Given That E Red = -0.23 V For Ni 2+/Ni At 25C, Find E And E For The Concentration Cell Expressed Using

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Understanding electrochemical cells and their potentials is fundamental in the field of chemistry, especially when analyzing the behavior of different metal ions and their respective electrodes. The given reduction potential of nickel (Ni) ions provides a foundation to explore how concentration variations influence cell potentials and how to calculate the electromotive force (EMF) of concentration cells. This article offers a comprehensive guide to determining the cell potential (E) and the standard electrode potential (E°) for a concentration cell based on the given data, emphasizing the importance of such calculations in both academic and industrial applications.

Introduction to Electrochemical Cells

Electrochemical cells are devices that convert chemical energy into electrical energy through redox reactions. These cells consist of two electrodes—an anode and a cathode—immersed in electrolytes containing ions of the respective metals. The potential difference between these electrodes drives the flow of electrons, generating an electric current.

There are two primary types of electrochemical cells:

- Voltaic (Galvanic) Cells: Generate electricity spontaneously.
- Electrolytic Cells: Require external energy to drive non-spontaneous reactions.

In this context, we are focusing on a galvanic cell involving nickel ions and metal nickel electrodes, where the cell potential depends on the concentration of ions in solution.

Standard Reduction Potential of Ni^{2+}/Ni

The standard reduction potential (E°) indicates the tendency of a species to gain electrons (be reduced) under standard conditions (1 M concentration, 25°C, 1 atm pressure). For nickel:

- Given Data: $E°_{(Ni^{2+}/Ni)}$ = -0.23 V at 25°C

This value serves as a reference point for calculating the cell potential under non-standard conditions, considering the actual ion concentrations.

Understanding Concentration Cells

A concentration cell is a type of galvanic cell where both electrodes are made of the same material, but the electrolyte solutions differ in concentration. The potential difference arises solely due to the concentration gradient of the ions, and the cell potential can be calculated using the Nernst equation.

Key features of concentration cells:

- Both electrodes are identical (e.g., Ni/Ni).
- The cell potential depends on the difference in ion concentrations.
- The overall cell potential is zero when concentrations are equal.

Fundamental Equations for Cell Potential Calculations

To analyze and compute the cell potential, we use the Nernst equation, which relates the electrode potential to ion concentrations:

The Nernst Equation

$$
E = E^{\circ} - \frac{RT}{nF} \ln Q
$$

Where:

- E: Cell potential under actual conditions
- E°: Standard electrode potential
- R: Universal gas constant (8.314 J/mol·K)
- T: Temperature in Kelvin (25°C = 298K)
- n: Number of electrons transferred in the redox reaction
- F: Faraday's constant (96485 C/mol)
- Q: Reaction quotient, expressed in terms of ion concentrations

For simplicity, at 25°C, the equation is often expressed as:

$$
E = E^{\circ} - \frac{0.0592}{n} \log Q
$$

Note: When dealing with concentration cells where both electrodes are identical, the cell potential simplifies to:

$$
E_{cell} = \frac{0.0592}{n} \log \frac{a_{oxidant}}{a_{reductant}}
$$

where $a_{oxidant}$ and $a_{reductant}$ are activities (approximated by concentrations for dilute solutions).

Step-by-Step Calculation of Cell Potential (E)

Suppose the concentration of Ni^{2+} ions in the two half-cells is different, with the following concentrations:

- Concentration at the anode (more dilute): $[Ni^{2+}]_{anode} = c_1$
- Concentration at the cathode (more concentrated): $[Ni^{2+}]_{cathode} = c_2$

The reaction involved:

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

- Number of electrons transferred, $n = 2$

Calculating the Cell Potential:

$$
E_{cell} = \frac{0.0592}{2} \log \frac{c_{cathode}}{c_{anode}} = 0.0296 \log \frac{c_2}{c_1}
$$

Example:

If $c_2 = 1\, M$ and $c_1 = 0.01\, M$, then:

$$
E_{cell} = 0.0296 \log \frac{1}{0.01} = 0.0296 \times 2 = 0.0592\, V
$$

This potential indicates the cell's EMF based on the concentration difference.

