

$$\frac{1.0 \times 10^3 \text{ g Al}}{1} \left| \frac{1 \text{ mole Al}}{26.98 \text{ g Al}} \right| \left| \frac{3 \text{ mole } e^-}{1 \text{ mole Al}} \right| \left| \frac{96,500 \text{ C}}{1 \text{ mole } e^-} \right| \left| \frac{1 \text{ s}}{100.0 \text{ C}} \right| =$$

$$1.1 \times 10^5 \text{ s} = 30 \text{ hr}$$



$$\frac{1.0 \text{ g Ni}}{1} \left| \frac{1 \text{ mole Ni}}{58.69 \text{ g Ni}} \right| \left| \frac{2 \text{ mole } e^-}{1 \text{ mole Ni}} \right| \left| \frac{96,500 \text{ C}}{1 \text{ mole } e^-} \right| \left| \frac{1 \text{ s}}{100.0 \text{ C}} \right| = 33 \text{ s}$$

79. a) $\frac{15 \text{ C}}{1 \text{ s}} \left| \frac{60 \text{ s}}{1 \text{ min}} \right| \left| \frac{60 \text{ min}}{1 \text{ hr}} \right| = 5.4 \times 10^4 \frac{\text{C}}{\text{hr}}$

$$\frac{5.4 \text{ C}}{96,500 \text{ C}} \left| \frac{1 \text{ mole } e^-}{2 \text{ mole } e^-} \right| \left| \frac{1 \text{ mole Co}}{58.93 \text{ g Co}} \right| = 16 \text{ g Co}$$

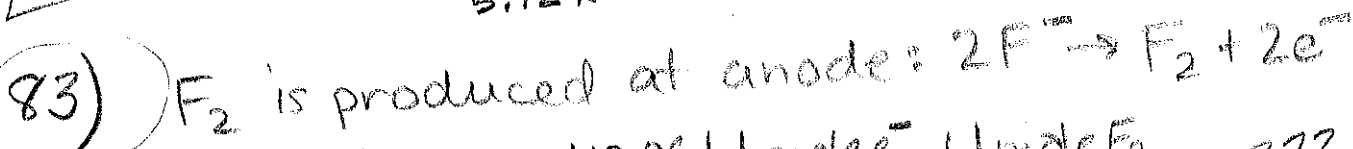
b) $\frac{5.4 \text{ C}}{96,500 \text{ C}} \left| \frac{1 \text{ mole } e^-}{2 \text{ mole } e^-} \right| \left| \frac{1 \text{ mole I}_2}{253.8 \text{ g I}_2} \right| = 71 \text{ g I}_2$



81) $\frac{74.1 \text{ s}}{1} \left| \frac{2.00 \text{ C}}{1 \text{ s}} \right| \left| \frac{1 \text{ mole } e^-}{96,500 \text{ C}} \right| \left| \frac{1 \text{ mole M}}{3 \text{ mole } e^-} \right| = 5.12 \times 10^{-4} \text{ moles}$

$$MM = \frac{.107 \text{ g M}}{5.12 \times 10^{-4} \text{ moles}} = 209$$

Bi



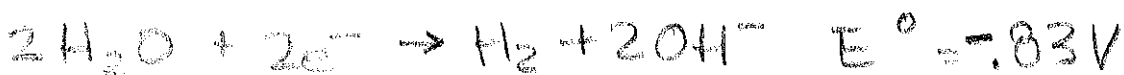
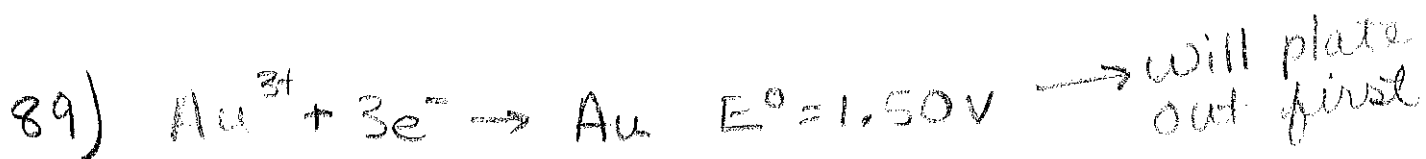
$$\frac{2.00 \text{ hr}}{1} \left| \frac{3600 \text{ s}}{1 \text{ hr}} \right| \left| \frac{10.0 \text{ C}}{1 \text{ s}} \right| \left| \frac{1 \text{ mole } e^-}{96,500 \text{ C}} \right| \left| \frac{1 \text{ mole F}_2}{2 \text{ mole } e^-} \right| = .373 \text{ moles}$$

