3.3 Oxidation and reduction 2005 Answers

QUESTION ONE: COBALT AND CHROMIUM IN REDOX REACTIONS

a $Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$

$3Co \longrightarrow 3Co^{2+} + 6e^-$

$Cr_2O_7^{2-} + 14H^+ + 3Co \longrightarrow 2Cr^{3+} + 7H_2O + 3Co^{2+}$

A = Correct equation.

b $Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$ $E° = + 1.36$ V

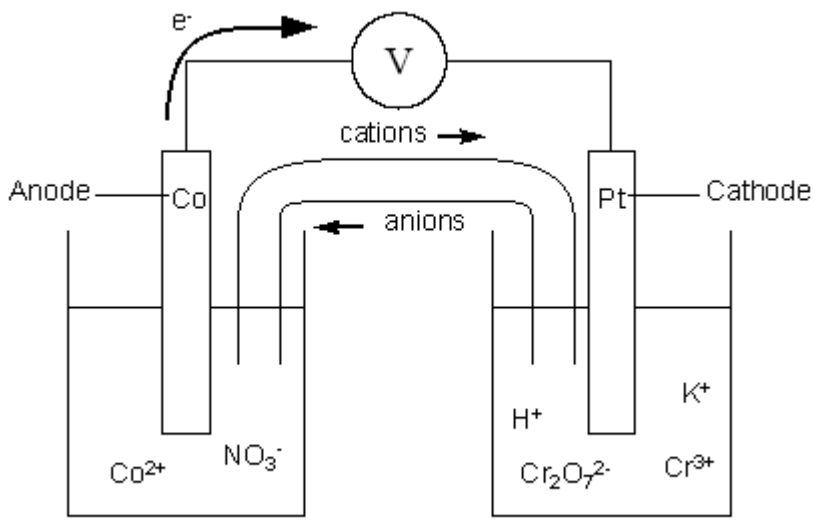
$3Co \longrightarrow 3Co^{2+} + 6e^-$ $E° = + 0.28$ V (reversed because Ox)

Therefore $E_{cell} = +1.36$ V + 0.28 V = + 1.64 V

OR $E_{cell} = E_{red} - E_{ox}$
$= 1.36$ V – (–0.28 V)
= +1.64 V

A = Correct answer

c



Cell drawing (beakers may be swapped over)
A = Salt bridge shown OR one half-cell correct (Co electrode with Co^{2+} or $Cr_2O_7^{2-}$ and Cr^{3+} with Pt or C electrode).
M = Salt bridge shown AND both half-cells correct (spectator ions optional).

Electron and ion flow:
A = Electron flow from Co to Pt OR anion flow toward Co OR cation flow toward Pt.
M= Electron flow AND both ion flow correct.

d $Co \mid Co^{2+} \parallel Cr_2O_7^{2-}, Cr^{3+} \mid Pt$

A = correct (any inert electrode may be used and state symbols may be included).

e We know, from **a** and **b** above, that $Cr_2O_7^{2-}$ will oxidise Co to Cr^{2+}, so the question is, will $Cr_2O_7^{2-}$ oxidise Cr^{2+} to Cr^{3+}?

$E_{cell} = E_{red} - E_{ox}$
$= +1.36\ V - (+1.82\ V)$
$= -0.46\ V$

Since the E_{cell} is negative, this reaction does not occur, therefore when Co is reacted with $Cr_2O_7^{2-}$ it is oxidised to Cr^{2+}, not Cr^{3+}.

A = correct answer (Cr^{2+} formed, not Cr^{3+}).
M = correct answer plus partial explanation that includes a relevant calculation.
E = correct answer plus full explanation including calculation and recognition of spontaneous formation of Co^{2+}.

f Method one

Reaction 1
Co^{3+} is reduced by U^{3+}. Co^{3+} has a stronger reduction potential than U^{4+}

Reaction 2
U^{4+} is not reduced by Fe^{2+}. U^{4+} has a weaker reduction potential than Fe^{3+}

Reaction 3
Co^{3+} is reduced by Fe^{2+}. Co^{3+} has a stronger reduction potential than Fe^{3+}.

Strongest reduction potential (Co^{3+}/Co^{2+}) then (Fe^{3+}/Fe^{2+}) and weakest is (U^{4+}/U^{3+})

The ion that is the strongest oxidant is Co^{3+}.

