Stoichiometry of Electrolytic Cells

In an electrolytic cell, the amount of electrons that pass through the cell determines the masses of substances that react or are produced at the electrodes. For example, when zinc is plated onto a steel pipe to galvanize it, two moles of electrons must be gained by one mole of zinc ions to deposit one mole of zinc atoms as metal.

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

This relationship establishes a mole ratio of electrons to zinc metal (2:1). Unfortunately, you cannot directly measure the number of electrons. Instead, we measure the electric current (I) being applied to the electrolytic cell and use it to calculate the amount of charge (Q) that passes through the cell in a certain amount of time (t).

$$Q = It$$

Current is measured in amps (A), charge in Coulombs (C), and time in seconds (s).

Example 1
A certain electrolytic cell uses 300 kA (kiloamps) of current. How much charge passes through this cell in 24 hours?

Faraday's Law

The relationship between electricity and electrochemical changes was first investigated by Michael Faraday in the 1830s. He discovered that the amount of an element produced or consumed at the electrode of an electrolytic cell was directly proportional to the time the cell operated, as long as the current was constant. This was called **Faraday's Law**.

Furthermore, he found that 96500 Coulombs of charge is transferred for every mole of electrons that flows through the cell. This value is known as Faraday's constant (*F*).

$$F = 96500\ C/mol$$

This constant can be used to determine the number of moles of electrons that flow through a cell.

$$moles\ of\ e^- = \frac{Q}{F} \qquad \text{or} \qquad moles\ of\ e^- = \frac{It}{F}$$

Example 2
What amount of electrons flows through an electrolytic cell that operates for $1.25\ h$ at a current of $0.15\ A$?

You can also use Faraday's law to determine the amount (moles or mass) of an element produced or consumed at the electrode of an electrolytic cell. The procedure is outlined below:

1. Write the half reaction that results in the production/consumption of the desired element.

2. Calculate the number of moles of electrons that pass through the cell.

$$moles\ of\ e^- = \frac{It}{F}$$

3. Convert the moles of electrons to moles of the desired element using a mole ratio.

4. Convert the moles of the desired element to mass, if necessary, using the molar mass.

Example 3
What is the mass of copper deposited at the cathode of a copper electrorefining cell operated at 12 A for 40 minutes?

This type of question may also be worked backwards, where you are given the mass of the element produced, and asked to determine the current required. The procedure to solve such a question would simply be the reverse of that used in Example 3.

1. Write the half reaction that results in the production/consumption of the desired element.

2. Convert the mass of the desired element to moles.

3. Convert the moles of the desired element into moles of electrons using a mole ratio.

4. Use Faraday's law to determine the required current.

$$I = \frac{F \cdot moles\ of\ e^{-}}{t}$$

Example 4
In a silver electroplating cell, $0.175\ g$ of silver is to be deposited from a silver cyanide solution in a time of 10 minutes. Predict the current required.

Worksheet

1. Calculate the charge transferred by a current of $1.5\ A$ flowing for $30\ s$.

2. How long, in minutes, does it take a current of $1.6\ A$ to transfer a charge of $375\ C$?

3. Convert a current of $1.74\ A$ for 10 minutes into an amount in moles of electrons.

4. How long, in minutes, will it take a current of $3.5\ A$ to transfer $0.1\ mol$ of electrons?

5. An electroplating cell operates for 35 minutes with a current of $1.9\ A$. Calculate the amount, in moles, of electrons transferred.

6. A cell transferred $0.146\ mol$ of electrons with a constant current of $1.24\ A$. How long, in hours, did this take?

7. Calculate the current required to transfer $0.015\ mol$ of electrons in a time of 20 minutes.

8. A student reconstructs Volta's battery using sheets of copper and zinc, and a current of $0.5\ A$ is produced for 10 minutes. Calculate the mass of zinc that is oxidized to form aqueous zinc ions.

9. A car bumper is plated with chromium using chromium(III) ions in solution. If a current of $54\ A$ flows in the cell for 45 minutes, determine the mass of chromium deposited on the bumper.