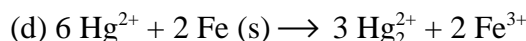
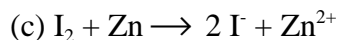
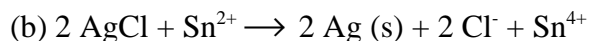
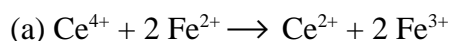


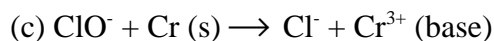
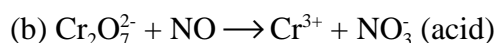
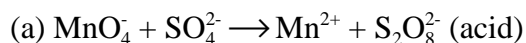
CHE152
Chapter 21 - Electrochemistry

1. Identify the element reduced and the element oxidized in each reaction.

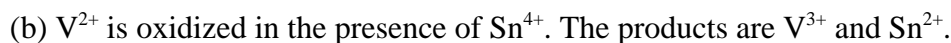
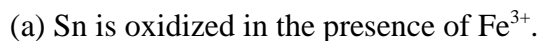


2. In problem 1 above identify the reducing agent and the oxidizing agent.

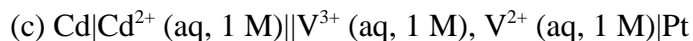
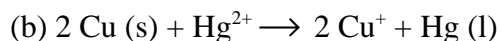
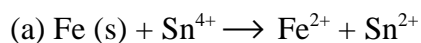
3. Balance the following redox reactions.



4. Diagram the following cells.



5. Determine E_{cell}° for the following.



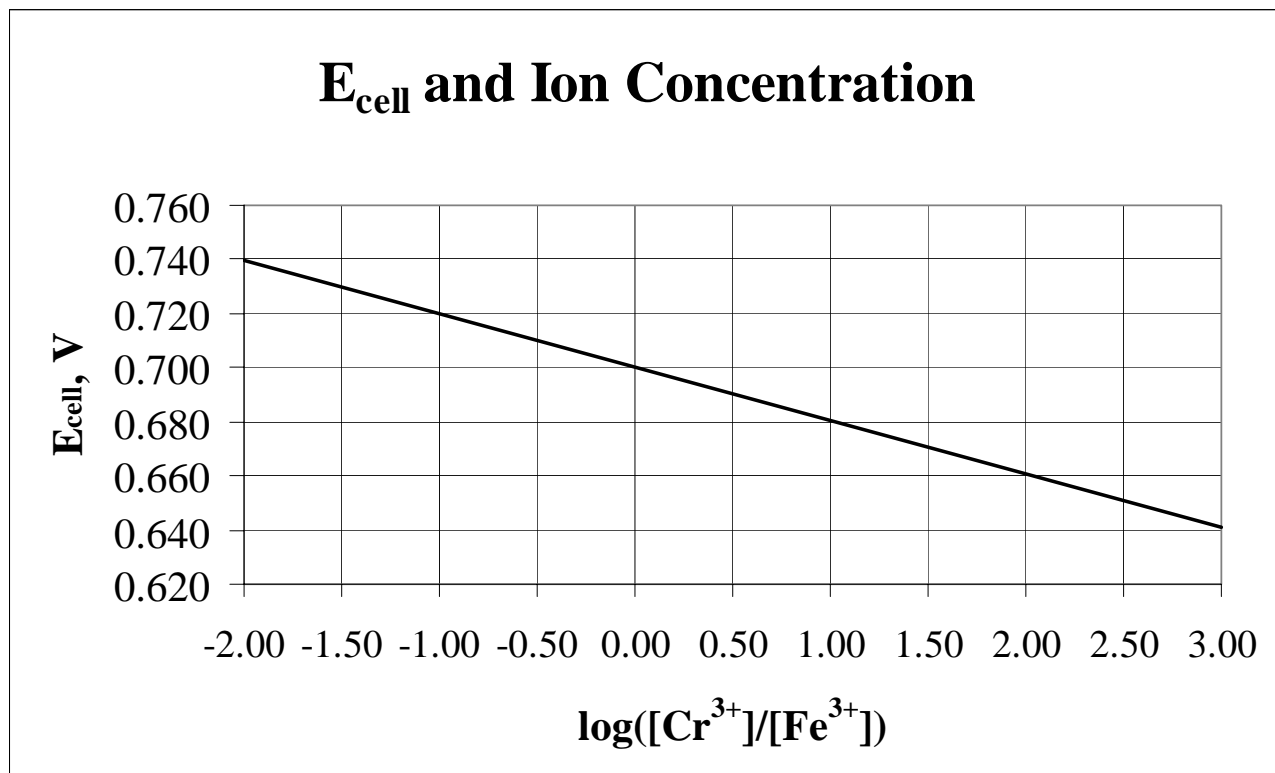
6. Which cell in question 5 above is spontaneous?

