Unit B: Electrochemical Changes – Solutions

Question 1

Assign oxidation numbers and identify the reducing agent (RA) and oxidizing agent (OA):

$$2\,\underset{RA}{\overset{0}{Na}}(g) + \underset{OA}{\overset{0}{Cl_2}}(g) \rightarrow 2\,\overset{+1}{Na}\,\overset{-1}{Cl}\,(g) + energy$$

It is the OA that undergoes reduction per this half-reaction:

D) $Reduction\!: \underset{OA}{Cl_2}(g) \rightarrow 2\,Cl^-(g) + 2\,e^-$

Question 2

Option B is incorrect. Oxidation is the process of losing electrons.

Option D is incorrect. Oxidation has nothing to do with the loss or gain of protons.

The half-reaction for the oxidation of zinc is as follows:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$$

Zn(s) loses electrons to form Zn^{2+} ions.

Option C is correct!

Question 3

$$\overset{+2}{Cu}\;\overset{-2}{O}\;(s) + 2\,\overset{+1}{H^+}(aq) \rightarrow \overset{+2}{Cu^{2+}}(aq) + \overset{+1}{H_2}\;\overset{-2}{O}\;(l)$$

By assigning oxidation numbers for all entities, there is **no change** in the oxidation number of copper.

Option D is correct! Neither oxidation nor reduction occurs.

Question 4

The operational definition of reduction is the process of producing metals from their compounds.

Options A, B, and C are incorrect. These options speak of metals or a compound reacting with oxygen. Oxygen is an oxidizing agent which will undergo reduction itself.

Option D is correct! Zinc metal is produced from a zinc compound (zinc sulfide ore).

Question 5

Options A and C are incorrect. They are acid-base reactions which are not redox reactions.

Option D is incorrect. It is a double replacement reaction which is not a redox reaction.

Option B is correct! It is a single replacement reaction which IS a redox reaction.

Question 6

$$\underset{RA}{Mg(s)} + 2\,\underset{OA}{H_2O(l)} \rightarrow Mg(OH)_2(s) + H_2(g)$$

Metal solids are ALWAYS reducing agents. The **oxidizing agent** in this reaction is **$H_2O(l)$ (i)**.

The half-reaction for the oxidation of Mg(s) is as follows:

$$Mg(s) \rightarrow Mg^{2+}(s) + 2\,e^-$$

2 moles of electrons are transferred when powdered magnesium (Mg(s)) reacts **(ii)**.

Option D is correct! $i = H_2O(l), ii = 2\,mol$

Question 7

The correct options should show a **pure metal solid on the reactants side**. Metal solids are always reducing agents which get oxidized.

The only options which do so are:

$$I)\ 2\,Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

$$II)\ 2\,Al(s) + 3\,Cl_2(g) \rightarrow 2AlCl_3(s)$$

Option A is correct! I and II only

Question 8

I) There is a change in the oxidation number of copper going from reactants to products:

$$\overset{0}{Cu}(s) + 4\,HNO_3(aq) \rightarrow \overset{+2}{Cu}(NO_3)_2(aq) + 2\,NO_2(g) + 2\,H_2O(l)\ \ \textbf{Redox}!$$

II) This is a double replacement reaction that is NOT a redox reaction.

$$Cu(NO_3)_2(aq) + 2\,NaOH(aq) \rightarrow Cu(OH)_2(s) + 2\,NaNO_3(aq)$$

III) There is no change in oxidation numbers from reactants to products.

$$\overset{+2}{Cu}\left(\overset{-2}{O}\ \overset{+1}{H}\right)_2(s) \rightarrow \overset{+2}{Cu}\ \overset{-2}{O}\ (s) + \overset{+1}{H_2}\ \overset{-2}{O}\ (l)$$

IV) This is an acid-base reaction which is NOT a redox reaction.

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

V) This is a single replacement reaction which IS a redox reaction!

$$\overset{+2}{Cu}SO_4(aq) + \overset{0}{Zn}(s) \rightarrow \overset{+2}{Zn}SO_4(aq) + \overset{0}{Cu}(s)\ \ \textbf{Redox}!$$

Option B is correct! I and V only

Question 9

The correct options should show a **pure metal solid on the reactants side**. **Metal solids** are always reducing agents which get oxidized and in doing so, **lose electrons**.

The only options which satisfy this are:

$$III)\ Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

$$IV)\ Cu(s) + 4\,HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\,NO_2(g) + 2\,H_2O(l)$$

Option D is correct! III and IV

Question 10

The goal is to build a redox table with oxidizing agents in the left column and reducing agents in the right column.

The strongest oxidizing agent is highest on the left and the strongest reducing agent is lowest on the right.

Each spontaneous reaction should have its oxidizing agent higher up in the table than its reducing agent.

One can easily eliminate options C and D because metal solids are NEVER oxidizing agents.

Start by labelling each reactant as an OA or RA:

$$(\mathbf{1})\ \underset{OA}{S^{2+}(aq)} + 2\,\underset{RA}{T(s)} \rightarrow 2\,T^{+}(aq) + S(s)$$

$$(\mathbf{2})\ \underset{OA}{R^{3+}(aq)} + \underset{RA}{T(s)} \rightarrow no\ reaction$$

$$(\mathbf{3})\ 2\,\underset{OA}{R^{3+}(aq)} + 3\,\underset{RA}{V(s)} \rightarrow 3\,V^{2+}(aq) + 2\,R(s)$$

Place oxidizing agents and reducing agents in a table (in **bold**), orienting them based on the spontaneity of their reactions.

	Oxidizing Agents	**Reducing Agents**	
Strongest OA	**S^{2+}(aq) (1)**		Weakest RA
		T(s) (1)	
	R^{3+}(aq) (2)		
Weakest OA		**V(s) (3)**	Strongest OA

Finish by filling in the blanks with each OA's or RA's conjugate (in *italics*) as seen on the products side of the reactions:

	Oxidizing Agents	Reducing Agents	
Strongest OA	**S^{2+}(aq)**	*S(s)*	Weakest RA
	T^{+}(aq)	**T(s)**	
	R^{3+}(aq)	*R(s)*	
Weakest OA	*V^{2+}(aq)*	**V(s)**	Strongest OA

Option A is correct! Listed from strongest to weakest, the correct OA order is: S^{2+}(aq), T^{+}(aq), R^{3+}(aq), V^{2+}(aq).

Question 11

Start by labelling each reactant as an OA or RA. Pay close attention to how the oxidation numbers are changing as each reactant is a metal ion, which could be either an OA or an RA.

$$(\mathbf{1})\ \underset{OA}{Ce^{4+}(aq)} + \underset{RA}{Cu^{+}(aq)} \rightarrow Ce^{3+}(aq) + Cu^{2+}(s)\ \ spontaneous$$

$$(\mathbf{2})\ \underset{RA}{Co^{2+}(aq)} + \underset{OA}{Ce^{4+}(aq)} \rightarrow Co^{3+}(aq) + Ce^{3+}(aq)\ \ nonspontaneous$$

$$(\mathbf{3})\ 2\ \underset{OA}{Cu^{2+}(aq)} + \underset{RA}{Mn(s)} \rightarrow 2\ Cu^{+}(aq) + Mn^{2+}(aq)\ \ spontaneous$$

Place oxidizing agents and reducing agents in a table (in **bold**), orienting them based on the spontaneity of their reactions.

	Oxidizing Agents	Reducing Agents	
Strongest OA		**Co^{2+}(aq) (2)**	Weakest RA
	Ce^{4+}(aq) (1)		
	Cu^{2+}(aq) (3)	**Cu^{+}(aq) (1)**	
Weakest OA		**Mn(s) (3)**	Strongest OA

Finish by filling in the blanks with each OA's or RA's conjugate (in *italics*) as seen on the products side of the reactions:

	Oxidizing Agents	Reducing Agents	
Strongest OA	*Co^{3+}(aq)*	**Co^{2+}(aq)**	Weakest RA
	Ce^{4+}(aq)	*Ce^{3+}(aq)*	
	Cu^{2+}(aq)	**Cu^{+}(aq)**	
Weakest OA	*Mn^{2+}(aq)*	**Mn(s)**	Strongest OA

Option C is correct! Listed from strongest to weakest, the correct OA order is: Co^{3+}(aq), Ce^{4+}(aq), Cu^{2+}(aq), Mn^{2+}(aq).

Question 12

It is important to first identify which entities are oxidizing agents. Remember that metal solids are always reducing agents. The oxidizing agents are highlighted in the table below in grey.