Method 2
Reaction 1 is spontaneous, so $E_{red} > E_{ox,}$ which means $E(Co^{3+}/Co^{2+}) > E(U^{4+}/U^{3+})$
Reaction 2 is not spontaneous, so $E_{red} < E_{ox}$ which means $E(U^{4+}/U^{3+}) < E(Fe^{3+}/Fe^{2+})$
Reaction 3 is spontaneous, so $E_{red} > E_{ox,}$ which means $E(Co^{3+}/Co^{2+}) > E(Fe^{3+}/Fe^{2+})$

Thus the order is $E(Co^{3+}/Co^{2+}) > E(Fe^{3+}/Fe^{2+}) > E(U^{4+}/U^{3+})$
and Co^{3+} is the strongest oxidant.

A = Correct order OR identifies strongest oxidant.
M = Correct order AND strongest oxidant but incomplete explanation.
E = Correct order AND strongest oxidant AND full explanation.

QUESTION TWO: CAR BATTERIES

a i Pb = 0 PbO_2 = +4 $PbSO_4$ = +2

ii The oxidant is the species that has been reduced – ie its oxidation number has decreased. The oxidant in this reaction is PbO_4 because the oxidation number of the Pb has decreased from +4 to +2.
The reductant is the species that has been oxidised – ie its oxidation number has increased. The reductant in this reaction is Pb because its oxidation number has increased from 0 to +2.

A = States the relationship between oxidant/reductant and oxidation number OR identifies the oxidant and reductant in the reactions OR identifies one of oxidant or reductant with oxidation number change.

M = Complete answer.

b Anode: $Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-$

Cathode $PbO_2(s) + 4H^+(aq) + 2e^- \longrightarrow Pb^{2+}(aq) + 2H_2O(l)$

(Equations may be written using $PbSO_4$, H_2SO_4 and SO_4^{2-})

A = One equation correct at correct electrode OR both equations correct but at wrong electrodes.
M = Both equations correct and at correct electrodes.

QUESTION THREE: ANALYSIS OF COPPER IN BRASS

a The yellow/brown screw is placed in colourless nitric acid. A brown gas is produced, the liquid turns green/blue, heat is produced and the metal slowly disappears. When all the gas has evolved the liquid turns blue.

A = At least two correct observations including at least one colour change.

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

$NO_3^-(aq) + 2H^+ + e^- \longrightarrow NO_2(g) + H_2O$

OR $2NO_3^-(aq) + 4H^+ + 2e^- \longrightarrow 2NO_2(g) + 2H_2O$

A = One half-equation correct (accept use of HNO_3 in place of NO_3^-).
M = Both half-equations correct.

c Step 2 In this reaction $I^-(aq)$ reacts with Cu^{2+}. The I^- is oxidised to I_2 while the Cu^{2+} is reduced to Cu^+.

$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$

$2Cu^{2+}(aq) + 2e^- \longrightarrow 2Cu^+(aq)$

The Cu^+ combines with I^- to form the white solid CuI(s).

Overall equation:

$2Cu^{2+}(aq) + 4I^-(aq) \longrightarrow 2CuI(s) + I_2(aq)$

Step 3 In this reaction yellow-brown $I_2(aq)$ is reduced by colourless $S_2O_3^{2-}(aq)$ to form colourless $I^-(aq)$. The $S_4O_6^{2-}(aq)$ also formed is colourless too.

$I_2(aq) + 2e^- \longrightarrow 2I^-(aq)$

$2S_2O_3^{2-}(aq) \longrightarrow S_4O_6^{2-}(aq) + 2e^-$

Overall equation

$I_2(aq) + 22S_2O_3^{2-}(aq) \longrightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

A = One observation from either step 2 or 3 linked to a species involved (name or formula).
M = Two observations linked to the appropriate half-equations.
E = One redox reaction (step 2 or step 3) completely identified with observations explained, the appropriate half-equations and a full balanced (ionic or full) equation written.

Judgement Statement
Achievement
SEVEN opportunities answered at Achievement level or higher.
7 × A

Achievement with Merit
EIGHT opportunities answered with at least FOUR at Merit level or higher.
4 × M *plus* 4 × A

Achievement with Excellence
NINE opportunities answered with at least TWO at Excellence level and FOUR at Merit level.
2 × E *plus* 4 × M *plus* 3 × A