

REDOX PAPER 1A (HL)

1. B
2. A
3. B
4. C
5. B

6. A
7. A
8. C
9. C
10. A

11. B
12. C
13. D
14. B
15. B

16. A
17. B
18. D
19. C
20. C

21. B
22. A
23. A
24. D
25. B

26. D
27. D
28. D
29. D
30. A

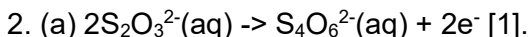
REDOX PAPER 1B (HL)

1. (a) $Q = I \times t$. $Q = 1.50 \times (40.0 \times 60) = 3600 \text{ C}$ [2].

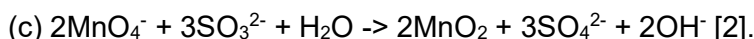
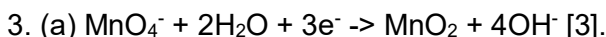
(b) Moles of $e^- = 3600 / 96500 = 0.0373 \text{ mol}$ [1]. Moles $\text{Ag} = 0.0373 \text{ mol}$ [1]. Mass = $0.0373 \times 107.87 = 4.02 \text{ g}$ [1].

(c) The gradient would be less steep (lower mass deposited per time unit) [1]. Cu^{2+} requires 2 moles of electrons for every 1 mole of copper metal, whereas Ag^+ only

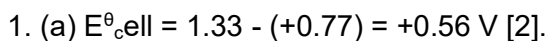
requires 1 [1]. Additionally, copper has a lighter atomic mass (63.55) than silver (107.87), significantly decreasing mass accumulation [1].



(b) Moles $\text{S}_2\text{O}_3 = 0.0100 \times 0.0245 = 2.45 \times 10^{-4}$ mol [1]. Ratio $\text{O}_2 : \text{S}_2\text{O}_3^{2-}$ is 1:4 [1].
Moles $\text{O}_2 = (2.45 \times 10^{-4}) / 4 = 6.13 \times 10^{-5}$ mol [2].

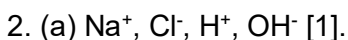


REDOX PAPER 2 (HL)

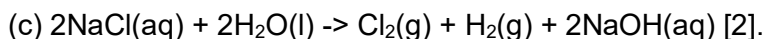


(b) $n = 6$ [1]. $\Delta G^\ominus = -nFE^\ominus = -6 \times 96500 \times 0.56$ [1]. $\Delta G^\ominus = -324240$ J = -324 kJ [1].

(c) Reaction: $2\text{Fe}^{3+} + 2\text{I}^- \rightarrow 2\text{Fe}^{2+} + \text{I}_2$. $E^\ominus_{\text{cell}} = 0.77 - 0.54 = +0.23$ V [1]. As E_{cell} is positive, the reaction is spontaneous [2].

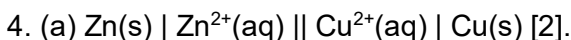


(b) Cathode: Hydrogen gas [1]. Anode: Chlorine gas [1]. E^\ominus of water reduction is less negative than Na^+ reduction (-0.83 vs -2.71 V), so H_2O preferentially discharges over Na^+ [1].



3. (a) Copper anode acts as an active electrode and oxidizes itself instead of the chloride/water ions [1]. $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ [1].

(b) Impure copper forms the anode and oxidizes to Cu^{2+} [1]. These ions migrate and preferentially reduce onto a pure copper cathode, leaving impurities behind as anode sludge [1].



(b) The standard hydrogen electrode serves as the universal reference assigned exactly 0.00 V [1]. The zinc half-cell is connected to the SHE; the voltmeter reads 0.76 V with Zn oxidizing, making its relative potential -0.76 V [2].

5. (a) $Q = 2.15 \times 9000\text{s} = 19350$ C [1]. $n(\text{e}^-) = 19350 / 96500 = 0.2005$ mol [1]. Moles Au = $0.2005 / 3 = 0.0668$ mol. Mass Au = $0.0668 \times 196.97 = 13.16$ g [1].

6. (a) PbSO_4 (Pb is +2). It oxidizes at the anode to PbO_2 (Pb +4) [1] and reduces at the cathode to Pb(s) (Pb 0) [1]. Disproportionation occurs across the battery [1].

7. (a) $3\text{Cu} + 2\text{NO}_3^- + 8\text{H}^+ \rightarrow 3\text{Cu}^{2+} + 2\text{NO} + 4\text{H}_2\text{O}$ [3].