	Ga(s)	Fe(s)	Zn(s)	Mg(s)
1 Ga^{3+}(aq)	✗	✗	✓	✓
2 Fe^{2+}(aq)	✓	✗	✓	✓
3 Zn^{2+}(aq)	✗	✗	✗	✓
4 Mg^{2+}(aq)	✗	✗	✗	✗

Ordering them from strongest to weakest can be easily done by looking at the number of spontaneous reactions each OA has:

(2) Fe^{2+}(aq) (3 reactions), **(1)** Ga^{3+}(aq) (2 reactions), **(3)** Zn^{2+}(aq) (1 reaction), **(4)** Mg^{2+}(aq) (0 reactions)

The correct answer is **2134**.

Question 13

In the table of standard electrode potentials in the data booklet, oxidizing agents are on the left and reducing agents are on the right. The oxidizing agents that will oxidize the reducing agent, Cr^{2+}(aq), are higher than it on the table on the left. Of the options listed, and shown in the table below, they are, F_2(g) (**2**), AgI(s) (**3**), Ni^{2+}(aq) (**6**), and O_2(g) + H_2O(l) (**7**).

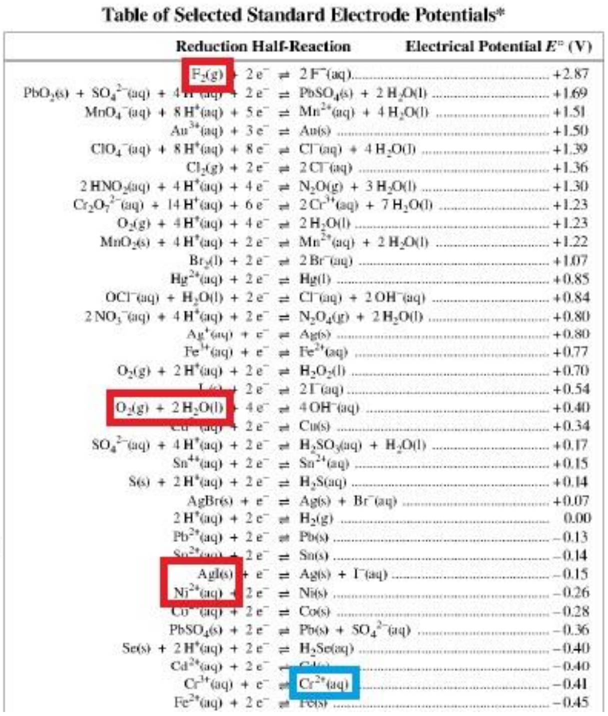


Table of Selected Standard Electrode Potentials*

Reduction Half-Reaction	Electrical Potential $E°$ (V)
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$	+1.69
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightleftharpoons Cl^-(aq) + 4H_2O(l)$	+1.39
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$2HNO_2(aq) + 4H^+(aq) + 4e^- \rightleftharpoons N_2O(g) + 3H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.22
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.07
$Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$	+0.85
$OCl^-(aq) + H_2O(l) + 2e^- \rightleftharpoons Cl^-(aq) + 2OH^-(aq)$	+0.84
$2NO_3^-(aq) + 4H^+(aq) + 2e^- \rightleftharpoons N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(l)$	+0.70
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45

The correct answer is **2367**.

Question 17

This question is tricky because reduction half-reactions given in the data booklet's table of standard reduction potentials must be considered as well as the reduction half-reactions shown.

Based on the order of electrical potentials, the booklet half-reactions (in **bold**) can be inserted as follows:

Reduction Half-Reaction	**Electrical Potential (E^{o})**
$\boxed{Am^{4+}(aq)} + e^{-} \rightleftharpoons \boxed{Am^{3+}(aq)}$	+2.60 V
$\boldsymbol{MnO_4^{-}(aq) + 8\,H^{+}(aq) + 5\,e^{-} \rightleftharpoons Mn^{2+}(aq) + 4\,H_2O(l)}$	**+1.51 V**
$\boxed{\boldsymbol{Cl_2(g)}} \boldsymbol{+ 2\,e^{-} \rightleftharpoons 2\,Cl^{-}(aq)}$	**+1.36 V**
$Tl^{3+}(aq) + 2\,e^{-} \rightleftharpoons \boxed{Tl^{+}(aq)}$	+1.25 V
$Ac^{3+}(aq) + 3\,e^{-} \rightleftharpoons Ac(s)$	-2.20 V
$\boldsymbol{Ca^{2+}(aq) + 2e^{-} \rightleftharpoons Ca(s)}$	**-2.87 V**
$Cs^{+}(aq) + e^{-} \rightleftharpoons \boxed{Cs(s)}$	-3.03 V

Oxidizing agents are to the left of the equilibrium arrows. Reducing agents are to the right of the equilibrium arrows.

The strongest oxidizing agent is highest on the left. The strongest reducing agent is lowest on the right.

For a spontaneous reaction, the OA and the RA must be in this arrangement:

$$\begin{array}{ccc} OA & & \\ & \rightleftharpoons & \\ & & RA \end{array}$$

For a non-spontaneous reaction, the OA and the RA must be in this arrangement:

$$\begin{array}{ccc} & & RA \\ & \rightleftharpoons & \\ OA & & \end{array}$$

Am^{4+}(aq) (**1**) will react spontaneously with Tl^{+}(aq).

Am^{3+}(aq) (**5**) will not react spontaneously with Cl_2(g).

Cs(s) (**8**) is a stronger reducing agent than Ca(s).

Am^{4+}(aq) (**1**) is a stronger oxidizing agent than MnO_4^{-}(aq) + 8 H^{+}(aq).

The correct answer is **1581**.

Question 21

In these sorts of questions all chemical species must be considered and whether they act as oxidizing or reducing agents.

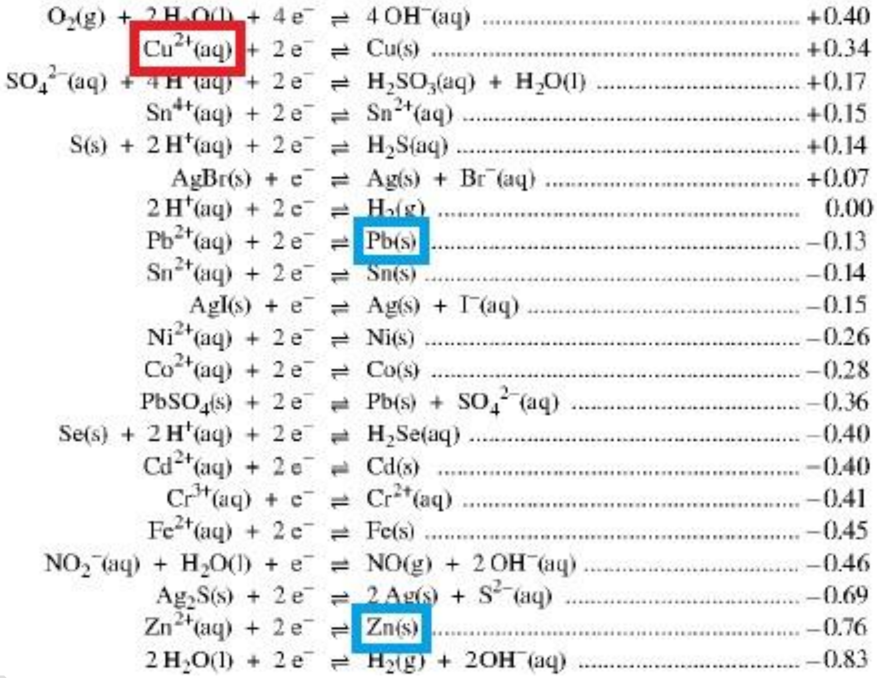
The metal solids in the test tubes, Zn(s), Pb(s), Cu(s), Au(s), and Ag(s) are all reducing agents.

H_2O(l) can be either an oxidizing agent or a reducing agent, but it is not the strongest oxidizing agent and based on its placement in the standard electrode potentials table, it is a weaker reducing agent than any of the metal solids.

NO_3^-(aq) can only be an oxidizing agent if it works with H^+(aq) so it need not be considered.

The strongest oxidizing agent present is Cu^{2+}(aq). It is highest on the left.

The only metals that will react are the ones that sit lower than Cu^{2+}(aq) in the table and they are Zn(s) (**I**) and Pb(s) (**II**).



$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83

The correct answer is A, I and II only.

