Chemistry 30 – Electrochemistry – Unit Homework

Topic	Textbook Reading	Textbook Questions
Redox Reactions Oxidation Numbers Identifying Redox Reactions	Section 20.1 (635-643)	#1, 2, 4,5
Half-Reactions Balancing Half-Reactions	Section 20.3 (650-653)	#24-26
Voltaic Cells Cell Notation Cell Potential	Section 21.1 (663-672)	#1-7

Oxidation Numbers

Oxidation Numbers

1. Identify the oxidation number for each element in the compound.
   a. $SnCl_4$
   b. Ca_3P_2
   c. SnO
   d. Ag_2S
   e. HI
   f. N_2H_4
   g. Al_2O_3
   h. S_8
   i. HNO_2
   j. O_2
   k. H_3O^+
   l. ClO_3^-
   m. $S_2O_3^{2-}$
   n. $KMnO_4$
   o. $(NH_4)_2SO_4$
2. Determine the oxidation number for carbon in each compound.
   a. CH_4
   b. CH_2O
   c. CO
   d. CO_2
3. Determine the oxidation number of nitrogen in each compound.
   a. N_2O (g)
   b. NO (g)
   c. NO_2 (g)
   d. NH_3 (g)
   e. N_2H_4 (g)
   f. $NaNO_3$ (s)
   g. N_2 (g)
   h. NH_4Cl (s)

Half-Reactions

Identifying Redox Reactions

4. For each of the following chemical reactions, assign oxidation numbers to each atom/ion and indicate whether the equation represents a redox reaction. If it does, identify the oxidation and reduction half-reactions.
   a. Cl_2 (aq) + 2 KI (aq) → I_2 (s) + 2 KCl (aq)
   b. 2 Al (s) + 3 Cl_2 (g) → 2 $AlCl_3$ (s)
   c. $Pb(NO_3)_2$ (aq) + 2 KI (aq) → PbI_2 (s) + 2 KNO_3 (aq)
   d. HCl (aq) + NaOH (aq) → H_2O (l) + NaCl (aq)
   e. 2 H_2O_2 (l) → 2 H_2O (l) + O_2 (g) *(do not write half-reactions)*

Writing Half-Reactions

5. Write a pair of balanced half-reaction equations, one showing a gain of electrons and one showing a loss, for each of the following reactions:
   a. Zn (s) + Cu^{2+} (aq) → Zn^{2+} (aq) + Cu (s)
   b. Mg (s) + 2 H^+ (aq) → Mg^{2+} (aq) + H_2 (g)
6. For each of the following, write the half-reactions. Indicate which is oxidation and which is reduction. Ignore spectator ions.
   a. Ni (s) + $Cu(NO_3)_2$ (aq) → Cu (s) + $Ni(NO_3)_2$ (aq)
   b. Pb (s) + $Cu(NO_3)_2$ (aq) → Cu (s) + $Pb(NO_3)_2$ (aq)
   c. Ca (s) + 2 HNO_3 (aq) → H_2 (g) + $Ca(NO_3)_2$ (aq)
   d. 2 Al (s) + Fe_2O_3 (s) → 2 Fe (l) + Al_2O_3 (s)

7. Has a redox reaction occurred here? Explain.
   $FeCl_3$ (aq) + 3 NaOH (aq) → $Fe(OH)_3$ (s) + 3 NaCl (aq)
8. For each reaction, write the half-reaction equation and identify if it is oxidation or reduction.
   a. dinitrogen oxide to nitrogen gas in an acidic solution
   b. nitrite ions to nitrate ions in a basic solution
   c. silver oxide to silver metal in a basic solution
   d. nitrate ions to nitrous acid in an acidic solution
   e. hydrogen gas to water in a basic solution
9. For each application, write the half-reaction equation and classify it as oxidation or reduction.
   a. bacterial action in soil: ammonia to nitrate ions in an acidic environment
   b. pulp and paper bleaching: hydrogen peroxide to water in an acidic solution
   c. alkaline battery: manganese(IV) oxide to manganese(III) oxide in a basic environment

<u>Balancing with Half-Reactions</u>

10. Balance the following redox equations.
    a. Ag (s) + $Cr_2O_7^{2-}$ (aq) + H^+ (aq) → Ag^+ (aq) + Cr^{3+} (aq) + H_2O (l)
    b. MnO_4^- (aq) + Br^- (aq) + H^+ (aq) → Mn^{2+} (aq) + Br_2 (l) + H_2O (l)
11. Balance the following redox equations that occur in <u>acidic solution</u>.
    a. Zn (s) + NO_3^- (aq) → NH_4^+ (aq) + Zn^{2+} (aq)
    b. Cl_2 (aq) + SO_2 (g) → Cl^- (aq) + SO_4^{2-} (aq)
    c. Mn^{2+} (aq) + $HBiO_3$ (aq) → Bi^{3+} (aq) + MnO_4^- (aq)
12. Balance the following redox equations that occur in <u>basic solution</u>.
    a. MnO_4^- (aq) + I^- (aq) → MnO_2 (s) + I_2 (s)
    b. CN^- (aq) + IO_3^- (aq) → CNO^- (aq) + I^- (aq)
    c. CrO_4^{2-} (aq) + Cl^- (aq) → Cr^{3+} (aq) + Cl_2 (g)

