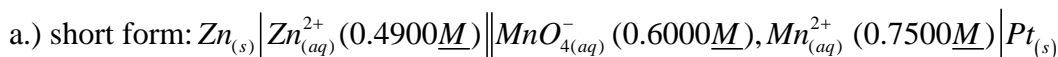
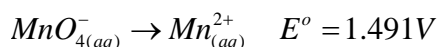
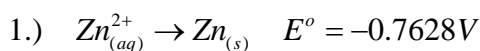


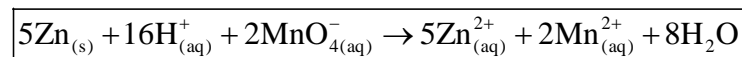
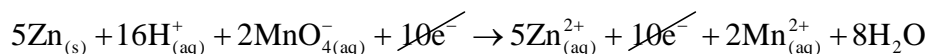
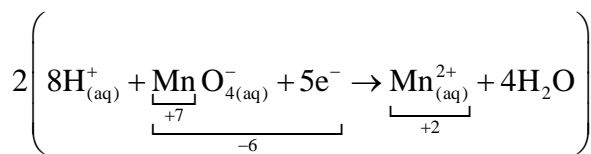
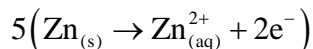
In Class Exercise for Chapter 20 – Electrochemistry

For the following four voltaic cells under **acidic** conditions (pH = 5.00 unless stated otherwise):

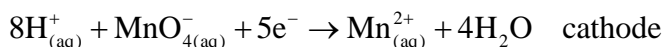
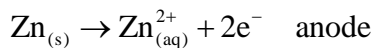
- write the short form of the cell
- write the balanced equation
- state which half reaction is the cathode and the anode
- state the species which is oxidized and reduced
- state the species that is the oxidizing reagent and reducing reagent
- determine the standard emf, E_{cell}°
- determine the emf, E_{cell} , for the given condition at 25°C
- is the reaction spontaneous?



b.) balanced equation:



c.) the cell is written in terms of electron flow: anode \rightarrow cathode



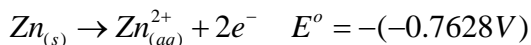
d.) oxidized species is Zn goes from 0 \rightarrow +2

reduced species is $\text{MnO}_{4(\text{aq})}^{-}$: Mn goes from +7 \rightarrow +2

e.) oxidizing agent: the one which is reduced $\text{MnO}_{4(\text{aq})}^{-}$

reducing agent: the one which is oxidized $\text{Zn}_{(\text{s})}$

f.) Standard emf



$$E_{\text{cell}}^{\circ} = 0.7628\text{V} + 1.491\text{V} = 2.2538\text{V} \sim 2.254$$

$$g.) E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \left(\frac{[\text{Zn}^{2+}]^5 [\text{Mn}^{2+}]^2}{[\text{MnO}_4^-]^2 [\text{H}^+]^{16}} \right)$$

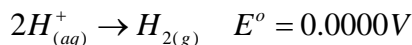
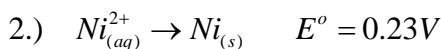
there are actually 10e-'s which travel from the anode to cathode as determined from the balanced equation - therefore, $n = 10$ the only reagents that are not under standard conditions are Zn^{2+} and MnO_4^- therefore they are the only ones which really effect our E_{cell} (if all reactants were 1 M then our emf would be standard $E_{\text{cell}} = E_{\text{cell}}^{\circ}$). Also, we place the product, Zn^{2+} , concentration raised to the power of its molar coefficient, 5, over the reactant concentration, MnO_4^- , raised to the power of its molar coefficient, 2, inside the log of the Nernst equation. Finally we include the pH as given from above.

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \left(\frac{[\text{Zn}^{2+}]^5 [\text{Mn}^{2+}]^2}{[\text{MnO}_4^-]^2 [\text{H}^+]^{16}} \right)$$

$$E_{\text{cell}} = 2.2538 \text{ V} - \frac{0.0592}{10} \log \left(\frac{[0.4900]^5 [0.7500]^2}{[0.6000]^2 [10^{-5}]^{16}} \right)$$

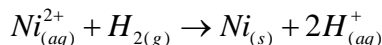
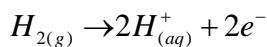
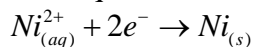
$$E_{\text{cell}} = 1.788 \text{ V}$$

i.) Since $E_{\text{cell}} > 0$ the reaction is spontaneous

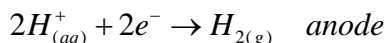
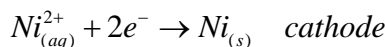


a.) short form: $\text{Pt}_{(s)} | \text{H}_{(aq)}^{+} (0.1500 \text{ M}) | \text{H}_{2(g)} || \text{Ni}_{(aq)}^{2+} (0.2500 \text{ M}) | \text{Ni}_{(s)}$

b.) balanced equation:



c.) anode and cathode:



d.) oxidized and reduced reagents:

