The Following Galvanic Cell Has A Potential Of 1.214 V At 25C:Hg(l)|Hg2Br2(s)|Br(0.10M)||MnO4(0.10M),Mn2+(0.10M),H+(0.10M)|Pt(s)Calculate

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Understanding the Components of the Galvanic Cell

Cell Description

The given galvanic cell can be broken down into its constituent half-cells:

- Anode (oxidation): **Hg2Br2(s)** in contact with Br^- ions (0.10 M)
- Cathode (reduction): $\mathbf{MnO_4^-}$ in contact with Mn^{2+} ions (0.10 M) and H^+ ions (0.10 M)
- Electrodes: Mercury (Hg(l)) and Platinum (Pt) as inert electrodes

The cell diagram is written as:
Hg(l) | Hg2Br2(s) | Br-(0.10 M) || MnO4-(0.10 M), Mn2+(0.10 M), H+(0.10 M) | Pt(s)

This notation indicates the phases and concentrations in each half-cell, with the double vertical line representing the salt bridge or junction.

Fundamental Principles for Calculating Cell Potential

Standard Electrode Potentials

The standard electrode potential (E°) is an intrinsic property of a half-cell, measured under standard conditions (1 M concentrations, 25°C, 1 atm). The overall cell potential (Ecell) depends on the standard potentials of each half-reaction and the conditions of the cell.

Cell Potential and 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 standard cell potential
- R is the gas constant (8.314 J/mol·K)
- T is the temperature in Kelvin (25°C = 298 K)
- n is the number of electrons transferred
- F is Faraday’s constant (96485 C/mol)
- Q is the reaction quotient

In practice, at 25°C, this simplifies to:

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

Step-by-Step Calculation of the Cell Potential

1. Identify and Write the Half-Reactions

Anode (oxidation):
The oxidation of mercury(I) bromide (Hg_2Br_2):

$$
\[

$$
\text{Hg}_2\text{Br}_2(s) + 2e^- \rightarrow 2 \text{Hg}(l) + 2 \text{Br}^-(aq)
$$

Cathode (reduction):
Reduction of permanganate ions in acidic solution:

$$
\text{MnO}_4^-(aq) + 8 \text{H}^+(aq) + 5 e^- \rightarrow \text{Mn}^{2+}(aq) + 4 \text{H}_2O
$$

---

2. Find Standard Electrode Potentials (E°)

- Mercury half-reaction:
Standard potential for Hg2Br2/Hg couple is not directly tabulated, but the standard potential for the Hg2Cl2/Hg couple is approximately +0.268 V. Since Hg2Br2 is similar, approximate E° can be sourced from literature or estimated. For this example, assume:

$$
E^{\circ}_{\text{Hg}_2\text{Br}_2/\text{Hg}} \approx +0.63\, V
$$

- Permanganate couple:
$$
E^{\circ}_{\text{MnO}_4^-/\text{Mn}^{2+}} = +1.51\, V
$$

---

3. Calculate the Standard Cell Potential (E°_{cell})

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

Since the anode involves oxidation of Hg2Br2, and the cathode involves reduction of MnO4-:

$$
E^{\circ}_{cell} = 1.51\, V - 0.63\, V = 0.88\, V
$$

---

4. Determine the Reaction Quotient (Q)

The overall cell reaction combines the two half-reactions, considering their electrons:

Combined reaction:

$$
\text{Hg}_2\text{Br}_2(s) + 2 \text{MnO}_4^-(aq) + 16 \text{H}^+(aq) + 10 \text{Br}^-(aq) \rightarrow 2 \text{Hg}(l) + 2 \text{Mn}^{2+}(aq) + 8 \text{H}_2O
$$

The reaction quotient Q depends on the concentrations of the aqueous species:

$$
Q = \frac{[\text{Mn}^{2+}]^2 \times [\text{H}^+]^8}{[\text{MnO}_4^-]^2 \times [\text{Br}^-]^{10}}
$$

Plugging in the concentrations:

$$
Q = \frac{(0.10)^2 \times (0.10)^8}{(0.10)^2 \times (0.10)^{10}} = \frac{0.01 \times 10^{-8}}{0.01 \times 10^{-10}} = \frac{10^{-10}}{10^{-12}} = 10^{2}
$$

---

5. Calculate the Cell Potential Using the Nernst Equation

Number of electrons transferred, n, is 5 (from the MnO4- reduction).

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

$$
E_{cell} = 0.88\, V - \frac{0.0592}{5} \times \log(10^{2})
$$

$$
E_{cell} = 0.88\, V - 0.01184 \times 2 = 0.88\, V - 0.02368\, V = 0.856\, V
$$

Final Result and Interpretation

The calculated cell potential under the given conditions is approximately 0.856 V, which is slightly less than the standard potential of 1.214 V originally provided — indicating that actual conditions, electrode kinetics, or approximations in standard potentials influence the observed emf.

Note: The initial potential of 1.214 V at 25°C suggests that the actual cell is operating under specific conditions or has different standard potentials. For precise calculations, experimental data for the exact half-reactions and their standard potentials would be necessary.

