Calculate The Equilibrium Constant For The Electrochemical Cell In Problem 10. Identify The Correct Answer.

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Introduction to Electrochemical Cells and Equilibrium Constants

Understanding the equilibrium constant (K) of an electrochemical cell is fundamental in electrochemistry as it provides insight into the cell's spontaneity and the extent to which a reaction proceeds. When an electrochemical cell reaches equilibrium, the forward and reverse reactions occur at equal rates, and the cell potential drops to zero. The equilibrium constant relates directly to the standard cell potential (E°) through the Nernst equation, allowing chemists to predict the behavior of electrochemical systems under various conditions.

In this article, we will explore the process of calculating the equilibrium constant for a specific electrochemical cell described in Problem 10, analyze the necessary concepts, and demonstrate the step-by-step approach to arrive at the correct answer.

Understanding the Given Data in Problem 10

Before delving into calculations, it is crucial to understand the typical data provided in such problems. Generally, Problem 10 might include:

- The standard cell potential, E° (in volts)
- The balanced overall cell reaction
- The number of electrons transferred (n)
- Conditions such as temperature (usually 25°C or 298 K)

Suppose the data from Problem 10 indicates:

- Standard cell potential, E° = 1.10 V
- The number of electrons transferred, n = 2
- Temperature, T = 298 K

With this data, we can proceed to determine the equilibrium constant.

Fundamental Concepts and Equations

The Relationship Between Cell Potential and Equilibrium Constant

The core equation linking the standard electrode potential to the equilibrium constant is derived from the thermodynamic relationship between Gibbs free energy (ΔG°) and the equilibrium constant (K):

$$
\Delta G^\circ = -RT \ln K
$$

where:

- R is the universal gas constant (8.314 J/(mol·K))
- T is the temperature in Kelvin
- K is the equilibrium constant

Additionally, the standard Gibbs free energy change relates to the standard cell potential via:

$$
\Delta G^\circ = -nFE^\circ
$$

where:

- n is the number of moles of electrons transferred
- F is the Faraday constant (96485 C/mol)
- E° is the standard cell potential

Combining these equations yields:

$$
-nFE^\circ = -RT \ln K
$$

Rearranged, the formula for K becomes:

$$
\ln K = \frac{nFE^\circ}{RT}
$$

or

$$
K = e^{\frac{nFE^\circ}{RT}}
$$

This is the primary equation used to calculate the equilibrium constant from the standard cell potential.

Step-by-Step Calculation of the Equilibrium Constant

Step 1: Gather Known Values

From Problem 10, we have:

- $E^\circ = 1.10\, \text{V}$
- $n = 2$
- $T = 298\, \text{K}$
- $R = 8.314\, \text{J/(mol·K)}$
- $F = 96485\, \text{C/mol}$

Step 2: Calculate the Exponent $\frac{nFE^\circ}{RT}$

Plugging in the values:

$$
\frac{nFE^\circ}{RT} = \frac{2 \times 96485\, \text{C/mol} \times 1.10\, \text{V}}{8.314\, \text{J/(mol·K)} \times 298\, \text{K}}
$$

Note: 1 Volt = 1 Joule/Coulomb, so units are compatible.

Calculating numerator:

$$
2 \times 96485 \times 1.10 = 2 \times 96485 \times 1.10 = 2 \times 106133.5 = 212267\, \text{J/mol}
$$

Calculating denominator:

$$
8.314 \times 298 = 2478.572\, \text{J/(mol)}
$$

Now, compute the ratio:

$$
\frac{212267}{2478.572} \approx 85.55
$$

Step 3: Calculate the Equilibrium Constant K

Using the exponential function:

$$
K = e^{85.55}
$$

This number is extremely large, indicating the reaction favors products heavily at equilibrium.

Interpreting the Result and Its Significance

The exponential of 85.55 yields a very high value, practically approaching infinity, which suggests that the reaction is essentially complete under standard conditions. In practical terms, the equilibrium constant is so large that the reactants are almost entirely converted to products at equilibrium.

This implies:

- The electrochemical cell is highly spontaneous.
- The cell potential is strong enough to drive the reaction forward to completion.
- Under standard conditions, the reaction's equilibrium lies heavily toward the products.

Identifying the Correct Answer

In multiple-choice scenarios, the options for K might be presented as:

- a) 1.2×10^{37}
- b) 2.5×10^{85}
- c) 3.2×10^{55}
- d) 5.0×10^{100}

Since our calculation yields $K \approx e^{85.55}$, which is approximately:

\[

$$K \approx \exp(85.55) \approx 3.16 \times 10^{37}$$

However, note that the exponent calculated was 85.55, leading to K ≈ $e^{85.55}$. But in the options, answer b) is 2.5×10^{85}, which indicates a much larger K. The discrepancy might be due to unit considerations or rounding.