Question 22

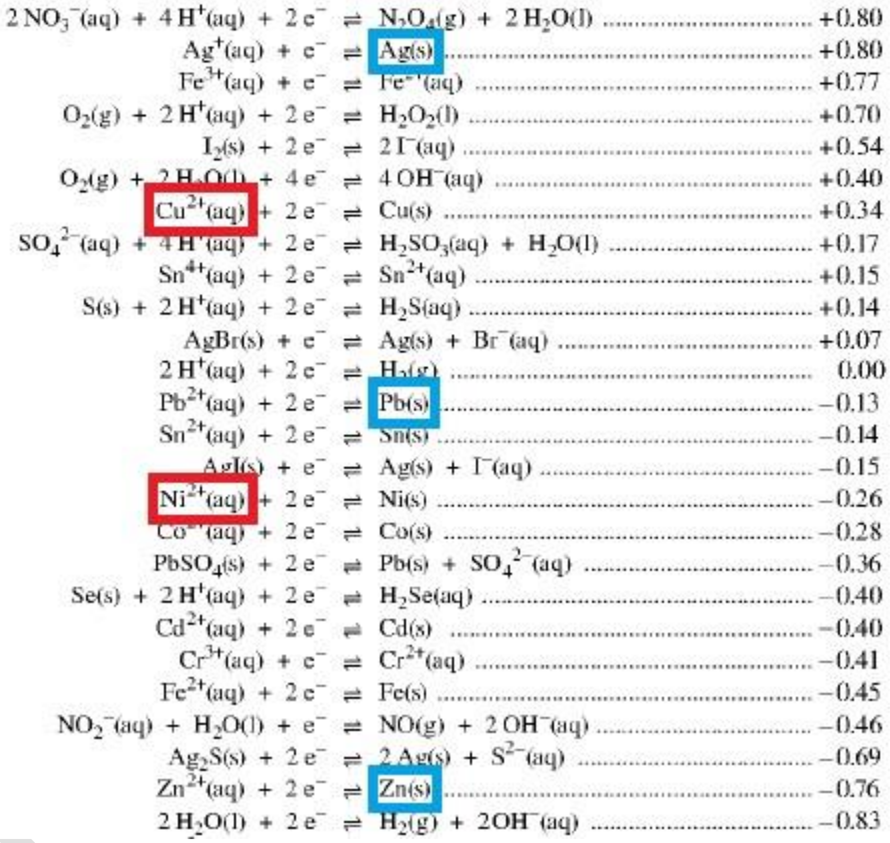
In these sorts of questions all chemical species must be considered and whether they act as oxidizing or reducing agents.

The metal solids, X(s), Y(s), and Z(s) are all reducing agents.

$H_2O(l)$ can be either an oxidizing agent or a reducing agent, but it is not the strongest oxidizing agent and based on its placement in the standard electrode potentials table, it is a weaker reducing agent than any of the metal solids.

$NO_3^-(aq)$ can only be an oxidizing agent if it works with $H^+(aq)$ so it need not be considered.

The strongest oxidizing agents present are $Cu^{2+}(aq)$ and $Ni^{2+}(aq)$. They are highest on the left.



$2NO_3^-(aq) + 4H^+(aq) + 2e^-$	$\rightleftharpoons$	$N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^-$	$\rightleftharpoons$	$\boxed{Ag(s)}$	+0.80
$Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2O_2(l)$	+0.70
$I_2(s) + 2e^-$	$\rightleftharpoons$	$2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^-$	$\rightleftharpoons$	$4OH^-(aq)$	+0.40
$\boxed{Cu^{2+}(aq)} + 2e^-$	$\rightleftharpoons$	$Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^-$	$\rightleftharpoons$	$Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2S(aq)$	+0.14
$AgBr(s) + e^-$	$\rightleftharpoons$	$Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$\boxed{Pb(s)}$	−0.13
$Sn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Sn(s)$	−0.14
$AgI(s) + e^-$	$\rightleftharpoons$	$Ag(s) + I^-(aq)$	−0.15
$\boxed{Ni^{2+}(aq)} + 2e^-$	$\rightleftharpoons$	$Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Co(s)$	−0.28
$PbSO_4(s) + 2e^-$	$\rightleftharpoons$	$Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Cd(s)$	−0.40
$Cr^{3+}(aq) + e^-$	$\rightleftharpoons$	$Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^-$	$\rightleftharpoons$	$NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^-$	$\rightleftharpoons$	$2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$\boxed{Zn(s)}$	−0.76
$2H_2O(l) + 2e^-$	$\rightleftharpoons$	$H_2(g) + 2OH^-(aq)$	−0.83

Based on table placement, both Pb(s) and Zn(s) will react with $Cu^{2+}(aq)$. Pb(s) will not react with $Ni^{2+}(aq)$, but Zn(s) WILL react with $Ni^{2+}(aq)$. Ag(s) will not react with either $Ni^{2+}(aq)$ or $Cu^{2+}(aq)$.

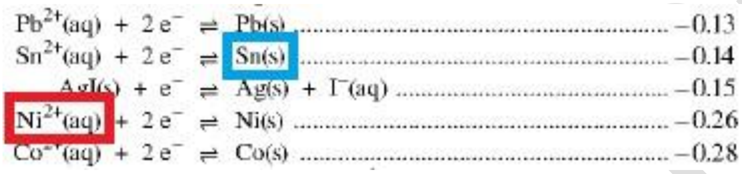
Option D is correct! $X(s) = lead, Y(s) = zinc, Z(s) = silver$

Question 26

A $Ni(NO_3)_2(aq)$ solution can be stored in a container made of a metal that will not react with the oxidizing agent, $Ni^{2+}(aq)$.

The RA must be in this orientation with respect to $Ni^{2+}(aq)$ on the table of standard reduction potentials:

$$\begin{array}{ccc} & & RA \\ & \rightleftharpoons & \\ Ni^{2+}(aq) & & \end{array}$$



$Pb^{2+}(aq) + 2\,e^-$	$\rightleftharpoons$	$Pb(s)$	....	−0.13
$Sn^{2+}(aq) + 2\,e^-$	$\rightleftharpoons$	$Sn(s)$	....	−0.14
$AgI(s) + e^-$	$\rightleftharpoons$	$Ag(s) + I^-(aq)$	....	−0.15
$Ni^{2+}(aq) + 2\,e^-$	$\rightleftharpoons$	$Ni(s)$	....	−0.26
$Co^{2+}(aq) + 2\,e^-$	$\rightleftharpoons$	$Co(s)$	....	−0.28

The correct answer is A, tin.

Question 28

This half-reaction is occurring in an "acidic environment" so it must be balanced in acidic conditions.

$$ClO_3^-(aq) \rightarrow ClO_2(g)$$

The major atoms are already balanced so the next step would be to balance oxygen by adding $H_2O(l)$ to one side:

$$ClO_3^-(aq) \rightarrow ClO_2(g) + \boldsymbol{H_2O(l)}$$

$H^+(aq)$ must be added to the other side to balance the hydrogens:

$$ClO_3^-(aq) + \mathbf{2}\ \boldsymbol{H^+(aq)} \rightarrow ClO_2(g) + H_2O(l)$$

The charge must then be balanced by adding electrons to the appropriate side. The overall charge on the left is +1 and the overall charge on the right is 0. Therefore, a single electron must be placed on the left side:

$$ClO_3^-(aq) + 2\,H^+(aq) + \boldsymbol{e^-} \rightarrow ClO_2(g) + H_2O(l)$$

Option B is correct!

Question 29

Options A and C are incorrect. Both half-reactions have electrons on the left side showing a gain of electrons which would make them reduction half-reactions.

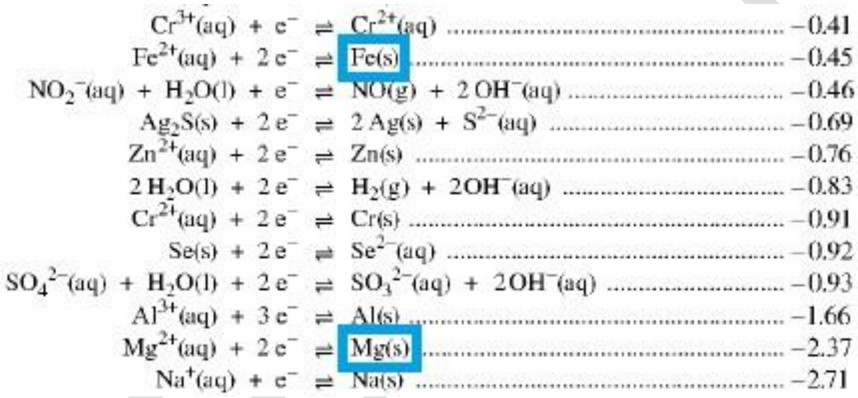
Option B is incorrect. The overall charge on the left (-2) does not match the overall charge on the right (-4).

Option D is correct! Everything balances!

Question 30

A sacrificial anode is a metal that is a stronger reducing agent, making it a more easily oxidized metal, than the metal it is protecting.

The correct option should be a metal that sits lower in the table of standard reduction potentials than iron.



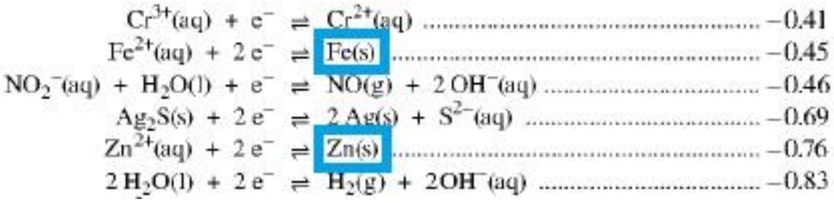
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons \boxed{Fe(s)}$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	−0.91
$Se(s) + 2e^- \rightleftharpoons Se^{2-}(aq)$	−0.92
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	−0.93
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons \boxed{Mg(s)}$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71

The correct answer is D, magnesium.

