For The Cell Ag(s) / Ag (aq) // Cr 2 (aq)/cr (s), What Is The Standard Cell Potential ?

Introduction

For The Cell Ag(s) / Ag (aq) // Cr^{2+}(aq)/Cr(s), What Is The Standard Cell Potential?
Understanding the standard cell potential is fundamental in electrochemistry, as it provides insight into the spontaneity and energy efficiency of electrochemical reactions. This particular cell involves a silver electrode and a chromium ion, creating a galvanic cell that can be analyzed by examining the standard reduction potentials of each half-reaction. Calculating the cell potential involves identifying the oxidation and reduction processes, referencing standard reduction potentials, and applying the Nernst equation or standard cell potential formulas. In this article, we will explore the electrochemical principles underlying this cell, determine its standard cell potential, and discuss the implications of these values in practical applications.

Fundamentals of Electrochemical Cells

What Is a Galvanic Cell?

A galvanic cell, also known as a voltaic cell, converts chemical energy into electrical energy through spontaneous redox reactions. It consists of two half-cells, each containing an electrode immersed in an electrolyte solution. Electrons flow from the anode (oxidation site) to the cathode (reduction site), generating an electric current.

Standard Electrode Potentials

Standard electrode potentials (E°) are tabulated values that measure the tendency of a species to be reduced under standard conditions (1 M concentration, 25°C, 1 atm pressure). These potentials are referenced against the standard hydrogen electrode (SHE), which is assigned an E° of 0.00 V.

Cell Potential and Spontaneity

The overall cell potential (E°_cell) is calculated by subtracting the anode potential from the cathode potential:

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

A positive E°_cell indicates a spontaneous reaction.

Identifying the Half-Reactions

Given Cell Components

The cell involves the following components:

- Anode: Solid silver (Ag(s))
- Cathode: Chromium(II) ions in aqueous solution (Cr^{2+}(aq))
- Species involved:
- Silver: Ag(s) / Ag^{+}(aq)
- Chromium: Cr^{2+}(aq) / Cr(s)

Half-Reactions

Based on the cell notation:

- Anode (oxidation): Silver metal loses electrons to form Ag^{+}(aq):

$$ \text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^- $$

- Cathode (reduction): Cr^{2+}(aq) gains electrons to form chromium metal:

$$ \text{Cr}^{2+}(aq) + 2e^- \rightarrow \text{Cr}(s) $$

Note: The number of electrons exchanged in each half-reaction must be balanced when combining.

Standard Reduction Potentials

Standard Electrode Potentials for Silver and Chromium

Referring to standard reduction potential tables (at 25°C):

- Silver:

$$ \text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s) \quad E^\circ = +0.80\, \text{V} $$

- Chromium:

$$ \text{Cr}^{2+}(aq) + 2e^- \rightarrow \text{Cr}(s) \quad E^\circ = -0.91\, \text{V} $$

Note: The value for Cr^{2+}/Cr is negative, indicating that chromium is less readily reduced than silver under standard conditions.

Important: The reduction potentials are for the reactions as written. To find the oxidation potential of silver (which is the reverse of the reduction), we need to reverse the half-reaction and change the

sign.

Calculating the Standard Cell Potential

Step 1: Assign Oxidation and Reduction Half-Reactions

- Silver acts as the anode (oxidation):

$$ \text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^- \quad E^\circ_{\text{oxidation}} = -0.80\, \text{V} $$

- Chromium acts as the cathode (reduction):

$$ \text{Cr}^{2+}(aq) + 2e^- \rightarrow \text{Cr}(s) \quad E^\circ_{\text{reduction}} = -0.91\, \text{V} $$

Note: The electrode potential for the oxidation of silver is the negative of its reduction potential.

Step 2: Balance the Electrons

Since the chromium half-reaction involves 2 electrons, and the silver half-reaction involves 1 electron, multiply the silver half-reaction by 2 to balance electrons:

- Oxidation (silver):

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

- Reduction (chromium):

$$ \text{Cr}^{2+}(aq) + 2e^- \rightarrow \text{Cr}(s) $$

Now, total electrons exchanged are balanced.

Step 3: Calculate the Standard Cell Potential

The standard cell potential is:

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

But since we have the reduction potentials for both half-reactions, and considering the oxidation of silver:

- Anode (oxidation):

$$ E^\circ_{\text{oxidation, Ag}} = -E^\circ_{\text{reduction, Ag}} = -0.80\, \text{V} $$

- Cathode (reduction, Cr^{2+}):

$$ E^\circ_{\text{reduction, Cr}} = -0.91\, \text{V} $$

Thus,

$$ E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.91\, \text{V}) - (-0.80\, \text{V}) = -0.91\, \text{V} + 0.80\, \text{V} = -0.11\, \text{V} $$

Result: The standard cell potential is -0.11 V.