To clarify, because the exponent is 85.55, the K should be:

$$K \approx e^{85.55} \approx \exp(85.55) \approx 3.16 \times 10^{37}$$

Therefore, the correct answer among typical options would be:

- a) 1.2×10^{37}
- b) 2.5×10^{85}
- c) 3.2×10^{55}
- d) 5.0×10^{100}

Given our calculation, the closest and most accurate estimate is option a), 1.2×10^{37}.

Conclusion and Final Remarks

Calculating the equilibrium constant from the standard cell potential involves understanding the thermodynamic relationships and applying the correct equations. The key steps include:

1. Gathering known data (E°, n, T).
2. Using the relation $K = e^{(nFE^\circ)/(RT)}$.
3. Carefully performing the calculations with proper unit consistency.
4. Interpreting the magnitude of K to understand the reaction's spontaneity.

In the context of Problem 10, the computed value of K indicates a highly spontaneous reaction, with an equilibrium constant on the order of 10^{37}. Identifying the correct answer among options requires matching this estimate, considering possible rounding and significant figures.

Understanding these concepts not only helps in solving specific problems but also deepens the appreciation of how thermodynamics and electrochemistry intertwine to describe the behavior of electrochemical systems.

Frequently Asked Questions

What is the first step in calculating the equilibrium constant (K) for the electrochemical cell described in Problem 10?

The first step is to write the balanced overall cell reaction based on the provided half-reactions and their standard potentials.

Which thermodynamic relationship relates the standard cell potential (E°) to the equilibrium constant (K)?

The relationship is $\Delta G° = -nFE°$, and at equilibrium, $\Delta G°$ is also related to K by $\Delta G° = -RT \ln K$, allowing the calculation of K from E°.

What is the formula to calculate the equilibrium constant (K) from the standard cell potential (E°)?

$K = e^{(nFE° / RT)}$, where n is the number of electrons transferred, F is Faraday's constant, R is the gas constant, and T is temperature in Kelvin.

How do you determine the number of electrons (n) involved in the cell reaction for Problem 10?

Identify the half-reactions and count the electrons transferred in the balanced overall reaction, which gives the value of n.

If the standard cell potential (E°) is given as 1.25 V at 25°C, how would you compute K?

Convert E° to volts, then use the formula $K = e^{(nFE° / RT)}$, plugging in the values for n, F, R, T, and E° to compute K.

What is the significance of the equilibrium constant (K) in electrochemical cells?

K indicates the extent of the reaction at equilibrium; a large K means the reaction favors products, while a small K favors reactants.

Why is it important to identify the correct number of electrons transferred in Problem 10?

Because the value of n directly affects the calculation of K; an incorrect n leads to an inaccurate equilibrium constant.

Additional Resources

Calculating the Equilibrium Constant for an Electrochemical Cell: A Comprehensive Guide

Understanding how to determine the equilibrium constant (K) of an electrochemical cell is fundamental in electrochemistry. It bridges the gap between thermodynamics and electrochemical processes, providing insight into the spontaneity and extent of redox reactions. In this review, we will explore the method to calculate the equilibrium constant for an electrochemical cell, particularly focusing on Problem 10, where the task is to identify the correct value of K based on given cell data. We will dive deep into the theoretical background, step-by-step calculation procedures, and practical considerations to ensure clarity and mastery of the concept.

Introduction to Electrochemical Cells and Equilibrium

Electrochemical cells are devices that convert chemical energy into electrical energy (galvanic cells) or use electrical energy to drive non-spontaneous chemical reactions (electrolytic cells). The fundamental principle underlying these cells is the redox (reduction-oxidation) reactions that occur at two electrodes: the anode (oxidation) and the cathode (reduction).

Key Concepts:
- Standard Electrode Potential (E°): The potential of a half-cell under standard conditions.
- Cell Potential (E°cell): The difference between cathode and anode potentials; indicates cell spontaneity.
- Reaction Quotient (Q): The ratio of concentrations or partial pressures of reactants and products at any point during the reaction.
- Equilibrium Constant (K): The ratio of product activities to reactant activities at equilibrium; a thermodynamic measure of the reaction's position.

The Relationship Between Cell Potential and Equilibrium Constant

The core thermodynamic relation that connects the electrochemical cell's potential to its equilibrium constant is derived from the standard Gibbs free energy change (ΔG°):

$$
\Delta G^\circ = -RT \ln K
$$

where:
- R = universal gas constant (8.314 J/mol·K),
- T = temperature in Kelvin,
- K = equilibrium constant.

From electrochemistry, the standard Gibbs free energy change relates to the standard cell potential via:

$$
\Delta G^\circ = -nFE^\circ_{cell}
$$

where:
- n = number of moles of electrons transferred,
- F = Faraday's constant (~96485 C/mol),
- E°_{cell} = standard cell potential.

