

# REDOX CHEMISTRY

## ACTIVITY OF METALS

### OBJECTIVES:

To investigate the relative reactivity of metals; specifically, their tendency to be oxidized; to develop an activity series for these metals.

### MATERIALS:

Small samples of metals (zinc, copper, nickel, magnesium and lead); 0.10 M solutions of  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{Zn}(\text{NO}_3)_2$ ,  $\text{Ni}(\text{NO}_3)_2$ ,  $\text{Mg}(\text{NO}_3)_2$ , and  $\text{Pb}(\text{NO}_3)_2$ ; a metal file or sandpaper to clean metal surfaces

### EQUIPMENT:

15 × 100 mm test tubes

### SAFETY:

Safety goggles should be worn at all times.

### WASTE DISPOSAL:

All test solutions should be collected and disposed of in the inorganic waste container. Solid metal samples can be collected for disposal or for recycling.

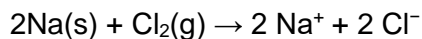
### REVIEW:

You should understand the difference between elements and compounds. You should be able to recognize evidence of chemical reactions, and know how to use equations to represent chemical reactions.

## INTRODUCTION

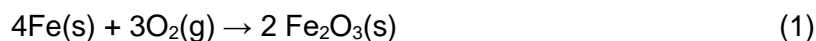
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All chemical reactions involve the breaking of some bonds and the formation of others. A common and important class of reactions involves the transfer of electrons between species. This transfer of electrons can be direct and obvious, as is the case in the formation of ionic compounds. In the reaction between Na and Cl below, for example, the Na atoms lose electrons and the Cl atoms gain electrons to form cations and anions.



In such reactions, the species that loses electrons is said to undergo **oxidation**, while the species that gains electrons undergoes **reduction**. Since both oxidation and reduction occur in tandem, these reactions are often called **redox** reactions.

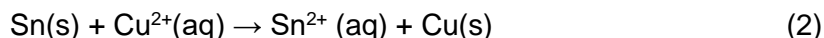
The transfer of electrons is not always obvious, however. Take, for example, the reaction between iron and oxygen to form iron oxide:



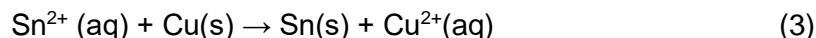
In this case it is not obvious which species has gained or lost electrons. The early definition of oxidation was based on the reaction of substances with oxygen; a substance was oxidized if it gained oxygen, while a substance was reduced if it lost oxygen. Using this definition, the iron in Eq. 1 is oxidized since it gains oxygen to form  $\text{Fe}_2\text{O}_3$ ; since the iron has been oxidized, the elemental  $\text{O}_2$  must have been reduced. Another generalization can be drawn from Eq. 1. In reactions involving metals and non-metals, metals tend to lose electrons (are oxidized) while non-metals tend to gain electrons (are reduced).

In this experiment you will examine the ability of several metals to be oxidized. The ease with which a metal loses electrons is often referred to as the **activity** of the metal; the more readily the metal loses electrons, the more active the metal. In general, metal activity increases as you move down and to the left in the periodic table.

One way to study activities is to use chemical reactions in which metal species compete for electrons. An elemental metal is placed in contact with a solution of a metal ion. If the elemental metal is more active (has a greater tendency to lose electrons) it will transfer electrons to the metal ion in solution. Evidence of a reaction is often observed by a change in color of the solution or the formation of a solid film on the surface of the elemental metal. For example, suppose a sample of tin metal is placed in contact with a solution of copper (II) ions. The possible reaction is presented in Equation 2.



The formation of a copper film on the surface of the tin metal and the disappearance of the blue color of the  $\text{Cu(II)}$  ion from solution would be evidence of a chemical reaction. The tin metal would have been oxidized and the copper(II) ion would have been reduced, and the tin would be identified as the more active metal. If no changes were observed in solution you might conclude that no reaction occurred. However, negative results are sometimes misleading and it would be useful to study the reverse reaction given in Equation 3.



In this case, elemental copper would be placed in a solution containing  $\text{Sn}^{2+}$  ion. Evidence of a reaction would include the formation of a solid film on the surface of the copper metal and the appearance of a blue color in solution as  $\text{Cu(II)}$  ions are formed. If evidence of this reaction was observed, copper would be identified as the more active metal.