7. Determine ΔG° for each cell in question 5 above.

8. Will Ni^{2+} displace Cr in an aqueous solution?

9. Will Mg^{2+} displace Pb metal in an aqueous solution?

10. Can $\text{S}_2\text{O}_8^{2-}$ be employed to oxidize Br^- ?
11. Can $\text{Cr}_2\text{O}_7^{2-}$ be employed to oxidize Mn^{2+} ?
12. Calculate K for each reaction in question 5 at 25°C .



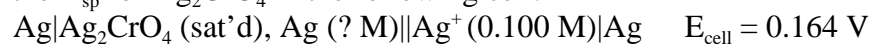
13. Determine the slope and intercept for the above graph. Based on the Nernst equation to what are the slope and intercept equal? $\text{Fe}^{3+} + \text{Cr} \rightarrow \text{Fe} + \text{Cr}^{3+}$.
14. Determine E_{cell} for the following.
- (a) $\text{Fe}(\text{s}) + \text{Sn}^{4+}(\text{aq}, 0.5000 \text{ M}) \rightarrow \text{Fe}^{2+}(\text{aq}, 0.007500 \text{ M}) + \text{Sn}^{2+}(\text{aq}, 0.0001250 \text{ M})$
- (b) $2 \text{Cu}(\text{s}) + \text{Hg}^{2+}(0.0003250 \text{ M}) \rightarrow 2 \text{Cu}^+(0.2500 \text{ M}) + \text{Hg}(\text{l})$
- (c) $\text{Cd}|\text{Cd}^{2+}(\text{aq}, 0.9500 \text{ M})||\text{V}^{3+}(\text{aq}, 0.0002750 \text{ M}), \text{V}^{2+}(\text{aq}, 0.001125 \text{ M})|\text{Pt}$
- (d) $\text{Mg}|\text{Mg}^{2+}(\text{aq}, 0.02500 \text{ M})||\text{Fe}^{3+}(\text{aq}, 0.08500 \text{ M})|\text{Fe}$
- (e) $\text{Sn}|\text{Sn}^{2+}(\text{aq}, 1.2500 \text{ M})||\text{Zn}^{2+}(\text{aq}, 2.2500 \text{ M})|\text{Zn}$
15. Determine E_{cell} for the following cell:



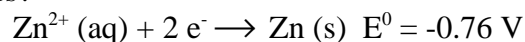
16. What is the minimum $[\text{Mg}^{2+}]$ at which the following cell is able to function as a voltaic cell?



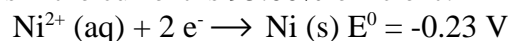
17. Determine the K_{sp} for Ag_2CrO_4 in the following cell.



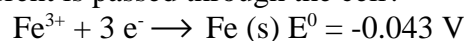
18. What mass of Zn will plate out from an aqueous solution when a current of 30.00 A is passed for 18.00 hours?



19. What mass of Ni will plate out from an aqueous solution when a current of 15.00 A is passed for 12.00 hours if the current is 95.00% efficient?

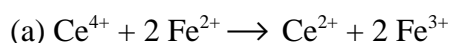


20. How long, in hours, will it take to plate out 125.000 g of Fe from an aqueous solution if a 35.00 amp current is passed through the cell?

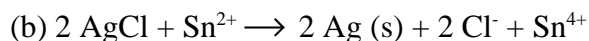


21. Determine how long, in hours, it will take to plate out the iron in question 20 if the current is 85.00% efficient.

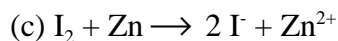
1. Identify the element reduced and the element oxidized in each reaction.



