Standard Electrode Potential

Outcomes:

- Define Standard electrode potential
- Calculate cell potentials given electrode potentials
- Predict the spontaneity of reactions using standard electrode potentials

Standard Reference Electrode:

Electrochemical cells are usually described in terms of their **VOLTAGE** →Can be predicted using your **REDUCTION** **POTENTIAL** table.

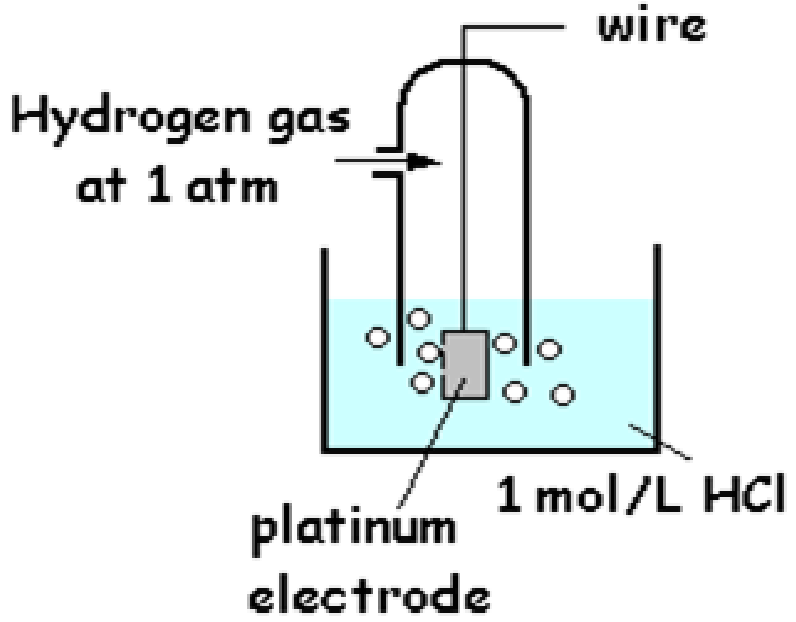
The electrode potentials (E°) were obtained by comparing half-reactions to a **STANDARD** (or **REFERENCE**).

Standard Reference Electrode:

Standard Reference Electrode:

- Chemists chose the **HYDROGEN** half-cell as the **STANDARD**.
- They assigned it a **REDUCTION POTENTIAL** of **E° = 0.00V.**
- They hooked the hydrogen half cell up to other half cells to find their potentials.
  - **POSITIVE** voltage = **GREATER** tendency to **ATTRACT e^-**

  - **NEGATIVE** voltage = **LESSER** tendency to **ATTRACT e^-**



Standard Electrode Potential:

Standard Electrode Potential:

- The **VOLTAGE** (**POTENTIAL**) produced by an electrode when the half-cell ion **CONCENTRATION** is **1.0M** at 1atm and 25°C.
- See table of reduction potentials.

Cell Potentials:

- The cell **POTENTIAL** (**VOLTAGE**) can be predicted from the table.
- It is the **SUM** of the potentials of the **OXIDATION HALF-CELL** ***(E°_{ox})*** and the **REDUCTION HALF-CELL** ***(E°_{red})***.

$$E^\circ_{cell} = E^\circ_{ox} + E^\circ_{red}$$

- Note that your table lists only potentials for **REDUCTION** reactions.
- Since the **OXIDATION** reaction is simply the **REVERSE** of reduction, the **OXIDATION** **POTENTIAL** is the **INVERSE** (**SWITCH SIGN**) of the **REDUCTION** potential.
- If the potential is:
  - **POSITIVE** (**E°CELL > 0**) → **SPONTANEOUS** Rxn
  - **NEGATIVE** (**E°CELL < 0**) → **NOT** Spontaneous

Standard Electrode Potential:

Net Reactions:

- Are the **SUM** of the **OXIDATION** and **REDUCTION** reactions.
- Note that the **ELECTRONS LOST** must **EQUAL ELECTRONS GAINED** before adding!
- The half-cell potential is **NOT MULTIPLIED** when we balance electrons lost & gained.

