

Concept QuickStart – Nernst Equation and Its Applications

Unit: Unit 2: Electrochemistry

Subject: For CBSE Class 12 Chemistry

Section 1: Understanding the Concept

Before we dive into the formal definitions and mathematical derivations, it's essential to build an intuitive grasp of the Nernst Equation. This section sets the stage by exploring what the equation represents at a conceptual level, why it is a necessary tool in electrochemistry, and how it connects the theoretical world of standard conditions to the practical reality of chemical reactions.

This initial understanding provides the context needed to appreciate the formal textbook definitions and examples that follow.

Section 2: What the Textbook Says (NCERT)

This section distills the core principles, equations, and solved examples related to the Nernst Equation as presented in the standard NCERT textbook. Mastering this material is crucial, as it forms the foundational knowledge required to solve problems accurately in board examinations and competitive tests. We will focus on the key formulas and their direct applications.

2.1 NCERT Key Statements

Based on the provided NCERT text, the Nernst Equation and its related concepts can be summarized through the following key principles:

- **Primary Purpose:** The Nernst equation provides a method to calculate the electrode potential or cell potential under non-standard conditions, specifically when the concentration of the species involved in the electrode reaction is not unity (1 M).
- **General Form for an Electrode:** For a general reduction half-reaction $M^{n+}(aq) + ne^{-} \rightarrow M(s)$, the Nernst equation relates the electrode potential $E(M^{n+}/M)$ to the standard electrode potential $E^{\circ}(M^{n+}/M)$ as follows, where the concentration of the solid metal $[M]$ is taken as unity (as the concentration or activity of a pure solid is considered constant): $E(M^{n+}/M) = E^{\circ}(M^{n+}/M) - (RT/nF) \ln(1/[M^{n+}])$ Here, R is the gas constant ($8.314 \text{ JK}^{-1} \text{ mol}^{-1}$), T is the temperature in Kelvin, n is the number of moles of electrons transferred in the reaction, and F is the Faraday constant (96487 C mol^{-1}).

- **Application to a Full Galvanic Cell:** To understand its application to a complete cell, we can use the Daniell cell as the primary example ($\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$). For this specific reaction, the Nernst equation is written as $E(\text{cell}) = E^\circ(\text{cell}) - (RT/2F) \ln ([\text{Zn}^{2+}]/[\text{Cu}^{2+}])$. For any general electrochemical reaction, the concentration term is expressed as the reaction quotient, Q . The general formula is: $E(\text{cell}) = E^\circ(\text{cell}) - (RT/nF) \ln Q$
- **Simplified Equation at 298 K (25°C):** By substituting the values for R , F , and T (298 K) and converting the natural logarithm (\ln) to a base-10 logarithm (\log), the equation for cell potential is simplified for practical calculations: $E(\text{cell}) = E^\circ(\text{cell}) - (0.059/n) \log Q$
- **Effect of Ion Concentration:** The cell potential $E(\text{cell})$ is directly dependent on the concentrations of the reacting ions. For the Daniell cell, the $E(\text{cell})$ increases when the concentration of the reactant ions (Cu^{2+}) increases or when the concentration of the product ions (Zn^{2+}) decreases.
- **Condition of Equilibrium:** When a galvanic cell's reaction reaches equilibrium, the cell can no longer perform work, and its potential $E(\text{cell})$ drops to zero. At this point, the reaction quotient Q becomes equal to the equilibrium constant, K_c .

2.2 NCERT Examples and Distinctions

The NCERT textbook clarifies the application of the Nernst equation through solved examples and by linking it to fundamental thermodynamic concepts like Gibbs free energy and the equilibrium constant.

A key solved example demonstrates the calculation of cell potential under non-standard conditions for the reaction: $\text{Mg(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{Ag(s)}$

Given the concentrations $[\text{Mg}^{2+}] = 0.130 \text{ M}$ and $[\text{Ag}^+] = 0.0001 \text{ M}$, and $E^\circ(\text{cell}) = 3.17 \text{ V}$, the actual cell potential $E(\text{cell})$ is calculated. The reaction quotient involves squaring the concentration of Ag^+ because its stoichiometric coefficient is 2. The formula is applied as: $E(\text{cell}) = E^\circ(\text{cell}) - (0.059/2) \log ([\text{Mg}^{2+}]/[\text{Ag}^+]^2)$ This calculation shows how significantly the cell potential can deviate from its standard value due to changes in ion concentrations.

