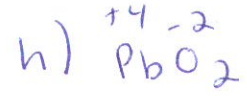
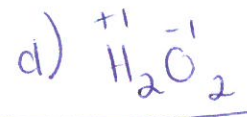
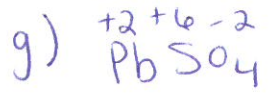
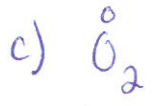
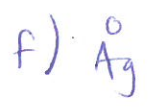
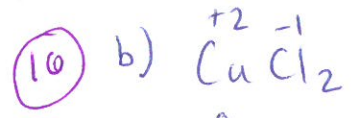
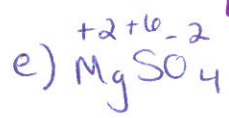
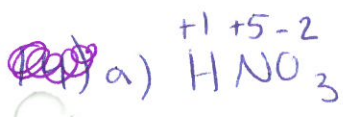


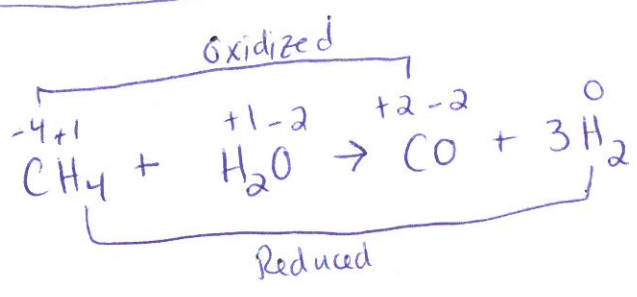
Chp. 18, Day 1 p. 877 (16, a-h, 17, 18, 19, 29 a-c, 31, 37, 41, 43, 45, 47, 49, 51, 53, 59 a)



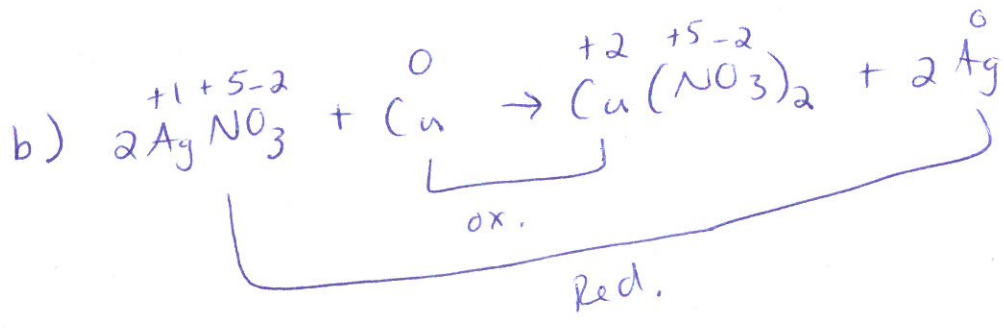
*Red. Pot. on p. 845!

17)

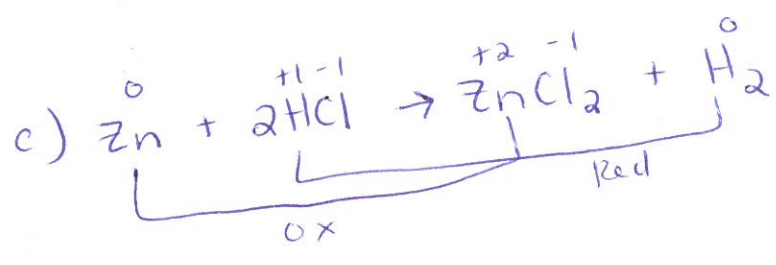
a)



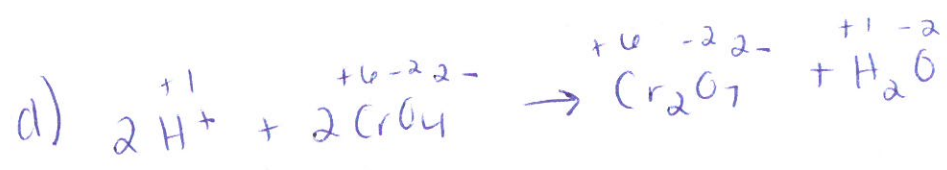
ox = C
 Red = H
 OA = CH₄ & H₂O
 RA = CH₄



