E-RESOURCE CENTER

MAJOR ADVISING: FORENSIC SCIENCE

Sample Assignment: Textbook page (Year 1 & Year 2)

18.5 ELECTROLYTIC CELLS

In an **electrolytic cell,** a nonspontaneous redox reaction is made to occur by pumping electrical energy into the system. A generalized diagram for such a cell is shown in Figure 18.8. The storage battery at the left provides a source of direct electric current. From the terminals of the battery, two wires lead to the electrolytic cell. This consists of two electrodes, A and C, dipping into a solution containing ions M^+ and X^-.

The battery acts as an electron pump, pushing electrons into the *cathode*, C, and removing them from the *anode*, A. To maintain electrical neutrality, some process within the cell must consume electrons at C and liberate them at A. This process is an oxidation-reduction reaction; when carried out in an electrolytic cell, it is called **electrolysis.** At the cathode, an ion or molecule undergoes reduction by accepting electrons. At the anode, electrons are produced by the oxidation of an ion or molecule.

Oxidation occurs at the anode, reduction at the cathode, in both voltaic and electrolytic cells.

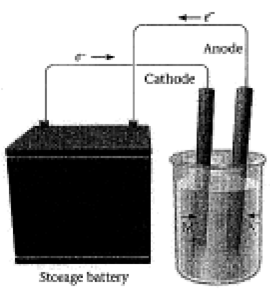

Figure 18.8
Diagram of an electrolytic cell. Electrons enter the cathode from an external source. Cations move to the cathode, where they are reduced, and anions move to the anode, where they are oxidized.

Quantitative Relationships

There is a simple relationship between the amount of electricity passed through an electrolytic cell and the amounts of substances produced by oxidation or reduction at the electrodes. From the balanced half-equations

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$
$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$
$$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$$

you can deduce that

$$1 \text{ mol of } e^- \longrightarrow 1 \text{ mol of Ag (107.9 g of Ag)}$$
$$2 \text{ mol of } e^- \longrightarrow 1 \text{ mol of Cu (63.55 g of Cu)}$$
$$3 \text{ mol of } e^- \longrightarrow 1 \text{ mol of Au (197.0 g of Au)}$$

Relations of this type, obtained from balanced half-equations, can be used in many practical calculations involving electrolytic cells. You will also need to become familiar with certain electrical units, including those of

quantity of electrical charge. The common unit here is the **coulomb, C.** The coulomb is related to the charge carried by a mole of electrons through the Faraday constant:

$$1 \text{ mol electrons} = 9.648 \times 10^4 \text{ C}$$

rate of current flow. Here the common unit is the **ampere, A.** When a current of one ampere passes through an electrical circuit, one coulomb passes a given point in the circuit in one second. That is,

$$1 \text{ A} = 1 \text{ C/s}$$

amount of electrical energy. This can be expressed in joules, J. When a charge of one coulomb (C) moves through a potential difference of one volt (V), it acquires an energy of one joule:

$$1 \text{ J} = 1 \text{ C} \cdot \text{V}$$

Table 18.3 Electrical Units

Quantity	Unit	Defining Relation	Conversion Factors
Charge	**coulomb (C)**	1 C = 1 A · s = 1 J/V	1 mol e^- = 9.648×10^4 C
Current	**ampere (A)**	1 A = 1 C/s	
Potential	**volt (V)**	1 V = 1 J/C	
Power	watt (W)	1 W = 1 J/s	
Energy	joule (J)	1 J = 1 V · C	1 kWh = 3.600×10^6 J

Excerpted from Zumhdal, S. & Zumhdal, S. *Chemistry* (6th ed)

Created in 20011 at the Center for English Language Support, John Jay College, on a U.S. Department of Education (Title V Collaborative) Grant awarded to John Jay College of Criminal Justice and Queensborough Community College.