Electrochemistry Worksheet

1. Assign oxidation numbers to each atom in the following:

   a. P_4O_6 b. BiO_3^- c. N_2H_4

   d. $Mg(BrO_4)_2$ e. $MnSO_4$ f. $Mn(SO_4)_2$

2. For each of the reactions below identify the oxidizing agent and the reducing agent.

   a. $2\ KCl + MnO_2 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + Cl_2 + H_2O$

   b. $SiCl_4 + 2\ Mg(s) \rightarrow 2\ MgCl_2 + Si$

3. Use the half-reaction method to balance each of the following oxidation-reduction reactions. Identify the oxidizing agent and the reducing agent.

   a. $Cu(s) + Ag^{+}(aq) \rightarrow Ag(s) + Cu^{2+}$

   b. $Al(s) + I_2(s) \rightarrow AlI_3(s)$

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

4. Balance each of the following oxidation-reduction reactions. Identify the oxidizing agent and the reducing agent.

   a. $S_2O_3^{2-} + OCl^{-} \rightarrow Cl^{-} + S_4O_6^{2-}$ in acid

b. $CH_3OH + MnO_4 \rightarrow HCOO^- + MnO_2$ in base

c. $NO_3^- + Zn \rightarrow NH_4^+ + Zn^{2+}$ in acid

d. $Br_2 \rightarrow Br^- + BrO_3^-$ in base

5. Complete and balance each of the following single displacement reactions. Use the emf table in the textbook to determine if the reaction occurs. If no reaction occurs, write NR instead of products.

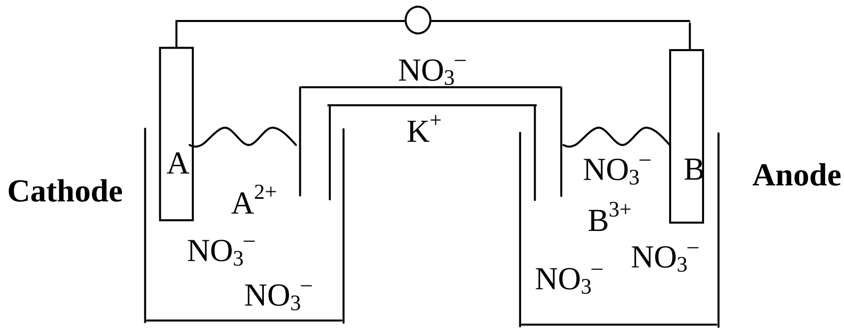
   a. $Ag(s) + Fe(NO_3)_2(aq) \rightarrow$

   b. $CuCl_2(aq) + Na(s) \rightarrow$

6. Write the cell notation for the voltaic cell that incorporates the following redox reaction.

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

7. Answer the questions below regarding the voltaic cell drawn.



a. Write both half-reactions:

Cathode Half Reaction: ______________________________

Anode Half Reaction: ______________________________

b. In which direction will the electrons flow?

c. Which electrode will be positively charged?

d. In which direction will the NO_3^- ions flow in the salt bridge?

e. Which electrode decreases in mass during the reaction?

f. Write the cell notation for this voltaic cell.

8. Draw a voltaic cell that is constructed with a Mn/Mn^{2+} electrode and a Cd/Cd^{2+} electrode. Use the emf table in the textbook to determine which electrode will be the cathode and which will be the anode. Your drawing should include all of the following components:
a. Label the location of each substance (Mn, Mn^{2+}, Cd, and Cd^{2+})
b. Label the cathode and the anode
c. Label the direction of electron flow
d. Label which electrode is positively charged and which is negatively charged.
e. Include a salt bridge with $NaNO_3$. Label the direction that each ion flows
f. Write the cell notation for this voltaic cell.
g. Calculate the cell potential of this cell.

9. Given the following half-reactions and half-cell potentials, write the balanced overall electrochemical reaction that would occur and calculate the cell potential of a voltaic cell incorporating these two half reactions.

   $O_2(g) + 2\ H_2O(l) + 4\ e^- \rightarrow 4\ OH^-(aq)$ $E° = +0.40\ V$
   $Cr^{3+}(aq) + 3\ e^- \rightarrow Cr(s)$ $E° = -0.74\ V$

10. Balance the following skeleton reaction, calculate $E°_{cell}$, and determine whether the reaction would be spontaneous as written. You will need to use the emf table in the textbook.

