Calculate The Equilibrium Constant At 25 C For The ReactionCo(s) + 2Ag+(aq) Co2+(aq) + 2Ag(s)Standard

Calculate The Equilibrium Constant At 25°C For The Reaction $Co(s) + 2Ag^+(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$ – Standard Conditions

Understanding how to calculate the equilibrium constant (K) for a chemical reaction at a specific temperature is fundamental in chemistry, especially in fields like electrochemistry, environmental science, and industrial processes. In this article, we will explore the detailed process of calculating the equilibrium constant at 25°C (which is 298 K) for the reaction:

$Co(s) + 2Ag^+(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$

This reaction involves solid metals and aqueous ions, and its equilibrium constant provides insight into the extent of the reaction's favorability under standard conditions.

Understanding the Reaction and Its Components

Before diving into calculations, it's crucial to understand the nature of the reaction, the species involved, and their standard potentials.

Reaction Overview

The reaction involves the oxidation of cobalt metal and the reduction of silver ions:

- Cobalt metal ($Co(s)$) is oxidized to cobalt ions (Co^{2+}).
- Silver ions (Ag^+) are reduced to metallic silver ($Ag(s)$).

The balanced reaction is:

$Co(s) + 2Ag^+(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$

Standard Electrode Potentials

The key to calculating the equilibrium constant is the standard reduction potentials (E°) for the involved half-reactions.

Half-Reaction	Standard Reduction Potential (E°)	Reference
$Ag^+ + e^- \rightarrow Ag(s)$	+0.80 V	Standard
$Co^{2+} + 2e^- \rightarrow Co(s)$	–0.28 V	Standard

Note: The potentials are obtained from standard electrochemical tables.

Step 1: Write the Half-Reactions and Find the Cell Potential

The overall reaction combines the two half-reactions. To determine the cell potential (E°_cell), we need to identify which species undergo oxidation and which undergo reduction.

Half-Reactions

- Silver reduction: $Ag^+ + e^- \rightarrow Ag(s)$ (E° = +0.80 V)
- Cobalt oxidation: $Co(s) \rightarrow Co^{2+} + 2e^-$ (E° = ?)

Since the standard reduction potential for $Co^{2+} + 2e^- \rightarrow Co(s)$ is –0.28 V, the oxidation of Co(s) is the reverse:

$Co(s) \rightarrow Co^{2+} + 2e^-$ (E°_oxidation = +0.28 V)

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

Calculating the Standard Cell Potential (E°_cell)

Combine the two half-reactions:

- Oxidation: $Co(s) \rightarrow Co^{2+} + 2e^-$ (E° = +0.28 V)
- Reduction: $2Ag^+ + 2e^- \rightarrow 2Ag(s)$ (E° = +0.80 V)

Since the reduction potential for silver is per $1e^-$, and the reaction involves $2Ag^+$, the reduction half-

reaction becomes:

$2Ag^+ + 2e^- \rightarrow 2Ag(s)$

The standard cell potential is:

$E°_{cell} = E°_{cathode} - E°_{anode}$

But in electrochemical notation, it's:

$E°_{cell} = E°_{(reduction, cathode)} + E°_{(oxidation, anode)}$

Alternatively, more straightforwardly:

$E°_{cell} = E°_{(reduction\ of\ Ag^+)} + |E°_{(reduction\ of\ Co^{2+})}|$

Given the oxidation potential of Co is +0.28 V, the cell potential is:

$E°_{cell} = E°_{(Ag^+/Ag)} + (-E°_{(Co^{2+}/Co)})$
$E°_{cell} = +0.80\ V + 0.28\ V = +1.08\ V$

Note: The positive value indicates the reaction is spontaneous under standard conditions.

Step 2: Relate Cell Potential to Equilibrium Constant

The Nernst equation provides a link between cell potential and the reaction quotient, but at standard conditions (activities or concentrations at 1 M and pure solids), the cell potential equals the standard potential:

$E°_{cell} = (RT/nF) \ln K$

Where:
- R = universal gas constant = 8.314 $J \cdot mol^{-1} \cdot K^{-1}$
- T = temperature in Kelvin = 298 K
- n = number of electrons transferred = 2
- F = Faraday's constant = 96485 $C \cdot mol^{-1}$
- K = equilibrium constant

Rearranged:

$K = e^{(nF E°_{cell})/(RT)}$

Step 3: Calculate the Equilibrium Constant (K) at 25°C

Now, substitute known values:

$K = e^{(2 \ 96485 \ C/mol \ 1.08 \ V) / (8.314 \ J/(mol·K) \ 298 \ K)}$

Calculate numerator:

$2 \ 96485 \ 1.08 \approx 208{,}373.8 \ J/mol$

Calculate denominator:

$8.314 \ 298 \approx 2477.7 \ J/mol$

Compute the exponent:

$208{,}373.8 / 2477.7 \approx 84.1$

Finally, calculate K:

$K = e^{84.1} \approx 2.69 \times 10^{36}$

Interpretation of the Result

The equilibrium constant $K \approx 2.69 \times 10^{36}$ indicates an extremely favorable reaction toward the formation of products (Co^{2+} and Ag(s)) under standard conditions at 25°C. This high value suggests that, at equilibrium, the reaction overwhelmingly favors the right-hand side, meaning cobalt metal readily reduces silver ions to metallic silver while oxidizing to cobalt ions.

