# Electrochemisty Unit: Assignment 11

## Lecture notes on The Galvanic Cell Part I

**DIRECTIONS:** view "Galvanic Cell Part I" video and answer the following questions. Place in project folder when finished.

Looking at the reaction taking place between the cooper metal placed in the silver nitrate solution, list three visible evidences that a reaction is happening:

What happened to the electrical energy generated by this REDOX reaction?

What was the idea to harness this otherwise lost energy?

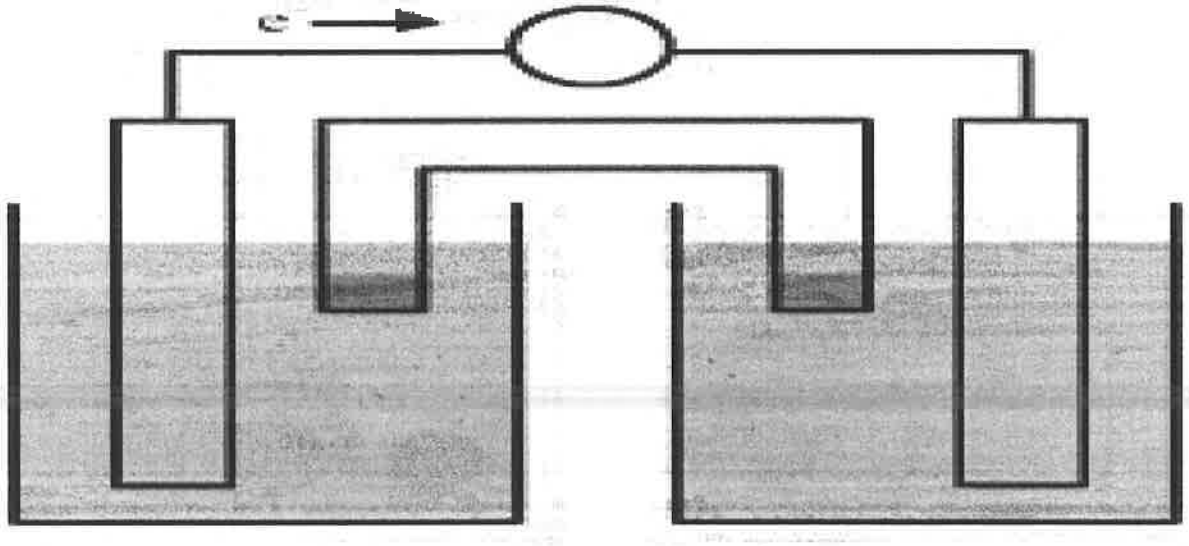
What are the two half reactions for the copper/silver reaction?

What must each compartment contain to maintain a ½ reaction?

What must we be careful of when making the solutions?

What is the ideal polyatomic ion? Why?

Label the galvanic cell for the copper/silver reaction:




*Continued on next page*

Make sure to show:

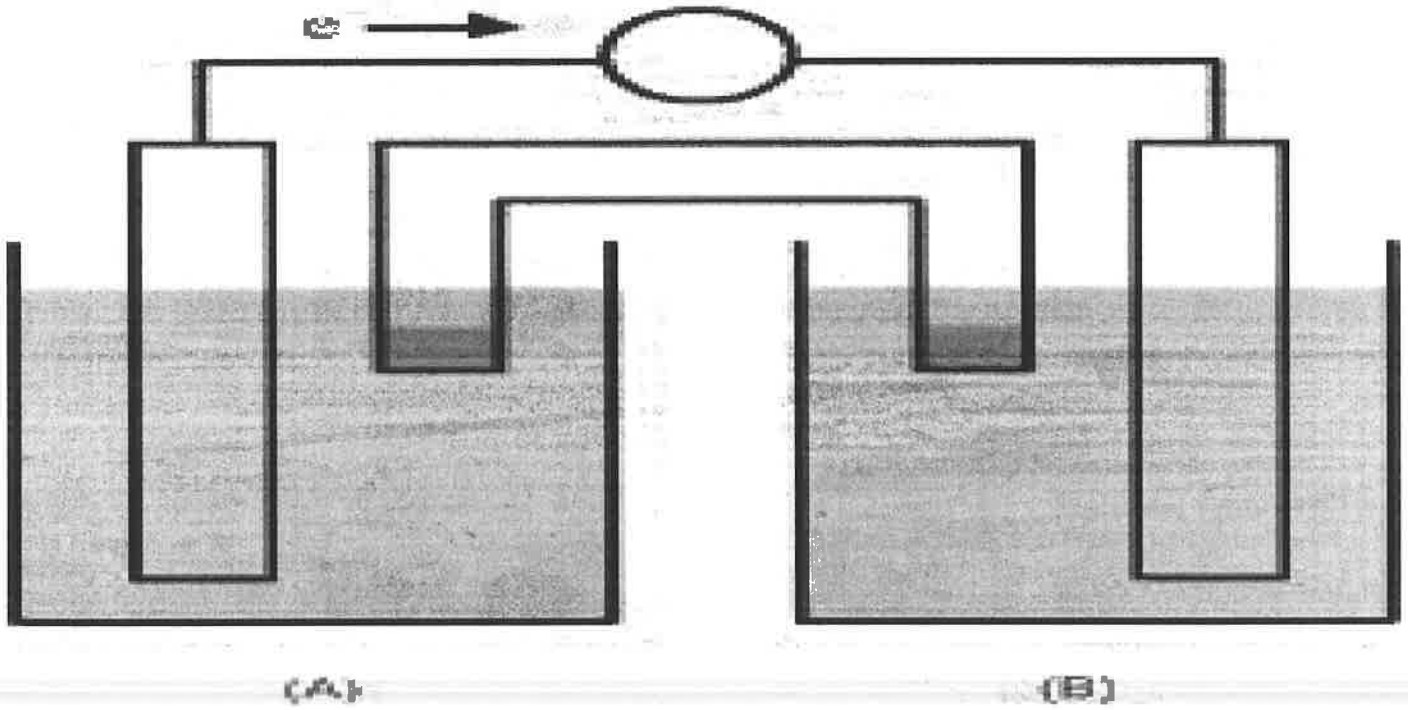
- What direction the electrons flow (from anode to cathode)
- What takes place in anode (oxidation)
- What takes place in cathode (reduction)
- Identity of the electrodes
- The specific ions in solution

