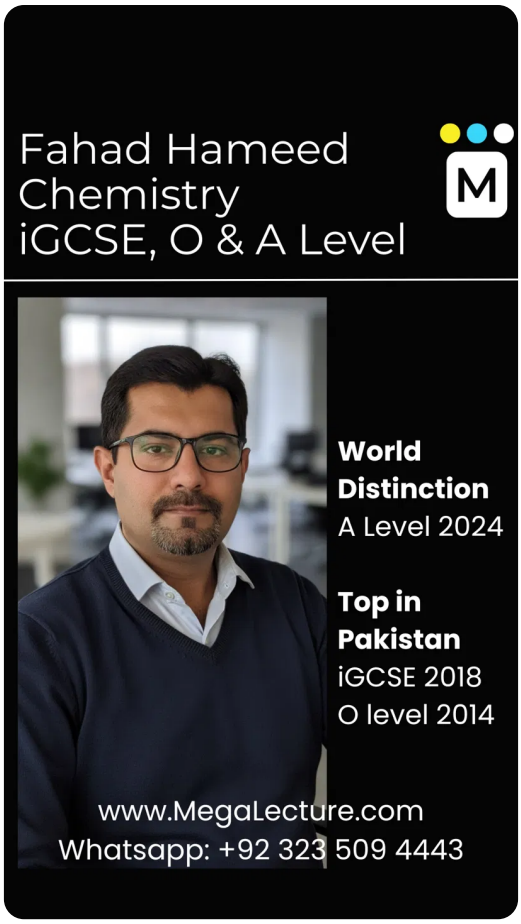
MEGA LECTURE · A LEVEL CHEMISTRY

Chapter 24: Redox and Electrode Potentials

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



SPECIFICATION	TOTAL MARKS	SUGGESTED TIME
3.3.5 Redox and electrode potentials: cells, E°, feasibility, applications	94	1 hour 53 minutes

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Instructions to candidates

- Answer **all** questions in the spaces provided. Show all working in calculations.
- Marks are shown in brackets [ ] at the end of each part.

Question 1 — Oxidation, reduction and half-equations 3.3.5

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

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

b. Using oxidation numbers, identify the species oxidised and the species reduced in the reaction above. **[2]**

c. Write the two half-equations for this reaction, including state symbols. **[2]**

d. State which species is the reducing agent in this reaction, and explain your answer. **[2]**

Question 1 total: 8 marks

Question 2 — Standard electrode potential and the standard hydrogen electrode 3.3.5

a. Define the term *standard electrode potential* of a half-cell. **[2]**

b. **[4]**

Sketch and label a standard hydrogen electrode connected to a metal/metal-ion half-cell, to show how a standard electrode potential is measured.

Sketch: standard hydrogen electrode, salt bridge, voltmeter, test half-cell

c. State the value assigned to the standard electrode potential of the standard hydrogen electrode, and explain why this value is used. **[2]**

d. State two conditions, other than ion concentration and gas pressure, that must be standard when measuring a standard electrode potential. **[2]**

Question 2 total: 10 marks

Question 3 — Constructing an electrochemical cell 3.3.5

A cell is made by connecting a $Cu^{2+}(aq)/Cu(s)$ half-cell ($E° = +0.34$ V) to a $Zn^{2+}(aq)/Zn(s)$ half-cell ($E° = -0.76$ V).

a. Sketch and label a diagram of this electrochemical cell, showing the salt bridge, voltmeter, and the direction of conventional current flow in the external circuit. **[4]**

Sketch: Cu^{2+}/Cu half-cell — salt bridge — Zn^{2+}/Zn half-cell, with voltmeter

b. State the purpose of the salt bridge in this cell. **[2]**

c. Identify which electrode is the negative electrode (anode) and which is the positive electrode (cathode), with a reason. **[2]**

d.

