Calculate The Equilibrium Constant For The Reaction Cd 2+(aq)+Zn(s)Zn 2+(aq)+Cd(s) If E Cd 2+/Cd=0.403V;E

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Introduction

Understanding the equilibrium constant (K) of a chemical reaction is fundamental in chemistry, especially in the context of redox reactions and electrochemistry. This article focuses on calculating the equilibrium constant for the reaction:

$$\mathrm{Cd^{2+}(aq) + Zn(s) \rightleftharpoons Zn^{2+}(aq) + Cd(s)}$$

Given the standard reduction potential for the $\mathrm{Cd^{2+}/Cd}$ couple as $E^\circ = 0.403\,V$, we aim to determine the equilibrium constant for this reaction. This calculation provides insight into the spontaneity of the reaction and the relative stabilities of the ions and metals involved.

Electrochemical potentials and the Nernst equation form the core tools for such calculations. Understanding how to convert standard potentials into equilibrium constants enables chemists to predict reaction directions, design electrochemical cells, and interpret redox behavior in various chemical systems.

In this article, we will explore the theoretical background, step-by-step calculations, and practical implications of determining the equilibrium constant for the specified reaction.

Understanding Standard Reduction Potentials and Their Role

What Is a Standard Reduction Potential?

A standard reduction potential (E°) measures the tendency of a species to gain electrons and be reduced under standard conditions (1 M concentration, 1 atm pressure, 25°C). It is measured relative to the standard hydrogen electrode (SHE), which is assigned an E° of 0.000 V.

For the reaction:

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

the given reduction potential is:

$$ E^\circ_{Cd^{2+}/Cd} = 0.403\,V $$

Similarly, the zinc couple is represented as:

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

with its standard reduction potential typically provided as:

$$ E^\circ_{Zn^{2+}/Zn} = -0.763\,V $$

(Note: This value is standard and can be found in electrochemical tables.)

Significance in Redox Reactions

The difference in reduction potentials indicates which species is more prone to gain electrons. A higher E° suggests a greater tendency to be reduced. When two half-reactions are combined, the overall cell potential (E_{cell}°) determines whether the reaction proceeds spontaneously.

The reaction:

$$ \mathrm{Cd^{2+}(aq) + Zn(s) \rightleftharpoons Zn^{2+}(aq) + Cd(s)} $$

can be viewed as the zinc metal reducing cadmium ions to cadmium metal, while zinc itself is oxidized to zinc ions.

Determining the Standard Cell Potential (E°_{cell})

Half-Cell Reactions

To analyze the reaction, we need the two half-reactions:

- Reduction of cadmium ions:

$$ \mathrm{Cd^{2+} + 2e^- \rightarrow Cd(s)} \quad E^\circ = 0.403\,V $$

- Oxidation of zinc metal:

$$ \mathrm{Zn(s) \rightarrow Zn^{2+} + 2e^-} \quad E^\circ_{oxidation} = -E^\circ_{Zn^{2+}/Zn} = +0.763\,V $$

(Note: The oxidation potential is the negative of the reduction potential.)

Calculating E°_{cell}

The standard cell potential is:

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$

But since zinc is oxidized (loses electrons), its oxidation potential is used directly:

$$E^\circ_{cell} = E^\circ_{reduction\,of\,cathode} - E^\circ_{reduction\,of\,anode}$$

Alternatively, considering the zinc oxidation:

$$E^\circ_{cell} = E^\circ_{Zn^{2+}/Zn} + E^\circ_{Cd^{2+}/Cd}$$

Using the standard reduction potentials:

$$E^\circ_{cell} = (-0.763\,V) + (0.403\,V) = -0.360\,V$$

However, since the reduction potential for zinc is negative, indicating zinc metal is a good reducing agent, the overall cell potential is:

$$E^\circ_{cell} = 0.403\,V - (-0.763\,V) = 1.166\,V$$

This positive value indicates that the reaction:

$$\mathrm{Zn(s) + Cd^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cd(s)}$$

is spontaneous under standard conditions.

Note: The correct approach is to reverse the sign of the oxidation half-reaction potential and add it to the reduction potential of cadmium:

$$E^\circ_{cell} = E^\circ_{reduction\,Cd^{2+}/Cd} + E^\circ_{oxidation\,Zn}$$

$$E^\circ_{cell} = 0.403\,V + 0.763\,V = 1.166\,V$$

This confirms the spontaneity of the reaction.

