Chemistry: Form WS10.1.1A

REDOX AND ELECTROCHEMISTRY

Name ______________________________

Date ________________ Period _____

Finding the Oxidation State

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	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

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 ____________________ Class______ Date__________

Activity 8-2
Oxidation and Reduction

Definitions

Define each of the following terms.

1. Oxidation number. ____________________

2. Oxidation. ____________________

3. Reduction. ____________________

4. Oxidizing agent. ____________________

5. Reducing agent. ____________________

Redox in direct combination (synthesis) reactions

6. What is a direct combination, or synthesis, reaction?
(See Activity 4-6.) ____________________

Questions 7-9. Apply steps *a, b,* and *c* below to each of the reactions in questions 7 to 9.

a. Below each equation, write the oxidation number of each element. As an example, we apply this step to the reaction for producing water from hydrogen and oxygen:

$$\underset{0}{2H_2} + \underset{0}{O_2} \rightarrow \underset{1+\;\;2-}{2H_2O}$$

b. Show the change in oxidation number for one atom of each element in the equation. For example:

$$\underset{0}{2H_2} + \underset{0}{O_2} \rightarrow \underset{1+\;\;2-}{2H_2O}$$

decrease of 2

increase of 1

c. Indicate which element is oxidized and which is reduced. For example, for the reaction above, H_2 is oxidized since its atoms increase in oxidation number, while O_2 is reduced since its atoms decrease in oxidation number.

For questions 7, 8, and 9, carry out steps *a, b,* and *c* above. Then balance each equation by inspection, and write the correct coeffecient in front of each substance of the equation.

7. ______ H_2 + ______ Br_2 → ______ HBr substance oxidized ______________ substance reduced ______________

8. ______ Al + ______ Cl_2 → ______ $AlCl_3$ substance oxidized ______________ substance reduced ______________

9. ______ Mg + ______ N_2 → ______ Mg_3N_2 substance oxidized ______________ substance reduced ______________

Redox in decomposition (analysis) reactions

10. What is a decomposition, or analysis, reaction?
(See Activity 3-10.) ____

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Name ______________________________ Class______ Date____________

Questions 11, 12, and 13. Carry out steps *a, b,* and *c* on page 272 for each of the following decomposition reactions. Balance each equation by inspection, and write the correct coefficient in front of each substance in the equation. Note that when atoms of one element in a substance are oxidized and atoms of another element *in the same substance* are reduced, the substance is said to undergo auto-oxidation. Auto-oxidation takes place in the following reactions.

11. ____ $HI \rightarrow$ ____ H_2 + ____ I_2 substance oxidized ____________

substance reduced ____________

12. ____ $Al_2O_3 \rightarrow$ ____ Al ____ O_2 substance oxidized ____________

substance reduced ____________

13. ____ $2H_2O_2 \rightarrow$ ____ $2H_2O$ + ____ O_2 substance oxidized ____________

substance reduced ____________

Redox in single replacement reactions

14. What is a single replacement reaction?
(See Activity 3-10). ____

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____

Questions 15, 16, 17. Carry out steps *a, b,* and *c* on page 272 for each of the following single replacement reactions. Balance each equation by inspection and write the correct coefficient in front of each substance in the equation.

15. ____ Zn + ____ $HCl \rightarrow$ ____ $ZnCl_2$ + ____ H_2 substance oxidized ____________

substance reduced ____________

16. ____ $FeBr_3$ + ____ $Cl_2 \rightarrow$ ____ $FeCl_3$ + ____ Br_2 substance oxidized ____________

substance reduced ____________

17. ____ Al + ____ $H_2SO_4 \rightarrow$ ____ $Al_2(SO_4)_3$ + ____ H_2

substance oxidized ____________

substance reduced ____________

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: $\underset{\text{reducing agent}}{\overset{\text{oxidized}}{Mg}} + \underset{\text{oxidizing agent}}{\overset{\text{reduced}}{Br_2}} \rightarrow MgBr_2$

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

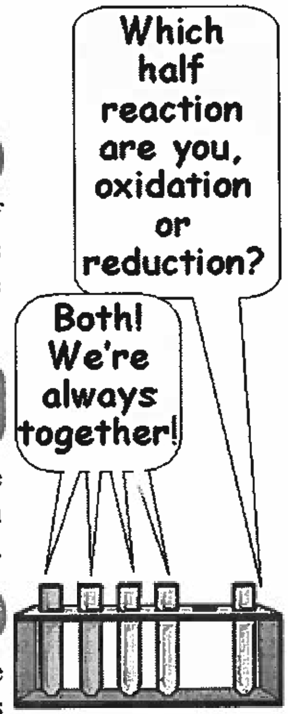
3. $2Al + 3Fe^{+2} \rightarrow 2Al^{+3} + 3Fe$