State the direction of electron flow in the external circuit of this cell. **[2]**

Question 3 total: 10 marks

Question 4 — Calculating standard cell potential (I) 3.3.5

a. State the formula used to calculate standard cell potential, $E°_{cell}$, from two standard electrode potentials. **[1]**

b. Calculate $E°_{cell}$ for the cell formed from the Cu^{2+}/Cu ($E° = +0.34$ V) and Zn^{2+}/Zn ($E° = -0.76$ V) half-cells. **[3]**

c. Write the overall cell (redox) equation for this reaction, and predict, with a reason, whether the forward reaction is feasible. **[3]**

d. A cell is made from Mg^{2+}/Mg ($E° = -2.37$ V) and Ag^{+}/Ag ($E° = +0.80$ V) half-cells. Calculate $E°_{cell}$ and write the overall (balanced) equation for the feasible reaction. **[3]**

Question 4 total: 10 marks

Question 5 — Using a data table of electrode potentials 3.3.5

Half-equation	E° / V
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$I_2(aq) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

a. Using the data above, determine which of Fe^{3+}/Fe^{2+} and I_2/I^- is reduced, and calculate $E°_{cell}$ for the feasible reaction between $Fe^{3+}(aq)$ and $I^-(aq)$. **[3]**

b. Write the overall ionic equation for the feasible reaction between $Fe^{3+}(aq)$ and $I^{-}(aq)$, ensuring electrons are balanced. **[3]**

c. Use the data table to determine whether acidified $MnO_4^{-}(aq)$ can oxidise $Cl^{-}(aq)$ to $Cl_2(g)$, and calculate $E°_{cell}$ for this reaction. **[3]**

d. Write the overall ionic equation for the reaction between acidified potassium manganate(VII) and chloride ions. **[3]**

Question 5 total: 12 marks

Question 6 — Further $E°_{cell}$ calculations

3.3.5

a. A cell is made from a Ni^{2+}/Ni half-cell ($E° = -0.25$ V) and a Pb^{2+}/Pb half-cell ($E° = -0.13$ V). Determine $E°_{cell}$ and the overall equation for the feasible reaction. **[4]**

b. Given $E°(Cl_2/Cl^{-}) = +1.36$ V and $E°(Br_2/Br^{-}) = +1.09$ V, explain, in terms of electrode potentials, why chlorine water can displace bromide ions from solution to form bromine, but bromine water cannot displace chloride ions. **[3]**

c. Given $E°(Ag^+/Ag) = +0.80$ V and $E°(Fe^{3+}/Fe^{2+}) = +0.77$ V, predict, with a calculation, whether silver metal will react with $Fe^{3+}(aq)$ ions. **[3]**

Question 6 total: 10 marks

Question 7 — Limitations of standard electrode potentials 3.3.5

a. Explain why a reaction with a positive $E°_{cell}$ may not actually be observed to occur, even though it is thermodynamically feasible. **[2]**

b. Explain why using standard electrode potentials to predict the feasibility of a reaction can be unreliable if the reaction takes place under non-standard conditions. **[3]**

c. The reaction between acidified $MnO_4^-(aq)$ and $Cl^-(aq)$ has a small positive $E°_{cell}$ (+0.15 V, from Question 5). Suggest reasons why reactions with only a small positive $E°_{cell}$ value are sometimes found not to proceed as predicted. **[3]**

Question 7 total: 8 marks

Question 8 — Commercial applications: primary and secondary cells 3.3.5

a. Distinguish between a primary cell and a secondary cell, giving one example of each. **[2]**

b.

Explain, in terms of electrode reactions, how recharging a secondary cell is possible. **[3]**

c. State three advantages of rechargeable (secondary) cells compared with non-rechargeable (primary) cells. **[3]**

d. State one disadvantage of secondary cells compared with primary cells. **[2]**

Question 8 total: 10 marks

Question 9 — Fuel cells 3.3.5

a. Write the half-equation for the oxidation of hydrogen at the negative electrode of a hydrogen–oxygen fuel cell operating under acidic conditions. **[2]**

b. Write the half-equation for the reduction of oxygen at the positive electrode of the same fuel cell under acidic conditions. **[2]**

c. Write the half-equation for the reduction of oxygen at the positive electrode of a hydrogen–oxygen fuel cell operating under alkaline conditions. **[2]**

d. Combine appropriate half-equations to give the overall equation for the reaction occurring in a hydrogen–oxygen fuel cell, and state the only product formed. **[2]**

e. State one advantage and one disadvantage of using hydrogen fuel cells, compared with rechargeable batteries or fossil fuels, in vehicles. **[2]**