$$V = \frac{nRT}{P} = \frac{(.373)(.0821)(298)}{1.00} = 9.13 \text{ L}$$

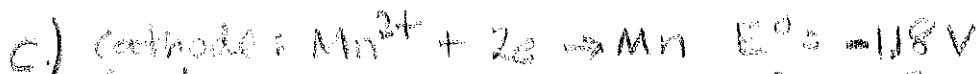
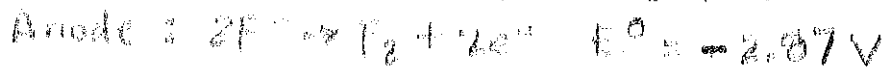
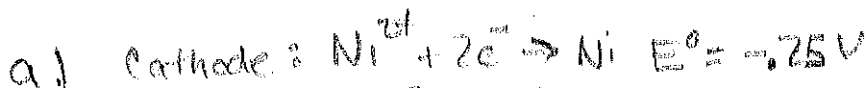
87)

$$\frac{2.30 \text{ min} \times 60 \text{ s} \times 2.00 \text{ C}}{1 \text{ min} \times 1 \text{ s}} \times \frac{1 \text{ mole } e^-}{96,500 \text{ C}} \times \frac{1 \text{ mole Ag}}{1 \text{ mole } e^-} =$$

$$M = \frac{\quad}{.250 \text{ L}} =$$



91)



will not interfere w/ plating process

82) Alkaline earth for +2 ions

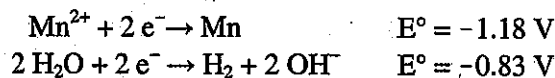
$$\frac{748 \text{ s} \times 5.00 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mole } e^-}{96,500 \text{ C}} \times \frac{1 \text{ mole M}}{2 \text{ mole } e^-} =$$

moles M

$$MM = \frac{.471 \text{ g M}}{\text{moles M}} =$$

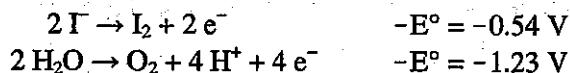
F^- and H_2O can be oxidized. The oxidation potentials are $-E^\circ = -2.87$ V for F^- and $-E^\circ = -1.23$ V for H_2O (assuming standard conditions). From the potentials, we would predict H_2O to be oxidized at the anode ($2 H_2O \rightarrow O_2 + 4 H^+ + 4 e^-$).

- c. Species present: Mn^{2+} , Γ^- and H_2O ; Mn^{2+} and H_2O can be reduced. The possible cathode reactions are:



Reduction of H_2O will occur at the cathode since $E_{H_2O}^\circ$ is most positive.

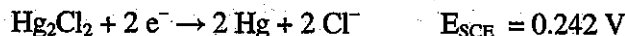
Γ^- and H_2O can be oxidized. The possible anode reactions are:



Oxidation of Γ^- will occur at the anode since $-E_{\Gamma^-}^\circ$ is most positive.

Additional Exercises

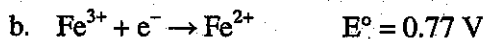
93. The half-reaction for the SCE is:



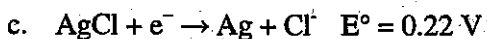
For a spontaneous reaction to occur, E_{cell} must be positive. Using the standard reduction potentials in Table 17.1 and the given the SCE potential, deduce which combination will produce a positive overall cell potential.