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

Chemistry: Form WS10.1.3A 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:

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

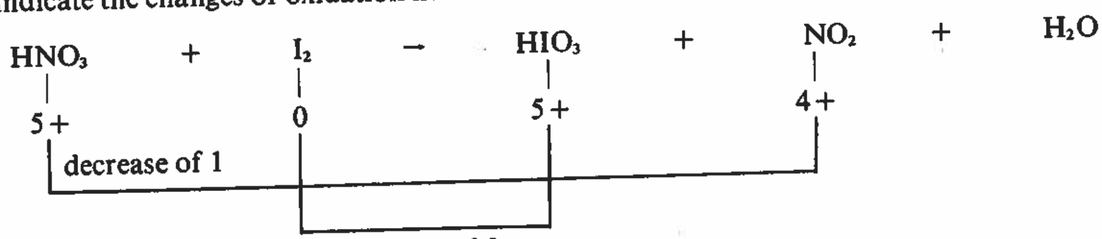
Solution There are five steps, *a-e* below, in the solution.

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

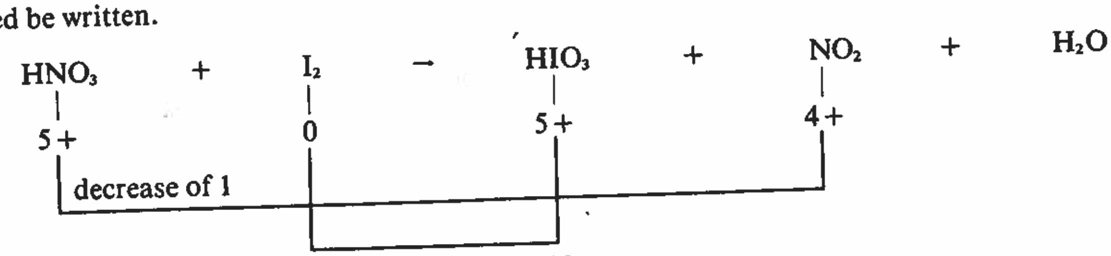
$$\overset{1+\ 5+\ 2-}{HNO_3} + \overset{0}{I_2} \rightarrow \overset{1+\ 5+\ 2-}{HIO_3} + \overset{4+\ 2-}{NO_2} + \overset{1+\ 2-}{H_2O}$$

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

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



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.



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.

	$10HNO_3$	+ I_2	→	$2HIO_3$	+	$10NO_2$	+	$4H_2O$
oxygen atoms	$10\times3=30$			$2\times3=6$	+	$10\times2=20$	+	4
				total = 30				

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.

	$3Cu + 8HNO_3 \rightarrow$	$3Cu(NO_3)_2$	$+ 2NO$	$+ 4H_2O$
oxygen atoms	$8 \times 3 = 24$	$3 \times 3 \times 2 = 18$	$+ 2$	$+ 4$
		total = 24		

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

7. ____$FeCl_3$ + ____H_2S $\rightarrow$ ____$FeCl_2$ + ____S + ____HCl

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

9. ____Cu + ____H_2SO_4 $\rightarrow$ ____$CuSO_4$ + ____SO_2 + ____H_2O

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

11. ____Zn + ____H_2SO_4 $\rightarrow$ ____$ZnSO_4$ + ____SO_2 + ____H_2O

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

13. ____$HMnO_4$ + ____HCl $\rightarrow$ ____$MnCl_2$ + ____Cl_2 + ____H_2O

14. ____Bi_2S_3 + ____HNO_3 $\rightarrow$ ____$Bi(NO_3)_3$ + ____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. $KIO_3 + H_2SO_3 \rightarrow KI + H_2SO_4$

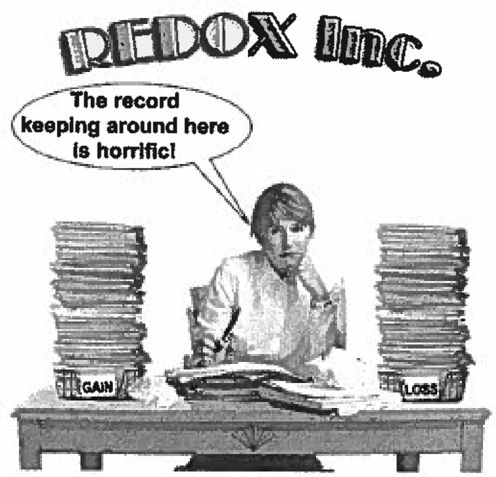
10. $K_2Cr_2O_7 + HCl \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$

Chemistry: Form WS10.1.4A Name ______________________

ELECTROCHEMISTRY Date ________________ Period _____

Balancing Redox Reactions

Redox equations are often too complex to balance by inspection alone. Instead, they are balanced by the *half-reaction method* or *ion-electron method*. In redox reactions, the number of electrons lost is always equal to the number of electrons gained. Keeping track of the electrons helps to balance the parts of the equation that can't be balanced by inspection. This is done by the procedure outlined below.



