GALVANIC CELLS AND ELECTRODE PROCESSES: PRACTICAL APPLICATIONS AND COMMON SYSTEMS

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Abstract:

Galvanic cells are devices that convert chemical energy into electrical energy through redox reactions. These systems form the basis for batteries and accumulators widely used in modern technology. This article explores the theory of electrode processes in galvanic elements, introduces the Nernst equation, and illustrates practical examples such as the Daniell cell. Tables and diagrams are provided to aid understanding.

Keywords:

Galvanic cell, Daniell cell, Nernst equation, electrode potential, electrochemical series, accumulator, battery

1. Introduction

Galvanic cells play a crucial role in converting chemical energy into electrical energy. They are employed in various technological devices including batteries, portable electronics, and energy storage systems. The principle behind a galvanic cell is the redox reaction occurring between two different electrodes immersed in electrolyte solutions.

2. The Structure and Working Principle of Galvanic Cells

A typical galvanic cell consists of two different metals (electrodes) placed in electrolyte solutions, connected externally by a conductor and internally via a salt bridge. When a redox reaction occurs, electrons flow from the anode (oxidation site) to the cathode (reduction site), generating electric current.

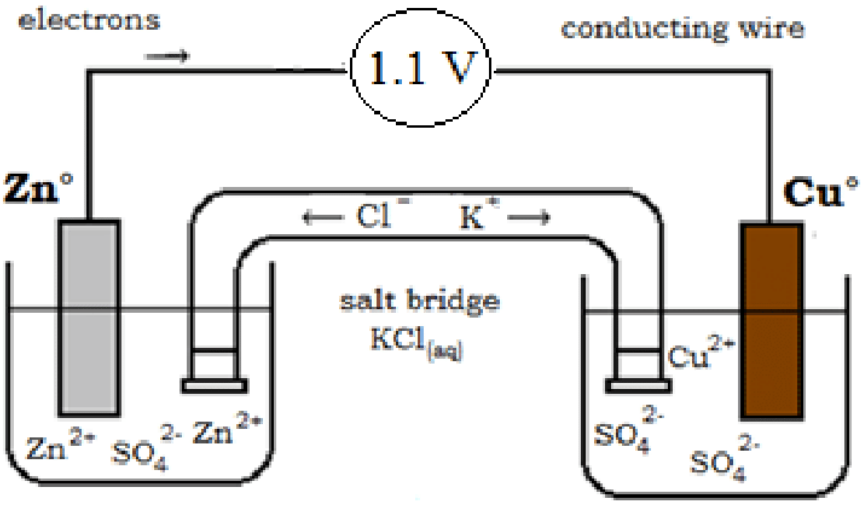

Figure 1. Schematic of a Daniell Cell:

Zn | ZnSO4 || CuSO4 | Cu

- Zn electrode (anode) undergoes oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$

- Cu^{2+} ions are reduced at the Cu electrode (cathode): $Cu^{2+} + 2e^- \rightarrow Cu$

Overall reaction: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

3. Electrode Potential and the Nernst Equation

When a metal is immersed in its ion solution, a potential difference is established between the metal and the solution due to the transfer of ions. This potential is termed the electrode potential.

Nernst Equation:

$$E = E° + (0.059/n) * \log([Ox]/[Red])$$

For a Daniell cell:

$$E = E°(Cu) - E°(Zn) + (0.059/2) * \log([Cu^{2+}]/[Zn^{2+}])$$

4. Types of Electrodes

Electrodes can be classified into:

- Primary electrodes: Made of active metals that participate in reactions.
- Secondary electrodes: Inert metals (e.g., platinum, gold) that provide surface

for reactions.

- Gas electrodes, amalgam electrodes, oxide-reduction electrodes, and glass electrodes are also common in laboratory and industrial setups.

5. Practical Galvanic Systems: Batteries and Accumulators

Galvanic principles are employed in various batteries:
- Lead-acid accumulator: Used in vehicles; based on Pb/PbO_2 and sulfuric acid.
- Alkaline batteries: Use zinc and manganese dioxide.
- Lithium-ion batteries: Found in electronics.

6. Reversibility and Charging of Galvanic Cells

Some galvanic cells, like accumulators, are reversible. When an external current is applied, the chemical reactions reverse, restoring the original materials. This principle is fundamental to rechargeable batteries.

7. Conclusion

Galvanic elements are fundamental to electrochemistry and practical applications. Understanding their working principles, including electrode potential and Nernst equation, allows efficient use in energy storage technologies. Advances in battery chemistry continue to improve the efficiency and sustainability of galvanic systems.

Table 1. Standard Electrode Potentials:

Electrode	Half-Reaction	E° (V)
Zn^{2+}/Zn	$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76
Cu^{2+}/Cu	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34

EMF of the Daniell Cell:
E_cell = 0.34 - (-0.76) = 1.10 V

Table 2. Comparison of Common Galvanic Cells:

Type	Anode	Cathode	Electrolyte	EMF (V)
Daniell Cell	Zn	Cu	$ZnSO_4$/$CuSO_4$	1.10
Lead-acid Battery	Pb	PbO_2	H_2SO_4	~2.0
Alkaline Battery	Zn	MnO_2	KOH	~1.5
Li-ion Battery	Li	Various metals	Li salt in organic sol.	3.6–3.7

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