3.3 Oxidation and reduction 2006 Answers

QUESTION ONE: REACTIONS IN A CELL

a

The oxidation number of Cr in $Cr_2O_7^{2-}$ is +6, while in Cr^{3+} it is +3.
Since the oxidation number of Cr has decreased in this reaction, the process is reduction.

A = correct oxidation numbers for each species AND statement that a decrease in oxidation number means reduction.

b i

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^- \quad \text{OR} \quad 3Cu(s) \longrightarrow 3Cu^{2+}(aq) + 6e^-$$

ii

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 3Cu(s) \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3Cu^{2+}(aq)$$

A = Two half-equations correctly balanced.
M = All three equations correct.

c In the left-hand cell the brown copper metal would slowly dissolve, and the colourless liquid would turn pale blue as the Cu(s) is slowly oxidised to Cu^{2+}(aq).
In the right-hand cell the orange solution would slowly turn green/blue as the $Cr_2O_7^{2-}$(aq) is reduced to Cr^{3+}.

A = Identifies one change occurring in each cell by both stating the observation and stating the species involved.

QUESTION TWO: REACTIONS OF MANGANESE SPECIES

a Combining the two equations to calculate the E_{cell} with MnO_4^- to be reduced, Mn^{2+} to be oxidised:

$$E_{cell} = E_{red} - E_{ox}$$
$$= 0.85 \text{ V} - (+1.49 \text{ V})$$
$$= -0.64 \text{ V}$$

Since this E_{cell} is negative, the reaction is not spontaneous.

A = Identifies that a negative E_{cell} indicates that the reaction is not spontaneous OR recognises that $E°(MnO_4^-/Mn^{3+})$ is less than $E°(Mn^{3+}/Mn^{2+})$.
M = Calculates E_{cell} as –0.64 V and links this value to the fact that the spontaneous reaction does not occur OR an explanation that correctly links reduction potentials and the strength of oxidant and reductant so that no reaction occurs.

b i In strongly acidic conditions the purple MnO_4^-(aq) will be reduced to colourless Mn^{2+}(aq), while the colourless SO_3^{2-}(aq) is oxidised to colourless SO_4^{2-}. You will see the purple solution turn colourless as it is added to the colourless solution in the beaker.

In neutral or weakly basic solution the purple MnO_4^-(aq) is reduced to the brown precipitate MnO_2(s), while the colourless SO_3^{2-}(aq) is oxidised to colourless SO_4^{2-} as before. You will see the purple solution form a brown precipitate.

In strongly basic conditions the purple MnO_4^-(aq) is reduced to dark green MnO_4^{2-}(aq) while the colourless SO_3^{2-}(aq) is oxidised to colourless SO_4^{2-} as before. You will see the purple solution turn dark green.

ii $MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$

$SO_3^{2-}(aq) + H_2O \longrightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

$2MnO_4^-(aq) + 6H^+(aq) + 5SO_3^{2-}(aq) \longrightarrow 2Mn^{2+}(aq) + 3H_2O(l) + 5SO_4^{2-}(aq)$

A = At least two correct observations linked to species involved
OR Correct observations for all three reactions of MnO_4^-
OR All four species correctly linked to their colours
OR Redox reactants and products identified: SO_3^{2-} and MnO_4^-, Mn^{2+} and SO_4^{2-}.
M = All observations correct and linked to species involved
OR Fully balanced equation correct.
E = All correct observations with links to species and correctly balanced equation.

QUESTION THREE: CONSTRUCTING A CELL

a Current cannot flow through the wire because the U-tube is filled with methanol – a non-electrolyte. With no method for ions to be transferred from one beaker to the other, electrons cannot move from one beaker to the other. Fill the U-tube with an electrolyte such as sulfuric acid solution or potassium nitrate solution to allow ions to be exchanged between the beakers.

b In this set-up the solutions are in the wrong beakers. No reaction will occur in the left-hand beaker, while in the right hand beaker electrons will be transferred from the Zn(s) to the Cu^{2+}(aq) and there will be no flow of electrons between the beakers. Swap the beakers over so that the copper metal is in the $Cu(NO_3)_2$ solution and the zinc metal is in the $Zn(NO_3)_2$ solution. This will cause electrons to flow through the wire and ions to be transferred through the salt bridge.

A = Correctly explains and fixes the problem in one cell.
M = Correctly explains and fixes the problems with both cells.

