

# Electrochemistry and the Nernst Equation (#4.1)

## The CCLI Initiative Computers in Chemistry Laboratory Instruction

### Learning Objectives

The objectives of this experiment are to...

- construct galvanic cells and develop an electrochemical series based on potential differences between half-cells.
- Measure and understand the origin of the Nernst Equation.

### Background

Any chemical reaction involving the transfer of electrons from one substance to another is an oxidation-reduction (redox) reaction. The substance losing electrons is **oxidized** and the substance gaining electrons is **reduced**. Let us consider the following redox reaction:



This redox reaction can be divided into an oxidation and a reduction half-reaction:



A **galvanic cell** (Figure 1) is a device used to separate a redox reaction into its two component half-reactions in such a way that the electrons are transferred through an external circuit rather than by direct contact of the oxidizing agent and the reducing agent. This transfer of electrons through an external circuit is electricity.

Each side of the galvanic cell is known as a **half-cell**. Each half-cell consists of an electrode (the metal of the half-reaction) and a solution containing the corresponding cation of the half-reaction. The electrodes of the half-cells are connected by a wire for electron flow and a **salt bridge** for ionic flow to maintain electrical neutrality. The electrode at which oxidation occurs is called anode and the electrode at which reduction occurs is called the **cathode**.

Different metals, such as zinc and lead, have different tendencies to oxidize; similarly their ions have different tendencies to be reduced. The cell potential is due to the difference in these tendencies of the two metals. The cell voltage measured is the result of the two half-reactions, and the magnitude of the potential depends on the concentrations of the ions, the temperature, and pressure of gases. Standard conditions are when the metal ion concentration is 1 molar, the temperature is 25 °C, and the pressure of any gas is 1 atmosphere then the half-cell voltage is the **standard reduction voltage**. We pick one half reaction, call it the standard, construct a cell, measure the cell voltage and report the potential relative to the standard. The standard that has been chosen by convention is:



Where  $\mathcal{E}^\circ$  is the standard electrode potential and is the potential of the measured cell versus the standard hydrogen electrode.

Tables of standard half-reaction potentials have been computed. The reactions by convention are written as reductions and hence the tables are called tables of standard reduction potentials.

### The Nernst Equation

Standard conditions are often difficult if not impossible to obtain. The **Nernst Equation** allows cell voltages to be predicted when the conditions are not standard. Walter Nernst developed the following equation the late 1800's while studying the thermodynamics of electrolyte solutions:

$$\mathcal{E}_{\text{cell}} = \mathcal{E}^\circ_{\text{cell}} - (2.303 \text{ RT/n}\mathcal{F}) \log Q \quad (1)$$

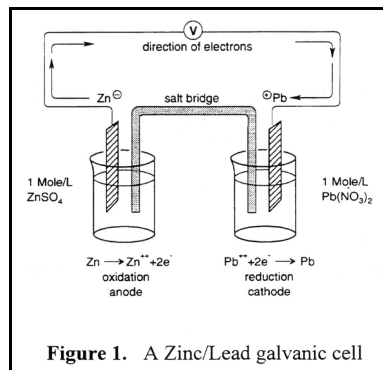


Figure 1. A Zinc/Lead galvanic cell

In equation (1),  $R$  is the gas constant ( $8.314 \text{ J mole}^{-1}\text{K}^{-1}$ ),  $T$  is the temperature (Kelvin),  $\mathcal{F}$  is Faraday's constant ( $96,485 \text{ coulombs/mole}$ ),  $n$  is the number of electrons transferred in the balanced oxidation/reduction reaction, and  $Q$  is the reaction quotient, or ( $[\text{products}]/[\text{reactants}]$ ). If the reactions are carried out at room temperature ( $25 \text{ }^\circ\text{C}$ ), the Nernst equation becomes

$$\mathcal{E}_{\text{cell}} = \mathcal{E}_{\text{cell}}^{\circ} - (0.05916/n) \log Q \quad (2)$$

### Experiments Performed

1. Metal/ion half-cell combinations will be tried for the various combinations of  $0.1 \text{ M}$  solutions of  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{ZnSO}_4$ ,  $\text{Fe}^{2+}/\text{Fe}^{3+}$ ,  $0.1 \text{ M}$   $\text{SnCl}_2$ , with and  $0.1 \text{ M}$  solution of  $0.1 \text{ M}$   $\text{Cu}(\text{NO}_3)_2$  as the salt bridge, the listing various elements and ions in order of their tendency to gain or lose electrons.
2. The Nernst Equation: The Nernst equation will be derived from a study of five solutions of  $\text{CuSO}_4$  ( $1.00$ ,  $0.100$ ,  $0.0100$ ,  $0.0010$ , and  $0.0001 \text{ M}$ ) referenced to a  $1.00 \text{ M}$   $\text{Zn}/\text{ZnSO}_4$  half cell.

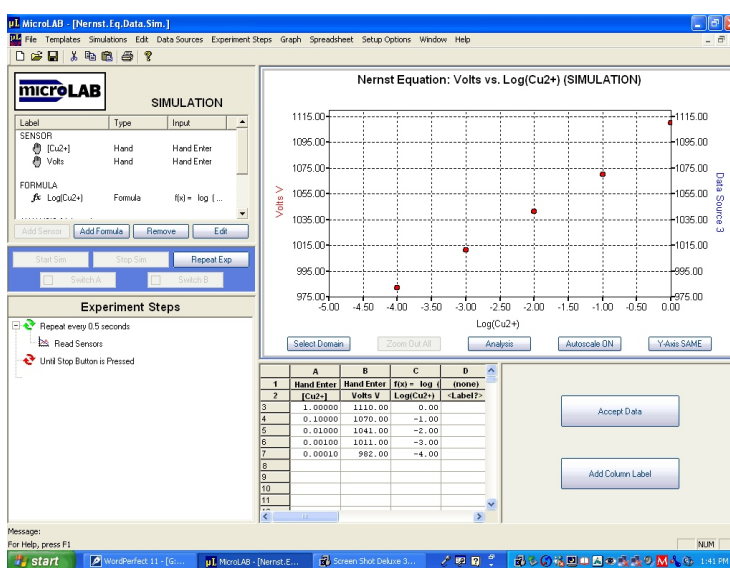
### Data Analysis

Guidance is given the student to help analyze the data to obtain the desired results.

Metal-ion Half-cell Table.

Sample	Red:	Black	Voltage(
1	Cu	Sn	610
2	Zn	Sn	-460
3	Fe	Sn	960
4	Fe	Cu	330
5	Fe	Zn	1300
6	Cu	Zn	1050

Graph of voltage versus  $\log[\text{Cu}^{2+}]$  showing slope of  $31.5 \text{ mV}/\log[\text{Cu}^{2+}]$



### Instructor Resources Provided

1. Sample Report Sheets providing the format to organize the data collection with sample data.
2. Questions to consider, answer and turn-in with suggested answers.
3. Tips and Traps section to assist the instructor with potential problems and solutions.
4. Sample *MicroLAB* screen shots and graphs.
5. Laboratory preparation per student station.

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