

Totally Epic AP Chem ³¹Review: Cell Potential!

Electrochemistry: a study of the interchange of electrical and chemical energy

❖ There are two types of electrochemical cells: galvanic (voltaic) and electrolytic.

- **Voltaic or Galvanic Cell (Battery):** a thermodynamically favorable (i.e. spontaneous) redox reaction which generates useful electrical energy in the form of an electric current
- **Electrolytic Cell:** requires electrical energy (direct current or DC power source) to drive a thermodynamically unfavorable (i.e. non-spontaneous) redox reaction.
- In short: galvanic (voltaic) cells produce current, while electrolytic cells use current!

Cell Potential (E_{cell}): a measure of the potential difference (how much voltage exists) between two half cells an electrochemical cell. The potential difference is caused by the ability of electrons to flow from one half-cell to the other.

→ The cell potential is a quantitative description of the driving force behind an electrochemical reaction that pushes electrons through the wire (or external circuit).

Standard Cell Potential (E_{cell}°): cell potential measured at standard conditions: **1 atm, 1 M solution, and 25°C**. A 'naught' sign ($^{\circ}$) is used to show standard conditions. Usually measured in Volts (1 V = 1 Joule/Coulomb = 1 J/C)

The cell potential can be easily calculated by adding the oxidation and reduction potentials.

$$E_{cell}^{\circ} = E_{ox}^{\circ} + E_{red}^{\circ}$$

- E_{cell}° is the standard cell potential.
 - E_{ox}° is the standard oxidation cell potential for the oxidation half-reaction.
 - E_{red}° is the standard reduction cell potential for the reduction half-reaction.
- The more positive the value of E_{cell}° , the greater the driving force of electrons through the system (under standard conditions), thus the more likely the reaction will proceed → more spontaneous/more thermodynamically favorable.

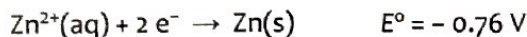
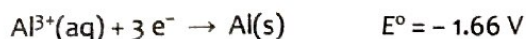
When you balance a redox reaction, don't change the cell potential!

Cell potential is an intensive property and thus does not depend on how many times a reaction occurs.

Because the values come from a chart of standard reduction potentials, you **MUST REVERSE** the sign of the E° of the **oxidized** species before adding to the E° of the reduced species.

1. For a spontaneous redox reaction to occur, the overall cell potential must be positive.
 - a) The metal with the **greater** (more positive) reduction potential will be **reduced!**
2. For a non-spontaneous redox reaction to occur, the overall cell potential must be negative.
 - a) The metal with the **greater** (more positive) reduction potential will be **oxidized!**

Example 1: Consider the half reactions shown below and the standard electrode reduction potentials that follow.

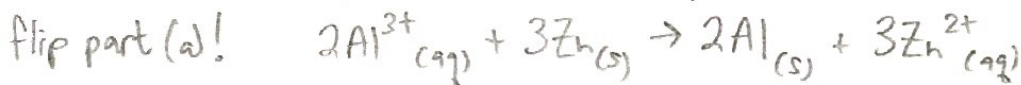


- a. Write the balanced redox reaction for copper and aluminum that is thermodynamically favorable (i.e. spontaneous). Calculate the standard cell potential of this reaction.



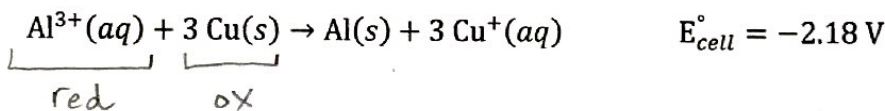
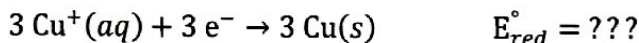
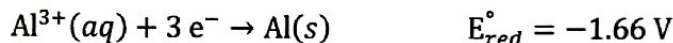
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = 1.66 - 0.76 = \boxed{0.90 \text{ V}}$$

- b. Write the balanced redox reaction for copper and aluminum that is not thermodynamically favorable (i.e. non-spontaneous). Calculate the standard cell potential of this reaction.



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = 0.76 - 1.66 = \boxed{-0.90 \text{ V}}$$

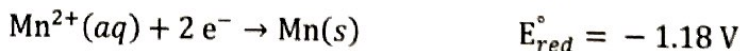
Example 2: use the provided information to find the reduction potential of the missing half-reaction



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = E^{\circ}_{\text{ox}} - 1.66 = -2.18$$

$$E^{\circ}_{\text{ox}} = 1.66 - 2.18 = -0.52 \text{ V}$$

$$\Rightarrow E^{\circ}_{\text{red}}(\text{Cu}^{+}) = \boxed{+0.52 \text{ V}}$$



Example 3: Based on the reduction potentials given above, which of the following reactions will be thermodynamically favored?