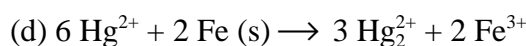
Ce^{4+} is reduced; Fe^{2+} is oxidized



Ag^+ is reduced; Sn^{2+} is oxidized



I_2 is reduced; Zn is oxidized



Hg^{2+} is reduced; Fe is oxidized

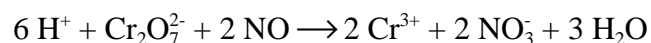
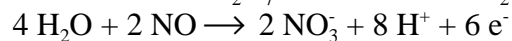
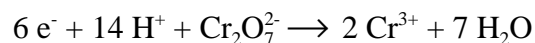
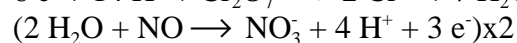
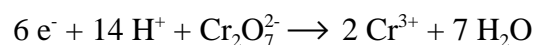
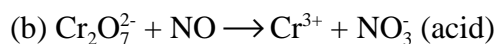
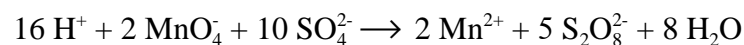
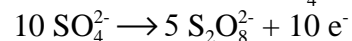
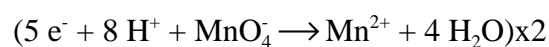
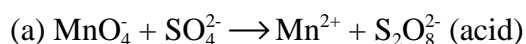
2. (a) Fe^{2+} is the reducing agent; Ce^{4+} is oxidizing agent

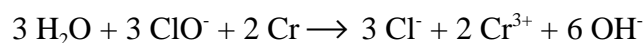
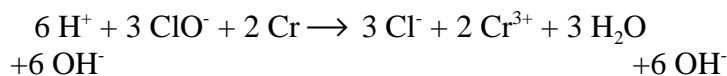
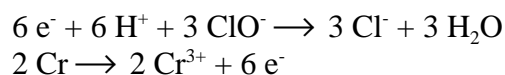
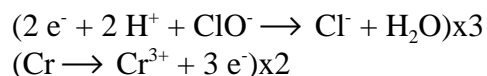
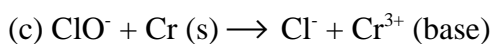
(b) Sn^{2+} is the reducing agent; Ag^+ is the oxidizing agent

(c) Zn is the reducing agent; I_2 is the oxidizing agent

(d) Fe is the reducing agent; Hg^{2+} is the oxidizing agent

3. Balance the following redox reactions.





4. Diagram the following cells.

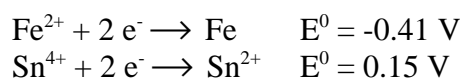
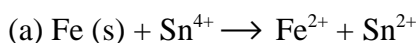
(a) Sn is oxidized in the presence of Fe^{3+} .



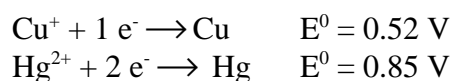
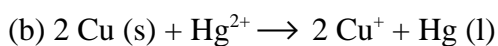
(b) V^{2+} is oxidized in the presence of Sn^{4+} . The products are V^{3+} and Sn^{2+} .



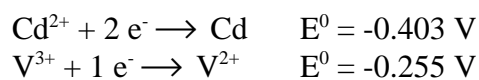
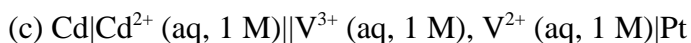
5. Determine E_{cell}° for the following.



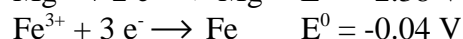
$$E_{\text{cell}} = 0.15 \text{ V} - (-0.41 \text{ V}) = 0.56 \text{ V}$$



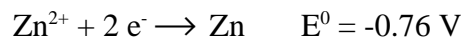
$$E_{\text{cell}} = 0.85 \text{ V} - 0.52 \text{ V} = 0.33 \text{ V}$$



$$E_{\text{cell}} = -0.255 \text{ V} - (-0.403 \text{ V}) = 0.148 \text{ V}$$



$$E_{\text{cell}} = -0.04 \text{ V} - (-2.38 \text{ V}) = 2.34 \text{ V}$$



$$E_{\text{cell}} = -0.76 \text{ V} - (-0.14 \text{ V}) = -0.62 \text{ V}$$

6. Which cell in question 5 above is spontaneous?

For a cell to be spontaneous $E_{\text{cell}} > 0$. Cells a, b, c and d are spontaneous; cell e is not spontaneous.

