What Is The Emf Of A Cell Consisting Of A Pb2 /pb Half-cell And A Pt/h /h2 Half-cell If [pb2] = 0.10

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Understanding the electromotive force (emf) of electrochemical cells is fundamental in electrochemistry, as it provides insight into the cell's ability to do electrical work. In this article, we will analyze the emf of a specific cell comprising a lead half-cell (Pb^{2+}/Pb) and a platinum electrode with hydrogen gas (Pt/H_2) under particular conditions, specifically when the concentration of lead ions, [Pb^{2+}], is 0.10 mol/L. We will explore the principles behind calculating emf, the relevant electrochemical equations, and the significance of the given concentration.

Introduction to Electrochemical Cells

Electrochemical cells convert chemical energy into electrical energy through redox reactions. They consist of two half-cells connected via an external circuit, each containing an electrode and an electrolyte solution. The potential difference between these electrodes determines the emf of the cell, which can be calculated using standard reduction potentials and the Nernst equation when conditions deviate from standard state.

Components of the Cell

The cell in question comprises:

- Pb^{2+}/Pb Half-Cell: A lead electrode immersed in a solution containing lead ions at a concentration of 0.10 mol/L.
- Pt/H_2 Half-Cell: A platinum electrode exposed to hydrogen gas at a certain pressure, usually 1 atm, serving as a standard hydrogen electrode (SHE) or a similar hydrogen half-cell.

The overall cell reaction involves the oxidation of lead and the reduction of hydrogen ions, depending on the conditions and potentials.

Understanding Standard Electrode Potentials

Each half-cell has an associated standard reduction potential (E°), which measures its tendency to gain electrons under standard conditions (1 M concentration, 1 atm pressure, 25°C).

- Pb^{2+}/Pb Electrode: The standard reduction potential is approximately -0.13 V.
- Standard Hydrogen Electrode (SHE): Defined as 0 V by convention.

However, when the concentration of Pb^{2+} differs from 1 M, the actual potential shifts according to the Nernst equation.

Calculating the Cell Potential (E_{cell})

The emf of the cell depends on the individual electrode potentials and can be calculated as:

Ecell = Ecathode - Eanode

In this specific setup:

- The Pb^{2+}/Pb electrode acts as either the anode or cathode depending on the direction of the reaction.
- The Pt/H_2 electrode (standard hydrogen electrode) acts as the reference, with potential 0 V under standard conditions.

Note: For simplicity, assume the Pb^{2+}/Pb half-cell acts as the cathode (reduction of Pb^{2+} to Pb), and the hydrogen half-cell acts as the anode or cathode based on the reaction.

Applying the Nernst Equation

The Nernst equation relates the electrode potential to the ion concentration:

$E = E° - (RT/nF) \ln([Red]/[Ox])$

Or, at 25°C (298 K):

$E = E° - (0.0592 \text{ V} / n) \log([Red]/[Ox])$

For the Pb^{2+}/Pb half-cell:

- The reduction reaction: $Pb^{2+} + 2e^- \rightarrow Pb(s)$
- Standard potential, E° = -0.13 V
- Number of electrons transferred, n = 2

The potential of the lead half-cell at a given concentration:

$E_{Pb^{2+}} = E° – (0.0592\ V / 2) \log([Pb^{2+}])$

Given $[Pb^{2+}] = 0.10$ mol/L, the calculation becomes:

$E_{Pb^{2+}} = -0.13\ V – (0.0296\ V) \log(0.10)$

Since $\log(0.10) = -1$:

$E_{Pb^{2+}} = -0.13\ V – (0.0296\ V)(-1) = -0.13\ V + 0.0296\ V = -0.1004\ V$

The hydrogen electrode, being a standard hydrogen electrode, has a potential of 0 V under standard conditions. When hydrogen gas is at 1 atm, the potential remains 0 V. If hydrogen gas is at a different pressure, the Nernst equation applies for the hydrogen electrode as well:

$E_{H_2} = 0 – (0.0592\ V) \log(P_{H_2} / 1\ atm)$

Assuming hydrogen gas is at 1 atm, $E_{H_2} = 0$ V.

Note: If the pressure differs, adjust accordingly.

Calculating the Cell emf

Assuming the Pb^{2+}/Pb half-cell acts as the cathode (reduction) and the hydrogen electrode as the anode (oxidation), the emf is:

$E_{cell} = E_{Pb^{2+}/Pb} – E_{H_2}$

Plugging in the values:

$E_{cell} = (–0.1004\ V) – (0\ V) = –0.1004\ V$

A negative emf indicates the cell, as constructed under these conditions, favors the reverse reaction. To get the emf in the conventional sense (positive value when the cell operates as written), we interpret the signs appropriately.

Final emf value:

- $E_{cell} \approx –0.10$ V

or, considering the direction of electron flow, the cell potential is approximately 0.10 V in the reduction direction.

