

CHEM 20B Week 9

Aidan Jan

March 4, 2024

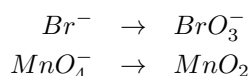
Electrochemistry

General Rules to Balance Oxidation-Reduction Equation:

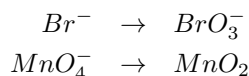
1. Write the two unbalanced half-reactions.
2. Balance both half-reactions for all atoms except O and H.
3. Balance each half-reaction for O by adding H_2O , and then balance for H by adding H^+ .
4. Balance each half-reaction for charge by adding electrons to the side with greater positive charge.
5. Multiply each half-reaction by a factor to make the electron count the same in both half-reactions.
6. Add the two balanced reactions together and cancel species that appear on both sides of the equation.
 - Stop here for acidic solutions
7. For a basic solution, "neutralize" the excess H^+ by adding OH^- and cancel any water (if possible).

Example: Balancing Reaction of MnO_4^- and Br^- in Basic Solution

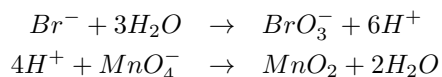
1. Write the two unbalanced half-reactions



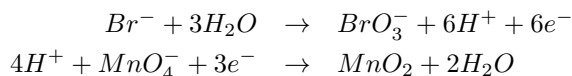
2. Balance both half-reactions for all atoms except O and H.



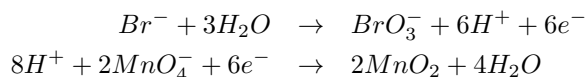
3. Balance each half-reaction for O by adding H_2O , and then balance for H by adding H^+ .



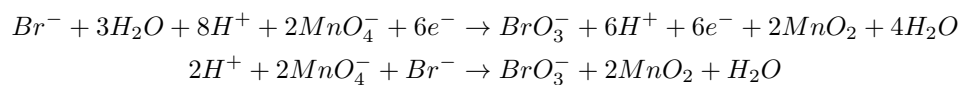
4. Balance each half-reaction for charge by adding electrons to the side with greater positive charge.



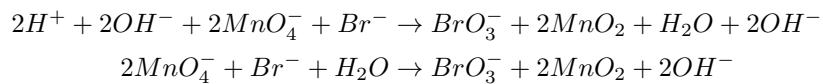
5. Multiply each half-reaction by a factor to make the electron count the same in both half-reactions.



6. Add the two balanced reactions together and cancel species that appear on both sides of the equation.



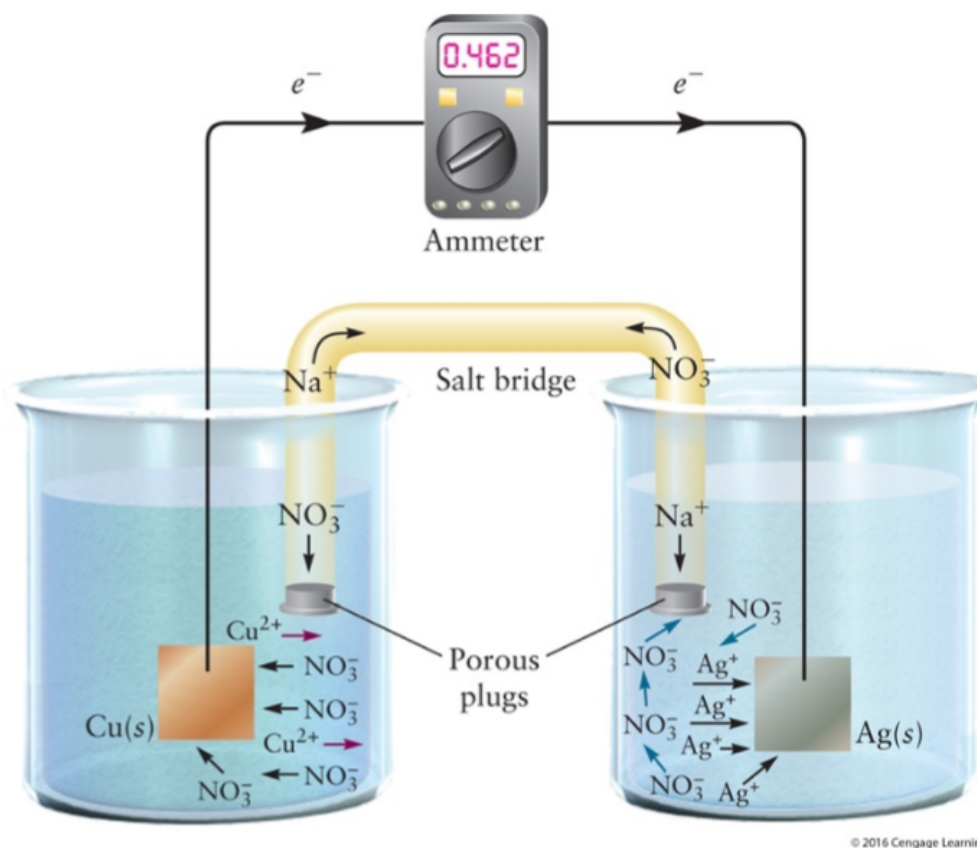
7. Since the reaction occurs in a basic solution, "neutralize" the excess H^+ by adding OH^- and cancel any water (if possible).



Electrochemical cells

An **electrochemical cell** is a device in which an electric current is either produced by a spontaneous chemical reaction or used to bring about a nonspontaneous reaction.

- **Galvanic cell:** electrochemical cell in which a spontaneous chemical reaction is used to generate an electric current.
- For example, batteries
 - If we simply mix two species that undergo a redox reaction, energy is released as heat, but no electricity is generated. However, if we separate the reactants and provide a pathway for the electrons to travel, the electrons can do work, generating an electric current.
- A galvanic cell consists of **two electrodes**:
 - **Anode:** the electrode at which oxidation (Loss of electrons) occurs.
 - **Cathode:** the electrode at which reduction (gain of electrons) occurs.
 - Electrons flow from the anode to cathode.



Voltage

- Electrostatic Potential: Φ (V)
- Electrostatic Potential Energy: E_p (J)
- The SI unit for potential in the volt:

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

$$\Delta E_p = q\Delta\Phi$$

Faraday's Laws

- The mass of a given substance that is produced or consumed in an electrochemical reaction is proportional to the quantity of electric charge passed.
- The magnitude of the charge on a single electron:

$$|-e| = e = 1.60 \times 10^{-19} C$$

- The magnitude of the charge of one mole of electrons:

$$Q = (6.022 \times 10^{23})(1.60 \times 10^{-19} C) = 96485.34 C$$

$$Q = e \cdot N$$

- Faraday constant: Denoted as F , the magnitude of the charge per mole of electrons:

$$F = 96,485.34 C \cdot mol^{-1}$$

- Electric Current: is the amount of charge that flows through a circuit per second, measured in the SI unit, the ampere (A):

$$1 C = (1 A)(1 s)$$

- *The amount of charge that has passed when a current of 1 amperes has flowed for t seconds is*

$$Q = I \cdot t$$

- The number of moles of electrons n transferred:

$$n = \frac{I \cdot t}{96485.34}$$