Chapter 18 Number 1 (18.1 – 18.5) (Unit 4) 23 March 2018

1. Balance the following oxidation-reduction (redox) reactions using the smallest whole number coefficients. All ionics are aqueous and acidic.

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

+5 −2; 0; +2 −2; +2

2e oxid, 6e; 3e red, 6e tot

LCM = 6

6+ 6+

$14H^+ + Cr_2O_7^{2-} + 4Cl^- \rightarrow 2Cr^{3+} + 3Cl_2(g) + 7H_2O$

LCM = 6

3e red / Cr, 6e total; 1e⁻ oxid, 6e tot

14 + 2− 6− = 6+ 6+

$3H_2O + 3S_2O_3^{2-} + 4IO_3^- \rightarrow 6SO_4^{2-} + 4I^- + 6H^+$

LCM = 24

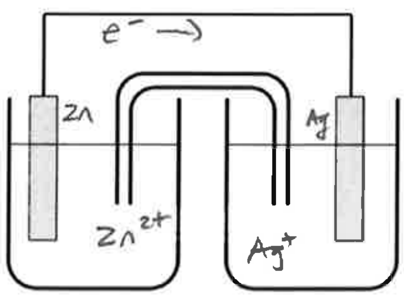
+2 +5 +6 −1

4e/S, 8e tot; 6e⁻

6− and 4− = 10− 12 − 4 − 6+ = 10−

2a. Sketch labels on the Galvanic cell shown below, with the zinc anode on the left. Labels should include all four constituents of the reaction. Indicate electron flow with an arrow.

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



2b. As the reaction proceeds, do you expect…

i. the mass of the anode electrode to increase? No

ii. the mass of the cathode electrode to increase? Yes

iii. $[Ag^+]$ to increase or decrease? No

iv. $[Zn^{2+}]$ to increase or decrease? Yes

2c. Does oxidation occurs at the anode or cathode?

2d. Balancing the reaction.

$2\,Ag^+ + Zn \rightarrow 2\,Ag + Zn^{2+}$

2e. Write the reaction in cell notation shorthand.

$Zn \,|\, Zn^{2+} \,||\, Ag^+ \,|\, Ag$

2f. Suppose the mass of the silver electrode was 2.00 g and the zinc electrode was 3.00 g and both solutions were initially 250 mL 1.00 M. After some time, the mass of the zinc electrode was found to be 1.20 g. What is the estimated mass of the silver electrode? What is $[Ag^+]$? What is $[Zn^{2+}]$?

$\Delta m_{Zn} = 3.00 - 1.20 = 1.80\,g\ Zn$

$n_{Zn} = \frac{1.80\,g}{65.38\,g}\frac{mol}{} = 0.0275\ mol\ Zn$

0.0551 mol Ag^+ ← 2

$2Ag^+ + Zn \rightarrow 2Ag + Zn^{2+}$

$m_{Ag} = 0.0551\ mol * 107.9\ g\ mol^{-1} = 5.94\,g\ Ag$; +2.00 = 7.94 g

$[Ag^+] = \frac{0.250\,mol - 0.0551\,mol}{0.250\,L} = 0.78\,M$

$[Zn^{2+}] = 1.11\ M$

Questions in final exam format:

3. Based on the balanced chemical equation shown below, determine the molarity of a solution containing Fe^{2+}, if 40.00 mL of the Fe^{2+} solution is required to completely react with 30.00 mL of a 0.125 M potassium bromate, $KBrO_3$, solution.

$6\,Fe^{2+} + BrO_3^- + 6\,H^+ \rightarrow 6\,Fe^{3+} + Br^- + 3\,H_2O$

A. 0.0156 M B. 0.0938 M
C. 0.562 M D. 1.00 M

4. During an electrochemical reaction, electrons move through the external circuit toward the ____ and positive ions in the cell move toward the ____.

A. anode, anode B. anode, cathode
C. cathode, anode D. cathode, cathode

5. What is the shorthand notation that represents the following galvanic cell reaction?

$Fe(s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$

A. $Fe(s) \,|\, Fe^{2+}(aq) \,||\, Cu^{2+}(aq) \,|\, Cu(s)$
B. $Cu(s) \,|\, Cu^{2+}(aq) \,||\, Fe^{2+}(aq) \,|\, Fe(s)$
C. $Fe(s) \,|\, NO_3^-(aq) \,||\, NO_3^-(aq) \,|\, Cu(s)$
D. $Cu(s) \,|\, Cu(NO_3)_2(aq) \,||\, Fe(NO_3)_2(aq) \,|\, Fe(s)$

Now try these problems from the book:
Section 18.1 – 18.5. Problems 1 – 11, 38, 46 – 68, even.