Importance of Calculating Cell Potential in Electrochemistry

Understanding how to calculate the cell potential is crucial for multiple applications:

- Designing batteries and fuel cells
- Predicting the spontaneity of electrochemical reactions
- Understanding corrosion processes
- Developing electrolysis processes for metal extraction

In this case, calculating the potential of the galvanic cell helps predict whether the cell can spontaneously generate electricity and how the concentrations of ions influence the emf.

Conclusion

Calculating the potential of a galvanic cell like the one described involves understanding the individual half-reactions, their standard potentials, and the effect of ion concentrations via the Nernst equation. While approximations are often necessary due to the lack of tabulated potentials for complex compounds like Hg_2Br_2, this process provides valuable insights into electrochemical processes. Accurate calculations enable chemists and engineers to optimize electrochemical systems for various industrial and scientific applications.

Frequently Asked Questions

What is the overall cell potential of the galvanic cell described at 25°C?

The overall cell potential given is 1.214 V at 25°C.

Which half-reactions are involved in this galvanic cell?

The cell involves the oxidation of mercury(I) bromide (Hg_2Br_2) to mercury (Hg) and bromide ions (Br^-), and the reduction of permanganate (MnO_4^-) to Mn^{2+} in acidic solution.

What is the significance of the standard potential in calculating the cell's behavior?

The standard potential indicates the tendency of the cell to gain or lose electrons; it helps predict the cell's voltage and spontaneity under standard conditions.

How do you determine the standard reduction potentials for the half-reactions involved?

Standard reduction potentials are obtained from standard electrochemical tables, which list the potentials for various half-reactions at 25°C under standard conditions.

What role does the concentration (0.10 M) of the reactants play in calculating the cell potential?

The concentrations influence the Nernst equation calculation, adjusting the cell potential from standard conditions to actual reaction conditions.

How is the Nernst equation applied to find the cell potential in this scenario?

The Nernst equation relates the standard potential to the reaction quotient (Q), accounting for reactant and product concentrations, to compute the actual cell potential at 25°C.

What is the expected direction of electron flow in this galvanic cell?

Electrons flow from the oxidation half-cell (Hg_2Br_2) to the reduction half-cell (MnO_4^-), indicating the anode to cathode direction.

How does the presence of Pt(s) as an inert electrode affect the cell operation?

Pt(s) serves as an inert conductor facilitating electron transfer without participating in the reaction, providing an electrode surface for reduction or oxidation.

Can this galvanic cell be used to drive non-spontaneous reactions? Why or why not?

No, because the positive cell potential (1.214 V) indicates the reaction is spontaneous under standard conditions, making it suitable for energy generation rather than driving non-spontaneous reactions.

What steps are involved in calculating the cell potential for this specific cell?

Identify the half-reactions and their standard potentials, write the overall cell reaction, use the Nernst equation with given concentrations, and compute the actual cell potential at 25°C.

Additional Resources

The Following Galvanic Cell Has a Potential of 1.214 V at 25°C: $Hg(l)|Hg_2Br_2(s)|Br^-(0.10\ M)\ ||\ MnO_4^-(0.10\ M), Mn^{2+}(0.10\ M), H^+(0.10\ M)|Pt(s)$ — An In-Depth Analytical Perspective

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Introduction

Galvanic cells, also known as voltaic cells, are electrochemical devices that convert chemical energy into electrical energy through spontaneous redox reactions. They are fundamental to various technological applications, from batteries powering portable electronics to corrosion prevention systems. The specific galvanic cell described here features a complex arrangement involving a mercury-based electrode, mercurous bromide, and multiple aqueous solutions containing permanganate, manganese ions, and protons, with platinum serving as an inert electrode.

This article aims to dissect this particular galvanic cell configuration, analyze its potential of 1.214 V at 25°C, and explore the underlying electrochemical principles governing its operation. By doing so, we will illuminate the intricacies of calculating cell potentials, understanding the roles of each component, and applying thermodynamic principles to interpret the cell's behavior.

Section 1: Understanding the Cell Components and Configuration

1.1 The Anode and Cathode Designation

In any galvanic cell, the anode is where oxidation occurs, and the cathode is where reduction takes place. Based on the cell notation:

Anode side: $Hg(l) \mid Hg_2Br_2(s) \mid Br^-$ (0.10 M)
Cathode side: MnO_4^- (0.10 M), Mn^{2+} (0.10 M), H^+ (0.10 M) | Pt(s)

The double vertical bar "||" indicates the salt bridge or porous junction separating the two half-cells.

Implication:

- The presence of mercury (Hg) and mercurous bromide (Hg_2Br_2) suggests a mercurous-based electrode.
- The cathode side involves permanganate reduction, manganese ions, and protons, all in aqueous solution, with platinum as an inert electrode facilitating electron transfer.

1.2 The Role of Mercury and Hg_2Br_2

Mercuric (Hg) and mercurous bromide (Hg_2Br_2) form a half-cell that involves the equilibrium of mercury species with bromide ions. This setup resembles a mercurous bromide electrode, which can function as a reference or an active electrode depending on the potential.

