1. A voltaic cell is an electrochemical cell in which a spontaneous chemical reaction is used to obtain electrical energy. An electrolytic cell is one in which electrical energy is used to drive an otherwise non-spontaneous reaction.

2. The cathode is the electrode in an electrochemical cell at which the reduction takes place, while the anode is the electrode at which the oxidation takes place. This description holds for both voltaic and electrolytic cells.

   In a voltaic cell the cathode has a positive charge and the anode has a negative charge. Therefore, cations migrate toward the anode and anions migrate toward the cathode.

   In an electrolytic cell, the cathode has a negative charge and the anode has a positive charge. Therefore, cations migrate toward the cathode and anions migrate toward the anode.

3. The SI unit of electric potential is the volt (V).

4. The faraday (F) is the magnitude of the charge on one mole of electrons. In terms of the SI unit of electric charge, the Coulomb (C), 1 F = 96,485 C to 5 S.F.

7. The SI unit of energy is the joule (J). 1 J = 1 C · V

8. $\Delta G° = -nFE°$ $\Delta G° = -RT \ln K$ $\ln K = \dfrac{nFE°}{RT}$

9. The corrosion of iron can be described by a series of reactions that starts with

$$2\,Fe\,(s) + O_2\,(g) + 2\,H_2O\,(l) \rightarrow 2\,Fe^{2+}\,(aq) + 4\,OH^-\,(aq)$$

   The Nernst Equation tells us:

$$E_{cell} = E°_{cell} - \frac{RT}{4F} \ln \frac{[Fe^{2+}]^2[OH^-]^4}{P_{O_2}}$$

   As the pH decreases, $[OH^-]$ decreases and E_{cell} becomes more positive indicating that the process becomes more favorable as the pH decreases.

13. Zinc is more active than iron, but iron is more active than tin. If iron is coated with zinc, the iron is not exposed and thus protected from corrosion. Even if the zinc coating is compromised and the iron is exposed, the zinc is oxidized in preference to the iron. We say the iron is cathodically protected by the zinc. If the iron is coated by tin, however, the tin protects the iron only as long as the coating remains intact. If the coating is damaged and the iron is exposed, the tin actually accelerates the oxidation of the iron because the tin is cathodic relative to the iron.

14. The sulfuric acid serves two purposes. Since H_2SO_4 is a strong acid, it dissociates into H^+ and HSO_4^- ions. The presence of the ions allows current to pass through the solution. In addition, electrolyzing an acidic aqueous solution decreases the potential required for the electrolysis.

$2\ H^+\ (1\ M) + 2\ e^- \rightarrow H_2\ (1\ atm)$	$E° = 0.000$ V	pH = 0
$2\ H_2O\ (l) + 2\ e^- \rightarrow H_2\ (1\ atm) + 2\ OH^-\ (1\ M)$	$E° = -0.83$ V	pH = 14

At pH = 7, the potential would be −0.41 V.

17. In molten NaCl, Na^+ is the only species that can be reduced at the cathode and Cl^- is the only substance that can be oxidized at the anode. However, in aqueous solutions, the reduction of water is much more favorable than the reduction of Na^+, and, except at relatively high Cl^- concentration, the oxidation of water is more favorable than the oxidation of Cl^-. Even if Na^+ could be formed, it would spontaneously react with water to produce H_2 anyway.

19. a. no change

    b. no change

    c. The zinc strip would gradually disintegrate, and the blue color characteristic of Cu^{2+} solutions would fade as the Cu^{2+} ions are reduced to copper metal.

    d. no change

21. The zinc is present to protect the iron from corrosion. The presence of a more active metal (acting as a sacrificial anode) in contact with the iron renders the iron cathodic and keeps it from corroding.

23. Assuming that the zinc is the limiting reactant in the process, with more zinc present, the heavy duty cell will operate for a longer period of time under the same load than the standard cell.

25. One choice is Ag/Ag^+ combined with Pb/Pb^{2+}

$$Pb \mid Pb^{2+}\ (1\ M) \parallel Ag^+\ (1\ M) \mid Ag$$

$Pb\ (s) \rightarrow Pb^{2+}\ (1M) + 2\ e^-$	E° = +0.13 V
$2\ Ag^+\ (1\ M) + 2\ e^- \rightarrow Ag\ (s)$	E° = +0.80 V
$2\ Ag^+\ (1\ M) + Pb\ (s) \rightarrow Pb^{2+}\ (1M) + Ag\ (s)$	E° = +0.93 V

27. For this cell, $E = E° - \frac{RT}{nF} \ln Q = 1.10\ V - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$

a. Increasing $[Cu^{2+}]$ would increase the potential and increase the lamp's intensity.

b. Increasing $[Zn^{2+}]$ would decrease the potential and decrease the lamp's intensity.

c. Adding water would decrease $[Cu^{2+}]$ and would decrease the lamp's intensity.

d. Removing the salt bridge would open the circuit, no current would flow and the light would go out.

29. a. $Cr_2O_7^{2-} + C_2O_4^{2-} \rightarrow Cr^{3+} + CO_2$

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ (reduction)

$3\ (\ C_2O_4^{2-} \rightarrow 2\ CO_2 + 2\ e^-\ )$ (oxidation)

$14\ H^+ + Cr_2O_7^{2-} + 3\ C_2O_4^{2-} \rightarrow 2\ Cr^{3+} + 6\ CO_2 + 7\ H_2O$

b. $Cu + NO_3^- \rightarrow Cu^{2+} + NO$

$3\ (\ Cu \rightarrow Cu^{2+} + 2\ e^-\ )$ (oxidation)

$2\ (\ 3\ e^- + 4\ H^+ + NO_3^- \rightarrow NO + 2\ H_2O\ )$ (reduction)

$8\ H^+ + 3\ Cu + 2\ NO_3^- \rightarrow 3\ Cu^{2+} + 2\ NO + 4\ H_2O$

c. $MnO_2 + HNO_2 \rightarrow Mn^{2+} + NO_3^-$

$2\ e^- + 4\ H^+ + MnO_2 \rightarrow Mn^{2+} + 2\ H_2O$ (reduction)

$H_2O + HNO_2 \rightarrow NO_3^- + 3\ H^+ + 2\ e^-$ (oxidation)

$H^+ + MnO_2 + HNO_2 \rightarrow Mn^{2+} + NO_3^- + H_2O$

d. $PbO_2 + Mn^{2+} + SO_4^{2-} \rightarrow PbSO_4 + MnO_4^-$

$5\ (\ 2\ e^- + 4\ H^+ + PbO_2 + SO_4^{2-} \rightarrow PbSO_4 + 2\ H_2O\ )$ (reduction)