    $Cr^{3+}(aq) + Cu(s) \rightarrow Cr(s) + Cu^{2+}(aq)$

11. Answer the question below regarding the reaction, $Cl_2(g) + Fe^{2+}(aq) \rightarrow Cl^{-}(aq) + Fe^{3+}(aq)$.

   a. Balance the reaction using the half-reaction method.

   b. Use the emf series in the textbook to calculate the $E°_{cell}$ and determine whether this reaction would occur in a voltaic cell or an electrolytic cell.

   c. Use $E°_{cell}$ to calculate K_c for this reaction at 25 °C.

   d. Use $E°_{cell}$ to calculate $\Delta G°$ for this reaction (the ° symbol denotes standard conditions; what temperature is standard conditions?)

12. Answer the question below regarding the reaction, $Pb^{2+}(aq) + Ag(s) \rightarrow Pb(s) + Ag^{+}(aq)$.

   a. Balance the reaction using the half-reaction method.

b. Use the emf table in the textbook to calculate the $E°_{cell}$ and determine whether this reaction would occur in a voltaic cell or an electrolytic cell.

c. Use $E°_{cell}$ to calculate K_c for this reaction at 25 °C.

d. Use $E°_{cell}$ to calculate ΔG° for this reaction (the ° symbol denotes standard conditions; what temperature is standard conditions?)

13. Calculate the cell potential for a voltaic cell with Pt/Pt^{2+} and Ag/Ag^{+} half-cells and the initial concentrations $[Pt^{2+}] = 0.90$ M and $[Ag^{+}] = 0.20$ M.

$Pt^{2+}(aq) + 2\ e^{-} \rightarrow Pt(s)$ $\quad$ E° = +1.20 V
$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$ $\quad$ E° = +0.80 V

14. An aluminum electrode weighing 54.98 g is used in an electrolysis reaction using a current of 1.2 A. After the reaction is stopped, the aluminum electrode weighs 54.09 g.

    a. Did the aluminum electrode described above act as the cathode or the anode?

    b. How many hours was the current applied to the electrolysis cell?

15. How many grams of silver can form by passing 19.8 A through an electrolytic cell containing silver ions for 13.2 min?

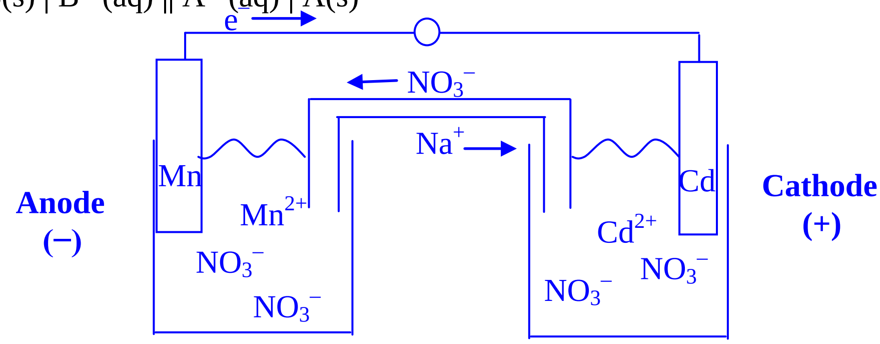
1. a. O −2, P +3 b. O −2, Bi +5 c. H +1, N -2
   d. Mg +2, O −2, Br +7 e. Mn +2, O −2, S +6 f. Mn +4, O −2, S +6
   Hint: On these ionic compounds, separate them into the ions first. For example, for $MnSO_4$, I separated into $Mn^{?+} + SO_4^{2-}$. That way, I can figure out the charge on Mn, and it's now easier to find the oxidation number of S.
2. a. Reducing Agent: KCl, Oxidizing Agent: MnO_2
   b. Reducing Agent: Mg, Oxidizing Agent: $SiCl_4$
3. a. $Cu(s) + 2\ Ag^+(aq) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$
   oxidizing agent: Ag^+ reducing agent: Cu(s)
   b. $2\ Al(s) + 3\ I_2(s) \rightarrow 2\ AlI_3(s)$
   oxidizing agent: $I_2(s)$ reducing agent: Al(s)
   c. $3\ Pb(s) + 2\ Fe^{3+}(aq) \rightarrow 3\ Pb^{2+}(aq) + 2\ Fe(s)$
   oxidizing agent: Fe^{3+}(as) reducing agent: Pb(s)
4. a. $2\ S_2O_3^{2-} + 2\ H^+ + OCl^- \rightarrow S_4O_6^{2-} + Cl^- + H_2O$
   $S_2O_3^{2-}$ is the reducing agent
   OCl^- is the oxidizing agent.
   Note: for full credit, you will need to show your work on the balancing redox equations. This means show me the two balanced half reactions as I have done in the next problem
   b.