Question 33

Options A and B are incorrect. They are not practical solutions. The pipeline is already built. Also, even if the new metals are more resistant to corrosion than iron, they are STILL susceptible to corrosion.

Placing metallic strips at appropriate intervals would be comparable to using a series of sacrificial anodes along the pipeline. The correct choice of metal is one which is a stronger reducing agent, i.e. more susceptible to oxidation, than iron.



$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons \boxed{Fe(s)}$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons \boxed{Zn(s)}$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83

Option D is correct! Zinc is a stronger reducing agent than iron.

Question 34

$$4\,FeS_2(s) + 11\,\overset{0}{O}_2(g) \rightarrow 2\,Fe_2\,\overset{-2}{O}\,_3(s) + 8\,S\,\overset{-2}{O}\,_2(g)$$

The oxidation number of elemental oxygen is 0. On the products side of the reaction, the oxidation number of oxygen in both compounds is -2 as expected.

The correct answer is A! The oxidation number of oxygen changes from 0 to -2.

Question 36

A disproportionation reaction is one in which a single chemical entity both oxidizes and reduces. This is evident when assigning oxidation numbers. One reactant element’s oxidation number will change to two different oxidation numbers in two different entities on the products side.

Option C is correct!

$$\overset{0}{Cl_2}(aq) + H_2O(l) \rightarrow HO\,\overset{+1}{Cl}\,(aq) + H^+(aq) + \overset{-1}{Cl^-}(aq)$$

The chlorine in Cl_2(aq) starts with an oxidation number of 0 and it disproportionates, giving chlorines with oxidation numbers of +1 in HOCl(aq) and -1 in Cl^-(aq).

Question 39

The reducing agent and the substance being oxidized are the SAME THING! The RA can be determined by observing the changes in oxidation numbers from reactants to products. Note: All the hydrogens and oxygens are present in compounds so they can be easily assigned oxidation numbers of +1 and -2 respectively.

$$16\;\overset{+1}{H}\,_2\,\underset{RA}{\overset{-2}{S}}\,(g) + 8\;\underset{OA}{\overset{+4}{S}}\;\overset{-2}{O}\,_2(g) \rightarrow 16\,H_2O(l) + 3\,\overset{0}{S_8}(g)$$

The sulfur in H_2S(g) starts with an oxidation number of -2 which changes to 0 in S_8(g).

Option A is correct! H_2S(g) is both the reducing agent (**i**) and the substance being oxidized (**ii**).

Question 42

Note: The equation should read:

$$\overset{0}{Cl_2}(aq) + 2\,OH^-(aq) \rightarrow HO\,\overset{+1}{Cl}\,(aq) + H_2O(l) + \overset{-1}{\boldsymbol{Cl^-}}(\boldsymbol{aq})$$

The chlorine in Cl_2(aq) starts with an oxidation number of 0 and it disproportionates, giving chlorines with oxidation numbers of +1 in HOCl(aq) and -1 in Cl^-(aq).

Option A is correct!

Question 43

It is important to recognize that $CaMoO_4$(s) is an ionic compound composed of the ions Ca^{2+} and MoO_4^{2-}. The calcium, being a monoatomic ion, can easily be assigned an oxidation number of +2.

One can now focus on the MoO_4^{2-} ion:

$$\overset{x}{Mo}\;\overset{-2}{O}{}_4^{\,2-}$$

There are four oxygen atoms with an oxidation number of -2. The overall charge of the ion is -2:

$$4(-2) + x = -2$$

$$x = +6$$

Option C is correct! The oxidation number of Mo in $CaMoO_4$(s) is +6.

Question 46

All the hydrogens and oxygens in the reaction below are present in compounds so they can be easily assigned oxidation numbers of +1 and -2 respectively. It is iodine's and chlorine's oxidation numbers that are changing.

$$3\ \underset{RA}{\overset{0}{I_2}}\ (s) + 5\ \underset{OA}{\overset{+5}{Cl}}\,O_3^-(aq) + 3\ H_2O(l) \rightarrow 6\ H\ \overset{+5}{I}\ \ O_3(aq) + 5\ \overset{-1}{Cl^-}(aq)$$

The oxidation number of iodine changes from 0 to +5. Therefore, I$_2$(s) is the species that loses electrons (**i**).

Six iodine atoms with an oxidation number of 0 become six iodine atoms with an oxidation number of +5:

$$6(+5) - 6(0) = +30$$

The total change in oxidation number is +30 which means that 30 electrons are lost (**ii**).

The correct answer is A! $i = I_2(s), ii = 30$

Question 47

Given an unbalanced, complete, redox reaction, this must be balanced using oxidation numbers.

All the hydrogens and oxygens in the reaction below are present in compounds, or monoatomic ions in the case of H$^+$(aq), so they can be easily assigned oxidation numbers of +1 and -2 respectively. It is nitrogen's and aluminum's oxidation numbers that are changing.

$$H_2O(l) +\ \overset{+3}{N}\ O_2^-(aq) + \overset{0}{Al}(s) \rightarrow\ \overset{-3}{N}\ H_3(g) +\ \overset{+3}{Al}\,O_2^-(aq) + H^+(aq)$$

There is no need to initially balance the atoms whose oxidation number is changing so the increase in oxidation number by aluminum must be balanced with the decrease in oxidation number by nitrogen.

$$\mathbf{2}\times(Al\ increase\ by\ 3) = N\ decrease\ by\ 6$$

Coefficients must be placed in front of Al(s) and AlO$_2^-$(aq) as follows:

$$H_2O(l) +\ \overset{+3}{N}\ O_2^-(aq) + \mathbf{2}\ \overset{0}{Al}(s) \rightarrow\ \overset{-3}{N}\ H_3(g) + \mathbf{2}\ \ \overset{+3}{Al}\,O_2^-(aq) + H^+(aq)$$

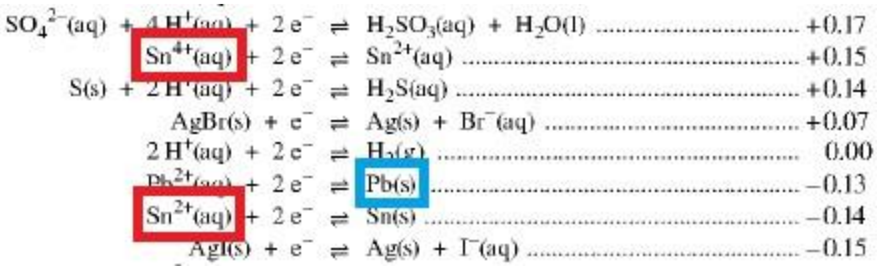
The rest of the atoms can now be balanced. When finished, be sure to check that the overall charge of the reactants balances with the overall charge of the products.

$$\mathbf{2}\ H_2O(l) +\ \overset{+3}{N}\ O_2^-(aq) + 2\ \overset{0}{Al}(s) \rightarrow\ \overset{-3}{N}\ H_3(g) + 2\ \ \overset{+3}{Al}\,O_2^-(aq) + H^+(aq)$$

The correct answer is **2121**.

Question 49

The correct option is the reducing agent that sits between the oxidizing agents, Sn^{4+}(aq) and Sn^{2+}(aq), on the table of standard electrode potentials.



$SO_4^{2-}(aq) + 4\,H^+(aq) + 2\,e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$ .... +0.17

$\boxed{Sn^{4+}(aq)} + 2\,e^- \rightleftharpoons Sn^{2+}(aq)$ .... +0.15

$S(s) + 2\,H^+(aq) + 2\,e^- \rightleftharpoons H_2S(aq)$ .... +0.14

$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$ .... +0.07

$2\,H^+(aq) + 2\,e^- \rightleftharpoons H_2(g)$ 0.00

$Pb^{2+}(aq) + 2\,e^- \rightleftharpoons \boxed{Pb(s)}$ .... −0.13

$\boxed{Sn^{2+}(aq)} + 2\,e^- \rightleftharpoons Sn(s)$ .... −0.14

$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$ .... −0.15

The correct answer is B, Pb(s).