$2\ (\ 4\ H_2O + Mn^{2+} \rightarrow MnO_4^- + 8\ H^+ + 5\ e^-\ )$ (oxidation)

$4\ H^+ + 5\ PbO_2 + 5\ SO_4^{2-} + 2\ Mn^{2+} \rightarrow 5\ PbSO_4 + 2\ MnO_4^- + 2\ H_2O$

e. $HNO_2 + Cr_2O_7^{2-} \rightarrow Cr^{3+} + NO_3^-$

$3\ (\ H_2O + HNO_2 \rightarrow NO_3^- + 3\ H^+ + 2\ e^-\ )$ (oxidation)

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ (reduction)

$5\ H^+ + 3\ HNO_2 + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 3\ NO_3^- + 4\ H_2O$

31. a. $Mn^{2+} + H_2O_2 \rightarrow MnO_2 + H_2O$

$4\ OH^- + 2\ H_2O + Mn^{2+} \rightarrow MnO_2 + 4\ H_2O + 2\ e^-$ (oxidation)

$2\ e^- + 2H_2O + H_2O_2 \rightarrow 2\ H_2O + 2\ OH^-$ (reduction)

$2\ OH^- + Mn^{2+} + H_2O_2 \rightarrow MnO_2 + 2\ H_2O$

b. $MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-$

$2\ (\ 3\ e^- + 4\ H_2O + MnO_4^- \rightarrow MnO_2 + 2\ H_2O + 4\ OH^-\ )$ (reduction)

$3\ (\ 2\ OH^- + H_2O + NO_2^- \rightarrow NO_3^- + 2\ H_2O + 2\ e^-\ )$ (oxidation)

$H_2O + 2\ MnO_4^- + 3\ NO_2^- \rightarrow 2\ MnO_2 + 3\ NO_3^- + 2\ OH^-$

c. $Mn^{2+} + ClO_3^- \rightarrow MnO_2 + ClO_2$

$4\ OH^- + 2\ H_2O + Mn^{2+} \rightarrow MnO_2 + 4\ H_2O + 2\ e^-$ (oxidation)

$2\ (\ e^- + 2\ H_2O + ClO_3^- \rightarrow ClO_2 + H_2O + 2\ OH^-\ )$ (reduction)

$Mn^{2+} + 2\ ClO_3^- \rightarrow MnO_2 + 2\ ClO_2$

d. $MnO_4^- + NO_2 \rightarrow MnO_2 + NO_3^-$

$3\ e^- + 4\ H_2O + MnO_4^- \rightarrow MnO_2 + 2\ H_2O + 4\ OH^-$ (reduction)

$3\ (\ 2\ OH^- + H_2O + NO_2 \rightarrow NO_3^- + 2\ H_2O + e^-\ )$ (oxidation)

$2\ OH^- + MnO_4^- + 3\ NO_2 \rightarrow MnO_2 + 3\ NO_3^- + H_2O$

e. $Cl_2 \rightarrow Cl^- + ClO_3^-$

$5\ (\ 2\ e^- + Cl_2 \rightarrow 2\ Cl^-\ )$ (reduction)

$12\ OH^- + 6\ H_2O + Cl_2 \rightarrow 2\ ClO_3^- + 12\ H_2O + 10\ e^-$ (oxidation)

$12\ OH^- + 6\ Cl_2 \rightarrow 10\ Cl^- + 2\ ClO_3^- + 6\ H_2O$

$6\ OH^- + 3\ Cl_2 \rightarrow 5\ Cl^- + ClO_3^- + 3\ H_2O$

33. a. $H_2S + NO_3^- \rightarrow + NO_2 + S_8$

$8\ H_2S \rightarrow S_8 + 16\ H^+ + 16\ e^-$ (oxidation)

$16\ (\ e^- + 2\ H^+ + NO_3^- \rightarrow NO_2 + H_2O\ )$ (reduction)

$16\ H^+ + 16\ NO_3^- + 8\ H_2S \rightarrow S_8 + 16\ NO_2 + 16\ H_2O$

b. $NO_3^- + Cu \rightarrow NO + Cu^{2+}$

$3\ (\ Cu \rightarrow Cu^{2+} + 2\ e^-\ )$ (oxidation)

$2\ (\ 3\ e^- + 4\ H^+ + NO_3^- \rightarrow NO + 2\ H_2O\ )$ (reduction)

$8\ H^+ + 3\ Cu + 2\ NO_3^- \rightarrow 3\ Cu^{2+} + 2\ NO + 4\ H_2O$

c. $MnO_4^- + SO_2 \rightarrow SO_4^{2-} + Mn^{2+}$

$2\,(\,5\,e^- + 8\,H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O\,)$ (reduction)

$5\,(\,2\,H_2O + SO_2 \rightarrow SO_4^{2-} + 4\,H^+ + 2\,e^-\,)$ (oxidation)

$2\,H_2O + 2\,MnO_4^- + 5\,SO_2 \rightarrow 5\,SO_4^{2-} + 2\,Mn^{2+} + 4\,H^+$

d. $Bi(OH)_3 + Sn(OH)_3^- \rightarrow Bi + Sn(OH)_6^{2-}$

$2\,(\,3\,e^- + Bi(OH)_3 \rightarrow Bi + 3\,OH^-\,)$ (reduction)

$3\,(\,3\,OH^- + Sn(OH)_3^- \rightarrow Sn(OH)_6^{2-} + 2\,e^-\,)$ (oxidation)

$2\,Bi(OH)_3 + 3\,Sn(OH)_3^- + 3\,OH^- \rightarrow 2\,Bi + 3\,Sn(OH)_6^{2-}$

35. a. $MnO_4^- + I^- \rightarrow MnO_2 + IO_3^-$

$2\,(\,3\,e^- + 4\,H_2O + MnO_4^- \rightarrow MnO_2 + 2\,H_2O + 4\,OH^-\,)$ (reduction)

$6\,OH^- + 3\,H_2O + I^- \rightarrow IO_3^- + 6\,H_2O + 6\,e^-$ (oxidation)

$H_2O + 2\,MnO_4^- + I^- \rightarrow 2\,MnO_2 + IO_3^- + 2\,OH^-$

b. $Cr_2O_7^{2-} + Cl^- \rightarrow Cr^{3+} + Cl_2$

$6\,e^- + 14\,H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$ (reduction)

$3\,(\,2\,Cl^- \rightarrow Cl_2 + 2\,e^-\,)$ (oxidation)