7. Determine ΔG° for each cell in question 5 above.

$$\Delta G^{\circ} = -nFE_{\text{cell}}^{\circ}$$

$$(a) \Delta G^{\circ} = -2 \text{ mol } e^{-} \times 96,500 \text{ J/V mol } e^{-} \times 0.56 \text{ V} = -1.1 \times 10^5 \text{ J} = -1.1 \times 10^2 \text{ kJ}$$

$$(b) \Delta G^{\circ} = -2 \text{ mol } e^{-} \times 96,500 \text{ J/V mol } e^{-} \times 0.33 \text{ V} = -6.4 \times 10^4 \text{ J} = -64 \text{ kJ}$$

$$(c) \Delta G^{\circ} = -2 \text{ mol } e^{-} \times 96,500 \text{ J/V mol } e^{-} \times 0.148 \text{ V} = -2.86 \times 10^4 \text{ J} = -28.6 \text{ kJ}$$

$$(d) \Delta G^{\circ} = -6 \text{ mol } e^{-} \times 96,500 \text{ J/V mol } e^{-} \times 2.34 \text{ V} = -1.35 \times 10^6 \text{ J} = -1.35 \times 10^3 \text{ kJ}$$

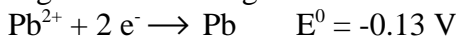
$$(e) \Delta G^{\circ} = -2 \text{ mol } e^{-} \times 96,500 \text{ J/V mol } e^{-} \times (-0.62 \text{ V}) = 1.2 \times 10^5 \text{ J} = 1.2 \times 10^2 \text{ kJ}$$

8. Will Ni^{2+} displace Cr in an aqueous solution?



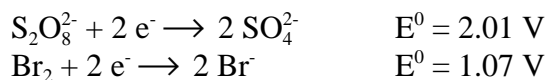
Ni^{2+} must be reduced in order to displace Cr. Since E° for Ni^{2+} is more positive than E° for Cr, it will displace it.

9. Will Mg^{2+} displace Pb metal in an aqueous solution?



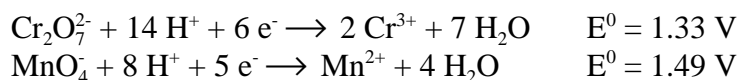
Mg^{2+} must be reduced in order to displace Cr. Since E° for Mg^{2+} is more negative than E° for Pb, it will not displace it.

10. Can $\text{S}_2\text{O}_8^{2-}$ be employed to oxidize Br^- ?



$\text{S}_2\text{O}_8^{2-}$ must be reduced in order to oxidize Br^- . Since E^0 for $\text{S}_2\text{O}_8^{2-}$ is more positive than E^0 for Br_2 , it will oxidize Br^- .

11. Can $\text{Cr}_2\text{O}_7^{2-}$ be employed to oxidize Mn^{2+} ?



$\text{Cr}_2\text{O}_7^{2-}$ must be reduced in order to oxidize Mn^{2+} . Since E^0 for MnO_4^- is more positive than E^0 for $\text{Cr}_2\text{O}_7^{2-}$, $\text{Cr}_2\text{O}_7^{2-}$ will not oxidize Mn^{2+} .

12. Calculate K for each reaction in question 5 at 25°C .

$$E^0 = \frac{0.0592}{n} \log(K); \log(K) = E^0 \left(\frac{n}{0.0592} \right)$$

$$(a) \quad \log(K) = 0.52 \left(\frac{2}{0.0592} \right) = 17.57; K = 3.7 \times 10^{17} \quad (10^{17})$$

