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Chemistry project on electrochemical cell pdf

A voltaic cell is a device that produces an electric current from energy released by a spontaneous redox reaction in two half-cells. Recall that reduction happens at the cathode and oxidation happens at the anode in a voltaic cell

Key Takeaways

Key Points

Oxidation describes the loss of electrons by a molecule, atom, or ion. Reduction describes the gain of electrons by a molecule, atom, or ion. The electrons always flow from the anode to the cathode. The half-cells are connected by a salt bridge that allows the ions in the solution to move from one half-cell to the other, so that the reaction can continue.

Key Terms

redox: A reversible chemical reaction in which one reaction is an oxidation and the reverse is a reduction.

half-cell: Either of the two parts of an electrochemical cell containing an electrode and an electrolyte.

voltaic cell: A cell, such as in a battery, in which an irreversible chemical reaction generates electricity; a cell that cannot be recharged.

An electrochemical cell is a device that produces an electric current from energy released by a spontaneous redox reaction. This kind of cell includes the galvanic, or voltaic, cell, named after Luigi Galvani and Alessandro Volta. These scientists conducted several experiments on chemical reactions and electric current during the late 18th century. Electrochemical cells have two conductive electrodes, called the anode and the cathode. The anode is defined as the electrode where oxidation occurs. The cathode is the electrode where reduction takes place. Electrodes can be made from any sufficiently conductive materials, such as metals, semiconductors, graphite, and even conductive polymers. In between these electrodes is the electrolyte, which contains ions that can freely move. The voltaic cell uses two different metal electrodes, each in an electrolyte solution. The anode will undergo oxidation and the cathode will undergo reduction. The metal of the anode will oxidize, going from an oxidation state of 0 (in the solid form) to a positive oxidation state, and it will become an ion. At the cathode, the metal ion in the solution will accept one or more electrons from the cathode, and the ion's oxidation state will reduce to 0. This forms a solid metal that deposits on the cathode. The two electrodes must be electrically connected to each other, allowing for a flow of electrons that leave the metal of the anode and flow through this connection to the ions at the surface of the cathode. This flow of electrons is an electrical current that can be used to do work, such as turn a motor or power a light.

Example Reaction

The operating principle of the voltaic cell is a simultaneous oxidation and reduction reaction, called a redox reaction. This redox reaction consists of two half-reactions. In a typical voltaic cell, the redox pair is copper and zinc, represented in the following half-cell reactions: Zinc electrode (anode): $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ Copper electrode (cathode): $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$

The cells are constructed in separate beakers. The metal electrodes are immersed in electrolyte solutions. Each half-cell is connected by a salt bridge, which allows for the free transport of ionic species between the two cells. When the circuit is complete, the current flows and the cell "produces" electrical energy.

A galvanic, or voltaic, cell: The cell consists of two half-cells connected via a salt bridge or permeable membrane. The electrodes are immersed in electrolyte solutions and connected through an electrical load. Copper readily oxidizes zinc; the anode is zinc and the cathode is copper. The anions in the solutions are sulfates of the respective metals. When an electrically conducting device connects the electrodes, the electrochemical reaction is: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

The zinc electrode produces two electrons as it is oxidized ($\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$), which travel through the wire to the copper cathode. The electrons then find the Cu^{2+} in solution and the copper is reduced to copper metal ($\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$). During the reaction, the zinc electrode will be used and the metal will shrink in size, while the copper electrode will become larger due to the deposited Cu that is being produced. A salt bridge is necessary to keep the charge flowing through the cell. Without a salt bridge, the electrons produced at the anode would build up at the cathode and the reaction would stop running. Voltaic cells are typically used as a source of electrical power. By their nature, they produce direct current.

A battery is a set of voltaic cells that are connected in parallel. For instance, a lead–acid battery has cells with the anodes composed of lead and cathodes composed of lead dioxide. Electrolysis uses electrical energy to induce a chemical reaction, which then takes place in an electrolytic cell. Recall the three components necessary to construct an electrolytic cell

Key Takeaways

Key Points

Electrometallurgy is the process of reducing metals from metallic compounds to obtain the pure form of the metal using electrolysis. Electrolysis can sometimes be thought of as running a non-spontaneous galvanic cell. Electrodes of metal, graphite, and semiconductor material are widely used in electrolysis. Other systems that utilize the electrolytic process are used to produce metallic sodium and potassium, chlorine gas, sodium hydroxide, and potassium and sodium chlorate. Cell notation is shorthand that expresses a certain reaction in an electrochemical cell. Produce the appropriate electrochemical cell notation for a given electrochemical reaction

Key Points

The cell anode and cathode (half-cells) are separated by two bars or slashes, which represent a salt bridge. The anode is placed on the left and the cathode is placed on the right. Individual solid, liquid, or aqueous phases within each half-cell are written separated by a single bar. Concentrations of dissolved species can be written in the parentheses after the phase notation (s, l, g, or aq).