$14\,H^+ + Cr_2O_7^{2-} + 6\,Cl^- \rightarrow 2\,Cr^{3+} + 3\,Cl_2 + 7\,H_2O$

c. $S_8 + NO_3^- \rightarrow SO_2 + NO$

$32\,(\,3\,e^- + 4\,H^+ + NO_3^- \rightarrow NO + 2\,H_2O\,)$ (reduction)

$3\,(\,16\,H_2O + S_8 \rightarrow 8\,SO_2 + 32\,H^+ + 32\,e^-\,)$ (oxidation)

$32\,H^+ + 3\,S_8 + 32\,NO_3^- \rightarrow 24\,SO_2 + 32\,NO + 16\,H_2O$

d. $H_2O_2 + MnO_4^- \rightarrow O_2 + MnO_2$

$2\,(3\,e^- + 4\,H_2O + MnO_4^- \rightarrow MnO_2 + 2\,H_2O + 4\,OH^-)$ (reduction)

$3\,(2\,OH^- + H_2O_2 \rightarrow O_2 + 2\,H_2O + 2\,e^-)$ (oxidation)

$3\,H_2O_2 + 2\,MnO_4^- \rightarrow 3\,O_2 + 2\,MnO_2 + 2\,OH^- + 2\,H_2O$

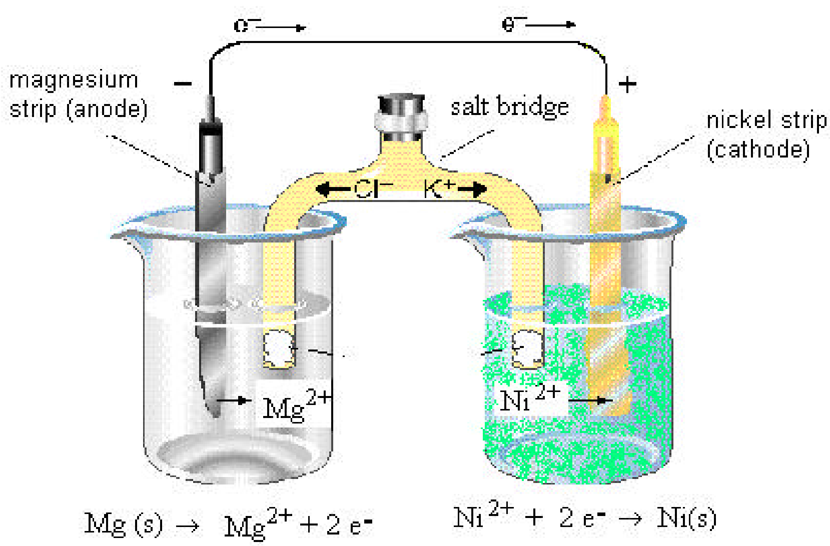
e. $Zn + NO_3^- \rightarrow Zn^{2+} + N_2$

$5\,(Zn \rightarrow Zn^{2+} + 2\,e^-)$ (oxidation)

$10\,e^- + 12\,H^+ + 2\,NO_3^- \rightarrow N_2 + 6\,H_2O$ (reduction)

$12\,H^+ + 5\,Zn + 2\,NO_3^- \rightarrow 5\,Zn^{2+} + N_2 + 6\,H_2O$

37.



39. The cell looks identical to that in question 35 with Zn replacing Mg and Ag replacing Ni. The half reactions are:

$Zn \rightarrow Zn^{2+} + 2\,e^-$ (anode)

$Ag^+ + e^- \rightarrow Ag$ (cathode)

41. $Zn\ (s) + 2\ OH^- (aq) \rightarrow Zn(OH)_2\ (s) + 2\ e^-$ (anode)

$Ag_2O\ (s) + H_2O\ (l) + 2\ e^- \rightarrow 2\ Ag\ (s) + 2\ OH^-\ (aq)$ (cathode)

$Zn\ (s) + Ag_2O\ (s) + H_2O\ (l) \rightarrow Zn(OH)_2\ (s) + 2\ Ag\ (s)$ (net reaction)

43. According to the half reactions given, Ni is the anode because it is oxidized and Pb is the cathode since Pb^{2+} is reduced to Pb. Checking the standard reduction potentials, Ni is more easily oxidized than Pb, so the combination of this cathode and anode do in fact result in a voltaic cell. The anode is always on the left and the cathode on the right in a cell diagram. Therefore we have:

$$Ni\ (s) |\ Ni^{2+}\ (aq) ||\ Pb^{2+}\ (aq)\ |\ Pb\ (s)$$

45. H^+ is more easily reduced than Ni^{2+}, so H^+ is reduced to H_2 at the cathode and Ni is oxidized to Ni^{2+} at the anode.

$$Ni\ (s) |\ Ni^{2+}\ (1\ M) ||\ H^+\ (1\ M)\ |\ H_2\ (g) |\ Pt$$

47. $Fe\ (s) \rightarrow Fe^{2+}\ (aq) + 2\ e^-$ (anode)

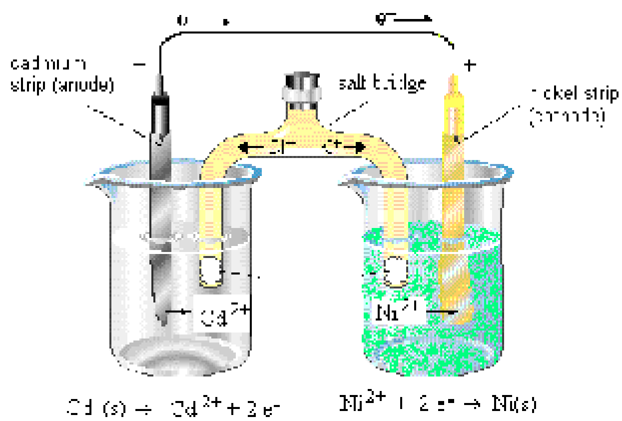
$2\ (\ Ag^+\ (aq) + e^- \rightarrow Ag\ (s))$ (cathode)

$Fe\ (s) + 2\ Ag^+\ (aq) \rightarrow 2\ Ag\ (s) + Fe^{2+}\ (aq)$

49. $Cd\ (s) \rightarrow Cd^{2+}\ (aq) + 2\ e^-$ (anode)

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

$Cd\ (s) + Ni^{2+}\ (aq) \rightarrow Ni\ (s) + Cd^{2+}\ (aq)$



51. For one mole of Fe, the equation can be written as:

Fe^{3+} (aq) + ½ Zn (s) → Fe^{2+} (aq) + ½ Zn^{2+} (aq)   so n = 1

$$w'_{max} = \Delta G = -nFE$$
$$w'_{max} = -nFE$$
$$w'_{max} = -1\ mol(96485\ \frac{J}{V\cdot mol})(0.72\ V)$$
$$w'_{max} = -6.9 \times 10^4\ J = -69\ kJ$$

