**Assessment Schedule 3.3 – NZIC 2011**
**Chemistry: Describe oxidation-reduction processes (AS 90696)**

| Q | Evidence | Achievement | Achievement with Merit | Achievement with Excellence |
|---|---|---|---|---|
| **ONE**<br>(a) (i)<br><br><br><br>(ii) | Electrode A = Silver (Ag)<br>Solution A = $Ag^+$<br><br>Electrode B = Graphite (C) or Platinum (Pt)<br>Solution B = $Br_2$, $Br^-$<br><br>$E°_{cell} = E°_{RHS} - E°_{LHS} = +1.08 - (+0.80) = +0.28$ V | **Three out of:**<br><br>One half cell correct<br>OR<br>Both solutions correct<br>OR<br>Both electrodes correct | **Three of:**<br><br>Both half cells correct | |
| (b) | Anions ($NO_3^-$) from the salt bridge migrate to the LHS to neutralise the extra positive charge in the cell compartment produced as Ag atoms are oxidised to $Ag^+$ ions.<br>Cations ($K^+$) in the salt bridge migrate to the RHS to neutralise the extra negative charges as $Br^-$ ions are produced in the solution. | Correct $E°_{cell}$ value for (a)(ii) OR (c)<br><br>Identifies that ions move from salt bridge | Correct movement of anions and cations from salt bridge | Correct movement of anions and cations from salt bridge |
| (c) | Since the $Ag^+$/Ag half cell has a more positive standard reduction potential than the $Fe^{2+}$/Fe half cell, electrons will move from the $Fe^{2+}$/Fe half cell to the $Ag^+$/Ag half cell (i.e. in the opposite direction to the original cell).<br>Observations and half equations are:<br>$Fe^{2+}$/Fe half cell:<br>• Iron electrode erodes / starts to disappear / decreases in mass<br>• Solution turns pale green<br>• $Fe \rightarrow Fe^{2+} + 2e^-$<br><br>$Ag^+$ / Ag half cell:<br>• Grey deposit builds up on silver electrode / silver electrode increases in mass<br>• $Ag^+ + e^- \rightarrow Ag$<br>*As described*<br>$E°_{cell} = E°_{RHS} - E°_{LHS} = -0.44 - (0.8) = -1.24$ V<br>Accept +1.24 V | Correct observations of what occurs at anode OR cathode<br>Both half equations correctly balanced | Correct movement of electrons with explanation<br><br>Correct observations of anode AND cathode, reactions linked to species involved | AND<br>Correct discussion of all points, including half equations |

| Q | Evidence | Achievement | Achievement with Merit | Achievement with Excellence |
|---|---|---|---|---|
| **TWO** | | **Two out of three:** | **Two of:** | |
| (a) (i) | Oxidation: $2H_2O + SO_2 \rightarrow SO_4^{2-} + 4H^+ + 2e^-$<br><br>Reduction: $5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$<br><br>Overall: $2H_2O + 5SO_2 + 2MnO_4^- \rightarrow 5SO_4^{2-} + 4H^+ + Mn^{2+}$ | Both half equations correctly balanced | Correct overall equation | |
| (ii) | In strongly basic conditions:<br>$E°_{cell} = E°_{RHS} - E°_{LHS}$<br>$E°(SO_4^{2-}/SO_2) = E°_{RHS} - E°_{cell} = +0.56 - (+0.11) = +0.45$ V | Correct process for calculating $E°(SO_4^{2-}/SO_2)$ value, but may have used wrong permanganate half cell | Correct $E°(SO_4^{2-}/SO_2)$ value | |
| (b) (i) | $E°(NO_3^-/NO_2) > E°(Fe^{3+}/Fe^{2+}) > E°(Cu^{2+}/Cu)$ | | Correct sequence and strongest oxidant identified but incomplete explanation | Correct sequence and strongest oxidant identified with full explanation |
| | $NO_3^-$ can oxidise both Cu and $Fe^{2+}$, so $E°(NO_3^-/NO_2)$ will be greater than both $E°(Fe^{3+}/Fe^{2+})$ and $E°(Cu^{2+}/Cu)$.<br>$Cu^{2+}$ cannot oxidise $Fe^{2+}$, so $E°(Fe^{3+}/Fe^{2+})$ will be greater than $E°(Cu^{2+}/Cu)$. | EITHER strongest oxidant recognised with incomplete explanation OR correct sequence with incomplete explanation | | |
| (ii) | $NO_3^-$ can oxidise both Cu and $Fe^{2+}$, so it is the strongest oxidant. | | | |

<table>
<tr><th>Q</th><th>Evidence</th><th>Achievement</th><th>Achievement with Merit</th><th>Achievement with Excellence</th></tr>
<tr><td>THREE</td><td></td><td>Three out of:</td><td>Two out of:</td><td></td></tr>
<tr><td>(a) (i)</td><td>Reaction 1:<br>Reduction occurs at the cathode</td><td>Reaction 1 circled with reason</td><td></td><td></td></tr>
<tr><td>(ii)</td><td>$Zn \,|\, Zn^{2+} \,\|\, MnO_2, Mn_2O_3 \,|\, C$<br>| in place of , OK</td><td>Correct cell diagram</td><td></td><td></td></tr>
<tr><td>(b) (i)</td><td>The oxidation number of the Cu in $Cu_2O$ decreases from +1 to 0 in Cu, so $Cu_2O$ is reduced. However, the oxidation number of the Cu in $Cu_2O$ increases from +1 to +2 in $CuSO_4$, so $Cu_2O$ is also oxidised.</td><td>Correctly assigns oxidation numbers to Cu species in reactants and products</td><td>Explains redox reaction using correct oxidation numbers</td><td>Explains redox reaction using correct oxidation numbers<br><br>AND</td></tr>
<tr><td>(ii)</td><td>Copper (I) ions can reduce and oxidise each other in solution according to the following balanced equation:<br>$2Cu^{+} \rightarrow Cu + Cu^{2+}$<br><br>OR:<br>$Cu^{+} + e^{-} \rightarrow Cu$<br>$Cu^{+} \rightarrow Cu^{2+} + e^{-}$<br><br>The cell potential for this reaction is:<br>$E°_{cell} = E°_{RHS} - E°_{LHS} = +0.52 - (+0.16) = +0.36$ V<br>Since the $E°_{cell}$ is positive, the reaction will be spontaneous, so copper (I) ions cannot exist in solution.</td><td>Correct half equations<br>OR<br>Balanced equation<br><br>Correct $E°_{cell}$ value</td><td>Recognises disproportionation reaction occurring in solution<br><br>Calculates $E°_{cell}$ and relates this to spontaneity of reaction</td><td>Full explanation of why copper (I) ions cannot exist in solution including balanced equation</td></tr>
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**Judgement Statement**

| Achieved | Merit | Excellence |
|---|---|---|
| Total of TWO opportunities answered at Achievement level or higher.<br><br>**2 x A** | Total of THREE opportunities answered with TWO at Merit level or higher.<br><br>**2 x M** | Total of THREE opportunities answered with TWO at Excellence level or higher.<br><br>**2 x E** |