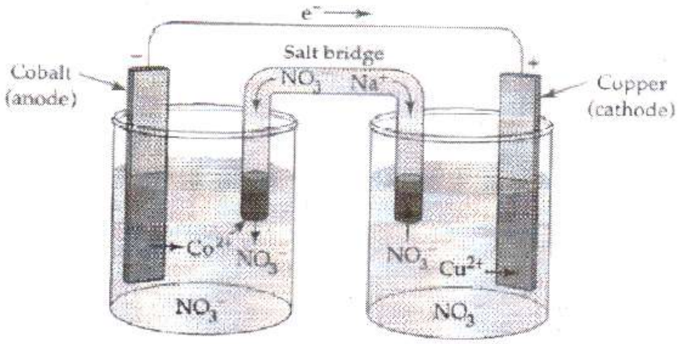
South Pasadena • AP Chemistry Name ______________________ Period___

21 • Electron Transfer Reactions

STATION 1 – CELL NOTATION

anode reaction $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$
cathode reaction $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
overall reaction $Co(s) + Cu^{2+}(aq) \rightarrow Co^{2+}(aq) + Cu(s)$



The "cell notation" for this electrochemical cell is $Co(s) \mid Co^{2+} \parallel Cu^{2+} \mid Cu(s)$

Use the above information to answer the following questions:

1. The left portion of the cell notation represents the ANODE ([✓ anode] / cathode).
2. The "||" represents the SALT BRIDGE (anode / cathode / [✓ salt bridge])
3. Write the cell notation for $Cl_2(g) + Zn(s) \rightarrow 2\,Cl^- + Zn^{2+}$ $Zn \mid Zn^{2+} \parallel Cl_2 \mid Cl^-$
4. Write the cell notation for $2Ag(s) + Pt^{2+} \rightarrow Pt(s) + 2Ag^+$ $Ag \mid Ag^+ \parallel Pt^{2+} \mid Pt$

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Standard Reduction Potentials (volts)

$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.535
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.337
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$	+0.15
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	-0.14
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$	-0.40
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	-0.763
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.828
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66

STATION 2 – E° VALUES

1. A cell is made from Sn in 1.0 M $Sn(NO_3)_2$ and Al in 1.0 M $Al(NO_3)_3$. The E° of the cell is +1.52 volts.
   $E^\circ_{cell} = E^\circ_{red} + E^\circ_{ox} = -.14 + (+1.66)$
2. A cell is made from Sn in 1.0 M $Sn(NO_3)_2$ and Cd in 1.0 M $Cd(NO_3)_2$. The E° of the cell is +.26 volts.
   $-.14 + (+.40)$
3. A cell is made from Ag in 1.0 M $AgNO_3$ and Cu in 1.0 M $Cu(NO_3)_2$. The E° of the cell is +.463 volts. (+.46)
   ~~+.337+~~
   $+.80 + (-.337)$
4. A cell is made from Zn in 1.0 M $Zn(NO_3)_2$ and Ag in 1.0 M $AgNO_3$. The E° of the cell is +1.563 volts. (+1.56)
   $+.80 + (+.763)$

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STATION 3 – NERNST EQUATION

Standard Reduction Potentials (volts)

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	-0.14
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66

$$E_{cell} = E° - \frac{RT}{nF} \ln Q$$

look at your equation sheet for R and F.
"n" is the moles of electrons gained or lost in a redox reaction.

1. $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$ E° = +0.78 V
   - a) What is n? 2 moles
   - b) If $[Cu^{2+}] = 0.10$ M and $[Fe^{2+}] = 1.5$ M,

     $Q = \frac{[1.5]}{[.10]} = 15$ $\ln(15) = 2.708$
   - c) Calculate the E_{cell}.

     $= +.78v - \frac{(8.31\ \frac{volt \cdot \cancel{C}}{\cancel{mol \cdot K}})(298\cancel{K})}{(2\cancel{mol})(96500\ \frac{\cancel{C}}{\cancel{mol}})}(2.708) =$

     $= +.745 = \boxed{.75\ v}$

2. A cell is made from Sn in .25 M $Sn(NO_3)_2$ and Al in 0.25 M $Al(NO_3)_3$ at 25°C.
   - a) The E° of the cell is 1.52 volts. $-.14 - (+1.66) =$
   - b) The reaction at the anode is: $Al \rightarrow Al^{3+} + 3e^-$
   - c) The reaction at the cathode is: $Sn^{2+} + 2e^- \rightarrow Sn$
   - d) The overall reaction is: $2Al + 3Sn^{2+} \rightarrow 2Al^{3+} + 3Sn$
   - e) The value of n is 6 moles.
   - f) $Q = \frac{[Al^{3+}]^2}{[Sn^{2+}]^3} = \frac{(.25)^2}{(.25)^3} = \frac{1}{.25} = 4$ $\ln(4) = 1.386$
   - g) Calculate the E_{cell}.

