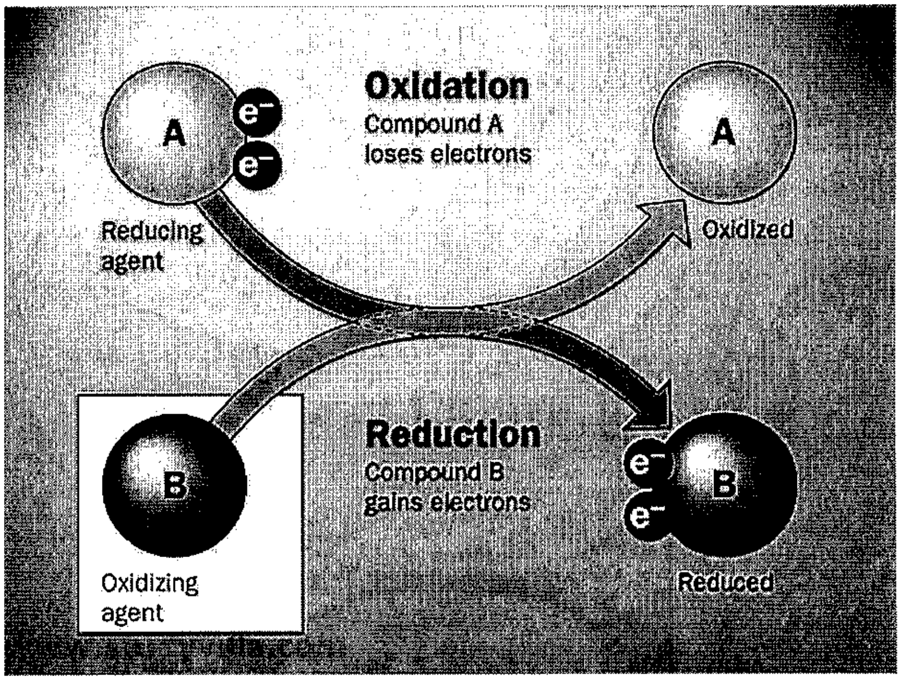
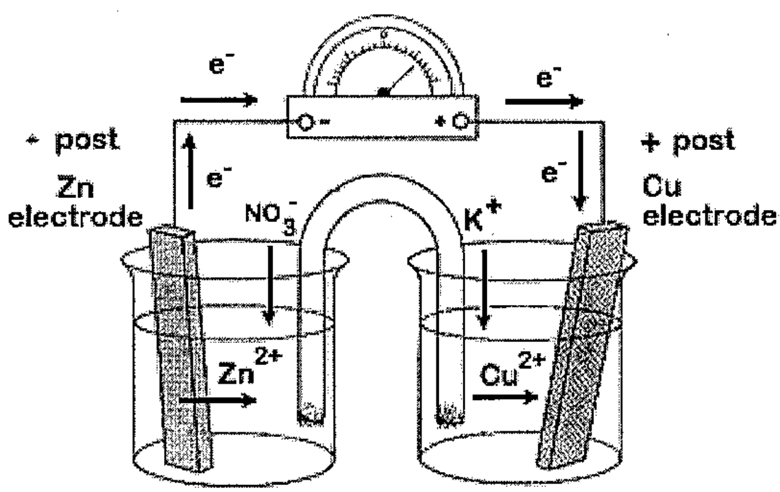
REDOX and Electrochemsitry





$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Anode oxidation

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

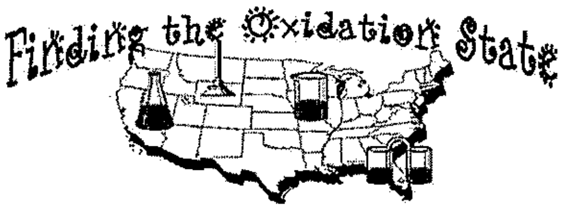
Cathode reduction

Chemistry: Form WS10.1.1A

REDOX AND ELECTROCHEMISTRY

Name ______________________

Date ______________ Period ______



Study the rules for assigning oxidation numbers and examine the sample problem below. Then determine the unknown oxidation state in each example.

RULES FOR ASSIGNING OXIDATION NUMBERS

1. Oxidation numbers for atoms that are free elements are always zero
2. The oxidation numbers of ions are the same as the charge on the ion
3. Some elements have only one oxidation state
   a. group 1 metals always form 1+ ions and always have a +1 oxidation state
   b. group 2 metals always form 2+ ions and always have a +2 oxidation state
4. Some elements usually have a particular oxidation state
   a. oxygen has a –2 oxidation state except in peroxides where it is –1 and in compounds with fluorine (OF_2) where it is +2
   b. hydrogen has a +1 oxidation state except in hydrides with group 1 and group 2 metals
5. the sum of the oxidation numbers
   a. in a compound it is always zero
   b. in a polyatomic ion it is equal to the charge on the ion

Sample Problem

Find the oxidation state of the elements in $K_2Cr_2O_7$.

Element	K	Cr	O	
Subscript	2	2	7	TOTAL
Oxidation state	+1	?	–2	
Sum of oxidation states	+2	??	–14	0

[a] potassium is a group one metal; its oxidation state is always +1
[b] oxygen usually has an oxidation state of –2
[c] the sum of oxidation states of each element is the product of the subscript and the oxidation state
[d] find the sum of the oxidation states of chromium (??) by setting the sum of all the oxidation states to zero

$$(+2) + ?? + (-14) = 0$$
$$?? = +12$$

[f] find the oxidation state of chromium (?) by dividing the sum (+12) by the subscript (2)

$$+12 \div 2 = +6$$

1. Chlorine in $KClO_4$ 1. ______
2. Nitrogen in $Ba(NO_3)_2$ 2. ______
3. Phosphorus in $Ca_3(PO_4)_2$ 3. ______
4. Manganese in $LiMnO_4$ 4. ______
5. Sulfur in Na_2SO_3 5. ______
6. Chromium in $CaCrO_4$ 6. ______
7. Sulfur in MgS_2O_3 7. ______
8. Nitrogen in $Zn(NO_2)_2$ 8. ______
9. Chlorine in $HClO_3$ 9. ______
10. Carbon in CaC_2O_4 10. ______
11. Sulfur in $KHSO_4$ 11. ______

ASSIGNING OXIDATION NUMBERS

Name ____________________

Assign oxidation numbers to all of the elements in each of the compounds or ions below.

1. HCl	11. H_2SO_3
2. KNO_3	12. H_2SO_4
3. OH^-	13. BaO_2
4. Mg_3N_2	14. $KMnO_4$
5. $KClO_3$	15. LiH
6. $Al(NO_3)_3$	16. MnO_2
7. S_8	17. OF_2
8. H_2O_2	18. SO_3
9. PbO_2	19. NH_3
10. $NaHSO_4$	20. Na

Name ______________________

Date ______________ Period ______

Competition for Electrons

Aim
- write equations for oxidation and reduction half reactions

Notes

Atoms compete for each other's electrons

★ When chemical bonds form, electrons are either lost, gained or shared

★ Oxidation-Reduction reactions (Redox reactions)
  - ☆ Metals
    - ★ lose electrons (OXIDATION)[*NOTE: as when metals combine with oxygen*]
    - ★ are oxidized
    - ★ are reducing agents
  - ☆ Nonmetals
    - ★ gain electrons reducing their oxidation states (REDUCTION)
    - ★ are reduced
    - ★ are oxidizing agents

Oxidation
Is
Loss

Reduction
Is
Gain

★ Example 1 - $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Mg	O_2
★ loses electrons ★ gets oxidized to Mg^{2+} ★ is the reducing agent for O_2	★ gains electrons ★ gets reduced to O^{2-} ★ is the oxidizing agent for Mg

☆ Half reactions — reaction showing either a gain or loss of electrons

★ $2Mg^0 \rightarrow 2Mg^{2+} + 4e^-$

★ $O_2^0 + 4e^- \rightarrow 2O^{2-}$

☆ Net equation (REDOX REACTION)— combination of the half reactions such that the number of electrons lost equals the number of electrons gained

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

★ Example 2 - More active metals replace less active metals in compounds by transferring electrons to them

☆ Sample Reaction:

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

☆ Half reactions — reaction showing either a gain or loss of electrons

★ $Zn^0 \rightarrow Zn^{2+} + 2e^-$

★ $Cu^{2+} + 2e^- \rightarrow Cu^0$

☆ Net equation — combination of the half reactions such that the number of electrons lost equals the number of electrons gained

$$Cu^{2+} + Zn^0 \rightarrow Zn^{2+} + Cu^0$$

☆ Spectator ions — ions that are present during a reaction but do not participate in the reaction:

$$2NO_3^-$$

Oxidation number (Oxidation state) - number assigned to keep track of electrons based on the arbitrary assumption that shared electrons belong to the more electronegative element

★ Rules for assigning oxidation numbers
  - ☆ Oxidation numbers for atoms that are free elements are always zero
  - ☆ The oxidation numbers of ions are the same as the charge on the ion
  - ☆ Some elements have only one oxidation state
    - ★ group 1 metals always form 1+ ions and always have a +1 oxidation state
    - ★ group 2 metals always form 2+ ions and always have a +2 oxidation state
  - ☆ Some elements usually have a particular oxidation state
    - ★ oxygen has a -2 oxidation state except in peroxides where it is -1 and in compounds with fluorine (OF_2) where it is +2
    - ★ hydrogen has a +1 oxidation state except in hydrides with group 1 and group 2 metals
  - ☆ the sum of the oxidation numbers
    - ★ in a compound it is always zero
    - ★ in a polyatomic ion it is equal to the charge on the ion

★ Finding oxidation numbers
  - ☆ apply the rules
  - ☆ construct a table if necessary

Sample Problem

Find the oxidation state of the elements in $K_2Cr_2O_7$.

Element	K	Cr	O	TOTAL
Subscript	2	2	7	
Oxidation state	+1	?	-2	
Sum of oxidation states	+2	??	-14	0

[a] potassium is a group one metal; its oxidation state is always +1

[b] oxygen usually has an oxidation state of -2

[c] the sum of oxidation states of each element is the product of the subscript and the oxidation state

[d] find the –sum of the oxidation states of chromium (??) by setting the sum of all the oxidation states to zero

(+2) + ?? + (-14) = 0

?? = +12

[f] find the oxidation state of chromium (?) by dividing the sum (+12) by the subscript (2)

+12 ÷ 2 = +6

Answer the questions below by circling the number of the correct response

1. In this reaction, the oxidation number (oxidation state) of C changes from: $2CO_2 \rightarrow 2CO + O_2$
(1) 0 to +4 (2) +2 to +4 (3) +3 to 0 (4) +4 to +2

2. In the reaction:
$2KMnO_4 + 3H_2SO_4 + 5H_2S \rightarrow 5S + 2MnSO_4 + K_2SO_4 + 8H_2O$
the oxidation number of sulfur changes from
(1) +5 to -5 (2) -5 to +5 (3) 0 to -2 (4) -2 to 0

3. What is the oxidation number of Cr in Na_2CrO_4?
(1) +1 (2) +2 (3) +3 (4) +6

4. What is the oxidation state of the chromium in $K_2Cr_2O_7$?
(1) +5 (2) +6 (3) +3 (4) +12

5. In the reaction $Pb + 2Ag^+ \rightarrow Pb^{+2} + 2Ag$, the reducing agent is
(1) Ag (2) Ag^+ (3) Pb (4) Pb^{+2}

6. Which is not an oxidation-reduction reaction?
(1) $4Na + O_2 \rightarrow 2Na_2O$
(2) $Fe + 2HCl \rightarrow FeCl_2 + H_2$
(3) $CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$
(4) $2H_2O \rightarrow 2H_2 + O_2$