Salt Bridge:

What happens to the charge in the anode and cathode compartments as the reaction proceeds: (show this on galvanic cell)

What would happen as a result of this "charge" buildup?

What salts make good salt bridges and why?



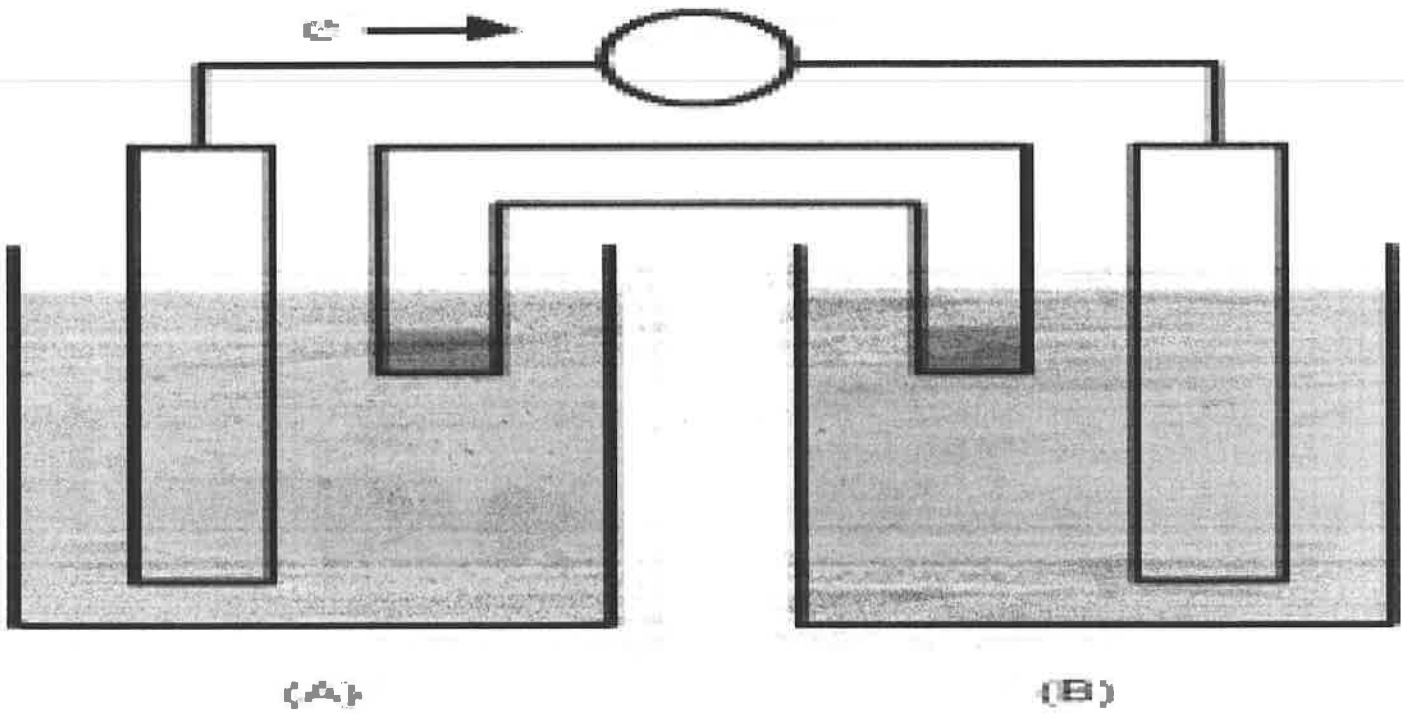
What specific direction will the cation and anion flow?