Chapter 18 Number 2 (18.6 – 18.9) (Unit 4) 4 April 2018

1. Balance the following redox reaction in basic solution.

$6OH^-$ $+3H_2O$ LCM $=10$

$+2$ P(s) + 5 SO_4^{2-} → 2 PO_4^{3-} + 5 SO_3^{2-} + $6H^+$ $+6OH^-$

0 +6 +5 +4

2e⁻ 10e⁻ 5e 10e⁻

$3\ 6H_2O$

$6OH^- + 2P + 5SO_4^{2-} \rightarrow 2PO_4^{3-} + 5SO_3^{2-} + 3H_2O$

16− 16−

2. Consider the reaction:

Cd(s) | Cd^{2+}(aq) || Ni^{2+}(aq) | Ni(s)

2a. Write the half reaction for the anode and look up its E° value.

$Cd \rightarrow Cd^{2+} + 2e^- \quad E^0_{ox} = +0.40\,V$

2b. Write the half reaction for the cathode and look up its E° value.

$Ni^{2+} + 2e^- \rightarrow Ni \quad E^0_{red} = -0.26\,V$

2c. Combine the two half reactions and determine the E° value for the galvanic reaction.

$Cd + Ni^{2+} \rightarrow Cd^{2+} + Ni \quad E^0 = +0.14\,V$

2d. Calculate ΔG^o_{rxn}

$\Delta G^0 = -nFE^0$

$= -2 * 96.5 * 0.14$

$= -27.0$ kJ

2e. Calculate E for

Cd(s) | Cd^{2+}(0.10 M) || Ni^{2+}(0.50 M) | Ni(s) $Q_c = \frac{[Cd^{2+}]}{[Ni^{2+}]}$

$E = E^0 - \frac{.0592}{n} \log Q_c \qquad Q_c = \frac{0.10}{0.50} = 0.20$

$= 0.14 - \frac{.0592}{2} \log(0.20)$

$= 0.16\,V$

2g. Calculate ΔG under these conditions.

$\Delta G = -nFE$

$= -2 \cdot 96.5 \cdot .16$

$= -30.9$ kJ

2g. Calculate K_c from E° and then from ΔG°

$E^0 = \frac{.0592}{n} \log K_c \Rightarrow K_c = 5.4 \times 10^4$

$\Delta G^0 = -RT \ln K_c \Rightarrow K_c = 5.4 \times 10^4$

Questions in final exam format:

3. In a galvanic cell Pb(s) | Pb^{+2}(aq) || Ag^+ | Ag(s), which electrode will gain mass?

A. the anode, Pb(s) B. the cathode, Pb(s)
C. the anode, Ag(s) D. the cathode, Ag(s)

4. The standard cell potential for the dry cell battery (such as a AA or AAA, etc.) is 1.56 V. What is the standard free energy change for this cell?

$Zn(s) + 2\ MnO_2(s) + 2\ NH_4^+ \rightarrow 2\ NH_3 + Mn_2O_3(s) + Zn^{2+} + H_2O(l)$

A. +151 kJ B. -151 kJ $\Delta G = -nFE$
C. -301 kJ D. -602 kJ

5. Doubling all the coefficients in the equation for the cell reaction...

A. doubles both E° and ΔG°.
B. doubles E°, but does not change ΔG°.
C. doubles ΔG°, but does not change E°.
D. does not change E° or ΔG°.

6. The standard potential for the following galvanic cell is +0.90 V:

$3\ Cu^{2+} + 2\ Ga(s) \rightleftharpoons 3\ Cu(s) + 2\ Ga^{3+}$

Given that the standard reduction potential for the Cu^{2+}/Cu half-cell is +0.34 V, what is the standard reduction potential for the Ga^{3+}/Ga half-cell?

A. -1.34 V B. -0.56 V
C. +0.56 V D) +1.36 V

7. A galvanic cell consists of these half-cells. What is formed at the cathode and anode, respectively?

$Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E° = +0.80\ V$

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

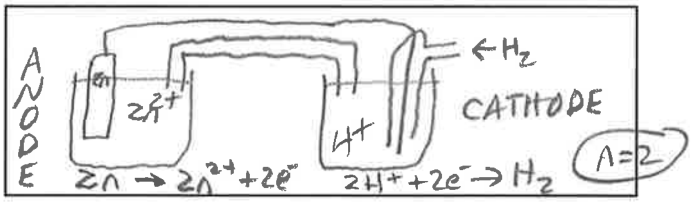
A. Ag(s) and Cu(s)
B. Ag(s) and Cu^{2+}
C. Cu(s) and Ag^+(aq)
D. Cu^{2+}(aq) and Cu(s)

Now try these problems from the book:
Section 18.6 – 18.9. Problems 12 – 23, 38, 70 – 110 even.

