Experiment 10 Name: ____________________________

ReDOX REACTIONS AND ACTIVITY SErIES

In this experiment, you will observe single displacement reactions that occur when a metal is oxidized by salts (or ionic compounds) and by acids. You will also learn to use the activity series to predict if a single displacement metal-salt or metal-acid reaction occurs or not.

Oxidation-Reduction (Redox) Reactions

Every atom, whether in a compound or in its elemental state, has an oxidation number. This oxidation number describes how many electrons it has compared to its neutral, elemental state.

In some redox chemical reactions, the number of electrons an atom has changes from the reactants to the products. This results in a change in oxidation number.

An atom is **oxidized** if the atom loses electrons (its oxidation number becomes more positive).

An atom is **reduced** if the atom gains electrons (its oxidation number becomes more negative).

The total number of electrons in a redox reaction does not change. Thus, if one atom gains 1 electron, then another atom in the reaction must lose 1 electron. Note that it is possible for one atom to gain 2 electrons and 2 atoms to lose 1 electron each.

To determine the oxidation number of each atom in a reaction, the following rules are used:

1. An atom in its <u>Elemental Form</u>: Always Zero
   Cu (s): oxidation number for Cu is zero, 0
   H_2 (g): oxidation number for H is zero, 0

2. <u>Monatomic Ion</u>: Equal to the ionic charge
   K^+ (aq): oxidation number for K is +1
   S^{-2} (aq): oxidation number for S is -2

3. <u>Nonmetals</u>: Typically have negative oxidation numbers
   a) Hydrogen is +1, Exception are hydrides: -1
   b) Oxygen is -2 , Exception are peroxides: -1
   c) Fluorine is -1, always
   d) Other halogens are -1, Exception when bonded to oxygen (oxyanions): positive

4. <u>The sum of oxidation numbers</u>:
   a) For a neutral compound: Sum equals zero
   b) Polyatomic Ion: Sum equals the charge of the ion

   If the oxidation number cannot be ascertained from these rules directly, then they must be calculated. The following example demonstrates how each atom in $KMnO_4$ can be ascertained.

$KMnO_4$ can first be broken into its respective ions: $KMnO_4 \rightarrow K^+ + MnO_4^-$

K^+ is a monoatomic ion, thus K has an oxidation state equal to its charge of +1

O is always -2, unless in a peroxide (which this compound is not), thus the oxidation state for O is -2

The oxidation state for Mn must be calculated. If you add up the oxidation state of Mn and each of the 4 Oxygens in MnO_4 it should equal the charge on the polyatomic ion of -1

$$Mn + 4 * Oxygen = -1$$

We know Oxygen = -2, so we substitute that in and solve for Mn.

$$Mn + 4\,(-2) = -1$$
$$Mn + (-8) = -1$$
$$Mn = -1 + 8 = +7$$

So, the oxidation state for each atom in $KMnO_4$ is:
K: <u>+1</u> Mn: <u>+7</u> O: <u>-2</u>

Oxidation of Metals

Single displacement reactions conform to the general pattern:

A	**+**	**BX**	→	**AX**	**+**	**B**
Metal	+	Salt	→	Salt	+	Metal
(More Active)						(Less Active)

The cation, B^{n+}, in solution is said to be displaced by the element A^0 as it oxidizes to A^{m+}.

When copper, a reddish-brown metal reacts with colorless silver nitrate solution, a gray solid deposits and the solution turns blue, due to the products silver and copper(II) nitrate. Cu displaces Ag.

$1\ Cu\ (s)$	+	$2\ AgNO_3\ (aq)$	→	$1\ Cu(NO_3)_2\ (aq)$	+	$2\ Ag\ (s)$	balanced chemical equation
reddish-brown metal		colorless solution		blue solution		gray metal	

The complete and net ionic equations for this reaction are:

$$1\ Cu^0\ (s) + 2\ Ag^+\ (aq) + 2\ NO_3^-\ (aq) \rightarrow 1\ Cu^{2+}\ (aq) + 2\ NO_3^-\ (aq) + 2\ Ag^0\ (s)$$ complete ionic equation

$$1\ Cu^0\ (s) + 2\ Ag^+\ (aq) \rightarrow 1\ Cu^{2+}\ (aq) + 1\ Ag^0\ (s)$$

The net ionic equation shows that this reaction is a redox reaction, with the oxidation number of Cu increasing from 0 to +2 and that of Ag decreasing from +1 to 0. Cu is oxidized by Ag^+ in the salt solution. Ag^+ is thus reduced. Note that as the reaction proceeds, the solution turns blue because of the production of Cu^{2+} (aq), which is blue.