Implications and Significance

Understanding the emf at non-standard concentrations allows chemists to predict how cells perform under various conditions. The decrease in emf from the standard value (which would be slightly more positive at 1 M) illustrates the influence of ion concentration on cell potential. This is vital in real-world applications where conditions are rarely ideal.

Summary of Key Points

- The emf of a cell depends on the standard reduction potentials of its half-cells and the actual conditions, especially ion concentrations.
- The Nernst equation enables calculation of the cell potential at non-standard conditions, factoring in concentration and pressure differences.
- In the given scenario, with $[Pb^{2+}]$ = 0.10 mol/L, the cell emf is approximately 0.10 V, slightly less than the standard potential.
- Understanding these calculations helps in designing batteries, electrolysis processes, and other electrochemical applications.

Conclusion

The emf of a lead and hydrogen electrode cell with a lead ion concentration of 0.10 mol/L is approximately 0.10 volts. This value illustrates how deviations from standard conditions affect cell potential. By applying the principles of electrochemistry and the Nernst equation, scientists and engineers can accurately predict and manipulate cell behavior for various technological applications, from energy storage to industrial electrolysis. Recognizing the relationship between ion concentration and emf is fundamental to advancing electrochemical science and optimizing practical systems.

Frequently Asked Questions

What is the overall emf of a cell with a Pb^{2+}/Pb half-cell and a Pt/H_2 half-cell when $[Pb^{2+}]$ = 0.10 M?

The emf can be calculated using the standard reduction potentials and the Nernst equation. It depends on the standard potentials of the half-cells and the concentration of Pb^{2+}; typically, for Pb^{2+}/Pb, the standard potential is -0.13 V.

How do you calculate the emf of a cell with non-

standard concentrations such as $[Pb^{2+}]$ = 0.10 M?

Use the Nernst equation: $E = E° - (0.0592/n) \log([Pb^{2+}])$, where E° is the standard reduction potential, and n is the number of electrons transferred (2 for Pb^{2+}/Pb).

What is the standard reduction potential for the Pb^{2+}/Pb half-cell?

The standard reduction potential for $Pb^{2+} + 2e^- \rightarrow Pb(s)$ is approximately -0.13 V.

How does changing the concentration of Pb^{2+} affect the emf of the cell?

Lowering $[Pb^{2+}]$ decreases the emf due to the Nernst equation, which accounts for concentration effects on cell potential.

What role does the Pt electrode play in the H_2/H^+ half-cell?

The Pt electrode acts as an inert conductor facilitating the transfer of electrons in the hydrogen half-cell, where H_2 gas is involved.

How do you determine the emf of the cell if $[Pb^{2+}]$ = 0.10 M?

Calculate E° for the Pb^{2+}/Pb half-cell, then use the Nernst equation with $[Pb^{2+}]$ = 0.10 M, and add this to the standard potential of the H_2/H^+ half-cell (which is 0 V at standard conditions).

Why is the Pt/H_2 half-cell considered a standard hydrogen electrode?

Because it involves H_2 gas at 1 atm pressure and H^+ ions, with a platinum electrode, serving as a reference electrode with zero potential under standard conditions.

What is the significance of the concentration of Pb^{2+} in calculating the emf of the cell?

The concentration directly influences the cell potential via the Nernst equation; a lower concentration results in a lower emf, reflecting the cell's electrochemical driving force.

Can the emf of the cell be negative under these conditions?

Yes, if the concentration of Pb^{2+} is sufficiently low and the standard potentials favor reduction at the Pt/H_2 electrode, the calculated emf could be negative, indicating the cell favors the reverse reaction.

Additional Resources

Emf of a Cell Consisting of a Pb^{2+}/Pb Half-Cell and a Pt/H_2 Half-Cell with [Pb^{2+}] = 0.10 M

Understanding the electromotive force (emf) of electrochemical cells is fundamental to electrochemistry. When dealing with a cell comprising a lead(II)/lead half-cell and a platinum/hydrogen half-cell, the emf provides insights into the cell's voltage potential, which is crucial for applications ranging from batteries to electroplating. This article offers an in-depth analysis of such a cell, focusing on calculating its emf under given conditions, exploring the underlying principles, and discussing practical considerations.

Introduction to the Cell Components

Electrochemical cells are composed of two half-cells, each involving a specific redox couple. In our case, the cell consists of:

- Pb^{2+}/Pb Half-Cell: A lead electrode immersed in a solution containing lead ions.
- Pt/H_2 Half-Cell: A platinum electrode in contact with hydrogen gas, often used as an inert electrode, serving as a reference or for hydrogen evolution/oxidation.

The combination forms a galvanic cell where spontaneous redox reactions generate an emf that can be harnessed for electrical work.