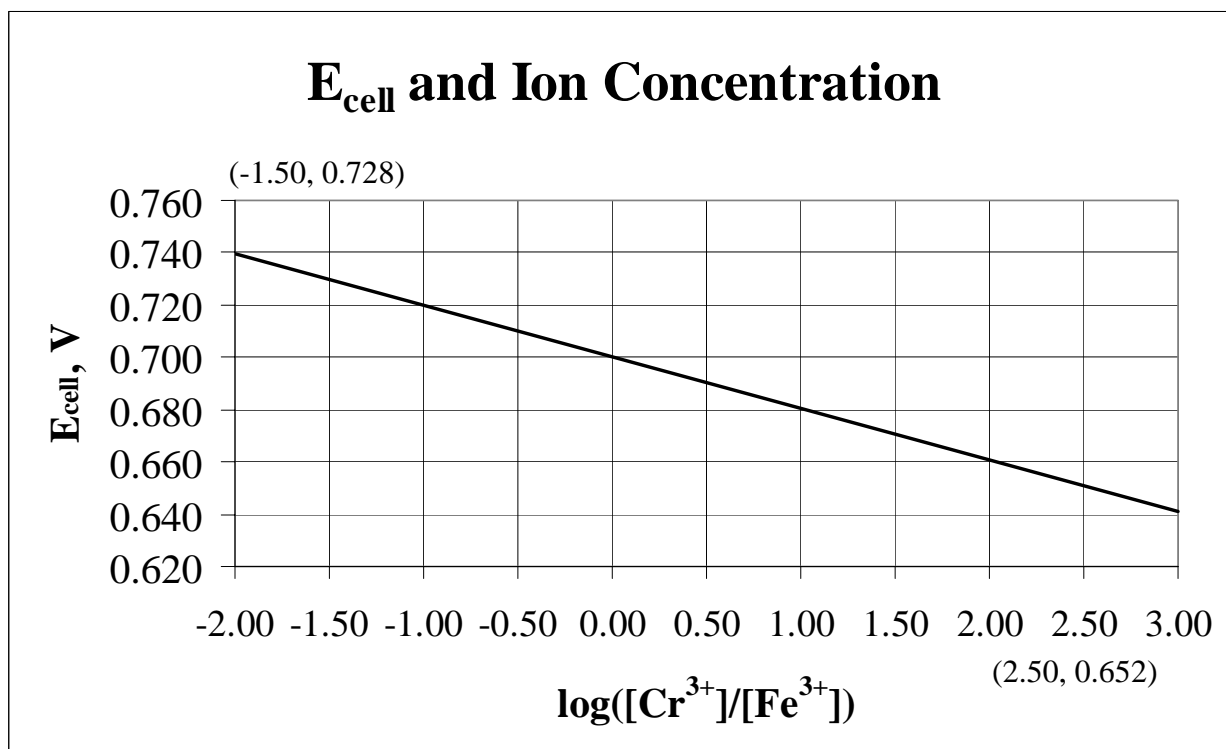
$$(b) \quad \log(K) = 0.33 \left(\frac{2}{0.0592} \right) = 11.15; K = 1.4 \times 10^{11} \quad (10^{11})$$

$$(c) \quad \log(K) = 0.148 \left(\frac{2}{0.0592} \right) = 5.00; K = 1.0 \times 10^5$$

$$(d) \quad \log(K) = 2.34 \left(\frac{6}{0.0592} \right) = 237; K = 10^{237}$$

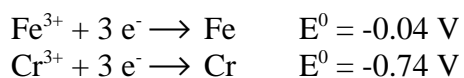
$$(e) \quad \log(K) = -0.62 \left(\frac{2}{0.0592} \right) = -20.9; K = 1.1 \times 10^{-21} \quad (10^{-21})$$

13.



$$m = \frac{(0.652 - 0.728)\text{V}}{(2.50 - (-1.50))} = -0.0190 \text{ V}$$

$$0.652 \text{ V} = -0.0190 \text{ V} \times 2.50 + b; b = 0.700 \text{ V}$$

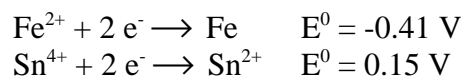
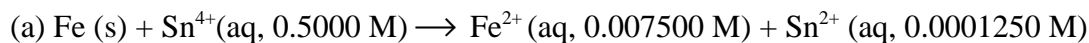


$$E_{\text{cell}}^{\circ} = -0.04 \text{ V} - (-0.74 \text{ V}) = 0.70 \text{ V} \text{ this is equal to the intercept}$$

$$-\frac{0.0592}{3} = -0.0197 \text{ this is close to the value of the slope } -0.0190$$

14. Determine E_{cell} for the following.

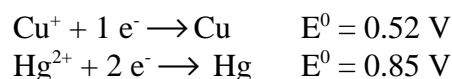
$$E = E^0 - \frac{0.0592}{n} \log(Q)$$



$$E_{\text{red}} = 0.15 - \frac{0.0592}{2} \log\left(\frac{0.0001250}{0.5000}\right) = 0.25_7 \text{ V}$$

$$E_{\text{ox}} = -0.41 - \frac{0.0592}{2} \log\left(\frac{1}{0.007500}\right) = -0.47_3 \text{ V}$$