53. Ni (s) → Ni^{2+} (aq) + 2 e^-

2 (Ag^+ (aq) + e^- → Ag (s))

Ni (s) + 2 Ag^+ (aq) → 2 Ag (s) + Ni^{2+} (aq) E = +0.97 V

$$w'_{max} = -2\ mol(96485\ \frac{J}{V\cdot mol})(0.97\ V)$$
$$w'_{max} = -1.87 \times 10^5\ J = -187\ kJ\ per\ mol\ Ni$$

$$30.0\ g\ Ni \times \frac{1\ mol\ Ni}{58.69\ g} \times \frac{-187\ kJ}{1\ mol\ Ni} = -96\ kJ$$

55. Oxidizing power increases as E_{red} increases. Under standard conditions we look at E°_{red}.

NO_3^- (aq) + 4 H^+(aq) + 3 e^- → NO (g) + 2 H_2O (l) E°_{red} = +0.96 V

O_2 (g) + 4 H^+(aq) + 4 e^- → 2 H_2O (l) E°_{red} = +1.23 V

MnO_4^- (aq) + 8 H^+(aq) + 5 e^- → Mn^{2+} (aq) + 4 H_2O (l) E°_{red} = +1.49 V

Oxidizing power increases from NO_3^- to O_2 to MnO_4^-.

57. Reducing power increases as the reduction potential of the oxidized form becomes more negative.

Cu^{2+} (aq) + e^- → Cu^+ (aq) E°_{red} = +0.16 V

Fe^{2+} (aq) + 2 e^- → Fe (s) E°_{red} = −0.41 V

Zn^{2+} (aq) + 2 e^- → Zn (s) E°_{red} = −0.76 V

Reducing power increases from Cu^+ to Fe to Zn

59. a. Sn^{4+} (aq) + 2 e^- → Sn^{2+} (aq) E°_{red} = +0.15 V

2 Fe^{2+} (aq) → 2 Fe^{3+} (aq) + 2 e^- E°_{ox} = −0.77 V

Sn^{4+} (aq) + 2 Fe^{2+} (aq) → Sn^{2+} (aq) + 2 Fe^{3+} (aq) E°_{cell} = −0.62 V

Since E°_{cell} < 0, the reaction is non-spontaneous as written under standard conditions.

b. 4 MnO_4^- (aq) + 32 H^+(aq) + 20 e^- → 4 Mn^{2+} (aq) + 16 H_2O (l) E°_{red} = +1.49 V

10 H_2O (l) → 5 O_2 (g) + 20 H^+(aq) + 20 e^- E°_{ox} = −1.23 V

4 MnO_4^- (aq) + 12 H^+(aq) → 4 Mn^{2+} (aq) + 5 O_2 (g) + 6 H_2O (l) E°_{cell} = +0.26 V

Since E°_{cell} > 0, the reaction is spontaneous as written under standard conditions.

61. The Cl_2 (g) will oxidize Br^- (aq) to Br_2 (l), but will not react with F^- (aq)

Cl_2 (g) + 2 e^- → 2 Cl^- (aq) E°_{red} = +1.36 V

2 Br^- (aq) → Br_2 (l) + 2 e^- E°_{ox} = −1.07 V

Cl_2 (g) + 2 Br^- (aq) → Br_2 (l) + 2 Cl^- (aq) E°_{cell} = +0.29 V

Reaction is spontaneous

Cl_2 (g) + 2 e^- → 2 Cl^- (aq) E°_{red} = +1.36 V

2 F^- (aq) → F_2 (g) + 2 e^- E°_{ox} = −2.87 V

Cl_2 (g) + 2 F^- (aq) → F_2 (g) + 2 Cl^- (aq) E°_{cell} = −1.51 V

Reaction is non-spontaneous

63. Cr (s)| Cr^{3+} (aq)|| Hg_2^{2+} (aq) | Hg (l)

2 Cr (s) → 2 Cr^{3+} (aq) + 6 e^- E°_{ox} = +0.74 V

3 Hg_2^{2+} (aq) + 6 e^- → 6 Hg (l) E°_{red} = +0.80 V

2 Cr (s) + 3 Hg_2^{2+} (aq) → 6 Hg (l) + 2 Cr^{3+} (aq) E°_{cell} = +1.54 V

65. $2\ Cr\ (s) \rightarrow 2\ Cr^{3+}\ (aq) + 6\ e^{-}$ $\qquad E^{\circ}_{ox} = +0.74\ V$

$3\ I_2\ (g) + 6\ e^{-} \rightarrow 6\ I^{-}\ (aq)$ $\qquad E^{\circ}_{red} = +0.54\ V$

$2\ Cr\ (s) + 3\ I_2\ (g) \rightarrow 2\ Cr^{3+}\ (aq) + 6\ I^{-}\ (aq)$ $\qquad E^{\circ}_{cell} = +1.28\ V$

67. $3\ (\ Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2\ e^{-}\ )$ $\qquad E^{\circ}_{ox} = -0.34\ V$

$2\ (\ 3\ e^{-} + 4\ H^{+}\ (aq) + NO_3^{-}\ (aq) \rightarrow NO\ (g) + 2\ H_2O\ (l))$ $\qquad E^{\circ}_{red} = +0.96\ V$

$8\ H^{+}\ (aq) + 3\ Cu\ (s) + 2\ NO_3^{-}\ (aq)$

$\rightarrow 3\ Cu^{2+}\ (aq) + 2\ NO\ (g) + 4\ H_2O\ (l)$ $\qquad E^{\circ}_{cell} = +0.62\ V$

$$\Delta G^{\circ} = -nFE^{\circ}$$
$$= -6\ mol\left(96485\ \frac{J}{V\cdot mol}\right)(0.62\ V)$$
$$= -3.6 \times 10^{5}\ J = -360\ kJ$$

69. $Cl_2\ (g) + 2\ e^{-} \rightarrow 2\ Cl^{-}\ (aq)$ $\qquad E^{\circ}_{red} = +1.36\ V$

$2\ I^{-}\ (aq) \rightarrow I_2\ (s) + 2\ e^{-}$ $\qquad E^{\circ}_{ox} = -0.54\ V$

$Cl_2\ (g) + 2\ I^{-}\ (aq) \rightarrow I_2\ (s) + 2\ Cl^{-}\ (aq)$ $\qquad E^{\circ}_{cell} = +0.82\ V$

$$\Delta G^{\circ} = -nFE^{\circ}$$
$$= -2\ mol\left(96485\ \frac{J}{V\cdot mol}\right)(0.82\ V)$$
$$= -1.6 \times 10^{5}\ J = -160\ kJ$$

