THE CONVERSION OF CHEMICAL ENERGY INTO ELECTRICAL ENERGY IN GALVANIC CELLS

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Abstract

Galvanic cells are electrochemical systems that convert chemical energy into electrical energy through redox reactions. This article explores the principles underlying galvanic cell operation, electrochemical potentials, electrode polarization phenomena, and applications of the Nernst equation for potential calculations. The Daniel-Jacobi and hydrogen-based galvanic cells are analyzed as examples, highlighting the processes of oxidation at the anode and reduction at the cathode. The impact of ion concentrations, temperature, and environmental factors on the electromotive force (EMF) is also discussed.

Keywords

Galvanic cell, redox reaction, electrode potential, Nernst equation, polarization, electromotive force (EMF)

1. Introduction

The transformation of chemical energy into electrical energy is fundamental in electrochemical power sources. Devices that perform this transformation using redox reactions are called galvanic cells. These systems rely on the transfer of electrons from a more reactive metal (anode) to a less reactive one (cathode) through an external circuit, while ionic migration occurs through an electrolyte or salt bridge.

The first galvanic cell, developed by Italian scientist Alessandro Volta, involved copper and zinc electrodes immersed in sulfuric acid. The observation of electrical current in this setup led to the creation of the Voltaic pile, the precursor of modern batteries.

2. Materials and Methods

To construct a galvanic cell, two different metal electrodes (e.g., zinc and copper) are immersed in their respective salt solutions (e.g., $ZnSO_4$ and $CuSO_4$). The metals are connected externally via a conductive wire, and the solutions are connected internally by a salt bridge or a porous membrane to allow ion exchange.

Example: Daniel-Jacobi Cell

This classic galvanic cell involves:

Anode (oxidation): $Zn \rightarrow Zn^{2+} + 2e^-$

Cathode (reduction): $Cu^{2+} + 2e^- \rightarrow Cu$

The overall reaction is:

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

Cell notation: (-) $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$ (+)

(-) $Zn \rightarrow Zn^{2+} + 2e^-$

Salt Bridge ||

$Cu^{2+} + 2e^- \rightarrow Cu$ (+)

3. Results and Discussion

3.1. Electromotive Force (EMF)

The EMF of a galvanic cell is calculated by subtracting the anode potential from the cathode potential:

$$EMF = E°(cathode) - E°(anode)$$

For the Daniel cell under standard conditions:

$$E°(Cu^{2+}/Cu) = +0.34\ V$$

$$E°(Zn^{2+}/Zn) = -0.76\ V$$

$$EMF = 0.34 - (-0.76) = 1.10\ V$$

3.2. Nernst Equation

When ion concentrations deviate from standard 1 mol/L, the Nernst equation is used to determine actual electrode potentials:

$$E = E° + (RT/nF) \ln C$$

At 25°C (298 K), this simplifies to:

$$E = E° + (0.059/n) \log C$$

3.3. Polarization in Galvanic Cells

During continuous operation, electrode potentials may deviate due to accumulation of reaction products and depletion of reactants. This phenomenon is called polarization and results in a decrease in the EMF.

4. Case Study: Ammonia Complex Formation

In an ammonia solution, copper can be oxidized by atmospheric oxygen forming the complex ion:

$$Cu + 8NH_3 + O_2 + 2H_2O \rightarrow 2[Cu(NH_3)_4]^{2+} + 4OH^-$$

The corresponding galvanic cell is:

$$(-)\ Cu \mid [Cu(NH_3)_4]^{2+} \parallel O_2 \mid OH^-\ (+)$$

Using standard potentials:

$$E°(O_2/OH^-) = +0.40\ V$$

$$E°([Cu(NH_3)_4]^{2+}/Cu) = -0.05\ V$$

$$EMF = 0.40 - (-0.05) = +0.45\ V$$

5. Conclusion

Galvanic cells represent a vital application of redox chemistry, converting chemical energy into electrical energy. Their efficiency depends on electrode material, ion concentration, and the minimization of polarization. Understanding and optimizing these variables, including through the Nernst equation and the use of complexing agents or oxidizers, is crucial for designing efficient electrochemical systems for power generation and sensing applications.

References

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