*Continued on next page*

Your Turn!

Set up the correct galvanic cell (including salt bridge!) for the following reaction:

$$Zn + Sn(NO_3)_2 \rightarrow Sn + Zn(NO_3)_2$$




After completing, place in unit folder

# Electrochemisty Unit: Assignment 12

## Lecture notes on The Galvanic Cell Part II

**DIRECTIONS:** view "Galvanic Cell Part II" video and answer the following questions. Place in project folder when finished.

Draw, at the particulate level, what is happening at the anode $Cu \rightarrow Cu^{2+} + 2e^-$

Draw, at the particulate level, what is happening at the cathode $Zn^{2+} + 2e^- \rightarrow Zn$

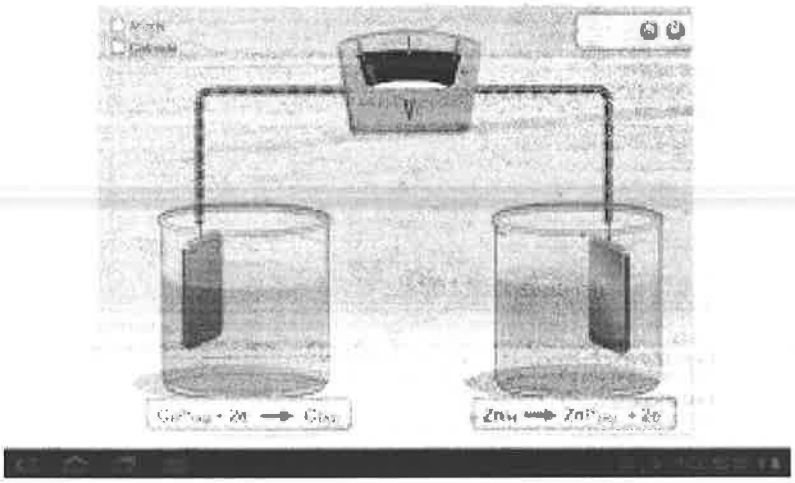
In summary:

- _____ gets _____ as neutral _____are converted into soluble _____
- _____gets _____as soluble _____are converted into neutral _____

What is the common misconception on how the galvanic cell is set up

$Cu(s) + 2Zn(NO_3)_2(aq) \rightarrow 2Zn(s) + Cu(NO_3)_2(aq)$ write the net ionic equation for this reaction.

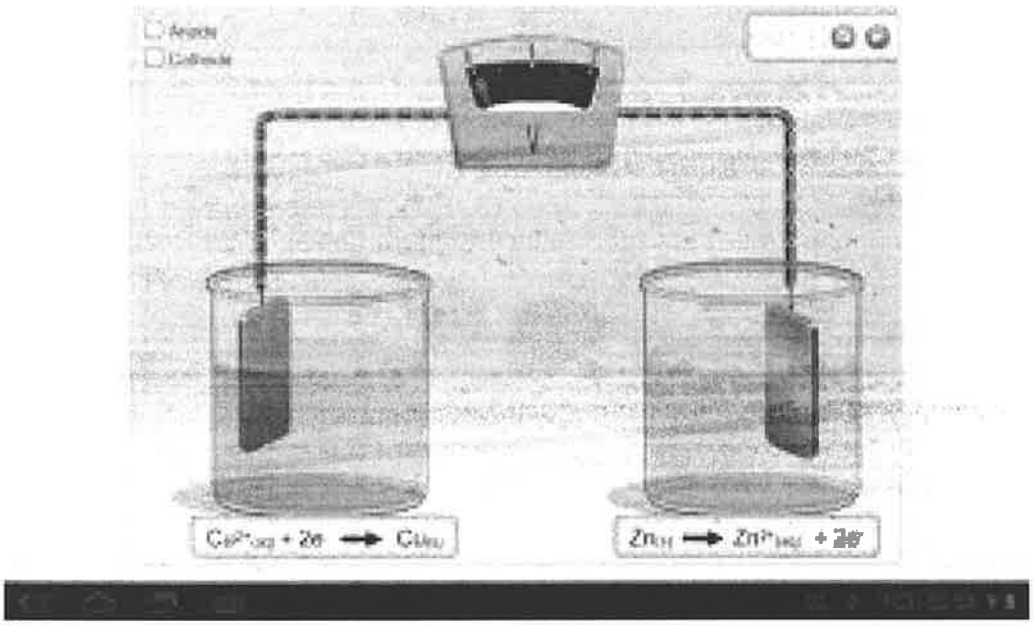
Show in the galvanic cell where the reactants and products are located



Cells, eventually die. Even though the electrodes are still there and the salt bridge still is replenishing charge, the cell will eventually die.