71. $Mg\ (s) + 2\ Ag^{+}\ (aq) \rightarrow Mg^{2+}\ (aq) + 2\ Ag\ (s)$

$\Delta G^{\circ} = [1\ mol\ (\Delta G^{\circ}_{f}\ (Mg^{2+}\ (aq))) + 2\ mol\ (\Delta G^{\circ}_{f}\ (Ag\ (s)))]$

$-\ [1\ mol\ (\Delta G^{\circ}_{f}\ (Mg\ (s))) + 2\ mol\ (\Delta G^{\circ}_{f}\ (Ag^{+}\ (aq)))]$

$\Delta G^{\circ} = [1\ mol\ (-454.8\ kJ \cdot mol^{-1}) + 2\ mol\ (0\ kJ \cdot mol^{-1})]$

$-[1\ mol\ (0\ kJ \cdot mol^{-1}) + 2\ mol\ (77.12\ kJ \cdot mol^{-1})]$

$\Delta G° = -609.0$ kJ = $-609,\underline{0}00$ J

$$E° = \frac{-\Delta G°}{nF}$$
$$= \frac{-(-609{,}000\ J)}{2\ mol(96485\ \frac{J}{V\cdot mol})}$$
$$= 3.156\ V$$

73. $Pb\ (s) + 2\ HSO_4^-\ (aq) + 2\ H^+\ (aq) + PbO_2\ (s) \rightarrow 2\ PbSO_4\ (s) + 2\ H_2O\ (l)$

$\Delta G° = [2\ mol\ (\Delta G°_f\ (PbSO_4\ (s))) + 2\ mol\ (\Delta G°_f\ (H_2O\ (l)))] - [1\ mol\ (\Delta G°_f\ (Pb\ (s)))$

$+ 2\ mol\ (\Delta G°_f\ (HSO_4^-\ (aq))) + 2\ mol\ (\Delta G°_f\ (H^+\ (aq))) + 1\ mol\ (\Delta G°_f\ (PbO_2\ (s)))]$

$\Delta G° = [2\ mol\ (-813.2\ kJ \cdot mol^{-1}) + 2\ mol\ (-237.1\ kJ \cdot mol^{-1})] - [1\ mol\ (0\ kJ \cdot mol^{-1})$

$+ 2\ mol\ (-756.0\ kJ \cdot mol^{-1}) + 2\ mol\ (0\ kJ \cdot mol^{-1}) + 1\ mol\ (-215.4\ kJ \cdot mol^{-1})]$

$\Delta G° = -373.2$ kJ = $-373,\underline{2}00$ J

$$E° = \frac{-\Delta G°}{nF}$$
$$= \frac{-(-373{,}200\ J)}{2\ mol(96485\ \frac{J}{V\cdot mol})}$$
$$= 1.93\ V$$

75. $2\ Fe^{3+}\ (aq) + 2\ e^- \rightarrow 2\ Fe^{2+}\ (aq)$ $\quad E°_{red} = +0.77$ V

$Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2\ e^-$ $\quad E°_{ox} = -0.34$ V

$Cu\ (s) + 2\ Fe^{3+}\ (aq) \rightarrow Cu^{2+}\ (aq) + 2\ Fe^{2+}\ (aq)$ $\quad E°_{cell} = -0.43$ V

$$\ln K = \frac{nFE°}{RT}$$
$$= \frac{(2\ mol)(96485\ \frac{J}{V\cdot mol})(0.43\ V)}{(8.314\ \frac{J}{mol\cdot K})(298\ K)}$$
$$= 3\underline{3}.5$$
$$K = e^{3\underline{3}.5} = 10^{14}$$

77. Cu^{+} (aq) + e^{-} → Cu (s) E°_{red} = +0.52 V

Cu^{+} (aq) → Cu^{2+} (aq) + e^{-} E°_{ox} = −0.16 V

2 Cu^{+} (aq) → Cu (s) + Cu^{2+} (aq) E°_{cell} = +0.36 V

$$\ln K = \frac{nFE^{\circ}}{RT}$$

$$= \frac{(1\ \text{mol})(96485\ \frac{J}{V\cdot mol})(0.36\ V)}{(8.314\ \frac{J}{mol\cdot K})(298\ K)}$$

$$= 14.0$$

$$K = e^{14.0} = 10^{6}$$

79. The half reactions are:

2 Cr (s) → 2 Cr^{3+} (1.0 x 10^{-3} *M*) + 6 e^{-} E°_{ox} = +0.74 V

3 Ni^{2+} (1.5 *M*) + 6 e^{-} → 3 Ni (s) E°_{red} = −0.23 V

2 Cr (s) + 3 Ni^{2+} (1.5 *M*) → 2 Cr^{3+} (1.0 x 10^{-3} *M*) + 3 Ni (s) E°_{cell} = +0.51 V

$$Q = \frac{[Cr^{3+}]^2}{[Ni^{2+}]^3} = \frac{(1.0 \times 10^{-3})^2}{(1.5)^3}$$

$$E = E^{\circ} - \frac{RT}{nF} \ln Q$$

$$= 0.51\ V - \frac{0.0591}{6} \log \frac{(1.0 \times 10^{-3})^2}{(2.0)^3}$$

$$= 0.51\ V - (-0.0643\ V) = 0.57\ V$$

81. $2\,(\,5\,e^- + 8\,H^+\,(1.0\,M) + MnO_4^-\,(0.010\,M) \rightarrow Mn^{2+}\,(0.15\,M) + 4\,H_2O\,)$ $\quad E^\circ_{red} = +1.49$ V

$5\,(\,2\,Br^-\,(0.010\,M) \rightarrow Br_2\,(l) + 2\,e^-\,)$ $\quad E^\circ_{ox} = -1.07$ V

$16\,H^+\,(1.0\,M) + 2\,MnO_4^-\,(0.010\,M) + 10\,Br^-\,(0.010\,M)$

$\rightarrow 5\,Br_2\,(l) + 2\,Mn^{2+}\,(0.15\,M) + 8\,H_2O$ $\quad E^\circ_{cell} = +0.42$ V

$$Q = \frac{[Mn^{2+}]^2}{[MnO_4^-]^2\,[H^+]^{16}\,[Br^-]^{10}} = \frac{(0.15)^2}{(0.010)^2\,(1.0)^{16}\,(0.010)^{10}} = 2.\underline{2}5 \times 10^{22}$$

$$E = E^\circ - \frac{RT}{nF}\ln Q$$
$$= 0.42\text{ V} - \frac{0.0591}{10}\log 2.\underline{2}5 \times 10^{22}$$
$$= 0.42\text{ V} - 0.1321\text{ V} = 0.29\text{ V}$$