     $E_{cell} = 1.52v - \frac{(8.31\ \frac{volt \cdot \cancel{C}}{\cancel{mol \cdot K}})(298K)}{(6\ mol)(96500\cancel{C})}(1.386) = \boxed{1.51\ volts}$

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STATION 4 – BALANCING REDOX EQ'S(ACIDIC)

Balance the following equations in acidic solution:

$$Cr_2O_7^{-2}(aq) + C_2O_4^{-2}(aq) \rightarrow Cr^{+3}(aq) + CO_2(g)$$

$\cancel{6e^-} + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

$3\,C_2O_4^{2-} \rightarrow \overset{6}{\cancel{2}}CO_2 + \overset{\cancel{6}}{\cancel{2e^-}}$

$Cr_2O_7^{2-} + 3C_2O_4^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 6CO_2 + 7H_2O$

$$MnO_4^-(aq) + SO_2(g) \rightarrow SO_4^{-2}(aq) + Mn^{+2}(aq)$$

$\overset{10}{\cancel{5e^-}} + \overset{16}{\cancel{8H^+}} + {}^{2}MnO_4^- \rightarrow {}^{2}Mn^{2+} + \cancel{\overset{8}{4}H_2O}$

$\overset{\cancel{10}}{2}\,\cancel{4}H_2O + {}^{5}SO_2 \rightarrow {}^{5}SO_4^{2-} \overset{4}{+} \overset{\cancel{8}0}{\cancel{4}}H^+ + \overset{10}{\cancel{2e^-}}$

$2MnO_4^- + 5SO_2 + 2H_2O \rightarrow 5SO_4^{2-} + 2Mn^{2+} + 4H^+$

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STATION 5 – BALANCING REDOX EQ'S (BASIC)

Balance the following equations in basic solution:

$$Mn^{+2}(aq) + ClO_3^-(aq) \rightarrow MnO_2(s) + ClO_2(g)$$

$$\cancel{2H_2O} + Mn^{2+} \rightarrow MnO_2 + \cancel{4H^+} + \cancel{2e^-}$$

$$\cancel{2e^-} + \overset{4}{\cancel{2H^+}} + {}^{2}ClO_3^- \rightarrow {}^{2}ClO_2 + {}^{2}\cancel{H_2O}$$

$$Mn^{2+} + 2ClO_3^- \rightarrow MnO_2 + 2ClO_2$$

$$Cl_2(g) \rightarrow Cl^-(aq) + ClO_3^-(aq)$$

$$\overset{10}{\cancel{2e^-}} + {}^{5}Cl_2 \rightarrow \overset{10}{\cancel{2}}Cl^-$$

$$6H_2O + Cl_2 \rightarrow 2ClO_3^- + 12H^+ + \cancel{10e^-}$$

$$6Cl_2 + 6H_2O \rightarrow 10Cl^- + 2ClO_3^- + 12H^+ + 12OH^-$$

$$+12OH^-$$

$$6Cl_2 + \cancel{6H_2O} + 12OH^- \rightarrow 10Cl^- + 2ClO_3^- + \overset{6}{\cancel{12}}H_2O$$

$$\boxed{6Cl_2 + 12OH^- \rightarrow 10Cl^- + 2ClO_3^- + 6H_2O}$$

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STATION 6 – ELECTROLYSIS

How long will it take to electroplate each of the following with a current of 100.0 A?