net ionic equation

When manganese is allowed to react with hydrobromic acid, bubbles appear due to the formation of hydrogen gas as Mn displaces H.

$$1\ Mn\ (s) + 2\ HBr\ (aq) \rightarrow 1\ MnBr_2\ (aq) + 1\ H_2\ (g)$$

balanced chemical equation

The complete and net ionic equations for this reaction are:

$$1\ Mn^0\ (s) + 2\ H^+\ (aq) + 2\ Br^-\ (aq) \rightarrow 1\ Mn^{2+}\ (aq) + 2\ Br^-\ (aq) + 1\ H_2{}^0\ (g)$$

complete ionic equation

$$1\ Mn^0\ (s) + 2\ H^+\ (aq) \rightarrow 1\ Mn^{2+}\ (aq) + 1\ H_2{}^0\ (g)$$

net ionic equation

Mn is oxidized by H^+ in the acid. The oxidation number of Mn increased from 0 to +2 and that of H decreased from +1 to 0. Therefore, H^+ is reduced.

Activity Series

Elements differ in their ability to displace other elements. For example, Cu can displace Ag but Ag cannot displace Cu.

$$1\ Cu(s) + 2\ Ag^+(aq) \rightarrow 1\ Cu^{2+}(aq) + 2\ Ag(s)$$

$$Ag(s) + Cu^{2+}(aq) \rightarrow \text{No reaction}$$

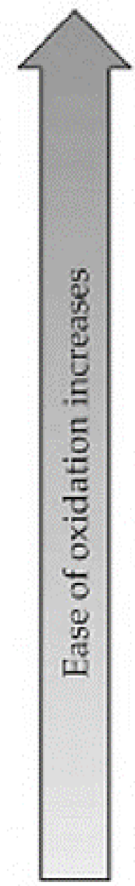
H can oxidize Mn but not Au. Therefore, bubbles will only be observed when Mn is placed in acid (H^+(aq)) and not when gold is placed in acid.

$$1\ Mn(s) + 2\ H^+(aq) \rightarrow 1\ Mn^{2+}(aq) + 1\ H_2(g)$$

$$Au(s) + H^+(aq) \rightarrow \text{No reaction}$$

The activity series allows for the prediction whether or not a single displacement reaction will occur. The activity series shown here lists the elements from the most active to the least active, that is, from the easiest to oxidize to the hardest:

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^{+}(aq) + e^{-}$
Potassium	$K(s) \longrightarrow K^{+}(aq) + e^{-}$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^{-}$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^{-}$
Sodium	$Na(s) \longrightarrow Na^{+}(aq) + e^{-}$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^{-}$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^{-}$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^{-}$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^{-}$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^{-}$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^{-}$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^{-}$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^{-}$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^{-}$
Hydrogen	$H_2(g) \longrightarrow 2\,H^{+}(aq) + 2e^{-}$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$
Silver	$Ag(s) \longrightarrow Ag^{+}(aq) + e^{-}$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^{-}$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^{-}$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^{-}$



An element can be oxidized by the ions of the elements below it in the activity series. For example, Ag is below Cu so Ag can oxidize Cu to Cu^{2+}. In other words, Cu displaces Ag in the salt solution because it is easier to oxidize Cu than Ag. H is below Mn in the series but above Au, thus it can oxidize Mn (to Mn^{2+}) but not Au.

PROCEDURE

I. Oxidation of Metals by Salts

1. Clean test tubes.

2. Make initial observations on the appearance of the metals Mg, Fe, Cu and Zn.

3. Make initial observations on the appearance of the solutions of $Mg(NO_3)_2$, $Fe(NO_3)_3$, $Cu(NO_3)_2$ and $Zn(NO_3)_2$.

4. Obtain four test tubes and to each one, add 5-10 drops of one of the solutions. Set these aside. They will be used as reference for solution color changes that may occur during reaction by placing the solutions against a white background.

> As a general overview of how to proceed, refer to the data table in Part B. Reactions of your worksheet. Combine each metal with each solution indicated in the data table.
>
> For instance, add one piece of magnesium to three test tubes. Add 5-10 drops of $Fe(NO_3)_3$ into the first test tube, $Cu(NO_3)_2$ into the second test tube, and $Zn(NO)_3)_2$ into the third. Record your observations in the data table. Repeat for the other metals.
>
> Do not do the combinations in the table indicated by an X. No reaction will take place when mixing Mg and $Mg(NO_3)_2$, for example since Mg will not displace itself.

