2Cr_2O_7$, is also an oxidising agent.

(i) Give the colour change that occurs when acidified potassium dichromate acts as an oxidising agent. [1]

(ii) When sodium hydroxide is added to a solution of potassium dichromate, a colour change occurs without a redox reaction occurring. Give the formula of the new chromium-containing ion and the colour of the solution formed. [2]

Total [20]

END OF QUESTION PACK

5 questions · 93 marks · ~2 h 29 min
Source: WJEC CH5 (2008 modular spec, Jun 2010 – Jun 2016)
Curated for WJEC Chemistry 2015 spec A2 Unit 3 — Topic 1 (3.1 & 3.2)