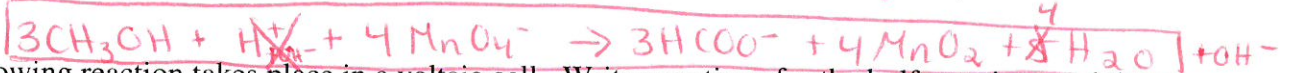
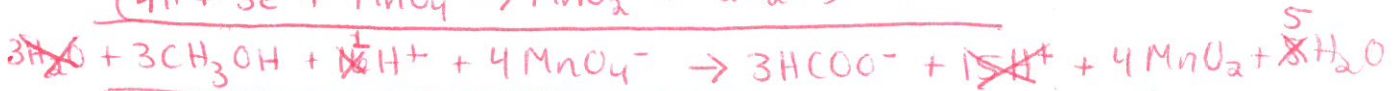
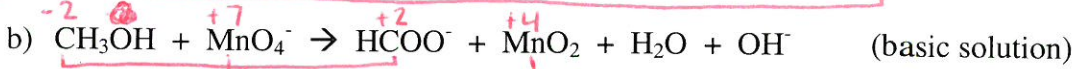
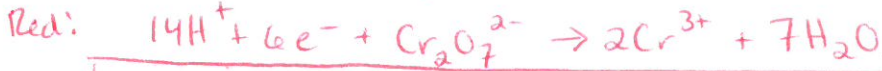


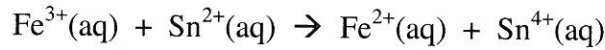


AP Chemistry Review  
Chapter 17 - Electrochemistry  
18

1. Use the half-reaction method to balance the following redox reactions:



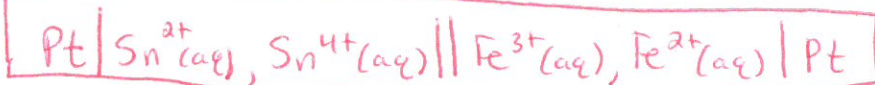
2. The following reaction takes place in a voltaic cell. Write equations for the half-reactions and the overall cell reaction. Write a cell diagram for the voltaic cell and calculate the value of  $E^\circ_{\text{cell}}$ .



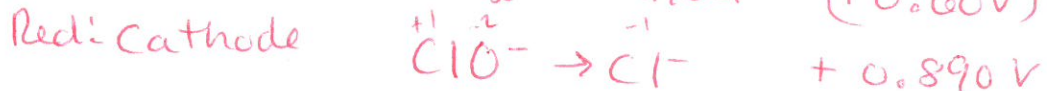
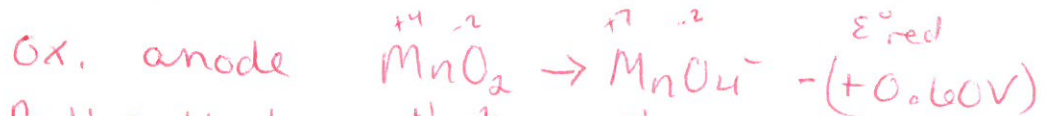
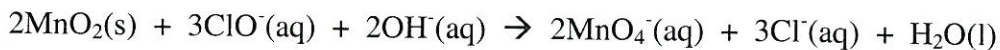
$E^\circ_{\text{red}}$   
- (0.154 V)  
+ (0.771 V)

$$E^\circ_{\text{cell}} = -0.154 + (0.771)$$

$$E^\circ_{\text{cell}} = 0.617 \text{ V}$$



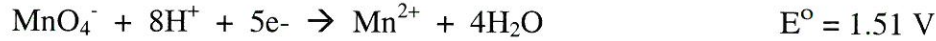
3. Predict whether the following reaction will be spontaneous in the forward direction. Assume that all reactants and products are in their standard states.



$$E^\circ_{\text{cell}} = 0.890 + (-0.60) = +0.290 \text{ V}$$

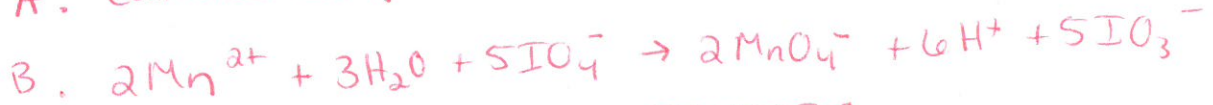
Spontaneous

4. Describe the galvanic cell below, based on the half-reaction. (Assume that all concentrations are 1.0 M and that all partial pressures are 1.0 atm.)



