Chapter 20 # 27, 37, 35, ...

Name _____ # ____

20.27 A voltaic cell similar to that shown in Figure 20.5 is constructed. One electrode half-cell consists of a silver strip placed in a solution of $AgNO_3$, and the other has an iron strip placed in a solution of $FeCl_2$. The overall cell reaction is

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

(a) What is being oxidized, and what is being reduced?

(b) Write the half-reactions that occur in the two half-cells.
(c) Which electrode is the anode, and which is the cathode?

(d) Indicate the signs of the electrodes. (e) Do electrons flow from the silver electrode to the iron electrode or from the iron to the silver?

(f) In which directions do the cations and anions migrate through the solution?

20.37 Using standard reduction potentials (Appendix E), calculate the standard emf for each of the following reactions:

> on #37, ALSO Calculate ΔG°_{rxn} for each Rxn.

(a) $Cl_2(g) + 2\,I^-(aq) \longrightarrow 2\,Cl^-(aq) + I_2(s)$

(c) $Fe(s) + 2\,Fe^{3+}(aq) \longrightarrow 3\,Fe^{2+}(aq)$

(d) $2\,NO_3^-(aq) + 8\,H^+(aq) + 3\,Cu(s) \longrightarrow 2\,NO(g) + 4\,H_2O(l) + 3\,Cu^{2+}(aq)$

20.35 A voltaic cell that uses the reaction

$$Tl^{3+}(aq) + 2\,Cr^{2+}(aq) \longrightarrow Tl^{+}(aq) + 2\,Cr^{3+}(aq)$$

has a measured standard cell potential of + 1.19 V. (a) Write the two half-cell reactions.

(b) By using data from Appendix E, determine E°_{red} for the reduction of $Tl^{3+}(aq)$ to $Tl^{+}(aq)$.

(c) Sketch the voltaic cell, label the anode and cathode, and indicate the direction of electron flow.

20.39 The standard reduction potentials of the following half-reactions are given in Appendix E: And here: E°_{red} (V)

$Ag^{+}(aq) + e^{-} \longrightarrow Ag(s)$ + 0.799

$Cu^{2+}(aq) + 2\,e^{-} \longrightarrow Cu(s)$ + 0.337

$Ni^{2+}(aq) + 2\,e^{-} \longrightarrow Ni(s)$ − 0.28

$Cr^{3+}(aq) + 3\,e^{-} \longrightarrow Cr(s)$ −0.74

(a) Determine which combination of these half-cell reactions leads to the cell reaction with the largest positive cell potential and calculate the value.

20.41 A 1 *M* solution of $Cu(NO_3)_2$ is placed in a beaker with a strip of Cu metal. A 1 *M* solution of $SnSO_4$ is placed in a second beaker with a strip of Sn metal. A salt bridge connects the two beakers, and wires to a voltmeter link the two metal electrodes. (a) Which electrode serves as the anode and which as the cathode? (b) Which electrode gains mass and which loses mass as the cell reaction proceeds? (c) Write the equation for the overall cell reaction. (d) What is the emf generated by the cell under standard conditions?

a)

b)

c)

d)

20.55 Using the standard reduction potentials listed in Appendix E, calculate the equilibrium constant for each of the following reactions at 298 K:

(a) $Fe(s) + Ni^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Ni(s)$

(b) $Co(s) + 2\,H^{+}(aq) \longrightarrow Co^{2+}(aq) + H_2(g)$

(c) $10\,Br^{-}(aq) + 2\,MnO_4^{-}(aq) + 16\,H^{+}(aq) \longrightarrow 2\,Mn^{2+}(aq) + 8\,H_2O(l) + 5\,Br_2(l)$

20.31 (a) Write the half-reaction that occurs at a hydrogen electrode in acidic aqueous solution when it serves as the cathode of a voltaic cell. (b) What is *standard* about the standard hydrogen electrode? (c) What is the role of the platinum foil in a standard hydrogen electrode?

a)

b)

c)

20.33 ~~(a) What is a *standard reduction potential*?~~ (b) What is the standard reduction potential of a standard hydrogen electrode?

____ Volts

✳ 20.59 A voltaic cell is based on the reaction

$$Sn(s) + I_2(s) \longrightarrow Sn^{2+}(aq) + 2\,I^-(aq)$$

Under standard conditions, what is the maximum electrical work, in joules, that the cell can accomplish if 75.0 g of Sn is consumed?