Additional Considerations and Practical Implications

Understanding the magnitude of the equilibrium constant is essential in various applications:

- Electrochemical cells: Such reactions form the basis for galvanic cells and batteries.
- Metal recovery: The high K value indicates efficient silver recovery via cobalt reduction.
- Corrosion and material stability: Knowledge of such reactions informs corrosion prevention strategies.

Factors influencing the equilibrium constant include:

- Temperature: Changes in temperature alter E°_cell and thus K.
- Concentrations: Deviations from standard conditions affect the reaction quotient Q.
- Activity coefficients: In real solutions, non-ideal behavior can modify effective concentrations.

Summary and Conclusion

Calculating the equilibrium constant at 25°C for the reaction $Co(s) + 2Ag^{+}(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$ involves understanding the electrochemical potentials, combining half-reactions to find the standard cell potential, and applying the Nernst equation. The resulting K value of approximately 2.69×10^{36} underscores the reaction's strong spontaneity and tendency to proceed toward product formation under standard conditions. This calculation not only exemplifies the fundamental principles of electrochemistry but also offers insights into practical applications in metal recovery, electrochemical device design, and materials science.

In summary:

- The reaction's standard cell potential at 25°C is +1.08 V.
- The equilibrium constant (K) at 25°C is approximately 2.69×10^{36}.
- The high K value indicates a reaction that strongly favors product formation under standard conditions.

By mastering these calculations, chemists and engineers can predict reaction behavior, design better electrochemical systems, and optimize industrial processes involving metal ions and solids.

Keywords: Equilibrium constant, standard conditions, electrochemistry, cell potential, Nernst equation, cobalt, silver, oxidation, reduction, thermodynamics, 25°C

Frequently Asked Questions

What is the expression for the equilibrium constant (K) for the reaction $Co(s) + 2Ag^+(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$?

The equilibrium constant expression is $K = [Co^{2+}] / [Ag^+]^2$, since solids are omitted from the expression.

Given standard reduction potentials, how can I calculate the equilibrium constant at 25°C for this reaction?

Use the relation $\Delta G° = -RT \ln K$, where $\Delta G°$ is derived from the standard reduction potentials, and then solve for K.

What is the standard cell potential (E°cell) for this reaction based on standard reduction potentials?

Calculate E°cell by subtracting the standard reduction potential of Co^{2+}/Co from that of Ag^+/Ag, considering their respective reduction potentials.

How do I find the standard reduction potentials for Co and Ag to compute K?

Use standard reduction potential tables: Ag^+/Ag has $E° \approx +0.80$ V, and Co^{2+}/Co has $E° \approx -0.28$ V.

What is the value of the equilibrium constant (K) at 25°C for the given reaction?

First, determine E°cell = 0.80 V - (-0.28 V) = 1.08 V. Then, use $\Delta G° = -nFE°cell$, where n=2, to find $\Delta G°$, and finally compute K using $K = e^{(-\Delta G°/RT)}$. The approximate value of K is 1.2×10^8.

Why do solids like Co(s) and Ag(s) not appear in the equilibrium expression?

Because the activities of pure solids and liquids are considered constant and are set to 1, they are omitted from the expression for the equilibrium constant.

How does temperature affect the calculation of the equilibrium constant

for this reaction?

The calculation assumes a temperature of 25°C (298 K). If temperature changes, E°cell and ΔG° change accordingly, affecting the value of K.

Additional Resources

Calculate The Equilibrium Constant At 25°C For The Reaction $Co(s) + 2Ag^{+}(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$

Introduction

The calculation of the equilibrium constant (K) for chemical reactions provides critical insight into the extent of reaction, spontaneity, and the relative stability of reactants and products under specified conditions. In this context, determining the equilibrium constant at 25°C (298 K) for the reaction:

$Co(s) + 2Ag^{+}(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$

serves as a quintessential example of integrating electrochemical data with thermodynamic principles. This particular reaction involves the oxidation of cobalt metal with silver ions, resulting in cobalt ions and metallic silver. Understanding the thermodynamics of this process requires a comprehensive analysis of the relevant standard electrode potentials, standard Gibbs free energies, and their interrelations.

Thermodynamic Foundations

To compute the equilibrium constant, it is essential to understand the underlying thermodynamic relationships:

Gibbs Free Energy and Equilibrium Constant

The standard Gibbs free energy change (ΔG°) for a reaction at temperature T relates directly to the equilibrium constant K via:

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

where:

- R = 8.314 J/(mol·K) (universal gas constant)
- T = temperature in Kelvin (for 25°C, T = 298 K)
- K = equilibrium constant

Rearranging, the equilibrium constant can be expressed as:

$$
K = e^{-\frac{\Delta G^\circ}{RT}}
$$

Standard Gibbs Free Energy and Standard Electrode Potentials

The standard Gibbs free energy change for an electrochemical reaction relates to the standard cell potential (E°) through:

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

where:

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

Thus, by calculating E°, we can determine ΔG°, and subsequently K.

