Name: ANSWER KEY

Chemistry 30

Unit 6: Redox Reactions and Electrochemistry

Assignment 3 Electrochemistry 26

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4 1. Use half-reaction potentials to predict whether the following reactions are spontaneous or nonspontaneous in aqueous solutions. If the reaction is spontaneous, write a balanced equation and calculate the total voltage.

a. $I_2 + Cl^- \rightarrow I^- + Cl_2$

$I_2 + 2\ e^- \rightarrow 2\ I^-$	***$E° = +0.54$***
$2\ Cl^- \rightarrow 2\ e^- + Cl_2$	***$\underline{E° = -1.36}$***
	$E° = -0.82$ not spontaneous

b. $MnO_4^- + Br^- + H^+ \rightarrow Mn^{2+} + Br_2 + H_2O$

$2 \times (MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O)$	$E° = +1.51$
$5 \times (2Br^- \rightarrow Br_2 + 2\ e^-)$	$\underline{E° = -1.07}$
	$E° = +0.44$ spontaneous

$2\ MnO_4^- + 10\ Br^- + 16H^+ \rightarrow 2Mn^2 + 5Br_2 + 8\ H_2O$ **balanced equation**

3 2. Compare an electrochemical cell with an electrolytic cell by completing the following table:

	Electrochemical Cell	Electrolytic Cell
Energy Conversion	**chemical → electrical**	**electrical → chemical**
Spontaneous Chemical Reaction?	**yes**	**no**
Value of E° (positive or negative)	**positive**	**negative**

2 3. Potassium reacts with chlorine to produce the ionic compound potassium chloride:

$2\ K\ (s) + Cl_2\ (g) \rightarrow 2\ KCl\ (s)$

a. Write a balanced half-reaction for the oxidation reaction.

$2\ K \rightarrow 2\ K^+ = 2\ e^-$

b. Write a balanced half-reaction for the reduction reaction.

$Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$

2 4. What reaction (oxidation or reduction) occurs at an anode of . . .

a. an electrochemical cell ***oxidation***

b. an electrolytic cell ***oxidation***

5 5. An iron bar is to be electroplated with zinc.

- Identify what will act as the two electrodes for the cell
- Identify each electrode as either the anode or cathode
- Write the half-reactions occurring at each electrode
- Identify a solution that would make a suitable electrolyte for this cell
- Identify which electrode will be attached to the negative post of the battery and which will be attached to the positive post, and explain.

The objective of electroplating is to place a thin layer of a desired metal (zinc) on another object.

Thus on the iron bar we want to form solid zinc from Zn^{2+}:

$Zn^{2+} + 2\ e^- \rightarrow Zn\ (s)$	This is reduction, making the iron bar the ***cathode*** (Red Cat). This reaction requires electrons, which must be supplied by a battery (or other source of current). The iron bar must be connected to the ***negative post*** of the battery because it is the negative battery post that provides electrons.

The cathode reaction will require Zn^{2+} ions, which come from the **electrolytic solution**. Thus this solution must contain Zn^{2+} - $Zn(NO_3)_2$ would be a suitable choice.

As the Zn^{2+} from the electrolytic solution get used up (at the cathode), they must be replaced. This is the function of the other electrode, which must be a bar of the plating metal – zinc.

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	The is the **oxidation reaction** occurring at the **anode** (An Ox). Here the solid zinc bar will disintegrate, producing Zn^{2+} ions which will be used to coat the iron bar. Electrons released here will return to the **positive post** of the battery.

NOTE: The iron undergoes no chemical reaction.

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10 6. The net equation for a given galvanic cell is:

$$Sn\ (s) + 2\ Ag^{+} \rightarrow Sn^{2+} + 2\ Ag\ (s)$$

a. Write the two half-reactions involved, and identify each in terms of (1) site of oxidation or reduction and (2) anode or cathode.

$Sn \rightarrow Sn^{2+} + 2\ e^{-}$	***oxidation***	***anode***	***$E° = + 0.14$***
$2\ Ag^{+} + 2\ e^{-} \rightarrow 2\ Ag$	***reduction***	***cathode***	***$E° = + 0.80$***

b. Calculate the net potential of the cell (the voltage), assuming standard conditions.

$E° = 0.94$ volts

c. Draw a fully labeled diagram of the electrochemical cell. Be sure to indicate the flow of electrons in the external circuit (through the wire and light bulb) and the flow of ions in the solution.