Use the following table of standard reduction potentials:
http://www.jesuitnola.org/upload/clark/refs/red_pot.htm

1. Consider the redox reaction: $VO^{2+}(aq) + MnO_4^-(aq) \rightarrow MnO_2(s) + VO_2^+(aq)$
   a. Balance it in acid
      $H_2O(l) + 3\ VO^{2+}(aq) + MnO_4^-(aq) \rightarrow MnO_2(s) + 3\ VO_2^+(aq) + 2\ H^+(aq)$
   b. Balance it in base
      $2\ OH^-(aq) + 3\ VO^{2+}(aq) + MnO_4^-(aq) \rightarrow MnO_2(s) + 3\ VO_2^+(aq) + H_2O(l)$
   c. Calculate $E°_{cell}$
      $E°_{cell} = E°_{cathode} - E°_{anode} = 0.59 - 1.00 = -0.41$

2. Consider the redox reaction: $BiO^+(aq) + 2Cr^{3+}(aq) \rightarrow Bi(s) + Cr_2O_7^{2-}(aq)$
   a. Balance it in acid
      $5\ H_2O(l) + 2\ BiO^+(aq) + 2Cr^{3+}(aq) \rightarrow 2\ Bi(s) + Cr_2O_7^{2-}(aq) + 10\ H^+(aq)$
   b. Balance it in base
      $10\ OH^- + 2\ BiO^+(aq) + 2Cr^{3+}(aq) \rightarrow 2\ Bi(s) + Cr_2O_7^{2-}(aq) + 5\ H_2O(l)$
   c. Calculate $E°_{cell}$
      $E°_{cell} = E°_{cathode} - E°_{anode} = 0.32 - 1.33 = -1.01$

3. Consider the redox reaction: $Cd(s) + Co^{3+}(aq) \rightarrow Cd^{2+}(aq) + Co^{2+}(aq)$
   a. Balance it
      $Cd(s) + 2\ Co^{3+}(aq) \rightarrow Cd^{2+}(aq) + 2\ Co^{2+}(aq)$
   b. Calculate $E°_{cell}$
      $E°_{cell} = E°_{cathode} - E°_{anode} = 1.842 - {}^-0.403 = 2.445$

4. Consider the redox reaction: $Fe^{2+}(aq) \rightarrow Fe(s) + Fe^{3+}(aq)$
   a. Balance it
      $3\ Fe^{2+}(aq) \rightarrow Fe(s) + 2\ Fe^{3+}(aq)$
   b. Calculate $E°_{cell}$
      $E°_{cell} = E°_{cathode} - E°_{anode} = -0.440 - 0.771 = -1.211$

5. What effect will raising the pH have on $E°_{cell}$ for
   a. the reacion in question 1 above?
      it will increase $E°_{cell}$
   b. the reacion in question 2 above?
      it will increase $E°_{cell}$

6. Consider the answer to 3b above, If $[Co^{3+}]$ and$[Co^{2+}]$ were both 1 M, what concentration of $[Cd^{2+}]$ could be used to make a 2.5 V battery based on this reaction?
   $E_{cell} = E°_{cell} - (0.05916/n_e)\log Q$
   $2.5 = 2.445 - (0.05916/2)\log([Cd^{2+}]\cdot 1/1)$
   $0.005 = -(0.05916/2)\log([Cd^{2+}])$
   $-0.169 = \log([Cd^{2+}])$
   $[Cd^{2+}] = 0.678\ M$

7. Consider the answer to 4b above. Using the Nernst equation, what ratio of $[Fe^{3+}]$ and $[Fe^{2+}]$ would be present if this cell were allowed to reach equilibrium at room temperature?
   $E_{cell} = E°_{cell} - (0.05916/n_e)\log Q = 0$ at equilibrium
   $0 = -1.211 - (0.05916/2)\log Q$
   $1.211 = -(0.05916/2)\log Q$

-40.94 = logQ

$Q = [Fe^{3+}]/[Fe^{2+}] = 1.15 \times 10^{-41}$

8. If $E°_{cell}$ for a particular reaction is negative,

a. what are two things one could do to make the E_{cell} positive?

You could lower the concentration of aqueous species or gases on the right side of the equation or raise the concentration of aqueous species or gases on the left side of the equation.

b. what is one thing you could do to make $E°_{cell}$ positive?

Reverse the cell.

9. Rank the following species from worst to best oxidizing agent: $PbSO_4$, Na^+, Fe^{3+}, Cl_2 and AgCl.

Na^+ $PbSO_4$ AgCl Fe^{3+} Cl_2

10. Rank the following species from best to worst reducing agent: Mn^{2+}, Ni, H_2, Al, Hg_2^{2+}.

Al Ni H2 Hg_2^{2+} Mn^{2+}

11. What will be produced at the cathode

a. for the reaction in question 1?

