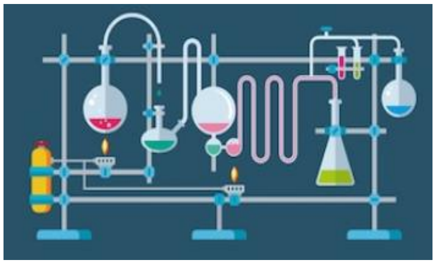

430 chem Practical

اعداد

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Experiment # 1

Title of experiment: Nernst Equation

Aim:

1- To verify Nernst Equation.
2- To calculate the concentration of an un unknown [Cu^{2+}] by constructing a Galvanic Cell

Introduction:

Chemical reactions involving the transfer of electrons from one reactant to

another are called oxidation-reduction reactions or redox reactions. In a redox reaction, two half-reactions occur; one reactant gives up electrons (undergoes oxidation) and another reactant gains electrons (undergoes reduction). When immersed in water, a piece of zinc going into a solution as zinc ions, with each Zn atom giving up 2 electrons. This is an example of an oxidation half-reaction.

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e \qquad (1)$$

Another example of reduction is the formation of solid copper from copper ions in solution.

$$Cu^{2+}(aq) + 2e \rightarrow + Cu(s) \qquad (2)$$

The redox reaction results when an oxidation and a reduction half-reaction are

combined to complete the transfer of electrons as in the following example:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s) \qquad (3)$$

A measure of the tendency for a reduction to occur is its reduction potential, E, measured in units of volts. Standard reduction potentials have been measured for many half-reactions and they are listed in tables.

A galvanic cell or voltaic cell is a device in which a redox reaction, such as the one in equation (3), occurs spontaneously and produces an electric current. In order for the transfer of electrons in a redox reaction to produce an electric current and be useful, the electrons are made to pass through an external conducting wire instead of being directly transferred between the oxidizing and reducing agents. The design of a galvanic cell (shown in Figure 1 for equation (3) reaction) allows this to occur. In a galvanic cell, two solutions, one containing the ions of the oxidation half-reaction and the other containing the ions of the reduction half-reaction, are placed in separated compartments called half-cells.

For each half-cell, the metal, which is called an electrode, is placed in the solution

connected to an external wire. The electrode at which oxidation occurs is called the and anode[Zn in equation (3)] and the electrode at which reduction occurs is called the cathode [Cu in equation (3)].

The two half-cells are connected by a salt-bridge which allows a “current” of

Electrons from one half-cell to the other to complete the circuit of electron current in the external wires. When the two electrodes are connected to an electric load (such as a light bulb or voltmeter) the circuit is completed, the oxidation-reduction reaction occurs, and electrons move from the anode (−) to the cathode (+), producing an electric current.

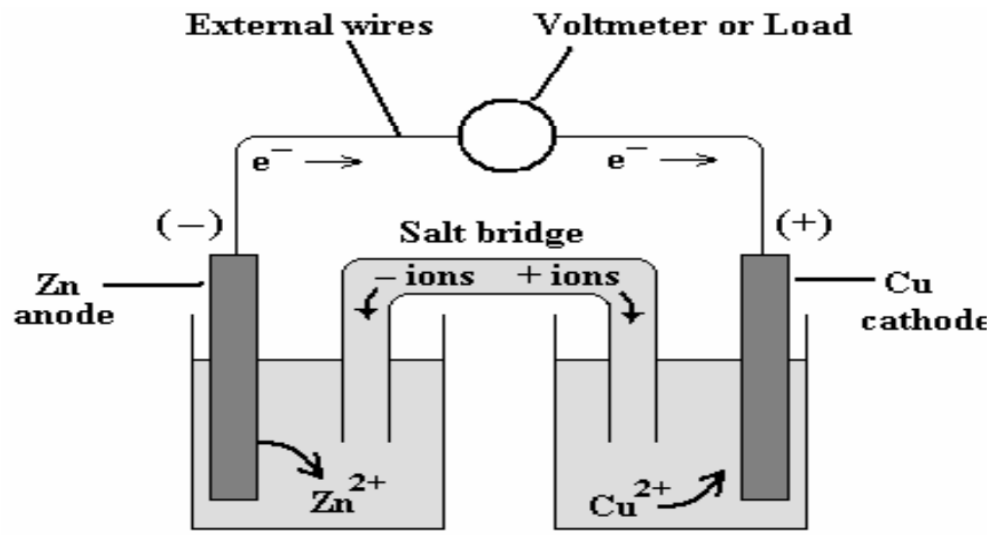

Figure 1. Galvanic cell (or battery) based on the redox reaction in equation (3).

