

# ELECTRON TRANSFER BETWEEN METALS

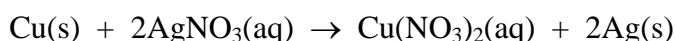
## LAB ELC 1.CALC

### INTRODUCTION

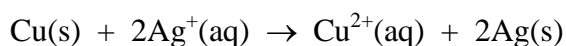
Chemical reactions involving electron transfer from one chemical substance to another are called **oxidation-reduction**, or **redox**, reactions. Electron transfer reactions are very important. You experience such reactions and their consequences everyday. Redox reactions occur in many common procedures such as: the rusting of iron, the bleaching of hair, the production of electricity in batteries, and many biological processes. This experiment will focus on electron transfer reactions that produce electrical energy, like those used in batteries.

In redox reactions, one species loses electrons and another gains electrons. When a substance loses electrons the process is called **oxidation**. When a substance gains electrons the process is called **reduction**. Metals vary in how easily they lose electrons. A metal that loses electrons easily is said to have a high **activity**. In a redox reaction, the more active metal species loses electrons and the less active metal species gains electrons.

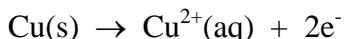
This is an example of a redox reaction:



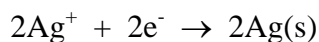
The transfer of electrons is more easily seen in the **net ionic equation**, which only shows the atoms undergoing changes in oxidation state.



In this reaction, Cu begins in the elemental form with a charge of zero and changes into  $\text{Cu}^{2+}$ . In this process, copper is oxidized, losing two electrons. This is indicated in the following half-reaction, which only shows the transformation of the copper species.



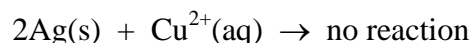
In contrast, Ag begins as  $\text{Ag}^+$  and changes into the elemental form, Ag, with a charge of zero. Therefore, silver is reduced, each silver ion gaining one electron, as shown in this half-reaction.



### REFERENCE

Eaker, C.W.; Eisenmann, M.A., *Electron Transfer Between Metals*. Presented at the 13<sup>th</sup> Annual T<sup>3</sup> International Conference, Columbus, OH, March, 2001.

In this redox reaction, copper is the more active metal because it loses electrons and silver is the less active metal because it gains electrons. Copper spontaneously transfers electrons to silver ions. Copper ions, however, will not spontaneously accept electrons from silver. This indicates that electrons will spontaneously flow in only one direction and that the reverse reaction will not occur of its own accord.



In electrochemical cells, such as batteries, the more active metal is the negative electrode and the less active metal is the positive electrode. The electrons flow from the negative electrode to the positive electrode, as shown in Figure 1. A salt bridge completes the electrical circuit, maintaining a balance in ionic charge as ions are added or removed during the chemical reaction. The **voltage** measured between the electrodes of an electrochemical cell indicates the relative ability of the metals to gain or lose electrons.

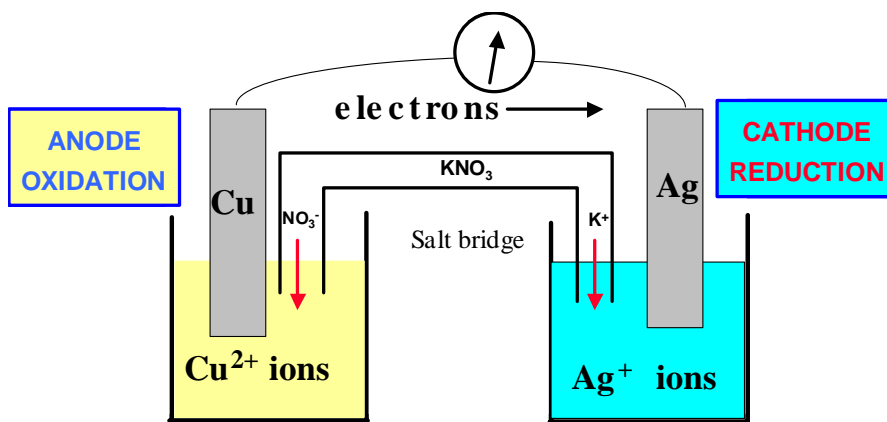


Figure 1

This experiment will allow you to investigate the activity of a variety of metals, to measure the voltage of a selection of batteries, and to construct half-cells and measure the voltage of electrons between them.

## PURPOSE

The purpose of this experiment is to study the electron transfer properties of metals and how those properties enable the production of electricity in electrochemical cells.

## EQUIPMENT/MATERIALS

LabPro	New and used batteries (AA, C, D, 9-volt)
TI Graphing Calculator	24- well culture plate
Voltage Probe with 2 paperclips	0.5 x 3.0 cm strips Cu, Pb,Zn,and Mg
5.0 cm pieces of string soaked in 1 M KNO <sub>3</sub> (or NaNO <sub>3</sub> )	30 mL dropper bottles of 1 M Cu <sup>2+</sup> , Pb <sup>2+</sup> , Zn <sup>2+</sup> , and Mg <sup>2+</sup> solutions
Sand paper or steel wool	

## SAFETY

- Always wear goggles and an apron in the lab.