7. Given: $2Al + 3Zn^{+2} \rightarrow 2Al^{+3} + 3Zn$. In this reaction, the oxidizing agent is (1) Al (2) Al^{+3} (3) Zn (4) Zn^{+2}

8. Given: $2Al + 3Zn^{+2} \rightarrow 2Al^{+3} + 3Zn$ In this reaction, electrons are transferred from (1) Al to Al^{+3} (2) Zn^{+2} to Zn (3) Al to Zn^{+2}
(4) Zn^{+2} to Al

9. What is the oxidation number of nitrogen in N_2O_3? (1) +1 (2) +2 (3) +3 (4) +6

10. In the reaction $3CO + Fe_2O_3 \rightarrow 3CO_2 + 2Fe$, the oxidation number of the iron changes from (1) +2 to 0 (2) +2 to +3
(3) +3 to +2 (4) +3 to 0

11. What is the oxidation number of Br in BrO_3^{-2}?
(1) +1 (2) +6 (3) +5 (4) +4

12. Which is the reducing agent in the following reaction?
$Cl_2(aq) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$
(1) Cl_2 (2) H_2O (3) K^+ (4) Br^-

13. What is the oxidation number of carbon in $C_2O_4^{-2}$?
(1) +1 (2) +2 (3) +3 (4) +4

14. Which is an oxidation-reduction reaction?
(1) $CaCO_3 \rightarrow CaO + CO_2$
(2) $KOH + HBr \rightarrow KBr + H_2O$
(3) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
(4) $Mg + Cl_2 \rightarrow MgCl_2$

15. $MnSO_4$ is a product in a reaction that contained $KMnO_4$ as a reactant. The oxidation number of the manganese changed from
(1) -2 to +5 (2) +7 to +2 (3) +5 to -2 (4) -7 to +2

16. Given the balanced equation:
$2HNO_3 + 3H_2S \rightarrow 4H_2O + 2NO + 3S$
Which is reduced? (1) S (2) S-2 (3) N+2 (4) N+5

17. During the reaction $Ca + H_2 \rightarrow CaH_2$, the oxidation number of the hydrogen changes from
(1) 0 to +1 (2) +1 to 0 (3) 0 to -1 (4) -1 to 0

18. In the reaction $Sn^{+4} + H_2(g) \rightarrow Sn^{+2} + 2H^+$, the reducing agent is
(1) Sn^{+4} (2) H_2 (3) Sn^{+2} (4) H^+

19. Given: $3Ag + 4HNO_3 \rightarrow NO + 3AgNO_3 + 2H_2O$. The reducing agent in this reaction is
(1) Ag (2) Ag^{+1} (3) H^{+1} (4) N^{+2}

20. The reaction $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ is an example of
(1) an oxidation reaction, only
(2) a reduction reaction, only
(3) both an oxidation and a reduction reaction
(4) neither an oxidation nor a reduction reaction

21. The oxidation number of manganese in $KMnO_4$ is
(1) +1 (2) +7 (3) +3 (4) +4

22. In the reaction $Sn^{+2} + 2Fe^{+3} \rightarrow Sn^{+4} + 2Fe^{+2}$, the reducing agent is
(1) Fe^{+2} (2) Fe^{+3} (3) Sn^{+2} (4) Sn

23. An oxidizing agent will always
(1) lose electrons
(2) increase in oxidation number
(3) be reduced
(4) increase in mass

NAME:______________________________ DATE:______SECTION____ACTIVITY

Hey Baby, Can I Get Your Number!

Background:

Because electrons are not shown in chemical equations, we must do some work to determine whether a reaction involves oxidation and reduction. The concept of oxidation numbers was devised as a simple way of keeping track of electrons in reactions. The oxidation number is a hypothetical charge assigned to the atom using a set of rules.

Rules:

1) For an atom in its elemental form, the oxidation number is always zero.

2) For any monoatomic ion, the oxidation number equals the charge on the ion.

3) The oxidation number of oxygen is usually –2. An exception is a peroxide O_2^{-2}, where its oxidation number is –1.

4) The oxidation number of hydrogen is +1 when bonded to nonmetals and –1 when bonded to metals.

5) The oxidation number of fluorine is always –1.

6) The algebraic sum of the oxidation numbers in the formula of a compound is zero.

7) The algebraic sum of the oxidation numbers in the formula for a polyatomic ion is equal to the charge on that ion.

8) All alkali metals (group 1) have an oxidation number of +1.

9) All alkaline earth metals (group 2) have an oxidation number of +2.

10) In any ionic compound (a metal and a nonmetal) the non metal has the negative charge.

11) The periodic table only shows the most common oxidation states for the elements, but an atom can have others not listed.

Problems: Determine the oxidation number of each element present.

1) NaBr

2) KClO

3) MgO

4) CsH

5) N_2O_5

6) S_8

7) $Li_2Cr_2O_7$

8) $CoCl_4{}^{2-}$

9) $C_2O_4{}^{2-}$

10) $AgNO_3$

11) Na

12) Zn^{2+}

13) H_2SO_4

Use the following balanced equation to answer questions 14 through 16. The coefficients in front of a species have no effect on the oxidation number.

$$Mg\,(s) + 2\,HCl\,(aq) \rightarrow MgCl_2\,(aq) + H_2\,(g)$$

14) In the boxes below the reaction place each species' oxidation number.

15) Which atom was oxidized (had its oxidation number increase)? ___________

16) Which species was reduced (had its oxidation number decrease)? ___________

17) Which species if any was a spectator (no change in oxidation number)? ___________

Reflection:

Describe your method to assign oxidation numbers to an atom. Describe how this is used to show an oxidation and reduction reaction.

NAME:_________________________________ DATE:_______SECTION_____LAB_____

Who Is Leo Ger?

Background:

A redox reaction will have at least one type of atom releasing electrons and another type of atom accepting electrons. How can you most easily tell if a reaction is redox? Label every atom on both the reactant and product side of the equation with its oxidation number. If there is a change in oxidation number from one side of the equation to the other of the same species of atom, it is a redox reaction. Each complete equation must have at least one atom species losing electrons and at least one atom species gaining electrons. The loss and gain of electrons will be reflected in the changes of oxidation number.

Oxidized Species

- **Reducing agent** - the reactant that gives up electrons.
- The reducing agent contains the element that is oxidized (looses electrons).
- If a substance gives up electrons easily, it is said to be a **strong reducing agent**.

Reduced Species

- **Oxidizing agent** - the reactant that gains electrons.
- The oxidizing agent contains the element that is reduced (gains electrons).
- If a substance gains electrons easily, it is said to be a **strong oxidizing agent**.

Procedure:

In the reactions below determine the species that is:

1) $Mg + HCl \rightarrow MgCl_2 + H_2$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

2) $Fe + V_2O_3 \rightarrow Fe_2O_3 + VO$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

3) $Na + Ag^+ \rightarrow Na^+ + Ag$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

4) $KClO_3 \rightarrow KCl + O_2$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

5) $Ca + Zn^{2+} \rightarrow Ca^{2+} + Zn$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

6) $Fe_2O_3 + CO \rightarrow Fe + CO_2$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

7) $HCl + KOH \rightarrow H_2O + KCl$

- Oxidized Species
- Reduced Species
- Oxidizing Agent
- Reducing Agent

Base your answer to questions 8 through 11 on the following redox reaction.

$$\_\_\_Li + \_\_\_Al^{3+} \rightarrow \_\_\_Li^{+} + \_\_\_Al$$

8) Determine which species is oxidized and reduced.

Oxidized Species:____________ Reduced Species:____________

9) Describe the location of the two species in the above equation can never be the oxidized and or reduced species.

10) Based on the charges of each species determine the total number of electrons transferred between the oxidized and reduced species.

Number of Electrons:______________

11) Based on the law of conservation of charge balance the above equation. How does the balanced equation relate to the number of electrons transferred?

Reflection: Describe how you can identify the oxidized and reduced species in a redox reaction.

REDOX REACTIONS

Name ____________________

For the equations below, identify the substance oxidized, the substance reduced, the oxidizing agent, the reducing agent, and write the oxidation and reduction half reactions.

Example:

	oxidized		reduced		
	Mg	+	Br_2	$\rightarrow$	$MgBr_2$
	reducing agent		oxidizing agent		

oxidation half reaction: $Mg^{\circ} \rightarrow Mg^{+2} + 2e^-$

reduction half reaction: $2e^- + Br_2^{\circ} \rightarrow 2Br^-$

1. $2H_2 + O_2 \rightarrow 2H_20$

2. $Fe + Zn^{2+} \rightarrow Fe^{2+} + Zn$

3. $2Al + 3Fe^{+2} \rightarrow 2Al^{+3} + 3Fe$

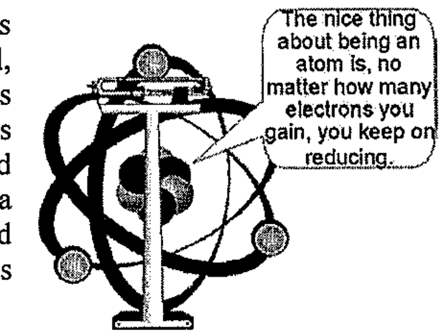
4. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

Chemistry: Form WS10.1.2A Name ____________

REDOX AND ELECTROCHEMISTRY Date ________ Period ____

Analyzing Oxidation-Reduction Reactions

When chemical bonds form, electrons are either lost, gained or shared. Metals lose electrons. This is what happens when iron rusts. When the iron, a metal, combines with oxygen, a non metal, to form rust, it loses electrons. This process is called oxidation even when the nonmetal is not oxygen. Nonmetals gain electrons causing their oxidation states to go down. This is called reduction. It is possible to tell what was oxidized and what was reduced in a chemical reaction by checking the oxidation states of the elements before and after the reaction. The element that has an increase in oxidation state was oxidized while the one that has a decrease in oxidation state was reduced.



Example

$$2FeCl_2 + Cl_2 \rightarrow 2FeCl_3$$

$Fe^{+2} \rightarrow Fe^{+3}$ Iron was oxidized

$Cl^{0} \rightarrow Cl^{-1}$ Chlorine was reduced

For each of the examples below, determine the oxidation states of the elements on both sides of the equation. Then determine which element was oxidized and which was reduced. Write your answer in the space provided.