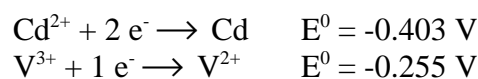
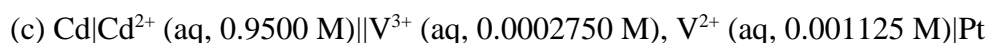
$$E_{\text{cell}} = 0.25_7 \text{ V} - (-0.47_3 \text{ V}) = 0.73 \text{ V}$$



$$E_{\text{red}} = 0.85 - \frac{0.0592}{2} \log\left(\frac{1}{0.0003250}\right) = 0.74_7 \text{ V}$$

$$E_{\text{ox}} = 0.52 - \frac{0.0592}{1} \log\left(\frac{1}{0.2500}\right) = 0.48_4 \text{ V}$$

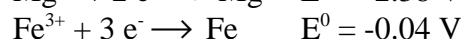
$$E_{\text{cell}} = 0.74_7 \text{ V} - 0.48_4 \text{ V} = 0.26 \text{ V}$$



$$E_{\text{red}} = -0.255 - \frac{0.0592}{1} \log\left(\frac{0.001125}{0.0002750}\right) = -0.291 \text{ V}$$

$$E_{\text{ox}} = -0.403 - \frac{0.0592}{2} \log\left(\frac{1}{0.9500}\right) = -0.404 \text{ V}$$

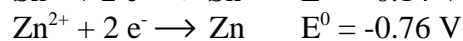
$$E_{\text{cell}} = -0.291 \text{ V} - (-0.404 \text{ V}) = 0.112 \text{ V}$$



$$E_{\text{red}} = -0.04 - \frac{0.0592}{2} \log\left(\frac{1}{0.08500}\right) = -0.07_2 \text{ V}$$

$$E_{\text{ox}} = -2.38 - \frac{0.0592}{3} \log\left(\frac{1}{0.02500}\right) = -2.41 \text{ V}$$

$$E_{\text{cell}} = -0.07_2 \text{ V} - (-2.41 \text{ V}) = 2.34 \text{ V}$$



$$E_{\text{red}} = -0.76 - \frac{0.0592}{2} \log\left(\frac{1}{2.2500}\right) = -0.75 \text{ V}$$

$$E_{\text{ox}} = -0.14 - \frac{0.0592}{2} \log\left(\frac{1}{1.2500}\right) = -0.13_7 \text{ V}$$

$$E_{\text{cell}} = -0.75 \text{ V} - (-0.13_7 \text{ V}) = -0.61 \text{ V}$$

15. Determine E_{cell} for the following cell:

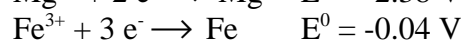


$$E_{\text{red}} = 0.34 - \frac{0.0592}{2} \log\left(\frac{1}{0.1250}\right) = 0.31 \text{ V}$$

$$E_{\text{ox}} = 0.34 - \frac{0.0592}{2} \log\left(\frac{1}{0.000001250}\right) = 0.17 \text{ V}$$

$$E_{\text{cell}} = 0.31 \text{ V} - 0.17 \text{ V} = 0.14 \text{ V}$$

16. What is the minimum $[\text{Mg}^{2+}]$ at which the following cell is able to function as a voltaic cell?



$$E_{\text{red}} = -0.04 - \frac{0.0592}{3} \log\left(\frac{1}{0.1250}\right) = -0.06 \text{ V}$$

$$E_{\text{red}} = -2.38 - \frac{0.0592}{2} \log\left(\frac{1}{[\text{Mg}^{2+}]}\right)$$

$$E_{\text{cell}} = 0; \text{ minimum point}$$

$$0 = -0.06 - \left(-2.38 - \frac{0.0592}{2} \log\left(\frac{1}{[\text{Mg}^{2+}]}\right)\right)$$

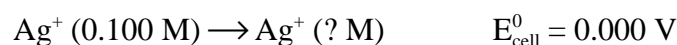
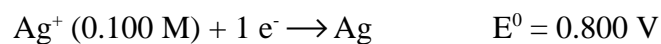
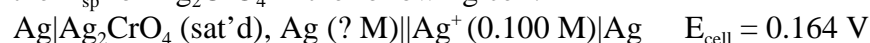
$$0.06 = 2.38 + \frac{0.0592}{2} \log\left(\frac{1}{[\text{Mg}^{2+}]}\right)$$

$$-2.32 = \frac{0.0592}{2} \log\left(\frac{1}{[\text{Mg}^{2+}]}\right); \log\left(\frac{1}{[\text{Mg}^{2+}]}\right) = -78.5$$

$$\frac{1}{[\text{Mg}^{2+}]} = 3.53 \times 10^{-79}; [\text{Mg}^{2+}] = 2.83 \times 10^{78} \text{ M}$$

Not possible!

