WKS Name_____

Oxidation and Reduction (Redox) Period_____

1) When a substance is oxidized, its charge **(increases, decreases).**
Thus, when a substance is oxidized it is **(gaining, losing)** electrons.

2) When a substance is reduced, its charge **(increases, decreases**).
Thus, when a substance is reduced it is **(gaining, losing)** electrons.

3) Label these changes as either oxidation or reduction:

_____a) Cu^{2+} (aq) → Cu (s)

_____b) Cl_2 (g) → Cl^- (aq)

_____c) Fe (s) → Fe^{3+} (aq)

_____d) Cr^{3+} (aq) → Cr^{6+} (aq)

4) Put in all charges (oxidation numbers) in the following substances.
(Remember: The charge on pure elements is zero. Also, when H is in a compound, its charge is +1)

a) KCl b) $Fe(NO_3)_2$ c) HCl d) H_2 e) $Co_2(SO_4)_3$ e) Na

5) Put in all charges. Then show what is being oxidized or reduced by connecting with brackets.

a) Cu (s) + 2 AgCl (aq) → $CuCl_2$ (aq) + 2 Ag (s)

b) 2 Na (s) + Br_2 (g) → 2 NaBr (aq)

c) 2 AuCl (aq) + 2 HCl (aq) → 2 $AuCl_3$ (aq) + H_2

6) Put in all charges and then write the net ionic equation. *(Only write substances whose charges change, not spectator ions.)*

Zn (s) + $Cu(NO_3)_2$ (aq) → Cu (s) + $Zn(NO_3)_2$ (aq)

Net Ionic Eq:

7) For the net ionic equations below, do the following:
- *Write the two half reactions (ox and red) -- make sure you write in electrons on correct side.*
- *Balance equation by balancing out the electrons.*

a) K (s) + Pb^{4+} (aq) → K^+ (aq) + Pb^{2+} (aq)

Oxidation:

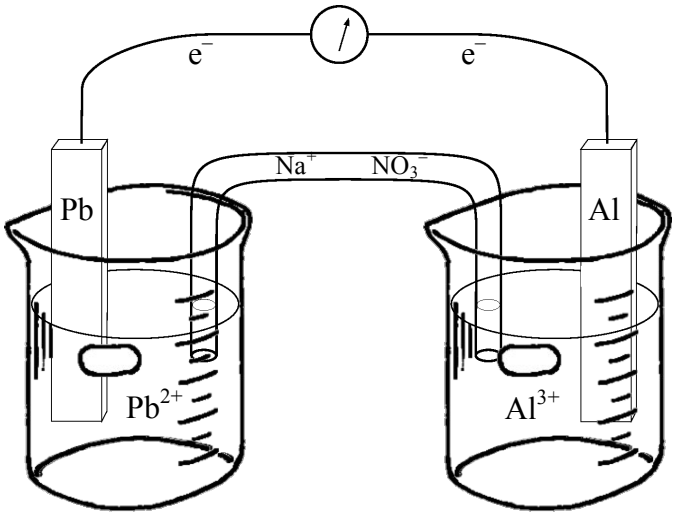
Reduction:

b) Cr^{3+} (aq) + Mg (s) → Cr (s) + Mg^{2+} (aq)

Oxidation:

Reduction:

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



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?_____

b) Thus, what will be oxidized? _____ c) What will be reduced?_____

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.**

Oxidation:

Reduction:_____

Overall reaction:

e) During which half reaction are electrons being lost? _____Which electrode is losing e^-? ____

f) During which half reaction are electrons being gained?____Which electrode is gaining e^-? ____

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) _____Why?

i) Which electrode will lose mass during the reaction? (Pb or Al)_____ Why?

j) Explain why the ions in the salt bridge moved the way you labeled them in your diagram.

k) What would happen if the salt bridge were removed? Explain why this happens.

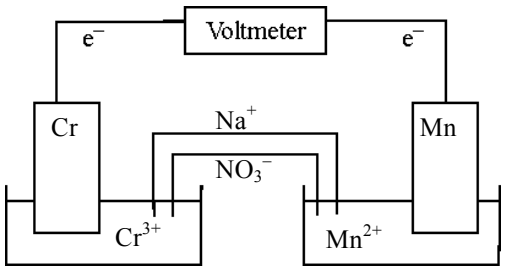
9) For the voltaic cell in #8 above (Pb^{2+}/Pb & Al^{3+}/Al), rewrite the two half reactions, determine their E° values, write the overall redox reaction and calculate the $E°_{net}$. Remember that cell voltages do NOT get multiplied.

Oxidation: $E^{\circ}_{ox} =$

Reduction: $E^{\circ}_{red} =$

Overall reaction: $E^{\circ}_{net} =$

10) 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? _____)*
   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^0net for the overall reaction.

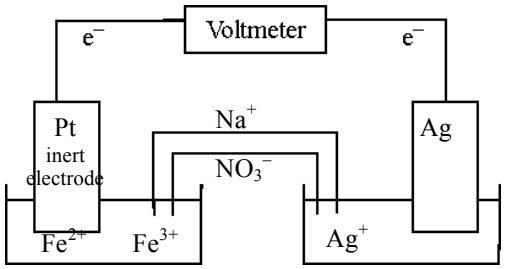


Ox: $E^{\circ} =$

Red: $E^{\circ} =$

Overall Rxn: $E^{\circ}_{net} =$

11) 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^0net for the overall reaction.



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

Ox: $E^{\circ} =$

Red: $E^{\circ} =$

$E^{\circ}_{net} =$

12) Will the following reaction be spontaneous in the forward direction? _____
Hint: Which is more easily oxidized? I^- or H_2? (Use reference chart to determine E°_{net}.)

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

13) This reaction is <u>spontaneous</u> in the forward direction.

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

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

14) The following reaction is <u>not spontaneous</u> in the forward direction, but is spontaneous in the reverse.

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

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

Answers: 9) $E°_{net} = +1.53$ V; 10) $E°_{net} = +0.44$ V; 11) $E°_{net} = +0.03$ V; 12) $E°_{net} = -0.54$ V