Question 59

The first thing that must be done is balance the equation using oxidation numbers:

$$\overset{+7}{MnO_4^-}(aq) + \mathbf{8}\,H^+(aq) + \mathbf{5}\,\overset{+2}{Fe^{2+}}(aq) \rightarrow \overset{+2}{Mn^{2+}}(aq) + \mathbf{4}\,H_2O(l) + \mathbf{5}\,\overset{+3}{Fe^{3+}}(aq)$$

Next, determine the moles of MnO_4^-(aq) used in the titration:

$$n_{MnO_4^-(aq)} = (V_{MnO_4^-(aq)})([MnO_4^-(aq)]) = (0.0550\ L)\left(0.100\ \frac{mol}{L}\right) = 0.00550\ mol$$

The moles of MnO_4^-(aq) can be converted to moles of Fe^{2+}(aq) using the coefficients in the balanced chemical equation:

$$n_{Fe^{2+}(aq)} = \left(\frac{5\ mol\ Fe^{2+}(aq)}{1\ mol\ MnO_4^-(aq)}\right)\left(0.00550\ MnO_4^-(aq)\right) = 0.0275\ mol\ Fe^{2+}(aq)$$

Finally, the mass of iron in the ore sample can be determined using the molar mass of iron:

$$m_{Fe^{2+}} = (0.0275\ mol\ Fe^{2+})\left(\frac{55.85\ g\ Fe^{2+}}{1\ mol\ Fe^{2+}}\right) \doteq \mathbf{1.54\ g\ Fe^{2+}}$$

Option D is correct!

Question 64

To calculate the amount of dichromate required, the three trial volumes of titrant must first be averaged:

$$\boldsymbol{I}: 27.2\ mL - 10.1\ mL = 17.1\ mL$$

$$\boldsymbol{II}: 44.5\ mL - 27.2\ mL = 17.3\ mL$$

$$\boldsymbol{III}: 30.1\ mL - 12.9\ mL = 17.2\ mL$$

$$Average\ Titrant\ Volume\ Used = \frac{17.1\ mL + 17.3\ mL + 17.2\ mL}{3} = 17.2\ mL = 0.0172\ L$$

The amount of dichromate can now be easily determined by multiplying the average volume by the concentration of the dichromate solution:

$$n_{Cr_2O_7^{2-}(aq)} = (0.0172\ L)\left(0.125\ \frac{mol}{L}\right) \doteq \boldsymbol{2.15\times10^{-3}\ mol}$$

Option D is correct!

Question 65

The balanced chemical equation for this titration can be put together using the half-reactions in the table of standard electrode potentials.

Reduction:

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

Oxidation (The half-reaction needs to be multiplied by 3 to balance the electrons lost with the electrons gained.)**:**

$$(Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2\ e^-)\times\mathbf{3}$$

Overall reaction:

$$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 3\ Sn^{2+}(aq) \rightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l) + 3\ Sn^{4+}(aq)$$

The moles of $Cr_2O_7^{2-}$(aq) from Question 64 can be converted to moles of Sn^{2+}(aq) using the coefficients in the balanced chemical equation:

$$n_{Cr_2O_7^{2-}(aq)} = \left(\frac{3\ mol\ Sn^{2+}(aq)}{1\ mol\ Cr_2O_7^{2-}(aq)}\right)\left(0.00215\ mol\ Cr_2O_7^{2-}(aq)\right) = 0.00645\ mol\ Sn^{2+}(aq)$$

Finally, the concentration of Sn^{2+}(aq) can be calculated by dividing the moles of Sn^{2+}(aq) by the 20.0 mL sample volume (Convert mL to L.):

$$[Sn^{2+}(aq)] = \frac{0.00645\ mol}{0.0200\ L} \doteq \mathbf{3.23\times10^{-1}}\ \frac{\boldsymbol{mol}}{\boldsymbol{L}}$$

$$\boldsymbol{a = 3}$$

$$\boldsymbol{b = 2}$$

$$\boldsymbol{c = 3}$$

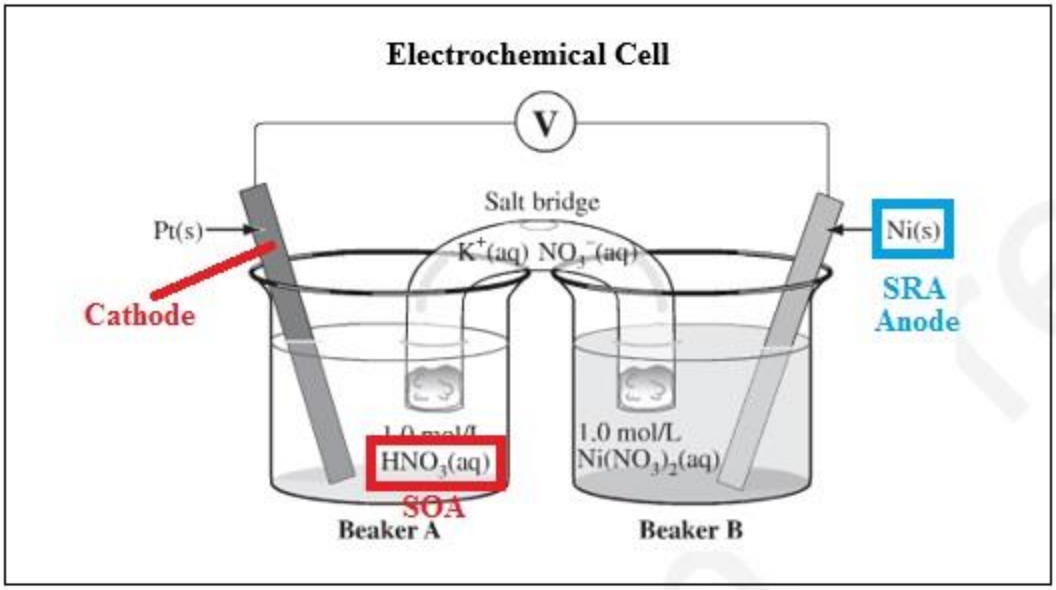
$$\boldsymbol{d = 1}$$

The correct answer is 3231.

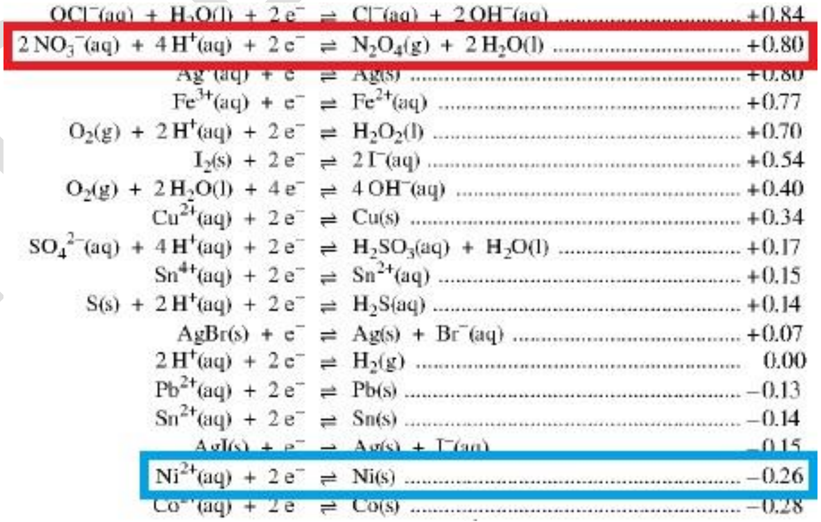
Question 70

If the standard silver reduction half-reaction had been chosen as the reference half-reaction instead of the hydrogen reduction half-reaction it would NOT change the relative positions of all the other reduction half-reactions. $E°_{\text{cell}}$ values can be calculated normally regardless of what is chosen as the reference half-reaction.

The overall redox reaction happens between the strongest oxidizing agent (SOA) and the strongest reducing agent (SRA).



When calculating the $E°_{\text{cell}}$ for the cell, **use the reduction potentials given in the table** of standard electrode potentials.