## PROCEDURE

### Part I: Relative Activity of Metals

1. Obtain a well plate. Sand pieces of Cu, Pb, Zn, and Mg to make sure they are clean.
2. Put 5 drops of a 1.0 M Cu<sup>2+</sup> solution in the first four wells of row 1 of the well plate.
3. Repeat for the Pb<sup>2+</sup>, Zn<sup>2+</sup>, and Mg<sup>2+</sup> solutions in rows 2,3, and 4, respectively.
4. Place the Cu metal strips in the wells in first column of the well plate.
5. Repeat for Pb, Zn, and Mg in columns 2,3, and 4, respectively.
6. Observe any changes, such as a darkening of the metal or the solution, that suggest a chemical reaction occurred.
7. Record in the Data Table whether a reaction takes place (a darkening is considered a chemical reaction).

### Part II: Battery Voltage

1. Plug the voltage probe into Channel 1 of the LabPro. Be sure that the link cable attaching the graphing calculator to the LabPro is firmly inserted into each unit.
2. Turn on the calculator. Start the DataMate program. Press [CLEAR] on the calculator to reset the program. The calculator should be displaying CH1: voltage. If not, tell your instructor.
3. Press [1] to select SETUP from the menu at the bottom of the screen (wait patiently).
4. Move the arrow cursor to MODE (to get to MODE quickly, arrow up). At MODE, press [ENTER] (wait patiently).
5. Select 4: SINGLE POINT. This will average several readings over a 10-sec time period and give you one value.
6. Press [1] for OK.

7. Touch the red voltage probe to the + pole of the battery and the black probe to the – pole. For the AA, C, and D batteries you may first want to attach a paper clip to each of the voltage probe leads to make it easier to make contact with the ends of the batteries. Press [2] to start the SINGLE POINT measurement. Be sure to keep the probes in contact with the battery for the entire 10-sec period.
8. Use your voltage probe to measure the voltage of at least 2 batteries. Record what you measured and what voltage is written on the side of the battery. You should also reverse the leads to see how this affects the measured voltage.

### Part III: Voltage Potentials between Half-Cells

1. Using the well plate, half fill (~2mL) four adjacent wells (see Figure 2) with the 1.0 M solutions of each of the metal compounds. Add a freshly cleaned, shiny (sanded) piece of the metal of the same type (e.g. Cu in the  $\text{Cu}^{2+}$  solution, Pb in the  $\text{Pb}^{2+}$  solution, etc.) that can stand in the well and can accept the clip of the voltage probe. These are the electrodes.
2. Insert a salt bridge (a piece of string that has been soaked overnight in  $\text{KNO}_3$ ) between the wells. This string should be wet and should have each end below the surface of the solutions.
3. Clip the red voltage probe to the top of the Cu electrode and clip the black probe to the Pb electrode. Measure the voltage and record your results in the data table.
4. Next, move the black probe to the Zn electrode and measure the voltage. Repeat for Mg.

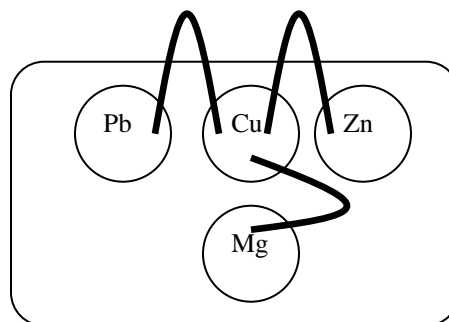


Figure 2

**DATA SHEET**

Name \_\_\_\_\_

Name \_\_\_\_\_

Period \_\_\_\_\_ Class \_\_\_\_\_

Date \_\_\_\_\_

**ELECTRON TRANSFER BETWEEN METALS**

**DATA TABLES**

**PART I**

<b>Metals</b>	<b>Observations</b>	<b>Reaction (Y or N)</b>
<b>Cu/Cu<sup>2+</sup></b>		
<b>Cu/Pb<sup>2+</sup></b>		
<b>Cu/Zn<sup>2+</sup></b>		
<b>Cu/Mg<sup>2+</sup></b>		
<b>Pb/Cu<sup>2+</sup></b>		
<b>Pb/Pb<sup>2+</sup></b>		
<b>Pb/Zn<sup>2+</sup></b>		
<b>Pb/Mg<sup>2+</sup></b>		
<b>Zn/Cu<sup>2+</sup></b>		
<b>Zn/Pb<sup>2+</sup></b>		
<b>Zn/Zn<sup>2+</sup></b>		
<b>Zn/Mg<sup>2+</sup></b>		
<b>Mg/Cu<sup>2+</sup></b>		
<b>Mg/Pb<sup>2+</sup></b>		
<b>Mg/Zn<sup>2+</sup></b>		
<b>Mg/Mg<sup>2+</sup></b>		