Reaction	**Element:**	
	Oxidized	**Reduced**
Example: $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$ $\overset{0}{Cu} + 2\overset{+1}{Ag}\overset{+5}{N}\overset{-2}{O_3} \rightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O_3})_2 + 2\overset{0}{Ag}$	Cu	Ag
1. $2Mg + O_2 \rightarrow 2MgO$		
2. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$		
3. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$		
4. $2K_2Cr_2O_7 + 2H_2O + 3S \rightarrow 4KOH + 2Cr_2O_3 + 3SO_2$		

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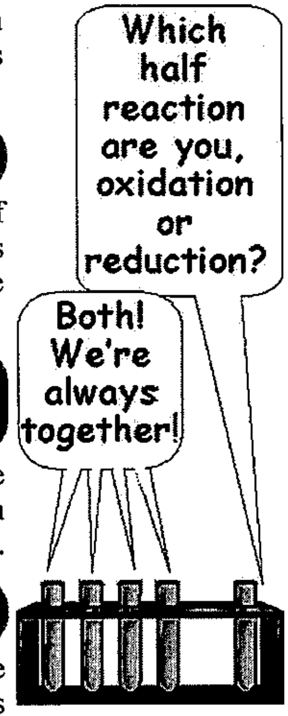
Reaction	Element:	
	Oxidized	Reduced
5. $2H_2O + O_2 \rightarrow 2H_2O_2$		
6. $2KClO_3 \rightarrow 2KCl + 3O_2$		
7. $4NaOH + Ca(OH)_2 + C + 4ClO_2 \rightarrow 4NaClO_2 + CaCO_3 + 3H_2O$		
8. $3P + 5HNO_3 + 2H_2O \rightarrow 5NO + 3H_3PO_4$		
9. $3Cu + 8HNO_3 \rightarrow 2NO + 3Cu(NO_3)_2 + 4H_2O$		
10. $2PbSO_4 + 2H_2O \rightarrow PbO_2 + Pb + 2H_2SO_4$		
11. $4HCl + MnO_2 \rightarrow MnCl_2 + 2H_2O + Cl_2$		
12. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$		
13. $16HCl + 2KMnO_4 \rightarrow 8H_2O + 2KCl + 2MnCl_2 + 5Cl_2$		
14. $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$		
15. $8HNO_3 + 6KI \rightarrow 6KNO_3 + 3I_2 + 2NO + 4H_2O$		
16. $I_2 + 5HClO + H_2O \rightarrow 2HIO_3 + 5HCl$		
17. $K_2Cr_2O_7 + 3SnCl_2 + 14HCl \rightarrow 2CrCl_3 + 3SnCl_4 + 2KCl + 7H_2O$		
18. $SnCl_2 + 2HgCl_2 \rightarrow SnCl_4 + Hg_2Cl_2$		

Chemistry: Form WS10.1.3A Name ______________________

REDOX AND ELECTROCHEMISTRY Date ______________ Period ____

Writing Half Reactions



During a redox reaction electrons are both lost and gained. The metal loses and the non metal gains. An equation showing either the gain or the loss of electrons but not both is called a half reaction. Consider the reaction below:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Magnesium loses electrons while oxygen gains. The reaction can be split into two half reactions showing each. The oxidation half reaction shows the loss of electrons. Electrons are shown on the product side of the equation. The reduction half reaction shows the electron gain. Electrons are shown on the reactant side of the equation.

Oxidation Half:			$2[Mg^0$	$\rightarrow$	Mg^{2+}	+	$2e^-]$
Reduction Half:	$2[O^0$	+	$2e^-$	$\rightarrow$	$O^{2-}]$		

The net equation, the redox reaction, is a combination of the half reactions such that the number of electrons lost equals the number of electrons gained. The electrons are not shown in the net equation because the electrons that were lost are the same ones that were gained.

Net Reaction: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

To write the half reactions, it is first necessary to determine the oxidation states of the elements on both sides of the equation so you know what was oxidized and what was reduced. Then write the oxidation and reduction halves as shown above, making sure the equation is balanced so the number of electrons lost equals the number gained.

Write the half reactions for each of the redox reactions below:

1. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO_2 + H_2O$

2. $CdS + I_2 + HCl \rightarrow CdCl_2 + HI + S$

3. $NaClO + H_2S \rightarrow NaCl + H_2SO_4$

☞ Go on to the next page.

4. $Sn + HNO_3 + H_2O \rightarrow H_2SnO_3 + NO$

5. $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + H_2O + Cl_2$

6. $Fe(OH)_2 + H_2O_2 \rightarrow Fe(OH)_3$

7. $Na + H_2O \rightarrow NaOH + H_2$

8. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO_2 + H_2O$

9. $H_2O_2 \rightarrow H_2O + O_2$

10. $K_2Cr_2O_7 + H_2O + S \rightarrow SO_2 + KOH + Cr_2O_3$

Name ______________________ Class______ Date______________

Activity 8-3

Balancing Redox Equations by Oxidation Numbers

Redox equations of the types studied in Activity 8-2 can easily be balanced by inspection. Equations for more complicated redox reactions are often not easy to balance by inspection. However, the total increase of oxidation numbers equals the total decrease of oxidation numbers in a correctly balanced equation. This fact is the basis for a method of balancing redox equations.

Sample Problem 1 Balance the following equation for a redox reaction between nitric acid and iodine:

$$HNO_3 + I_2 \rightarrow HIO_3 + NO_2 + H_2O$$

Solution There are five steps, *a-e* below, in the solution.

a. Assign oxidation numbers to each atom in the equation, and determine which atoms are changing oxidation numbers.

$$\overset{1+\;5+\;2-}{HNO_3} + \overset{0}{I_2} \rightarrow \overset{1+\;5+\;2-}{HIO_3} + \overset{4+\;2-}{NO_2} + \overset{1+\;2-}{H_2O}$$

We see that the atoms changing oxidation numbers are nitrogen (N) and iodine (I).

b. Indicate the changes of oxidation number below the equation.

$$\underset{5+}{HNO_3} + \underset{0}{I_2} \rightarrow \underset{5+}{HIO_3} + \underset{4+}{NO_2} + H_2O$$

decrease of 1

increase of 5

c. Multiply the increase and decrease by the subscripts of the atoms undergoing changes in oxidation number. Only those subscripts—and, therefore, multipliers—that are greater than 1 need be written.

$$\underset{5+}{HNO_3} + \underset{0}{I_2} \rightarrow \underset{5+}{HIO_3} + \underset{4+}{NO_2} + H_2O$$

decrease of 1

(increase of 5) $\times 2 = 10$

d. Pick coefficients that make the total increase and total decrease of oxidation number the same. If 10 atoms of N decrease by 1, the total decrease will be 10, the same as the total increase already determined for I. Therefore, we use 10 as the coefficient for both HNO_3 and NO_2. And we use 2 for HIO_3 since there are 2 atoms of I in I_2 that must be balanced.

$$10HNO_3 + I_2 \rightarrow 2HIO_3 + 10NO_2 + H_2O$$

e. By inspection, adjust the coefficients for the rest of the substances, leaving oxygen for the last. Here, we see that since we have 10 hydrogen atoms on the left side, we need $5H_2O$ to make 10 hydrogen atoms on the right side. Finally, check the result by finding the total number of oxygen atoms on each side of the equation. If the total on the left equals the total on the right, the equation is correctly balanced. Thus, the count of oxygen atoms serves as a check on the balancing of the equation.

$$\begin{array}{lllllllll} & 10HNO_3 & + & I_2 & \rightarrow & 2HIO_3 & + & 10NO_2 & + & 4H_2O \\ \text{oxygen atoms} & 10\times3=30 & & & & \underbrace{2\times3=6 & + & 10\times2=20 & + & 4}_{\text{total}=30} \end{array}$$

Use steps *a-e* as in Sample Problem 1 to balance the following equations. In the final balanced equation, draw a circle around the formula of the oxidizing agent.

1. ____ $CrCl_3$ + ____ MnO_2 + ____ H_2O → ____ $MnCl_2$ + ____ H_2CrO_4

2. ____ H_2S + ____ HNO_3 → ____ NO_2 + ____ S + ____ H_2O

3. ____ HNO_3 + ____ S → ____ NO + ____ H_2SO_4

4. ____ $KMnO_4$ + ____ Na_2SnO_2 + ____ H_2O → ____ MnO_2 + ____ Na_2SnO_3 + ____ KOH

5. ____ PbO_2 + ____ H_2MnO_3 + ____ HNO_3 → ____ $Pb(NO_3)_2$ + ____ $HMnO_4$ + ____ H_2O

6. ____ HNO_3 + ____ P → ____ H_3PO_4 + ____ NO_2 + ____ H_2O

Sample Problem 2 Balance the following equation for a redox reaction between copper and nitric acid:

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$$

Solution Carry out steps *a-d* as in Sample Problem 1. As the result of step *d*, we have the following:

$$3Cu + 2HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + H_2O \qquad \text{(not balanced yet)}$$

Name ________________ Class______ Date__________

Note that nitrogen (N) appears in *two* formulas on the right side of the equation. In one of these, $Cu(NO_3)_2$, the oxidation number of nitrogen remains unchanged. Therefore, we must increase by 6 the coefficient of HNO_3 on the left side of the equation in order to provide for the 6 atoms of nitrogen in $3Cu(NO_3)_2$ on the right.

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + H_2O$$

Now we carry out step *e* as in Sample Problem 1. Check the result by finding the total number of oxygen atoms on each side of the equation.

$$\begin{array}{llll} & 3Cu + 8HNO_3 \rightarrow & 3Cu(NO_3)_2 \quad + 2NO \quad + 4H_2O \\ \text{oxygen atoms} & 8\times3=24 & \underbrace{3\times3\times2=18 \ + 2 \qquad + 4}_{\text{total} = 24} \end{array}$$

Balance the following equations. Draw a circle around the oxidizing agent in each balanced equation.

7. ____$FeCl_3$ + ____H_2S → ____$FeCl_2$ + ____S + ____HCl

8. ____PbS + ____HNO_3 → ____$Pb(NO_3)_2$ + ____NO + ____S + ____H_2O

9. ____Cu + ____H_2SO_4 → ____$CuSO_4$ + ____SO_2 + ____H_2O

10. ____Cu + ____HNO_3 → ____$Cu(NO_3)_2$ + ____NO_2 + ____H_2O

11. ____Zn + ____H_2SO_4 → ____$ZnSO_4$ + ____SO_2 + ____H_2O

12. ____Cu + ____HNO_3 → ____$Cu(NO_3)_2$ + ____NO + ____H_2O

13. ____$HMnO_4$ + ____HCl → ____$MnCl_2$ + ____Cl_2 + ____H_2O

14. ____Bi_2S_3 + ____HNO_3 → ____$Bi(NO_3)_3$ + ____NO + ____S + ____H_2O

BALANCING REDOX EQUATIONS

Name ____________

Balance the equations below using the half-reaction method.

1.	$Sn^{\circ} + Ag^{+} \rightarrow Sn^{+2} + Ag^{\circ}$
2.	$Cr^{\circ} + Pb^{2+} \rightarrow Cr^{+3} + Pb^{\circ}$
3.	$KClO_3 \rightarrow KCl + O_2$
4.	$NH_3 + O_2 \rightarrow NO + H_2O$
5.	$PbS + H_2O_2 \rightarrow PbSO_4 + H_2O$
6.	$H_2S + HNO_3 \rightarrow S + NO + H_2O$
7.	$MnO_2 + H_2C_2O_4 + H_2SO_4 \rightarrow MnSO_4 + CO_2 + H_2O$
8.	$H_2S + H_2SO_3 \rightarrow S + H_2O$
9.	$KIO_3 + H_2SO_3 \rightarrow KI + H_2SO_4$
10.	$K_2Cr_2O_7 + HCl \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$

Chemistry: Form Ls10.2A Name ______________________

REDOX AND ELECTROCHEMISTRY Date ______________ Period ____

Activity Series

Aim

- compare the activity of metals to hydrogen

Notes

★ More active metals can replace less active metals
★ Metals that are more active than hydrogen can replace hydrogen
★ Hydrogen is used as a standard for comparing the activity of metals
- ☆ Lithium (*MOST ACTIVE*)
- ☆ Rubidium
- ☆ Potassium
- ☆ Cesium
- ☆ Barium
- ☆ Strontium
- ☆ Calcium
- ☆ Sodium
- ☆ Magnesium
- ☆ Aluminum
- ☆ Titanium
- ☆ Manganese
- ☆ Zinc
- ☆ Chromium
- ☆ Iron
- ☆ Nickel
- ☆ Tin
- ☆ Lead
- ☆ **HYDROGEN**
- ☆ Copper
- ☆ Mercury
- ☆ Silver
- ☆ Platinum
- ☆ Gold (*LEAST ACTIVE*)

★ Acids release hydrogen when they react with active metals
★ Active metals corrode easily
- ☆ Definition: CORROSION — loss of metallic properties due to action of air, water, and chemicals
- ☆ Examples
  - ✯ Rust: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
  - ✯ Action of Acids: $2Fe + 6HCl \rightarrow 2FeCl_3 + 3H_2$

★ Spontaneous reactions - replacement of a less active metal by a more active metal occurs spontaneously

Answer the questions below by circling the number of the correct response

1. Which element is used as a standard for comparing the activity of metals? (1) gold (2) iron (3) francium (4) hydrogen

2. In which of the following pairs of metals is the more active metal listed first? (1) iron/sodium (2) copper/tin (3) lithium/platinum (4) zinc/magnesium

3. Based on the activity series, which of the following reactions is likely to occur? (1) $2Fe + 6HCl \rightarrow 2FeCl_3 + 3H_2$ (2) $MgSO_4 + Zn \rightarrow ZnSO_4 + Mg$ (3) $3BaCl_2 + 2Al \rightarrow 2AlCl_3 + 3Ba$ (4) $H_2 + 2LiOH \rightarrow 2Li + 2H_2O$

4. Based on the activity series, which of the following metals could ***NOT*** replace any of the others? (1) Calcium (2) Sodium (3) Magnesium (4) Aluminum

5. The standard on which the activity series is based is (1) fluorine, (2) lithium, (3) hydrogen, (4) oxygen.

6. Of the following, which is ***NOT*** a way to prevent corrosion? (1) painting (2) galvanizing (3) electroplating (4) coating with acid.

