Abstract

In this lab, two things were achieved: a galvanic cell was made from different mixed metal components, allowing students to determine the identity of an unknown metal, and the value of the Nernst constant was calculated through experimental data from a concentration cell. A galvanic cell with 1.5M Copper Nitrate solution at the cathode and a 1.5M Zinc Nitrate solution at the anode was made. A copper wire was used for both electrodes. A measured electrochemical cell potential for this system was found to be -0.61V, giving a 19.7% error when compared against the theoretical. A concentration cell with a Copper Nitrate solution at the anode was made. A copper wire was used for both electrodes. A measured electrochemical cell potential for this system was found to be -7.69% error when compared against the theoretical. A concentration cell using 1.5M Copper Nitrate aqueous solution and 1.5M Lead Nitrate aqueous was also made, resulting in a 7.69% error when compared against the theoretical.

Introduction/ Background

\[ E_{\text{cell}} = E^\circ - (RT/nF) \ln Q \]

Equation 1. The Nernst Equation. RT/F = 0.0592

\[ \Delta G^\circ = nF \Delta E^\circ \]

Equation 2. Gibbs Free Energy Equation

\[ \Delta G = \Delta G^\circ + RT \ln Q \]

Equation 3. Gibbs Free Energy Equation (a) standard and (b) non-standard conditions

\[ E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ \]

Equation 4. Concentration cell data of \( \log[\text{Cu}^{2+}\text{ox}/\text{Cu}^{2+}\text{red}] \) measured \( E_{\text{cell}} \) for a copper concentration cell

Results & Discussion

<table>
<thead>
<tr>
<th>Galvanic Cell Type (with a copper cathode)</th>
<th>Measured ( E_{\text{cell}} ) (V)</th>
<th>Accepted ( E_{\text{cell}} ) from standard reduction potential (V)</th>
<th>Calculated ( E_{\text{cathode}} ) (V)</th>
<th>Accepted ( E_{\text{cathode}} ) from standard reduction potential (V)</th>
<th>% Error for ( E_{\text{cathode}} ) (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn⁺</td>
<td>0.95</td>
<td>0.34</td>
<td>-0.12</td>
<td>-0.16</td>
<td>-19.7</td>
</tr>
<tr>
<td>Pb²⁺</td>
<td>0.46</td>
<td>0.34</td>
<td>-0.12</td>
<td>-0.13</td>
<td>-7.69</td>
</tr>
<tr>
<td>Fe³⁺</td>
<td>-0.47</td>
<td>0.34</td>
<td>0.81</td>
<td>0.77</td>
<td>5.19</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>1.58</td>
<td>0.34</td>
<td>-1.24</td>
<td>-2.37</td>
<td>-47.7</td>
</tr>
<tr>
<td>Sn²⁺</td>
<td>0.53</td>
<td>0.34</td>
<td>-0.19</td>
<td>-0.14</td>
<td>35.7</td>
</tr>
</tbody>
</table>

Table 1. Galvanic Cell data with measured \( E_{\text{cell}} \) values, Accepted \( E_{\text{cell}} \) values, Calculated \( E_{\text{cathode}} \) values, and % error between the accepted and calculated \( E_{\text{cathode}} \) values.

Concentration Cell

Experimental:

Figure 4. Concentration cell data of \( \log[\text{Cu}^{2+}\text{ox}/\text{Cu}^{2+}\text{red}] \) measured \( E_{\text{cell}} \) for a copper concentration cell

Theoretical (n=2):

\[ E_{\text{cell}} = 0 - (0.0296) \ln Q \]

Equation 5. \( E_{\text{cell}} \) = 0 - (0.0296) \ln Q

Percent Error: -28.96%

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References