The Nernst equation also allows us to draw crucial distinctions and connections between electrochemical potential and other thermodynamic properties:

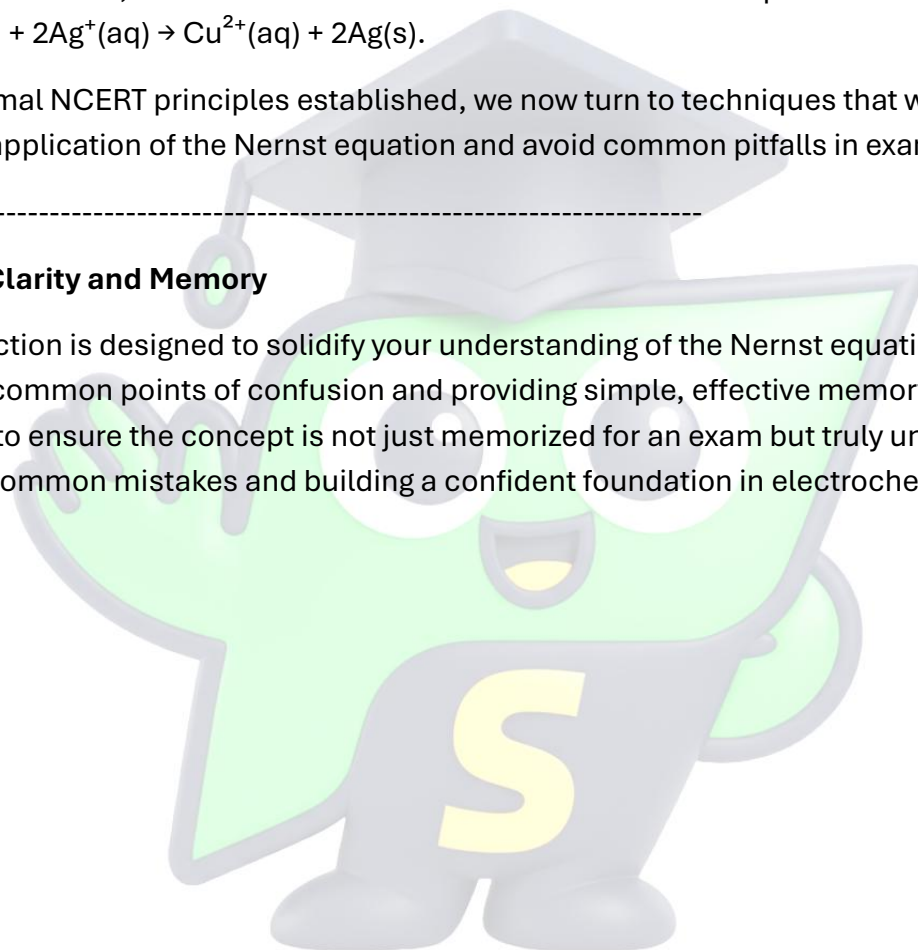
- **Cell Potential and Gibbs Energy:** The maximum electrical work a galvanic cell can do (reversible work) is equal to the decrease in its Gibbs free energy ($\Delta_r G$). This relationship is expressed by the formula: $\Delta_r G = -nFE(\text{cell})$. A positive cell potential ($E(\text{cell}) > 0$) results in a negative Gibbs free energy ($\Delta_r G < 0$), which signifies a spontaneous reaction. Under standard conditions, the relationship is $\Delta_r G^\circ = -nFE^\circ(\text{cell})$.

- **Standard Potential and Equilibrium Constant:** At equilibrium, two conditions are met: the cell potential $E(\text{cell})$ is zero, and the reaction quotient Q becomes the equilibrium constant K_c . Substituting these into the Nernst equation gives $0 = E^\circ(\text{cell}) - (RT/nF) \ln K_c$. Rearranging this shows a direct link between the *standard* cell potential and the equilibrium constant: $E^\circ(\text{cell}) = (2.303RT/nF) \log K_c$. This powerful relationship allows for the calculation of equilibrium constants from standard potential data, which is demonstrated in the textbook's Example 2.2 for the reaction $\text{Cu(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{Ag(s)}$.

With the formal NCERT principles established, we now turn to techniques that will help you master the application of the Nernst equation and avoid common pitfalls in examinations.

Section 3: Clarity and Memory

This final section is designed to solidify your understanding of the Nernst equation by addressing common points of confusion and providing simple, effective memory aids. The objective is to ensure the concept is not just memorized for an exam but truly understood, preventing common mistakes and building a confident foundation in electrochemistry.



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