Interpreting the Result

Spontaneity of the Cell

Since the calculated E°_{cell} is negative, the overall redox process is non-spontaneous under standard conditions. This means that, as written, the electrochemical cell would not generate electrical energy spontaneously.

Implications of the Negative Cell Potential

- The cell would require an external source of energy (electrical input) to drive the reaction in the opposite direction.
- The reaction favors the formation of silver metal and chromium ions in the reverse direction.

Summary of Key Points

- The half-reactions involved are:
    - Oxidation at anode: $Ag(s) \rightarrow Ag^{+}(aq) + e^-$
    - Reduction at cathode: $Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$
- Standard reduction potentials are:
    - Ag^{+}/Ag: +0.80 V
    - Cr^{2+}/Cr: -0.91 V
- The calculated standard cell potential is -0.11 V, indicating non-spontaneity.

Practical Applications and Considerations

Electrochemical Cells in Industry

Understanding the sign and magnitude of standard cell potentials guides the design of batteries, electroplating processes, and corrosion prevention strategies.

Limitations of Standard Conditions

Real-world conditions often differ from standard state, affecting cell potential. Factors such as temperature, concentration, and electrode surface area influence actual cell voltages.

Further Exploration

- Investigate other chromium oxidation states (e.g., Cr^{3+}) and their potentials.
- Explore standard potentials of Ag^{+}/Ag in different environments.
- Analyze how changing ion concentrations affects cell potential via the Nernst equation.

Conclusion

The electrochemical cell involving solid silver and aqueous chromium(II) ions has a standard cell potential of approximately -0.11 V under standard conditions. This negative value indicates that the reaction, as written, is non-spontaneous, and external energy input would be necessary to drive the process in the forward direction. Understanding these electrochemical principles provides foundational knowledge for various technological applications, from energy storage to corrosion control, emphasizing the importance of standard electrode potentials in predicting and manipulating chemical reactions.

Frequently Asked Questions

What is the standard cell potential for the cell involving Ag(s)/Ag^{+}(aq) and Cr^{2+}(aq)/Cr(s)?

The standard cell potential is approximately +1.05 V, calculated using standard reduction potentials of Ag^{+}/Ag and Cr^{2+}/Cr couples.

How do you determine the standard cell potential for the

Ag(s)/Ag^{+}(aq) and Cr^{2+}(aq)/Cr(s) cell?

You subtract the standard reduction potential of the anode from that of the cathode, using their standard reduction potentials: E°cell = E°cathode - E°anode.

Which electrode acts as the cathode and which as the anode in the Ag/Cr cell?

In this cell, silver (Ag) acts as the cathode (reduction), and chromium (Cr) acts as the anode (oxidation).

Why is the standard cell potential important in galvanic cells like the Ag/Cr cell?

It indicates the cell's voltage and spontaneity; a positive E°cell means the reaction proceeds spontaneously, providing electrical energy.

What is the significance of standard reduction potentials in calculating the cell potential for Ag and Cr electrodes?

Standard reduction potentials allow us to predict the direction of electron flow and calculate the overall cell potential accurately.

Additional Resources

For The Cell Ag(s) / Ag (aq) // Cr 2 (aq)/cr (s), What Is The Standard Cell Potential?

Understanding the intricacies of electrochemical cells is fundamental to fields ranging from electrochemistry research to practical applications like batteries and corrosion prevention. Among the key parameters that define the behavior of these cells is the standard cell potential, often denoted as E°_cell. This value indicates the maximum potential difference between two electrodes under standard conditions and predicts the spontaneity of the electrochemical reaction. In this article, we will explore in detail how to determine the standard cell potential for the cell comprising silver and chromium electrodes, specifically for the cell configuration:

Ag(s) / Ag^{+}(aq) // Cr^{3+}(aq) / Cr(s)

We will delve into the fundamental concepts, identify the relevant half-reactions, utilize standard reduction potentials, and ultimately calculate the cell's standard potential.

Understanding Electrochemical Cells and Standard Cell Potential

Before focusing on our specific cell, it’s essential to grasp foundational concepts surrounding electrochemical cells and their potentials.

What Is an Electrochemical Cell?

An electrochemical cell converts chemical energy into electrical energy through redox reactions occurring at two electrodes: an anode (oxidation) and a cathode (reduction). The overall cell potential reflects the driving force behind electron flow and is influenced by the tendencies of the redox couples involved.

Standard Conditions and Standard Electrode Potentials

The standard cell potential is measured under standard conditions:

- 1 M concentration for aqueous ions
- 1 atm pressure for gases
- The temperature typically at 25°C (298 K)

Standard reduction potentials (E°) are tabulated values that indicate how readily a species gains electrons (is reduced) compared to the standard hydrogen electrode (SHE). These potentials are crucial for predicting the spontaneity and calculating the cell potential.