Question 9 total: 10 marks

Question 10 — Extended response

Synoptic

A student proposes using standard electrode potentials to predict whether hydrogen fuel cells are a viable, environmentally friendly replacement for petrol/diesel engines and lithium-ion batteries in vehicles. Using ideas about electrode potentials, cell feasibility, and the limitations of E° values in predicting real-world behaviour, discuss the extent to which standard electrode potential data alone can be used to judge the suitability of hydrogen fuel cells as an energy source for transport. **[6]**

This question is marked using a level of response.

Question 10 total: 6 marks

PAPER TOTAL: 94 MARKS

MARK SCHEME — Chapter 24: Redox and Electrode Potentials

Award marks independently unless stated. ECF applies to calculations.

Q1 (a) [2]

- Oxidation is the loss of electrons (1); reduction is the gain of electrons (1).

Q1 (b) [2]

- Zn is oxidised — its oxidation number increases from 0 to +2 (1).
- Cu^{2+} is reduced — its oxidation number decreases from +2 to 0 (1).

Q1 (c) [2]

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

Q1 (d) [2]

- Zn is the reducing agent (1); because it reduces Cu^{2+} (donates electrons to Cu^{2+}) and is itself oxidised (1).

Q2 (a) [2]

- The e.m.f. (voltage) measured when a half-cell is connected to a standard hydrogen electrode (1), under standard conditions (298 K, 100 kPa gas pressure, 1 mol dm^{-3} ion concentration) (1).

Q2 (b) [4]

- Platinum electrode (coated in platinum black) in contact with H_2 gas at 100 kPa and 1 mol dm^{-3} $H^+(aq)$ (1).
- Salt bridge (e.g. containing KNO_3 solution) connecting the two half-cells (1).
- High-resistance voltmeter connected between the two electrodes, completing the external circuit (1).
- All components correctly labelled and both half-cells shown at 298 K (1).

Q2 (c) [2]

- 0.00 V (1).
- It is used as an arbitrary reference (zero) point against which all other standard electrode potentials are measured, since absolute electrode potentials cannot be measured directly (1).

Q2 (d) [2]

- Temperature must be 298 K (1).
- Any solid electrodes/metals must be in their standard states (e.g. a platinum electrode is used where no solid metal is present in the half-cell) (1).

Q3 (a) [4]

- Two beakers, each containing a metal electrode dipped in a 1 mol dm^{-3} solution of its own ions (1).
- Salt bridge (e.g. filter paper soaked in $KNO_3(aq)$) connecting the two solutions (1).
- Voltmeter connected in the external circuit between the two electrodes (1).
- Conventional current shown flowing through the external circuit from the copper (positive) electrode to the zinc (negative) electrode (1).

Q3 (b) [2]

- Completes the circuit by allowing ions to flow between the two half-cells (1).
- Maintains electrical neutrality in each half-cell as electrons flow around the external circuit (1).

Q3 (c) [2]

- Zinc is the negative electrode (anode), as it has the more negative electrode potential/is oxidised, releasing electrons (1).
- Copper is the positive electrode (cathode), as it has the more positive electrode potential/is reduced, accepting electrons (1).

Q3 (d) [2]

- Electrons flow from the zinc electrode to the copper electrode (i.e. from the more negative to the more positive electrode) through the external circuit (2).