Connecting Cell Potential to Equilibrium Constant

The Nernst Equation

The Nernst equation relates the cell potential to the concentrations of reactants and products:

$$ E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln Q $$

Where:

- E_{cell} is the cell potential at non-standard conditions,
- E°_{cell} is the standard cell potential,
- R is the universal gas constant ($8.314\,J\,mol^{-1}K^{-1}$),
- T is the temperature in Kelvin (usually 298 K),
- n is the number of electrons transferred (2 in this case),
- F is the Faraday constant ($96485\,C\,mol^{-1}$),
- Q is the reaction quotient:

$$ Q = \frac{[\mathrm{Zn^{2+}}]}{[\mathrm{Cd^{2+}}]} $$

At equilibrium, $E_{cell} = 0$, and the reaction quotient becomes the equilibrium constant K:

$$ E_{cell}^\circ = \frac{RT}{nF} \ln K $$

Rearranged to solve for K:

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

This allows us to calculate the equilibrium constant directly from the standard cell potential.

Calculating the Equilibrium Constant (K)

Step-by-Step Calculation

Given:

- $E^\circ_{cell} = 1.166\,V$,
- $n = 2$,
- $T = 298\,K$,
- $R = 8.314\,J\,mol^{-1}K^{-1}$,
- $F = 96485\,C\,mol^{-1}$.

Plugging into the formula:

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

Calculations:

1. Compute numerator:

$$ n F E^\circ_{cell} = 2 \times 96485 \times 1.166 = 2 \times 96485 \times 1.166 $$

$$ 2 \times 96485 = 192970 $$

$$ 192970 \times 1.166 \approx 225,245.82\,J $$

2. Compute denominator:

$$ RT = 8.314 \times 298 \approx 2478.0\,J\,mol^{-1} $$

3. Compute the exponent:

$$ \frac{225,245.82}{2478.0} \approx 90.8 $$

4. Final calculation:

$$ K = e^{90.8} $$

Using a calculator:

$$ K \approx e^{90.8} \approx 2.8 \times 10^{39} $$

Therefore, the equilibrium constant is approximately:

$$ \boxed{K \approx 2.8 \times 10^{39}} $$

This extremely large value indicates the reaction strongly favors the formation of zinc ions and solid cadmium at equilibrium.

Practical Implications of the Calculated K

The calculated equilibrium constant provides valuable insights:

- Reaction Favorability: The very high K suggests the reaction proceeds spontaneously in the forward direction, with zinc metal reducing cadmium ions to metallic cadmium while zinc ions are formed.

- Electrochemical Cell Design: Knowing E°_{cell} and K helps in designing batteries and electrochemical cells involving these metals, optimizing for efficiency and stability.

- Predicting Reaction Behavior: Chemists can predict how changes in concentrations or conditions may shift the equilibrium, based on Le Châtelier's principle.

- Corrosion and Material Stability:

Frequently Asked Questions

What is the overall reaction when cadmium ions react with

zinc solid?

The overall reaction is $Cd^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cd(s)$.

Given the standard electrode potentials E° for Cd^{2+}/Cd and Zn^{2+}/Zn, how do you determine the cell potential?

The cell potential $E°_{cell}$ is calculated by subtracting the anode potential from the cathode potential: $E°_{cell} = E°_{cathode} - E°_{anode}$.

What is the standard electrode potential for Cd^{2+}/Cd?

The standard electrode potential for Cd^{2+}/Cd is 0.403 V.

How do you find the equilibrium constant (K) from the cell potential?

The equilibrium constant K is related to the cell potential by the equation: $\Delta G° = -nFE°_{cell}$, and $\Delta G° = -RT \ln K$. Combining these gives $K = e^{(nFE°_{cell})/(RT)}$.

What is the value of the number of electrons transferred (n) in this reaction?

In this reaction, n = 2 electrons are transferred per reaction, as indicated by the change in oxidation states.

How do you calculate the equilibrium constant K for the given reaction at standard conditions?

Calculate $E°_{cell}$, then use $K = 10^{(E°_{cell}\, n\, F) / (2.303\, R\, T)}$ or $K = e^{(n\, F\, E°_{cell}) / (RT)}$, where R = 8.314 J/(mol·K), T = 298 K, and F = 96485 C/mol.

