Name ______________________ Date ________________ Class ______________

34 DETERMINATION OF AN ACTIVITY SERIES

Small-Scale Experiment for text Section 20.1

OBJECTIVES

- **Derive** an activity series from experimental data.
- **Describe** oxidation–reduction reactions by writing net ionic equations and half-reactions.
- **Formulate** a functional definition for *oxidation* and *reduction.*

INTRODUCTION

You have seen that magnesium reacts with HCl much more readily than does zinc. Copper appears not to react at all with hydrochloric acid. The half-reactions that describe these processes reveal that electrons are lost by the metals and gained by the hydrogen ions. For example, the half-reactions for the reaction of magnesium with hydrochloric acid are

$$2H^+ + 2e^- \rightarrow H_2(g) \qquad \text{reduction}$$

$$Mg(s) \rightarrow Mg^{2+} + 2e^- \qquad \text{oxidation}$$

The relative ease with which a metal loses electrons is called its *oxidation potential.* You have seen, for example, that magnesium is more reactive toward acids than is zinc. This means that magnesium loses electrons more readily than zinc does. In other words, magnesium has a higher oxidation potential.

Conversely, the relative ease with which a metal cation gains electrons is called its *reduction potential.* For example, we have seen that copper ions react with zinc metal according to the following half-reactions and net ionic equation:

Reduction	$Cu^{2+} + 2e^- \rightarrow Cu(s)$
Oxidation	$Zn(s) \rightarrow Zn^{2+} + 2e^-$
Net ionic	$Cu^{2+} + Zn(s) \rightarrow Cu(s) + Zn^{2+}$

Because copper ions react with zinc metal by taking away its electrons, we can say that copper has the greater reduction potential and zinc has the greater oxidation potential.

Name ______________________ Date ________________ Class ________________

PURPOSE

In this experiment, you will carry out a series of reactions in which electrons are gained and lost between metals and metal ions. You will organize the experiments so that metals actually compete for one another's electrons. You will observe which metals lose electrons more readily and which metal ions gain electrons more readily. You can compare the relative tendencies of metals to gain and lose electrons by counting the number of reactions each metal and each metal ion undergoes. The metal involved in the most reactions is the one that loses electrons (is oxidized) the most readily. It has the highest oxidation potential. The metal involved in the least number of reactions loses electrons the least readily. It has the lowest oxidation potential. A list of metals ranging from those most easily oxidized to those least easily oxidized is called an activity series. From this data, you will derive an activity series for common metals.

SAFETY

- Wear safety goggles.
- Use full small-scale pipets only for the controlled delivery of liquids.
- Don't chew gum, drink, or eat in the laboratory. Never taste a chemical in the laboratory.

MATERIALS

Small-scale pipets of the following solutions:
copper(II) sulfate ($CuSO_4$)
iron(II) sulfate ($FeSO_4$)
magnesium sulfate ($MgSO_4$)
silver nitrate ($AgNO_3$)
zinc chloride ($ZnCl_2$)

Four pieces each of the following solid metals:
zinc (Zn)
silver (Ag)
magnesium (Mg)
iron (Fe)
copper (Cu)

EQUIPMENT

small-scale reaction surface

Name ______________________ Date ________________ Class ________________

EXPERIMENTAL PAGE

Add two drops of each salt solution to one piece of each of the indicated metals.

One piece each	$CuSO_4$ (Cu^{2+})	$FeSO_4$ (Fe^{2+})	$MgSO_4$ (Mg^{2+})	$AgNO_3$ (Ag^{+})	$ZnCl_2$ (Zn^{2+})
Zn(*s*)					
Ag(*s*)					
Mg(*s*)					
Fe(*s*)					
Cu(*s*)					

Place this side of the Experimental Page facedown. Use the other side under your small-scale reaction surface.

Name ______________________ Date __________________ Class ________________

EXPERIMENTAL DATA

Record your results in Table 34.1 or in a copy of the table in your notebook.

Table 34.1 Metal Activity

	$CuSO_4$ (Cu^{2+})	$FeSO_4$ (Fe^{2+})	$MgSO_4$ (Mg^{2+})	$AgNO_3$ (Ag^{+})	$ZnCl_2$ (Zn^{2+})	
Zn(*s*)						
Ag(*s*)						
Mg(*s*)						
Fe(*s*)						
Cu(*s*)						)

Write oxidation and reduction half-reactions for each metal in the spaces provided next to the table.

CLEANING UP

Avoid contamination by cleaning up in a way that protects you and your environment. Carefully clean and dry each piece of leftover metal and place it in the appropriate recycling container. Clean the small-scale reaction surface by absorbing the contents onto a paper towel, rinse it with a damp paper towel, and dry it. Dispose of the paper towels in the waste bin. Wash your hands thoroughly with soap and water.

Name ______________________ Date ________________ Class ________________

QUESTIONS FOR ANALYSES

Use what you learned in this experiment to answer the following questions.

1. Read across each row of Table 34.1. Count the number of reactions that each metal undergoes with the metal salts. Order them according to reactivity. This order is called an activity series.

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2. Reading down the rows of Table 34.1, count the number of reactions that each metal ion undergoes. Order them according to reactivity.

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3. Compare the reactivity of the metal ions to that of the metals. What are the similarities and differences?

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4. For each reaction you observe, a metal is oxidized; that is, it gives away its electrons. In doing so, the metal becomes a metal ion. Write an oxidation half-reaction for each reaction you observed. Write them in order of reactivity, with the most reactive first.

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5. For each reaction you observe, a metal ion is reduced; that is, the metal ion gains electrons. In doing so, the metal ion becomes a metal atom. Write a reduction half-reaction for each reaction you observed. Write them in order of reactivity, with the most reactive first.

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6. Examine your two lists of half-reactions and describe how each list can be used to predict whether a reaction will occur between a metal and a metal ion.

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Name ______________________ Date ________________ Class ______________

7. Add the half-reactions to obtain the net ionic equations for each of the reactions you observed. Use either list and choose two half-reactions. Reverse the one lower on the list and add them together. Make sure the number of electrons gained equals the number of electrons lost. For example,

Oxidation	$Mg(s) \rightarrow Mg^{2+} + 2e^-$
Reduction	$2\,[Ag^+ + e^- \rightarrow Ag(s)]$
Net Ionic Equation	$Mg(s) + 2Ag^+ \rightarrow Mg^{2+} + 2Ag(s)$

Notice that silver metal does not give its electrons to a zinc ion:

$Ag^0(s) + Zn^{2+} \rightarrow$ No Visible Reaction

NOW IT'S YOUR TURN!

1. Repeat the experiment, using common household objects in place of the metals. For example, use galvanized nails or zinc-plated washers for zinc, silverware for silver, staples for iron, and pennies for copper. How do your results compare with the original experiment?