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Department of Mechanical and Electromechanical Engineering Process Engineering 2nd year

**Solution Chemistry practical’s Works**



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**Experiment 5 : Experimental verification of Nernst’s law**

We will learn in this experiment how construct electrochemical cells; identify anode and cathode and explain what occurs at each (oxidation or reduction), and measure E0cell for different pairs of metals and metal ion solutions; compare this value by that calculated according to Nernst's law.

# introduction

An oxidation-reduction or redox reaction is a chemical reaction in which one or more electrons from one molecule or atom are transferred to another. Thermodynamics can predict if electrons would prefer to be transferred from one species to another based on the free energy change of the system.

**Δ*G* = −*nF*Δ*E*, or Δ*E* = −Δ*G*/*nF***

n: The number of moles of electrons transferred. F is Faraday's constant ***F* = 96485coulomb mole-1**

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A process will spontaneously occur if ΔG < 0. this implies spontaneity is associated with a positive **Δ**E.

 The Nernst equation includes a term that reflects the potential difference between products and reactants under standard conditions (E°) and a term to account for non-standard conditions

***E* = *E*° − *RT/ nF* ln*****Q* = *E*° − *0,0591/ nF* ln*****Q***

E is the voltage measured for the cell. E° is the voltage measured when *Q* = 1.

***E*°(cell) = *E*°1/2(reduction) − *E*°1/2(oxidation) = *E*°1/2(cathode) − *E*°1/2(anode)**

# reagents

- Solutions of 0.1 M CuSO4, 0.1 M Zn(NO3)2, 0.1 M Pb(NO3)2, 0.1 M FeSO4 and 0.1 M KNO3

# Procedure:

Zn2+/Zn (*E*° = −0.76); Pb2+ / Pb (*E*° = −0.13); Cu2+ / Cu (*E*° = +0.34); Fe2+/Fe (*E*° = -0,44).

**Part 1: The Zn-Cu Redox Reaction**

1. Clean the Zn metal with sandpaper.

2. Place the clean metal in a small test tube.

3. Fill the test tube with 0.1M CuSO4 solution is sure that the Zn metal is completely submerged.

 4. After 2.5 minutes, record your observations.

5. After 5 minutes, record your observations.

6. Clean the Zn metal strip with sandpaper and dry using paper towels and return to the container.

7. Retain the CuSO4 solution for use in Part 2.

**Part 2: The Pb-Cu Redox** **Reaction**

1. Clean the Pb metal with sandpaper.

 2. Place the clean metal in a small test tube.

3. Fill the test tube with 0.1M CuSO4 solution; be sure that the Pb metal is completely submerged. 4. After 2.5 minutes, record your observations.

 5. After 5 minutes, record your observations.

 6. Clean the Pb metal with sandpaper and dry using paper towels and return to the container

7. Pour the CuSO4 solution into waste beaker.

 **Part 3: The Zn-Pb Redox Reaction**

1. Clean the Zn metal with sandpaper.

2. Place the clean metal in a small test tube.

 3. Fill the test tube with 0.1M Pb(NO3)2 solution be sure that the Zn metal is completely submerged.

 4. After 2.5 minutes, record your observations.

5. After 5 minutes, record your observations.

6. Clean the Zn metal with sandpaper and dry using paper towels and return to the container

7. Discard the Pb(NO3)2 solution into waste beaker

**Part 4: Electrochemical Half-Cell Reactions**

1. Dispense 10 mL of solutions 0.1 M CuSO4, 0.1 M Zn(NO3)2, 0.1 M Pb(NO3)2, 0.1 M FeSO4 and 0.1 M KNO3 into beakers of 30 mL.

2. Clean the copper, zinc, lead, and iron electrodes using sandpaper and rinse with deionized water. 3. Place each metal electrode in its corresponding ionic solution; e.g. copper strip goes into the CuSO4 solution. It is important that the correct metal is in the correct solution or your cell will not work properly.

4. Obtain small strips of filter paper to be used as salt bridges. Completely wet one strip in the beaker containing 0.1 M KNO3.

5. Carefully remove the completely wet strip and place one end in the CuSO4 solution and the other in the Zn(NO3)2 solution. The salt bridge should not touch the electrodes.

 6. Attach one alligator clip from the positive terminal on the voltmeter to the Cu electrode and the second clip on the negative terminal to the Zn electrode.

7. Record the voltage of the electrochemical cell. Record the copper as the cathode and the zinc as the anode.

8. Repeat for the remaining cells in the data table, recording the positive cell voltages and the anode (- terminal) and cathode (+ terminal) for each cell. Use a new wet piece of filter paper as a salt bridge for each determination.

9. Clean the metal strips with sandpaper and dry each strip using paper towels.

10. Dispose of all solutions into the appropriate waste container.

11. Return all metal pieces to their original container clean and dry.

12. Clean and dry your work area with water. Wash all glassware with soap then rinse 3 times with tap water, and once with distilled water.

# QUESTIONS :

1. Give the principle of this experiment
2. give the pictograms of the products used
3. identify which metal is reduced and which is oxidized based on your experimental parameters. Write the oxidation reaction, the reduction reaction and the overall cell reaction for each of the metal combinations for parts 1 through 3.
4. Part 4: -First, identify which metal is reduced and which is oxidized based on your identification of the cathode and anode on your data sheet.
* Reduction takes place at the cathode and oxidation at the anode. Write the overall cell reaction for each of the metal combinations for part4.
* Calculating Cell Potentials for Part 4 Locate both metals in the reduction potential table. Identify which of the two metals is oxidized.
* Then, calculate the cell voltage E°cell, and the Percent Difference between the Calculated Cell Potential and the Measured Cell Potential.
1. Calculated the equilibrium constant for each reaction.