$$E_{cell} = 0.34 - 0.242 = 0.10 \text{ V}; \text{ SCE is the anode.}$$



$$E_{cell} = 0.77 - 0.242 = 0.53 \text{ V}; \text{ SCE is the anode.}$$



$$E_{cell} = 0.242 - 0.22 = 0.02 \text{ V}; \text{ SCE is the cathode.}$$



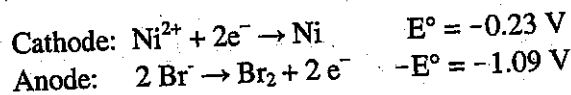
$$E_{cell} = 0.242 + 1.66 = 1.90 \text{ V}; \text{ SCE is the cathode.}$$



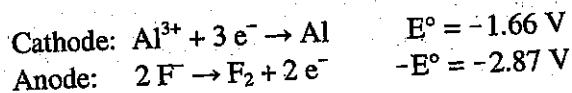
$$E_{cell} = 0.242 + 0.23 = 0.47 \text{ V}; \text{ SCE is the cathode.}$$

91. Reduction occurs at the cathode, and oxidation occurs at the anode. First, determine all the species present, then look up pertinent reduction and/or oxidation potentials in Table 17.1 for all these species. The cathode reaction will be the reaction with the most positive reduction potential, and the anode reaction will be the reaction with the most positive oxidation potential.

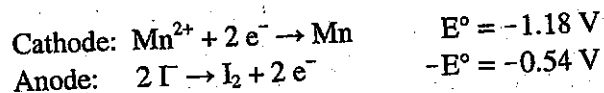
- a. Species present: Ni^{2+} and Br^- ; Ni^{2+} can be reduced to Ni, and Br^- can be oxidized to Br_2 (from Table 17.1). The reactions are:



- b. Species present: Al^{3+} and F^- ; Al^{3+} can be reduced, and F^- can be oxidized. The reactions are:

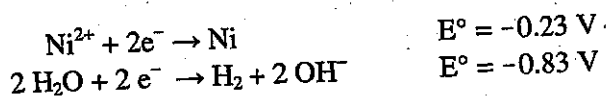


- c. Species present: Mn^{2+} and I^- ; Mn^{2+} can be reduced, and I^- can be oxidized. The reactions are:



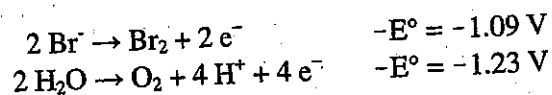
92. These are all in aqueous solutions, so we must also consider the reduction and oxidation of H_2O in addition to the potential redox reactions of the ions present. For the cathode reaction, the species with the most positive reduction potential will be reduced, and for the anode reaction, the species with the most positive oxidation potential will be oxidized.

- a. Species present: Ni^{2+} , Br^- and H_2O . Possible cathode reactions are:



Because it is easier to reduce Ni^{2+} than H_2O (assuming standard conditions), Ni^{2+} will be reduced by the above cathode reaction.

Possible anode reactions are:



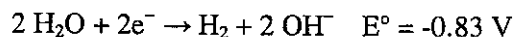
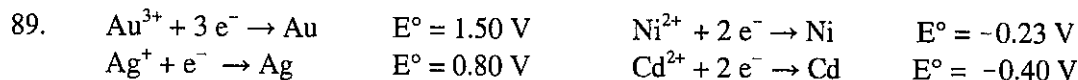
Because Br^- is easier to oxidize than H_2O (assuming standard conditions), Br^- will be oxidized by the above anode reaction.

- b. Species present: Al^{3+} , F^- and H_2O ; Al^{3+} and H_2O can be reduced. The reduction potentials are $E^\circ = -1.66 \text{ V}$ for Al^{3+} and $E^\circ = -0.83 \text{ V}$ for H_2O (assuming standard conditions). H_2O will be reduced at the cathode ($2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$).

88. $0.50 \text{ L} \times 0.010 \text{ mol Pt}^{4+}/\text{L} = 5.0 \times 10^{-3} \text{ mol Pt}^{4+}$

To plate out 99% of the Pt^{4+} , we will produce $0.99 \times 5.0 \times 10^{-3} \text{ mol Pt}$.

$$0.99 \times 5.0 \times 10^{-3} \text{ mol Pt} \times \frac{4 \text{ mol e}^-}{\text{mol Pt}} \times \frac{96,485 \text{ C}}{\text{mol e}^-} \times \frac{1 \text{ s}}{4.00 \text{ C}} \times \frac{1 \text{ mol Ag}}{\text{mol e}^-} = 480 \text{ s}$$



Au(s) will plate out first since it has the most positive reduction potential, followed by Ag(s) , which is followed by Ni(s) , and finally Cd(s) will plate out last since it has the most negative reduction potential of the metals listed. Water will not interfere with the plating process.