Q4 (a) [1]

- E°_{cell} = E°(reduction/cathode, more positive) − E°(oxidation/anode, more negative) (1)

Q4 (b) [3]

- $E°_{cell} = E°(Cu^{2+}/Cu) - E°(Zn^{2+}/Zn)$ (1)
- $= (+0.34) - (-0.76)$ (1)
- $= +1.10$ V (1)

Q4 (c) [3]

- $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ (1)
- $E°_{cell}$ is positive (+1.10 V) (1)
- So the reaction as written is (thermodynamically) feasible (1)

Q4 (d) [3]

- $E°_{cell} = E°(Ag^{+}/Ag) - E°(Mg^{2+}/Mg) = (+0.80) - (-2.37) = +3.17$ V (1)
- Overall equation: $Mg(s) + 2Ag^{+}(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$ (1)
- Correctly balanced ($2Ag^{+}/2Ag$ with electrons balanced against Mg losing 2 electrons) (1)

Q5 (a) [3]

- Fe^{3+}/Fe^{2+} (+0.77 V) has a more positive E° than I_2/I^- (+0.54 V), so Fe^{3+} is reduced and I^- is oxidised (1).
- $E°_{cell} = 0.77 - 0.54$ (1)
- $= +0.23$ V, confirming the reaction is feasible (1)

Q5 (b) [3]

- Half-equations: $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (× 2) and $2I^- \rightarrow I_2 + 2e^-$ (1)
- Electrons balanced (2 electrons transferred in each half-equation) (1)
- Overall: $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$ (1)

Q5 (c) [3]

- $E°(MnO_4^-/Mn^{2+}) = +1.51$ V is more positive than $E°(Cl_2/Cl^-) = +1.36$ V, so MnO_4^- can oxidise Cl^- to Cl_2 (1).
- The reaction is feasible (1).
- $E°_{cell} = 1.51 - 1.36 = +0.15$ V (1)

Q5 (d) [3]

- Half-equations: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (× 2) and $2Cl^- \rightarrow Cl_2 + 2e^-$ (× 5) (1)
- Electrons balanced at 10 (1)
- Overall: $2MnO_4^-(aq) + 16H^+(aq) + 10Cl^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Cl_2(g)$ (1)

Q6 (a) [4]

- Pb^{2+}/Pb (−0.13 V) is more positive than Ni^{2+}/Ni (−0.25 V), so Pb^{2+} is reduced and Ni is oxidised (1).
- $E°_{cell} = -0.13 - (-0.25) = +0.12$ V (1)
- Overall equation: $Ni(s) + Pb^{2+}(aq) \rightarrow Ni^{2+}(aq) + Pb(s)$ (1)
- Reaction is feasible since $E°_{cell}$ is positive (1)

Q6 (b) [3]

- $E°(Cl_2/Cl^-) = +1.36$ V is more positive than $E°(Br_2/Br^-) = +1.09$ V, so Cl_2 is a stronger oxidising agent than Br_2 (1).
- Cl_2 can oxidise Br^- to Br_2, giving a positive $E°_{cell}$ for this direction (1).
- The reverse reaction (Br_2 oxidising Cl^-) would have a negative $E°_{cell}$ and is therefore not feasible (1).

Q6 (c) [3]

- For Ag to react, Ag would need to be oxidised (reverse of the given half-equation) while Fe^{3+} is reduced to Fe^{2+} (1).
- $E°_{cell} = E°(Fe^{3+}/Fe^{2+}) - E°(Ag^{+}/Ag) = 0.77 - 0.80 = -0.03$ V (1)
- $E°_{cell}$ is negative, so the reaction is not (thermodynamically) feasible — silver will not react with $Fe^{3+}(aq)$ (1)

Q7 (a) [2]

- $E°_{cell}$ (like ΔG) only indicates thermodynamic feasibility, not reaction rate (1).
- The reaction may have a high activation energy, so it occurs too slowly to be observed/is kinetically controlled (1).

Q7 (b) [3]

- Standard electrode potentials are only valid at standard conditions (298 K, 100 kPa, 1 mol dm^{-3}) (1).
- If concentrations, temperature or pressure differ from standard, the actual electrode potential (and hence the actual cell potential) will differ from the standard value (1).
- This could mean a reaction predicted feasible from E° values is not actually feasible under real (non-standard) conditions, or vice versa (1).

