In Class Exercise for Chapter 20 – Electrochemistry

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

a.) write the short form of the cell
b.) write the balanced equation
c.) state which half reaction is the cathode and the anode
d.) state the species which is oxidized and reduced
e.) state the species that is the oxidizing reagent and reducing reagent
f.) determine the standard emf, E^o_{cell}
g.) determine the emf, E_{cell}, for the given condition at 25°C
h.) is the reaction spontaneous?

1.) $Zn^{2+}_{(aq)} \rightarrow Zn_{(s)} \quad E^o = -0.7628V$

$MnO^-_{4(aq)} \rightarrow Mn^{2+}_{(aq)} \quad E^o = 1.491V$

a.) short form: $Zn_{(s)} \left| Zn^{2+}_{(aq)} (0.4900\underline{M}) \right\| MnO^-_{4(aq)} (0.6000\underline{M}), Mn^{2+}_{(aq)} (0.7500\underline{M}) \left| Pt_{(s)} \right.$

b.) balanced equation:

$$5\left(Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-\right)$$

$$2\left(8H^+_{(aq)} + \underbrace{\underbrace{Mn}_{+7}\, O^-_{4(aq)} + 5e^- \rightarrow \underbrace{Mn^{2+}_{(aq)}}_{+2}}_{-6} + 4H_2O\right)$$

$$5Zn_{(s)} + 16H^+_{(aq)} + 2MnO^-_{4(aq)} + \cancel{10e^-} \rightarrow 5Zn^{2+}_{(aq)} + \cancel{10e^-} + 2Mn^{2+}_{(aq)} + 8H_2O$$

$$\boxed{5Zn_{(s)} + 16H^+_{(aq)} + 2MnO^-_{4(aq)} \rightarrow 5Zn^{2+}_{(aq)} + 2Mn^{2+}_{(aq)} + 8H_2O}$$

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

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \quad$ anode

$8H^+_{(aq)} + MnO^-_{4(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O \quad$ cathode

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

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

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

reducing agent: the one which is oxidized $Zn_{(s)}$

f.) Standard emf

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \quad E^o = -(-0.7628V)$$

$$8H^+_{(aq)} + MnO^-_{4(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O \quad E^o = 1.491V$$

$$E^o_{cell} = 0.7628V + 1.491V = 2.2538V \sim 2.254$$

g.) $E_{cell} = E^{o}_{cell} - \frac{0.0592}{n}\log\left(\frac{[Zn^{2+}]^5\left[Mn^{2+}\right]^2}{[MnO_4^-]^2\left[H^+\right]^{16}}\right)$

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 $E_{cell} = E^{o}_{cell}$). 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.

$$E_{cell} = E^{o}_{cell} - \frac{0.0592}{n}\log\left(\frac{[Zn^{2+}]^5\left[Mn^{2+}\right]^2}{[MnO_4^-]^2\left[H^+\right]^{16}}\right)$$

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

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

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

2.) $Ni^{2+}_{(aq)} \rightarrow Ni_{(s)} \quad E^{o} = 0.23V$

$2H^{+}_{(aq)} \rightarrow H_{2(g)} \quad E^{o} = 0.0000V$

a.) short form: $Pt_{(s)}\left|H^{+}_{(aq)}(0.1500\underline{M})\right|H_{2(g)}\left\|Ni^{2+}_{(aq)}(0.2500\underline{M})\right|Ni_{(s)}$

b.) balanced equation:

$$Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$$

$$\underline{H_{2(g)} \rightarrow} 2H^{+}_{(aq)} + 2e^-$$

$$Ni^{2+}_{(aq)} + H_{2(g)} \rightarrow Ni_{(s)} + 2H^{+}_{(aq)}$$

c.) anode and cathode:

$Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$ *cathode*

$2H^{+}_{(aq)} + 2e^- \rightarrow H_{2(g)}$ *anode*

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

$$Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)} \quad E^o = 0.23V$$
$$\underline{H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^- \quad E^o = 0.0000V}$$
$$E^o_{cell} = +0.23V$$