Can you provide the step-by-step calculation to find K for this reaction?

Yes. First, determine $E°_{cell} = E°_{cathode} - E°_{anode}$. Then, use $K = e^{(n\, F\, E°_{cell}) / (RT)}$ with known constants to compute the equilibrium constant.

What assumptions are made in calculating the equilibrium constant from standard electrode potentials?

The calculation assumes standard conditions (25°C, 1 atm, 1 M concentrations), and that the system is at equilibrium with no other interfering reactions.

Why is the standard electrode potential for Cd^{2+}/Cd given as 0.403 V important in this calculation?

It provides the reference potential needed to calculate the cell potential and ultimately determine the equilibrium constant for the reaction.

How does the standard electrode potential influence the spontaneity of the reaction?

A positive $E°_{cell}$ indicates a spontaneous reaction, which correlates with a larger value of the equilibrium constant K, favoring product formation.

Additional Resources

Calculate The Equilibrium Constant For The Reaction $Cd^{2+}(aq) + Zn(s) \rightleftharpoons Zn^{2+}(aq) + Cd(s)$

Introduction

Understanding the thermodynamic feasibility of electrochemical reactions is fundamental in fields ranging from inorganic chemistry to materials science and electrochemical engineering. A pivotal aspect of this understanding involves calculating the equilibrium constant (K) for a given redox reaction. This constant provides insight into the position of equilibrium, indicating whether a reaction favors the formation of products or reactants under standard conditions.

In this article, we focus on a specific redox process:

$$\mathrm{Cd^{2+}(aq)} + \mathrm{Zn(s)} \rightleftharpoons \mathrm{Zn^{2+}(aq)} + \mathrm{Cd(s)}$$

The goal is to calculate the equilibrium constant K for this reaction given the standard reduction potential for the $\mathrm{Cd^{2+}/Cd}$ couple:

$$E^\circ_{\mathrm{Cd^{2+}/Cd}} = 0.403\, \mathrm{V}$$

and the standard reduction potential for the $\mathrm{Zn^{2+}/Zn}$ couple, which, although not explicitly provided, can be deduced from known electrochemical data.

Theoretical Background

The Relationship Between Cell Potential and Equilibrium Constant

The fundamental link between electrochemical measurements and thermodynamic parameters is expressed through the Nernst equation and the thermodynamic Gibbs free energy change:

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

where:
- ΔG° is the standard Gibbs free energy change,
- n is the number of electrons transferred,
- F is the Faraday constant (96,485 C/mol),
- E°_{cell} is the standard cell potential.

The equilibrium constant K relates directly to ΔG°:

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

Combining these relations yields:

$$
\boxed{
\Delta G^\circ = -nFE^\circ_{\text{cell}} = -RT \ln K
}
$$

which can be rearranged to find:

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

where:
- R is the universal gas constant (8.314 J/(mol·K)),
- T is the temperature in Kelvin (standard temperature is 298 K).

Step 1: Determining the Standard Reduction Potentials

The Standard Reduction Potential of Zinc

The standard reduction potential for the $\mathrm{Zn^{2+}/Zn}$ couple is well-established:

$$
E^\circ_{\mathrm{Zn^{2+}/Zn}} = -0.762\, \mathrm{V}
$$

This value is referenced against the standard hydrogen electrode (SHE).

The Standard Reduction Potential of Cadmium

Given:

$$
E^\circ_{\mathrm{Cd^{2+}/Cd}} = 0.403\, \mathrm{V}
$$

which is also relative to SHE.