17. Determine the K_{sp} for Ag_2CrO_4 in the following cell.



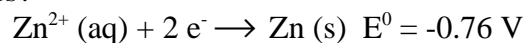
$$0.164 = 0.000 - \frac{0.0592}{1} \log\left(\frac{[\text{Ag}^+]}{0.100}\right)$$

$$\log\left(\frac{[\text{Ag}^+]}{0.100}\right) = -2.77; \frac{[\text{Ag}^+]}{0.100} = 1.7 \times 10^{-3}; [\text{Ag}^+] = 1.7 \times 10^{-4} \text{ M}$$



$$K_{\text{sp}} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}] = (1.7 \times 10^{-4})^2 \times 8.5 \times 10^{-5} = 2.4 \times 10^{-12}$$

18. What mass of Zn will plate out from an aqueous solution when a current of 30.00 A is passed for 18.00 hours?

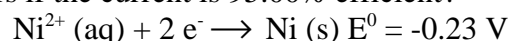


$$18.00\text{ h} \times \frac{60\text{ min}}{1\text{ h}} \times \frac{60\text{ s}}{1\text{ min}} = 6.480 \times 10^4\text{ s}$$

$$30.00\text{ amp} \times 6.480 \times 10^4\text{ s} = 1.944 \times 10^6\text{ As} = 1.944 \times 10^6\text{ C}$$

$$1.944 \times 10^6\text{ C} \times \frac{1\text{ mol e}^{-}}{96,500\text{ C}} \times \frac{1\text{ mol Zn}}{2\text{ mol e}^{-}} \times \frac{65.39\text{ g Zn}}{1\text{ mol Zn}} = 658.6\text{ g Zn}$$

19. What mass of Ni will plate out from an aqueous solution when a current of 15.00 A is passed for 12.00 hours if the current is 95.00% efficient?



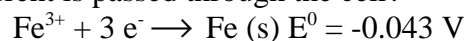
$$12.00\text{ h} \times \frac{60\text{ min}}{1\text{ h}} \times \frac{60\text{ s}}{1\text{ min}} = 4.320 \times 10^4\text{ s}$$

$$15.00\text{ A passed} \times \frac{95.00\text{ A}}{100.0\text{ A passed}} = 14.25\text{ A}$$

$$14.25\text{ A} \times 4.320 \times 10^4\text{ s} = 6.156 \times 10^5\text{ As} = 6.156 \times 10^5\text{ C}$$

$$6.156 \times 10^5\text{ C} \times \frac{1\text{ mol e}^{-}}{96,500\text{ C}} \times \frac{1\text{ mol Ni}}{2\text{ mol e}^{-}} \times \frac{58.69\text{ g Ni}}{1\text{ mol Ni}} = 187.2\text{ g Ni}$$

20. How long, in hours, will it take to plate out 125.000 g of Fe from an aqueous solution if a 35.00 amp current is passed through the cell?



$$\frac{125.000\text{ g Fe}}{1} \times \frac{1\text{ mol Fe}}{55.847\text{ g Fe}} \times \frac{3\text{ mol e}^{-}}{1\text{ mol Fe}} \times \frac{96,500\text{ C}}{1\text{ mol e}^{-}} = 1.2960 \times 10^5\text{ C}$$

$$= 1.2960 \times 10^5\text{ As}$$

$$1.2960 \times 10^5\text{ As} = 35.00\text{ A} \times t; t = 3703\text{ s} = 1.029\text{ h}$$

21. Determine how long, in hours, it will take to plate out the iron in question 20 if the current is 85.00% efficient.

$$1.2960 \times 10^5\text{ As} = 35.00\text{ A passed} \times \frac{85.00\text{ A}}{100.0\text{ A passed}} \times t; t = 4356\text{ s} = 1.210\text{ h}$$