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

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

$$\mathrm{E_{cell}} = 0.23\,\mathrm{V} - \frac{0.0592}{2}\log\left(\frac{[0.1500]^2}{0.2500}\right)$$
$$\mathrm{E_{cell}} = 0.26\,\mathrm{V}$$

h.) it is spontaneous, $\mathrm{E^o_{cell}} > 0$

3.) $S_2O^{2-}_{8(aq)} \rightarrow 2SO^{2-}_{4(aq)} \quad 2.01V$

$Hg^{2+}_{2(aq)} \rightarrow 2Hg_{(l)} \quad 0.80V$

a.) short form: $Pt_{(s)}\left|Hg_{(l)}\right|Hg^{2+}_{2(aq)}(0.2500\underline{M})\left\|S_2O^{2-}_{8(aq)}(0.0300\underline{M}), SO^{2-}_{4(aq)}(0.2000\underline{M})\right|Pt_{(s)}$

b.) balanced equation:

$$S_2O^{2-}_{8(aq)} + 2e^- \rightarrow 2SO^{2-}_{4(aq)}$$
$$\underline{2Hg_{(l)} \rightarrow Hg^{2+}_{2(aq)} + 2e^-}$$
$$S_2O^{2-}_{8(aq)} + 2Hg_{(l)} \rightarrow SO^{2-}_{4(aq)} + Hg^{2+}_{2(aq)}$$

c.) anode and cathode:

$$S_2O^{2-}_{8(aq)} + 2e^- \rightarrow 2SO^{2-}_{4(aq)} \quad cathode$$
$$2Hg_{(l)} \rightarrow Hg^{2+}_{2(aq)} + 2e^- \quad anode$$

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

$S_2O^{2-}_{8(aq)}$ is the oxidizing agent

f.) standard emf

$$S_2O^{2-}_{8(aq)} + 2e^- \rightarrow 2SO^{2-}_{4(aq)} \quad E^\circ = 2.01V$$
$$\underline{2Hg_{(l)} \rightarrow Hg^{2+}_{2(aq)} + 2e^-} \quad E^\circ = -0.80V$$
$$E^\circ_{cell} = 1.21V$$

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

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

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

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

h.) it is spontaneous, $E^o_{cell} > 0$

4.) $Mg^{2+}_{(aq)} \rightarrow Mg_{(s)} \quad -2.37V$

$Co^{3+}_{(aq)} \rightarrow Co^{2+}_{(aq)} \quad 1.81V$

a.) short form: $Mg_{(s)} \left| Mg^{2+}_{(aq)}(0.5000\underline{M}) \right\| Co^{3+}_{(aq)}(1.5000\underline{M}), Co^{2+}_{(aq)}(0.6000\underline{M}) \left| Pt_{(s)} \right.$

b.) balanced equation:

$$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^-$$

$$\underline{Co^{3+}_{(aq)} + e^- \rightarrow Co^{2+}_{(aq)}}$$

$$Mg_{(s)} + 2Co^{3+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Co^{2+}_{(aq)}$$

c.) anode and cathode:

$$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^- \quad anode$$

$$Co^{3+}_{(aq)} + e^- \rightarrow Co^{2+}_{(aq)} \quad anode$$

d.) oxidized and reduced reagents:

Mg is oxidized from 0 → +2

Co is reduced from +3 → +2

e.) oxidizing and reducing agents:

$Mg_{(s)}$ is the reducing agent since it is oxidized

Co^{3+} is the oxidizing agent

f.) standard emf

$$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^- \quad E^o = -(-2.37V)$$

$$\underline{Co^{3+}_{(aq)} + e^- \rightarrow Co^{2+}_{(aq)} \quad E^o = 1.81V}$$

$$E^o_{cell} = +4.18V$$

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

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

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

$$E_{cell} = 4.21V$$

h.) it is spontaneous, $E^o_{cell} > 0$