To explain why, we must review concepts of equilibrium. Write the equilibrium expression for the reaction:

Show what is happening in the cell as the reaction progresses over time.



What is happening to the equilibrium ratio as the reaction continues? What is the "response" to re-establish the equilibrium?

After completing, place in unit folder

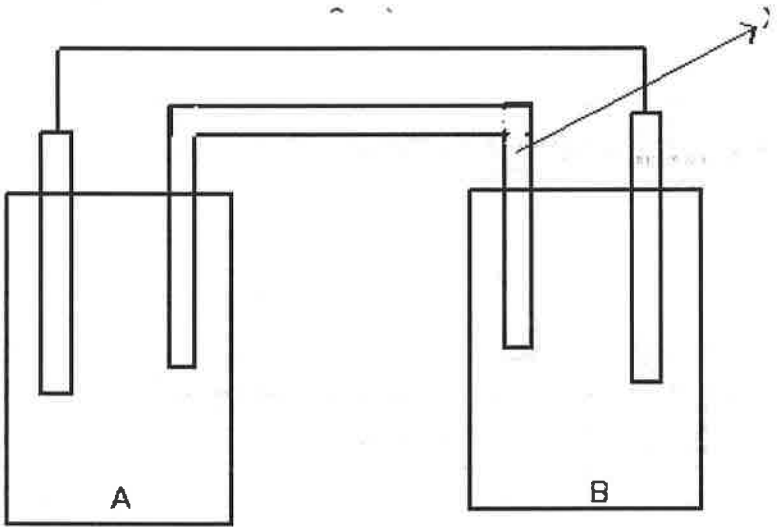
# Electrochemistry Unit: Assignment 13

## Galvanic Cell Practice (practice from lecture in assignment 11 and 12)

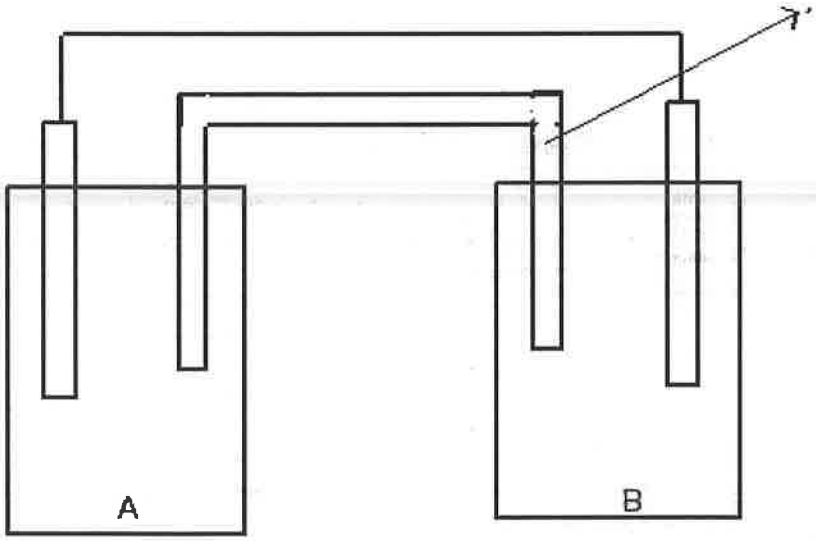
Directions: for the following reactions:

- label the anode and the cathode
- label which metal is the anode and which metal is the cathode
- identify the anode compartment and the cathode compartment
- identify the specific solutions that are in the compartments
- under each compartment write the correct half reaction occurring in the compartment
- show the direction the electrons flow during cell operation
- identify the specific slat in the slat bridge
- identify the direction the ions in the salt bridge are flowing

$Ca + Zn(NO_3)_2 \rightarrow Zn + Ca(NO_3)_2$




Iron is placed in a solution of cobalt (II) nitrate and forms cobalt metal and iron (II) nitrate




*continued on next page*

Particulate level drawing of what is happening to the size of the anode and cathode as the galvanic cell is running.

draw a rectangle for electrode and circles for atoms/ions to show what happens to size of electrode before and during the running of a galvanic cell of this reaction:

$$Cu + Zn(NO_3)_2 \rightarrow Zn + Cu(NO_3)_2$$

Anode

Cathode

Write a balanced chemical equation for single replacement reaction that could be used to set up a galvanic cell. Write the equilibrium expression for this reaction. Using the principles of equilibrium and La Chateleir, explain why a galvanic cell cannot run forever, will eventually "die".

