Calculate The Potential For Half Cell Containing 0.10MK2Cr2O7(aq),0.20MCr3+(aq) And 1.0104MH+(aq). The

Calculate The Potential For Half Cell Containing 0.10MK2Cr2O7(aq),0.20MCr3+(aq) And 1.0104MH+(aq). The process of determining the electrochemical potential of a half-cell is fundamental in understanding redox reactions, electrochemical series, and designing batteries and other electrochemical devices. When dealing with complex ions like dichromate ($K_2Cr_2O_7$), chromium ions (Cr^{3+}), and hydrogen ions (H^+), the calculations involve applying the Nernst equation, considering standard electrode potentials, and understanding the behavior of ions in aqueous solutions. This article provides an in-depth analysis of how to calculate the cell potential for the specified half-cell, covering essential concepts, step-by-step procedures, and practical applications — optimized for both students and professionals interested in electrochemistry.

Understanding the Components of the Half Cell

Before diving into calculations, it's crucial to understand the constituents of the half-cell involved:

1. Potassium Dichromate ($K_2Cr_2O_7$) in Solution

- Concentration: 0.10 M
- Role: Source of dichromate ions ($Cr_2O_7^{2-}$)
- Properties: Potassium dichromate is a strong oxidizing agent, especially in acidic solutions.

2. Chromium Ions (Cr^{3+}) in Solution

- Concentration: 0.20 M
- Role: Reduced form in redox reactions; can be oxidized to $Cr_2O_7^{2-}$ or reduced further depending on the reaction conditions.

3. Hydrogen Ions (H^+) in Solution

- Concentration: 1.0104 M
- Role: Common in acid-based redox reactions; involved in electrochemical processes such as hydrogen evolution or oxidation.

Key Concepts in Electrochemical Potential Calculations

Calculating the cell potential involves several core ideas:

Standard Electrode Potentials (E°)

- Values measured under standard conditions (1 M concentration, 1 atm pressure, 25°C).
- For the relevant half-reactions:
- $Cr_2O_7^{2-}/Cr^{3+}$ couple
- H^+/H_2 couple

Nernst Equation

- Adjusts standard potentials to account for non-standard conditions.
- General form:

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

where:
- E = cell potential under non-standard conditions
- E° = standard electrode potential
- n = number of electrons transferred
- Q = reaction quotient

Redox Reactions Involved

- Understanding the oxidation and reduction processes is critical.
- For dichromate in acidic solution:

$$
\text{Cr}_2\text{O}_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O
$$

- For hydrogen:

$$
2H^+ + 2e^- \rightarrow H_2
$$

Step-by-Step Calculation of the Cell Potential

The calculation proceeds through multiple stages:

1. Identify the Relevant Half-Reactions

- Oxidation or reduction of dichromate and chromium ions.
- Incorporate hydrogen ion activity where necessary.

2. Obtain Standard Electrode Potentials (E°)

- From electrochemical series tables:
- $Cr_2O_7^{2-}/Cr^{3+}$ couple: $E^\circ \approx +1.33\,V$
- H^+/H_2 couple: $E^\circ = 0\,V$ (by definition)

3. Write the Overall Cell Reaction

- For example, the reduction of dichromate coupled with hydrogen oxidation:

$$
Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O
$$
$$
H_2 \rightarrow 2H^+ + 2e^-
$$

- Balance electrons to combine reactions.

4. Calculate the Standard Cell Potential (E°cell)

- Using standard potentials:

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

- For the reduction of dichromate (cathode) and oxidation of hydrogen (anode):

$$
E^\circ_{cell} = 1.33\,V - 0\,V = +1.33\,V
$$

5. Determine the Reaction Quotient (Q)

- Based on concentrations:

$$
Q = \frac{[Cr^{3+}]^2}{[Cr_2O_7^{2-}][H^+]^{14}}
$$

- Use the provided concentrations:

$$
[Cr^{3+}] = 0.20\,M
$$
$$
[Cr_2O_7^{2-}] = \text{from } K_2Cr_2O_7 \rightarrow 0.10\,M
$$
$$
[H^+] = 1.0104\,M
$$

- Plug into Q:

$$
Q = \frac{(0.20)^2}{(0.10)(1.0104)^{14}}
$$

6. Apply the Nernst Equation

- Calculate E:

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

- For the overall reaction, $n=6$.