5. Obtain three more test tubes. To each, add a small sample of the metal (a piece of metal or a spatula-tip full for powder). To each test tube, add 5-10 drops of one solution. Observe any changes to the metal and solution, referring to the unreacted metal and to the reference solutions prepared in Step 3 if needed. Record the color change of the solution and solid. Write NR for no reaction.

> Keep in mind the reaction that is taking place (displacement of metal). What is being produced in these reactions?
>
> Ignore gas (bubbles) being produced. This is a minor side reaction or trapped air in the test tube. Keep in mind the chemical reaction of the single displacement reaction for two metals, where no gas is being produced.

6. Dispose of the wastes in the proper waste container. Wash the test tubes, shake off any excess water then proceed to the next metal. The test tubes do not need to be completely dry to move on since water will not affect the reaction. Work with one metal at a time to ensure you do not lose track of the test tube contents for proper waste disposal.

7. Upon completion of all the metal-salt solution combinations, dispose of the reference solutions and wash all test tubes.

II. Oxidation of Metals by Hydrochloric Acid

1. Obtain four test tubes. To each, add a small sample of each metal.

2. Add 5-10 drops of 6 M HCl solution to each test tube. Watch for appearance of bubbles. Note the intensity and reaction rate (fast, intermediate, or slow). Write NR for no reaction.

> Caution: Metal + acid could be a violent reaction and can get very hot. Add the acid while the test tube is sitting in the rack and do not hold onto the bottom of the test tube to avoid burning yourself.

CLEAN-UP

- Dispose of wastes in designated containers in the front hood.
- Wash all test tubes. Shake off excess water and return to the test tube racks on your work station.

Name: ______________________________ **Date:** ____________________

Partner's Name: ______________________________

ACTIVITY SERIES

DATA AND OBSERVATIONS

I. Oxidation of Metals by Salts

A. Initial Observations

Metals

Mg ______________________________

Fe ______________________________

Cu ______________________________

Zn ______________________________

Solutions

$Mg(NO_3)_2$ ______________________________

$Fe(NO_3)_3$ ______________________________

$Cu(NO_3)_2$ ______________________________

$Zn(NO_3)_2$ ______________________________

B. Reaction Observations

Metals	Solutions			
	$Mg(NO_3)_2$	$Fe(NO_3)_3$	$Cu(NO_3)_2$	$Zn(NO_3)_2$
Mg	X	Solution: Solid: Metal Formed:	Solution: Solid: Metal Formed:	Solution: Solid: Metal Formed:
Fe	Solution: Solid: Metal Formed:	X	Solution: Solid: Metal Formed:	Solution: Solid: Metal Formed:
Cu	Solution: Solid: Metal Formed:	Solution: Solid: Metal Formed:	X	Solution: Solid: Metal Formed:
Zn	Solution: Solid: Metal Formed:	Solution: Solid: Metal Formed:	Solution: Solid: Metal Formed:	X

II. Oxidation of Metals by Acid

	Observations	Speed
Mg + HCl	______________________	__________
Fe + HCl	______________________	__________
Cu + HCl	______________________	__________
Zn + HCl	______________________	__________

POST-LAB QUESTIONS

1. Based on your experimental results, list the four metals from the most active one to the least active one.

    1. ____________________ (most active)
    2. ____________________
    3. ____________________
    4. ____________________ (least active)

2. Write the balanced chemical equation, the complete and net ionic equations. Identify which element is being oxidized and which element is being reduced.
   For the metals that did not react, write NR.
   Note: A charge of zero (neutral) may be indicated by a ^{0}. For example: $Zn^0(s)$

    a. Mg and HCl

    Balanced chemical equation:

    Complete ionic equation:

    Net ionic equation:

    Element Oxidized: ______________ Element Reduced: ______________

b. HBr and Cu (Note if copper reacts, it will become Cu (I) in solution)

Balanced chemical equation:

Complete ionic equation:

Net ionic equation:

Element Oxidized: ___________________ Element Reduced: ___________________

c. $Co(NO_3)_2$ and Al

Balanced chemical equation:

Complete ionic equation:

Net ionic equation:

Element Oxidized: ___________________ Element Reduced: ___________________

d. Au and $NiCl_2$ (Note if gold reacts, it will become Au (III) in solution)

Balanced chemical equation:

Complete ionic equation:

Net ionic equation:

Element Oxidized: ___________________ Element Reduced: ___________________

3. What are the oxidation numbers of each element in each of the following compounds? (Hint: you may have to separate a compound into individual polyatomic ions first, before you can solve the oxidation numbers for each element)

   a. NO_3^{-1}

      N: ____________ O: ____________

   b. $(NH_4)_2CO_3$

      N: ____________ H: ____________ C: ____________ O: ____________

   c. NaH

      Na: ____________ H: ____________