Balance the following: $K_2Cr_2O_7 + H_2S + HCl \rightarrow CrCl_3 + KCl + S + H_2O$

Step 1: Write the ionic equation.

$2K^+ + Cr_2O_7^{2-} + 2H^+ + S^{2-} + H^+ + Cl^- \rightarrow Cr^{3+} + Cl^- + K^+ + Cl^- + S + H_2O$

Step 2: Determine the oxidation states.

$2K^+ + Cr_2O_7^{2-} + 2H^+ + S^{2-} + H^+ + Cl^- \rightarrow Cr^{3+} + Cl^- + K^+ + Cl^- + S + H_2O$

+1 +6 −2 +1 −2 +1 −1 +3 −1 +1 −1 0 +1 −2

Step 3: Write oxidation half reaction, balancing atoms and charge.

$H_2S \rightarrow 2H^+ + S^0 + 2e^-$

Step 4: Write reduction half reaction, balancing atoms and charge.

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

Step 5: Conserve charge (electrons lost = electrons gained).

$3H_2S \rightarrow 6H^+ + 3S^0 + 6e^-$

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

Step 6: Combine half reactions.

$Cr_2O_7^{2-} + 3H_2S + 8H^+ \rightarrow 2Cr^{3+} + 3S^0 + 7H_2O$

Step 7: Combine ions to form compounds in original equation.

$K_2Cr_2O_7 + 3H_2S + 8HCl \rightarrow 2CrCl_3 + 2KCl + 3S + 7H_2O$

In **Step 1** the ions are separated making the spectators easier to identify. In **Step 2** the oxidation states are determined so it is possible to tell what was oxidized and what was reduced. In **Steps 3** and **4**, half reactions are written showing the number of electrons transferred. Note that in the oxidation half, $2H^+$ are needed to balance the hydrogen in hydrogen sulfide. In the reduction half, $7H_2O$ are needed to balance the oxygen in the dichromate ion, and as a result $14H^+$ are needed on the reactant side. In **Step 5**, the half reactions are multiplied by the correct coefficients to make the number of electrons lost equal the number of electrons gained. In **Step 6**, note that the H^+ ions remaining are the net from the two half reactions where they are on opposite sides of the equation.

Balance the equations below by following the procedure above.

1. $H_2S(aq) + HNO_3(aq) \rightarrow S(s) + NO_2(g) + H_2O(\ell)$

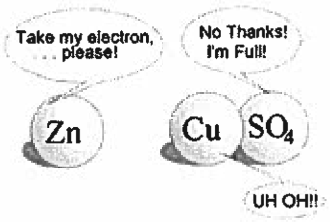
2. $LiNO_3(aq) + FeCl_2(aq) + HCl(aq) \rightarrow NO(g) + LiCl(aq) + FeCl_3(aq) + H_2O(\ell)$

3. $Na_2Cr_2O_7(aq) + HI(aq) \rightarrow CrI_3(aq) + NaI(aq) + I_2(s) + H_2O(\ell)$

4. $NaClO_3 + HCl \rightarrow ClO_2 + NaClO_4 + NaCl + H_2O$

5. $PbS(s) + HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + S(s) + NO(g) + H_2O(\ell)$

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		
	Au		
Least			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]

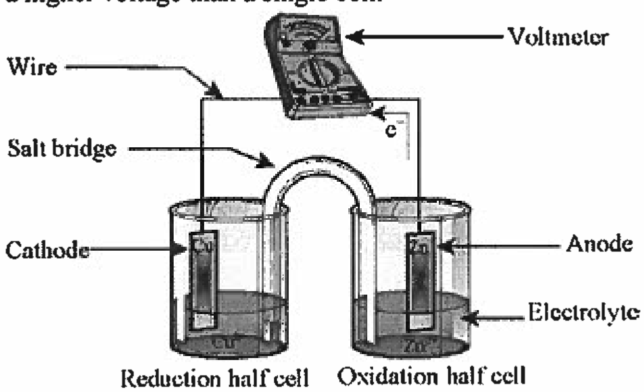
1. $Mg(s) + HCl(aq) \rightarrow$ ______________________
2. $Ag(s) + Cu(NO_3)_2(aq) \rightarrow$ ______________________
3. $Zn(s) + Mn(CH_3COO)_7(aq) \rightarrow$ ______________________
4. $Al(s) + HCl(aq) \rightarrow$ ______________________
5. $Cu(s) + HBr(aq) \rightarrow$ ______________________
6. $Cu(s)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