After completing, place in your unit folder

# Electrochemisty Unit: Assignment 14

## Lecture notes on EMF

What does "EMF" stand for?

What is the symbol?

What are "standard conditions"?

What is the unit of electrical energy?

What is the reason for the movement?

What is meant by SRP?

What are the table values in reference to? (Hint: SHE)

Since the flow of electrons is from the anode to the cathode, the anode must have a lower/higher SPR than the cathode.

Which set up would work, and why?

1 2

$Cu + Zn(NO_3)_2 \rightarrow Zn + Cu(NO_3)_2$ vs $Zn + Cu(NO_3)_2 \rightarrow Cu + Zn(NO_3)_2$

Copper oxidized zinc oxidized

Zinc reduced copper reduced

SRP answer:

Cu: + 0.34V

Zn: - 0.76V

Your turn:

- If you made a galvanic cell with iron and nickel which would need to be the anode?
- Write the equation that reflects this

Write the equation that calculates the $E^o$ of the cell. Write specifically what each "E" symbolizes

Calculate the $E^o$ of the Fe/Ni cell

*Continued on next page*

What does the value of $E^o$ tell us (+ vs -)

Calculate the value of the Cu/Zn cell mentioned earlier in your notes

What is the similarities and differences of AAA, AA, C, and D batteries

What is an analogy to explain the relationship between voltage and current?

After completing, place in unit folder

# Electrochemistry Unit: Assignment 15

## EMF Practice (practice from lecture in assignment 14)

Would the following reactions work as written?

- If YES, justify with $E^o$ calculation
- If NO, then rewrite the equation so that it will work, then justify with $E^o$ calculation

$Sr + LiNO_3 \rightarrow Li + Sr(NO_3)_2$

$Fe + Sn(NO_3)_2 \rightarrow Sn + Fe(NO_3)_2$

$Ag + Be(NO_3)_2 \rightarrow Be + AgNO_3$

$Cr + Ca(NO_3)_2 \rightarrow Ca + Cr(NO_3)_3$

Which combination of metals would produce the largest voltage? The smallest voltage?
Write the equation and calculate the $E^o$ of your two choices to prove your statement.

$\underline{E_{red}}$:

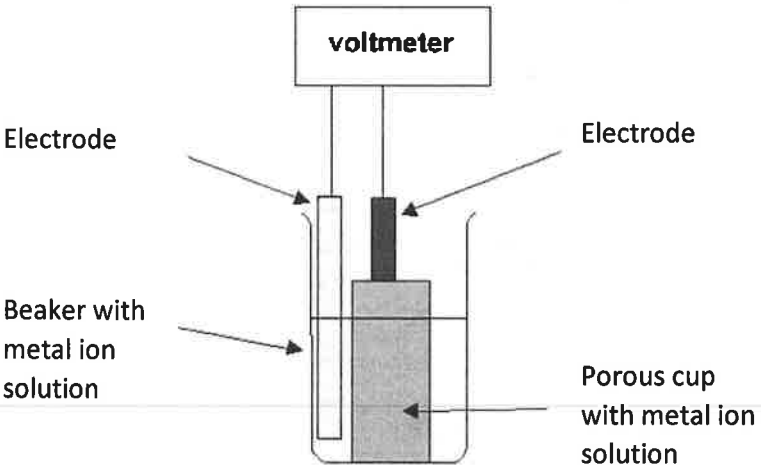
$Ba/Ba^{2+}$
$Cu/Cu^{2+}$
$Cr/Cr^{3+}$
$Au/Au^{3+}$

If you wanted to compete with name brand batteries that use Cu/Zn , is there another combination of metals that would give an $E^o$ of approximately 1.00-1.75 V? Uisng your understanding of properties of metals and cost of metals, what would be the pros and cons of your combination?

If AAA, AA, C, D sized batteries all produce the same voltage, why are there 4 different sizes?

After completing, place in your unit folder

## Electrochemistry Unit: Assignment 16 Galvanic Cell Lab (Based on lessons 11-14)




**Pre-Lab**

Make a galvanic cell with two of the following metals and their solutions:

Cu, Zn, Mg, Al, Fe

1. Write an equation for the cell you are choosing to make (make sure it works according to SRP's)

2. Determine what is oxidized and what is reduced

3. Determine what needs to be in anode compartment and cathode compartment

4. Calculate the voltage your galvanic cell should generate according to SRP's

**Lab:**

5. Make your galvanic cell in the following way:
   a. Use a 250 mL beaker as the cathode compartment
   b. Place around 50 mL of the "cathode" solution into the beaker
   c. Place porous cup into the beaker. The porous cup will be the anode compartment (the porous cup allows migration of ions thus negating the need for a salt bridge)
   d. Fill the porous cup about 2/3 full with the "anode" solution
   e. Place the metal electrodes in the appropriate ½ cell
   f. Attach one end of an alligator clip to each electrode (so two separate alligator clips)
   g. Make sure the alligator clips are attached to voltmeter correctly: (-) anode and (+) cathode
   h. Read the voltmeter and record the voltage (instructions on voltmeter settings on board)
6. How does your actual voltage compare to the value calculated in step 4?