Ni^{2+} is reduced from 2 \rightarrow 0

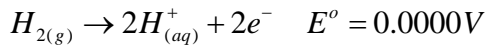
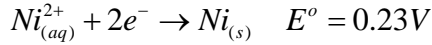
H_2 is oxidized from 0 \rightarrow +1

e.) oxidizing and reducing agents:

H_2 is the reducing agent since it is oxidized

Ni^{2+} is the oxidizing agent

f.) standard emf



$$E_{cell}^\circ = +0.23V$$

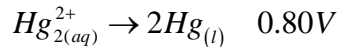
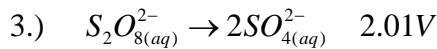
$$g.) E_{cell} = E_{cell}^\circ - \frac{0.0592}{n} \log \left(\frac{[H^+]^2}{[Ni^{2+}]} \right)$$

This time there are 2e-'s flowing from anode to cathode - n = 2

$$E_{cell} = 0.23V - \frac{0.0592}{2} \log \left(\frac{[0.1500]^2}{0.2500} \right)$$

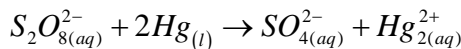
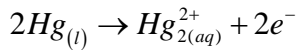
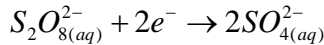
$$E_{cell} = 0.26V$$

h.) it is spontaneous, $E_{cell}^\circ > 0$

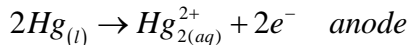
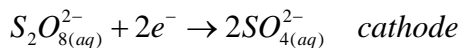


a.) short form: $Pt_{(s)} | Hg_{(l)} | Hg_{2(aq)}^{2+} (0.2500M) || S_2O_{8(aq)}^{2-} (0.0300M), SO_{4(aq)}^{2-} (0.2000M) | Pt_{(s)}$

b.) balanced equation:



c.) anode and cathode:



d.) oxidized and reduced reagents:

S is reduced from +7 \rightarrow +6

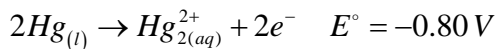
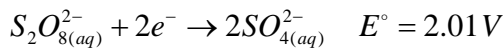
Hg is oxidized from 0 \rightarrow +1

e.) oxidizing and reducing agents:

$2Hg_{(l)}$ is the reducing agent since it is oxidized

$S_2O_{8(aq)}^{2-}$ is the oxidizing agent

f.) standard emf



$$E_{cell}^\circ = 1.21V$$

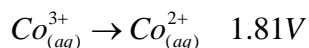
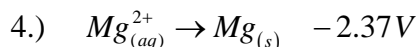
$$g.) E_{cell} = E_{cell}^\circ - \frac{0.0592}{n} \log \left(\frac{[SO_4^{2-}][Hg_2^{2+}]}{[S_2O_8^{2-}]} \right)$$

This time there are 2e-'s flowing from anode to cathode - n = 2

$$E_{\text{cell}} = 1.21 \text{ V} - \frac{0.0592}{2} \log \left(\frac{[0.2000][0.2500]}{[0.0300]} \right)$$

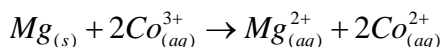
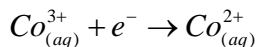
$$E_{\text{cell}} = 1.203 \text{ V} \sim 1.20 \text{ V}$$

h.) it is spontaneous, $E_{\text{cell}}^{\circ} > 0$

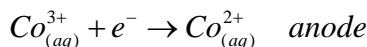
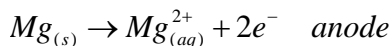


a.) short form: $\text{Mg}_{(s)} \mid \text{Mg}_{(aq)}^{2+} (0.5000 \text{ M}) \parallel \text{Co}_{(aq)}^{3+} (1.5000 \text{ M}), \text{Co}_{(aq)}^{2+} (0.6000 \text{ M}) \mid \text{Pt}_{(s)}$

b.) balanced equation:



c.) anode and cathode:



d.) oxidized and reduced reagents:

Mg is oxidized from 0 \rightarrow +2

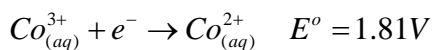
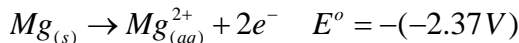
Co is reduced from +3 \rightarrow +2

e.) oxidizing and reducing agents:

$\text{Mg}_{(s)}$ is the reducing agent since it is oxidized

Co^{3+} is the oxidizing agent

f.) standard emf



$$E_{\text{cell}}^{\circ} = +4.18 \text{ V}$$

g.) $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \left(\frac{[\text{Mg}^{2+}][\text{Co}^{2+}]^2}{[\text{Co}^{3+}]^2} \right)$

This time there are 2e-'s flowing from anode to cathode - n = 2

$$E_{\text{cell}} = 4.18 \text{ V} - \frac{0.0592}{2} \log \left(\frac{[0.5000][0.6000]^2}{[1.5000]^2} \right)$$

$$E_{\text{cell}} = 4.21 \text{ V}$$

h.) it is spontaneous, $E_{\text{cell}}^{\circ} > 0$