Electrochemistry Name________________________________

Cell Voltage and Batteries

Read from **Lesson 2: Galvanic Cells** in the **Chemistry Tutorial** Section**, Chapter 18** of **The Physics Classroom:**
Part c: [Cell Voltage] **Part d:** [Batteries and Commercial Cells]

1. Non-Standard Conditions & Nernst Equation

- Real cells, including the ones that you build in chem lab, almost never operate under standard conditions. (1.00 M concentrations, 1 atm pressure, 25°C)
- Concentrations shift as the reaction proceeds, and those shifts change the cell's voltage.
- The **Nernst equation** lets us calculate the *actual* cell potential (**$E°_{cell}$**), based on the concentrations of reactants and products.

$$E = E° - \frac{R \cdot T}{n \cdot F} \cdot \ln Q$$

- **R** = 8.314 J/mol/K
- **T** = Kelvin temperature
- **n** = number of moles of *e-* transferred in the balanced chemical equation
- **F** is 96485 C/mol (*known as Faraday's constant*)
- **Q** is the reaction quotient.
- At a standard temperature of **298 K**, **R•T/F** simplifies to **0.0592** and the Nernst equation becomes:

$$E = E° - 0.0592 / n \cdot \log Q$$

2. How Concentration Affects Cell Voltage

- More reactants → higher voltage
  - If reactant concentrations increase, ***Q*** becomes smaller.
  - Since ***log Q*** becomes more negative, the subtraction becomes "minus a negative," which raises **E_{cell}**.
  - More reactants = stronger push toward products = stronger electrical output.
- More products → lower voltage
  - If product concentrations increase, ***Q*** becomes larger.
  - ***log Q*** becomes positive, so the Nernst term subtracts a positive number, lowering **E_{cell}**.
  - This is why batteries "run down": as products accumulate, voltage drops.

3. Gibbs Free Energy and Cell Voltage

- **The Equation: $\Delta G = - n \cdot F \cdot E_{cell}$**
  - **ΔG** = Gibbs free energy change (J or kJ)
  - **F** = Faraday's constant (96485 C/mol)
  - **n** = moles of electrons transferred
  - **E_{cell}** = cell voltage (V)
- If **E_{cell}** is positive…
  - ΔG is negative
  - Reaction is spontaneous
  - The cell can do useful work
  - Electrons flow from anode → cathode naturally
- If **E_{cell}** is negative…
  - ΔG is positive
  - Reaction is non-spontaneous
  - The cell will not do work without external energy
  - This is how electrolytic cells behave

4. Batteries and Commercial Cells

- All batteries are packaged galvanic cells that convert spontaneous redox reactions into usable electrical energy.
  - Alkaline dry cells: common household cells
  - Lead-acid batteries: automotive batteries with reversible reaction; sulfuric acid electrolyte.
  - Li-ion batteries: high energy and rechargeability for portable electronics.

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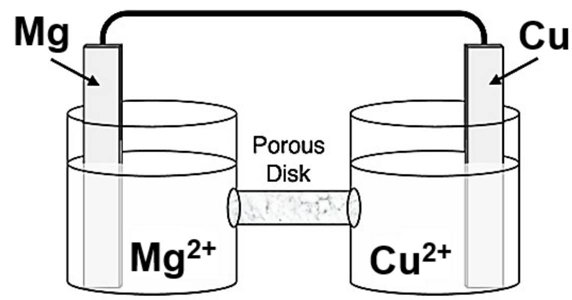
Questions

For all questions, use the Standard Reduction Potential Values found on the Chemistry Tutorial Reference Section: **Standard Reduction Potential Values** – E values for several half-cells.

1. Observe the Galvanic cell diagram below. It shows the electrodes and the solutions in each half-cell at standard conditions. Using this diagram and the Table of Standard Reduction Potentials, answer the questions that follow.



   a. Which half-reaction occurs at the anode?

   b. Which half-reaction occurs at the cathode?

   c. What is the overall reaction for this cell?

   d. What is the cell voltage for this Galvanic cell?

   e. Another Galvanic cell is made by replacing the 1.0 M solution of Cu^{2+} (aq) with a 0.50 M solution of Cu^{2+} (aq). (The Mg^{2+}(aq) remains 1.0 M) Would the cell potential of the new cell be greater than, less than, or equal to the cell potential of the original Galvanic cell? Explain your answer and show calculations to support your answer.

   f. Calculate the value of ΔG for the reaction. Does this value support the spontaneous reaction of the cell?

   g. Which Galvanic cell, the original with 1.0 M solutions, or the altered cell, reach equilibrium first? How do you know?

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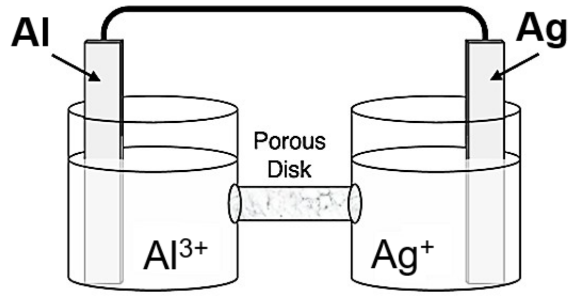
2. Observe the Galvanic cell diagram below. It shows the electrodes and the solutions in each half-cell at standard conditions. Using this diagram and the Table of Standard Reduction Potentials, answer the questions that follow.



a. Which half-reaction occurs at the anode?

b. Which half-reaction occurs at the cathode?

c. What is the overall reaction for this cell?

d. What is the cell voltage for this Galvanic cell? Explain why you don't multiply reduction potentials by stoichiometric coefficients even when balancing electrons.

e. Another Galvanic cell is made by replacing the 1.0 M solutions with 2.0 M solutions of Al^{3+} (aq) and Ag^{+} (aq). Would the cell potential of the new cell be greater than, less than, or equal to the cell potential of the original Galvanic cell? Explain your answer and show calculations to support your answer.

f. Calculate the value of ΔG for the reaction. Does this value support the spontaneous reaction of the cell?

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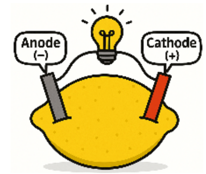
3. A lemon battery is a simple Galvanic cell that creates electricity using a lemon as an electrolyte, a zinc nail (anode), and a copper coin or wire (cathode) to power a light bulb or other device. How will the voltage of a lemon battery change over time as the chemical reaction continues? Explain your reasoning in terms of ion concentration and how the cell operates to support your answer.



4. **Alkaline dry cells** are among the most common household batteries. Describe the half-reactions that occur in an alkaline cell and explain how a voltage of ~1.50 V is generated.

5. Lead-acid batteries are rechargeable. Describe what happens chemically when the battery is recharged.