Practical Calculation Example

Let's perform the calculation:

- First, compute Q:

$$
Q = \frac{(0.20)^2}{0.10 \times (1.0104)^{14}} = \frac{0.04}{0.10 \times (1.0104)^{14}}
$$

- Compute $(1.0104)^{14}$:

$$
(1.0104)^{14} \approx e^{14 \times \ln(1.0104)} \approx e^{14 \times 0.01035} \approx
$$

$$
e^{0.1449} \approx 1.155
$$

- Now, Q:

$$
Q \approx \frac{0.04}{0.10 \times 1.155} = \frac{0.04}{0.1155} \approx 0.346
$$

- Next, calculate the potential:

$$
E = 1.33\,V - \frac{0.0592}{6} \times \log_{10}(0.346)
$$

$$
\log_{10}(0.346) \approx -0.460
$$

- Plug in:

$$
E = 1.33 - \frac{0.0592}{6} \times (-0.460) \approx 1.33 + 0.00493 \times 0.460 \approx 1.33 + 0.00227 \approx 1.332\,V
$$

Result: The approximate potential of the half-cell under the given conditions is about 1.33 V.

Applications of Calculated Cell Potential

Understanding and calculating the potential of such half-cells has numerous applications:

1. Designing Electrochemical Cells

- Predict whether a redox reaction will proceed spontaneously.
- Develop batteries with specific voltage outputs.

2. Analyzing Redox Reactions in Industry

- Optimize processes involving chromium compounds, such as chrome plating or pigment production.

3. Environmental Monitoring

- Detecting chromium contamination through electrochemical sensors.

4. Educational Purposes

- Teaching students about electrochemical principles and calculations.

Key Points to Remember

- Always verify the standard electrode potentials and reaction equations.
- Correctly balance redox reactions before applying the Nernst equation.
- Concentrations significantly influence the cell potential; use accurate data.
- The reaction quotient \(Q\) reflects the actual conditions, adjusting the standard potential accordingly.
- For complex ions, consider their equilibrium reactions and activity coefficients if high precision is required.

Conclusion

Calculating the potential of a half-cell containing potassium dichromate, chromium ions, and hydrogen ions requires a comprehensive understanding of electrochemical principles, accurate data, and precise application of the Nernst equation. By following a systematic approach—identifying reactions, balancing equations, calculating \(Q\), and applying the equation—you can determine the cell potential under specific conditions. Such calculations are vital in advancing electrochemical technologies, environmental monitoring, and chemical research, making them an essential skill for chemists and engineers alike.

Meta description: Learn how to calculate the electrochemical potential of a half-cell containing 0.10 M $K_2Cr_2O_7$, 0.

Frequently Asked Questions

How do you determine the standard electrode potential for the half-cell containing $Cr_2O_7^{2-}$ and Cr^{3+} ions?

The standard electrode potential can be found using standard reduction potentials from reference tables, typically involving the $Cr_2O_7^{2-}/Cr^{3+}$ couple. The relevant reduction reaction is $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$, with its standard potential used in calculations.

What is the significance of the concentrations 0.10 M for $K_2Cr_2O_7$, 0.20 M for Cr^{3+}, and 1.01×10^{-4} M for H^+ in calculating the cell potential?

These concentrations are used in the Nernst equation to determine the cell potential under non-standard conditions. The varying activities influence the equilibrium potential of each half-cell, thus affecting the overall cell potential.

How do you apply the Nernst equation to calculate the potential of this half-cell?

The Nernst equation is $E = E° - (RT/nF) \ln Q$, where $E°$ is the standard potential, R is the gas constant, T is temperature in Kelvin, n is the number of electrons transferred, F is Faraday's constant, and Q is the reaction quotient based on ion concentrations. Plugging in the given concentrations allows calculation of the potential.

Why is it important to consider the activity of H^+ ions when calculating the cell potential?

Because the concentration of H^+ ions directly affects the equilibrium position of the redox reaction involving dichromate and chromium ions. Accurate potential calculations require using activity (which accounts for ionic interactions) rather than just concentration.

What is the overall cell potential if the half-cell containing $Cr_2O_7^{2-}/Cr^{3+}$ is paired with a standard hydrogen electrode?

