Ch 19.6-19.7: Batteries and Corrosion
Homework 19-5: pg 834 #36, 38, 40, 41, 42

36 Discuss the advantages and disadvantages of fuel cells over conventional power plants in producing electricity.

Advantages of a Fuel Cell compared to conventional power plants
-- More efficient(70% instead of 40%)
-- no noise, no vibration , no heat transfer, no thermal pollution
Disadvantages of Fuel Cell:
--Needed electrocatalysts are unable to function efficiently over long periods of time without getting contaminated

38 Calculate the standard emf of the propane fuel cell discussed on p. 820 at 25°C, given that ΔG_f^0 for propane is –23.5 kJ/mol.

We can calculate the standard free energy change, $\Delta G°$, from the standard free energies of formation, ΔG_f° using Equation (18.12) of the text. Then, we can calculate the standard cell emf, E_{cell}°, from $\Delta G°$.

The overall reaction is:

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$$

$$\Delta G_{rxn}^\circ = 3\Delta G_f^\circ[CO_2(g)] + 4\Delta G_f^\circ[H_2O(l)] - \{\Delta G_f^\circ[C_3H_8(g)] + 5\Delta G_f^\circ[O_2(g)]\}$$

$$\Delta G_{rxn}^\circ = (3)(-394.4\ kJ/mol) + (4)(-237.2\ kJ/mol) - [(1)(-23.5\ kJ/mol) + (5)(0)] = -2108.5\ kJ/mol$$

We can now calculate the standard emf using the following equation:

$$\Delta G° = -nFE_{cell}^\circ$$

so

$$E_{cell}^\circ = \frac{-\Delta G°}{nF}$$

Check the half-reactions on p. 820 of the text to determine that 20 moles of electrons are transferred during this redox reaction.

$$E_{cell}^\circ = \frac{-(-2108.5 \times 10^3\ J/mol)}{(20)(96500\ J/V \cdot mol)} = \mathbf{1.09\ V}$$

Does this suggest that, in theory, it should be possible to construct a galvanic cell (battery) based on any conceivable spontaneous reaction?

40 "Galvanized iron" is steel sheet that has been coated with zinc; "tin" cans are made of steel sheet coated with tin. Discuss the functions of these coatings and the electrochemistry of the corrosion reactions that occur if an electrolyte contacts the scratched surface of a galvanized iron sheet or a tin can.

Galvanized iron sheet: The Zn protects the iron from oxidizing because the Zn is more easily oxidized than the Fe. Thus, the Zn will be the anode (ox) and the Fe acts as the cathode (red). Thus, the iron will not rust even if the Zn is scratched away at spots.
Tin can: The tin coating on the iron only protects the iron from oxidizing as long as the tin coating keeps the iron from being exposed. If the iron is exposed, the iron will oxidize because it is more easily oxidized than the tin.

41 **Tarnished silver contains Ag_2S. The tarnish can be removed by placing silverware in an aluminum pan containing an inert electrolyte solution, such as NaCl. Explain the electrochemical principle for this procedure. [The standard reduction potential for the half-cell reaction $Ag_2S(s) + 2e^- \rightarrow 2\ Ag(s) + S^{2-}(aq)$ is –0.71 V.]**

Anode (ox): $Al \rightarrow Al^{3+} + 3\ e^-$ E°_{anode}: –1.66 V

Cathode (red): $Ag_2S\ (s) + 2\ e^- \rightarrow 2\ Ag\ (s) + S^{2-}$ E°_{cath}: –0.71 V

Al is oxidized and Ag_2S is reduced. The E° = –0.71 V – (–1.66 V) = 0.95 V, so it is a spontaneous reaction.

42 **How does the tendency of iron to rust depend on the pH of the solution?**

Iron has a higher tendency to rust in more acidic solutions (lower pH). This is because E° for the reduction of O_2 gets more positive (hence more spontaneous) as $[H^+]$ increases and pH decreases. See partial reproduction of Table 19.1 of the text at right. This can also be explained by the consumption of OH^- by the H^+ as it is produced, reducing the concentration of the products of the reaction and causing the reaction to become more spontaneous.

$Cl_2\ (g) + 2e^-$	→	$2Cl^-\ (aq)$	+1.36
$O_2\ (g) + 4H^+\ (aq) + 4e^-$	→	$2H_2O\ (l)$	+1.23
$Pt^{2+}\ (aq) + 2e^-$	→	$Pt\ (s)$	+1.18
$Br_2\ (l) + 2e^-$	→	$2Br^-\ (aq)$	+1.07
$2Hg^{2+}\ (aq) + 2e^-$	→	$Hg_2^{2+}\ (aq)$	+0.92
$Hg^{2+}\ (aq) + 2e^-$	→	$Hg\ (l)$	+0.85
$Ag^+\ (aq) + e^-$	→	$Ag\ (s)$	+0.80
$Fe^{3+}\ (aq) + e^-$	→	$Fe^{2+}\ (aq)$	+0.77
$I_2\ (s) + 2e^-$	→	$2I^-\ (aq)$	+0.54
$Cu^+\ (aq) + e^-$	→	$Cu\ (s)$	+0.52
$O_2\ (g) + 2H_2O\ (l) + 4e^-$	→	$4OH^-\ (aq)$	+0.40
$Cu^{2+}\ (aq) + 2e^-$	→	$Cu\ (s)$	+0.34