$OCl^-(aq) + H_2O(l) + 2e^-$ ⇌	$Cl^-(aq) + 2OH^-(aq)$	+0.84
$2NO_3^-(aq) + 4H^+(aq) + 2e^-$ ⇌	$N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^-$ ⇌	$Ag(s)$	+0.80
$Fe^{3+}(aq) + e^-$ ⇌	$Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^-$ ⇌	$H_2O_2(l)$	+0.70
$I_2(s) + 2e^-$ ⇌	$2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^-$ ⇌	$4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^-$ ⇌	$Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$ ⇌	$H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^-$ ⇌	$Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^-$ ⇌	$H_2S(aq)$	+0.14
$AgBr(s) + e^-$ ⇌	$Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^-$ ⇌	$H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^-$ ⇌	$Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^-$ ⇌	$Sn(s)$	−0.14
$AgI(s) + e^-$ ⇌	$Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^-$ ⇌	$Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^-$ ⇌	$Co(s)$	−0.28

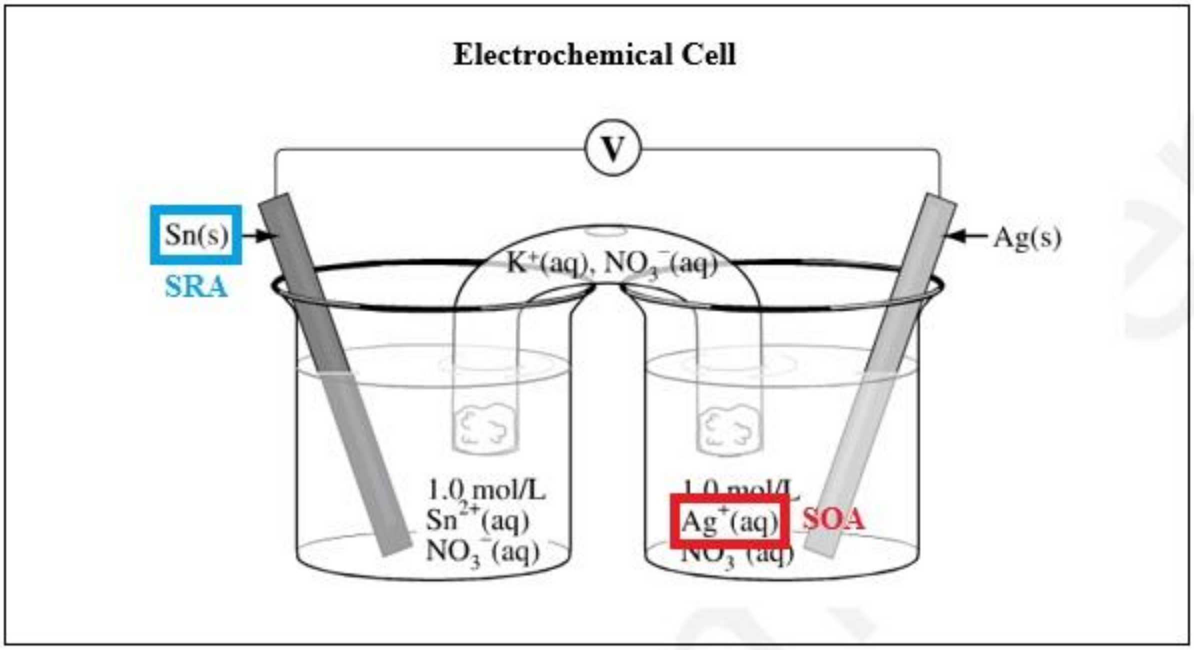
$$E°_{cell} = E°_{r_{cathode}} - E°_{r_{anode}} = (+0.80\ V) - (-0.26\ V) = \mathbf{+1.06\ V}$$

The correct answer is **1.06**.

Question 71

There is a typo in this question.

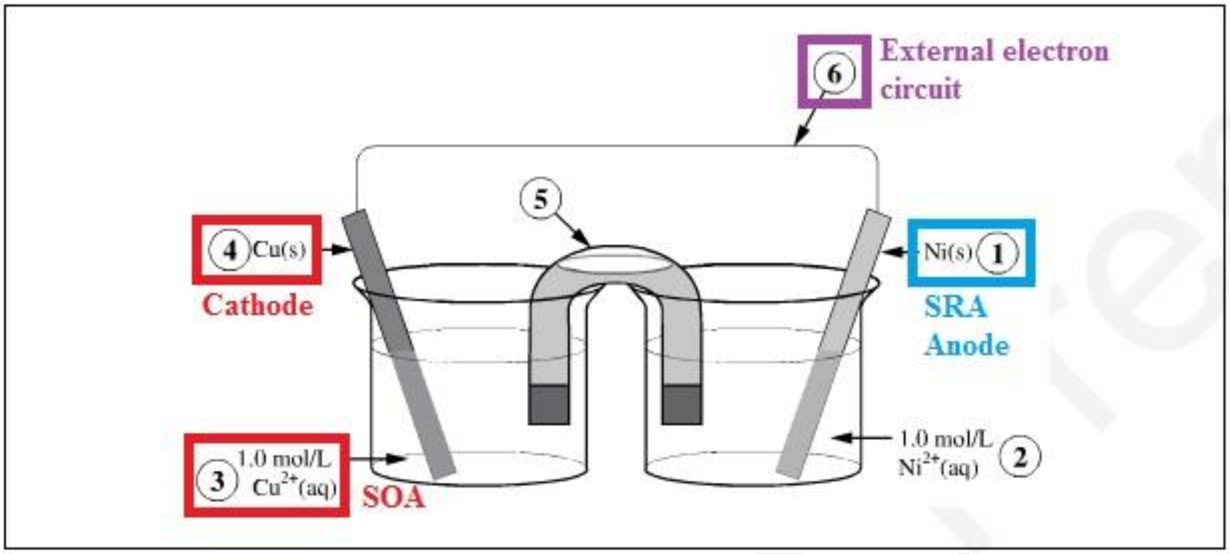
The overall redox reaction happens between the strongest oxidizing agent (SOA) and the strongest reducing agent (SRA).



The correct answer is said to be D and it SHOULD read:

$$\boldsymbol{D}.\ 2\,Ag^{+}(aq) + Sn(s) \rightarrow 2\,Ag(s) + Sn^{2+}(aq)$$

Question 73



Remember that both the cathode and the anode are the actual electrodes that are in contact with the strongest oxidizing agent (SOA) and the strongest reducing agent (SRA). In the case of the anode, it can be made of the metal that is the strongest reducing agent.

$$Cathode = \mathbf{4}$$

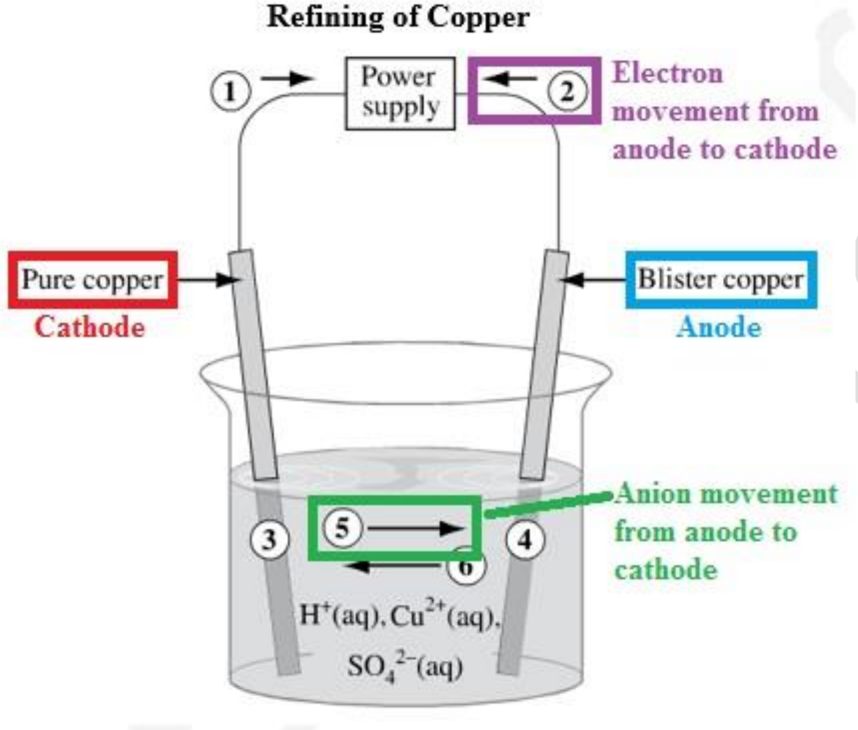
$$External\ electron\ circuit = \mathbf{6}$$

$$Oxidizing\ agent = \mathbf{3}$$

$$Anode = \mathbf{1}$$

The correct answer is **4631**.

Question 74



The information in the question states that the blister copper is the anode and the pure copper is the cathode.

It is at the cathode where reduction occurs (**3**).

Anions move to the anode (**5**).

Electrons flow from the anode to the cathode (**2**).

It is the cathode that increases in mass because positive metal ions gain electrons to for pure solid metal which plates at the cathode (**3**).

The correct answer is **3523**.

Question 77

The given reduction half-reactions' order should be reversed. It is the positive reduction potentials that are at the top of the table.

$$Ag_2O(s) + H_2O(l) + 2\,e^- \rightarrow 2\,Ag(s) + 2\,OH^-(aq) \quad E° = +0.34\,V$$

$$Zn(OH)_2(s) + 2\,e^- \rightarrow Zn(s) + 2\,OH^-(aq) \quad E° = -1.25\,V$$

With the correct order, it can easily be seen that the strongest reducing agent is Zn(s) which is the substance oxidized.

Option A is correct! Zn(s)

Question 81

It is **oxidation** that occurs at the **anode** so options B and D can be eliminated as those half-reactions are examples of reductions with the electrons on the reactants side.

$$NiO_2(s) + \underset{SRA}{Cd(s)} + 2\,H_2O(l) \rightarrow Cd(OH)_2(s) + Ni(OH)_2(s)$$

In the equation above, Cd(s) is the only pure metal solid on the reactants side and is the strongest reducing agent based on its placement in the table of standard electrode potentials. The correct option should show the oxidation of Cd(s) with electrons on the products side.