83. The net reaction is: $Cd\,(s) + Ni^{2+}\,(1.0\,M) \rightarrow Cd^{2+}\,(c) + Ni\,(s)$ $\quad E^\circ_{cell} = +0.170$ V

$$0.240\text{ V} = 0.170\text{ V} - \frac{0.0591}{2}\log\frac{c}{1.0}$$
$$\frac{0.0591}{2}\log\frac{c}{1.0} = 0.170\text{ V} - 0.240\text{ V} = -0.070\text{ V}$$

$$\log c = \frac{2\,(-0.070\text{ V})}{0.0591} = 2.\underline{3}7$$

$$c = 10^{-2.\underline{3}7} = 4 \times 10^{-3}\,M$$

85. a. cathode half reaction: $Ca^{2+}\,(l) + 2\,e^- \rightarrow Ca\,(l)$

anode half reaction: $S^{2-}\,(l) \rightarrow S\,(l) + 2\,e^-$

b. cathode half reaction: $Cs^+\,(l) + e^- \rightarrow Cs\,(l)$

anode half reaction: $4\,OH^-\,(l) \rightarrow O_2\,(g) + 2\,H_2O\,(l) + 4\,e^-$

87. a. Na_2SO_4 (aq) Water and (H^+ is reduced to H_2 (g) at the cathode and water is oxidized to O_2 (g) at the anode. (The reduction of SO_4^{2-} to SO_3^{2-} is also possible and is only a little less favorable than the reduction of water.)

$$2\,H_2O\,(l) \rightarrow 2\,H_2\,(g) + O_2\,(g)$$

b. KBr (aq) Water and (H^+ is reduced to H_2 (g) at the cathode and Br^- is oxidized to Br (l) at the anode. However, in dilute KBr, it is possible that water could be reduced to O_2 (g).

$$2\ H_2O\ (l) + 4\ Br^-\ (aq) \rightarrow H_2\ (g) + 2\ Br_2\ (l) + 2\ OH^-\ (aq)$$

89. $$3.61\text{ kg Al} \times \frac{1000\text{ g}}{1\text{ kg}} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} \times \frac{3\text{ F}}{1\text{ mol Al}} \times \frac{96{,}485\text{ C}}{1\text{ F}} = 3.87 \times 10^7\text{ C}$$

91. $$5.00 \times 10^3\text{ C} \times \frac{1\text{ F}}{96{,}485\text{ C}} \times \frac{1\text{ mol Li}}{1\text{ F}} \times \frac{6.941\text{ g Li}}{1\text{ mol Li}} = 0.360\text{ g Li}$$

93. a. anode: $3\ (\ Fe\ (s) \rightarrow Fe^{2+}\ (aq) + 2\ e^-\ )$

cathode: $2\ (\ NO_3^-\ (aq) + 4\ H^+\ (aq) + 3\ e^- \rightarrow NO\ (g) + 2\ H_2O\ (l)\ )$

net: $3\ Fe\ (s) + 2\ NO_3^-\ (aq) + 8\ H^+\ (aq) \rightarrow 3\ Fe^{2+}\ (aq) + 2\ NO\ (g) + 4\ H_2O\ (l)$

b. anode: $3\ (\ Fe^{2+}\ (aq) \rightarrow Fe^{3+}\ (aq) + e^-\ )$

cathode: $NO_3^-\ (aq) + 4\ H^+\ (aq) + 3\ e^- \rightarrow NO\ (g) + 2\ H_2O\ (l)$

net: $3\ Fe^{2+}\ (aq) + NO_3^-\ (aq) + 4\ H^+\ (aq) \rightarrow 3\ Fe^{3+}\ (aq) + NO\ (g) + 2\ H_2O\ (l)$

c. anode: $Fe\ (s) \rightarrow Fe^{3+}\ (aq) + 3\ e^-$

cathode: $NO_3^-\ (aq) + 4\ H^+\ (aq) + 3\ e^- \rightarrow NO\ (g) + 2\ H_2O\ (l)$

net: $Fe\ (s) + NO_3^-\ (aq) + 4\ H^+\ (aq) \rightarrow Fe^{3+}\ (aq) + NO\ (g) + 2\ H_2O\ (l)$

Note that this is the same as adding the net reactions of a and b and dividing through by 3.

$3\ Fe\ (s) + 2\ NO_3^-\ (aq) + 8\ H^+\ (aq) \rightarrow 3\ Fe^{2+}\ (aq) + 2\ NO\ (g) + 4\ H_2O\ (l)$

$3\ Fe^{2+}\ (aq) + NO_3^-\ (aq) + 4\ H^+\ (aq) \rightarrow 3\ Fe^{3+}\ (aq) + NO\ (g) + 2\ H_2O\ (l)$

$3\ Fe\ (s) + 3\ NO_3^-\ (aq) + 12\ H^+\ (aq) \rightarrow 3\ Fe^{3+}\ (aq) + 3\ NO\ (g) + 6\ H_2O\ (l)$

$Fe\ (s) + NO_3^-\ (aq) + 4\ H^+\ (aq) \rightarrow Fe^{3+}\ (aq) + NO\ (g) + 2\ H_2O\ (l)$

95. a. $16\ (3\ e^- + 4\ H_2O + MnO_4^- \rightarrow MnO_2 + 2\ H_2O + 4\ OH^-)$ (reduction)

$3\ (8\ S^{2-} \rightarrow S_8 + 16\ e^-)$ (oxidation)

$32\ H_2O + 16\ MnO_4^- + 24\ S^{2-} \rightarrow 3\ S_8 + 16\ MnO_2 + 64\ OH^-$

b. $6\ e^- + 6\ H^+ + IO_3^- \rightarrow I^- + 3\ H_2O$ (reduction)

$3\ (H_2O + HSO_3^- \rightarrow SO_4^{2-} + 3\ H^+ + 2\ e^-)$ (oxidation)

$IO_3^- + 3\ HSO_3^- \rightarrow 3\ SO_4^{2-} + 3\ H^+ + I^-$

c. $3\ e^- + 4\ H_2O + CrO_4^{2-} \rightarrow Cr(OH)_4^- + 4\ OH^-$ (reduction)

$3\ (Fe(OH)_2 + OH^- \rightarrow Fe(OH)_3 + e^-)$ (oxidation)

