Chem 160: Chapter 18 Electrochemistry

1. Define Oxidation: ____
2. Define Reduction: ____
3. Define Anode: ____
4. Define Cathode: ____

5. True or False; Correct the False Statements:
A. Oxidation always occurs at the cathode.
B. Electrons always flow from the anode to the cathode.
C. The salt bridge is to maintain charge balance.
D. A galvanic cell may have a negative or positive cell potential.
E. Cations always flow toward the anode.
F. The shorthand notation of a galvanic cell specifies the anode on the left, the cathode on the right and the reactants in the half-cell compartments.

6. Think about the steps used to balance a redox reaction by the half reaction method?
   a. Balance the reaction below in acidic conditions.

$$H_2O_2 \quad + \quad MnO_4^- \longrightarrow O_2 \quad + \quad Mn^{2+}$$

   b. Balance the same reaction in basic conditions.

7. Write the overall balanced reaction and calculate the cell potential of a galvanic cell that uses the two half reactions shown below:

	E^o (V)
$Cu^{2+} + 2e^- \longrightarrow Cu$	+0.34 V
$Na^+ + e^- \longrightarrow Na$	-2.71

Solutions

1. Define Oxidation: Loss of electrons during a reaction
2. Define Reduction: Gain of electrons during a reaction
3. Define Anode: positively charged electrode by which the electrons leave a device
4. Define Cathode: negatively charged electrode by which electrons enter an electrical device

5. True or False; Correct the False Statements:
   a. Oxidation always occurs at the cathode. False. Oxidation occurs at the anode
   b. Electrons always flow from the anode to the cathode. True
   c. The salt bridge is to maintain charge balance. True
   d. A galvanic cell may have a negative or positive cell potential. False. A galvanic cell may have positive cell potential or zero potential, but never negative.
   e. Cations always flow toward the anode. False. Cations flow toward the cathode.
   f. The shorthand notation of a galvanic cell specifies the anode on the left, the cathode on the right and the reactants in the half-cell compartments. True

6. Think about the steps used to balance a redox reaction by the half reaction method?
   a. Balance the reaction below in acidic conditions.

$$H_2O_2 + MnO_4^- \longrightarrow O_2 + Mn^{2+}$$

Step 1. Write unbalanced net ionic equation

$H_2O_2 + MnO_4^- \longrightarrow O_2 + Mn^{2+}$ (acidic solution)

Step 2. Decide which atoms are oxidized and which are reduced, and write two unbalanced half reactions for the separate processes

$$MnO_4^- \longrightarrow Mn^{2+}$$

$$H_2O_2 \longrightarrow O_2$$

Step 3. Balance both half reactions for all atoms other than H and O

N/A

Step 4. Balance the oxygens in the half reaction by adding the correct number of H_2O molecules on the appropriate side of the half reaction

$$MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$$

Step 5. Balance the hydrogens in the half reaction by adding the correct number of protons (H^+) on the appropriate side of the half reaction

$8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$

$H_2O_2 \longrightarrow O_2 + 2H^+$

**If solution is acidic proceed to Step 6. If the solution is basic, then add the correct number of OH^- groups on each side of the half reaction to neutralize the H^+ previously added

Step 6. Balance the charges by adding the correct number of electrons on the appropriate side of the half reaction

$5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$

$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$

** you want to cancel the electrons on both sides, you can do so by multiplying by the least common factor

$2 \times (5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O)$

$5 \times (H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-)$

$10e^- + 16H^+ + 2MnO_4^- \longrightarrow 2Mn^{2+} + 8H_2O$

$5H_2O_2 \longrightarrow 5O_2 + 10H^+ + 10e^-$

**now cancel electrons/protons and combine the two half reactions

$6H^+ + 5H_2O_2 + 2MnO_4^- \longrightarrow 2Mn^{2+} + 5O_2 + 8H_2O$

b. Balance the same reaction in basic conditions.

$5H_2O_2 + 2MnO_4^- \longrightarrow 2Mn^{2+} + 5O_2 + 6OH^- + 2H_2O$

7. Write the overall balanced reaction and calculate the cell potential of a galvanic cell that uses the two half reactions shown below

	E^o (V)
$Cu^{2+} + 2e^- \longrightarrow Cu$	+0.34 V
$Na^+ + e^- \longrightarrow Na$	-2.71

Based on the two half reactions and given cell potentials, we need to adjust reactions so that the electrons are not on the same side of both reactions. For one half reaction the electrons need to be on the right side and for the other on the left side. Also make sure that the overall cell potential is positive. In order to do so, reverse the 2nd reaction and leave the 1st reaction the way that it is. Multiply by two so the number of electrons are the same on both sides do not multiply the cell potential this value stays the same.

					E^o (V)
Cu^{2+} +	2e-	$\longrightarrow$	Cu		+0.34 V
2Na	$\longrightarrow$	$2Na^+$ +	$2e^-$		+2.71

The electrons will cancel, and you can combine the reaction and calculate the total cell potential by adding the values. E^0cell = 0.34 + 2.71 + 3.05 $Cu^{2+} + 2Na \longrightarrow Cu + 2Na^+$