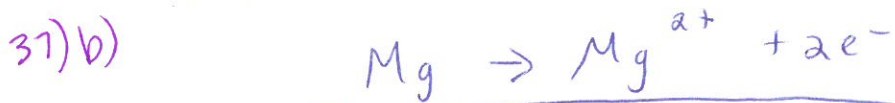
ox = Cu
 Red = Ag
 OA = AgNO₃
 RA = Cu



ox = Zn
 Red = H
 OA = HCl
 RA = Zn



Non Redox



cathode: Cu electrode; Cu^{2+} in solution

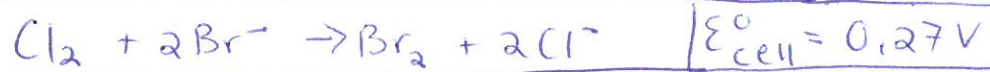
anode: Mg electrode; Mg^{2+} in solution

39) To get \mathcal{E}° for the overall cell rxn, add the standard reduction potential to the standard oxidation potential ($\mathcal{E}_{\text{cell}}^\circ = \mathcal{E}_{\text{red}}^\circ + \mathcal{E}_{\text{ox}}^\circ$). Use Table 17.1. Remember $\mathcal{E}_{\text{ox}}^\circ = -\mathcal{E}_{\text{red}}^\circ$ and standard potentials are not multiplied by the integer used to balance the rxn.

(37) a) $\mathcal{E}_{\text{cell}}^\circ = \mathcal{E}_{\text{Cl}_2 \rightarrow \text{Cl}^-}^\circ + \mathcal{E}_{\text{Cr}^{3+} \rightarrow \text{Cr}_2\text{O}_7^{2-}}^\circ = 1.36\text{V} + (-1.33\text{V}) = \boxed{0.03\text{V}}$

(37) b) $\mathcal{E}_{\text{cell}}^\circ = \mathcal{E}_{\text{Cu}^{2+} \rightarrow \text{Cu}}^\circ + \mathcal{E}_{\text{Mg} \rightarrow \text{Mg}^{2+}}^\circ = 0.34\text{V} + 2.37\text{V} = \boxed{2.71\text{V}}$

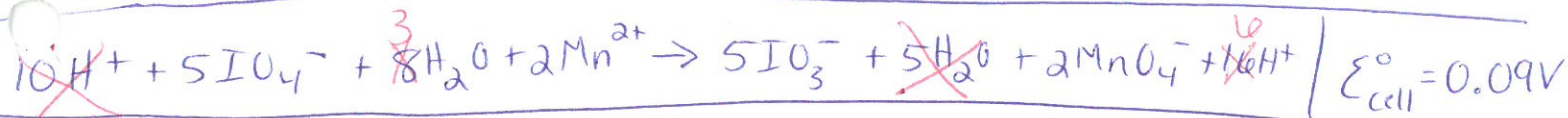
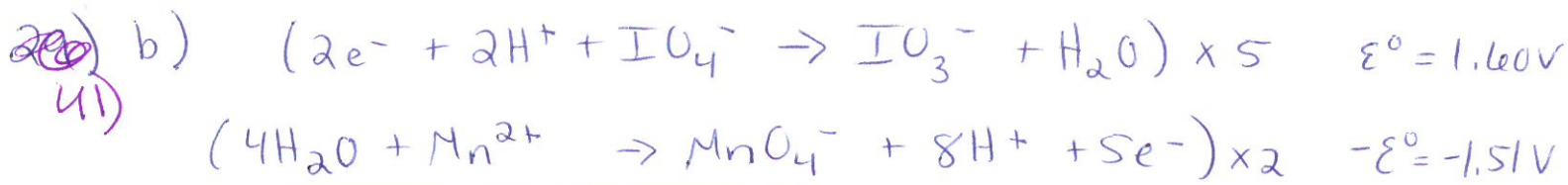
41) $\frac{1}{2}$ rxn w/ largest standard red. potential = cathode, $\frac{1}{2}$ rxn w/ smallest red. potential is reversed = anode. This will give a spontaneous rxn (+ overall potential).