$Fe(OH)_2 + + 4\ H_2O + CrO_4^{2-} \rightarrow Fe(OH)_3 + Cr(OH)_4^- + OH^-$

d. $Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$

$Cl_2 + 2\ H_2O + 4\ OH^- \rightarrow 2\ ClO^- + 4\ H_2O + 2\ e^-$

$2\ Cl_2 + 4\ OH^- \rightarrow 2\ ClO^- + 2\ H_2O + 2\ Cl^-$

$Cl_2 + 2\ OH^- \rightarrow ClO^- + Cl^- + H_2O$

97. $4\ (Fe(OH)_2 + OH^- \rightarrow Fe(OH)_3 + e^-)$ (oxidation)

$O_2 + 4\ H_2O + 4\ e^- \rightarrow 2\ H_2O + 4\ OH^-$ (reduction)

$4\ Fe(OH)_2 + O_2 + 2\ H_2O \rightarrow 4\ Fe(OH)_3$

99. Ca (s) | Ca^{2+} (aq) || Cl^- (aq) | Cl_2 (g) | Pt

Mg (s) → Mg^{2+} (aq) + 2 e^- $E^\circ_{ox} = +2.76$ V

Cl_2 (g) + 2 e^- → 2 Cl^- (aq) $E^\circ_{red} = +1.36$ V

Mg (s) + Cl_2 (g) → Mg^{2+} (aq) + 2 Cl^- (aq) $E^\circ_{cell} = +4.12$ V

101. a. $2\ Fe^{3+}\ (aq) + 2\ e^- \rightarrow 2\ Fe^{2+}\ (aq)$ $E^\circ_{red} = +0.77$ V

$Ni\ (s) \rightarrow Ni^{2+}\ (aq) + 2\ e^-$ $E^\circ_{ox} = +0.23$ V

$Ni\ (s) + 2\ Fe^{3+}\ (aq) \rightarrow 2\ Fe^{2+}\ (aq) + Ni^{2+}\ (aq)$ $E^\circ_{cell} = +1.00$ V

Since $E^\circ_{cell} > 0$, the reaction is spontaneous under standard conditions.

$2\ Fe^{3+}\ (aq) + 2\ e^- \rightarrow 2\ Fe^{2+}\ (aq)$ $E^\circ_{red} = +0.77$ V

$Sn^{2+}\ (s) \rightarrow Sn^{4+}\ (aq) + 2\ e^-$ $E^\circ_{ox} = -0.15$ V

$Ni\ (s) + 2\ Fe^{3+}\ (aq) \rightarrow 2\ Fe^{2+}\ (aq) + Ni^{2+}\ (aq)$ $E^\circ_{cell} = +0.62$ V

Since $E^\circ_{cell} > 0$, the reaction is spontaneous under standard conditions.

103. $$[Pb^{2+}] = \frac{K_{sp}}{[SO_4^{2-}]} = \frac{1.7 \times 10^{-8}}{1.0} = 1.7 \times 10^{-8}\ M$$

$Pb\ (s) \rightarrow Pb^{2+}\ (1.7 \times 10^{-8}\ M) + 2\ e^-$ $E^\circ_{ox} = +0.13$ V

$2\ H^+\ (1.0\ M) + 2\ e^- \rightarrow H_2\ (1\ atm)$ $E^\circ_{red} = +0.00$ V

$2\ H^+\ (1.0\ M) + Pb\ (s) \rightarrow Pb^{2+}\ (1.7 \times 10^{-8}\ M) + H_2\ (1\ atm)$ $E^\circ_{cell} = +0.13$ V

$$E = E^\circ - \frac{0.0591}{n} \log \frac{[Pb^{2+}]\,P_{H_2}}{[H^+]^2}$$
$$= 0.13\ V - \frac{0.0591}{2} \log \frac{(1.7 \times 10^{-8})(1.0)}{(1.0)^2}$$
$$= 0.13\ V - (-0.230\ V) = 0.36\ V$$

Note that this corresponds to E°_{ox} for the reaction

$Pb\ (s) + SO_4^{2-} \rightarrow PbSO_4\ (s) + 2\ e^-$ $E^\circ_{ox} = +0.355$ V

tabulated in more extensive tables.

105. a. $$\ln K = \frac{nFE°}{RT}$$

$$\ln k = \frac{(2\ mol)(96485\ \frac{J}{V \cdot mol})(0.010\ V)}{(8.314\ \frac{J}{mol \cdot K})(298\ K)}$$

ln k = 0.7$\underline{7}$8

k = $e^{0.7\underline{7}8}$ = 2.2

b.

	Sn (s) +	Pb^{2+} (aq)	⇌	Sn^{2+} (aq)	+ Pb (s)
$[\ ]_i$		1.0		0	
$\Delta[\ ]$		−x		+x	
$[\ ]_{eq}$		1.0 − x		x	

$$K = \frac{[Sn^{2+}]}{[Pb^{2+}]} = \frac{x}{1.0 - x} = 2.2$$

3.2 x = 2.2

x = 0.69

$[Pb^{2+}]$ = 1.0 − 0.69 = 0.3 *M*

107. a. $Na^+ + e^- \rightarrow Na$

$$1.0\ mol\ Na^+ \times \frac{1\ F}{1\ mol\ Na^+} = 1.0\ F$$

$$1.0\ F \times \frac{96{,}485\ C}{1\ F} = 9.6 \times 10^4\ C$$

b. $Cu^{2+} + 2\ e^- \rightarrow Cu$

$$1.0\ \text{mol}\ Cu^{2+}\ x = \frac{2\ F}{1\ \text{mol}\ Cu^{2+}} = 2.0\ F$$

$$2.0\ F\ x\ \frac{96{,}485\ C}{1\ F} = 1.9\ x\ 10^5\ C$$

c. $2\ H_2O\ (l) \rightarrow O_2\ (g) + 4\ H^+(aq) + 4\ e^-$

$$1.0\ g\ H_2O\ x\ \frac{1\ \text{mol}\ H_2O}{18.02\ g}\ x\ \frac{4\ F}{2\ \text{mol}\ H_2O} = 0.11\ F$$

$$0.11\ F\ x\ \frac{96{,}485\ C}{1\ F} = 1.1\ x\ 10^4\ C$$

d. $2\ Cl^- \rightarrow Cl_2 + 2e^-$

$$1.0\ g\ Cl^-\ x\ \frac{1\ \text{mol}\ Cl^-}{35.453\ g}\ x\ \frac{2\ F}{2\ \text{mol}\ Cl^-} = 0.028\ F$$

$$0.028\ F\ x\ \frac{96{,}485\ C}{1\ F} = 2.7\ x\ 10^3\ C$$

109. $$65.4\ s\ x\ 10.5\ mA\ x\ \frac{1\ A}{1000\ mA}\ x\ \frac{1\ C}{1\ A\cdot s}\ x\ \frac{1\ F}{96{,}485\ C} = 7.11\ x\ 10^{-6}\ F$$