Combining these equations gives:

$$
-nFE^\circ_{cell} = -RT \ln K
$$

which simplifies to:

$$
\boxed{
\ln K = \frac{nFE^\circ_{cell}}{RT}
}
$$

and hence,

$$
K = e^{\frac{nFE^\circ_{cell}}{RT}}
$$

This fundamental relation enables us to calculate the equilibrium constant directly from measurable cell potential data, provided we know the number of electrons involved and the reaction conditions.

Step-by-Step Approach to Calculate K in Problem 10

Suppose Problem 10 provides:
- The standard cell potential, E°_{cell},
- The temperature at which the reaction occurs (commonly 25°C or 298 K),
- The number of electrons transferred in the redox process, n.

Let's outline the step-by-step process:

1. Identify the Data Provided

Ensure all necessary data are extracted:

- Cell potential, E°_{cell}: Usually given under standard conditions.
- Temperature, T: Often 298 K unless specified otherwise.
- Number of electrons transferred, n: Derived from the balanced redox equation.

2. Convert Units Appropriately

- Confirm E°_{cell} is in volts (V).
- Temperature should be in Kelvin (K).
- Use the constants:
- $R = 8.314\, \mathrm{J/mol \cdot K}$
- $F = 96485\, \mathrm{C/mol}$

3. Apply the Equation

Using the relation:

$$
K = e^{\frac{n F E^\circ_{cell}}{RT}}
$$

calculate the exponent:

$$
\text{Exponent} = \frac{n \times F \times E^\circ_{cell}}{R \times T}
$$

which is dimensionless.

4. Calculate K

Compute the exponential to find K:

$$
K = e^{\text{Exponent}}
$$

The value of K indicates the position of equilibrium:
- $K \gg 1$ implies a reaction that favors products.
- $K \ll 1$ indicates a reaction favoring reactants.
- $K \approx 1$ suggests near-equilibrium conditions.

Practical Application: Example Calculation

Suppose the problem states:

- The standard cell potential, $E^\circ_{cell} = 1.10\, \mathrm{V}$,
- The reaction involves 2 electrons ($n=2$),

- Temperature is 25°C (298 K).

Calculation:

$$
\text{Exponent} = \frac{2 \times 96485 \times 1.10}{8.314 \times 298} \approx \frac{212,267}{2477.172} \approx 85.66
$$

So,

$$
K = e^{85.66} \approx 1.37 \times 10^{37}
$$

This enormous value indicates the reaction strongly favors products at equilibrium.

Interpreting the Results and Choosing the Correct Answer

Suppose in Problem 10, multiple-choice answers are provided, such as:

- A) 1.2×10^{10}
- B) 3.5×10^{15}
- C) 1.4×10^{37}
- D) 5.0×10^{5}

Given our calculation (~1.37×10^{37}), the correct answer is C.

This highlights the importance of precise calculation and understanding the magnitude of the equilibrium constant. Large K values indicate the reaction proceeds almost to completion, which is typical for cells with high E°.

Additional Considerations and Common Pitfalls

- Standard Conditions: Ensure the potential value corresponds to standard conditions; deviations may require adjustments.
- Sign of E°_{cell}: A negative value indicates non-spontaneity under standard conditions; the formula can still be used to determine the equilibrium constant, but interpret with caution.
- Temperature Effect: The equilibrium constant is temperature-dependent; calculations should reflect the actual temperature if different from 25°C.
- Electrode Potentials: When using non-standard potentials, adjust with the Nernst equation to find \(

E_{cell} \) at specific conditions before calculating K.
- Number of Electrons \(n \): Always carefully balance the redox reaction to determine the correct \(n \).

Conclusion

Calculating the equilibrium constant for an electrochemical cell involves a robust understanding of thermodynamics, electrochemical principles, and meticulous calculation. By leveraging the relation:

$$
K = e^{\frac{n F E^\circ_{cell}}{RT}}
$$

and carefully plugging in the correct values, we can accurately determine whether a reaction favors products or reactants and gauge its extent at equilibrium.

In the context of Problem 10, the key steps are to identify the provided data, apply the correct formulas, and interpret the resulting value within the broader chemical context. Mastery of this process enhances comprehension of electrochemical systems and their thermodynamic properties, enabling chemists and students to analyze electrode reactions critically and accurately.

In summary:

- Recognize the fundamental thermodynamic relation linking cell potential and K.
- Use precise values for constants and parameters.
- Carefully perform exponential calculations to avoid errors.
- Interpret the magnitude of K in terms of reaction spontaneity.
- Apply this knowledge to diverse electrochemical problems for robust understanding.

By mastering these concepts, students and professionals can confidently calculate and interpret equilibrium constants for electrochemical cells, a cornerstone skill in electrochemistry.

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