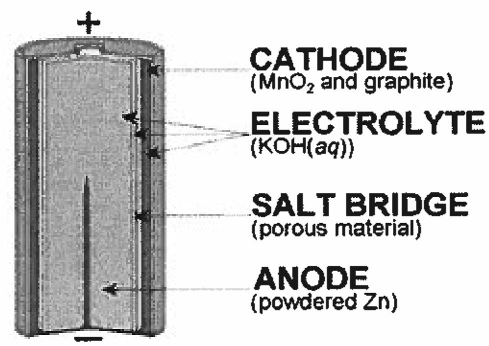
Chemistry: Form WS10.3.1A   Name ______________________

REDOX AND ELECTROCHEMISTRY   Date ______________ Period _____

A Salt and Battery

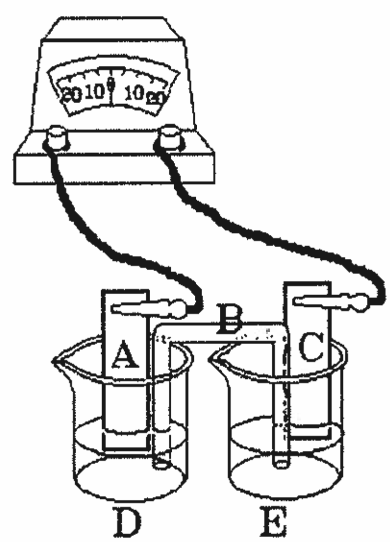
Portable electronic devices run on batteries. The electricity generated by a battery comes from a chemical reaction known as an oxidation-reduction reaction. During an a single replacement, a type of oxidation-reduction reaction, more active metals transfer electrons to less active metals. As a result, 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?

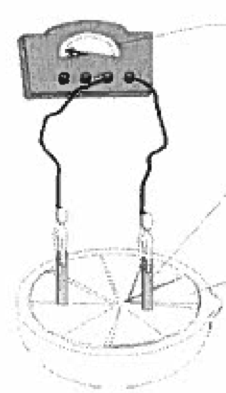
____

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

Name ______________________

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*.

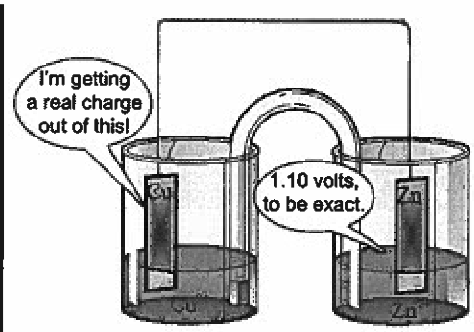
Procedure

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

Name ______________________

Date ______________ Period _____

Activity and Electricity

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

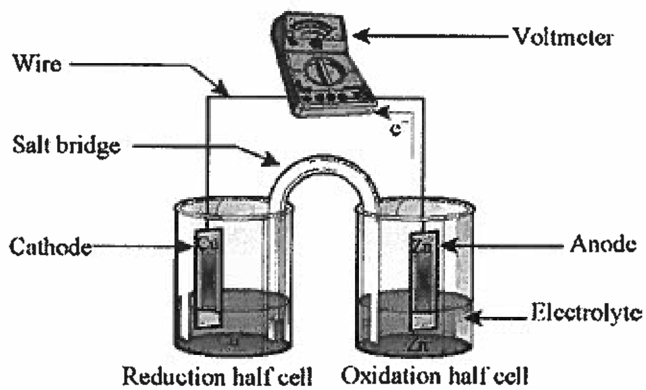
Notes

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

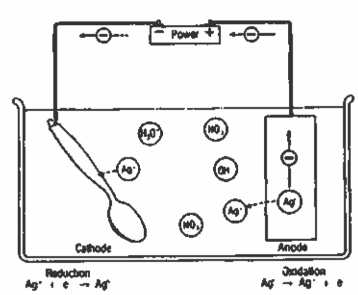
★ Parts of an electrochemical cell
  ☆ electrodes
    ★ anode — place where oxidation occurs
    ★ cathode — place where reduction occurs
  ☆ half cells — separate containers in which oxidation and reduction half reactions occur