90. To begin plating out Pd:

$$E = 0.62 - \frac{0.0591}{2} \log \frac{[\text{Cl}^-]^4}{[\text{PdCl}_4^{2-}]} = 0.62 - \frac{0.0591}{2} \log \frac{(1.0)^4}{0.020}$$

$$E = 0.62 \text{ V} - 0.050 \text{ V} = 0.57 \text{ V}$$

When 99% of Pd has plated out, $[\text{PdCl}_4^{2-}] = \frac{0.020}{100} = 0.00020 \text{ M}$.

$$E = 0.62 - \frac{0.0591}{2} \log \frac{(1.0)^4}{2.0 \times 10^{-4}} = 0.62 \text{ V} - 0.11 \text{ V} = 0.51 \text{ V}$$

To begin Pt plating: $E = 0.73 \text{ V} - \frac{0.0591}{2} \log \frac{(1.0)^4}{0.020} = 0.73 - 0.050 = 0.68 \text{ V}$

When 99% of Pt plated: $E = 0.73 - \frac{0.0591}{2} \log \frac{(1.0)^4}{2.0 \times 10^{-4}} = 0.73 - 0.11 = 0.62 \text{ V}$

To begin Ir plating: $E = 0.77 \text{ V} - \frac{0.0591}{3} \log \frac{(1.0)^4}{0.020} = 0.77 - 0.033 = 0.74 \text{ V}$

When 99% of Ir plated: $E = 0.77 - \frac{0.0591}{3} \log \frac{(1.0)^4}{2.0 \times 10^{-4}} = 0.77 - 0.073 = 0.70 \text{ V}$

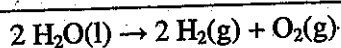
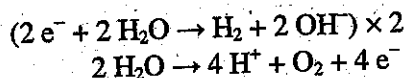
Yes, because the range of potentials for plating out each metal do not overlap, we should be able to separate the three metals. The exact potential to apply depends on the oxidation reaction. The order of plating will be Ir(s) first, followed by Pt(s) and finally Pd(s) as the potential is gradually increased.

$$\frac{(0.373 \text{ mol})(0.08206 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol})(298 \text{ K})}{1.00 \text{ atm}} = 9.12 \text{ L F}_2$$

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

$$0.746 \text{ mol e}^- \times \frac{1 \text{ mol K}}{\text{mol e}^-} \times \frac{39.10 \text{ g K}}{\text{mol K}} = 29.2 \text{ g K}$$

84. The half-reactions for the electrolysis of water are:



Note: $4 \text{ H}^+ + 4 \text{ OH}^- \rightarrow 4 \text{ H}_2\text{O}$ and $n = 4$ for this reaction as it is written.

$$15.0 \text{ min} \times \frac{60 \text{ s}}{\text{min}} \times \frac{2.50 \text{ C}}{\text{s}} \times \frac{1 \text{ mol e}^-}{96,485 \text{ C}} \times \frac{2 \text{ mol H}_2}{4 \text{ mol e}^-} = 1.17 \times 10^{-2} \text{ mol H}_2$$

At STP, 1 mole of an ideal gas occupies a volume of 22.42 L (see Chapter 5 of the text).

$$1.17 \times 10^{-2} \text{ mol H}_2 \times \frac{22.42 \text{ L}}{\text{mol H}_2} = 0.262 \text{ L} = 262 \text{ mL H}_2$$

$$1.17 \times 10^{-2} \text{ mol H}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \times \frac{22.42 \text{ L}}{\text{mol O}_2} = 0.131 \text{ L} = 131 \text{ mL O}_2$$

$$\begin{aligned} 85. \quad & \frac{150. \times 10^3 \text{ g C}_6\text{H}_8\text{N}_2}{\text{h}} \times \frac{1 \text{ h}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}} \times \frac{1 \text{ mol C}_6\text{H}_8\text{N}_2}{108.14 \text{ g C}_6\text{H}_8\text{N}_2} \times \frac{2 \text{ mol e}^-}{\text{mol C}_6\text{H}_8\text{N}_2} \times \frac{96,485 \text{ C}}{\text{mol e}^-} \\ & = 7.44 \times 10^4 \text{ C/s or a current of } 7.44 \times 10^4 \text{ A} \end{aligned}$$