Pt electrode

cathode: $\text{Cl}_2(\text{g})$ bubbled in, Cl^- in solution

anode: Pt electrode, Br_2 & Br^- in solution



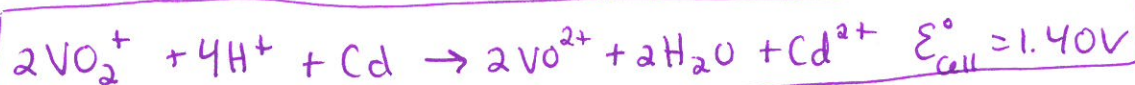
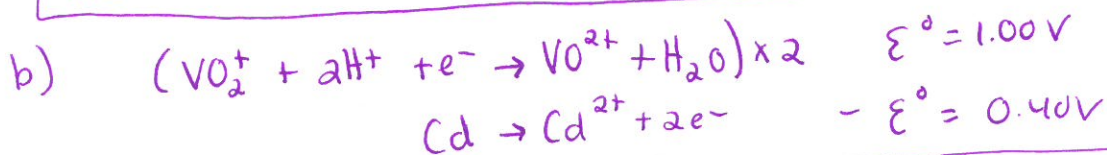
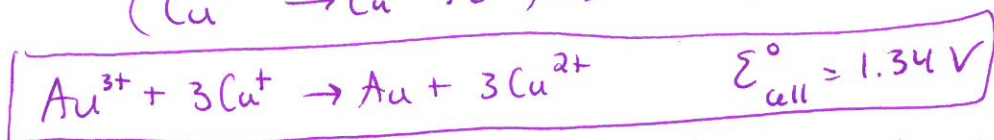
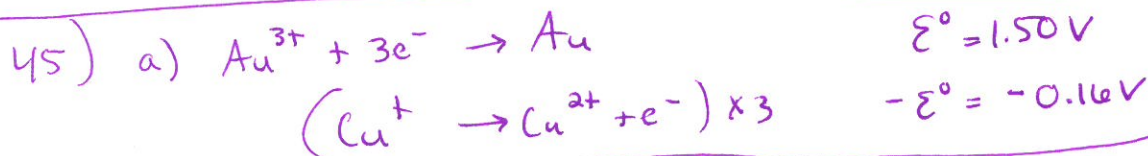
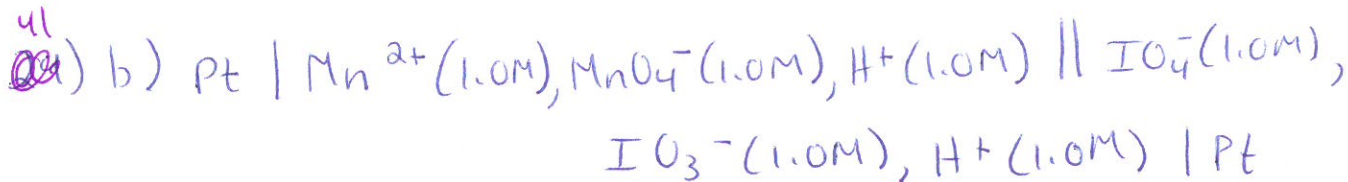
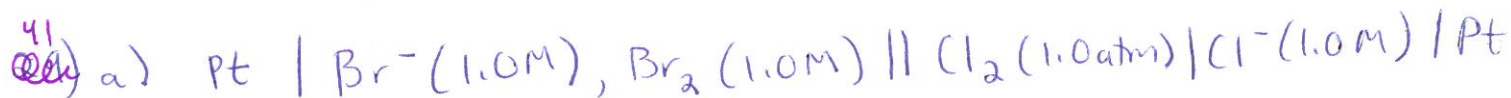
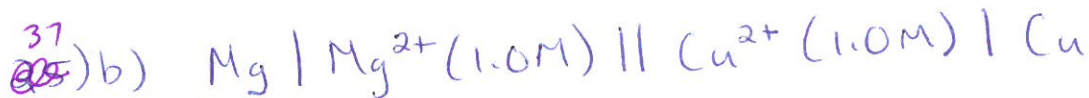
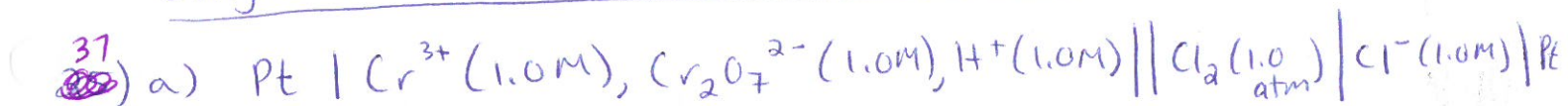
simplifies!