Important points:

- The sign of E_{cell} indicates the direction of electron flow.
- The more concentrated solution acts as the cathode, where reduction occurs.
- As the concentrations approach equality, the cell potential approaches zero.

Calculating the Standard Electrode Potential (E°) for the Cell

In a concentration cell, the standard electrode potential E° is zero because the electrodes are identical. However, if the cell involves different species or an external standard, the overall cell potential relates to the standard potentials of the electrodes.

Given the data:

- $E_{Red} = -0.23\, V$ for Ni^{2+}/Ni at 25°C

The standard reduction potential for Ni^{2+}/Ni is already provided, so:

$$
E^{\circ} = -0.23\, V
$$

If the question seeks to find the standard potential of the entire cell based on the measured potential under specific concentrations, the Nernst equation adjustments are necessary.

In summary:

$$
E_{cell} = E^{\circ} - \frac{0.0592}{n} \log Q
$$

where:

$$
Q = \frac{1}{[Ni^{2+}]} \quad \text{(for Ni/Ni concentration cell)}
$$

or, more generally, the ratio of concentrations at the electrodes.

Application: Calculating E and E° for a Nickel Concentration Cell

Suppose the following conditions:

- Electrode 1 (anode): Ni metal immersed in 0.01 M Ni^{2+} solution
- Electrode 2 (cathode): Ni metal immersed in 1 M Ni^{2+} solution
- Temperature: 25°C

Step 1: Calculate the cell potential E_{cell}:

$$
E_{cell} = 0.0592 \times \log \frac{c_{cathode}}{c_{anode}} = 0.0592 \times \log \frac{1}{0.01} = 0.0592 \times 2 = 0.1184\, V
$$

Step 2: Determine the standard electrode potential E°:

Since the electrodes are identical, the standard potential is zero:

$$
E^{\circ} = 0\, V
$$

Step 3: Find the actual electrode potentials:

- At the anode:

$$
E_{anode} = E^{\circ} - \frac{0.0592}{2} \log [Ni^{2+}] = -0.23\, V - 0.0296 \times \log 0.01
$$

$$
\log 0.01 = -2
$$

$$
E_{anode} = -0.23\, V - 0.0296 \times (-2) = -0.23\, V + 0.0592\, V = -0.1708\, V
$$

- At the cathode:

$$
E_{cathode} = E^{\circ} - \frac{0.0592}{2} \log 1 = -0.23\, V - 0.0296 \times 0 = -0.23\, V
$$

\]

Note: The difference between these potentials matches the cell potential calculated earlier (~0.1184 V).

Implications and Applications in Real-World Scenarios

Understanding how to compute the cell potential for nickel-based concentration cells has practical applications across various fields:

- Corrosion Prevention: Monitoring the potential difference can help predict corrosion rates in nickel-containing alloys.
- Electroplating and Metal Recovery: Precise calculations of cell potentials guide the electroplating process, ensuring uniform coatings.
- Battery Design: Nickel-based batteries, such as nickel-cadmium or nickel-metal hydride, rely on similar principles for optimizing performance.
- Environmental Monitoring: Detecting nickel ion concentrations in water sources assists in assessing pollution levels.

Conclusion

Frequently Asked Questions

What is the significance of the given E° value for Ni^{2+}/Ni in calculating cell potentials?

The E° value (-0.23 V) represents the standard reduction potential for the Ni^{2+}/Ni couple at 25°C, serving as a baseline to determine the cell potential under specific conditions using the Nernst equation.

How do you calculate the cell potential (E) for a concentration cell using the Nernst equation?

The cell potential can be calculated using $E = E° - (0.0592/n) \log(Q)$, where Q is the reaction quotient based on ion concentrations, and n is the number of electrons transferred.

What is a concentration cell, and how does it differ from a standard electrochemical cell?

A concentration cell consists of two half-cells with the same electrode material but different ion concentrations, generating a potential difference solely due to concentration gradients, unlike standard cells where different materials are involved.

How do you express the cell potential (E) for the concentration cell in terms of ion concentrations?