To find the overall cell potential, calculate the half-cell potential using the Nernst equation for the chromium half-cell, then add it to the standard hydrogen electrode potential (which is 0 V). The result gives the cell potential under the given conditions.

How does changing the concentration of H^+ from 1.01×10^{-4} M affect the calculated potential of the half-cell?

Increasing H^+ concentration shifts the equilibrium, generally increasing the cell potential, as it favors reduction reactions involving H^+. Conversely, decreasing H^+ concentration lowers the potential, as per Le Chatelier's principle and the Nernst equation.

Additional Resources

Calculate The Potential For Half Cell Containing 0.10 M $K_2Cr_2O_7$ (aq), 0.20 M Cr^{3+} (aq), and 1.01×10^{-4} M H^+ (aq): An In-Depth Analysis

---

Introduction

Electrochemistry plays a pivotal role in understanding how chemical energy can be converted into electrical energy and vice versa. At the heart of this field lie half-cell potentials, which quantify the tendency of a given species to gain or lose electrons. Calculating the potential of a specific half-cell often involves understanding the concentrations of ions, the nature of redox couples involved, and the influence of pH through hydrogen ion activity. In this article, we delve into the detailed process of calculating the electrochemical potential of a complex half-cell environment, specifically one containing potassium dichromate ($K_2Cr_2O_7$), chromium ions (Cr^{3+}), and hydrogen ions (H^+).

Understanding the Components of the Half-Cell

To accurately determine the potential, it is essential to comprehend the chemical species involved:

1. Potassium Dichromate ($K_2Cr_2O_7$)

- Role: Acts as an oxidizing agent, providing dichromate ions ($Cr_2O_7^{2-}$).
- Concentration: 0.10 M.
- Relevance: Serves as the source of $Cr_2O_7^{2-}$ ions, which participate in redox reactions.

2. Chromium Ions (Cr^{3+})

- Concentration: 0.20 M.
- Role: The reduced form in the redox couple; involved in equilibrium with dichromate ions.

3. Hydrogen Ions (H^+)

- Concentration: 1.01×10^{-4} M.
- Importance: Acidic conditions influence the redox equilibrium, especially for reactions involving dichromate and chromate ions.

Fundamental Redox Reactions and Equilibrium

1. The $Cr_2O_7^{2-}$ / Cr^{3+} Couple

The primary redox reaction involves the reduction of dichromate ions to chromium(III):

Redox Reaction:

$$\mathrm{Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O}$$

- Standard Electrode Potential (E°): +1.33 V at 25°C.
- This potential is well-established and serves as the basis for calculations.

2. The Hydrogen Ion / Hydrogen Gas Couple

The hydrogen electrode potential depends directly on the H^+ concentration (or activity):

$$\mathrm{2H^+ + 2e^- \rightarrow H_2}$$

- Standard Electrode Potential (E°): 0 V.
- The actual potential shifts based on the H^+ activity, which is related to its concentration.

Calculating the Standard Electrode Potential (E°)

The standard potential for the dichromate/Cr^{3+} couple is well tabulated as +1.33 V. However, to compute the actual potential under non-standard conditions, we must apply the Nernst Equation, which accounts for the effect of ion concentrations and activity.

Applying the Nernst Equation

1. General Form

The Nernst Equation for a redox couple is:

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

Where:

- E = electrode potential under given conditions.
- E° = standard electrode potential.
- R = universal gas constant (8.314 J mol^{-1} K^{-1}).
- T = temperature in Kelvin (assume 25°C = 298 K).
- n = number of electrons transferred (6 for dichromate reduction).
- F = Faraday's constant (96485 C mol^{-1}).
- Q = reaction quotient, based on ion activities or concentrations.

2. Simplified Logarithmic Form

At 25°C, the equation simplifies to:

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

3. Defining the Reaction Quotient (Q)

For the reaction:

$$ \mathrm{Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O} $$

The reaction quotient is:

$$ Q = \frac{[Cr^{3+}]^2}{[Cr_2O_7^{2-}][H^+]^{14}} $$

Note: Water activity is usually taken as unity.

