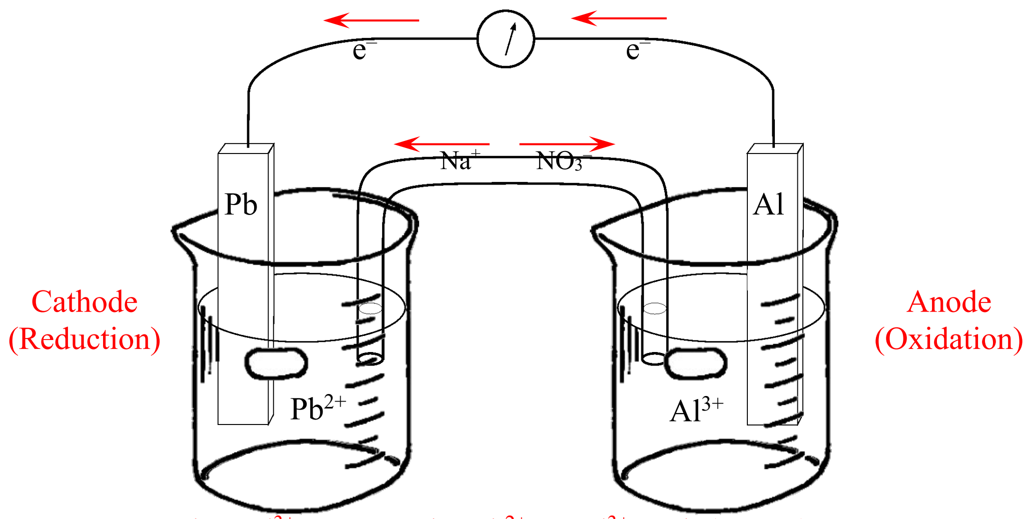
WKS 16-1
Voltaic Cells: Setup +Voltage

Name_Answer Key___________________
Period__________ Date __________

1) A Voltaic cell is an electrochemical cell which produces electricity. Given the following voltaic cell, answer the following questions and label diagram as directed.



$Pb + Al^{3+} \rightarrow$ NR; $Al + Pb^{2+} \rightarrow Al^{3+} + Pb$ (occurs)

a) In a single beaker, if Pb(s) is placed into a solution of Al^{3+}, no reaction occurs, but Al(s) in Pb^{2+} (aq) will react and Pb(s) will precipitate. Which metal is most easily oxidized? Al

b) Thus, what will be oxidized? Al c) What will be reduced? Pb^{2+}

d) Write the two half reactions and write the overall redox reaction. Make sure you balance out the electrons so that your overall redox reaction is **balanced.** (no voltages needed)

Oxidation: $[\ Al \rightarrow Al^{3+} + 3\ e^- \ ] \times 2$

Reduction: $[\ Pb^{2+} + 2\ e^- \rightarrow Pb\ ] \times 3$

Overall reaction: $2\ Al + 3\ Pb^{2+} \rightarrow 2\ Al^{3+} + 3\ Pb$

e) During which half reaction are electrons being lost? Ox Which electrode is losing e^-? Anode

f) During which half reaction are electrons being gained? Red Which electrode is gaining e^-? Cath

g) Label your diagram with the following: anode and cathode, draw arrows showing the flow of electrons in the wire & draw arrows showing the flow of ions in the salt bridge.

h) Which electrode will gain mass during the reaction? (Pb or Al) Pb Why?
Pb is being formed as Pb^{2+} is reduced, which "plates out" onto the Pb electrode.

i) Which electrode will lose mass during the reaction? (Pb or Al) Al Why?
Al is being oxidized into Al^{3+}, which dissolve into solution off of the electrode.

j) Explain why the ions in the salt bridge moved the way you labeled them in your diagram.
The Na^+ moves to the cathode since the reduction would build up excess – charge, attracting the + cations; the NO_3^- moves to the anode since the oxidation would build up excess + charge, attracting the – anions.

k) What would happen if the salt bridge were removed? Explain why this happens.
The voltage would go to 0 since the circuit has been interrupted; the build-up of charge prevents further reactions from occurring.

2) For the voltaic cell above (Pb^{2+}/Pb & Al^{3+}/Al), rewrite the two half reactions, determine their E° values, write the overall redox reaction and calculate the $E°_{net}$. Remember that cell voltages do NOT get multiplied.