$MnO_2(s)$

b. for the reaction in question 2?

Bi(s)

c. for the reaction in question 3?

$Co^{2+}(aq)$

d. for the reaction in question 4?

Fe(s)

12. What will be consumed at the anode

a. for the reaction in question 1?

$VO^{2+}(aq)$

b. for the reaction in question 2?

$2Cr^{3+}(aq)$

c. for the reaction in question 3?

Cd(s)

d. for the reaction in question 4?

$Fe^{2+}(aq)$

12. Calculate ΔG

a. for the reaction in question 1

$\Delta G = -nFE = -3 \cdot 96{,}485 \cdot -0.41 = 118{,}677\ J \cdot mol^{-1} = 118.7\ kJ \cdot mol^{-1}$

b. for the reaction in question 2

$\Delta G = -nFE = -6 \cdot 96{,}485 \cdot -1.01 = 584{,}699\ J \cdot mol^{-1} = 584.7\ kJ \cdot mol^{-1}$

c. for the reaction in question 3

$\Delta G = -nFE = -2 \cdot 96{,}485 \cdot 2.455 = -473{,}741\ J \cdot mol^{-1} = -473.7\ kJ \cdot mol^{-1}$

d. for the reaction in question 4

$\Delta G = -nFE = -2 \cdot 96{,}485 \cdot -1.211 = 233{,}687\ J \cdot mol^{-1} = 233.7\ kJ \cdot mol^{-1}$

13. Calculate K

a. for the reaction in question 1

$K = e^{(nFE/RT)} = e^{(3 \cdot 96{,}485 \cdot -0.41/8.314 \cdot 298)} = 1.58 \times 10^{-21}$

b. for the reaction in question 2

$K = e^{(nFE/RT)} = e^{(6 \cdot 96{,}485 \cdot -1.01/8.314 \cdot 298)} = 3.22 \times 10^{-103}$

c. for the reaction in question 3

K = e^(nFE/RT) = e^(2·96,485·2.455/8.314·298) = 1.10×10^{83}

d. for the reaction in question 4

K = e^(nFE/RT) = e^(2·96,485·-1.211/8.314·298) = 1.09×10^{-41}

14. If a current of 1 ampere can plate 3 moles of a certain product in 10 hours,

a. how much of the same product could it plate in 20 minutes?

(3 moles / 10 hours) x 1/3 hours = 0.1 moles

b. what current would be needed to plate 6 moles in just 2 hours?

1 ampere x (10 hours / 3 moles) x (6 moles / 2 hours) = 10 amperes

15. If a 5.5845 g sample of Fe is plated from an unknown solution in 482 minutes at a constant current of 1 ampere, what was the oxidation state of the Fe ion is solution?

5.5845 g Fe x (1 mol / 55.845 g) = 0.1 moles Fe

482 minutes x (60 seconds / minute) = 28,920

$(I \cdot t)/(n_e \cdot F)$ = moles

$(1 \cdot 28{,}920)/(n_e \cdot 96485) = 0.1$

$n_e = 3$, Therefore, the solution contained Fe^{3+}.

16. Complete the table below from memory

	E	ΔG	K	Reduction	Oxidation	Cathode	Anode
Battery	+	-	K>1	Cathode	Anode	+	-
Electrolytic	-	+	0<K<1	Cathode	Anode	-	+

17. When electroplating metal (producing a solid metal from an ion by an electrolytic reaction), the amount of metal produced is

a. directly proportional to
current (I) and time (t)

b. inversely proportional to
number of electrons (n)

18. Using the table mentioned at the top of this document, what species could

a. reduce $O_2(g)$ to $H_2O(l)$ bot not $NO_3^-(aq)$ to NO(g)

$Br^-(aq)$

b. reduce $Cu^{2+}(aq)$ to $Cu^+(aq)$ but not $Pb^{2+}(aq)$ to Pb(s)

$Sn^{2+}(aq)$ or $H_2(g)$

c. oxidize Ag(s) to $Ag^+(aq)$ but not $H_2O_2(aq)$ to $O_2(g)$

$Fe^{3+}(aq)$

19. Determine the number of electrons gained/lost by the carbon atom in each reaction below

a. $CH_4 + 2\ O_2 \square + CO_2 + 2\ H_2O$

8 electrons lost

b. $CH_4 + O_2 \square + CH_2O + H_2O$

4 electrons lost

c. $2\ CH_4 + O_2 \square + 2\ CH_3OH$

2 electrons lost

20. Consider the answer to question 19 above. How does a breaking a C-H bond and forming a C-O bond effect the oxidation state of C?

It riases the oxidation number of the carbon by two.