The cell potential, E_{ce}ll, is a measure of the voltage that the battery can provide. E_{cell} is calculated from the half-cell reduction potentials: $E_{cell} = E_{\ cathode} – E_{\ anode}$

At standard conditions, the standard cell potential, $E°_{cell}$, is based upon the standard reduction potentials, as shown in equation (4).

$$E°_{cell} = E°_{cathode} – E°_{anode} \qquad (4)$$

Based on the values for the standard reduction potentials for the two half-cells in equation (4)

[+0.34 V for copper cathode and –0.76 V for zinc anode and], the standard cell potential, E°cell, for the galvanic cell in Figure 1 would be:

$$E°cell = +0.34\ V – (–0.76\ V) = +1.10\ V$$

The positive value of the voltage for E° cell indicates that at standard conditions the reaction is spontaneous. Recall that $\Delta G^o = - nFE^o$

cell, so that a positive $E^o$of a cell results in a negative ΔG^o. Thus the redox reaction in equation (3) would produce an electric current when it is set up as a galvanic cell.

When conditions are not standard, the Nernst equation (5), is used to calculate the

potential of a cell. In the Nernst equation, R is the universal gas constant with a value of 8.314 J/(K·mol), T is the temperature in K, . F is the Faraday constant with a known value of 96,500 J/(V·mol) and n is the number of electrons transferred in the redox reaction, for example, 2 electrons in equation (3). Q is the reaction quotient for the [ion products]/[ion reactants] of the cell.

$$E_{\mathrm{cell}} = E^{0}_{\mathrm{cell}} - \left(\frac{RT}{nF}\right)(\ln Q) \qquad (5)$$

For our equation (3) example, $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

$$Q = \frac{[\mathrm{Zn}^{2+}]}{[\mathrm{Cu}^{2+}]} \qquad \text{and} \qquad \ln Q = \ln\,[\mathrm{Zn}^{2+}] - \ln\,[\mathrm{Cu}^{2+}]$$

For Nernst equation in equation (3), n= 2,and this redox reaction becomes:

$$E_{\mathrm{cell}} = E^{0}_{\mathrm{cell}} - \left(\frac{RT}{2F}\right)\ln[\mathrm{Zn}^{2+}] + \left(\frac{RT}{2F}\right)\ln[\mathrm{Cu}^{2+}] \qquad (6)$$

Thus we perform a series of galvanic cells, in which $[Zn^{2+}]$ is kept constant while $[Cu^{2+}]$ is varied, E_{cell} can be measured and it will be found to vary with $\ln[Cu^{2+}]$.

A plot of the data obtained in which y is E_{cell} and x is $\ln[Cu^{2+}]$ will result in a straight line: y = mx + b. For equation (6), the terms E°cell and −[RT/2F]ln $[Zn^{2+}]$ are constant and together they equal the intercept, b, of the line. [RT/2F] will be the constant slope, m, provided the temperature is constant.