1.0 g of Al(s) from aqueous Al^{+3}.

$$\frac{100.0\ \cancel{Amp}}{1.0\ g\ \cancel{Al}} \times \frac{1\ \cancel{C}}{1\ amp\cdot s} \times \frac{1\ \cancel{mole\ e^-}}{96500\ \cancel{C}} \times \frac{1\ \cancel{mole\ Al}}{3\ \cancel{mole\ e^-}} \times \frac{27.0\ \cancel{g\ Al}}{1\ \cancel{mol\ Al}} = .009326\ s^{-1}$$

$$\frac{1}{ans} = \boxed{107.2\ sec}$$

← Two ways to Set up the equation

1.0 g of Ni(s) from aqueous Ni^{+2}.

$$\frac{1.0\ g\ \cancel{Ni}}{100.0\ \cancel{Amp}} \times \frac{1\ \cancel{mole\ Ni}}{58.69\ \cancel{g}} \times \frac{2\ \cancel{mole\ e^-}}{1\ \cancel{mole\ Ni}} \times \frac{96500\ \cancel{C}}{1\ \cancel{mol\ e^-}} \times \frac{1\ amp\cdot s}{1\ \cancel{C}} = 32.88\ sec$$

$$\boxed{33\ sec}$$

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STATION 7 – REACTIVITY

Consider the following half-reactions and E° values:

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E^\circ = 0.80$ V
$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$ $E^\circ = 0.34$ V
$Pb^{+2}(aq) + 2e^- \rightarrow Pb(s)$ $E^\circ = -0.13$ V

1. Which of these metals or ions is the strongest **oxidizing agent**? Ag^+ (most likely to be reduced)
2. Which is the strongest **reducing agent**? Pb(s) (most likely to get oxidized)

Predict whether each of the following reactions will occur as written:

3. $Cu^{2+} + Pb^\circ \rightarrow Pb^{2+} + Cu^\circ$ Y
4. $Pb^{2+} + 2Ag^\circ \rightarrow 2Ag^+ + Pb^\circ$ N (opposite reaction would occur)
5. $2Ag^+ + Pb^{2+} \rightarrow 2Ag^\circ + Pb^\circ$ N (both reduced)

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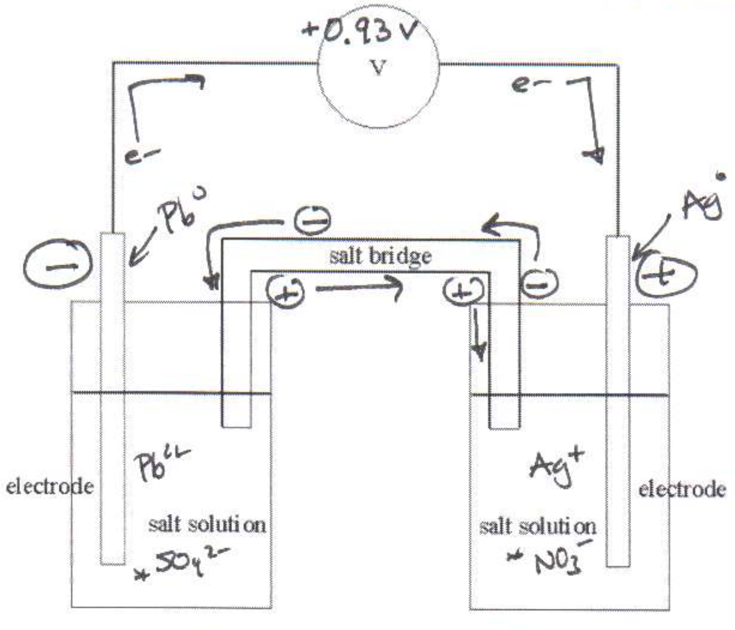
STATION 8 – SKETCH A CELL



Anode oxidation reaction

$Pb \rightarrow Pb^{2+} + 2e^-$

Cathode reduction reaction

$Ag^+ + e^- \rightarrow Ag$

$Pb + 2Ag^+ \rightarrow 2Ag + Pb^{2+}$

Consider these half-reactions & E° values:

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E^\circ = 0.80$ V
$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$ $E^\circ = 0.34$ V
$Pb^{+2}(aq) + 2e^- \rightarrow Pb(s)$ $E^\circ = -0.13$ V

Which two metals and 1.0 M solutions would give the greatest voltage? Ag Pb

Label:

- ✓ the anode reaction
- ✓ the cathode reaction
- ✓ the overall reaction
- ✓ the metals used for each electrode
- ✓ the ions in solution * could be different
- ✓ the expected voltage $E^\circ_{cell} = +.80 + (+.13)$
- ✓ the direction of flow of electrons ANODE → CATHODE
- ✓ the flow of ions in the salt bridge
- ✓ the charge on each electrode (+ or –)
- ✓ ions you might use in the salt bridge K^+ NO_3^-
- ✓ the observed changes in the electrodes

ANODE GETS SMALLER
CATHODE GETS LARGER