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



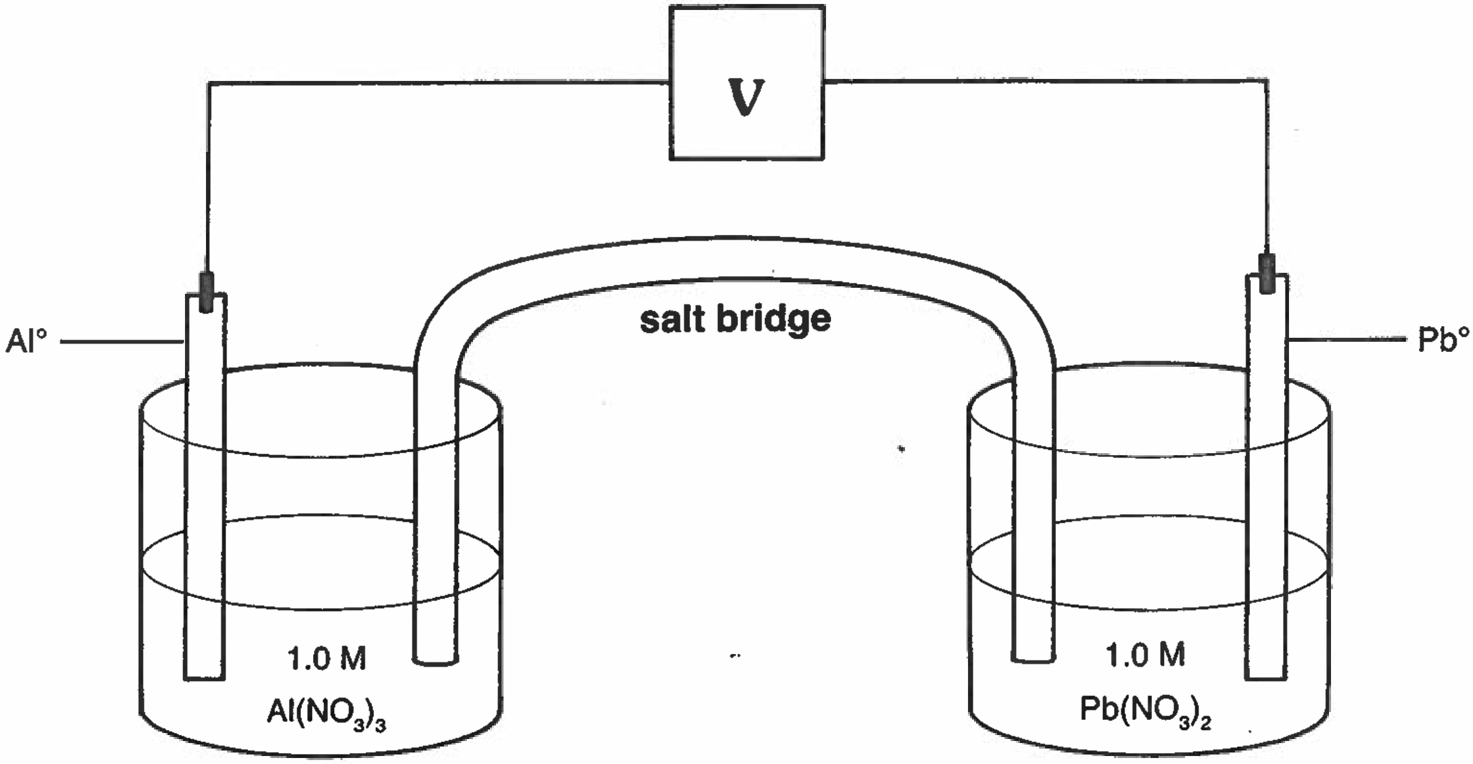
  ☆ U-tube or salt bridge — lets ions travel between half cells to complete the circuit

★ 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



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.