In a similar fashion you will investigate the relative activity of a series of metals by placing samples of the elemental metals in contact with solutions of other metal ions. By observing whether or not a reaction occurs you will be able to identify the more active metal. Once you have completed the set of reactions you will create an activity series by listing the metals in order from most active to least active.

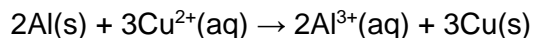
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## PRE-LAB QUESTIONS

1. Define each of the following terms:

- a. Oxidation:
- b. Reduction:
- c. Redox reaction:
- d. Activity:

2. In the following reaction identify the substance that is oxidized and the substance that is reduced. Which metal is more active?



Oxidized =

Reduced =

More active =

3. Why is it necessary to clean the surfaces of the metal samples before testing them in the metal ion solutions?

4. You will not be testing the activity of sodium (Na) or potassium (K) in this lab. Based on the periodic trends for activity, would you expect these metals to be *more active* or *less active* than the metals studied in this lab? Explain.

## PROCEDURE

1. Obtain small samples of the metals copper, zinc, magnesium, lead and nickel. Many metals form oxide layers on the surface that will affect your test results. Use the file or sandpaper to clean the surfaces of these metals—the cleaned samples should exhibit the luster typical of metals.
2. Pour about 2 mL of  $\text{Cu}(\text{NO}_3)_2$  into each of four small test tubes. Place a small piece of zinc into one of the test tubes, a small piece of magnesium into the second test tube, a small piece of lead in the third test tube, and a small piece of nickel in the fourth test tube. You may need to cut the metal samples into smaller pieces in order to fit them easily in the test tubes. Label the test tubes with the symbol for the metal being tested. (Question: Why don't you need to test a sample of elemental copper?)
3. Wait 5–10 minutes to allow time for reactions to occur. Record your observations on your Data Sheet and indicate if a reaction occurred between the metal sample and the metal ion in solution. If there is no evidence of a reaction, write NR in the appropriate box on your Data Sheet.
4. Decant the solutions in the test tubes into the sink and flush with plenty of water. Collect the solid metal samples for disposal or recycling as directed by your TA.
5. Rinse the test tubes with distilled water.
6. Repeat Steps 1–5 using  $\text{Zn}(\text{NO}_3)_2$  solution and small samples of the other metals. **Note:** You do not need to test zinc metal in this case.
7. Repeat Steps 1–5 with solutions of  $\text{Mg}(\text{NO}_3)_2$ ,  $\text{Ni}(\text{NO}_3)_2$ , and  $\text{Pb}(\text{NO}_3)_2$ . Record all observations on the Data Sheet.

NAME: \_\_\_\_\_ SECTION: \_\_\_\_\_ DATE: \_\_\_\_\_

## REDOX CHEMISTRY – ACTIVITY OF METALS

# DATA SHEET

Complete the table below with your observations, and indicate if reaction has occurred. If there is no evidence of reaction write NR in the appropriate box in the table.

Metal	Solution				
	$\text{Cu}(\text{NO}_3)_2$	$\text{Zn}(\text{NO}_3)_2$	$\text{Mg}(\text{NO}_3)_2$	$\text{Ni}(\text{NO}_3)_2$	$\text{Pb}(\text{NO}_3)_2$
Cu					
Zn					
Mg					
Ni					
Pb					

Based on your observations, list the metals from most active to least active:

\_\_\_\_\_ > \_\_\_\_\_ > \_\_\_\_\_ > \_\_\_\_\_  
 (Most active) (Least active)

Explain your reasoning

NAME: \_\_\_\_\_ SECTION: \_\_\_\_\_ DATE: \_\_\_\_\_

## POST-LAB QUESTIONS

1. Write chemical equations that represent the potential reactions of copper metal with the four metal ion solutions. Based on your observations, which reactions occurred and which did not?
2. Galvanized steel, used in construction and infrastructure, consists of steel (mostly iron) coated with an outer layer of zinc metal. Based on your observations, how does galvanizing steel protect the steel from corrosion?
3. Why is it not necessary to test the reactivity of the elemental metals with solutions of the same metal ion? To illustrate your explanation, write the equation for the potential chemical reaction using one of the metals studied in this lab.
4. Some metals form crusty oxide layers that can flake off to expose clean metal surfaces to moisture and air. Other metals form oxide layers that are durable and relatively impermeable. Use this fact to explain why aluminum, a more active metal than iron, is less susceptible to corrosion.