Q7 (c) [3]

- The reaction may proceed very slowly due to a high activation energy (kinetic barrier), even though it is thermodynamically feasible (1).
- As the reaction proceeds, concentrations of reactants fall (and products rise), which changes the actual electrode potentials away from standard, and could reduce or reverse the (now non-standard) cell potential (1).
- A small positive $E°_{cell}$ gives only a small thermodynamic "driving force," so such reactions are more easily halted or reversed by non-standard changes than a reaction with a large positive $E°_{cell}$ (1).

Q8 (a) [2]

- A primary cell cannot be recharged — once the reactants are used up it is "dead" (e.g. an alkaline zinc/manganese dioxide battery) (1).
- A secondary cell can be recharged by passing a current through it in reverse (e.g. a lithium-ion or lead–acid battery) (1).

Q8 (b) [3]

- During discharge, spontaneous redox reactions occur at each electrode (oxidation at the anode, reduction at the cathode), producing a current (1).
- During recharging, an external power source forces a current through the cell in the opposite direction (1).
- This reverses the electrode reactions, reforming the original reactants, so the cell can be discharged again (1).

Q8 (c) [3]

- Any three from: cheaper in the long term (not repeatedly bought/replaced); produce less solid waste, as they are reused many times rather than discarded; often deliver a more consistent voltage; more efficient/environmentally friendly overall despite the energy needed to recharge. (3 × 1, max 3)

Q8 (d) [2]

- Any valid explained disadvantage, e.g. secondary cells often have a lower energy density than some primary cells (1); or they require access to an external power supply/charger and gradually lose capacity over repeated charge/discharge cycles (1). (2 marks for one fully explained point, or 1 + 1 for two separate valid points)

Q9 (a) [2]

- $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ (2) — (1 for species/balancing, 1 for correct state symbols/electron balance)

Q9 (b) [2]

- $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ (2)

Q9 (c) [2]

- $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$ (2)

Q9 (d) [2]

- $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ (1)
- Water is the only product (1)

Q9 (e) [2]

- Advantage: the only product is water, so no CO_2 or other pollutants are released at the point of use, and fuel cells can be refuelled quickly, unlike batteries which need long recharging times (1).
- Disadvantage: hydrogen is difficult/costly to store and transport safely (requires high pressure or cryogenic storage), and hydrogen is often produced from fossil fuels (e.g. steam reforming of methane), which may release CO_2 elsewhere, offsetting the environmental benefit (1).

Q10 [6] — Level of response

- **Level 3 (5–6):** Explains clearly that E° data alone confirm the fuel cell reaction is thermodynamically feasible and give a maximum cell voltage, but discusses fully why this cannot alone judge "suitability" (rate, practical/environmental factors, hydrogen production and storage), reaching a balanced conclusion.
- **Level 2 (3–4):** Explains the feasibility/voltage role of E° data and identifies at least one relevant limitation, but the discussion is less complete or one-sided.
- **Level 1 (1–2):** Makes general statements about fuel cells or E° values with limited linkage between the two, or limited discussion of limitations.

Indicative content: *The half-equations for H_2 oxidation and O_2 reduction combine to give a large positive $E°_{cell}$, confirming the H_2/O_2 reaction is strongly thermodynamically feasible and indicating the maximum theoretical voltage obtainable from the cell — this is useful information from E° data. However, E° values say nothing about practical factors: the actual rate of reaction/power output depends on electrode/catalyst design and is not predicted by E°; E° values apply to standard 1 mol dm^{-3}/100 kPa conditions, not the real operating conditions of a vehicle fuel cell; and "suitability" as a transport fuel depends on much more than electrode potentials, including how the hydrogen fuel itself is produced (from fossil fuels, releasing CO_2, versus electrolysis using renewable electricity), how safely and compactly it can be stored/transported, refuelling infrastructure, and cost, none of which is addressed by E° data. A balanced answer should conclude that E° values are useful for confirming that the fuel cell reaction is energetically favourable and comparing cell voltages, but cannot alone be used to judge the overall practical or environmental suitability of hydrogen fuel cells for transport.*

END OF MARK SCHEME · 94 marks