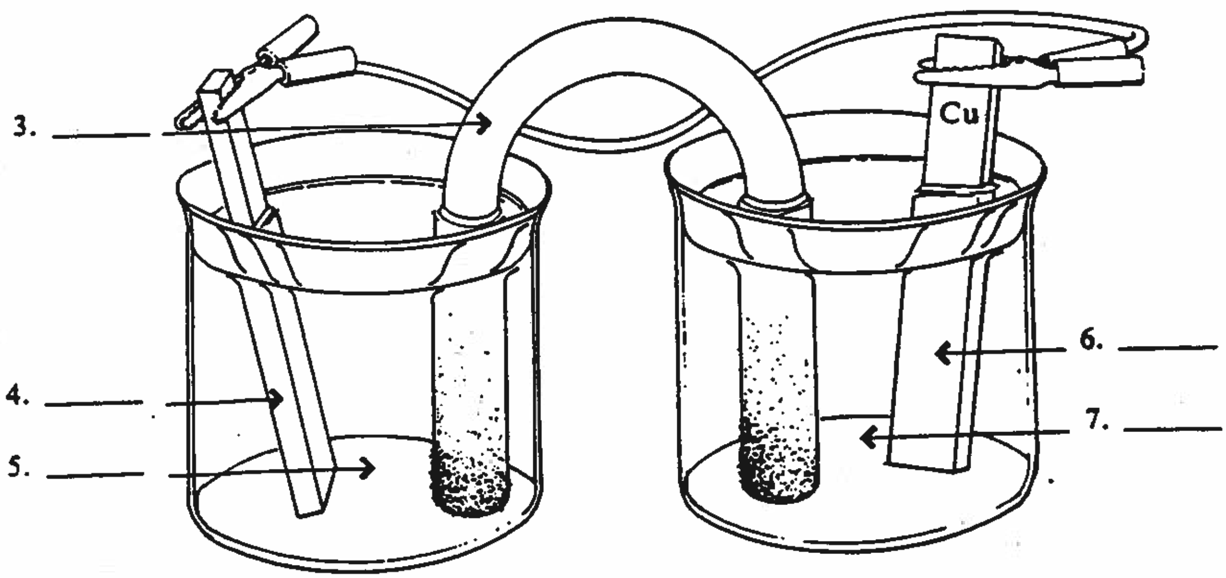
1. What names applied to electrochemical cells recognize the contributions of two Italian scientists? ______________ and ______________
2. What four substances are used to make a Daniell cell? ______________________________

____

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

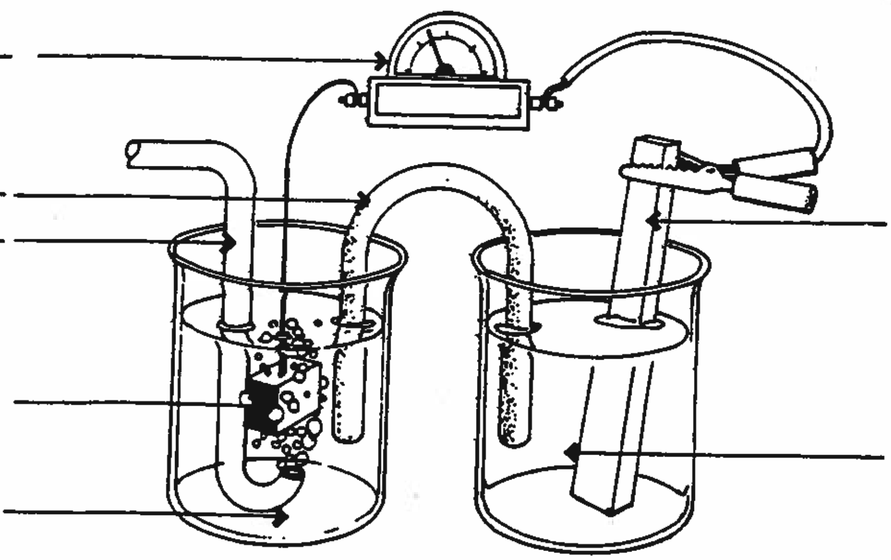
A half-cell (or electrode) consists of a metal strip immersed in a container of an electrolyte. Provision is made for the ions of the electrolyte to move either through the walls of a porous cup or through a salt bridge.

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

13. ______
14. ______
15. ______
16. ______
17. ______



18. ______
19. ______

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? ____________________

____

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.

____

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

____

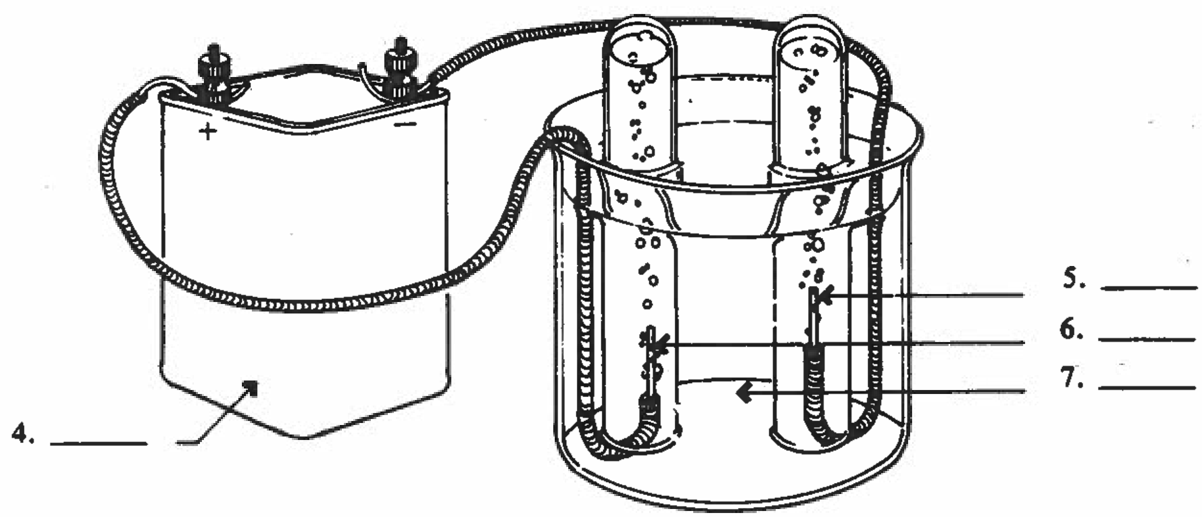
3. Write the equation for the cathode half-reaction.