Oxidation:	[$Al \rightarrow Al^{3+} + 3\,e^-$] × 2	$E°_{ox}$ = – (–1.66 V)
Reduction:	[$Pb^{2+} + 2\,e^- \rightarrow Pb$] × 3	$E°_{red}$ = –0.13 V
Overall reaction:	$2\,Al + 3\,Pb^{2+} \rightarrow 2\,Al^{3+} + 3\,Pb$	$E°_{net}$ = +1.53 V

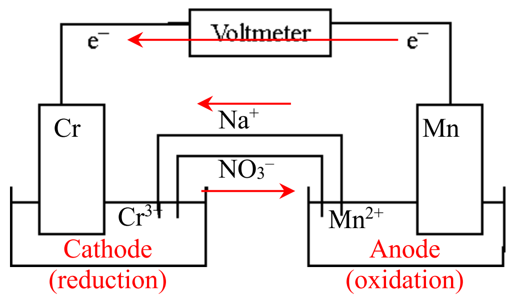
3) Look at the voltaic cell set up below and answer the following questions:

a) Use your chart of reduction potentials to determine which metal (Cr or Mn) will be oxidized. *(Which metal is most easily oxidized?* Mn *) Mn is* *lower* *on the Reduction potential table*

b) In the diagram, label the anode and cathode, show flow of electrons in wire, and show flow of ions in the salt bridge

c) Write the oxidation and reduction half reactions below. Make sure to balance the electrons. Write balance overall reaction.

d) Calculate $E°_{net}$ for the overall reaction.



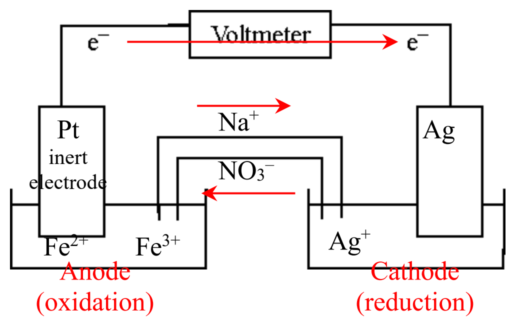
Ox:	[$Mn \rightarrow Mn^{2+} + 2\,e^-$] × 3	E° = –(–1.18 V)
Red:	[$Cr^{3+} + 3\,e^- \rightarrow Cr$] × 2	E° = –0.74 V
Overall Rxn:	$3\,Mn + 2\,Cr^{3+} \rightarrow 3\,Mn^{2+} + 2\,Cr$	$E°_{net}$ = +0.44 V

4) A voltaic cell (with an inert platinum electrode in the Fe^{3+}/Fe^{2+} cell) is constructed using the overall reaction and setup as shown below. Answer the following questions concerning this voltaic cell:

a) In the diagram, label the anode and cathode, show flow of electrons in wire, and show flow of ions in the salt bridge

b) Write the oxidation and reduction half reactions below. Make sure to balance the electrons and write in needed coefficients into the overall reaction.

c) Calculate $E°_{net}$ for the overall reaction.



$$Fe^{2+}\,(aq) + Ag^{+}\,(aq) \longrightarrow Fe^{3+}\,(aq) + Ag\,(s)$$

Ox:	$Fe^{2+} \rightarrow Fe^{3+} + e^-$	E° = –(+0.77 V)
Red:	$Ag^+ + e^- \rightarrow Ag$	E° = +0.80 V
		$E°_{net}$ = + 0.03 V

5) Will the following reaction be spontaneous in the forward direction? No

Hint: Which is more easily oxidized? I^- or H_2? (Use reference chart to determine $E°_{net}$.)

H_2 is more easily oxidized (more negative E°)

$$2\,I^-\,(aq) + 2\,H^+\,(aq) \longrightarrow I_2\,(s) + H_2\,(g)$$

Ox: $2\,I^-\,(aq) \rightarrow I_2\,(s) + 2\,e^-$	E° = –(+0.54 V)
Red: $2\,H^+ + 2\,e^- \rightarrow H_2\,(g)$	E° = 0.00 V
	$E°_{net}$ = –0.54 V
	No Spontaneous

6) This reaction is spontaneous in the forward direction.

ox ox

$$Tl^{3+}\,(aq) + Cr^{2+}\,(aq) \longrightarrow Tl^{+}\,(aq) + 2\,Cr^{3+}\,(aq)$$

Thus, which must be more easily oxidized? Cr^{2+} or Tl^+?
(Tl^{3+}/Tl^+ is not on reference chart!!)

7) The following reaction is not spontaneous in the forward direction, but is spontaneous in the reverse.

ox ox

$$3\,Ce^{3+}\,(aq) + Cr^{3+}\,(aq) \longrightarrow 3\,Ce^{4+}\,(aq) + Cr\,(s)$$

Thus, which must be more easily oxidized? Ce^{3+} or Cr?
(Ce^{4+}/Ce^{3+} is not on reference chart.)