Chapter 18 Number 3 (18.12 – 18.14) (Unit 4) 6 April 2018

1a. Sketch the Galvanic cell:

$Zn \mid Zn^{2+}(1.0\ M) \parallel H^+(?) \mid H_2(1.0\ atm) \mid Pt$.



1b. Determine the pH for this cell if E = +0.58 V

$E^o = +0.76\,V$

$E = E^o - \frac{0.0592}{n} \log \frac{P_{H_2} * [Zn^{2+}]}{[H^+]^2}$

$0.58 = 0.76 - \frac{0.0592}{2} \log \frac{1}{[H^+]^2}$

$\log \frac{1}{[H^+]^2} = 6.08 \Rightarrow pH = 3.04$

2. Use the table of reduction potentials from your book to answer these questions. 2a. Which is the best oxidizing agent from this list?

Cd(s)	$Cd^{2+}(aq)$	$Ni^{2+}(aq)$	Ni(s)
Cu(s)	$Cu^{2+}(aq)$	$Ag^+(aq)$	Ag(s)

2b. Which is the best reducing agent from the list?

Cd

2c. Which is most easily oxidized from the list?

Cd

2d. Which is most easily reduced from the list?

Ag^+

2e. Determine E° for the most spontaneous combination of half cells from the list – which pair gives the most spontaneous value.

$2\,Ag^+ + 2e^- \rightarrow Ag \quad E^o_{red} = +0.80\,V$

$Cd \rightarrow Cd^{2+} + 2e^- \quad E^o_{ox} = +0.40\,V$

$E^o = 1.20\,V$

3. Complete:

If E° > 0, then ΔG° < 0, and K > 1

If E > 0, then ΔG < 0, and Q $< K_c$

4. In the electrolysis of an aqueous sodium sulfate solution, only hydrogen and oxygen are obtained. How many moles of hydrogen are expected if a current of 2.00 amps were applied for 1.00 hour? Start by writing the reaction and determining n.

Charge = 2.00 amps * 3600 s
= 7200 coul

$n_{e^-} = 0.0746$ mol e^-

$2H^+ + 2e^- \rightarrow H_2$

$n_{H_2} = 0.0373$ mol H_2

Questions in final exam format:

3. Using the following standard reduction potentials

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

$Pb^{2+}(aq) + 2\,e^- \rightarrow Pb(s) \quad E° = -0.13\ V$

calculate the standard cell potential for the cell reaction given below, and determine whether or not this reaction is spontaneous under standard conditions.

$Pb^{2+}(aq) + 2\,Fe^{2+}(aq) \rightarrow 2\,Fe^{3+}(aq) + Pb(s)$

A. *E°* = -0.90 V, nonspontaneous
B. *E°* = -0.90 V, spontaneous
C. *E°* = +0.90 V, nonspontaneous
D. *E°* = +0.90 V, spontaneous

4. Given:

$Ag^+(aq) + e^- \rightarrow Ag(s) \quad E° = +0.799\ V$

$AgI(s) + e^- \rightarrow Ag(s) + I^-(aq) \quad E° = -0.152\ V$

$Ni^{2+}(aq) + 2\,e^- \rightarrow Ni(s) \quad E° = -0.267\ V$

Which of the following reactions should be spontaneous under standard conditions?

I. $2\,AgI(s) + Ni(s) \rightarrow 2\,Ag(s) + 2\,I^-(aq) + Ni^{2+}(aq)$
II. $2\,Ag^+(aq) + Ni(s) \rightarrow Ag(s) + Ni^{2+}(aq)$

A. I and II are both nonspontaneous.
B. I is nonspontaneous and II is spontaneous.
C. I is spontaneous and II is nonspontaneous.
D. I and II are both spontaneous.

5. Consider the galvanic cell:

$Pt(s)|H_2(1\ atm)|H^+(1\ M)||Cl^-(1\ M)|Hg_2Cl_2\ (s)|Hg(l)$

Which one of the following changes to the cell would cause the cell potential to increase (become more positive)?

A. decrease mass of Pt
B. increase mass of Pt
C. decrease the pH
D. increase the pH

Now try these problems from the book:
Section 18.12 – 18.14. Problems 28 – 30, 122 - 132 even.