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## CHE1031 Module 5 summary: Electrochemistry

Openstax chapter 17 (Notice that we're covering only select topics from this chapter.)

### 5.1: Galvanic cells (aka voltaic cells)

**Electrochemical cells** typically consist of two **half-cells**. The half-cells separate the oxidation half-reaction from the reduction half-reaction and make it possible for current to flow through an external wire. One half-cell, normally depicted on the left side in a figure, contains the anode. Oxidation occurs at the **anode**. The anode is connected to the cathode in the other half-cell, often shown on the right side in a figure. Reduction occurs at the **cathode**. Adding a **salt bridge** completes the circuit allowing current to flow. Anions in the salt bridge flow toward the anode and cations in the salt bridge flow toward the cathode. The movement of these ions completes the circuit and keeps each half-cell electrically neutral. Electrochemical cells can be described using **cell notation**. In this notation, information about the reaction at the anode appears on the left and information about the reaction at the cathode on the right. The salt bridge is represented by a double line, ||. The solid, liquid, or aqueous phases within a half-cell are separated by a single line, |. The phase and concentration of the various species is included after the species name. Electrodes that participate in the oxidation-reduction reaction are called **active electrodes**. Electrodes that do not participate in the oxidation-reduction reaction but are there to allow current to flow are inert electrodes. **Inert electrodes** are often made from platinum or gold, which are unchanged by many chemical reactions.

### 5.2: Standard reduction potentials

Assigning the potential of the **standard hydrogen electrode** (SHE) as zero volts allows the determination of standard reduction potentials,  $E^\circ$ , for half-reactions in electrochemical cells. As the name implies, standard reduction potentials use **standard states** (1 bar or 1 atm for gases; 1 M for solutes, often at 298.15 K) and are written as reductions (where electrons appear on the left side of the equation). The reduction reactions are reversible, so standard cell potentials can be calculated by subtracting the standard reduction potential for the reaction at the anode from the standard reduction for the reaction at the cathode. When calculating the standard cell potential, the standard reduction potentials are not scaled by the stoichiometric coefficients in the balanced overall equation.

### 5.3: Batteries and Fuel Cells

**Batteries** are galvanic cells, or a series of cells, that produce an electric current. When cells are combined into batteries, the potential of the battery is an integer multiple of the potential of a single cell. There are two basic types of batteries: primary and secondary. **Primary batteries** are “single use” and cannot be recharged. Dry cells and (most) alkaline batteries are examples of primary batteries. The second type is rechargeable and is called a **secondary battery**. Examples of secondary batteries include nickel-cadmium (NiCd), lead acid, and lithium ion batteries. **Fuel cells** are similar to batteries in that they generate an electrical current, but require continuous addition of fuel and oxidizer. The hydrogen fuel cell uses hydrogen and oxygen from the air to produce water, and is generally more efficient than internal combustion engines.



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#### 5.4: Corrosion

**Corrosion** is the degradation of a metal caused by an electrochemical process. Large sums of money are spent each year repairing the effects of, or preventing, corrosion. Some metals, such as aluminum and copper, produce a protective layer when they corrode in air. The thin layer that forms on the surface of the metal prevents oxygen from coming into contact with more of the metal atoms and thus “protects” the remaining metal from further corrosion. Iron corrodes (forms rust) when exposed to water and oxygen. The rust that forms on iron metal flakes off, exposing fresh metal, which also corrodes. One way to prevent, or slow, corrosion is by coating the metal. **Coating** prevents water and oxygen from contacting the metal. Paint or other coatings will slow corrosion, but they are not effective once scratched. **Zinc-plated or galvanized** iron exploits the fact that zinc is more likely to oxidize than iron. As long as the coating remains, even if scratched, the zinc will oxidize before the iron. Another method for protecting metals is **cathodic protection**. In this method, an easily oxidized and inexpensive metal, often zinc or magnesium (the sacrificial anode), is electrically connected to the metal that must be protected. The more active metal is the sacrificial anode, and is the anode in a galvanic cell. The “protected” metal is the cathode, and remains unoxidized. One advantage of cathodic protection is that the sacrificial anode can be monitored and replaced if needed.