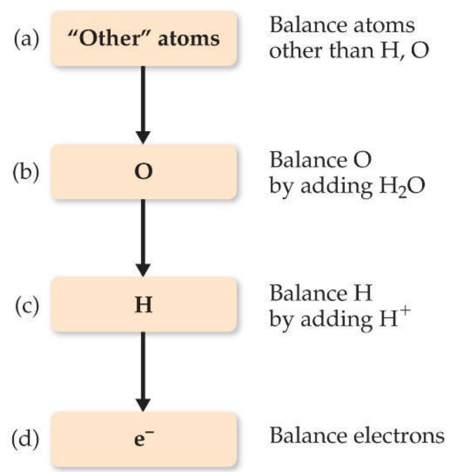
Chapter 20 Electrochemistry

- **Electrochemistry:** study of relationship between electricity and chemical reaction.
- **Oxidation Number:** keep track of which to determine whether oxidation-reduction reaction occurs.
    1. The oxidation number increases for the elements that lose electrons (The element is oxidized).
    2. The oxidation number decreases for the elements that gain electrons (The element is reduced).
- **Half-reaction:** used to balance redox equation

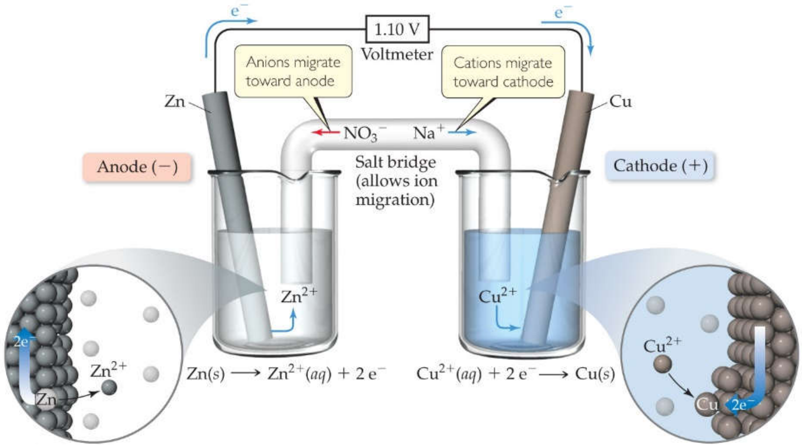
    For example:

$$\textit{Oxidation:} \quad Sn^{2+}(aq) \longrightarrow Sn^{4+}(aq) + 2e^-$$
$$\textit{Reduction:} \quad 2\,Fe^{3+}(aq) + 2e^- \longrightarrow 2\,Fe^{2+}(aq)$$

- **The Half-Reaction Method:**
    1. Divide the equation into one oxidation half-reaction and one reduction half-reaction.
    2. Balance atoms other than O and H. Then, balance O and H using H_2O/H^+
    3. Add electrons to balance charges.
    4. Multiply by common factor to make electrons in half-reactions equal.
    5. Add the half-reactions.
    6. Simplify by dividing by common factor or converting H^+ to OH^- if basic.
    7. Double-check atoms and charges balance.
    8. In basic solution, add OH^- to each side to "neutralize" the H^+ in the equation if exists.



- **Voltaic Cell:** a device in which the transfer of electrons takes place through an external pathway. Electrons leave the anode and flow through the wire to the cathode.
    1. The oxidation occurs at the anode.
    2. The reduction occurs at the cathode.



- **Electromotive Force** (emf, E_{cell} : 1 V =1 J/C): The potential difference between the anode and cathode in a cell.
- **Standard Hydrogen Electrode:** the reduction potential for hydrogen is 0V**.**