7. In the space below, draw your galvanic cell in the "classic fashion" labeling all the pertinent points discussed in lectures:

Turn in for a separate grade by May 15th

# Electrochemisty Unit: Assignment 17

## Lecture notes on The Electrolytic Cell

**DIRECTIONS:** view "electrolytic Cell" video and answer the following questions. Place in project folder when finished.

What is the driving force of the galvanic cell?

What type of elements are normally oxidized? Why

What type of elements are normally reduced? Why?

In a galvanic cell why would sodium HAVE to be oxidized and chlorine HAVE to be reduced?

What is the driving force of an electrolytic cell?

What is important to know with an electrolytic cell?

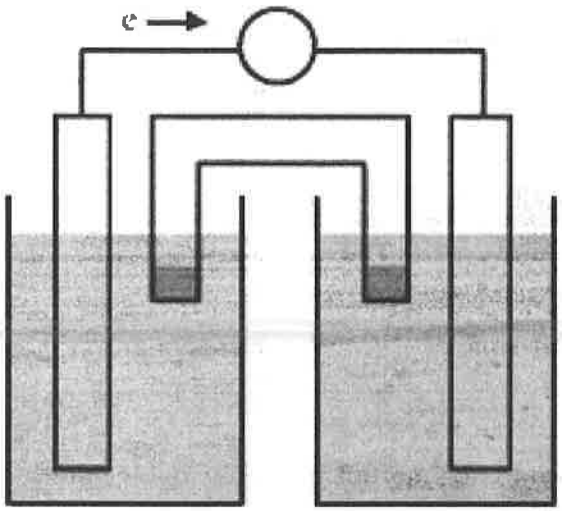
From the video, write the comparison of the two cells:

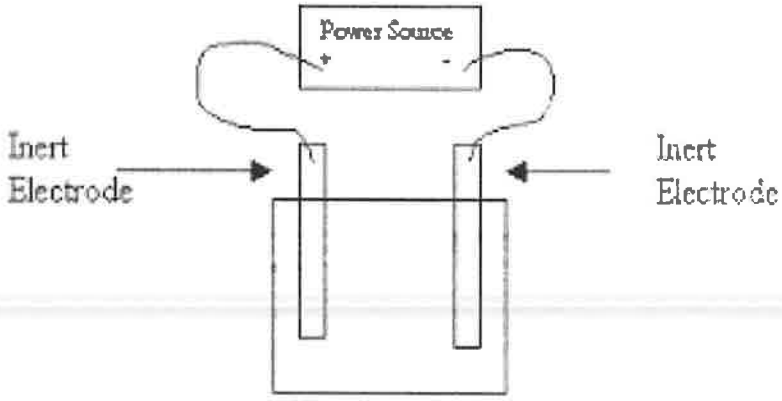
$Rb + 1e^- \rightarrow Rb \quad E^o = -2.92V$
$Cu + 2e^- \rightarrow Cu \quad E^o = +0.34V$

Galvanic $E^o = +3.26V$

$Na^+ + 1e^- \rightarrow Na \quad E^o = -2.71V$
$Cl_2 + 2e^- \rightarrow 2Cl^- \quad E^o = +1.36V$