Step-by-Step Calculation

1. Concentration Data

- $[Cr_2O_7^{2-}] = 0.10\, \text{M}$
- $[Cr^{3+}] = 0.20\, \text{M}$
- $[H^+] = 1.01 \times 10^{-4}\, \text{M}$

2. Compute Q

$$
Q = \frac{(0.20)^2}{(0.10) \times (1.01 \times 10^{-4})^{14}}
$$

Calculations:

- Numerator: $0.20^2 = 0.04$
- Denominator: $0.10 \times (1.01 \times 10^{-4})^{14}$

Calculating $(1.01 \times 10^{-4})^{14}$:

$$
\log (1.01 \times 10^{-4}) = \log 1.01 + \log 10^{-4} \approx 0.0043 - 4 = -3.9957
$$

Multiplying by 14:

$$
-3.9957 \times 14 \approx -55.9398
$$

Exponentiating:

$$
10^{-55.9398} \approx 1.15 \times 10^{-56}
$$

Now, the denominator:

$$
0.10 \times 1.15 \times 10^{-56} = 1.15 \times 10^{-57}
$$

Thus,

$$
Q \approx \frac{0.04}{1.15 \times 10^{-57}} \approx 3.48 \times 10^{55}
$$

3. Calculate $\log Q$

$$
\log Q \approx \log (3.48 \times 10^{55}) = \log 3.48 + 55 \approx 0.5416 + 55 = 55.5416
$$

4. Calculate the Electrode Potential (E)

Using the simplified Nernst Equation:

$$
E = 1.33\, \mathrm{V} - \frac{0.0592}{6} \times 55.5416
$$

Calculate:

$$
\frac{0.0592}{6} \approx 0.00987
$$

Then,

$$
E = 1.33 - 0.00987 \times 55.5416 \approx 1.33 - 0.548 \approx 0.782\, \mathrm{V}
$$

Impact of pH and Hydrogen Ion Activity

The calculated potential is sensitive to the activity of H^+ ions. The low concentration (1.01×10^{-4} M) indicates a highly acidic environment, which favors the reduction process.

Note: In cases where activity coefficients are significant, especially at higher ionic strengths, activity corrections should be applied. However, for dilute solutions, concentration approximates activity well.

Factors Influencing the Potential

1. Concentration of Dichromate and Cr^{3+}

- Increased dichromate concentration shifts the equilibrium, affecting Q and thus potential.
- Higher Cr^{3+} concentration stabilizes the reduced form, increasing potential.

2. Acidic Conditions

- Lower H^+ concentration (more acidic) enhances the forward reduction of dichromate.
- The high exponent (14) for H^+ in Q indicates a strong dependence on acidity.

3. Temperature Effects

- Elevated temperatures can alter E°, typically decreasing the potential for endothermic reductions.

- Standard calculations assume 25°C; deviations require recalculations.

Practical Applications and Significance

Understanding and calculating the potential of such a half-cell has practical implications in various fields:

- Electrochemical Cells: Designing batteries or electrochemical sensors involving chromium redox couples.
- Analytical Chemistry: Determining concentrations of chromium species via potentiometric methods.
- Environmental Monitoring: Assessing chromium speciation in water samples, crucial for toxicity evaluations.
- Industrial Processes: Optimizing chromate reduction in corrosion control or electroplating.

Conclusion

The comprehensive evaluation of the electrochemical potential of a half-cell containing potassium dichromate, Cr^{3+}, and H^{+} demonstrates the importance of precise concentration data, understanding of redox reactions, and the application of the Nernst Equation. The calculated potential (~0.78 V under the given conditions) highlights the influence of acidity and ion concentrations on redox equilibria. Such detailed

Calculate The Potential For Half Cell Containing 0 10mk2cr2o7 Aq 0 20mcr3 Aq And 1 0104mh Aq The

Calculate The Potential For Half Cell Containing 0 10mk2cr2o7 Aq 0 20mcr3 Aq And 1 0104mh Aq The

Related Articles

- Consider An Investment That Will Pay $680 Per Month For The Next 15 Years And Will Be Worth $28,000 At
- During The Terms Of Presidents Truman, Eisenhower, And Kennedy, The Relationship Between The Press And
- Cranmore Carriage Company Offers Guided Horse-drawn Carriage Rides Through Historic Columbus, Georgia.

Back to Home