Balance half reactions:	$H_2O + CH_3OH \rightarrow HCOO^- + 5\ H^+ + 4\ e^-$
	$4\ e^- + 4\ H^+ + MnO_4 \rightarrow MnO_2 + 2\ H_2O$
add together & cancel:	$CH_3OH + MnO_4 \rightarrow HCOO^- + MnO_2 + H_2O + H^+$
add OH^- to cancel H^+	$+ OH^-$ $\qquad + OH^-$
$OH^- + H^+ \rightarrow H_2O$:	$OH^- + CH_3OH + MnO_4 \rightarrow HCOO^- + MnO_2 + 2\ H_2O$

   CH_3OH is reducing agent (hint: look for the one that lost e^- in the half reactions. That's the one that got oxidized, so it's the reducing agent. The other one must be the opposite. No need to determine oxidation numbers!)
   MnO_4 is oxidizing agent
   c. $10\ H^+ + NO_3^- + 4\ Zn \rightarrow NH_4^+ + 3\ H_2O + 4\ Zn^{2+}$
   Zn is the reducing agent; NO_3^- is the oxidizing agent
   d. $12\ OH^- + 6\ Br_2 \rightarrow 10\ Br^- + 2\ BrO_3^- + 6\ H_2O$
   OR $6\ OH^- + 3\ Br_2 \rightarrow 5\ Br^- + BrO_3^- + 3\ H_2O$
   Br_2 is the reducing agent; Br_2 is the oxidizing agent
5. a. $2\ Ag(s) + Fe(NO_3)_2(aq) \rightarrow$ ~~$2\ AgNO_3(aq) + Fe(s)$~~ NR
   No reaction occurs because $E°_{cell}$ is negative
   b. $CuCl_2(aq) + 2\ Na(s) \rightarrow Cu(s) + 2\ NaCl(aq)$
   Theoretically, this reaction would occur. As you probably know, though, Na is so reactive that it would react with the water that the $CuCl_2$ was dissolved in.
6. $Mg(s)\ |\ Mg^{2+}(aq)\ ||\ Sn^{2+}(aq)\ |\ Sn(s)$
7. a. Cathode Half Reaction: $A^{2+}(aq) + 2e^- \rightarrow A(s)$
   Anode Half Reaction: $B(s) \rightarrow B^{3+}(aq) + 3e^-$
   b. Towards the left (anode to cathode)
   c. A (cathode)

d. Towards the right (opposite from e^-)
e. B (getting oxidized into B^{3+}(aq))
f. B(s) | B^{3+}(aq) || A^{2+}(aq) | A(s)

8. a-e:



f. Mn(s) | Mn^{2+}(aq) || Cd^{2+}(aq) | Cd(s)
g. 0.78 V

9. $3\ O_2(g) + 6\ H_2O(l) + 4\ Cr(s) \rightarrow 12\ OH^-(aq) + 4\ Cr^{3+}(aq)$ $E^\circ_{cell} = +1.14$ V
10. $2\ Cr^{3+}(aq) + 3\ Cu(s) \rightarrow 2\ Cr(s) + 3\ Cu^{2+}(aq$ $E^\circ_{cell} = -1.08$ V
The reaction is not spontaneous as written because E°_{cell} is negative
11. a. $Cl_2(g) + 2\ Fe^{2+}(aq) \rightarrow 2\ Cl^-(aq) + 2\ Fe^{3+}(aq)$
b. $E^\circ_{cell} = +0.59$ V
Voltaic cell because E°_{cell} is positive, so reaction is spontaneous
c. $K_c = 8.85 \times 10^{19}$ (notice the very high K value for a spontaneous reaction)
d. $\Delta G^\circ = -1.1 \times 10^2$ kJ (notice the negative ΔG value for a spontaneous reaction)
12. a. $Pb^{2+}(aq) + 2\ Ag(s) \rightarrow Pb(s) + 2\ Ag^+(aq)$
b. $E^\circ_{cell} = -0.93$ V
Electrolytic cell because E°_{cell} is negative, so reaction is not spontaneous
c. $K_c = 3.6 \times 10^{-32}$ (notice the very low K value for a nonspontaneous reaction)
d. $\Delta G^\circ = +1.8 \times 10^2$ kJ (notice the positive ΔG value for a nonspontaneous reaction)
13. $E_{cell} = +0.44$ V
14. a. Aluminum was the anode (decreased in mass, so it must have oxidized to Al^{3+})
b. 2.2 hrs
15. 17.5 g Ag