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 x t. Q = 1.50 × (40.0 * 60) = 3600 C [2].

(b) Moles of e- = 3600 / 96500 = 0.0373 mol [1]. Moles Ag = 0.0373 mol [1]. Mass = 0.0373 × 107.87 = 4.02 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].

2. (a) $2S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) + 2e^-$ [1].

(b) Moles $S_2O_3 = 0.0100 \times 0.0245 = 2.45 \times 10^{-4}$ mol [1]. Ratio $O_2 : S_2O_3^{2-}$ is 1:4 [1]. Moles $O_2 = (2.45 \times 10^{-4}) / 4 = 6.13 \times 10^{-5}$ mol [2].

3. (a) $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ [3].

(b) $SO_3^{2-} + 2OH^- \rightarrow SO_4^{2-} + H_2O + 2e^-$ [2].

(c) $2MnO_4^- + 3SO_3^{2-} + H_2O \rightarrow 2MnO_2 + 3SO_4^{2-} + 2OH^-$ [2].

REDOX PAPER 2 (HL)

1. (a) $E^\theta_{cell} = 1.33 - (+0.77) = +0.56$ V [2].

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

(c) Reaction: $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$. $E^\theta_{cell} = 0.77 - 0.54 = +0.23$ V [1]. As E_{cell} is positive, the reaction is spontaneous [2].

2. (a) Na^+, Cl^-, H^+, OH^- [1].

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

(c) $2NaCl(aq) + 2H_2O(l) \rightarrow Cl_2(g) + H_2(g) + 2NaOH(aq)$ [2].

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

4. (a) $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$ [2].

(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 × 9000s = 19350 C [1]. n(e-) = 19350 / 96500 = 0.2005 mol [1]. Moles Au = 0.2005 / 3 = 0.0668 mol. Mass Au = 0.0668 × 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) $3Cu + 2NO_3^- + 8H^+ \rightarrow 3Cu^{2+} + 2NO + 4H_2O$ [3].