$$2\,H^+(aq,\,1M) + 2e^- \longrightarrow H_2(g,\,1\text{ atm})$$

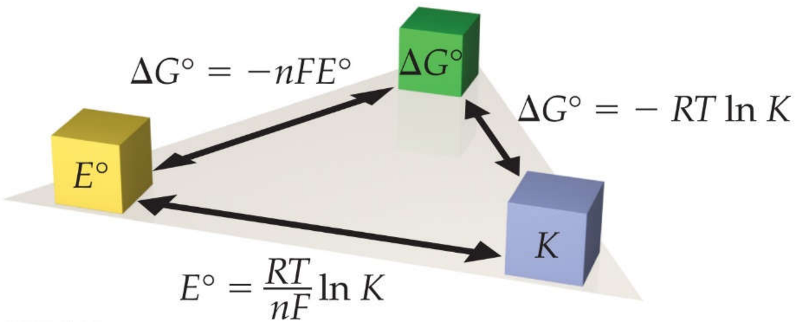
- **Standard Reduction Potential** (E^0_{red}): The tendency relative to standard hydrogen electrode (which is 0V by definition) for a species to be reduced at standard state.
- **Standard Cell Potentials:** standard reduction potential of the cathode reaction minus the standard reduction potential of the cathode reaction.

$$E^\circ_{cell} = E^\circ_{red}\,(\text{cathode}) - E^\circ_{red}\,(\text{anode})$$

- **Oxidizing and Reducing Agents:** The strongest oxidizers have the most positive reduction potentials, and the strongest reducers have the most negative reduction potentials.
- Relationship between emf and free-energy change:

$$\Delta G = -nFE$$

Where F is Faraday's constant, F = 96485 C/ mol = 96485 J/V-mol



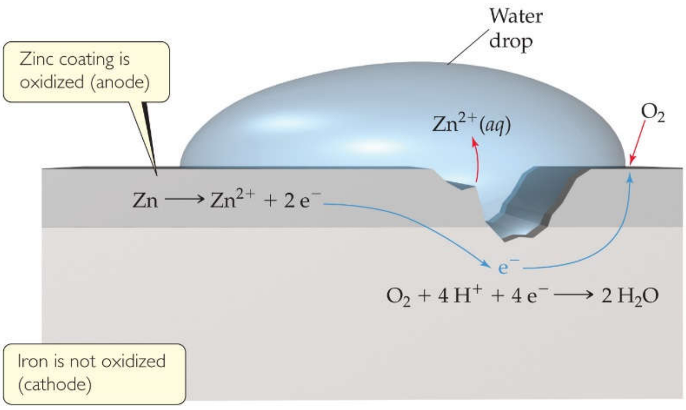
- **Nernst Equation:** to find the emf E produced by a cell under nonstandard conditions.

$$E = E^{\circ} - \frac{2.303\,RT}{nF}\log Q$$

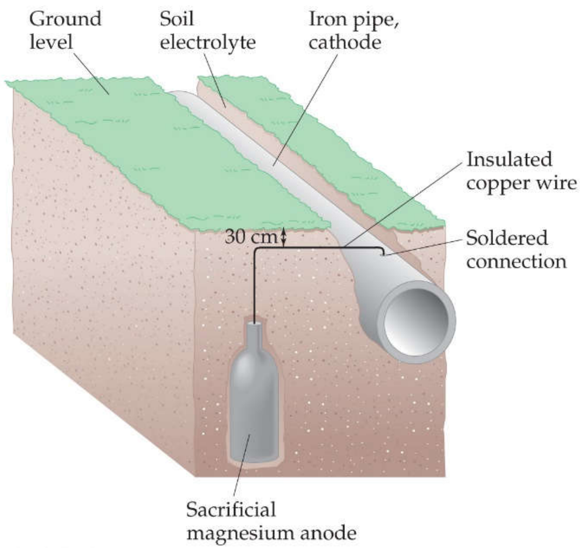
At T = 298K,

$$E = E^{\circ} - \frac{0.0592\text{ V}}{n}\log Q \qquad (T = 298\text{ K})$$

- **Concentration Cell:** same substance with different concentrations at the electrodes.
- **Applications of Cells:**
  1. Batteries: Lead-acid battery, Alkaline battery, Ni-Cd and Ni-metal hydride battery. Lithium-ion batteries.
  2. Fuel cells: convert the energy released by burning fuels to electrical energy more efficiently.
  3. Preventing Corrosion:
     a. coating iron with a mental that is more readily oxidized, like zinc.



     b. Underground pipes: sacrificial anode is oxidized before the pipes.



- **Electrolysis:** use electrical energy to create chemical reactions that are nonspontaneous.

  Charge Q (C) = I*t = nF,

  where I=current (A),

t = time(s),

n = moles of electrons pass through the wire in given time,

F = Faraday's constant