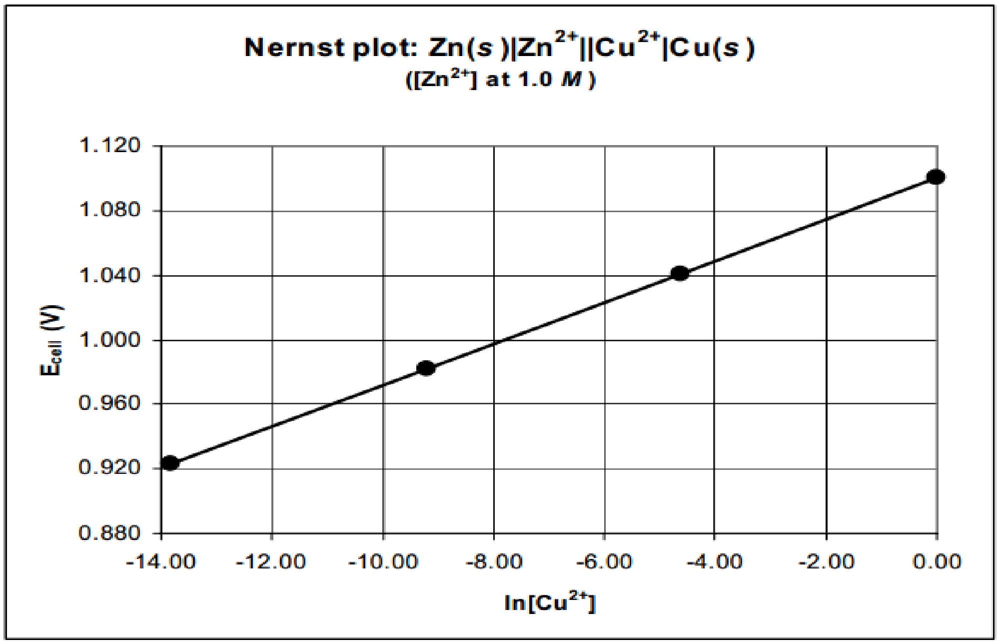

Figure 2. Nernst plot of E_{cell} vs. ln $[Cu^{2+}]$ with $[Zn^{2+}]$ constant at 1.0 *M*. Note the standard cell notation in the graph title for the galvanic cell.

Materials:

Volt meter (pH meter). 50 mL Beakers (5) ,two connection leads with Alligator clips , Sand paper

Zinc sheet (or electrode), solution of $Zn(NO_3)_2$ =1 M.

Copper strips (or electrode) , Solutions of $[Cu^{2+}]$ 1M, 0.1M,0.01M,0.001M, and unknown solution of $[Cu^{2+}]$, salt bridge filled with 3M KNO_3

Procedure:

1- Clean the Zinc with sand paper then with tab water followed with distilled water and dry. The Cu electrode is washed with immersing in 2M KCl for 1minute then with distilled water several times then dried.
2- Prepare the Cu half- cell by immersing the Cu sheet (or Cu electrode) in a beaker filled with 30 mL solution of $[Cu^{2+}]$ =0.001M.
3- Prepare the Zn half- cell by immersing the Zinc sheet in a beaker filled with 30 mL of 1.0 M $Zn(NO_3)_2$ next to the Cu half-cell. Connect the two half-cells with a freshly prepared salt bridge. Connect the copper and zinc electrodes to the correct voltage probe leads in the pH meter adjusted to read potential.
4- Measure and record the cell potential in your laboratory notebook using the same technique (5-10 second immersion) with the voltage probe

Repeat the previous procedure by changing the concentration of Solutions of $[Cu^{2+}]$ in order of increasing concentration. Then, in the same way, measure the E_{cell} for the unknown Cu^{2+}

to 0.01M, 0.1M,1 M, and the unknown solution of $[Cu^{2+}]$. Be careful to wash the Cu electrode each time with distilled water and dry.

Results:

1. Record the measured voltage values in a table as shown below in your notebook.

$[Cu^{2+}]$	1	2	3	4	5	Unknown
solution	0.0001M	0.001M	0.01M	0.1M	1.0 M	??
Voltage (V)						
Notes						

2. Plot the Nernst equation, E_{cell} vs $\ln[Cu^{2+}]$.
3. Record the voltage value of an unknown solution of $[Cu^{2+}]$.
4. Use the data to find from the constructed graph of $[Cu^{2+}]$ the concentration of the unknown solution

Reference:

1. Advanced Chemistry with Vernier: Experiments for AP, IB, and College General Chemistry, Jack Randall, Vernier Software and Technology,2004, 20-1UCCS Chem 106 Laboratory Manual