* The answer to #59 is 81900 J (the answer in the textbook is incorrect)

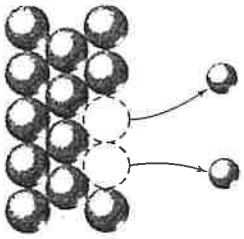
20.7 Consider a redox reaction for which $E°$ is a negative number.
(a) What is the sign of $\Delta G°$ for the reaction?
(b) Will the equilibrium constant for the reaction be larger or smaller than 1?
(c) Can an electrochemical cell based on this reaction accomplish work on its surroundings? [Section 20.5]

a) _____

b) _____

c) _____

20.3 The diagram that follows represents a molecular view of a process occurring at an electrode in a voltaic cell.



(a) Does the process represent oxidation or reduction?

(b) Is the electrode the anode or cathode?

(c) Why are the atoms in the electrode represented by larger spheres than the ions in the solution? [Section 20.3]

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20.44 From each of the following pairs of substances, use data in Appendix E to choose the one that is the stronger oxidizing agent:

(a) $Cl_2(g)$ or $Br_2(l)$

(b) $Zn^{2+}(aq)$ or $Cd^{2+}(aq)$

20.43 From each of the following pairs of substances, use data in Appendix E to choose the one that is the stronger reducing agent:

(a) Fe(*s*) or Mg(*s*)

(b) Ca(*s*) or Al(*s*)

20.6 Consider the following table of standard electrode potentials for a series of hypothetical reactions in aqueous solution:

Reduction Half-Reaction	$E°$(V)
$A^+(aq) + e^- \longrightarrow A(s)$	1.33
$B^{2+}(aq) + 2e^- \longrightarrow B(s)$	0.87
$C^{3+}(aq) + e^- \longrightarrow C^{2+}(aq)$	−0.12
$D^{3+}(aq) + 3e^- \longrightarrow D(s)$	−1.59

(a) Which substance is the strongest oxidizing agent? Which is weakest?

(b) Which substance is the strongest reducing agent? Which is weakest?

(c) Which substance(s) can oxidize C^{2+}?

20.89 (a) What is *electrolysis*?

(b) Are electrolysis reactions thermodynamically spontaneous? Explain.

(c) What process occurs at the anode in the electrolysis of molten NaCl?

(d) Why is sodium metal not obtained when an aqueous solution of NaCl undergoes electrolysis? Support your answer with voltages.

20.91 (a) A $Cr^{3+}(aq)$ solution is electrolyzed, using a current of 7.60 A. What mass of Cr(*s*) is plated out after 2.00 days?

(b) What amperage is required to plate out 0.250 mol Cr from a Cr^{3+} solution in a period of 8.00 h?

[20.108] (a) How many coulombs are required to plate a layer of chromium metal 0.25 mm thick on an auto bumper with a total area of 0.32 m^2 from a solution containing CrO_4^{2-}? The density of chromium metal is 7.20 g/cm^3.

(b) What current flow is required for this electroplating if the bumper is to be plated in 10.0 s?

20.63 What is the effect on the emf of the cell shown in Figure 20.9, which has the overall reaction $Zn(s) + 2\,H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$, for each of the following changes? (a) The pressure of the H_2 gas is increased in the cathode half-cell. (b) Zinc nitrate is added to the anode half-cell. (c) Sodium hydroxide is added to the cathode half-cell, decreasing $[H^+]$. (d) The surface area of the anode is doubled.

a.

b.

c.

d.