7. During a single displacement reaction, which of the following is true? (1) The more active metal steals electrons from the less active metal. (2) The more active metal is oxidized. (3) The more active metal is reduced. (4) The less active metal loses electrons.

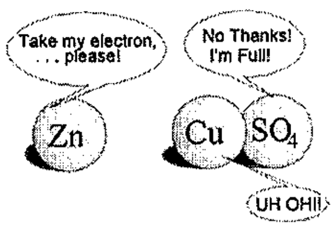
8. During the reaction
   $AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$,
   which ion was reduced? (1) Ag^{+1} (2) Na^{+1} (3) Cl^{-1} (4) none of these
9. Which metal will react with 1.0 M $Pb^{+2}(aq)$ but not with 1.0 M Mg^{+2}? 1. Ba 2. Al 3. Cu 4. Ag

10. If the reaction $X + Zn^{+2} \rightarrow X^{+2} + Zn$ is spontaneous, then X may be 1. Mg 2. Pb 3. Cu 4. Sn

11. Which metal can reduce Pb^{+2}? (1) Cu (2) Hg (3) Fe (4) Ag

12. Which ion can be most easily reduced? 1. Cu^{+2} 2. Zn^{+2} 3. Ni^{+2} 4. Ca^{+2}

Applying the Activity Series



During a single replacement reaction, one element takes the place of another in a compound. Many compounds, such as the copper II sulfate, consist of two parts, a metal (copper) and a nonmetal (sulfate). When a metal such as zinc is dropped into a solution containing copper II sulfate, its natural tendency is to combine with the sulfate by giving electrons to it. The sulfate's outer shell is already full, however, because it has already gained electrons from the copper. As a result, however, the copper has room for zinc's electrons. If zinc can force copper to take its electrons, zinc can become a cation and take copper's place in the compound. Whether or not the zinc can take the copper's place depends upon which metal has the greater tendency to lose electrons. Scientists have determined by experimentation which metals can replace each other in aqueous solution. This resulted in the development of the *Activity Series* as shown in Chart J to the right. The most active metals and nonmetals are shown toward the top of the chart. Elements at the top of the activity series can replace those below them.

Table J
Activity Series**

	Metals	Nonmetals	
Most	Li	F_2	Most
	Rb	Cl_2	
	K	Br_2	
	Cs	I_2	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	**H_2		
	Cu		
	Ag		
Least	Au		Least

**Activity Series based on hydrogen standard

For each example below, if a reaction will occur based on the elements' positions in the *Activity Series*, complete the equation and balance it. If there is no reaction, write no reaction. [*NOTE:* for metals, the format for single replacement reactions is AB + C → CB + A; for nonmetals the format is AB + D → AD + B]

1. $Mg(s) + HCl(aq) \rightarrow$ ______________________
2. $Ag(s) + Cu(NO_3)_2(aq) \rightarrow$ ______________________
3. $Zn(s) + Mn(CH_3COO)_7(aq) \rightarrow$ ______________________
4. $Al(s) + HCl(aq) \rightarrow$ ______________________
5. $Cu(s) + HBr(aq) \rightarrow$ ______________________
6. $Cu(s) + AgCH_3COO(aq) \rightarrow$ ______________________
7. $Sn(s) + H_2SO_4(aq) \rightarrow$ ______________________
8. $Mg(s) + Pb(NO_3)_2(aq) \rightarrow$ ______________________
9. $Pb(s) + AuCl(aq) \rightarrow$ ______________________
10. $Au(s) + LiCl(aq) \rightarrow$ ______________________

Chemistry: Form Ls10.3A Name ______________________

REDOX AND ELECTROCHEMISTRY Date ______________ Period _____

Activity and Electricity

Aim

- describe an electrochemical cell
- describe voltaic cells and electrolytic cells

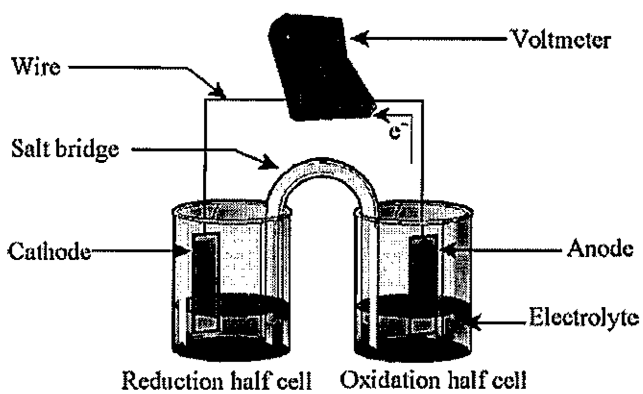
Notes

Electrochemical cells

★ Functioning of the electrochemical cell
- ☆ During a single replacement reaction, more active metals transfer electrons to less active metals
  - ✯ the more active metal is oxidized
  - ✯ the less active metal is reduced
- ☆ If the oxidation and reduction half reactions are physically separated and attached by a wire, electrons will flow through the wire during the reaction

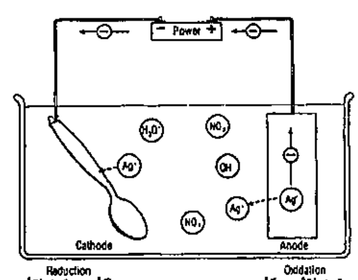
★ Parts of an electrochemical cell
- ☆ electrodes
  - ✯ anode — place where oxidation occurs
  - ✯ cathode — place where reduction occurs
- ☆ half cells — separate containers in which oxidation and reduction half reactions occur

The Electrode Zoo	
AN OX	– ANode = OXidation
RED CAT	– CAThode = REDuction



- ☆ U-tube or salt bridge — lets ions travel between half cells to complete the circuit

★ Examples of electrochemical cells
- ☆ Voltaic Cells (Spontaneous Reactions)
  - ✯ Definition — a system that uses a chemical reaction to produce electricity
  - ✯ Examples
    - ✬ lead acid storage battery (automobile battery)
    - ✬ dry cell (zinc container anode, carbon center post cathode)
- ☆ Electrolytic cells (Nonspontaneous Reactions)
  - ✯ Definition — a system that uses electricity to cause a chemical reaction
  - ✯ Examples
    - ✬ recharging a car battery: $2PbSO_4 + 2H_2O \rightarrow PbO_2 + Pb + 2H_2SO_4$
    - ✬ electrolysis of molten sodium chloride $2NaCl \rightarrow 2Na^0 + Cl_2^0$
    - ✬ electroplating



Answer the questions below by circling the number of the correct response

1. Which reaction will take place in a 1.0 molar aqueous solution?
   1. $Cu + Ag^+ \rightarrow$
   2. $Ag + Mn^{+2} \rightarrow$
   3. $Co + Zn^{+2} \rightarrow$
   4. $Sn + Fe^{+2} \rightarrow$

2. Which reaction occurs at the positive electrode during the electrolysis of molten sodium chloride?
   1. chloride ions are reduced
   2. sodium ions are reduced
   3. chloride ions are oxidized
   4. sodium ions are oxidized

3. Strips of zinc are placed in solutions of the salts listed below. In which solution will a redox reaction take place?
   1. $Ca(NO_3)_2$
   2. $Mg(NO_3)_2$
   3. $Ni(NO_3)_2$
   4. $Sr(NO_3)_2$

4. When the reaction of a chemical cell reaches equilibrium, the potential difference of the cell
   1. decreases 2. increases 3. remains the same

5. When electroplating with silver, the mass of the positive electrode (1) decreases (2) increases (3) remains the same

6. When electroplating with silver, the mass of the negative electrode (1) decreases (2) increases (3) remains the same

7. Which of the following half cells is used as the standard?
   1. $F_2 + 2e^- = 2F^-$
   2. $Li^+ + e^- = Li(s)$
   3. $2H^+ + 2e^- = H_2$
   4. $Ag^+ + e^- = Ag$

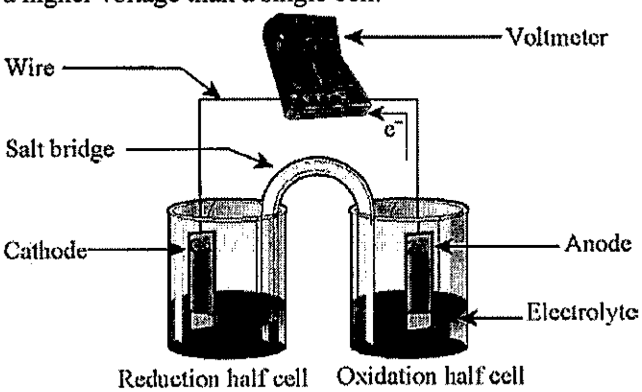
8. Oxygen and copper are produced during the electrolysis of a $CuSO_4$ solution. Which reaction occurs at the negative electrode?
   1. the copper atom is oxidized
   2. the copper ion is reduced
   3. the oxygen atom is oxidized
   4. the oxygen ion is reduced

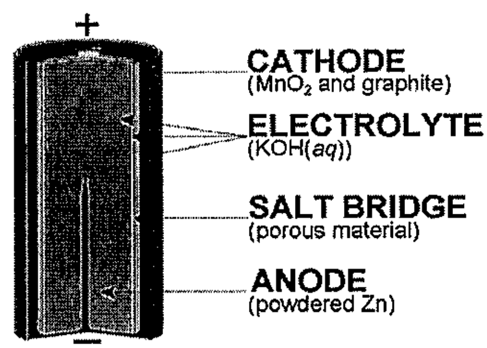
9. Oxidation will occur in the Ni, Ni^{2+}(1 M) half-cell when it forms a cell with
   1. Al, Al^{+3} (1 M)
   2. Au, Au^{+3} (1 M)
   3. Sr, Sr^{+2} (1 M)
   4. Zn, Zn^{+2} (1 M)

10. In the electrolysis of fused $CaCl_2$, the species that reacts at the negative electrode is (1) Ca (2) Ca^{+2} (3) Cl_2 (4) Cl^-

A Salt and Battery

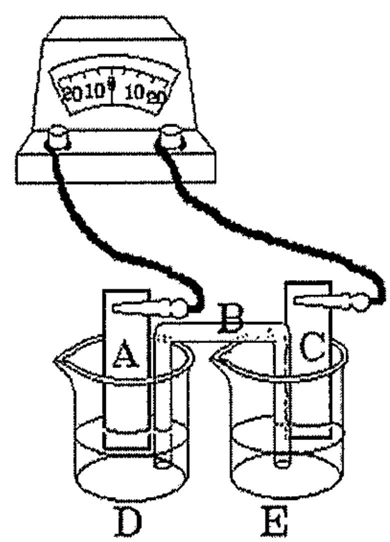
Portable electronic devices run on batteries. The electricity generated by a battery comes from a chemical reaction known as an oxidation-reduction reaction. During an a single replacement, a type of oxidation-reduction reaction, more active metals transfer electrons to less active metals. As a result, the more active metal is oxidized, and the less active metal is reduced. If the oxidation and reduction half reactions are physically separated and attached by a wire, electrons will flow through the wire during the reaction and can be used to power our portable electronics. This is done by putting electrolytes, usually aqueous acids, bases, or salts, into separate containers. The separate containers are called half cells because the half reactions are isolated in them. They are connected by a salt bridge which lets ions travel between half cells. Electrodes are immersed into the electrolytes. The electrodes are merely metals with differing activity. Completing the circuit by connecting the electrodes enables electrons to flow from the more active metal to the less active metal, reducing it. The electrode where reduction occurs is called the **cathode**. The electrode where oxidation occurs is called the **anode**. The device that produces electric current from a chemical reaction is called a **voltaic cell**. Several voltaic cells attached together form a battery of cells. A **battery**, produces a higher voltage than a single cell.





