REDOX PAPER 1A

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

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

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

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

REDOX PAPER 1B

1. (a) Moles $S_2O_3^{2-}$ = 0.050 * 0.01280 = 6.4×10^{-4} mol [1].

(b) Moles O_2 in 100cm3 = $(6.4\times10^{-4}) / 4 = 1.6 \times 10^{-4}$ mol [1]. Moles in 1 dm3 = 1.6×10^{-3} mol. Mass = 1.6×10^{-3} * 32.00 = 0.0512 g [1]. Mass in mg = 51.2 mg (ppm) [1].

(c) Day 5 moles $S_2O_3^{2-}$ = 0.050 * 0.00520 = 2.6×10^{-4} mol. Day 5 O_2 = $((2.6\times10^{-4}) / 4$ * 10) * 32.00 * 1000 = 20.8 ppm [1]. BOD = Day 0 ppm - Day 5 ppm [1]. BOD = 51.2 - 20.8 = 30.4 ppm [1].

(d) Very poor quality / heavily polluted (BOD > 5 ppm) [1].

2. (a) Y > W > Z > X [2]. (1 mark for inverted order)

(b) $W(s) + 2XNO_3(aq) \rightarrow W(NO_3)_2(aq) + 2X(s)$ [2]. (Award 1 mark if charges/stoichiometry incorrect but species correct).

(c) Metal Y [1]. It was able to reduce (react with) all the other metal ions, indicating it is the most easily oxidized / gives away electrons most readily [1].

REDOX PAPER 2

1. (a) SO_2 = +4 [1]. SO_4^{2-} = +6 [1]. H_2S = -2 [1].

(b) Fe^{2+} -> Fe^{3+} + e^- [1]. $Cr_2O_7^{2-}$ + $14H^+$ + $6e^-$ -> $2Cr^{3+}$ + $7H_2O$ [1].

(c) Multiply iron equation by 6. $Cr_2O_7^{2-}$ + $14H^+$ + $6Fe^{2+}$ -> $2Cr^{3+}$ + $7H_2O$ + $6Fe^{3+}$ [2].

2. (a) Electrode A is Zinc. It must be the anode (oxidation occurs there) as it is more reactive [1]. Electrode B is Copper [1].

(b) From Electrode A (Zinc) to Electrode B (Copper) [1].

(c) Cu^{2+} + $2e^-$ -> Cu(s) [1].

(d) As Zinc is oxidized, Zn^{2+} enters Solution A (net positive charge buildup) [1]. NO_3^- ions flow into the zinc half-cell to balance this positive charge [1].

3. (a) Fe -> Fe^{2+} + $2e^-$ [1]. O_2 + $2H_2O$ + $4e^-$ -> $4OH^-$ [1].

(b) $2Fe$ + O_2 + $2H_2O$ -> $2Fe^{2+}$ + $4OH^-$ [2]. (Award 1 for correct reactants/products but incorrect balance).

(c) +3 [1].

(d) Magnesium is more reactive than iron [1] and acts as a sacrificial anode, oxidizing in preference to the iron it is attached to [1].

4. (a) Moles = 0.0200 * 0.0182 = 3.64×10^{-4} mol [1].

(b) Moles Fe^{2+} = 5 * 3.64×10^{-4} = 1.82×10^{-3} mol [1]. Concentration = 1.82×10^{-3} / 0.0250 = 0.0728 mol dm^{-3} [1].

(c) Colorless to pale pink [1].

5. (a) Colorless to yellow/orange/brown [1]. $Cl_2(aq)$ + $2Br^-(aq)$ -> $2Cl^-(aq)$ + $Br_2(aq)$ [2].

(b) No reaction [1]. Iodine is a weaker oxidizing agent than chlorine, so it cannot oxidize chloride ions [1].