Galvanic Cells

<u>Galvanic vs. Electrolyrtic Cells</u>

13. Current still flows from anode to cathode in an electrolytic cell, and oxidation occurs at the anode. How are they labelled differently for an electrolytic and galvanic cell?
14. Identify four ways that galvanic and electrolytic cells differ.

<u>Cell Notation</u>

15. Write each of the following in cell notation.
    a. Co (s) + Cu^{2+} (aq) → Co^{2+} (aq) + Cu (s)
    b. 3 Pb (s) + 8 H^+ (aq) + 2 NO_3^- (aq) → 2 NO (g) + 3 Pb^{2+} (aq) + 4 H_2O (l)
    c. Zn (s) + 2 H^+ (aq) + 2 MnO_2 (s) → Zn^{2+} (aq) + Mn_2O_3 (s) + H_2O (l)
    d. Co (s) + $S_2O_8^{2-}$ (aq) → 2 SO_4^{2-} (aq) + Co^{2+}(aq)

<u>Cell Potential</u>

16. Given these half-reactions:

Fe^{3+} (aq) + e^- → Fe^{2+} (aq)	$E°_{red}$ = 0.77 V
$S_2O_6^{2-}$ (aq) + 4 H^+ (aq) + 2 e^- → 2 H_2SO_3 (aq)	$E°_{red}$ = 0.60 V
N_2O (g) + 2 H^+ (aq) + 2 e^- → N_2 (g) + H_2O (l)	$E°_{red}$ = -1.77 V
VO_2^+ (aq) + 2 H^+ (aq) + e^- → VO^{2+} (aq) + H_2O (l)	$E°_{red}$ = 1.00 V

    a. Write balanced chemical equations for the oxidation of Fe^{2+} (aq) by $S_2O_6^{2-}$ (aq), by N_2O (aq) and by VO_2^+ (aq).
    b. Calculate $E°_{cell}$ for each reaction in (a). Which are spontaneous?

17. For each of the following cells, write the equations for the reactions occurring at the cathode and anode, and an equation for the overall or net cell reaction. Calculate the standard cell potential.
    a. Cr (s) | Cr^{2+} (aq) || Sn^{2+} (aq) | Sn (s)
    b. Co (s) | Co^{2+} (aq) || Ag^{+} (aq) | Ag (s)
18. A galvanic cell is created using gold (making Au^{3+}) and magnesium half-cells. Determine which half-cell will undergo oxidation and which will undergo reduction, identify anode and cathode, and calculate the voltage for the cell.

<u>Cell Diagrams</u>

19. An electrochemical cell undergoes the following unbalanced reaction:

$$Mg\ (s) + HCl\ (aq) \rightarrow H_2\ (g) + MgCl_2\ (aq)$$

    a. Use oxidation numbers to determine which substance is being oxidized and which is being reduced.
    b. Balance this equation using half-reactions.
    c. Determine the cathode and anode, then write the cell notation.
    d. Draw a diagram of each cell, labelling the electrodes, electrolytes, direction of electron flow and direction of ion movement.
    e. Calculate the cell potential. Is this cell spontaneous?
20. For each of the following cells, use the given cell notation to write chemical equations to represent the cathode, anode and net cell reactions. Draw a diagram of each cell, labelling the electrodes, electrolytes, direction of electron flow and direction of ion movement.
    a. Zn (s) | Zn^{2+} (aq) || Ag^{+} (aq)| Ag (s)
    b. Al (s) | Al^{3+} (aq) || Au^{3+} (aq) | Au (s)
21. For each set of half-cells in standard conditions:
    - write the two half-reactions
    - label each half-reaction as oxidation or reduction
    - calculate the voltage of the galvanic cell
    - write the net balanced redox equation
    - diagram the cell, indicating the electrodes in appropriate electrolytic solutions, label the cathode and anode, the direction of the flow of electrons, an appropriate salt bridge and the direction of the flow of ions from the salt bridge

    a. iron-iron(II) ion (Fe|Fe^{2+}) and lead-lead(II) ion (Pb|Pb^{2+})
    b. chromium-chromium(III) ion (Cr|Cr^{3+}) and rubidium-rubidium ion (Rb|Rb^{+})
    c. copper-copper(I) ion (Cu|Cu^{+}) and aluminum-aluminum ion (Al|Al^{3+})
    (NOTE: Be sure to use the Cu^{1+} half-reaction, not Cu^{2+})