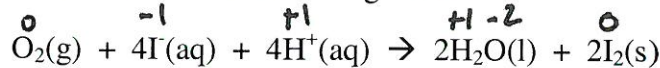
- A. Identify the cathode and anode.  
 B. Give the overall balanced reaction.  
 C. Determine  $E^\circ$  for the galvanic cells.

A. Cathode:  $\text{IO}_4^-$  half-rxn      Anode:  $\text{MnO}_4^-$  half-rxn



C.  $E^\circ_{\text{cell}} = 1.60 \text{ V} - 1.51 \text{ V} = \boxed{0.09 \text{ V}}$

5. Determine the values of  $E^\circ_{\text{cell}}$  and  $\Delta G^\circ$  for the following reaction:

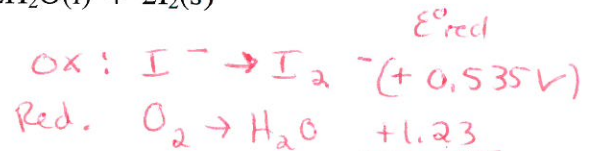


$$\Delta G^\circ = -nF E^\circ_{\text{cell}}$$

$$= -4 \left( 96,485 \frac{\text{C}}{\text{s}} \right) (0.694 \frac{\text{J}}{\text{C}})$$

$$\Delta G^\circ = -267,842 \text{ J}$$

$$\boxed{\Delta G^\circ = -268 \text{ kJ}}$$



$$E^\circ_{\text{cell}} = 1.23 - 0.535 \text{ V}$$

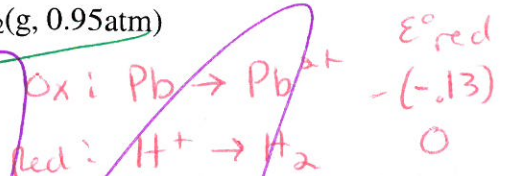
$$\boxed{E^\circ_{\text{cell}} = 0.694 \text{ V}}$$

6. What is the value of  $E_{\text{cell}}$  for the following reaction carried out in a voltaic cell?



$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log Q$$

$$E_{\text{cell}} = 0.13 \text{ V} - \frac{0.0591}{2} \log \frac{P_{\text{H}_2} [\text{Pb}^{2+}]}{[\text{H}^+]^2}$$



$$E^\circ_{\text{cell}} = 0 + 0.13 = \boxed{0.13 \text{ V}}$$

$$E_{\text{cell}} = 0.13 - \frac{0.0591}{2} \log \left[ \frac{(0.95)(0.85)}{(0.0025)^2} \right] = \boxed{-0.0263 \text{ V}}$$

7. Electrolysis of a molten metal chloride ( $\text{MCl}_3$ ) was done in a lab, using a current of 6.50 A for 23.28 minutes. At the end of the experiment, 1.41 g of the metal was deposited at the cathode. What is the metal?

$$\text{Ampere} = \frac{\text{C}}{\text{sec.}}$$

$$\frac{23.28 \text{ min} \cdot 60 \text{ sec}}{1 \text{ min.}} = 1397 \text{ sec.}$$

$$6.50 = \frac{\text{C}}{1397}$$

$$\text{MM} = \frac{\text{g}}{\text{mol}}$$

$$\text{C} = 9.1 \times 10^3 \text{ C} \left| \frac{1 \text{ mol e}^-}{96,485 \text{ C}} \right| \left| \frac{1 \text{ mol M}}{3 \text{ mol e}^-} \right| = 0.0314 \text{ mol}$$

$$\text{MM} = \frac{1.41 \text{ g}}{0.0314 \text{ mol}}$$

$$\boxed{\text{MM} = 44.9 \frac{\text{g}}{\text{mol}} \rightarrow \text{Scandium}}$$