1.3 The Electrolyte Solutions

The solutions in the cathode half-cell contain:

- Permanganate ions (MnO_4^-): Strong oxidizing agents capable of being reduced to Mn^{2+}.
- Manganese ions (Mn^{2+}): The reduced form after MnO_4^- reduction.
- Protons (H^+): Facilitate acidic reduction conditions.

The concentrations are equal (0.10 M), indicating a controlled environment for standard potential calculations.

1.4 Inert Electrode: Platinum (Pt)

Platinum acts as an inert conductor, providing a surface for electrode reactions without participating chemically, essential especially for redox reactions involving aqueous ions.

Section 2: Calculating the Cell Potential

2.1 Standard Cell Potential (E°)

The cell potential provided is 1.214 V at 25°C. This value corresponds to the cell's electromotive force (EMF) under standard conditions, considering all components are at 1 M concentrations and 25°C.

2.2 Nernst Equation Overview

To analyze the cell's behavior under non-standard conditions, the Nernst equation is used:

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

Where:

- E = cell potential under given conditions
- E° = standard cell potential
- R = universal gas constant (8.314 J·mol⁻¹·K⁻¹)
- T = temperature in Kelvin (here, 25°C = 298 K)
- n = number of electrons transferred
- F = Faraday's constant (96485 C·mol⁻¹)
- Q = reaction quotient

Section 3: Determining the Standard Electrode Potentials

3.1 The Anode Reaction: Mercury / Mercurous Bromide

Reaction:

$$
\text{Hg}_2\text{Br}_2(s) + 2e^- \rightarrow 2\text{Hg}(l) + 2\text{Br}^-
$$

This is a reduction; the reverse (oxidation) occurs at the anode:

$$
2\text{Hg}(l) \rightarrow \text{Hg}_2\text{Br}_2(s) + 2e^-
$$

The standard potential for this half-cell, based on standard reduction potentials, is associated with the mercurous bromide electrode.

Standard potential of Hg_2Br_2 (Mercurous bromide) electrode:

From electrochemical tables, the potential of the Hg_2Br_2 / Hg electrode is approximately +0.64 V vs. SHE (Standard Hydrogen Electrode).

3.2 The Cathode Reaction: Permanganate Reduction

In acidic solution:

$$
\text{MnO}_4^- + 8H^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4H_2O
$$

Standard reduction potential:

$$
E^\circ_{\text{MnO}_4^-/\text{Mn}^{2+}} = +1.51\, \text{V}
$$

since the reaction occurs in an acidic medium (H^+ present).

3.3 Calculating the Standard Cell Potential (E°)

The overall cell potential:

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

Using the standard potentials:

$$
E^\circ_{\text{cell}} = (+1.51\, V) - (+0.64\, V) = +0.87\, V
$$

This is the theoretical potential assuming ideal standard conditions.

Note: The measured potential is higher (1.214 V), indicating that actual conditions, concentrations, and electrode specifics influence the cell potential.

Section 4: Applying the Nernst Equation

4.1 Calculating the Reaction Quotient Q

The overall cell reaction involves:

$$
\text{MnO}_4^- + 8H^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4H_2O
$$

and the mercurous bromide half-reaction:

$$
\text{Hg}_2\text{Br}_2 + 2e^- \rightarrow 2\text{Hg} + 2\text{Br}^-
$$

The electrons are balanced over combined reactions, but for calculation purposes, in the overall cell, the net reaction involves 5 electrons (from permanganate reduction) and 2 electrons (from mercurous bromide oxidation). To balance electrons:

$$
\text{Multiply the mercurous bromide reaction by 5}:
$$

\]
\[
5 \text{Hg}_2\text{Br}_2 + 10 e^- \rightarrow 10 \text{Hg} + 10 \text{Br}^-
\]

and the permanganate reduction by 2:

\[
2 \text{MnO}_4^- + 16 H^+ + 10 e^- \rightarrow 2 \text{Mn}^{2+} + 8 H_2O
\]

Overall reaction:

\[
2 \text{MnO}_4^- + 16 H^+ + 5 \text{Hg}_2\text{Br}_2 \rightarrow 2 \text{Mn}^{2+} + 8 H_2O + 10 \text{Hg} + 10 \text{Br}^-
\]

The reaction quotient \(Q\) involves the concentrations of the aqueous ions:

\[
Q = \frac{[\text{Mn}^{2+}]^2}{[\text{MnO}_4^-]^2 [H^+]^{16} [\text{Br}^-]^{10}}
\]

Given all concentrations are 0.10 M, the calculation simplifies.

4.2 Calculation of Cell Potential at Non-Standard Conditions

Using the Nernst equation:

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

- \(E^\circ = 1.214\, V\) (given)
- \(n = 10\) electrons transferred in the overall balanced reaction
- \(T = 298\, K\)

Substituting known values:

\[
\frac{RT}{nF} = \frac{8.314 \times 298}{10 \times 96485} \approx 0.00255\, V
\]

Since all concentrations are 0.10 M, their ratios are unity, so:

\[
Q = 1
\]

thus,

\[
E = E^\circ - (0.002

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