____

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.

____

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.

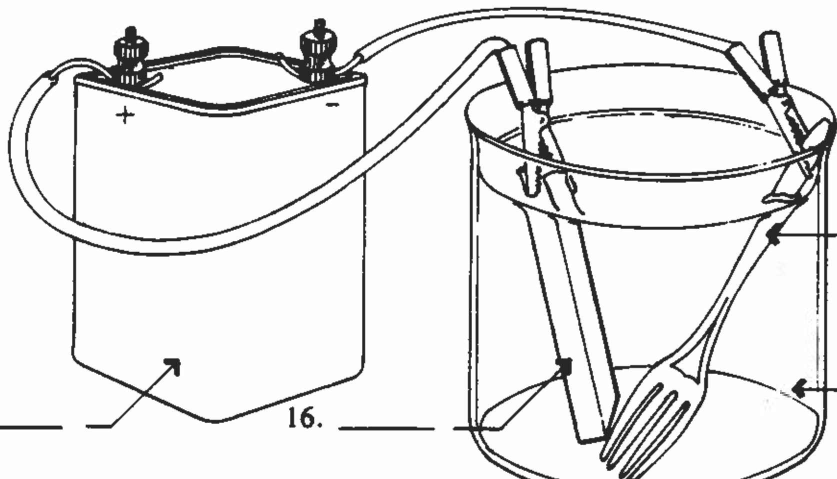
	Quantity	Change	Reason
11.	Na^+		
12.	Cl^-		
13.	OH^-		
14.	pH		