QUESTION FOUR: A CELL IN ACTION

a i Electron flow is from left to right (opposite direction to flow of anions).
Left-hand electrode is negative, right-hand electrode is positive (since electrons flow from negative to positive).

ii $E_{cell} = E_{RHE} - E_{LHE}$
The right hand electrode is the standard hydrogen electrode, which, by definition, has an $E°$ of 0.00 V.
$E_{LHE} = 0.00 \text{ V} - E_{cell}$
$= 0.00 \text{ V} - 0.74 \text{ V}$
So $E(X^{3+}/X) = -0.74$ V.

Oxidation is occurring in the left hand cell, meaning that X is being oxidised to X^{3+}, and that H^+ is being reduced to H_2 in the right hand cell. We would expect a positive value for $E°(X^{3+}/X)$ if it was reduced with respect to the hydrogen half-cell, but since it is being oxidised, the negative value is correct.

A = Direction of electron flow consistent with sign on electrodes (ie correct, or both electron flow and sign on electrodes reversed) OR $E(X^{3+}/X) = -0.74$ V with no working OR recognition that $E°_{cell}$ is difference between $E(H^+/H_2) = 0.00$ V and $E(X^{3+}/X)$ to give a value of 0.74 V.
M = $E°(X^{3+}/X) = -0.74$ V is calculated or explained correctly by comparison with $E°(H^+/H_2) = 0.00$ V AND either direction of e^- flow or labelling of electrodes is correct. OR Direction of electron flow and signs of electrodes both reversed AND comparison with $E°(H^+/H_2) = 0.00$ V gives $E°(X^{3+}/X) = +0.74$ V. OR Direction of electron flow AND reduction potential correct, but no or limited explanation/calculation.
E = All correct with valid explanation / calculation for electrode potential.

b $Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^-$

$X^{3+} + 3e^- \longrightarrow X(s)$

$$n(\text{Mn}) = \frac{m}{M} = \frac{200 \text{ g}}{54.9 \text{ g mol}^{-1}} = 3.643 \text{ mol}$$

Mn releases 2 electrons per mole

$$n(e^-) = 2 \times n(\text{Mn}) = 2 \times 3.643 \text{ mol} = 7.286 \text{ mol}$$

Each mol of X formed requires 3 moles of electrons.

$$n(\text{X}) = \frac{n(e^-)}{3} = \frac{7.286 \text{ mol}}{3} = 2.43 \text{ mol}$$

A = Method correct but mathematical error. This includes correct balanced equation but use inverted ratio to get n(X) =5.46 mol. OR incorrectly balanced eqn but application of ratio in that eqn correct.
M = Answer correct.

QUESTION FIVE: METALS AS REDUCTANTS

a Strongest Zn Ga Fe

b Zn reduces both Ga^{3+} and Fe^{2+}, so Zn is a stronger reductant than Ga or Fe; Ga reduces Fe^{2+}, so Ga is a stronger reductant than Fe.
The strongest reductant will be the cell with the most negative *E*°.
Thus $E°(Fe^{2+}/Fe) > E°(Ga^{3+}/Ga) > E°(Zn^{2+}/Zn)$

A = Order of metals as reductants (in part a) correct without valid justification. OR In part b identifies order of $E°(Fe^{2+}/Fe) > E°(Ga^{3+}/Ga) > E°(Zn^{2+}/Zn)$
M = Order of metals as reductants correct AND linked to information in table in terms of species that react and don't react. OR Links info in table to order of $E°(Fe^{2+}/Fe) > E°(Ga^{3+}/Ga) > E°(Zn^{2+}/Zn)$ but lists reductants as ions (ie order of oxidants $Fe^{2+} > Ga^{3+} > Zn^{2+}$.
E = Answer shows correct order of metals as reductants AND discussion clearly relates the order to the relevant *E*° values. eg best reductant will have most negative *E*° value and will be metal reacting with both of the other metal ions.

Judgement Statement

Achievement
FIVE questions answered correctly.
Minimum of 5 × A

Achievement with Merit
SIX questions answered correctly, including at least FOUR at Merit level.
Minimum of 4 × M + 2 × A

Achievement with Excellence
SEVEN questions answered correctly, including at least THREE at Merit level and at least TWO at Excellence level.
Minimum of 2 × E + 3 × M + 2 × A