Key Terms

half-cell: Either of the two parts of an electrochemical cell containing an electrode and an electrolyte.

electrode: The terminal through which electric current passes between metallic and nonmetallic parts of an electric circuit. In electrolysis, the electrodes are placed in the solution separately. Recall that standard cell potentials can be calculated from potentials E°_{cell} for both oxidation and reduction reactions. A positive cell potential indicates that the reaction proceeds spontaneously in the direction in which the reaction is written. Conversely, a reaction with a negative cell potential proceeds spontaneously in the reverse direction.

$$E^\circ_{\text{cell}} = E^\circ_{\text{reduction}} + E^\circ_{\text{oxidation}}$$

Cell notations are a shorthand description of voltaic or galvanic (spontaneous) cells. The reaction conditions (pressure, temperature, concentration, etc.), the anode, the cathode, and the electrode components are all described in this unique shorthand. Recall that oxidation takes place at the anode and reduction takes place at the cathode. When the anode and cathode are connected by a wire, electrons flow from anode to cathode. A typical galvanic cell: A typical arrangement of half-cells linked to form a galvanic cell. Using the arrangement of components, let's put a cell together. One beaker contains 0.15 M $\text{Cd(NO}_3)_2$ and a Cd metal electrode. The other beaker contains 0.20 M AgNO_3 and a Ag metal electrode. The net ionic equation for the reaction is written: $2\text{Ag}^+(\text{aq}) + \text{Cd(s)} \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{Ag(s)}$ In the reaction, the silver ion is reduced by gaining an electron, and solid Ag is the cathode. The cadmium is oxidized by losing electrons, and solid Cd is the anode. The anode reaction is: $\text{Cd(s)} \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{e}^-$ The cathode reaction is: $2\text{Ag}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Ag(s)}$

Cell Notation Rules

1. The anode half-cell is described first; the cathode half-cell follows. Within a given half-cell, the reactants are specified first and the products last. The description of the oxidation reaction is first, and the reduction reaction is last; when you read it, your eyes move in the direction of electron flow. Spectator ions are not included.

2. A single vertical line (|) is drawn between two chemical species that are in different phases but in physical contact with each other (e.g., solid electrode | liquid with electrolyte). A double vertical line (||) represents a salt bridge or porous membrane separating the individual half-cells.

3. The phase of each chemical (s, l, g, aq) is shown in parentheses. If the electrolytes in the cells are not at standard conditions, concentrations and/or pressure, they are included in parentheses with the phase notation. If no concentration or pressure is noted, the electrolytes in the cells are assumed to be at standard conditions (1.00 M or 1.00 atm and 298 K). Using these rules, the notation for the cell we put together is: $\text{Cd(s)} | \text{Cd}^{2+}(\text{aq}, 0.15\text{ M}) || \text{Ag}^+(\text{aq}, 0.20\text{ M}) | \text{Ag(s)}$

Learn how to set up a galvanic cell and determine the EMF

This helpful video from the University of Birmingham links electrochemistry to metals more broadly and identifies displacement reactions as a useful precursor to electrochemical cells. The Daniell cell is a particular type of galvanic cell using copper and zinc cathodes in beakers of copper sulfate and zinc sulfate respectively. The following two videos from amritacreate and SpaceyScience show the same experiment using slightly different apparatus. Comparing and contrasting the different set-ups can be a helpful way to probe students' understanding of the underlying chemistry and why one set-up might be preferable to another. Questions you can ask your students: What is the purpose of the salt bridge? What can be used instead of an inverted U-tube for the salt bridge? Why are the experimenters using solutions with concentrations of 1 mol dm⁻³? What would the voltmeter reading be if you switched the connections, connecting the copper rod to the negative terminal and the zinc rod to the positive terminal? Write down the half equation taking place at each cathode and write the overall redox equation for the cell. It's helpful for students to determine the EMF for various electrochemical cells and then compare their results with the calculated E°_{cell} using tables of standard electrode potentials. Why might their experimental results be different? $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ It's also worth being aware of microscale approaches to electrochemical cells, with one possible technique being demonstrated in the video below from the NCSSM. This microscale approach allows you to easily, quickly and inexpensively compare multiple 'cells' using a variety of metal pieces as cathodes. Also check out... Nailing corrosion demonstrations – Declan Fleming presents an experiment to illustrate the electrochemistry of rusting using phenolphthalein and hexacyanoferrate(III) as indicators. Explaining electrolytes explicitly – the latest research indicates that students' grasp of electrolyte chemistry may not be as well developed as we assume. How to teach displacement reactions – this CPD article offers advice on how to help students develop an understanding of the reactivity of metals; a useful precursor to electrochemistry. Battery power – the perfect resource for when you need an overview of the many different types of batteries and their histories. Chemistry Project Report on Finding EMF of Electrochemical Cell Class 12th Chemistry Project on Electrochemical Cells Class 12th Chemistry Project on Electrochemical Cells Chemistry Project Report on Rusting of Iron Chemistry Project Report Soap.docx Chemistry Project on Medicines Chemistry project on nicotine Chemistry Project On Spectroscopy Investigatory Project on Chemistry Chemistry project on nicotine chemistry project on polymer Cell Phone Detector Project Report Project Report on project financing Chemistry Report on Rate of Reaction Chemistry Project on Foaming Capacity of Soaps.pdf Chemistry Report on Rate of Reaction Chemistry Investigatory Project on Consituents of Alloys Chemistry Project on Foaming capacity of soaps Project Report on CocaCola Project Report on Hexamine

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