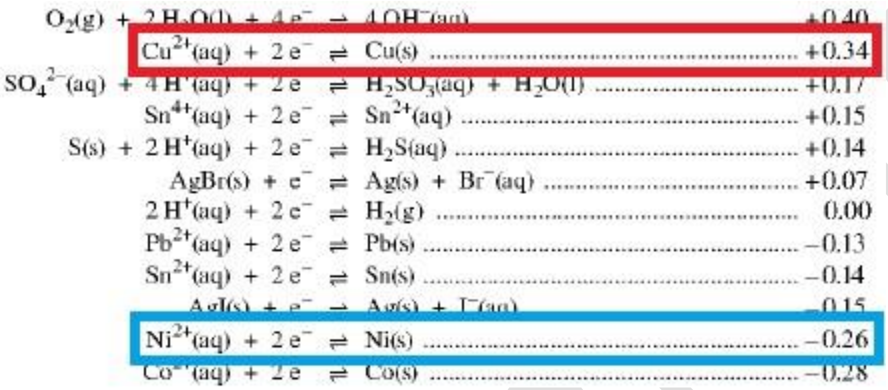
Option A is correct!

$$Cd(s) + 2\,OH^-(aq) \rightarrow Cd(OH)_2(s) + 2\,e^-$$

Question 86

If a different reduction half-reaction is chosen for the standard reference half-cell the electrical potential for any other half-reaction can simply be calculated by:

$$E°_{r_{new}} = E°_{r_{old}} - E°_{r_{ref}}$$



$O_2(g) + 2\,H_2O(l) + 4\,e^- \rightleftharpoons 4\,OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2\,e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4\,H^+(aq) + 2\,e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2\,e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2\,H^+(aq) + 2\,e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2\,H^+(aq) + 2\,e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2\,e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2\,e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2\,e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2\,e^- \rightleftharpoons Co(s)$	−0.28

The half-reaction chosen as the new reference half-reaction is that of $Cu^{2+}(aq)/Cu(s)$.

$$E°_{r_{ref}} = +0.34\ V$$

The reduction potential of the $Ni^{2+}(aq)/Ni(s)$ half-reaction will change from its original reduction potential.

$$E°_{r_{old}} = -0.26\ V$$

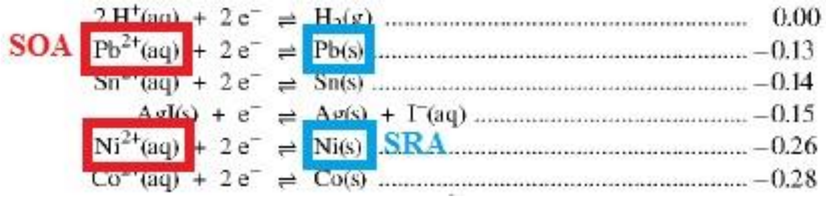
When the reduction potential of $Cu^{2+}(aq)/Cu(s)$ is set to 0.00 V, the new reduction potential for $Ni^{2+}(aq)/Ni(s)$ will be:

$$E°_{r_{new}} = (-0.26\ V) - (+0.34\ V) = \mathbf{-0.60\ V}$$

Option D is correct!

Question 93

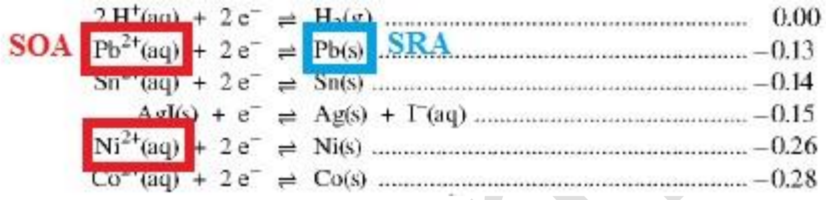
In the first cell:



$2\ H^+(aq)$	+	$2\ e^-$	⇌	$H_2(g)$	0.00
$Pb^{2+}(aq)$	+	$2\ e^-$	⇌	$Pb(s)$	−0.13
$Sn^{2+}(aq)$	+	$2\ e^-$	⇌	$Sn(s)$	−0.14
$AgI(s)$	+	e^-	⇌	$Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq)$	+	$2\ e^-$	⇌	$Ni(s)$	−0.26
$Co^{2+}(aq)$	+	$2\ e^-$	⇌	$Co(s)$	−0.28

The strongest oxidizing agent (SOA) and the strongest reducing agent (SRA) are in the necessary orientation on the table for a spontaneous reaction to occur.

In the second cell:

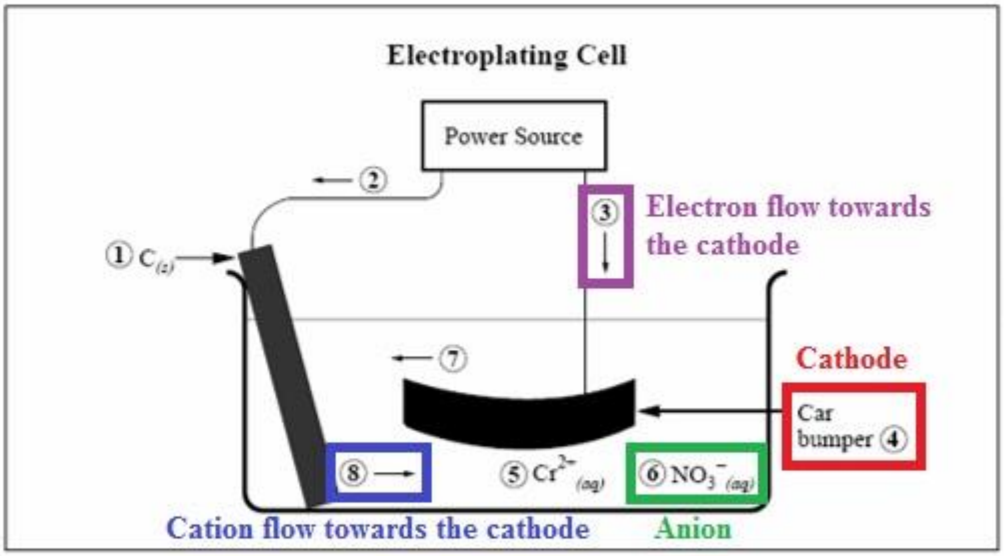


$2\ H^+(aq)$	+	$2\ e^-$	⇌	$H_2(g)$	0.00
$Pb^{2+}(aq)$	+	$2\ e^-$	⇌	$Pb(s)$	−0.13
$Sn^{2+}(aq)$	+	$2\ e^-$	⇌	$Sn(s)$	−0.14
$AgI(s)$	+	e^-	⇌	$Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq)$	+	$2\ e^-$	⇌	$Ni(s)$	−0.26
$Co^{2+}(aq)$	+	$2\ e^-$	⇌	$Co(s)$	−0.28

There is no Ni(s) electrode present. The strongest reducing agent is Pb(s). The SOA and SRA are NOT in the necessary orientation on the table for a spontaneous reaction to occur.

The correct answer is D! The reaction in the first cell is spontaneous and the reaction in the second cell is nonspontaneous.

Question 96



In electroplating cells, the object to be electroplated is the cathode (**4**).

Even in electrolytic cells, the electrons flow from the anode to the cathode (**3**).

Cations flow toward the cathode (**8**).

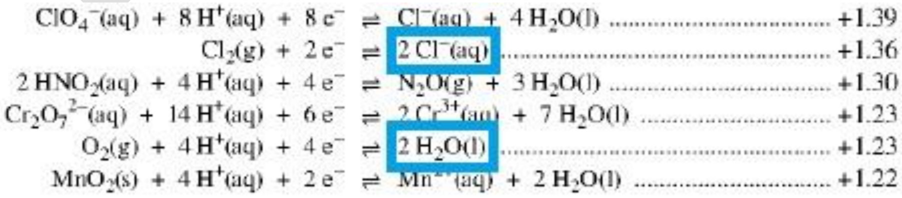
This is the tricky part of the question. It ONLY asks for an **anion**, not anion movement. The only anion present is $NO_3^-(aq)$ (**6**).

The correct answer is 4386.

Question 98

This is an example of the chloride anomaly in the electrolysis of a sodium chloride solution.

As per the table of standard electrode potentials, $H_2O(l)$ should be the strongest reducing agent present in an aqueous sodium chloride solution.



