Calculate The Mass Of Cl2 Consumed If The Battery Delivers A Constant Current Of 957 A For 72.0 Min.

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Understanding how to determine the mass of chlorine gas (Cl_2) consumed during electrolysis processes is fundamental in chemistry, especially in industrial applications such as chlorine production. When a battery or power source supplies a steady current, it drives the electrochemical reactions, leading to the formation or consumption of chemical species. In this article, we will explore how to calculate the mass of Cl_2 gas produced or consumed when a constant current is applied over a specific duration. The principles outlined here are essential for students, chemists, and engineers involved in electrochemical processes.

Introduction to Electrolysis and Chlorine Production

Electrolysis is a process that uses an electric current to drive a non-spontaneous chemical reaction. In the context of chlorine production, electrolysis of sodium chloride (NaCl) solution or brine results in the formation of chlorine gas at the anode, sodium hydroxide at the cathode, and hydrogen gas under certain conditions.

The relevant half-reaction at the anode (oxidation of chloride ions) is:

\[2Cl^- (aq) \rightarrow Cl_2 (g) + 2e^- \]

This reaction illustrates that two chloride ions produce one molecule of chlorine gas, releasing two electrons.

Understanding the Basic Concepts

Before delving into calculations, it's important to understand the core concepts involved:

- Current (I): The flow of electric charge, measured in amperes (A).

- Time (t): Duration of electrolysis, measured in seconds.

- Faraday's Laws of Electrolysis: These laws relate the amount of substance deposited or liberated at an electrode to the quantity of electricity passed through the electrolyte.

- Faraday's Constant (F): The charge of one mole of electrons, approximately 96,485 coulombs (C).

- Moles of Electrons (n): The number of moles of electrons transferred during the process.

- Molar mass of Cl_2: Approximately 70.90 g/mol.

Step-by-Step Calculation Process

To determine the mass of chlorine gas consumed or produced, follow these steps:

Step 1: Convert Time from Minutes to Seconds

Since current is in amperes (coulombs per second), convert the given time:

$$ t = 72.0\, \text{minutes} \times 60\, \text{seconds/minute} = 4320\, \text{seconds} $$

Step 2: Calculate Total Charge Passed (Q)

Using the formula:

$$ Q = I \times t $$

where:

- $ I = 957\, \text{A} $
- $ t = 4320\, \text{seconds} $

Thus,

$$ Q = 957\, \text{A} \times 4320\, \text{s} = 4,139,040\, \text{C} $$

Step 3: Determine Moles of Electrons Transferred

Using Faraday's constant:

$$ \text{moles of electrons} = \frac{Q}{F} $$

where:

- $ F = 96,485\, \text{C/mol} $

So,

$$ \text{moles of electrons} = \frac{4,139,040\, \text{C}}{96,485\, \text{C/mol}} \approx 42.9\, \text{mol} $$

Step 4: Relate Moles of Electrons to Moles of Cl_2

From the half-reaction:

$$ 2Cl^- \rightarrow Cl_2 + 2e^- $$

- 2 moles of electrons produce 1 mole of Cl_2.
- Therefore, the moles of Cl_2 produced or consumed are:

$$ \text{moles of Cl}_2 = \frac{\text{moles of electrons}}{2} $$

$$ \text{moles of Cl}_2 = \frac{42.9}{2} \approx 21.45\, \text{mol} $$

Step 5: Calculate the Mass of Cl_2

Using molar mass:

$$ \text{Mass} = \text{moles} \times \text{molar mass} $$

$$ \text{Mass of Cl}_2 = 21.45\, \text{mol} \times 70.90\, \text{g/mol} \approx 1,521\, \text{g} $$

Final Result

The mass of Cl_2 consumed (or produced) during the electrolysis under the specified conditions is approximately 1,521 grams.

This calculation assumes 100% efficiency, meaning all of the current contributes directly to chlorine production without any side reactions or losses.

Additional Factors and Considerations

While the above calculation provides a theoretical value, real-world conditions may differ due to various factors:

- Electrolyte purity: Impurities can affect current efficiency.

- Electrode efficiency: Not all electrons necessarily produce chlorine; some may lead to side reactions.

- Temperature and pressure: Variations can influence gas volume but generally do not affect the mass unless gas escapes.

- Current efficiency: Often less than 100%, which means actual chlorine produced could be less than theoretical calculations suggest.