For a Ni^{2+}/Ni concentration cell, $E = (0.0592/n) \log([Ni^{2+}]_{outside} / [Ni^{2+}]_{inside})$, assuming the standard potential is zero for identical electrodes, or adjusted based on the given E° value.

Given $E^\circ = -0.23$ V for Ni^{2+}/Ni, how does this affect the calculation of the cell potential at non-standard concentrations?

Since E° is negative, the cell potential will depend on the ratio of ion concentrations, and the Nernst equation accounts for this by adjusting E based on the logarithmic term, reflecting the deviation from standard conditions.

What are typical units used for expressing the concentrations in the Nernst equation for this context?

Concentrations are typically expressed in molarity (M), which is mols of solute per liter of solution, used directly in the Nernst equation calculations.

How does temperature influence the calculation of E and E for the concentration cell?

Temperature affects the Nernst equation through the RT/F term; at 25°C, the equation simplifies with the factor 0.0592 V, but at other temperatures, this value needs adjustment accordingly.

What is the practical importance of calculating E and E for concentration cells in electrochemistry?

Calculating E and E helps predict the cell's voltage under specific conditions, crucial for designing batteries, sensors, and understanding electrochemical processes in real-world applications.

Additional Resources

Given That E Red = -0.23 V For Ni 2+/Ni At 25°C, Find E And E For The Concentration Cell Expressed Using

In the realm of electrochemistry, understanding how voltage measurements relate to ion concentrations is essential for grasping how batteries function, how corrosion occurs, and how electrochemical sensors operate. When provided with a standard reduction potential, such as E°_red for a nickel half-cell, scientists and engineers can determine the actual cell potential under various conditions. This article delves into the process of calculating the cell potential (E) for a nickel-based concentration cell, starting with

the given standard reduction potential, and explores how to express the cell potential in terms of ion concentrations. Through a detailed, step-by-step explanation, we aim to make this complex topic accessible for students, researchers, and enthusiasts alike.

Understanding Standard Reduction Potential ($E°_{red}$) and Its Significance

Before diving into calculations, it is vital to understand what the standard reduction potential ($E°_{red}$) signifies. In electrochemistry, $E°_{red}$ measures the tendency of a chemical species to gain electrons and be reduced under standard conditions (1 M concentration, 25°C, 1 atm pressure). For the nickel ion Ni^{2+}/Ni system, the given $E°_{red}$ is -0.23 V, indicating that nickel ions are relatively less eager to accept electrons compared to more positive potentials like those of noble metals.

Key points:

- $E°_{red}$ (Standard Reduction Potential): A

constant value indicating the inherent tendency of a half-cell to gain electrons.
- At 25°C: The temperature at which standard potentials are referenced.
- For Ni^{2+}/Ni: The half-reaction is $Ni^{2+} + 2e^- \rightarrow Ni$ (s).

This value serves as a baseline for calculating the cell potential under non-standard conditions, where ion concentrations differ from 1 M.

From Standard Potential to Cell Potential: The Nernst Equation

The core tool for converting standard potentials into actual cell potentials under specific conditions is the Nernst Equation. It relates the reduction potential to the activities (or concentrations) of the ions involved.

Nernst Equation (at 25°C):

$$E = E°_{red} - (RT / nF) \ln(Q)$$

In practical units, at 25°C, the equation simplifies to:

$E = E°_{red} - (0.0592\ V / n) \log(Q)$

Where:
- E is the cell potential under the given conditions.
- $E°_{red}$ is the standard reduction potential.
- n is the number of electrons transferred in the half-reaction.
- Q is the reaction quotient, reflecting concentrations.

For the Ni^{2+}/Ni system:
- n = 2 (since 2 electrons are involved).
- $Q = [Ni^{2+}]/[Ni(s)]$ (but pure solid Ni has activity = 1, so Q reduces to $[Ni^{2+}]$).

Constructing the Concentration Cell: Conceptual Framework

A concentration cell consists of two half-cells with the same electrode material but differing ion concentrations. It operates based on the concentration gradient, generating a voltage

that depends solely on the difference in concentrations.