Answer the questions below based on your reading above and on your knowledge of chemistry.

*Answer questions1-4 by referring to the diagram to the right showing an electrochemical cell. The metal at electrode **A** is silver. The metal at electrode **C** is lead. The electrolytes at locations **B**, **D**, and **E** are potassium nitrate, silver nitrate, and lead nitrate respectively.*



1. In what direction do electrons flow in the electrochemical cell pictured to the right (A to C *or* C to A)? ____________________
2. What type of chemical change is taking place in the half-cell contained in the beaker at location *E*? ____________________
3. At which location are electrons being gained? ____________________
4. Which metal is being replaced during the reaction in this electrochemical cell?

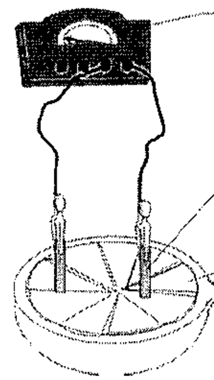
____

Continue ☞

Answer questions 5-16 by referring to Table J. For each of the electrode pairs, which would be the anode in an electrochemical cell?

5. Cu/Zn .... ________	9. Au/Pb .... ________	13. Co/Ni .... ________
6. Pb/Sn .... ________	10. Mn/Zn .... ________	14. H_2/Ag .... ________
7. K/Al .... ________	11. Fe/Zn .... ________	15. Cu/Mg .... ________
8. Ba/Li .... ________	12. Co/Ca .... ________	16. Zn/Al .... ________

Answer questions17-19 by referring to t he setup shown to the right using a lemon and metal strips. It actually produces measurable electricity.



17. Explain how the lemon battery works? ________________________________

____

____

____

____

____

18. What parts of a typical voltaic cell are missing in the lemon battery? What effect does this have on how well it functions?

Explain. ____

____

____

19. If the metal strip on the right is iron and the metal strip on the left is aluminum, in what direction will electricity flow?

____

20. What happens at the anode of an electrochemical cell? ________________________________

____

21. There are two voltaic cells pictured on the previous page. The one on the left is called a wet cell, while the one at the left is called a dry cell. The one at the right is also called an alkaline cell. What is the difference between these cells that

accounts for the difference in the way they are named? ________________________________

____

____

STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			$E°$(V)
$F_2(g) + 2e^-$	$\rightarrow$	$2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow$	Co^{2+}	1.82
$Au^{3+} + 3e^-$	$\rightarrow$	$Au(s)$	1.50
$Cl_2(g) + 2e^-$	$\rightarrow$	$2Cl^-$	1.36
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow$	$2H_2O(l)$	1.23
$Br_2(l) + 2e^-$	$\rightarrow$	$2Br^-$	1.07
$2Hg^{2+} + 2e^-$	$\rightarrow$	$Hg_2{}^{2+}$	0.92
$Hg^{2+} + 2e^-$	$\rightarrow$	$Hg(l)$	0.85
$Ag^+ + e^-$	$\rightarrow$	$Ag(s)$	0.80
$Hg_2{}^{2+} + 2e^-$	$\rightarrow$	$2Hg(l)$	0.79
$Fe^{3+} + e^-$	$\rightarrow$	Fe^{2+}	0.77
$I_2(s) + 2e^-$	$\rightarrow$	$2I^-$	0.53
$Cu^+ + e^-$	$\rightarrow$	$Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow$	$Cu(s)$	0.34
$Cu^{2+} + e^-$	$\rightarrow$	Cu^+	0.15
$Sn^{4+} + 2e^-$	$\rightarrow$	Sn^{2+}	0.15
$S(s) + 2H^+ + 2e^-$	$\rightarrow$	$H_2S(g)$	0.14
$2H^+ + 2e^-$	$\rightarrow$	$H_2(g)$	0.00
$Pb^{2+} + 2e^-$	$\rightarrow$	$Pb(s)$	−0.13
$Sn^{2+} + 2e^-$	$\rightarrow$	$Sn(s)$	−0.14
$Ni^{2+} + 2e^-$	$\rightarrow$	$Ni(s)$	−0.25
$Co^{2+} + 2e^-$	$\rightarrow$	$Co(s)$	−0.28
$Cd^{2+} + 2e^-$	$\rightarrow$	$Cd(s)$	−0.40
$Cr^{3+} + e^-$	$\rightarrow$	Cr^{2+}	−0.41
$Fe^{2+} + 2e^-$	$\rightarrow$	$Fe(s)$	−0.44
$Cr^{3+} + 3e^-$	$\rightarrow$	$Cr(s)$	−0.74
$Zn^{2+} + 2e^-$	$\rightarrow$	$Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	$\rightarrow$	$H_2(g) + 2OH^-$	−0.83
$Mn^{2+} + 2e^-$	$\rightarrow$	$Mn(s)$	−1.18
$Al^{3+} + 3e^-$	$\rightarrow$	$Al(s)$	−1.66
$Be^{2+} + 2e^-$	$\rightarrow$	$Be(s)$	−1.70
$Mg^{2+} + 2e^-$	$\rightarrow$	$Mg(s)$	−2.37
$Na^+ + e^-$	$\rightarrow$	$Na(s)$	−2.71
$Ca^{2+} + 2e^-$	$\rightarrow$	$Ca(s)$	−2.87
$Sr^{2+} + 2e^-$	$\rightarrow$	$Sr(s)$	−2.89
$Ba^{2+} + 2e^-$	$\rightarrow$	$Ba(s)$	−2.90
$Rb^+ + e^-$	$\rightarrow$	$Rb(s)$	−2.92
$K^+ + e^-$	$\rightarrow$	$K(s)$	−2.92
$Cs^+ + e^-$	$\rightarrow$	$Cs(s)$	−2.92
$Li^+ + e^-$	$\rightarrow$	$Li(s)$	−3.05

Name ______________________________

ELECTROCHEMISTRY

Date ________________ Period _____

Determining the Voltage of Electrochemical Cells

Chemical reactions often involve the movement of electrons. The driving force that moves the electrons can be measured. It is the voltage. The voltage of an electrochemical cell can be determined using the *Standard Reduction Table*.

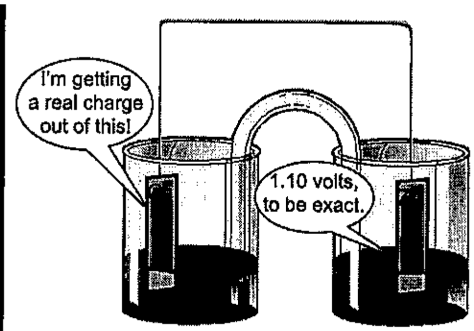
Procedure

1. All half reactions on the *Standard Reduction Potentials Table* are compared to hydrogen ($E^0 = 0$)
2. All half reactions can be read in reverse as oxidations in which case the sign of the voltage, E^0, is changed
3. The net voltage is the sum of the voltages of the oxidation half reactions and the reduction half reactions (see chart)

Example

What voltage is associated with the reaction $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$?

		Zn^0	$\rightarrow$	Zn^{2+}	+	$2e^-$	$E^0 = 0.76v$
Cu^{2+}	+	$2e^-$	$\rightarrow$			Cu^0	$E^0 = 0.34v$
Cu^{2+}	+	Zn^0	$\rightarrow$	Zn^{2+}	+	Cu^0	$E^0 = 1.10v$



Existential discussions in voltaic cells

Write the half reactions for each of the following reactions, balance them, and determine the voltage (E^0) associated with the reaction by using the *Standard Reduction Table.*

1. $Cu + AgNO_3 \rightarrow Ag + Cu(NO_3)_2$

2. $K_2Cr_2O_7 + SnCl_2 + HCl \rightarrow CrCl_3 + SnCl_4 + KCl + H_2O$

3. $SnCl_2 + HgCl_2 \rightarrow SnCl_4 + Hg_2Cl_2$

4. $Sn + HNO_3 + H_2O \rightarrow H_2SnO_3 + NO$

5. $KBr + Fe_2(SO_4)_3 \rightarrow Br_2 + K_2SO_4 + FeSO_4$

6. $Fe + CuSO_4 \rightarrow Cu + Fe_2(SO_4)_3$

7. $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + H_2O + Cl_2$

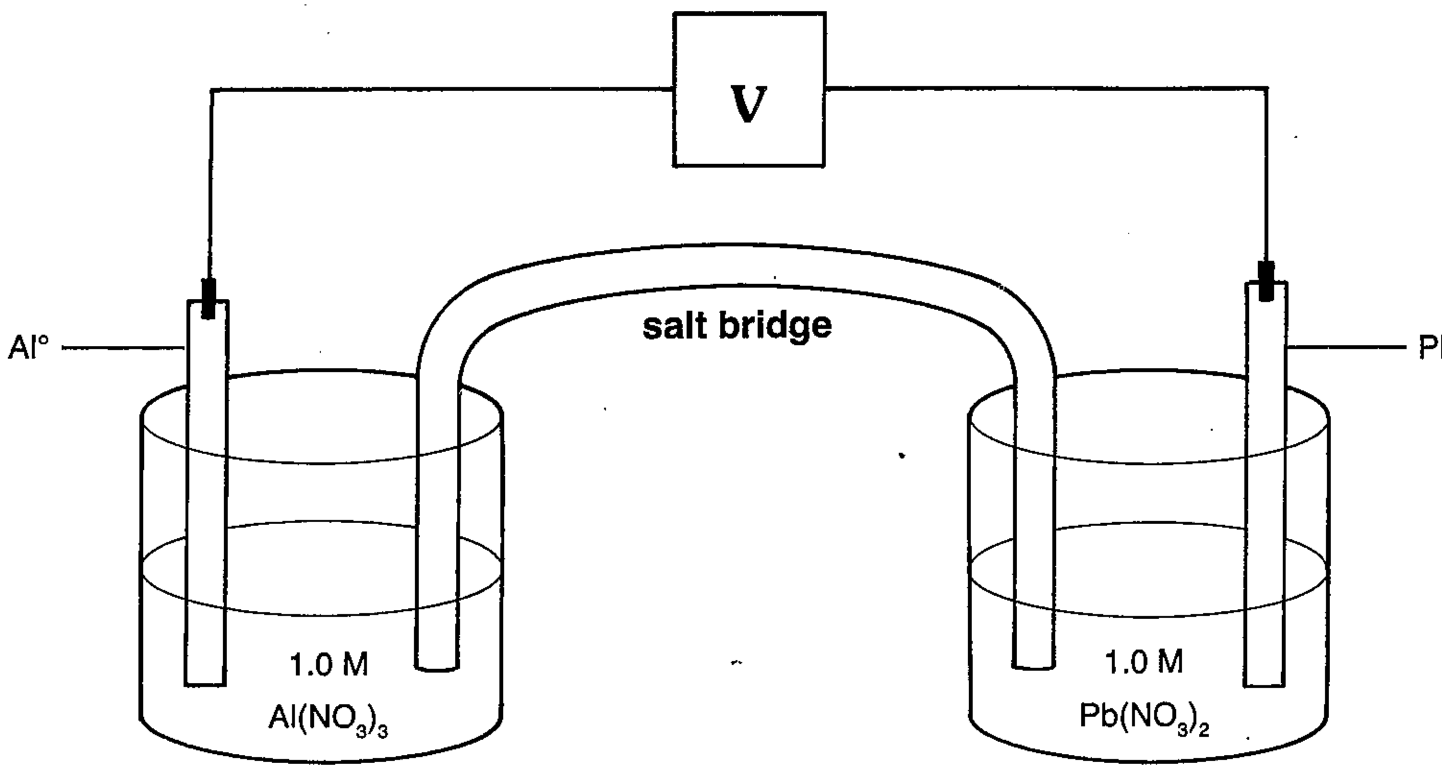
8. $Na + H_2O \rightarrow NaOH + H_2$

9. $HBr + MnO_2 \rightarrow MnBr_2 + H_2O + Br_2$

10. $HCl + K_2SO_4 \rightarrow KCl + SO_2 + H_2O + Cl_2$

THE ELECTROCHEMICAL CELL

Name ____________________



Answer the questions below referring to the above diagram and a Table of Standard Electrode Potentials.