Applications of the Calculation

Understanding how to compute the mass of chlorine produced from a given current and time is vital in several industrial and laboratory contexts:

- Chlorine manufacturing: Ensuring proper scaling of production equipment.

- Electrochemical research: Designing experiments and understanding reaction yields.

- Environmental safety: Estimating emissions during electrolysis processes.

- Quality control: Verifying that electrolysis systems operate within specified parameters.

Practice Problems for Reinforcement

1. If a battery supplies a current of 500 A for 3 hours, what is the mass of Cl_2 produced?

2. How does increasing the current from 957 A to 1000 A affect the amount of Cl_2 produced over the same time?

3. If the actual yield of Cl_2 is only 80% of the theoretical yield, what is the actual mass produced in the original problem?

Conclusion

Calculating the mass of chlorine gas involved in electrolysis processes is a straightforward application of fundamental electrochemistry principles and Faraday's laws. By converting current and time into charge, then relating that charge to moles of electrons, and finally to moles and mass of Cl_2, chemists can estimate production quantities accurately. Such calculations are essential for optimizing industrial processes, ensuring safety, and designing efficient electrochemical systems.

Keywords: electrolysis, chlorine gas, Cl_2, current, Faraday's law, molar mass, electrochemical reaction, mass calculation, industrial chemistry, electrochemical process

Frequently Asked Questions

How do you calculate the mass of Cl_2 gas produced from a given current and time?

You can calculate the mass of Cl_2 by first determining the total charge (Q) using Q = current (I) × time (t), then using Faraday's law to find the moles of Cl_2 produced, and finally converting moles to mass using the molar mass of Cl_2.

What is the significance of the 957 A current in calculating Cl_2 mass in electrolysis?

The current indicates the rate of electron flow, which directly relates to the amount of substance being reduced or oxidized during electrolysis, enabling calculation of the amount of Cl_2 gas generated.

How do you convert 72.0 minutes into seconds for this calculation?

Multiply 72.0 minutes by 60 seconds per minute: $72.0 \times 60 = 4320$ seconds, to use in the charge calculation.

What is the molar mass of chlorine gas (Cl_2), and why is it important here?

The molar mass of Cl_2 is approximately 70.9 g/mol, and it’s essential for converting moles of Cl_2 produced into the mass in grams.

What is the step-by-step formula to find the mass of Cl_2 consumed in this electrolysis process?

First, calculate total charge: $Q = I \times t$; then, determine moles of Cl_2: moles = $(Q / (n \times F))$, where n is the number of electrons exchanged per molecule (2 for Cl_2), and F is Faraday’s constant (~96485 C/mol); finally, multiply moles by molar mass (70.9 g/mol) to get the mass.

Additional Resources

Calculate The Mass Of Cl_2 Consumed If The Battery Delivers A Constant Current Of 957 A For 72.0 Min.

Introduction

Electrolysis, a fundamental process in modern chemistry and industrial applications, involves the use of electrical energy to drive non-spontaneous chemical reactions. Among the many reactions facilitated by electrolysis, the production of chlorine gas (Cl_2) from aqueous sodium chloride (NaCl) solutions is one of the most significant due to its vast industrial utility, including disinfectants, plastics, and other chemicals.

This article aims to explore a detailed, investigative approach to calculating the mass of chlorine gas produced during electrolysis under specified conditions. Specifically, we analyze a scenario where a battery delivers a steady current of 957 amperes over a duration of 72.0 minutes. Such an analysis is vital not only for understanding the underlying electrochemical principles but also for optimizing industrial processes, ensuring safety, and maintaining economic efficiency.

Background and Theoretical Foundations

The Process of Electrolysis of Sodium Chloride Solution

In aqueous solution, electrolysis of sodium chloride proceeds according to the following half-reactions:

- At the anode (oxidation):
$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

- At the cathode (reduction):
$2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

The overall reaction thus involves the conversion of chloride ions into chlorine gas and water into hydrogen gas at the respective electrodes.

The key to calculating the mass of Cl_2 generated is understanding the relationship between the electric charge passed and the amount of substance produced, which is rooted in Faraday's laws of electrolysis.

Faraday's Laws of Electrolysis

Faraday's first law states that the mass (m) of a substance altered at an electrode during electrolysis is directly proportional to the total electric charge (Q) passing through the electrolyte:

\[m = \frac{Q \times M}{n \times F} \]

Where:
- m = mass of substance produced (grams)
- Q = total electric charge (coulombs)
- M = molar mass of the substance (g/mol)
- n = number of electrons exchanged per ion (for Cl_2, n=2)
- F = Faraday's constant (~96485 C/mol)

This fundamental relation allows us to connect the measurable electrical parameters (current and time) to the amount of chemical product formed.