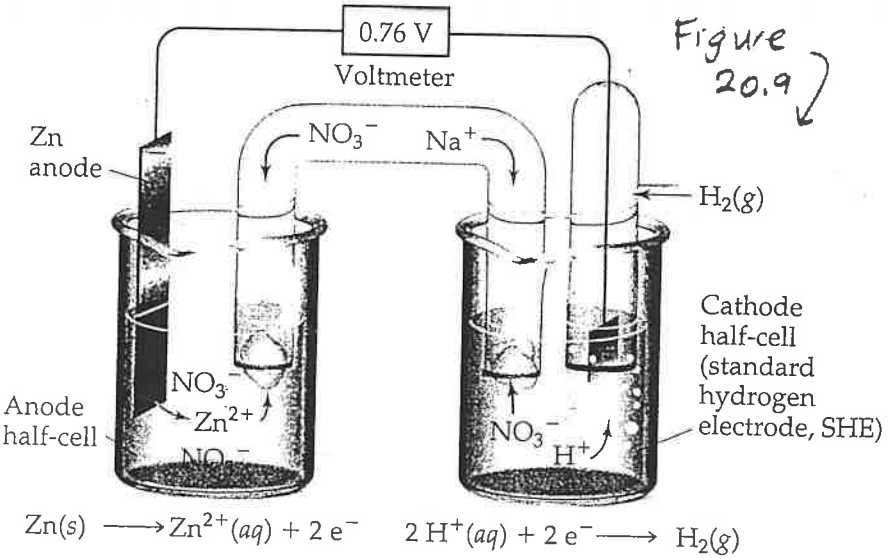

Figure 20.9

20.66 A voltaic cell utilizes the following reaction and operates at 298 K:

$$3\,Ce^{4+}(aq) + Cr(s) \longrightarrow 3\,Ce^{3+}(aq) + Cr^{3+}(aq)$$

(a) What is the emf of this cell under standard conditions?

(b) What is the emf of this cell when $[Ce^{4+}] = 3.0\,M$, $[Ce^{3+}] = 0.10\,M$, and $[Cr^{3+}] = 0.010\,M$?

(c) What is the emf of the cell when $[Ce^{4+}] = 0.010\,M$, $[Ce^{3+}] = 2.0\,M$, and $[Cr^{3+}] = 1.5\,M$?

20.72 A voltaic cell is constructed that is based on the following reaction:

$$Sn^{2+}(aq) + Pb(s) \longrightarrow Sn(s) + Pb^{2+}(aq)$$

(a) If the concentration of Sn^{2+} in the cathode half-cell is 1.00 M and the cell generates an emf of +0.22 V, what is the concentration of Pb^{2+} in the anode half-cell?

(b) If the anode half-cell contains $[SO_4^{2-}] = 1.00\ M$ in equilibrium with $PbSO_4(s)$, what is the K_{sp} of $PbSO_4$?

20.102 A voltaic cell is constructed from an $Ni^{2+}(aq)$-$Ni(s)$ half-cell and an $Ag^+(aq)$-$Ag(s)$ half-cell. The initial concentration of $Ni^{2+}(aq)$ in the Ni^{2+}-Ni half-cell is $[Ni^{2+}] = 0.0100\ M$. The initial cell voltage is +1.12 V. (a) By using data in Table 20.1, calculate the standard emf of this voltaic cell. or Appendix E

(b) Will the concentration of $Ni^{2+}(aq)$ increase or decrease as the cell operates?

(c) What is the initial concentration of $Ag^+(aq)$ in the Ag^+-Ag half-cell?

20.69 A voltaic cell is constructed with two Zn^{2+}–Zn electrodes. The two half-cells have $[Zn^{2+}] = 1.8\ M$ and $[Zn^{2+}] = 1.00 \times 10^{-2}\ M$, respectively. (a) Which electrode is the anode of the cell?

(b) What is the standard emf of the cell? ———

(c) What is the cell emf for the concentrations given?

(d) For each electrode, predict whether $[Zn^{2+}]$ will increase, decrease, or stay the same as the cell operates.

anode: _____ cathode: _____

Problem X. (Optional review for the redox balancing quiz!)

For each reaction, Write the balanced oxidation half reaction.
Write the balanced reduction half reaction.
Write the overall balanced redox reaction.

a. $Cr + Fe^{+3} \text{ -----> } Cr^{+2} + Fe$

oxidation: _____

reduction: _____

overall: _____

b. $Fe + Fe^{+3} \text{ ----> } Fe^{+2}$ (the Fe and the Fe^{+3} each react to form Fe^{+2}, as in $Fe + FeCl_3 \text{ ---> } FeCl_2$)

oxidation: _____

reduction: _____

overall: _____

c. (in ACID): $MnO_4^{-1} + SO_2 \text{ ------> } Mn^{+2} + HSO_4^{-1}$

oxidation: _____

reduction: _____

overall: _____

d. (in BASE) $CO + I_2O_5 \text{ -----> } CO_2 + I_2$

oxidation: _____

reduction: _____

overall: _____

e. (in BASE) $I_2 + CrO_4^{-2} \text{ -----> } Cr(OH)_3 + IO_2^{-1}$

oxidation: _____

reduction: _____

overall: _____

f. (in ACID) $HCOOH + SbO_2^{-1} \text{ -----> } CO_2 + Sb$

oxidation: _____

reduction: _____

overall: _____

g. (in BASE) $MoO_4^{-2} + H_2O_2 \text{ -----> } O_2 + Mo$

oxidation: _____

reduction: _____

overall: _____

h. $Mg + Rh^{+6} \text{ -----> } Mg^{+2} + Rh^{+2}$

oxidation: _____

reduction: _____

overall: _____