1. Which is more easily oxidized, metal, aluminum or lead? ____________________
2. What is the balanced equation showing the spontaneous reaction that occurs?

   ____
3. What is the maximum voltage that the above cell can produce? ____________________
4. What is the direction of electron flow in the wire? ____________________
5. What is the direction of positive ion flow in the salt bridge? ____________________
6. Which electrode is decreasing in size? ____________________
7. Which electrode is increasing in size? ____________________
8. What is happening to the concentration of aluminum ions? ____________________
9. What is happening to the concentration of lead ions? ____________________
10. What is the voltage in this cell when the reaction reaches equilibrium? ____________________
11. Which is the anode? ____________________
12. Which is the cathode? ____________________
13. What is the positive electrode? ____________________
14. What is the negative electrode? ____________________

Name ________________________________ Class______ Date__________

Activity 8-6
Electrochemical Cells

Introduction

The apparatus for a redox reaction can be designed so that the transfer of electrons from the reducing agent to the oxidizing agent takes place through an external wire circuit rather than by direct contact of the substances. Such an arrangement is called an electrochemical cell, or simply a chemical cell.

1. What names applied to electrochemical cells recognize the contributions of two Italian scientists? __________________ and __________________
2. What four substances are used to make a Daniell cell? ______________________________

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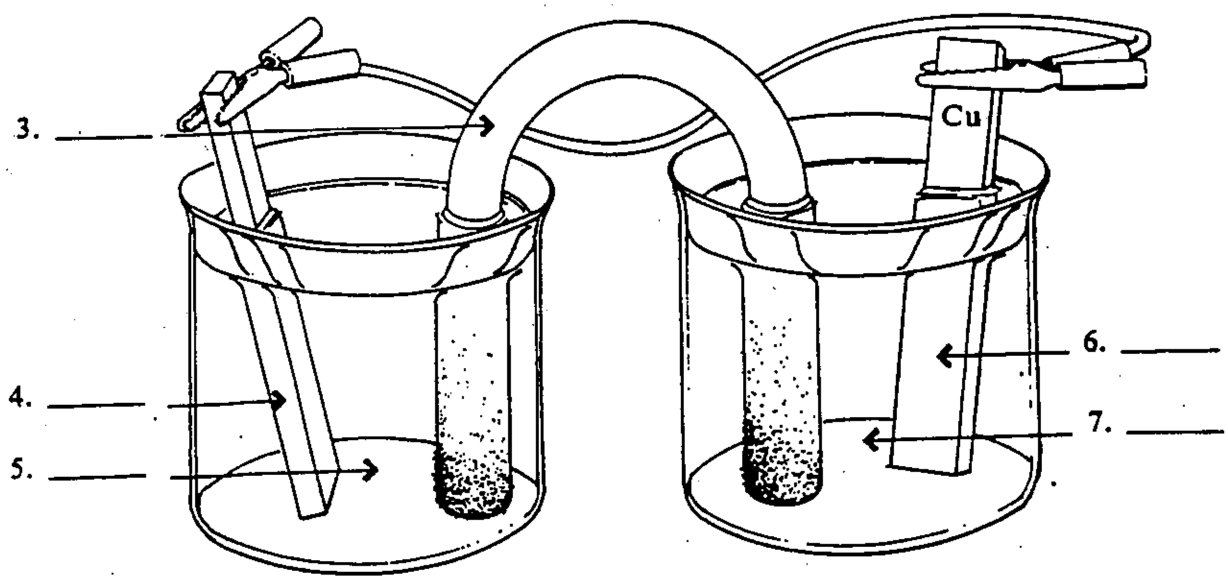
The following diagram shows a Daniell cell. On each numbered line, write the letter of the appropriate label from the list below.

Labels

A. salt bridge
B. anode
C. cathode
D. Cu^{2+}, SO_4^{2-}
E. Zn^{2+}, SO_4^{2-}



8. Which electrode is made of zinc metal? __________________ (*anode/cathode*)
9. Which electrode is made of copper metal? __________________ (*anode/cathode*)
10. On the diagram, draw an arrow that shows the direction of electron flow and label it *e*. Draw an arrow that shows the direction of the negative ion movement and label it *neg. ions.*
11. Which kind of reaction occurs at the anode of an electrochemical cell? __________ (*oxidation/reduction*)
12. Which kind of reaction occurs at the cathode of an electrochemical cell? __________ (*oxidation/reduction*)

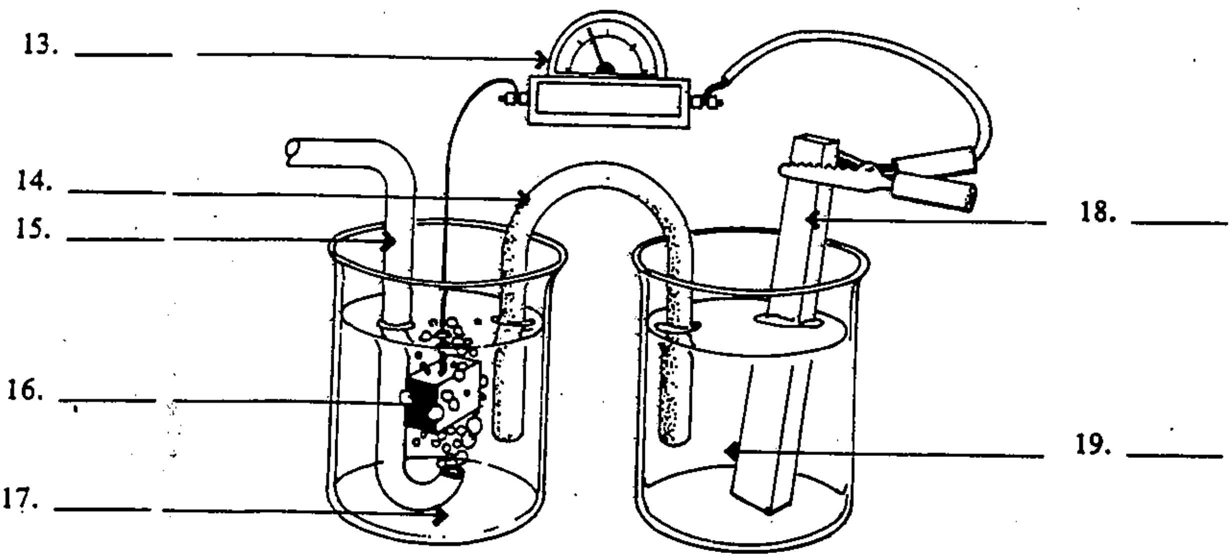
The half-cell

A half-cell (or electrode) consists of a metal strip immersed in a container of an electrolyte. Provision is made for the ions of the electrolyte to move either through the walls of a porous cup or through a salt bridge.

A chemical cell consists of two half-cells connected by an external circuit with provision for movement of ions between the half-cells. The following diagram shows a cell in which a standard hydrogen half-cell is connected to a standard silver half-cell. H_2 gives off e^-'s more readily than Ag°. Therefore, oxidation occurs at the H_2 electrode. On each numbered line of the diagram, write the letter of the appropriate label from the list below.

Labels

A. anode
B. cathode
C. salt bridge
D. source of H_2
E. solution containing H^+
F. solution containing Ag^+, NO_3^-
G. voltmeter



20. Which electrode is made of platinum? ____________ *(anode/cathode)*
21. Which electrode is made of silver? ____________ *(anode/cathode)*
22. On the diagram, draw an arrow that shows the direction of electron flow, and label it *e*.
23. Why is platinum chosen as an electrode in this cell? ____________

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NAME:______________________ DATE:_______SECTION______LAB______

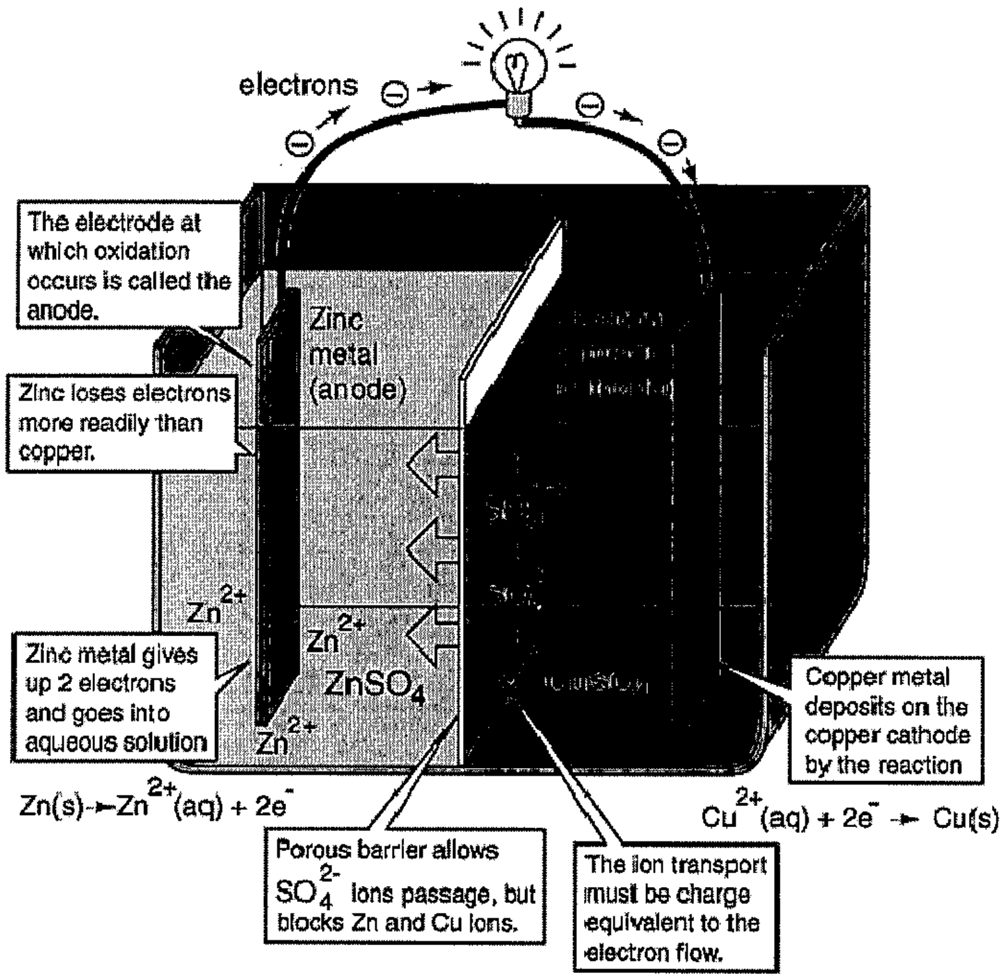
I HAVE AN OX AND A RED CAT

Electrochemical Cells

An extremely important class of oxidation and reduction reactions are used to provide useful electrical energy in batteries. A simple electrochemical cell can be made from copper and zinc metals with solutions of their sulfates. In the process of the reaction, electrons can be transferred from the zinc to the copper through an electrically conducting path as a useful electric current.