Understanding the Cell Reactions

To analyze the emf, we need to identify the relevant half-reactions:

Lead half-cell (reduction):
$$ \text{Pb}^{2+} + 2e^- \rightarrow \text{Pb} $$
Standard reduction potential:
$$ E^\circ_{\text{Pb}^{2+}/\text{Pb}} = -0.13\, \text{V} $$

Hydrogen half-cell (standard):
$$ 2H^+ + 2e^- \rightarrow H_2 $$
Standard reduction potential:
$$ E^\circ_{H^+/H_2} = 0.00\, \text{V} $$

In the cell, the lead electrode will undergo oxidation, and the hydrogen electrode will undergo reduction (or vice versa) depending on the emf.

Calculating the Standard Cell Potential

The standard emf ($E^\circ_{(cell)}$) is calculated as:

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

Typically, the cathode is where reduction occurs. For the lead half-cell, if it acts as the anode:

$$ \text{Pb} \rightarrow \text{Pb}^{2+} + 2e^- \quad (E^\circ_{ox} = +0.13\, V) $$

And the hydrogen electrode acts as the cathode:

$$ 2H^+ + 2e^- \rightarrow H_2 \quad (E^\circ = 0.00\, V) $$

Therefore, the standard cell potential:

$$ E^\circ_{cell} = 0.00\, V - (-0.13\, V) = +0.13\, V $$

This reflects the maximum emf under standard conditions (1 M concentration, 1 atm pressure).

Effect of Concentration: Nernst Equation

Since the problem specifies $[Pb^{2+}] = 0.10\, M$, the actual emf will deviate from the standard potential. The Nernst equation allows us to compute the emf under non-standard conditions:

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

Where:
- $R = 8.314\, \text{J mol}^{-1}\text{K}^{-1}$
- $T = 298\, \text{K}$ (assuming room temperature)
- $n = 2$ electrons transferred
- $F = 96485\, \text{C mol}^{-1}$
- Q is the reaction quotient

For the lead half-reaction (oxidation), the reaction quotient involves lead ion concentration:

$$ Q = \frac{1}{[Pb^{2+}]} $$

Since the hydrogen electrode is at standard conditions, its activity remains 1.

Calculating the emf with Nernst Equation

Using the simplified form at 25°C (298 K):

$$ E_{cell} = E^\circ_{cell} - \frac{0.0592\, \text{V}}{n} \log Q $$

Plugging in the values:

$$ E_{cell} = 0.13\, V - \frac{0.0592\, V}{2} \log \left(\frac{1}{0.10} \right) $$

$$ E_{cell} = 0.13\, V - 0.0296\, V \times \log(10) $$

Since $\log(10) = 1$:

$$ E_{cell} = 0.13\, V - 0.0296\, V \times 1 = 0.1004\, V $$

Therefore, the emf of the cell under the given concentration is approximately 0.10 V.

Features and Significance of the Result

Features:

- The emf decreases from the standard 0.13 V to about 0.10 V due to the decreased concentration of Pb^{2+} ions.
- The cell potential is directly influenced by ion concentration, illustrating Le Chatelier’s principle in action.
- The potential remains positive, indicating the cell can perform work spontaneously under these conditions.

Significance:

- This calculation exemplifies how concentration impacts cell emf, critical for designing batteries and electrochemical sensors.
- Understanding the Nernst equation helps in predicting cell behavior during electrolysis or corrosion processes.
- The relatively low emf suggests this cell is suitable for low-voltage applications or as a reference electrode.

Practical Considerations and Limitations

Pros:

- The lead electrode is inexpensive and stable, making this setup versatile for educational demonstrations.
- The cell's emf can be easily tuned by altering ion concentrations.
- The Pt electrode's inertness allows for reactions involving gases without electrode degradation.

Cons:

- Lead is toxic; handling and disposal require caution.
- The emf is relatively low, limiting high-power applications.
- The emf calculation assumes ideal behavior; real-world factors like activity coefficients and overpotentials can cause deviations.

Features to Optimize:

- Maintaining stable concentrations to ensure consistent emf.
- Using high-purity reagents to minimize interference.
- Ensuring proper electrode surface area and contact for accurate measurements.

Conclusion

The emf of a cell consisting of a Pb^{2+}/Pb half-cell and a Pt/H_2 half-cell with $[Pb^{2+}] = 0.10\, M$ is approximately 0.10 V. This value stems from the standard potential adjusted for the concentration of lead ions via the Nernst equation, illustrating the fundamental relationship between ion activity and cell potential. Such calculations are vital for understanding electrochemical systems, optimizing battery performance, and designing sensors. While the cell offers simplicity and cost-effectiveness, considerations around toxicity, low emf, and real-world conditions should guide practical applications.

In summary:

- The standard emf for the cell is around 0.13 V.
- At $[Pb^{2+}] = 0.10\, M$, the emf reduces to approximately 0.10 V.
- The Nernst equation effectively predicts the impact of ion concentrations on cell potential.
- Practical use involves balancing cost, safety, and desired voltage output.

This detailed exploration underscores the importance of electrochemical principles in both theoretical and applied chemistry, providing a foundation for further study and innovation in energy storage and conversion technologies.

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