$ClO_4^-(aq) + 8\,H^+(aq) + 8\,e^-$	$\rightleftharpoons$	$Cl^-(aq) + 4\,H_2O(l)$	+1.39
$Cl_2(g) + 2\,e^-$	$\rightleftharpoons$	$\boxed{2\,Cl^-(aq)}$	+1.36
$2\,HNO_2(aq) + 4\,H^+(aq) + 4\,e^-$	$\rightleftharpoons$	$N_2O(g) + 3\,H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,e^-$	$\rightleftharpoons$	$2\,Cr^{3+}(aq) + 7\,H_2O(l)$	+1.23
$O_2(g) + 4\,H^+(aq) + 4\,e^-$	$\rightleftharpoons$	$\boxed{2\,H_2O(l)}$	+1.23
$MnO_2(s) + 4\,H^+(aq) + 2\,e^-$	$\rightleftharpoons$	$Mn^{2+}(aq) + 2\,H_2O(l)$	+1.22

However, the reducing agent that oxidizes in the cell is $Cl^-(aq)$, which is contrary to what we would expect from what is shown on the table.

The correct answer is B! $2\,Cl^-(aq) \rightarrow Cl_2(g) + 2\,e^-$

Question 99

It should be noticed in the diagram that the electrons are flowing towards the purified copper sheet. Electrons flow from the anode to the cathode. The purified copper sheet is the cathode!

Therefore, options A and C are incorrect.

Option D is incorrect. SO_4^{2-}(aq) can only be an oxidizing agent when it is in combination with other entities, like H^+(aq).

Option B is correct! This is an apparatus for the purification of copper. The Cu^{2+}(aq) reduces to give pure Cu(s).

Question 101

The given half reactions are as follows.

Oxidation (Anode):

$$(CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6\,H^+(aq) + 6\,e^-)\times\mathbf{2}$$

Reduction (Cathode):

$$(O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l))\times\mathbf{3}$$

The above half reactions are multiplied by the necessary numbers to balance the electrons lost with the electrons gained.

Combining the two half-reactions and cancelling the appropriate quantities of H_2O(l) and H^+(aq) that appear on both sides will give:

$$\mathbf{2}\,CH_3OH(l) + \mathbf{3}\,O_2(g) \rightarrow \mathbf{2}\,CO_2(g) + \mathbf{4}\,H_2O(l)$$

It is simply the balanced chemical equation for the complete combustion of methanol!

The correct answer is **3422**.

Question 102

The question states that the $E°_{\text{cell}}$ for the methanol fuel cell is +0.80 V.

From the table of standard electrode potentials:

$$E°_{r_{cathode}} = +1.23\ V$$

Remember:

$$E°_{cell} = E°_{r_{cathode}} - E°_{r_{anode}}$$

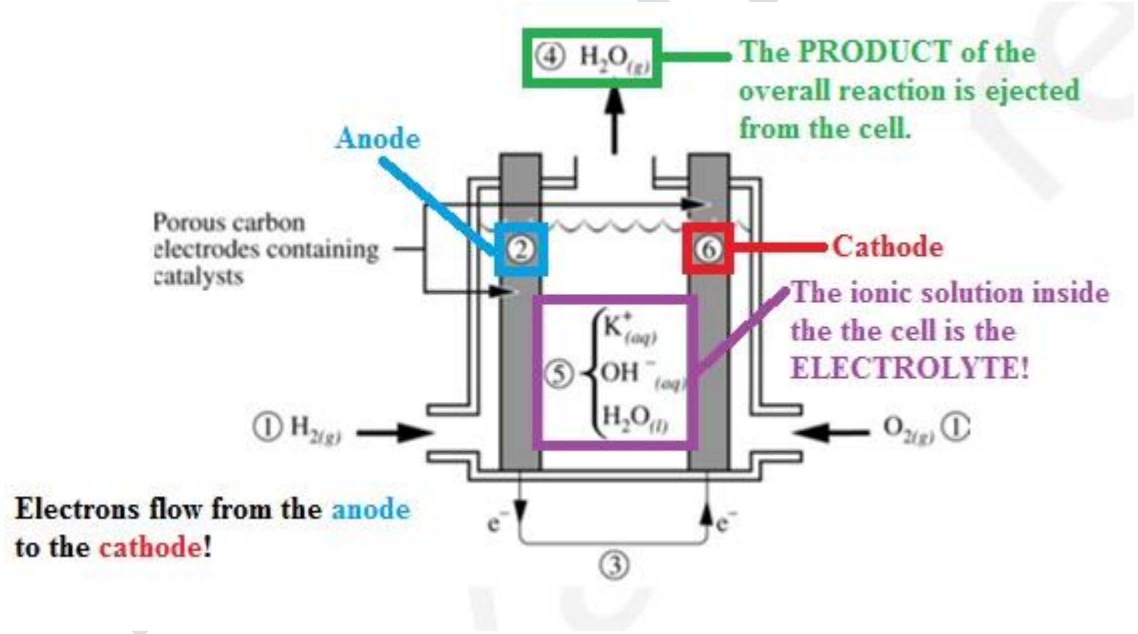
Substituting given values and solving for the unknown gives:

$$+0.80\ V = (+1.23\ V) - E°_{r_{anode}}$$

$$E°_{r_{anode}} = \mathbf{+0.43\ V}$$

The correct answer is **0.43**.

Question 105



The correct answer is **2654.**

Question 106

The anode is the site at which oxidation occurs. The solid metal of which the anode is made oxidizes and positive metal ions are released into the electrolytic solution. This results in **a decrease in the mass of the anode**. Therefore, options A and C can be eliminated.

Which metal the anode is will be the strongest reducing agent of the two metal solids and can be determined using the table of standard electrode potentials. It is the lowest on the right.

$$\underset{\substack{SRA \\ \boldsymbol{ANODE}!}}{Cd(s)}\ \left|\, Cd^{2+}(aq) \,\right\|\, Ag^{+}(aq) \,\left|\, Ag(s)\right.$$

Unfortunately, the formula for cell stoichiometry is not given in the data booklet. However, there is a simple mnemonic device that can be used to determine the gram mass of metal lost or gained.

$$m_{metal} = \frac{"Mint"}{"Funny"} = \frac{M_{metal} \cdot I \cdot n_{metal} \cdot t}{F \cdot n_{e^-}}$$

M_{metal} is the molar mass of the metal in g/mol. n_{metal} and n_{e^-} are the respective balancing coefficients in front of the metal and the electrons in the half-reaction. Remember that units of **seconds** must be used for **time**.

$$M_{Cd(s)} = 112.41\ g/mol$$

$$t = 2.00\ h = 7200\ sec$$

$$Cd(s) \rightarrow Cd^{2+}(aq) + 2\ e^{-}$$

$$m_{Cd(s)}\ lost = \frac{\left(112.41\ \frac{g\ Cd(s)}{mol\ Cd(s)}\right)(6.00\ A)(1\ mol\ Cd(s))(7200\ s)}{\left(9.65\times10^{4}\ \frac{C}{mol\ e^{-}}\right)(2\ mol\ e^{-})} \doteq \boldsymbol{25.2\ g\ Cd(s)\ lost}$$

The correct answer is B, a 25.2 g decrease!

Question 109

This is a simple Chem 20-type stoichiometry question which uses the half-reaction for the oxidation of zinc.

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^{-}$$

$$n_{e^-} = 200\ g\ Zn(s) \cdot \frac{1\ mol\ Zn(s)}{65.41\ g\ Zn(s)} \cdot \frac{2\ mol\ e^{-}}{1\ mol\ Zn(s)} \doteq 6.12\ mol\ e^{-}$$

The correct answer is C, 6.12 mol.

Question 115

This is a simple conversion of mass to moles using the molar mass.

$$n_{Cr^{2+}(aq)} = 112\ g\ Cr^{2+}(aq) \cdot \frac{1\ mol\ Cr^{2+}(aq)}{52.00\ g\ Cr^{2+}(aq)} = 2.15\ mol\ Cr^{2+}(aq)$$

The correct answer is **2.15**.

Question 116

The solution requires a rearrangement of the mint/funny formula to isolate for time.

$$m_{Cr(s)} = \frac{M_{Cr(s)} \cdot I \cdot n_{Cr(s)} \cdot t}{F \cdot n_{e^-}}$$

$$t = \frac{m_{Cr(s)} \cdot F \cdot n_{e^-}}{M_{Cr(s)} \cdot I \cdot n_{Cr(s)}}$$

The half-reaction for the reduction of Cr^{2+}(aq) is:

$$Cr^{2+}(aq) + 2\ e^- \rightarrow Cr(s)$$

So:

$$t = \frac{\left(112\ g\ Cr(s)\right)\left(9.65 \times 10^4 \frac{C}{mol\ e^-}\right)\left(2\ mol\ e^-\right)}{\left(52.00\ \frac{g\ Cr(s)}{mol\ Cr(s)}\right)\left(2000\ A\right)\left(1\ mol\ Cr(s)\right)} \doteq 207.846\ s \doteq \mathbf{3.46\ min}$$

Don’t forget to convert the time in seconds to time in minutes!

The correct answer is **3.46**.