An electrochemical cell can be created by placing metallic electrodes into an electrolyte where a chemical reaction either uses or generates an electric current. Electrochemical cells which generate an electric current are called voltaic cells or galvanic cells, and common batteries consist of one or more such cells. In other electrochemical cells an externally supplied electric current is used to drive a chemical reaction which would not occur spontaneously. Such cells are called electrolytic cells.

An electrochemical cell, which causes external electric current flow, can be created using any two different metals since metals differ in their tendency to lose electrons. Zinc more readily loses electrons than copper, so placing zinc and copper metal in solutions of their salts can cause electrons to flow through an external wire which leads from the zinc to the copper.



As a zinc atom provides the electrons, it becomes a positive ion and goes into aqueous solution, decreasing the mass of the zinc electrode. On the copper side, the two electrons received allow it to convert a copper ion from solution into an uncharged copper atom which deposits on the copper electrode, increasing its mass. The two reactions are typically written

$$Zn(s) \rightarrow Zn^{+2}(aq) + 2e^-$$

$$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$$

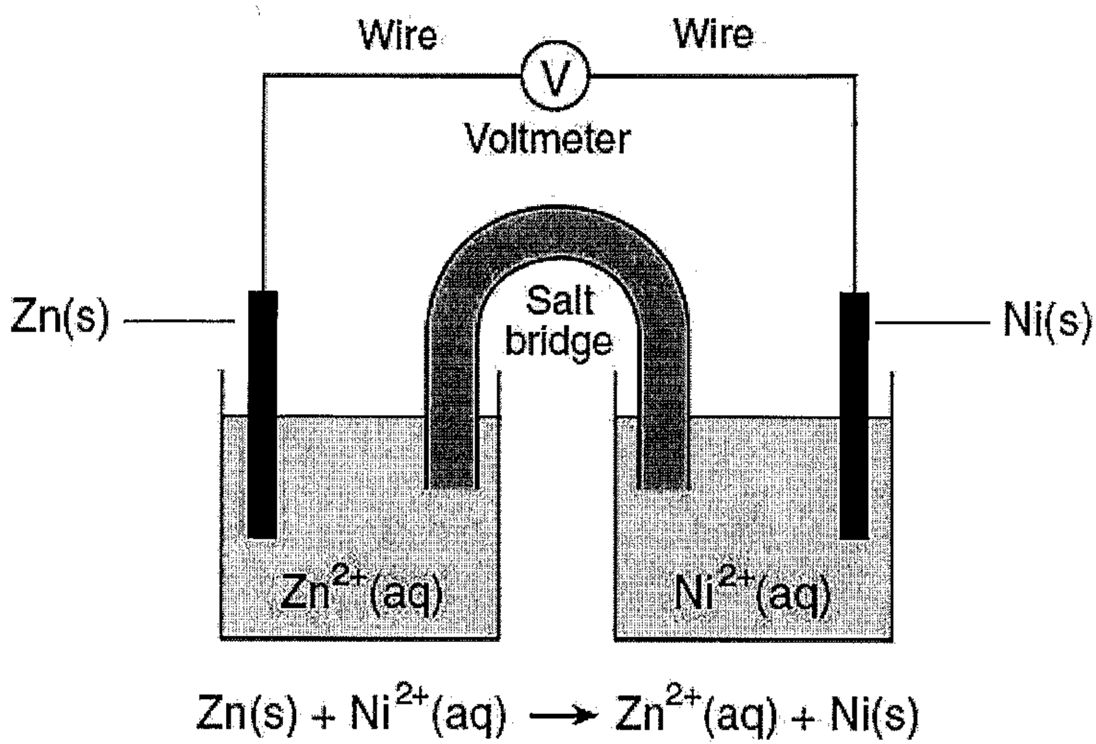
The letters in parentheses are just reminders that the zinc goes from a solid (s) into a water solution (aq) and vice versa for the copper. It is typical in the language of electrochemistry to refer to these two processes as "half-reactions" which occur at the two electrodes.

The zinc "half-reaction" is classified as oxidation since it loses electrons. The terminal at which oxidation occurs is called the "anode". For a battery, this is the negative terminal.

The copper "half-reaction" is classified as reduction since it gains electrons. The terminal at which reduction occurs is called the "cathode". For a battery, this is the positive terminal.

In order for the voltaic cell to continue to produce an external electric current, there must be a movement of the sulfate ions in solution from the right to the left to balance the electron flow in the external circuit. The metal ions themselves must be prevented from moving between the electrodes, so some kind of porous membrane or salt bridge must provide for the selective movement of the negative ions in the electrolyte from the right to the left.

A)



1) Oxidized species:______________ Reduced species:_______________

2) Anode:______________________ Cathode:_______________

3) Uses arrows and e^- on the diagram above to show the direction of electron flow.

4) Write a sentence stating the direction of electron flow. Include the identity of each species.

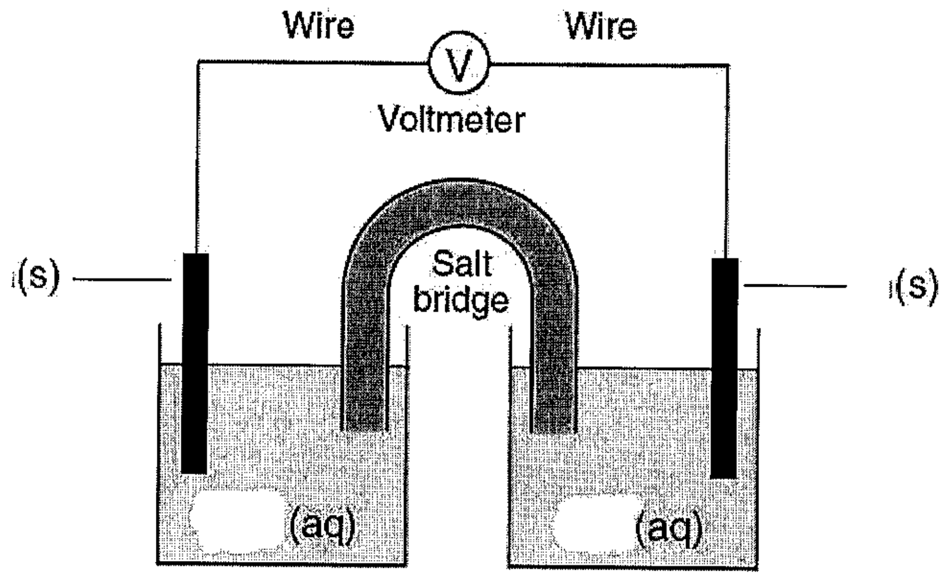
5) Use arrows and the word ions to show the direction of ion flow.

6) Half-Reaction for oxidation:

7) Half-Reaction for reduction:

8) A correctly balanced redox reaction.

B)



$$Sn^{2+}_{(aq)} + Na_{(s)} \rightarrow Sn_{(s)} + Na^{+}_{(aq)}$$

9) Oxidized species:______________ Reduced species:______________

10) Anode:______________________ Cathode:______________

11) Uses arrows and e⁻ on the diagram above to show the direction of electron flow.

12) Write a sentence stating the direction of electron flow. Include the identity of each species.

13) Use arrows and the word ions to show the direction of ion flow.

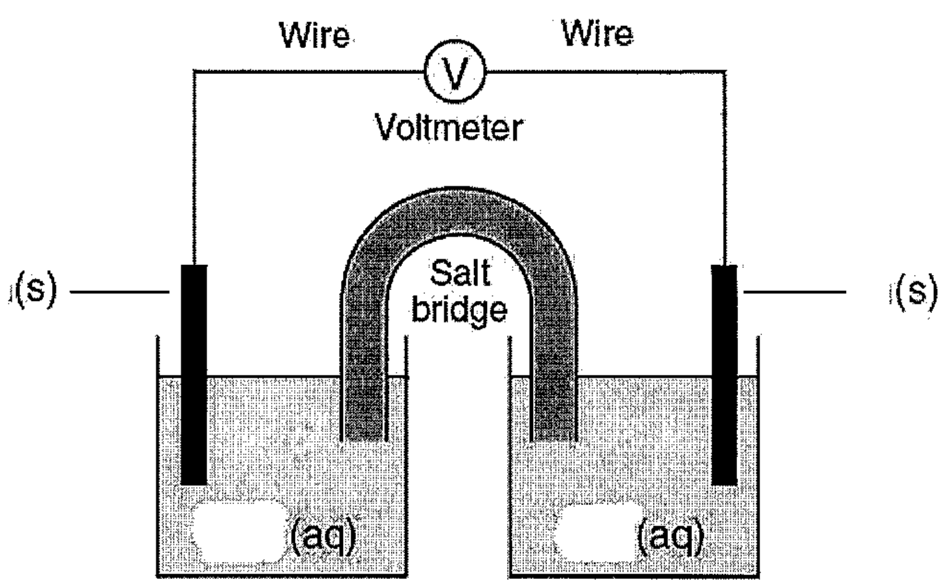
14) Half-Reaction for oxidation:

15) Half-Reaction for reduction:

16) A correctly balanced redox reaction.

17) How does Table J show which species will be oxidized?

C)



$$Sr_{(s)} + Al^{3+}_{(aq)} \rightarrow Sr^{2+}_{(aq)} + Al_{(s)}$$

18) Oxidized species:______________ Reduced species:______________

19) Anode:______________________ Cathode:______________

20) Uses arrows and e⁻ on the diagram above to show the direction of electron flow.

21) Write a sentence stating the direction of electron flow. Include the identity of each species.

22) Use arrows and the word ions to show the direction of ion flow.

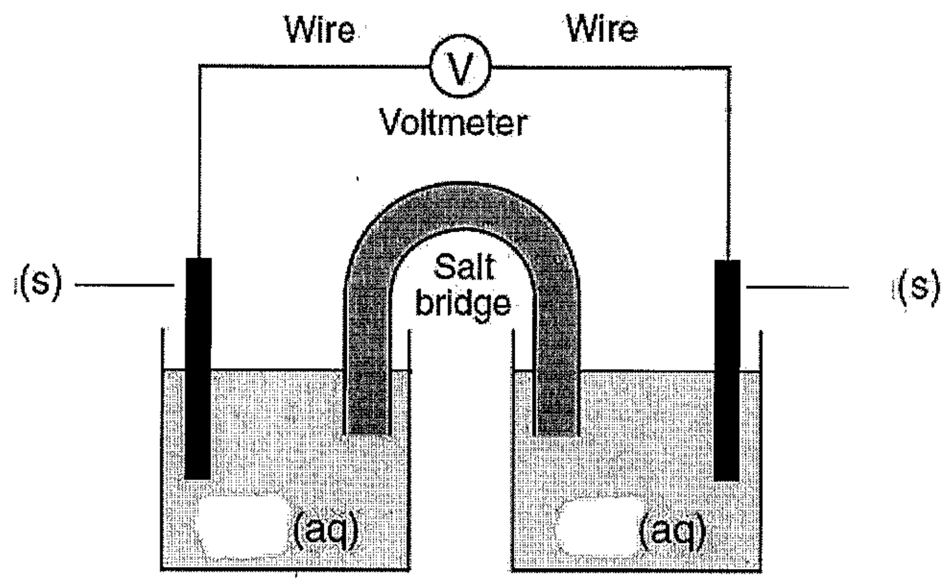
23) Half-Reaction for oxidation:

24) Half-Reaction for reduction:

25) A correctly balanced redox reaction.