86. $\text{Al}^{3+} + 3 \text{ e}^- \rightarrow \text{Al}$; 3 mol e^- are needed to produce Al from Al^{3+}

$$2000 \text{ lb Al} \times \frac{453.6 \text{ g}}{\text{lb}} \times \frac{1 \text{ mol Al}}{26.98 \text{ g}} \times \frac{3 \text{ mol e}^-}{\text{mol Al}} \times \frac{96,485 \text{ C}}{\text{mol e}^-} = 1 \times 10^{10} \text{ C of electricity needed}$$

$$\frac{1 \times 10^{10} \text{ C}}{24 \text{ h}} \times \frac{1 \text{ h}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}} = 1 \times 10^5 \text{ C/s} = 1 \times 10^5 \text{ A}$$

$$87. \quad 2.30 \text{ min} \times \frac{60 \text{ s}}{\text{min}} = 138 \text{ s}; 138 \text{ s} \times \frac{2.00 \text{ C}}{\text{s}} \times \frac{1 \text{ mol e}^-}{96,485 \text{ C}} \times \frac{1 \text{ mol Ag}}{\text{mol e}^-} = 2.86 \times 10^{-3} \text{ mol Ag}$$

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

79. $15 \text{ A} = \frac{15 \text{ C}}{\text{s}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{60 \text{ min}}{\text{h}} = 5.4 \times 10^4 \text{ C}$ of charge passed in 1 hour

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

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

c. $2 \text{ I}^- \rightarrow \text{I}_2 + 2 e^-$; $5.4 \times 10^4 \text{ C} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} \times \frac{1 \text{ mol I}_2}{2 \text{ mol } e^-} \times \frac{253.8 \text{ g I}_2}{\text{mol I}_2} = 71 \text{ g I}_2$

d. $\text{CrO}_3(\text{l}) \rightarrow \text{Cr}^{6+} + 3 \text{ O}^{2-}$; 6 mol e^- are needed to produce 1 mol Cr from molten CrO_3 .

$$5.4 \times 10^4 \text{ C} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} \times \frac{1 \text{ mol Cr}}{6 \text{ mol } e^-} \times \frac{52.00 \text{ g Cr}}{\text{mol Cr}} = 4.9 \text{ g Cr}$$

80. Al is in the +3 oxidation in Al_2O_3 , so 3 mol e^- are needed to convert Al^{3+} into Al(s).

$$2.00 \text{ h} \times \frac{60 \text{ min}}{\text{h}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{1.00 \times 10^6 \text{ C}}{\text{s}} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} \times \frac{1 \text{ mol Al}}{3 \text{ mol } e^-} \times \frac{26.98 \text{ g Al}}{\text{mol Al}} = 6.71 \times 10^5 \text{ g}$$

81. $74.1 \text{ s} \times \frac{2.00 \text{ C}}{\text{s}} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} \times \frac{1 \text{ mol M}}{3 \text{ mol } e^-} = 5.12 \times 10^{-4} \text{ mol M}$ where M = unknown metal

$$\text{Molar mass} = \frac{0.107 \text{ g M}}{5.12 \times 10^{-4} \text{ mol M}} = \frac{209 \text{ g}}{\text{mol}}; \text{ The element is bismuth.}$$

82. Alkaline earth metals form +2 ions, so 2 mol of e^- are transferred to form the metal, M.

$$\text{mol M} = 748 \text{ s} \times \frac{5.00 \text{ C}}{\text{s}} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} \times \frac{1 \text{ mol M}}{2 \text{ mol } e^-} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} = 1.94 \times 10^{-2} \text{ mol M}$$

$$\text{molar mass of M} = \frac{0.471 \text{ g M}}{1.94 \times 10^{-2} \text{ mol M}} = 24.3 \text{ g/mol}; \text{ MgCl}_2 \text{ was electrolyzed.}$$