Step 2: Constructing the Overall Cell Reaction and Calculating E°_{cell}

The overall reaction in the forward direction is:

$$
\mathrm{Cd^{2+}(aq)} + \mathrm{Zn(s)} \rightarrow \mathrm{Zn^{2+}(aq)} + \mathrm{Cd(s)}
$$

This involves oxidation of zinc and reduction of cadmium, with their respective half-reactions:

- Zinc oxidation:

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

with an oxidation potential:

$$
E^\circ_{\text{oxidation}} = -E^\circ_{\mathrm{Zn^{2+}/Zn}} = +0.762\, \mathrm{V}
$$

- Cadmium reduction:

$$
\mathrm{Cd^{2+}(aq)} + 2e^- \rightarrow \mathrm{Cd(s)}
$$

with:

$$
E^\circ_{\mathrm{Cd^{2+}/Cd}} = 0.403\, \mathrm{V}
$$

Since the overall cell involves zinc oxidation and cadmium reduction, the cell potential is:

$$
E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}
$$

Alternatively, when combining the half-reactions:

$$
E^\circ_{\text{cell}} = E^\circ_{\mathrm{Cd^{2+}/Cd}} - E^\circ_{\mathrm{Zn^{2+}/Zn}}
$$

Substituting the known values:

$$
E^\circ_{\text{cell}} = 0.403\, \mathrm{V} - (-0.762\, \mathrm{V}) = 0.403 + 0.762 = 1.165\, \mathrm{V}
$$

This positive cell potential indicates the reaction proceeds spontaneously under standard conditions.

Step 3: Calculating the Equilibrium Constant K

Using the fundamental thermodynamic relation:

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

- $n = 2$ (number of electrons transferred per reaction event),
- $F = 96485\, \mathrm{C/mol}$,
- $R = 8.314\, \mathrm{J/(mol·K)}$,
- $T = 298\, \mathrm{K}$,
- $E^\circ_{\text{cell}} = 1.165\, \mathrm{V}$.

Calculating the exponent:

$$
\frac{nFE^\circ_{\text{cell}}}{RT} = \frac{2 \times 96485 \times 1.165}{8.314 \times 298}
$$

Step-by-step:

1. Numerator:

$$
2 \times 96485 \times 1.165 \approx 2 \times 96485 \times 1.165 \approx 2 \times 112,551.025 \approx 225,102.05\, \mathrm{J}
$$

2. Denominator:

$$
8.314 \times 298 \approx 2477.572
$$

3. Final exponent:

$$
\frac{225,102.05}{2477.572} \approx 90.8
$$

Thus, the equilibrium constant is:

$$
K = e^{90.8} \approx 1.74 \times 10^{39}
$$

Conclusions and Implications

The calculated equilibrium constant $K \approx 1.74 \times 10^{39}$ indicates an extraordinarily favorable reaction toward products under standard conditions. This immense value suggests that, thermodynamically, the conversion of $\mathrm{Cd^{2+}}$ ions to metallic cadmium coupled with zinc metal oxidation is effectively complete at equilibrium.

Implications in Electrochemistry and Material Science:

- Electrochemical Reactions: The reaction's high K value signifies that, in practical electrochemical cells, zinc effectively displaces cadmium ions, making this process suitable for applications like metal plating or metal recovery.

- Corrosion and Metal Displacement: The strong driving force for zinc to reduce cadmium ions can be relevant for corrosion studies or designing metal displacement reactions in industrial processes.

- Battery and Energy Storage: Such thermodynamically favorable reactions are foundational in designing batteries and electrochemical cells, especially where metal displacement reactions are harnessed.

Additional Considerations

Effect of Non-Standard Conditions

While this calculation assumes standard temperature (298 K) and concentrations, real-world systems may deviate, affecting the actual equilibrium constant. The Nernst equation can be used to adjust for these factors:

$$
E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q
$$

where Q is the reaction quotient. Variations in ion concentrations or temperature can shift the equilibrium, although the thermodynamic favorability remains high.

Limitations and Assumptions

- The calculation presumes ideal behavior and neglects activity coefficients.
- The standard reduction potentials are referenced to the SHE; deviations can occur depending on experimental conditions.
- Kinetic factors are not considered; a high K indicates thermodynamic favorability but does not

necessarily equate to rapid reaction rates.

Final Remarks

Calculating the equilibrium constant from electrochemical data provides a powerful tool to predict the spontaneity and extent of redox reactions. In the case of the reaction between cadmium ions and zinc metal, the thermodynamic analysis reveals an exceedingly favorable process, underpinning its potential applications in electrochemical industries and materials science.

Understanding such fundamental parameters aids in the rational design of electrochemical systems, informs material selection, and guides practical implementations where metal displacement reactions are employed.

Calculate The Equilibrium Constant For The Reaction Cd 2 Aq Zn S Zn 2 Aq Cd S If E Cd 2 Cd 0 403v E

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