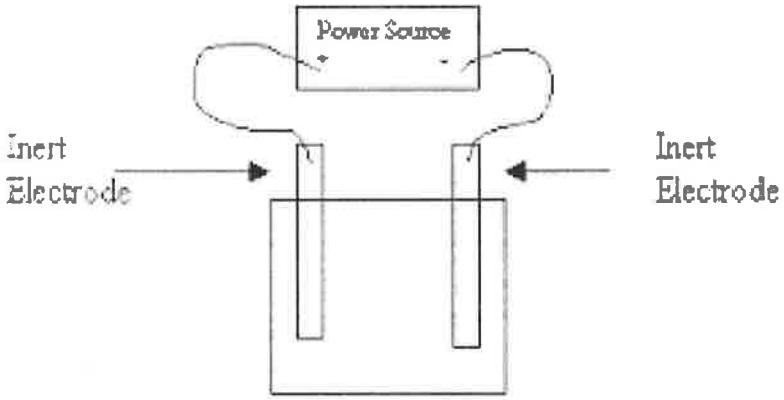
Electrolytic $E^o = -4.06V$



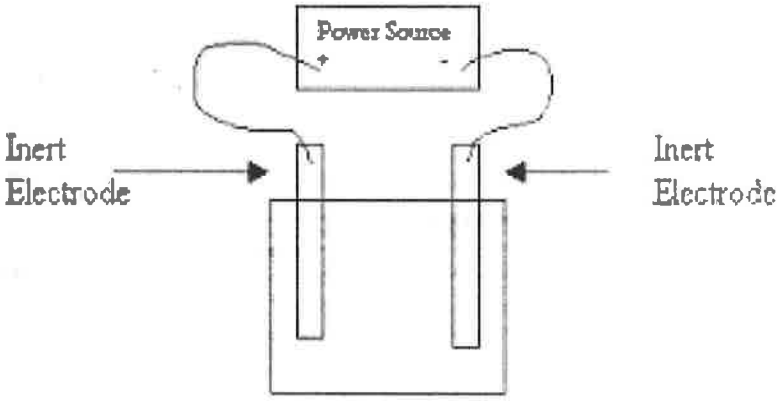

Con't

Now look closer at just the electrolytic cell that has molten sodium chloride inside the compartment




You try: electrolytic cell using potassium bromide:

- Look at SPR/s, what will be reduced/oxidized
- Calculate $E^o$
- Label the cell using previous example as template




Comparison of galvanic and electrolytic

| **Galvanic** | **electrolytic** |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |
| | |
| | |
| | |
| | |
| | |

After completing, place in unit folder

# Electrochemistry Unit: Assignment 18

## Electrolytic Cell practice (practice from lecture in assignment 17)

**Directions**: answer the following questions, draw appropriate cell when needed

1. Bullet the ways in which a galvanic and electrolytic cell differ

2. Bullet the ways in which a galvanic cell and electrolytic cell are similar

3. What reactions take place at the anode and cathode when each of the following is electrolyzed?
   a. Molten $NiBr_2$ b. molten $AlF_3$ c. molten $MnCl_2$

4. Starting with solid copper(II) iodide. Construct and correctly label all pertinent points of an electrolytic cell. Determine what is reduced and what is oxidized, determine the voltage that must be provided.

After completing, staple your work to this page and place in your unit folder

# Electrochemisty Unit: Assignment 19

## Lecture notes on Stoichiometry and Electrolysis

**DIRECTIONS:** view "Stoichiometry and Electrolysis" video and answer the following questions. Place in project folder when finished.

Define the following terms:

Coulomb (C)

Amphere/amp (A)

Faraday ($\mathcal{F}$)

Draw and label the stoichiometry triangle:

**Use the stoichiometry triangle to solve the following problems:**

how many grams of Cu would be deposited from a solution of copper(II) sulfate by a current of 3.00 amps flowing for 6.00 hours?

how long does it take to deposit 15.6 grams of nickel, when a current of 2.5 amps is run through a solution of nickel (II) chloride?

A total of 69,500 coulombs of electricity was required in the electrolytic reduction of 12.5 g of a metal from a solution of its tripositive ions. what is the metal?

*After completing, place in unit folder*

# Electrochemistry Unit: Assignment 20

## Stoichiometry and Electrolysis practice (practice from lecture in assignment 19)

**Directions**: answer the following questions, show all work!

1. Draw and label the electrolytic triangle

2. A $Cr^{3+}$ solution is electrolyzed using a current of 7.60 A. What mass of Cr is plated out after 2 hr?

Ans: 9.8 g

3. What amperage is required to plate out 2.3 grams of Au from a $Au^{2+}$ solution in a period of 3.5 hours?

Ans: .177 amps

4. Metallic magnesium can be made by electrolysis of molten $MgCl_2$. How many minutes are needed to plate out 10.0 grams Mg using 3.50 A current?

*Ans: 378 min*

5. Calculate the mass of Fe formed by electrolysis of $FeCl_3$ by a current of $7.5 \times 10^4$ A flowing for a period of 10 hours. Assume the electrolytic process is 85% efficient.

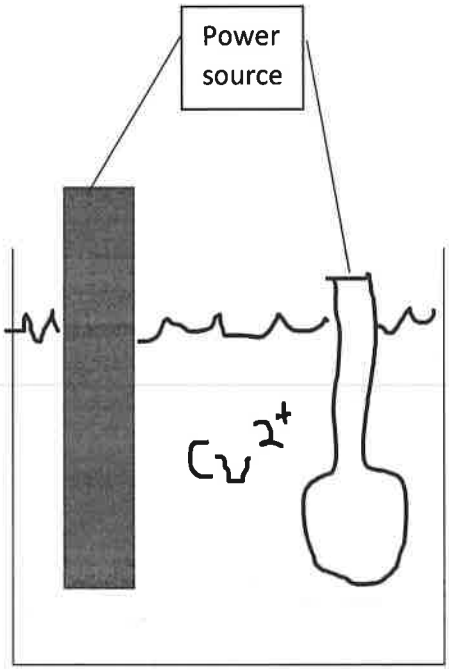
*Ans: 1.28 x10$^6$g*

6. An unknown metal, M, is electrolyzed. It took 9.55 min for a current of 2.00 A to plate out .107g from a solution containing $M(NO_3)_3$. Identify the metal.