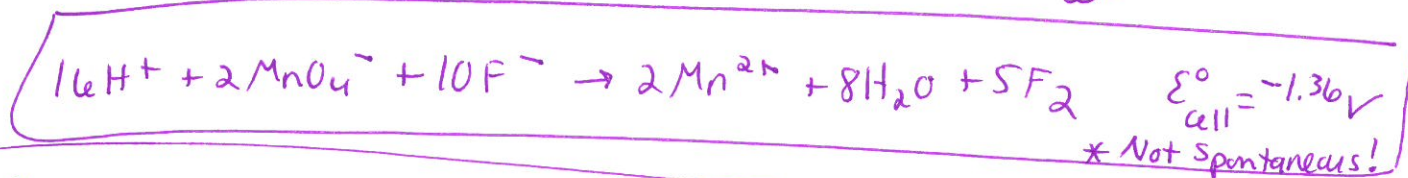
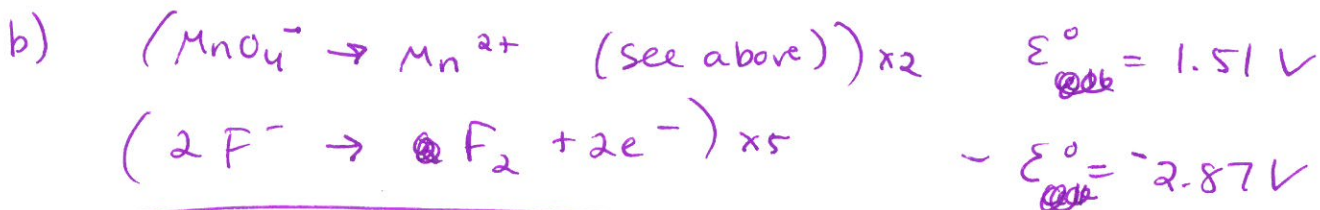
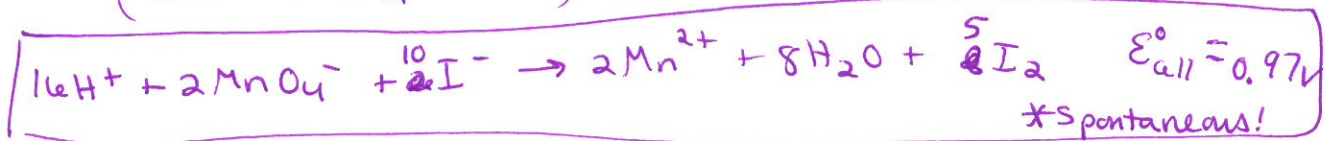
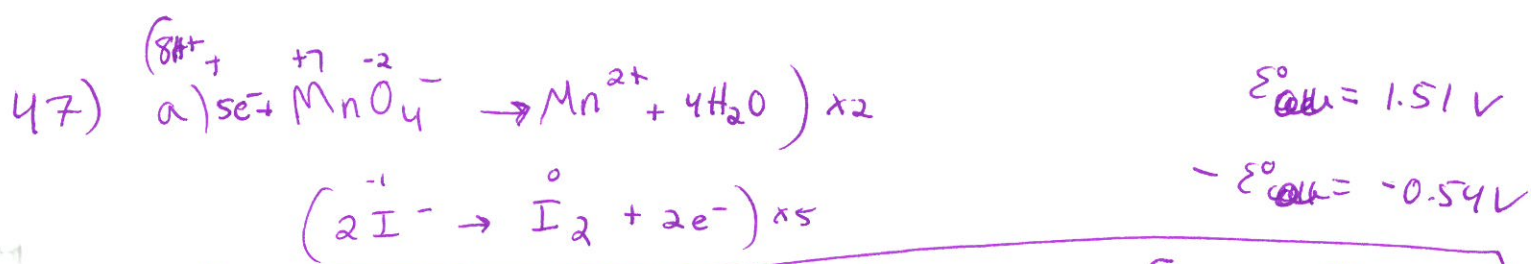