26) Determine the total number of electrons transferred in the reaction. __________

D)



27) Oxidized species:____________ Reduced species:_____________

28) Anode:____________________ Cathode:_____________

29) Uses arrows and e⁻ on the diagram above to show the direction of electron flow.

30) Write a sentence stating the direction of electron flow. Include the identity of each species.

31) Use arrows and the word ions to show the direction of ion flow.

32) Half-Reaction for oxidation:

33) Half-Reaction for reduction:

34) A correctly balanced redox reaction. Fill this reaction in the box bellow the cells.

35) Describe why you choose which species will be oxidized and reduced.

Name ______________________ Class______ Date__________

Activity 8-5
Electrolytic Cells II

Electrolysis of a solution of sodium chloride

1. Write the equation for the electrolysis of a solution of sodium chloride. Note that the potential reactants available are $Na^+(aq)$, $Cl^-(aq)$, and H_2O molecules.

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2. Write the equation for the anode half-reaction.

____

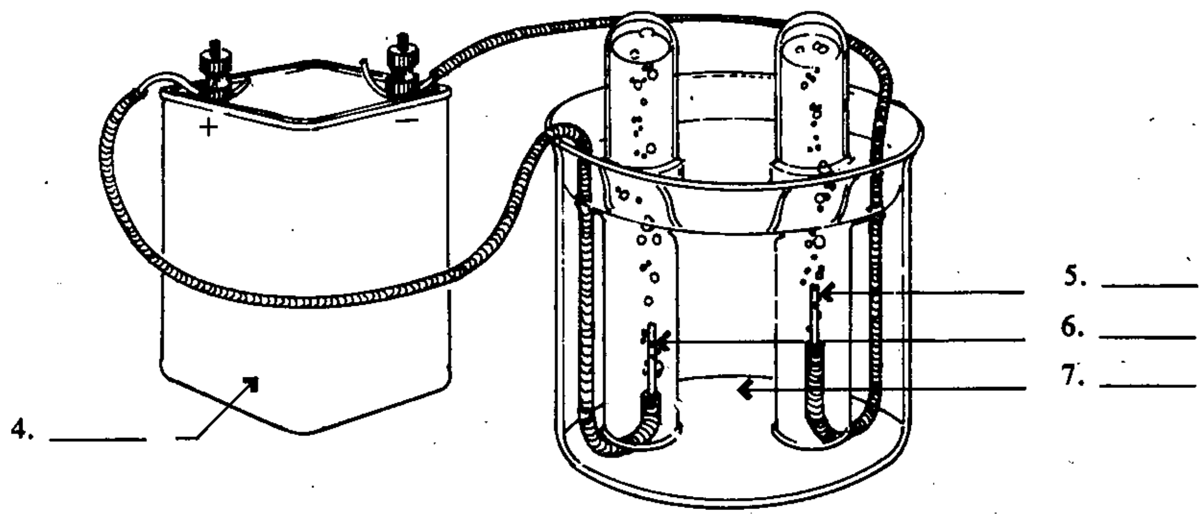
3. Write the equation for the cathode half-reaction.

____

The apparatus in the following diagram can be used to electrolyze a solution of sodium chloride. On each numbered line, write the letter of the appropriate label from the list below.

Labels

A. direct current source B. anode C. cathode D. ions in solution



8. On the diagram, draw an arrow that shows the direction of the electron flow and label it *e*. Draw arrows that show the direction of anion flow and label them Cl^- and OH^-. Draw an arrow that shows the direction of cation flow and label it Na^+.
9. Write the full ionic equation, including the spectator ions, for the overall reaction.

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10. Write the net ionic equation for the overall reaction.

____

As the electrolysis of sodium chloride solution proceeds, how do the quantities listed in the following table change? To complete the table, write:

I—for increases D—for decreases R—for remains the same

Give a reason for each answer.

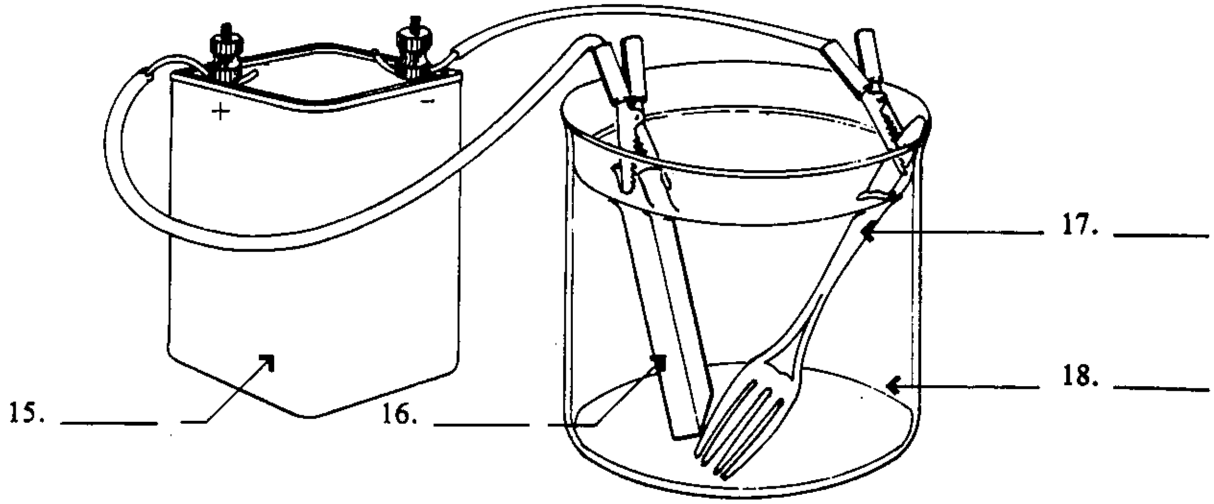
	Quantity	Change	Reason
11.	Na^+		
12.	Cl^-		
13.	OH^-		
14.	pH		

Electroplating

A thin layer of metal can be applied to the surface of another metal by means of electrolysis. This process is called electroplating. The following diagram shows a simplified process for plating silver onto a fork made of a less expensive metal. On each numbered line in the diagram, write the letter of the appropriate label from the list below.

Labels

A. anode B. ions in solution C. cathode D. direct current source



19. On the diagram, draw an arrow that shows the direction of electron flow and label it *e*. Draw an arrow that shows the direction of cation flow and label it Ag^+.
20. Write the equation for the anode half-reaction. ______________________
21. Write the equation for the cathode half-reaction. ______________________

As the electroplating process proceeds, how do the quantities listed in the following table change? To complete the table, write:

I—for increases D—for decreases R—for remains the same

Give a reason for each answer.

	Quantity	Change	Reason
22.	Mass of anode		
23.	Mass of cathode		
24.	$[Ag^+]$		

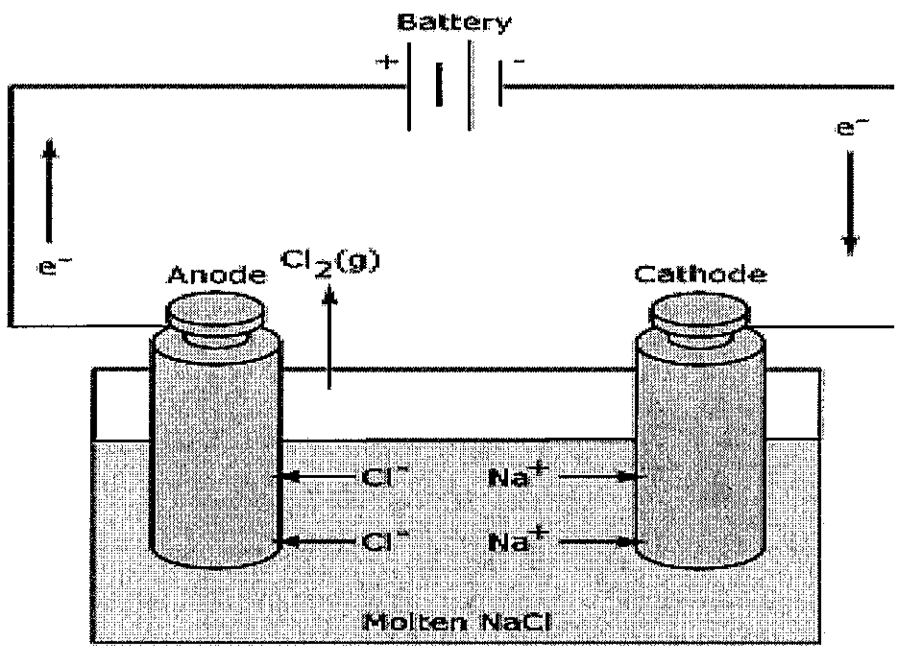
NAME:________________________________ DATE:________SECTION_______LAB_______

Why Are You Forcing Me To Do This?

Electrolytic Cells

The redox reaction in an electrolytic cell is nonspontaneous. Electrical energy is required to induce the electrolysis reaction. An example of an electrolytic cell is shown below, in which molten NaCl is electrolyzed to form liquid sodium and chlorine gas. The sodium ions migrate toward the cathode, where they are reduced to sodium metal. Similarly, chloride ions migrate to the anode and are oxided to form chlorine gas. This type of cell is used to produce sodium and chlorine. The chlorine gas can be collected surrounding the cell. The sodium metal is less dense than the molten salt and is removed as it floats to the top of the reaction container.

$$2Na^+ (\ell) + 2Cl^- (\ell) \rightarrow 2Na\,(s) + Cl_2\,(g)$$



Electrolytic Cell

1) Oxidized species:_____________ Reduced species:______________

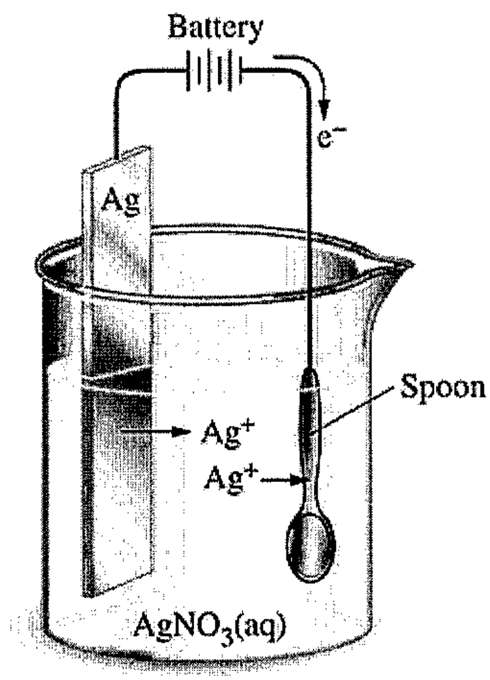
2) Half-Reaction for oxidation:

3) Half-Reaction for reduction:

4) Based on your half reactions discuss why this process is non-spontaneous; why does it need an external power source to drive it.

Electroplating is the coating of an electrically conductive object with a layer of metal using electrical current. The result is a thin, smooth, even coat of metal on the object.

The process used in electroplating is called **electrodeposition** and is analogous to an electrochemical cell acting in reverse. The item to be coated is placed into a container containing a solution of one or more metal salts. The item is connected to an electrical circuit, forming the cathode (negative) of the circuit while an electrode typically of the same metal to be plated forms the anode (positive). When an electrical current is passed through the circuit, metal ions in the solution take up excess electrons at the item. The result is a layer of metal on the item. However, considerable skill and craft-technique is required to ensure an evenly-coated finished product.



5) Based on the diagram above which object is acting as the anode and cathode. Explain your choice.

Anode:________________ Cathode:________________

6) Oxidized species:______________ Reduced species:______________

Reflection:

Explain how an electrolytic cell differs from an electrochemical cell.