$2\ I^- \rightarrow I_2 + 2e^-$

$$7.11\ x\ 10^{-6}\ F\ x\ \frac{1\ \text{mol}\ I_2}{2\ F}\ x\ \frac{1\ \text{mol}\ H_3AsO_3}{1\ \text{mol}\ I_2}\ x\ \frac{1\ \text{mol}\ As}{1\ \text{mol}\ H_3AsO_3}\ x\ \frac{74.92\ g\ As}{1\ \text{mol}\ As}$$

$= 2.66\ x\ 10^{-4}$ g As

111. a. $Cd\ (s) \rightarrow Cd^{2+}\ (1.0\ M) + 2\ e^-$ $E^\circ_{ox} = +0.40$ V

$2\ Ag^+\ (1.0\ M) + 2\ e^- \rightarrow 2\ Ag$ $E^\circ_{red} = +0.80$ V

$Cd\ (s) + 2\ Ag^+\ (1.0\ M) \rightarrow 2\ Ag\ (s) + Cd^{2+}\ (1.0\ M)$ $E^\circ_{cell} = +1.20$ V

b. The addition of S^{2-} would drastically reduce the Cd^{2+} concentration, making E_{ox} and therefore E_{cell} more positive.

c. The size of the electrode does not affect the potential.

113. The decrease in the concentration of Cu^{2+} and increase in the mass of the electrode implies that Cu^{2+} is being reduced to Cu (s). This electrode is therefore the cathode. The anode reaction is the oxidation of water to O_2 (g) which has H^+ as a product accounting for the increase in the hydronium ion concentration.

Cathode: $Cu^{2+}\text{(aq)} + 2\,e^- \rightarrow Cu\text{(s)}$

Anode: $2\,H_2O\text{(l)} \rightarrow O_2\text{(g)} + 4\,H^+\text{(aq)} + 4\,e^-$

115. a. $$2.48\text{ g Ag} \times \frac{1\text{ mol Ag}}{107.87\text{ g}} \times \frac{1\text{ F}}{1\text{ mol Ag}} \times \frac{96{,}485\text{ C}}{1\text{ F}} = 2.2\underline{1}8 \times 10^3\text{ C}$$

$$t = \frac{Q}{i} = \frac{2.2\underline{1}8 \times 10^3\text{ C}}{1.50\ \frac{\text{C}}{\text{s}}} = 1.48 \times 10^3\text{ s}$$

b. $$2.2\underline{1}8 \times 10^3\text{ C} \times \frac{1\text{ F}}{96{,}485\text{ C}} \times \frac{1\text{ mol Cr}}{3\text{ F}} \times \frac{51.996\text{ g Cr}}{1\text{ mol Cr}} = 0.398\text{ g Cr}$$

117. a. $$2.75\text{ A} \times 3.50\text{ hr} \times \frac{3600\text{ s}}{1\text{ hr}} \times \frac{1\text{ C}}{1\text{ A}\cdot\text{s}} \times \frac{1\text{ F}}{96{,}485\text{ C}} \times 0.900 = 0.323\text{ F}$$

The factor of 0.900 accounts for the 90.0% current efficiency.

b. $$21.221\text{ g Au} \times \frac{1\text{ mol Au}}{196.9665\text{ g}} = 0.10774\text{ mol Au}$$

$$n = \frac{0.323\text{ F}}{0.10774\text{ mol Au}} = 3.00\ \frac{\text{F}}{\text{mol Au}}$$

The half reaction is: $Au^{n+}\text{(aq)} + n\,e^- \rightarrow Au\text{(s)}$ Since n = 3, the charge on the gold ions is 3+

119. a. Cu^{2+} (0.0100 M) + 2 e^- → Cu (s) E°_{red} = +0.34 V

Zn (s) → Zn^{2+} (0.20 M) + 2 e^- E°_{ox} = +0.76 V

Cu^{2+} (0.0100 M) + Zn (s) → Cu (s) + Zn^{2+} (0.20 M) E°_{cell} = +1.10 V

$$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{0.200}{0.0100} = 20.0$$

$$E = E^\circ - \frac{RT}{nF}\ln Q$$
$$= 1.10\text{ V} - \frac{0.0591}{2}\log 20.0$$
$$= 1.10\text{ V} - (0.0332\text{ V}) = 1.06\text{ V}$$

b. $$1.00\text{ A} \times 225\text{ s} \times \frac{1\text{ C}}{1\text{ A}\cdot\text{s}} \times \frac{1\text{ F}}{96{,}485\text{ C}} \times \frac{1\text{ mol Cu}}{2\text{ F}} = 1.17 \times 10^{-3}\text{ mol Cu}$$

$$1.00\text{ L} \times \frac{0.0100\text{ mol Cu}}{1\text{ L}} = 0.0100\text{ mol Cu}$$ initially

moles Cu remaining = 0.0100 − 1.17 x 10^{-3} = 0.008$\underline{8}$ mol

$$[Cu^{2+}] = \frac{0.008\underline{8}\text{ mol Cu}}{1.00\text{ L}} = 0.0088\ M\ \ Cu^{2+}$$

121. a. Co^{2+} (aq) + Cd (s) → Co (s) + Cd^{2+} (aq)

ΔG° = [1 mol (ΔG°_f (Cd^{2+} (aq))) + 1 mol (ΔG°_f (Co (s)))]

− [1 mol (ΔG°_f (Cd (s))) + 1 mol (ΔG°_f (Co^{2+} (aq)))]

ΔG° = [1 mol (0 kJ · mol^{-1}) + 1 mol (−77.6 kJ · mol^{-1})]

−[1 mol (0 kJ · mol^{-1}) + 1 mol (−54.4 kJ · mol^{-1})]

ΔG° = −23.2 kJ = −23,$\underline{2}$00 J b.

$$E° = \frac{-\Delta G°}{nF} = \frac{-(-23{,}200\ J)}{2\ mol\left(96485\ \frac{J}{V \cdot mol}\right)} = 0.120\ V$$

cathode:	$Co^{2+}\ (aq) + 2\ e^- \rightarrow Co\ (s)$	$E°_{red}$ = ? V
anode:	$Cd\ (s) \rightarrow Cd^{2+}\ (aq) + 2\ e^-$	$E°_{ox}$ = +0.40 V
net:	$Co^{2+}\ (aq) + Cd\ (s) \rightarrow Co\ (s) + Cd^{2+}\ (aq)$	$E°_{cell}$ = +0.120 V

0.120 V = 0.40 V + $E°_{red}$

$E°_{red}$ = 0.120 V − 0.40 V = −0.28 V