Key features:

- Electrode: Nickel, in this case, serving as both anode and cathode in different compartments.
- Different concentrations: For example, one half-cell with a higher [Ni^{2+}], and the other with a lower [Ni^{2+}].
- Electrochemical potential difference: Drives electron flow from the lower to the higher concentration side.

The voltage of the concentration cell can be expressed using the Nernst equation, considering the concentration difference.

Calculating the Cell Potential (E) for the Concentration Cell

Given that the standard reduction potential for Ni^{2+}/Ni is -0.23 V, and assuming the cell operates at 25°C, the potential difference arises solely from the difference in

concentrations.

Step-by-step calculation:

1. Identify the half-reactions:

- Oxidation (anode): $Ni\ (s) \rightarrow Ni^{2+} + 2e^-$
- Reduction (cathode): $Ni^{2+} + 2e^- \rightarrow Ni\ (s)$

2. Express the cell potential:

Since both electrodes are nickel, the overall cell potential hinges on the difference in ion concentrations.

3. Express the cell potential using Nernst:

For the overall cell:

$E_{cell} = E°_{cell} + (0.0592\ V / n) \log([Ni^{2+}]_{oxidant} / [Ni^{2+}]_{reductant})$

But because the electrodes are identical, the standard cell potential ($E°_{cell}$) is zero—there's no difference in standard potentials.

Therefore:

$E_{cell} = (0.0592\ V / 2) \log([Ni^{2+}]_{high} /$

$[Ni^{2+}]_{low}$)

4. Final expression:

$E_{cell} = 0.0296\ V \log([Ni^{2+}]_{high} / [Ni^{2+}]_{low})$

This formula indicates that the cell potential depends solely on the ratio of the concentrations of Ni^{2+} in each half-cell.

Expressing E and E° for the Concentration Cell

Given the above, the potential of the concentration cell is:

$E = 0.0296\ V \log([Ni^{2+}]_{high} / [Ni^{2+}]_{low})$

where:

- $[Ni^{2+}]_{high}$ is the higher concentration of Ni^{2+} (on the cathode side).
- $[Ni^{2+}]_{low}$ is the lower concentration of Ni^{2+} (on the anode side).

Since the standard reduction potential $E°_{red}$ is -0.23 V, it serves as a baseline to

understand how the actual cell potential deviates from standard conditions due to concentration differences.

Practical Implications and Applications

Understanding how to calculate the potential of a nickel concentration cell has multiple practical implications:

- Battery Design: Optimizing concentration gradients to maximize voltage.
- Corrosion Prevention: Recognizing how ion concentrations influence metal degradation.
- Electrochemical Sensors: Detecting ion concentrations in solutions based on measured potentials.
- Industrial Processes: Managing electrolysis and plating operations where concentration gradients are manipulated.

Numerical Example: Calculating the Cell Potential for Specific Concentrations

Suppose we have a nickel concentration cell where:

- The Ni^{2+} concentration in the cathode compartment is 0.10 M.
- The Ni^{2+} concentration in the anode compartment is 0.01 M.

Using the formula:

E = 0.0296 V log(0.10 / 0.01) = 0.0296 V log(10)

Since log(10) = 1,

E = 0.0296 V 1 = 0.0296 V

Interpretation:

- The cell potential is approximately 29.6 mV.
- The positive voltage indicates electron flow from the lower concentration (anode) to the higher concentration (cathode), consistent with electrochemical principles.

Conclusion: The Power of Quantitative Electrochemistry

By starting with the given standard reduction potential for Ni^{2+}/Ni, we've demonstrated how to determine the actual cell potential for a concentration cell operating at 25°C. The key lies in applying the Nernst equation, which captures the influence of ion concentrations on the voltage. This approach not only deepens our understanding of electrochemical principles but also provides practical insights into designing and analyzing systems where ion concentrations drive electrical potentials.

In essence, the ability to express E and E° for such cells empowers scientists and engineers to harness electrochemical reactions effectively—be it in energy storage, corrosion control, or analytical chemistry. As we continue to explore the fascinating world of electrochemistry, these fundamental calculations form the backbone of innovations that shape our modern technological landscape.

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