Dissecting the Cell: $Ag(s) / Ag^{+}(aq) // Cr^{3+}(aq) / Cr(s)$

The cell configuration indicates the following:

- Anode (oxidation side): Silver metal (Ag(s)) is oxidized to Ag^{+} ions in solution.
- Cathode (reduction side): Chromium ions (Cr^{3+}) in solution are reduced to metallic chromium (Cr(s)).

The double slashes "//" separate the two half-cells, which are connected via a conductive pathway and a salt bridge to maintain electrical neutrality.

Step 1: Write the Half-Reactions

Oxidation Half-Reaction (Anode):

Since the anode involves oxidation of silver:

$Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$

This represents solid silver losing an electron to form silver ions.

Reduction Half-Reaction (Cathode):

Chromium ions are reduced:

$Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$

Here, chromium ions gain three electrons to become solid chromium metal.

Step 2: Obtain Standard Reduction Potentials

The next step involves consulting standard reduction potential tables:

Half-Reaction	Standard Reduction Potential, E° (V)
$Ag^+ + e^- \rightarrow Ag(s)$	+0.80
$Cr^{3+} + 3e^- \rightarrow Cr(s)$	–0.74

Note: These values are approximate and sourced from standard electrochemical tables, typically at 25°C.

Step 3: Determine the Relevant Half-Reactions

The standard potentials are given for reduction. To determine the overall cell potential, we need to ensure the oxidation and reduction reactions are correctly oriented:

- The cathode is where reduction occurs; for chromium, the reduction is straightforward.
- The anode involves oxidation of silver, which is the reverse of the reduction half-reaction.

Since the reduction potentials are given, and electrochemical principles state:

$E°_{cell} = E°_{cathode\ (reduction)} – E°_{anode\ (reduction)}$

But because the anode involves oxidation, we reverse its half-reaction and change the sign of its E° value.

Step 4: Balance the Electrons in Half-Reactions

The silver half-reaction involves 1 electron, while chromium involves 3 electrons. To balance electrons:

- Multiply the silver half-reaction by 3:

$3\ Ag(s) \rightarrow 3\ Ag^+(aq) + 3\ e^-$

- The chromium half-reaction remains:

$Cr^{3+}(aq) + 3\ e^- \rightarrow Cr(s)$

Now, the electrons cancel out when combining the two half-reactions.

Step 5: Calculate the Standard Cell Potential

Using the formula:

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

- Cathode (reduction): $Cr^{3+} + 3e^- \rightarrow Cr(s)$, $E° = –0.74$ V
- Anode (oxidation): $Ag(s) \rightarrow Ag^+ + e^-$, $E°$ (reduction) = +0.80 V

Since the anode involves oxidation, reverse the reduction potential:

$E°_{anode}$ (oxidation) = –(+0.80 V) = –0.80 V

Now, applying the formula:

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

$E°_{cell} = (–0.74\ V) – (–0.80\ V) = (–0.74\ V) + 0.80\ V = +0.06\ V$

Final Result: The Standard Cell Potential

The calculated standard cell potential for the $Ag(s)/Ag^+(aq)$ // $Cr^{3+}(aq)/Cr(s)$ cell is +0.06 volts.

This positive value indicates that, under standard conditions, the cell reaction is spontaneous, albeit with a relatively small driving force.

Significance of the Standard Cell Potential

A standard cell potential of +0.06 V suggests that:

- The cell can produce a small amount of electrical energy.
- The reaction favors the formation of chromium metal and silver ions, but only marginally.
- The cell’s efficiency and voltage output are modest, which could influence its application in practical devices.

Broader Context and Practical Implications

While this specific cell may not be ideal for high-voltage applications, understanding its potential provides insight into metal reactivity and electrochemical behavior:

- Corrosion and Metallurgy: Chromium's tendency to gain electrons reflects its corrosion resistance, especially in stainless steel.
- Electroplating: Silver’s oxidation behavior is exploited in plating processes.
- Battery Design: Small potentials like +0.06 V could be combined with other cells to develop batteries with specific voltage requirements.

Additional Considerations

- Standard conditions: Deviations in temperature, ion concentration, or pressure can influence the actual cell potential.
- Electrode surface conditions: Surface impurities or electrode shape can also impact the measured potential.
- Thermodynamic vs. Kinetic factors: A positive E°_cell indicates spontaneity, but reaction rates depend on activation energy and catalysts.

Conclusion

Determining the standard cell potential for the Ag(s)/Ag^{+}(aq) and Cr^{3+}(aq)/Cr(s) cell involves understanding the fundamental principles of electrochemistry, accurately identifying pertinent half-reactions, utilizing standard reduction potentials, and applying the appropriate calculations. The resulting value of +0.06 volts highlights a spontaneous reaction with modest energy output, exemplifying the nuanced interplay of metal reactivity and electrochemical dynamics. Such analyses underpin the development and optimization of electrochemical technologies across various scientific and industrial domains.

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