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.	$[Ag^+]$		

Chemistry: Form Ls10.1A Name ______________________

REDOX AND ELECTROCHEMISTRY 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

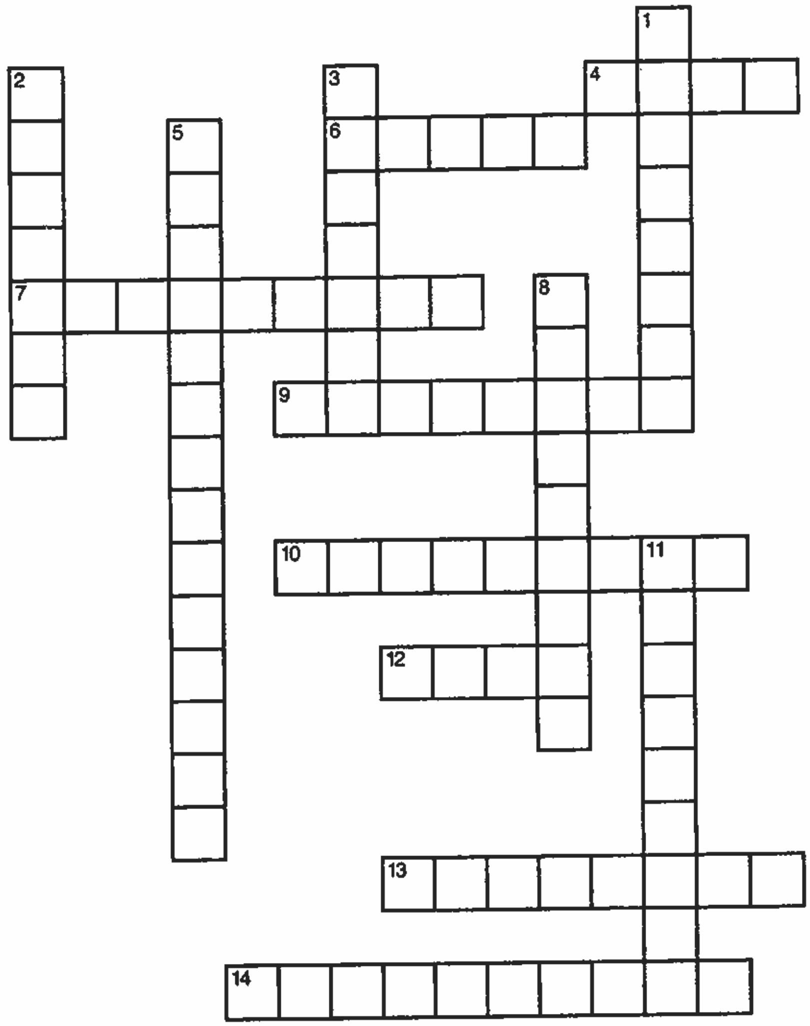
?? = +12

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

+12 ÷ 2 = +6

ELECTROCHEMISTRY CROSSWORD

Name ____________________



ACROSS

4. Unit of electrical potential
6. Electrode where oxidation takes place
7. Both atoms and _____ must be balanced in a redox equation.
9. The anode in an electrochemical cell has this charge.
10. Gain of electrons
12. Voltage of an electrochemical cell when it reaches equilibrium
13. A substance that is oxidized is the _____ agent.
14. Allows the flow of ions in an electrochemical cell

DOWN

1. The anode in an electrolytic cell has this charge.
2. Another word for an electrochemical cell
3. Electrode where reduction takes place
5. Process of layering a metal onto a surface in an electrolytic cell
8. Loss of electrons
11. A substance that is reduced is the _____ agent.

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

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^-

Name ______________________ Class______ Date____________

Unit 8 Redox and Electrochemistry

Activity 8-1
Oxidation Number

Introduction

Categories of reactions have been established to make the study of chemistry more efficient. One category is that of oxidation-reduction reactions, called redox for short. Each atom in a redox reaction can be assigned an oxidation number. The oxidation number, or oxidation state, is the apparent charge assigned to an atom in a particular molecule or ion according to a certain set of rules.

Rules for finding oxidation number

Choose words or numbers from the word list to fill in the blanks in the following statements, which summarize the rules for finding oxidation numbers.

Word List

charge	1+
elements	1−
sum	2+
0	2−

1. The oxidation number of a one-atom ion is equal to the ________________ of the ion.
2. The oxidation number of an element in the elemental (uncombined) state is ______.
3. The oxidation number of oxygen in most compounds is ______. However, in peroxides, such as Na_2O_2 and H_2O_2, the oxidation number of oxygen is ______. When oxygen is combined with the more electronegative element, fluorine, in OF_2, the oxidation number of oxygen is ______ and the oxidation number of fluorine is ______.
4. The oxidation number of hydrogen in most compounds is ______. In the compounds of active metals, such as NaH and BeH_2, the oxidation number of hydrogen is ______.
5. In neutral compounds, oxidation numbers are assigned to atoms of any other __________ so that the sum of the oxidation numbers for all atoms in the compound is equal to ______.
6. Oxidation numbers are also assigned to atoms of any other ________________ in polyatomic ions, so that the ________________ of the oxidation numbers for all atoms in the ion is equal to the ________________ on the ion.

Assign the correct oxidation number to each atom in each of the following. Write the numbers directly above the symbols in each formula, as in the samples.

Samples

1+ 2−	3− 1+	1+ 6+ 2−
H_2O	NH_4^+	Na_2SO_4
(water)	(ammonium ion)	(sodium sulfate)

A. 7. K 8. RbCl 9. Na_2O 10. H_2O_2 11. $MgBr_2$

12. CaS 13. $K_2Cr_2O_7$ 14. H_3PO_4 15. $(NH_4)_3PO_4$ 16. $(NH_4)_2S$

17. $FeCl_2$ 18. $FeCl_3$ 19. FeO 20. N_2 21. N_2O

B. 22. NO 23. N_2O_4 24. NO_2 25. N_2O_3 26. N_2O_5

27. Ag 28. $HC_2H_3O_2$ 29. $CaCO_3$ 30. CO_2 31. CO

32. $Na_2S_2O_3$ 33. S_8 34. Na_2SO_3 35. $MnSO_4$ 36. $KMnO_4$

C. 37. OH^- 38. CO_3^{2-} 39. NO_2^- 40. CrO_4^{2-} 41. $Cr_2O_7^{2-}$

42. IO_3^- 43. IO_4^- 44. SCN^- 45. HSO_4^- 46. PO_4^{3-}

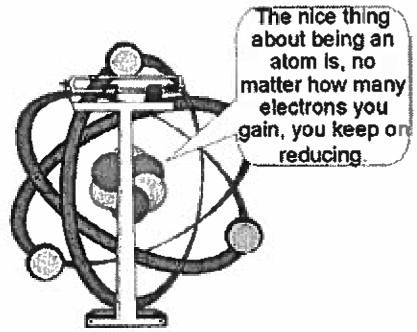
47. HCO_3^- 48. Hg_2^{2+} 49. ClO^- 50. ClO_3^- 51. $S_2O_3^{2-}$

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. $SnCl_2 + 2HgCl_2 \rightarrow SnCl_4 + Hg_2Cl_2$		