Cathode: Pt electrode; IO_4^- , IO_3^- & H_2SO_4 (H^+) in Soln.

Anode: Pt electrode; Mn^{2+} , MnO_4^- & H_2SO_4 in soln.

43) anode 1st, Double line to separate, electrodes on ends, single line shows phase Δ , include conc.





$\epsilon^{\circ} = 1.36 V$



$-\epsilon^{\circ} = -0.954 V$



$\epsilon_{cell}^{\circ} = 0.41 V = 0.41 \frac{J}{C}$

$\Delta G^{\circ} = -nF\epsilon_{cell}^{\circ} = -(2 \text{ mol } e^-)(96,485 \frac{C}{\text{mol } e^-})(0.41 \frac{J}{C})$

$\Delta G^{\circ} = -7.9 \times 10^4 J = \boxed{-79 \text{ kJ}}$

51) Since cells are @ standard conditions,

$w_{max} = \Delta G = \Delta G^{\circ} = -nF\epsilon_{cell}^{\circ}$

(45) a) $w_{max} = -(3 \text{ mol } e^-)(96,485 \frac{C}{\text{mol } e^-})(1.34 \frac{J}{C})$

$w_{max} = -3.88 \times 10^5 J = \boxed{-388 \text{ kJ}}$

(45) b) $w_{max} = -(2 \text{ mol } e^-)(96,485 \frac{C}{\text{mol } e^-})(1.40 \frac{J}{C})$

$w_{max} = -2.70 \times 10^5 J = \boxed{-270. \text{ kJ}}$



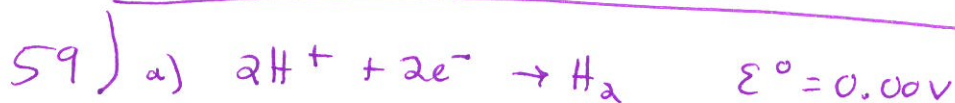
$$\Delta G^\circ = \sum n \Delta G^\circ_{\text{prod.}} - \sum n \Delta G^\circ_{\text{react.}}$$

