Electrochemistry **Name**______________________________

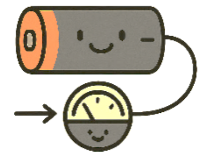
Reduction Potential and Cell Voltage

Read from **Lesson 2: Galvanic Cells** in the **Chemistry Tutorial** Section**, Chapter 18** of **The Physics Classroom:**
Part b: Reduction Potentials **Part c:** Cell Voltage

1. Reduction Potentials (E°)

Reduction potentials measure how strongly a substance *attracts electrons*. Comparing these values allows us to predict:

- Which species is reduced (cathode)
- Which species is oxidized (anode)
- Direction of electron flow
- Whether a reaction is spontaneous



2. Electric Potential

- A galvanic cell produces an **electric potential (voltage)** — a measure of its ability to produce current.
- The greater the difference in electron-acquisition tendencies between two half-cells, the **larger the voltage**.

3. Standard Reduction Potential ($E°_{red}$)

- Measured relative to the **Standard Hydrogen Electrode (SHE)**, defined as **0.00 V** at these conditions: 1.00 M solutes, 1.00 atm gases, 298 K
- **How measurement works**
  - Test half-cell connected to SHE.
  - A **positive voltmeter reading** means the test electrode is the **anode** (oxidation occurs), so its reduction potential is **negative**.
  - A **negative reading** means the test electrode is the **cathode**, so its reduction potential is **positive**.

4. Interpreting E° Values

- **More positive $E°_{red}$ → stronger oxidizing agent → more likely to be reduced**
- **More negative $E°_{red}$ → stronger reducing agent (when reversed) → more likely to be oxidized**
- **Example**
  - $Zn^{2+} + 2e^- \rightarrow Zn(s)$: **E° = –0.76 V**
  - $Ag^+ + e^- \rightarrow Ag(s)$: **E° = +0.80 V**
  - Ag^+ is a stronger oxidizing agent than H^+; Zn is a strong reducing agent.

5. Predicting Spontaneity

- A reaction is spontaneous if: the **oxidizing agent (left side) is lower on the table than the reducing agent (right side).**
- Example: $Cu^{2+} + Zn(s) \rightarrow Cu(s) + Zn^{2+}$ is spontaneous, but the reverse is not.

6. Cell Voltage ($E°_{cell}$)

- Cell voltage arises from the difference in reduction potentials between two half-cells.

7. Identifying Half-Reactions

- The **more positive $E°_{red}$** becomes the **reduction** half-reaction (cathode).
- The other half-reaction is reversed and becomes **oxidation** (anode).

8. Calculating Standard Cell Voltage

- **Equation: $E°_{cell} = E°_{red} + E°_{ox}$**
- Example (Zn/Cu cell): $Cu^{2+} + 2e^- \rightarrow Cu(s)$: **+ 0.34 V** $Zn(s) \rightarrow Zn^{2+} + 2e^-$: **+ 0.76 V** (oxidation potential)
- **$E°_{cell} = E°_{red} + E°_{ox}$** = = 0.34 + 0.76 = +1.10 V
- A **positive $E°_{cell}$** always indicates a **spontaneous** galvanic cell.

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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. Explain why a half-reaction with a highly positive $E°_{red}$ is a strong oxidizing agent.

2. What is the **Standard Hydrogen Electrode (SHE)**, and why do chemists use it to measure reduction potentials?

3. Rank the following in order of increasing strength as oxidizing agents. Explain how you determined this ranking**.** $\mathbf{Sn^{2+}}$**,** $\mathbf{Zn^{2+}}$**,** $\mathbf{Pb^{2+}}$**,** $\mathbf{Fe^{2+}}$

4. Rank the following in order of increasing strength as reducing agents. Explain how you determined this ranking**. Sn, Zn, Pb, Fe**

5. Which of the following metals (**Sn, Zn, or Pb**) would most effectively reduce $\mathbf{Fe^{3+}}$ to $\mathbf{Fe^{2+}}$? Explain your reasoning**.**

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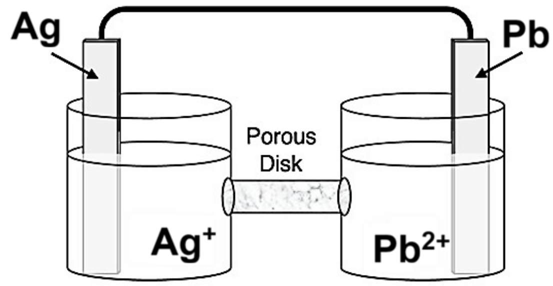
6. Observe the Galvanic cell diagram below. It shows the electrodes and the solutions in each half-cell. Using this diagram and the Table of Standard Reduction Potentials, answer the questions that follow.

   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? Assume standard conditions.

For the following requested Galvanic cell combinations, use the reduction half-reactions below,

$\mathbf{Ag^+ (aq) + e^- \rightarrow Ag (s)}$	$\mathbf{Fe^{2+} (aq) + 2e^- \rightarrow Fe(s)}$
$\mathbf{Cu^{2+} (aq) + 2e^- \rightarrow Cu (s)}$	$\mathbf{Cr^{3+} (aq) + 3e^- \rightarrow Cr (s)}$

And answer the following questions:

a. Which combination will generate the highest positive cell voltage?
b. Which half-reaction occurs at the anode?
c. Which half-reaction occurs at the cathode?
d. Write the balanced overall redox reaction for the cell.
e. Determine the standard cell potential for the Galvanic cell. Assume standard conditions.

7. a. Which combination will generate the highest positive cell voltage? Explain your reasoning.

   b. Which half-reaction occurs at the anode?

   c. Which half-reaction occurs at the cathode?

   d. Write the balanced overall redox reaction for the cell.

   e. Determine the standard cell potential for the Galvanic cell. Assume standard conditions.

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8. a. Which combination will generate the least positive cell voltage? Explain your reasoning.

   b. Which half-reaction occurs at the anode?

   c. Which half-reaction occurs at the cathode?

   d. Write the balanced overall redox reaction for the cell.

   e. Determine the standard cell potential for the Galvanic cell. Assume standard conditions.

9. Determine whether each reaction is **spontaneous** under standard conditions. Write **YES** (spontaneous) or **NO** (not spontaneous). Explain your reasoning.
   a. $2\ Fe^{3+}\ (aq) + Sn\ (s) \rightarrow 2\ Fe^{2+}\ (aq) + Sn^{2+}\ (aq)$

   b. $2\ Cr\ (s) + 3\ Pb^{2+}\ (aq) \rightarrow 2\ Cr^{3+}\ (aq) + 3\ Pb\ (s)$

   c. $Cu\ (s) + Br_2\ (l) \rightarrow Cu^{2+}\ (s) + 2\ Br^{-}\ (aq)$

   d. $Pb^{2+}\ (aq) + 2\ Fe^{2+}(aq) \rightarrow Pb\ (s) + 2\ Fe^{3+}\ (aq)$