Electrochemical Data for the Reaction Components

The reaction involves the oxidation of cobalt metal and the reduction of silver ions:

- Oxidation: $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$
- Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

To evaluate the overall cell potential, we need the standard reduction potentials for the half-reactions:

Half-Reaction	Standard Reduction Potential, E° (V)
$Ag^+ + e^- \rightarrow Ag(s)$	+0.80
$Co^{2+} + 2e^- \rightarrow Co(s)$	–0.28

Note: Standard potentials are referenced against the Standard Hydrogen Electrode (SHE).

Constructing the Cell Reaction and Calculating E°

The overall reaction involves the oxidation of cobalt and the reduction of silver:

$$
\text{Oxidation: } \text{Co(s)} \rightarrow \text{Co}^{2+}(aq) + 2e^-
$$
$$
\text{Reduction: } 2\text{Ag}^+(aq) + 2e^- \rightarrow 2\text{Ag(s)}
$$

By combining these half-reactions, the net cell reaction becomes:

$$
\text{Co(s)} + 2\text{Ag}^+(aq) \rightarrow \text{Co}^{2+}(aq) + 2\text{Ag(s)}
$$

The overall standard cell potential (E°_cell) is calculated as:

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

Where:

- Cathode (reduction): Ag^+/Ag, E° = +0.80 V
- Anode (oxidation): Co/Co^{2+}, E° = –0.28 V (since reduction potential is given, and the reaction proceeds in the opposite direction)

However, for the anode, the oxidation potential is the negative of the reduction potential:

$$
E^\circ_{\text{oxidation}} = –(–0.28\ V) = +0.28\ V
$$

Thus:

$$
E^\circ_{\text{cell}} = 0.80\,V – (–0.28\,V) = 0.80\,V + 0.28\,V = 1.08\,V
$$

\]

Alternatively, considering oxidation as the reverse of the reduction reaction, the calculation simplifies to:

\[
E^\circ_{\text{cell}} = E^\circ_{\text{reduction, cathode}} - E^\circ_{\text{reduction, anode}} = 0.80\,V - (-0.28\,V) = 1.08\,V
\]

Calculating ΔG° and K

Given the standard cell potential:

\[
E^\circ = 1.08\,V
\]

and the number of electrons transferred:

\[
n = 2
\]

the standard Gibbs free energy change:

\[
\Delta G^\circ = -nFE^\circ = -2 \times 96485\,\mathrm{C/mol} \times 1.08\,V
\]

Calculating:

\[
\Delta G^\circ = -2 \times 96485 \times 1.08 \approx -208,489\, \mathrm{J/mol}
\]

Now, applying the relationship to find K:

\[
K = e^{-\frac{\Delta G^\circ}{RT}} = e^{-\frac{-208,489}{8.314 \times 298}} = e^{\frac{208,489}{2477.772}} \approx e^{84.09}
\]

This yields:

$$
K \approx e^{84.09} \approx 1.58 \times 10^{36}
$$

Interpretation of the Equilibrium Constant

The enormous magnitude of K (~10^36) indicates that at 25°C, the reaction strongly favors the formation of products—namely, Co^{2+} ions and solid silver. This suggests that under standard conditions, the oxidation of cobalt by silver ions proceeds essentially to completion, with negligible reverse reaction.

This thermodynamic insight aligns with the expected electrochemical behavior: silver ions are reduced readily to metallic silver, and cobalt metal is oxidized to Co^{2+} with high spontaneity.

Factors Influencing the Equilibrium Constant

While the calculation above provides a theoretical value based on standard potentials, real-world conditions may introduce variations. Key factors include:

- Concentration of ions: The calculation assumes standard conditions (1 M for ions), but deviations influence the actual equilibrium.
- Temperature fluctuations: Deviating from 25°C alters E° and, consequently, K.
- Activity coefficients: Non-ideal solution behavior can affect effective concentrations.
- Kinetic barriers: Thermodynamic favorability does not guarantee rapid reaction rates.

Broader Implications and Applications

Understanding the equilibrium constant for such redox reactions has practical applications across various fields:

- Electroplating and metal recovery: Leveraging spontaneous reactions to deposit metals onto surfaces.
- Corrosion science: Predicting metal stability in aqueous environments.
- Electrochemical cell design: Engineering batteries and sensors based on redox potentials.
- Material synthesis: Controlling reaction conditions for desired product yields.

Conclusion

Calculating the equilibrium constant at 25°C for the reaction $Co(s) + 2Ag^{+}(aq) \rightleftharpoons Co^{2+}(aq) + 2Ag(s)$ reveals a thermodynamically favored process with an extraordinarily large K (~10^36). This underscores the reaction's spontaneity and completeness under standard conditions, driven by the favorable electrochemical potentials of silver and cobalt.

The methodology exemplifies the integration of electrochemical data with thermodynamic principles to derive fundamental reaction parameters. Such calculations form the backbone of understanding and designing electrochemical processes, informing advances in metallurgy, energy storage, and corrosion prevention.

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