$$\Delta G^\circ = 2(-157) - [2(-237)] = 160. \text{ kJ}$$

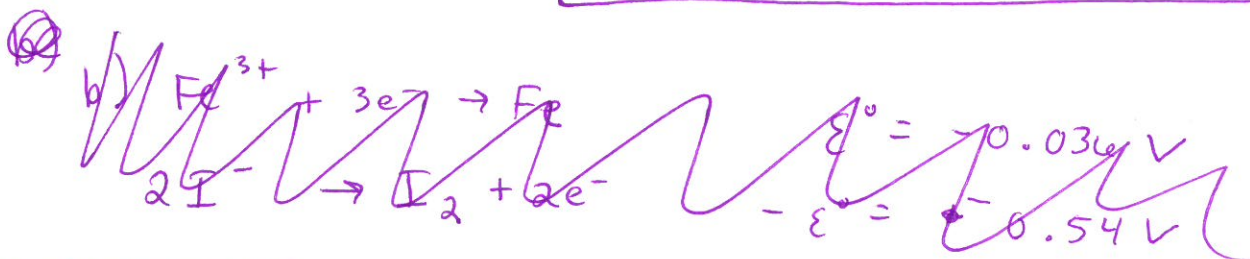
$$\Delta G^\circ = -nF\varepsilon^\circ, \quad \varepsilon^\circ = \frac{-\Delta G^\circ}{nF} = \frac{-1.60 \times 10^5 \text{ J}}{(2 \text{ mole e}^-)(96,485 \frac{\text{C}}{\text{mole e}^-})}$$

$$\varepsilon^\circ = -0.829 \frac{\text{J}}{\text{C}} = -0.829 \text{ V}$$

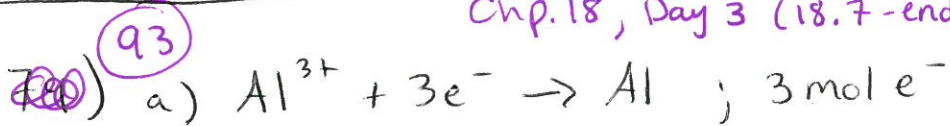
This agrees w/ value in Table 18.



$\varepsilon^\circ_{\text{cell}} = -0.34 \text{ V}$. No H^+ cannot oxidize Cu to Cu^{2+} @ stand. conditions b/c $\varepsilon^\circ_{\text{cell}} < 0$.



~~cell~~



$$\frac{1.0 \times 10^3 \text{ g Al}}{26.98 \text{ g Al}} \times \frac{1 \text{ mol Al}}{1 \text{ mol Al}} \times \frac{3 \text{ mole } e^-}{1 \text{ mol Al}} \times \frac{96,485 \text{ C}}{1 \text{ mol } e^-} \times \frac{1 \text{ s}}{100.0 \text{ C}}$$

= $1.07 \times 10^5 \text{ s}$ = 30. hours

b) $\frac{1.0 \text{ g Ni}}{58.69 \text{ g Ni}} \times \frac{1 \text{ mol Ni}}{1 \text{ mol Ni}} \times \frac{2 \text{ mole } e^-}{1 \text{ mol Ni}} \times \frac{96,485 \text{ C}}{1 \text{ mol } e^-} \times \frac{1 \text{ s}}{100.0 \text{ C}}$ = 33 sec.

c) $\frac{5.0 \text{ mol Ag}}{1 \text{ mol Ag}} \times \frac{1 \text{ mole } e^-}{1 \text{ mole } e^-} \times \frac{96,485 \text{ C}}{1 \text{ mol } e^-} \times \frac{1 \text{ s}}{100.0 \text{ C}}$ = $4.8 \times 10^3 \text{ s}$
 = 1.3 hours

95) $15 \text{ A} = \frac{15 \text{ C}}{\text{s}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{60 \text{ min}}{1 \text{ h}} = 5.4 \times 10^4 \text{ C}$ of charge passed in 1 hour,

a) $\frac{5.4 \times 10^4 \text{ C}}{96,485 \text{ C}} \times \frac{1 \text{ mole } e^-}{1 \text{ mole } e^-} \times \frac{1 \text{ mol Co}}{2 \text{ mole } e^-} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}}$ = 16 g Co

b) $\frac{5.4 \times 10^4 \text{ C}}{96,485 \text{ C}} \times \frac{1 \text{ mole } e^-}{1 \text{ mole } e^-} \times \frac{1 \text{ mol Hf}}{4 \text{ mole } e^-} \times \frac{178.5 \text{ g Hf}}{1 \text{ mol Hf}}$ = 25 g Hf