Methodology for Calculating Chlorine Mass

The calculation involves several steps:

1. Converting Current and Time into Total Charge (Q).
2. Applying the Electrolysis Equation to Find Moles of Cl_2.
3. Calculating the Mass of Cl_2 from Moles.

Let's explore these steps systematically.

Step 1: Convert Current and Time into Total Charge

The total charge Q delivered by the battery can be calculated as:

$$Q = I \times t$$

Where:
- I = current in amperes (A)
- t = time in seconds (s)

Given:
- $I = 957\, \text{A}$
- $t = 72.0\, \text{min}$

First, convert minutes to seconds:

$$t = 72.0\, \text{min} \times 60\, \text{s/min} = 4320\, \text{s}$$

Now, calculate Q:

$$Q = 957\, \text{A} \times 4320\, \text{s} = 4,136,640\, \text{C}$$

This is the total electric charge passed through the system.

Step 2: Determine Moles of Chlorine Gas Produced

Using Faraday's law, the number of moles of Cl_2 produced:

$$\text{Moles of Cl}_2 = \frac{Q}{n \times F}$$

Where:
- $n = 2$ (since 2 electrons are involved per Cl_2 molecule)
- $F = 96485\, \text{C/mol}$

Plugging in the values:

$$\text{Moles of Cl}_2 = \frac{4,136,640}{2 \times 96485}$$

Calculate denominator:

$$2 \times 96485 = 192,970$$

Calculate moles:

$$\text{Moles of Cl}_2 = \frac{4,136,640}{192,970} \approx 21.44\, \text{mol}$$

This indicates approximately 21.44 moles of chlorine gas are generated under the specified conditions.

Step 3: Calculate the Mass of Chlorine Gas

The molar mass of chlorine gas (Cl_2) is:

$$M_{Cl_2} = 2 \times M_{Cl} = 2 \times 35.45\, \text{g/mol} = 70.90\, \text{g/mol}$$

Using the number of moles:

$$\text{Mass} = \text{moles} \times M_{Cl_2}$$

$$\text{Mass} = 21.44\, \text{mol} \times 70.90\, \text{g/mol} \approx 1,521.6\, \text{g}$$

Final Result:

The mass of Cl_2 consumed (or produced) during this electrolysis process is approximately 1521.6 grams.

Factors Affecting Accuracy and Practical Considerations

While the calculation appears straightforward, several real-world factors can influence the actual amount of chlorine gas produced:

- Electrolyte Concentration: Variations in NaCl concentration can alter the efficiency.

- Electrode Conditions: Surface area, material, and cleanliness impact current efficiency.
- Side Reactions: Hydrogen evolution or oxygen evolution might reduce the yield of Cl_2.
- Temperature and Pressure: Changes can influence gas volume but not mass; however, in practical settings, gas collection may be affected.

Understanding these factors is crucial for industrial applications where precision and efficiency are paramount.

Industrial Implications and Safety Considerations

The production of chlorine gas is a highly controlled process owing to its toxicity and corrosiveness. Accurate calculations of the amount produced are essential for:

- Designing Safety Protocols: Preventing leaks and managing exposure.
- Optimizing Production: Ensuring cost-effective operation.
- Environmental Management: Handling and neutralizing by-products properly.

Industrial facilities utilize continuous current monitoring and process controls to maintain consistent gas production levels, aligning with calculated values to ensure safety and efficiency.

Conclusion

The detailed analysis of the electrolysis process under the specified conditions demonstrates that delivering a constant current of 957 A over 72.0 minutes results in the production of approximately 1521.6 grams of chlorine gas. This calculation underscores the importance of precise electrical measurements, fundamental electrochemical principles, and an understanding of practical variables affecting industrial electrolysis processes.

By thoroughly understanding and applying Faraday’s laws, chemists and engineers can accurately predict product yields, optimize operational parameters, and ensure safety in chemical manufacturing environments. This investigation not only elucidates a specific calculation but also highlights the critical interplay between electrical energy and chemical transformation—a cornerstone of modern electrochemical technology.

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This comprehensive review emphasizes the importance of understanding electrochemical principles for practical applications and provides a clear, methodical approach to calculating the mass of chlorine gas produced during electrolysis.

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