83. F_2 is produced at the anode: $2 \text{ F}^- \rightarrow \text{F}_2 + 2 e^-$

$$2.00 \text{ h} \times \frac{60 \text{ min}}{\text{h}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{10.0 \text{ C}}{\text{s}} \times \frac{1 \text{ mol } e^-}{96,485 \text{ C}} = 0.746 \text{ mol } e^-$$

$$0.746 \text{ mol } e^- \times \frac{1 \text{ mol F}_2}{2 \text{ mol } e^-} = 0.373 \text{ mol F}_2; \text{ PV} = n\text{RT}, \text{ V} = \frac{n\text{RT}}{\text{P}}$$

Handwritten notes on the right margin:

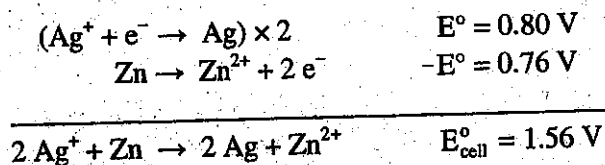
$$\frac{15 \text{ A} \times 3600 \text{ s}}{1 \text{ h}} \times \frac{1 \text{ mol } e^-}{96485 \text{ C}} \times \frac{1 \text{ mol I}_2}{2 \text{ mol } e^-} \times \frac{253.8 \text{ g I}_2}{\text{mol I}_2}$$

Handwritten notes on the right margin:

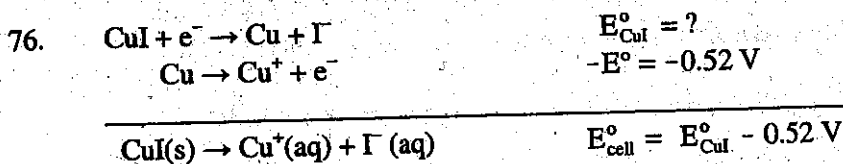
$$\frac{74.1 \text{ s} \times 2.00 \text{ C}}{96485 \text{ C}} \times \frac{1 \text{ mol M}}{3 \text{ mol } e^-}$$

$$\log K = \frac{nE^\circ}{0.0591} = \frac{3(0.41)}{0.0591} = 20.81, K = 10^{20.81} = 6.5 \times 10^{20}$$

75. NO_3^- is a spectator ion. The reaction that occurs is Ag^+ reacting with Zn.



$$\log K = \frac{nE^\circ}{0.0591} = \frac{2(1.56)}{0.0591}, K = 10^{52.792} = 6.19 \times 10^{52}$$



For this overall reaction, $K = K_{\text{sp}} = 1.1 \times 10^{-12}$:

$$E^\circ_{\text{cell}} = \frac{0.0591}{n} \log K_{\text{sp}} = \frac{0.0591}{1} \log (1.1 \times 10^{-12}) = -0.71 \text{ V}$$

$$E^\circ_{\text{cell}} = -0.71 \text{ V} = E^\circ_{\text{CuI}} - 0.52, E^\circ_{\text{CuI}} = -0.19 \text{ V}$$

Electrolysis

77. a. $\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$; 3 mol e^- are needed to produce 1 mol Al from Al^{3+} .

$$1.0 \times 10^3 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol } e^-}{\text{mol Al}} \times \frac{96,485 \text{ C}}{\text{mol } e^-} \times \frac{1 \text{ s}}{100.0 \text{ C}} = 1.07 \times 10^5 \text{ s} = 30. \text{ hours}$$

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

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

78. The oxidation state of bismuth in BiO^+ is +3 because oxygen has a -2 oxidation state in this ion. Therefore, 3 moles of electrons are required to reduce the bismuth in BiO^+ to Bi(s) .

$$10.0 \text{ g Bi} \times \frac{1 \text{ mol Bi}}{209.0 \text{ g Bi}} \times \frac{3 \text{ mol } e^-}{\text{mol Bi}} \times \frac{96,485 \text{ C}}{\text{mol } e^-} \times \frac{1 \text{ s}}{25.0 \text{ C}} = 554 \text{ s} = 9.23 \text{ min}$$