97) $\frac{74.1 \text{ s}}{1 \text{ s}} \times \frac{2.00 \text{ C}}{96,485 \text{ C}} \times \frac{1 \text{ mole } e^-}{1 \text{ mole } e^-} \times \frac{1 \text{ mol M}}{3 \text{ mole } e^-} = 0.00051 \text{ mol M}$

$\frac{0.107 \text{ g}}{0.00051 \text{ mol}} =$ ~~209~~ 209 g/mol \rightarrow Bi

85) F_2 is produced at the anode: $2F^- \rightarrow F_2 + 2e^-$

99

$$\frac{2.00 \text{ h} \mid 60 \text{ min.} \mid 60 \text{ sec.} \mid 10.0 \text{ C} \mid 1 \text{ mole } e^-}{1 \text{ hr.} \mid 1 \text{ min.} \mid 1 \text{ sec.} \mid 96,485 \text{ C}} = 0.746 \text{ mol } e^-$$

$$\frac{0.746 \text{ mol } e^- \mid 1 \text{ mol } F_2}{2 \text{ mol } e^-} = 0.373 \text{ mol } F_2$$

$$PV = nRT, \quad V = \frac{nRT}{P} = \frac{(0.373 \text{ mol})(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}})(298 \text{ K})}{1.00 \text{ atm}}$$

$$\boxed{V = 9.12 \text{ L } F_2}$$

K is produced at the cathode: $K^+ + e^- \rightarrow K$

~~0.746~~
$$\frac{0.746 \text{ mole } e^- \mid 1 \text{ mol } K \mid 39.10 \text{ g } K}{\text{mole } e^- \mid \text{mol } K} = \boxed{29.2 \text{ g } K}$$

87)

~~$$\frac{150. \times 10^3 \text{ g } C_6H_8N_2 \mid 1 \text{ hr} \mid 1 \text{ min.} \mid 1 \text{ mol } C_6H_8N_2 \mid 2 \text{ mole } e^- \mid 96,485 \text{ C}}{1 \text{ hr} \mid 60 \text{ min.} \mid 60 \text{ sec.} \mid 108.14 \text{ g } C_6H_8N_2 \mid \text{mol } C_6H_8N_2 \mid \text{mole } e^-}$$~~
~~$$= \boxed{7.44 \times 10^4 \text{ C/s or a current of } 7.44 \times 10^4 \text{ A}}$$~~

89)

103

$$\frac{2.30 \text{ min} \mid 60 \text{ s} \mid 2.00 \text{ C} \mid 1 \text{ mole } e^- \mid 1 \text{ mol } Ag}{1 \text{ min.} \mid \text{s} \mid 96,485 \text{ C} \mid \text{mole } e^-} = 2.86 \times 10^{-3} \text{ mol } Ag$$

$$[Ag^+] = \frac{2.86 \times 10^{-3} \text{ mol } Ag^+}{0.250 \text{ L}} = \boxed{1.14 \times 10^{-2} \text{ M}}$$



Au will plate out 1st (most (+) red. pot.), followed by Ag, Ni, ~~Cd~~ and Cd last (most (-) potential).

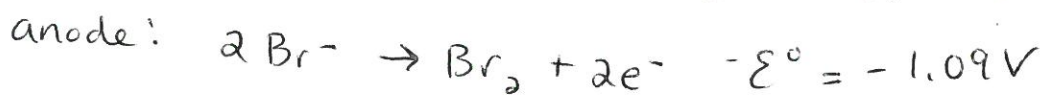
93) Reduction occurs @ cathode, oxidation @ anode.

109) First \rightarrow species present. Then red. potentials.

Cathode = rxn w/ most (+) red. pot.

Anode = rxn w/ most (+) ox. pot.

a) Ni^{2+} & Br^- ; Ni^{2+} gets reduced $\Rightarrow Ni$
 Br^- gets oxidized $\Rightarrow Br_2$



b) Al^{3+} , F^- , Al^{3+} gets reduced to Al
 F^- gets oxidized to F_2