*Ans: Aluminum*

*Continued on next page*

7. Copper can be plated onto a spoon by placing the spoon in an acidic solution of $CuSO_4$ and connecting it to a copper strip via a power source as illustrated below




a. Label the anode and cathode, and describe the direction of the electron flow
b. If you need $4.7 \times 10^{22}$ atoms of copper to coat the spoon, what combination of amp and time would be sufficient? Justify with calculations.

Ans: may vary, however amp x sec = *15068C*

After completing place in your unit folder

## STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

| Half-reaction | | | $E°$(V) |
|---|---|---|---|
| $Li^+ + e^-$ | → | $Li(s)$ | −3.05 |
| $Cs^+ + e^-$ | → | $Cs(s)$ | −2.92 |
| $K^+ + e^-$ | → | $K(s)$ | −2.92 |
| $Rb^+ + e^-$ | → | $Rb(s)$ | −2.92 |
| $Ba^{2+} + 2\,e^-$ | → | $Ba(s)$ | −2.90 |
| $Sr^{2+} + 2\,e^-$ | → | $Sr(s)$ | −2.89 |
| $Ca^{2+} + 2\,e^-$ | → | $Ca(s)$ | −2.87 |
| $Na^+ + e^-$ | → | $Na(s)$ | −2.71 |
| $Mg^{2+} + 2\,e^-$ | → | $Mg(s)$ | −2.37 |
| $Be^{2+} + 2\,e^-$ | → | $Be(s)$ | −1.70 |
| $Al^{3+} + 3\,e^-$ | → | $Al(s)$ | −1.66 |
| $Mn^{2+} + 2\,e^-$ | → | $Mn(s)$ | −1.18 |
| $Zn^{2+} + 2\,e^-$ | → | $Zn(s)$ | −0.76 |
| $Cr^{3+} + 3\,e^-$ | → | $Cr(s)$ | −0.74 |
| $Fe^{2+} + 2\,e^-$ | → | $Fe(s)$ | −0.44 |
| $Cr^{3+} + e^-$ | → | $Cr^{2+}$ | −0.41 |
| $Cd^{2+} + 2\,e^-$ | → | $Cd(s)$ | −0.40 |
| $Tl^+ + e^-$ | → | $Tl(s)$ | −0.34 |
| $Co^{2+} + 2\,e^-$ | → | $Co(s)$ | −0.28 |
| $Ni^{2+} + 2\,e^-$ | → | $Ni(s)$ | −0.25 |
| $Sn^{2+} + 2\,e^-$ | → | $Sn(s)$ | −0.14 |
| $Pb^{2+} + 2\,e^-$ | → | $Pb(s)$ | −0.13 |
| $2\,H^+ + 2\,e^-$ | → | $H_2(g)$ | 0.00 |
| $S(s) + 2\,H^+ + 2\,e^-$ | → | $H_2S(g)$ | 0.14 |
| $Sn^{4+} + 2\,e^-$ | → | $Sn^{2+}$ | 0.15 |
| $Cu^{2+} + e^-$ | → | $Cu^+$ | 0.15 |
| $Cu^{2+} + 2\,e^-$ | → | $Cu(s)$ | 0.34 |
| $Cu^+ + e^-$ | → | $Cu(s)$ | 0.52 |
| $I_2(s) + 2\,e^-$ | → | $2\,I^-$ | 0.53 |
| $Fe^{3+} + e^-$ | → | $Fe^{2+}$ | 0.77 |
| $Hg_2^{2+} + 2\,e^-$ | → | $2\,Hg(l)$ | 0.79 |
| $Ag^+ + e^-$ | → | $Ag(s)$ | 0.80 |
| $Hg^{2+} + 2\,e^-$ | → | $Hg(l)$ | 0.85 |
| $2\,Hg^{2+} + 2\,e^-$ | → | $Hg_2^{2+}$ | 0.92 |
| $Br_2(l) + 2\,e^-$ | → | $2\,Br^-$ | 1.07 |
| $O_2(g) + 4\,H^+ + 4\,e^-$ | → | $2\,H_2O(l)$ | 1.23 |
| $Cl_2(g) + 2\,e^-$ | → | $2\,Cl^-$ | 1.36 |
| $Au^{3+} + 3\,e^-$ | → | $Au(s)$ | 1.50 |
| $Co^{3+} + e^-$ | → | $Co^{2+}$ | 1.82 |
| $F_2(g) + 2\,e^-$ | → | $2\,F^-$ | 2.87 |