

Concept QuickStart – Electrochemical Cells

Unit: Unit 2: Electrochemistry

Subject: For CBSE Class 12 Chemistry

SECTION 1: UNDERSTANDING THE CONCEPT

1.1 What Is an Electrochemical Cell? (Core Idea and Anchor Definition)

An electrochemical cell is a device that serves as the bridge between chemical energy and electrical energy. It operates on a fundamental principle of energy conversion:

- It can generate electricity from the energy released during a spontaneous chemical reaction (a **galvanic cell** or **voltaic cell**).
- Conversely, it can use electrical energy from an external source to drive a chemical reaction that would not happen on its own (an **electrolytic cell**).

At its core, an electrochemical cell facilitates redox (reduction-oxidation) reactions, where the transfer of electrons between chemical species is harnessed or forced, resulting in the flow of electric current.

1.2 Why Electrochemical Cells Matter

The study of electrochemical cells is critical for both theoretical understanding and immense practical applications. The principles of electrochemistry are not confined to the laboratory; they are integral to modern life and industry.

- **Industrial Production:** Many essential chemicals and metals, such as sodium hydroxide, chlorine, fluorine, and aluminum, are produced on a massive scale using electrochemical methods.
- **Portable Energy:** All forms of batteries and fuel cells are practical applications of electrochemical cells, converting stored chemical energy into electrical energy to power countless devices.
- **Sustainable Technology:** Electrochemical reactions can be highly energy-efficient and less polluting than conventional methods, making them vital for developing eco-friendly technologies.
- **Biological Processes:** Nature itself is an expert in electrochemistry. The transmission of sensory signals through our nerve cells to the brain and the communication between cells are based on electrochemical principles.

1.3 Why This Concept Exists

The concept of the electrochemical cell exists to explain and control the interconversion of chemical and electrical energy. It provides a framework for understanding how a spontaneous chemical process, which naturally releases energy, can be configured to produce a useful electrical current instead of just heat. Fundamentally, the electrochemical cell provides a way to harness the Gibbs free energy (ΔG) of a spontaneous reaction as useful electrical work, rather than letting it dissipate merely as heat. It also explains how we can supply electrical energy to "reverse" a non-spontaneous reaction, forcing chemicals to form products they otherwise wouldn't. This duality is central to the entire field of electrochemistry.

1.4 Analogies and Mental Image

To visualize an electrochemical cell, picture the **Daniell cell**, a classic example. It is not a single container but a system of two connected halves.

- **Two Separate Beakers (Half-Cells):** Imagine two beakers. In one, a zinc metal strip (an *electrode*) is submerged in a solution of zinc sulfate (*electrolyte*). In the other, a copper electrode is submerged in a solution of copper sulfate.
- **The External Connection:** A metal wire connects the zinc and copper electrodes. This wire acts as a path for electrons to travel from one half-cell to the other.
- **The Internal Connection:** The two beakers are connected by a U-shaped tube filled with an inert electrolyte, known as a *salt bridge*.

The "spontaneous" reactions create a flow of electrons. The zinc electrode has a stronger tendency to dissolve into Zn^{2+} ions, leaving electrons behind on the metal strip. The copper electrode has a weaker tendency to do this. This difference creates an "electron pressure," or potential difference, that "pushes" electrons from the zinc electrode through the wire to the copper electrode, where they are consumed by Cu^{2+} ions from the solution.

The **salt bridge** is essential for sustaining this flow. As the reaction proceeds, positive Zn^{2+} ions build up in the anode beaker and positive Cu^{2+} ions are depleted from the cathode beaker. This charge imbalance would quickly halt the electron flow. The salt bridge prevents this by allowing negative ions (anions) to flow into the anode beaker and positive ions (cations) into the cathode beaker, neutralizing the charge and completing the electrical circuit.

1.5 Everyday Context and Applications

The most direct and relatable application of electrochemical cells is in **batteries and fuel cells**. Every time you use a remote control, a laptop, a smartphone, or start a car, you are using a galvanic cell. These devices are compact, portable power sources that package

spontaneous chemical reactions, ready to convert their chemical energy into electrical energy on demand.

SECTION 2: WHAT THE TEXTBOOK SAYS (NCERT)

The NCERT textbook provides the authoritative framework for understanding electrochemical cells as required by the CBSE curriculum. This section distills the core definitions, principles, and examples presented in the textbook, forming the essential knowledge base for your studies and examinations.

2.1 NCERT Key Statements

Based on the official text, the following fundamental principles define the operation and analysis of galvanic cells:

- **Defining a Galvanic Cell:** A galvanic (or voltaic) cell is a specific type of electrochemical cell that transforms the chemical energy from a spontaneous redox reaction into usable electrical energy.
- **The Half-Cell:** The basic unit of a galvanic cell is the half-cell, which is constructed by dipping a metallic electrode into a solution of its own ions (the electrolyte).
- **Anode and Cathode Roles:** The **anode** is the electrode where oxidation (loss of electrons) occurs; in a galvanic cell, it has a negative potential. The **cathode** is the electrode where reduction (gain of electrons) occurs, and it has a positive potential. This potential arises from a competition at the electrode-electrolyte interface: metal atoms tend to dissolve and leave electrons behind (making the electrode negative), while metal ions in solution tend to deposit and take electrons (making it positive). In the Daniell cell, zinc's tendency to leave electrons is stronger, making the anode negative, while at the copper electrode, the deposition of Cu^{2+} ions dominates, making the cathode positive relative to the anode.
- **Direction of Flow:** In the external circuit, electrons always flow from the negative electrode (anode) to the positive electrode (cathode). By convention, the direction of current flow is opposite, from the cathode to the anode.
- **The Universal Reference:** The potential of a single half-cell cannot be measured in isolation. Therefore, the Standard Hydrogen Electrode (SHE) is used as a universal reference, and its potential is arbitrarily defined as exactly zero volts at all temperatures.
- **Calculating Cell Potential (EMF):** The electromotive force (emf) of a cell, defined as the potential difference when no current is drawn, is calculated as the difference between the reduction potentials of the two half-cells. By convention, it is the

potential of the right-hand electrode (cathode) minus the potential of the left-hand electrode (anode):

2.2 NCERT Examples and Distinctions

The NCERT text uses a primary example and a key distinction to clarify these principles.

The Daniell Cell: A Primary Example

The Daniell cell is the quintessential model of a galvanic cell. It is constructed with a zinc electrode in a zinc sulfate solution and a copper electrode in a copper sulfate solution. When connected, a spontaneous redox reaction occurs. At the anode, zinc metal is oxidized ($\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$), and at the cathode, copper ions are reduced and deposited as copper metal ($\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$). This transfer of electrons generates an electrical potential. Under standard conditions (1 mol dm^{-3} , which is equivalent to 1 M or 1 mole per litre, for both solutions), the cell has a potential of **1.1 volts**.

Key Distinction: Galvanic vs. Electrolytic Cells

The function of an electrochemical cell can be reversed by applying an opposing external voltage (E_{ext}). Using the Daniell cell (potential = 1.1 V) as an example, its behavior changes dramatically:

- **When $E_{\text{ext}} < 1.1 \text{ V}$:** The cell functions normally as a **galvanic cell**. Electrons flow from zinc to copper, and the spontaneous chemical reaction proceeds, generating current.
- **When $E_{\text{ext}} = 1.1 \text{ V}$:** The external voltage perfectly balances the cell's potential. There is **no flow of electrons** or current, and the chemical reaction stops completely. The system is at equilibrium.
- **When $E_{\text{ext}} > 1.1 \text{ V}$:** The external voltage overpowers the cell's natural potential and forces the reaction to run in reverse. The cell now functions as an **electrolytic cell**. Electrons are forced to flow from copper to zinc. **Zinc is deposited at the zinc electrode and copper dissolves at the copper electrode**—a non-spontaneous reaction.

This distinction highlights that the same physical setup can act as either a galvanic or an electrolytic cell, depending on whether it is producing electricity or consuming it to drive a reaction.

This framework from the NCERT textbook provides the essential vocabulary and formulas. The next section focuses on simplifying these concepts for better clarity and long-term memory.

SECTION 3: CLARITY AND MEMORY

3.1 Key Clarity Lines

To master this topic, focus on these core, non-negotiable statements derived directly from electrochemical principles.

Galvanic Cell: Converts *spontaneous* chemical reaction energy into electrical energy.

Electrolytic Cell: Uses electrical energy to drive a *non-spontaneous* chemical reaction.

Within a **Galvanic Cell**:

- The **Anode** is where **Oxidation** occurs and is the **Negative** electrode.
- The **Cathode** is where **Reduction** occurs and is the **Positive** electrode.
- **Electron Flow:** From Anode to Cathode (negative to positive).
- **Current Flow:** From Cathode to Anode (positive to negative).
- **Cell Potential Formula:** $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$

3.2 How to Remember Electrochemical Cells

Instead of memorizing isolated facts, follow the official IUPAC conventions explained in the NCERT text. These rules provide a logical structure for analyzing any galvanic cell.

1. **The Left-Right Convention:** When representing a cell, the **anode is always written on the left** and the **cathode is always on the right**.
 - Example: $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu(s)}$
 - Here, Zinc is the anode (left) and Copper is the cathode (right).
2. **Identify the Chemical Process:** Oxidation always occurs at the anode, and reduction always occurs at the cathode. Following the left-right rule, the reaction on the left is oxidation, and the reaction on the right is reduction.
3. **Determine the Flow:** Since electrons are lost in oxidation (at the anode) and gained in reduction (at the cathode), they must flow from the **anode to the cathode** in the external wire. This corresponds to a flow from **left to right** in the cell notation.
4. **Calculate the Potential:** The formula $E_{\text{cell}} = E_{\text{right}} - E_{\text{left}}$ is a direct consequence of this convention. You simply take the standard reduction potential of the half-cell written on the right and subtract the standard reduction potential of the half-cell written on the left.

A positive E_{cell} value is the electrochemical signature of a spontaneous reaction (one that can proceed on its own and do work). This is why a galvanic cell, which relies on a spontaneous reaction, must have a positive overall potential.

Essential Memory Aids (Mnemonics)

- **"OIL RIG"**: **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain (of electrons).
- **"AN OX"** and **"RED CAT"**: The **A**node is for **O**xidation; **R**eduction occurs at the **C**athode.
- **"LOAN"** (for galvanic cells): **L**eft side is **O**xidation, **A**node, and **N**egative electrode.



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