Examples:

What is the cell potential for a silver-copper cell?

sign flipped

LEO = $Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$ $\qquad E^\circ_{ox} = -0.34\ V$

GER = $2Ag^+ + 2e^- \rightarrow 2Ag_{(s)}$ $\qquad E^\circ_{red} = +0.80\ V$

$E^\circ_{cell} = +0.46\ V$

Standard Electrode Potential:

Examples:

A cell is made using zinc and ~~gold~~ Copper(II) metal as electrodes.

a) What is the cathode and what is the anode?

b) What is the net reaction?

c) What is the cell potential?

a) Cathode (GER) = Copper

Anode (LEO) = Zinc

b) LEO: $Zn_{(s)} \rightarrow Zn^{2+} + 2e^-$ $\quad E^\circ_{ox} = +0.76V$

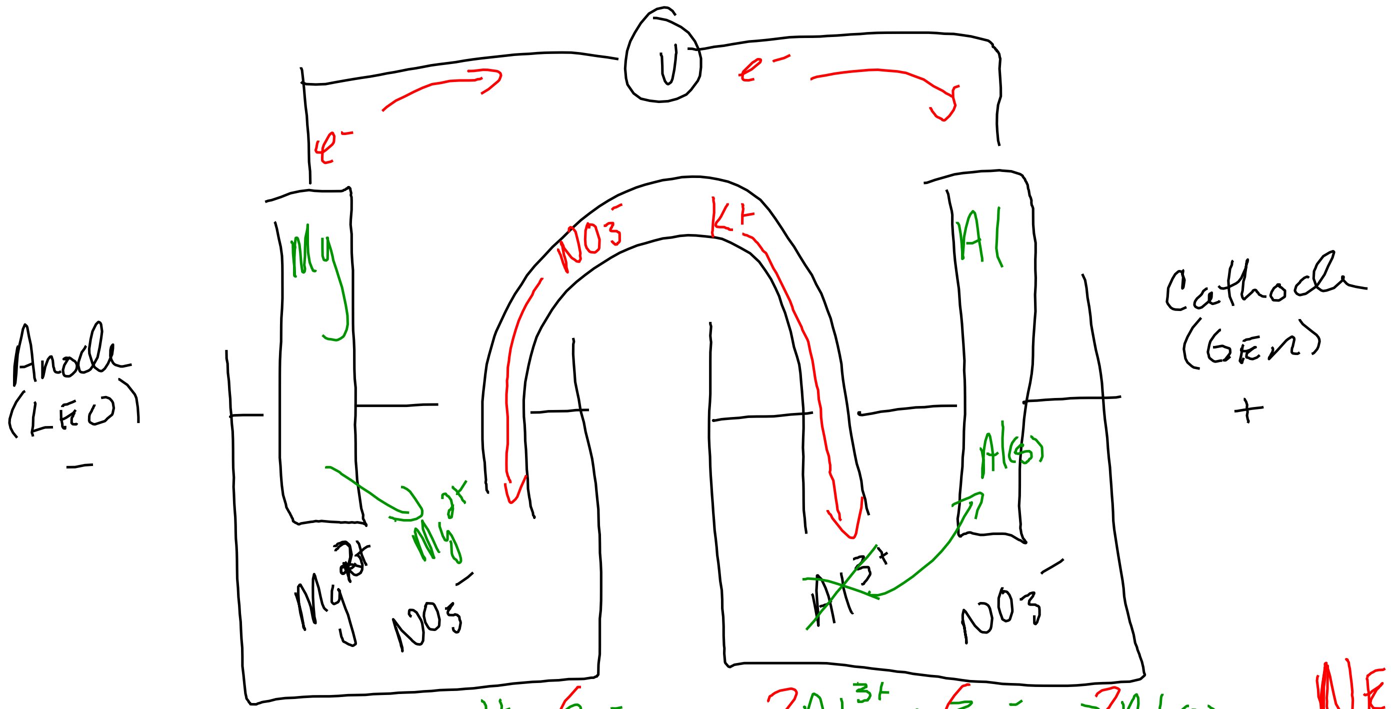
GER: $Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$ $\quad E^\circ_{red} = +0.34V$

c) $E^\circ_{cell} = 1.1\ V$

Putting it all together:

Example:

Draw a magnesium-aluminum cell. Show the movement of electrons, ions, and contents of all solutions. Label the electrodes and write the reaction at each. Determine the net reaction and overall cell potential.



$3\,Mg_{(s)} \rightarrow 3Mg^{2+} + 6e^-$

$E^\circ_{ox} = +2.37\,V$

$2Al^{3+} + 6e^- \rightarrow 2Al_{(s)}$

$E^\circ_{red} = -1.66\,V$

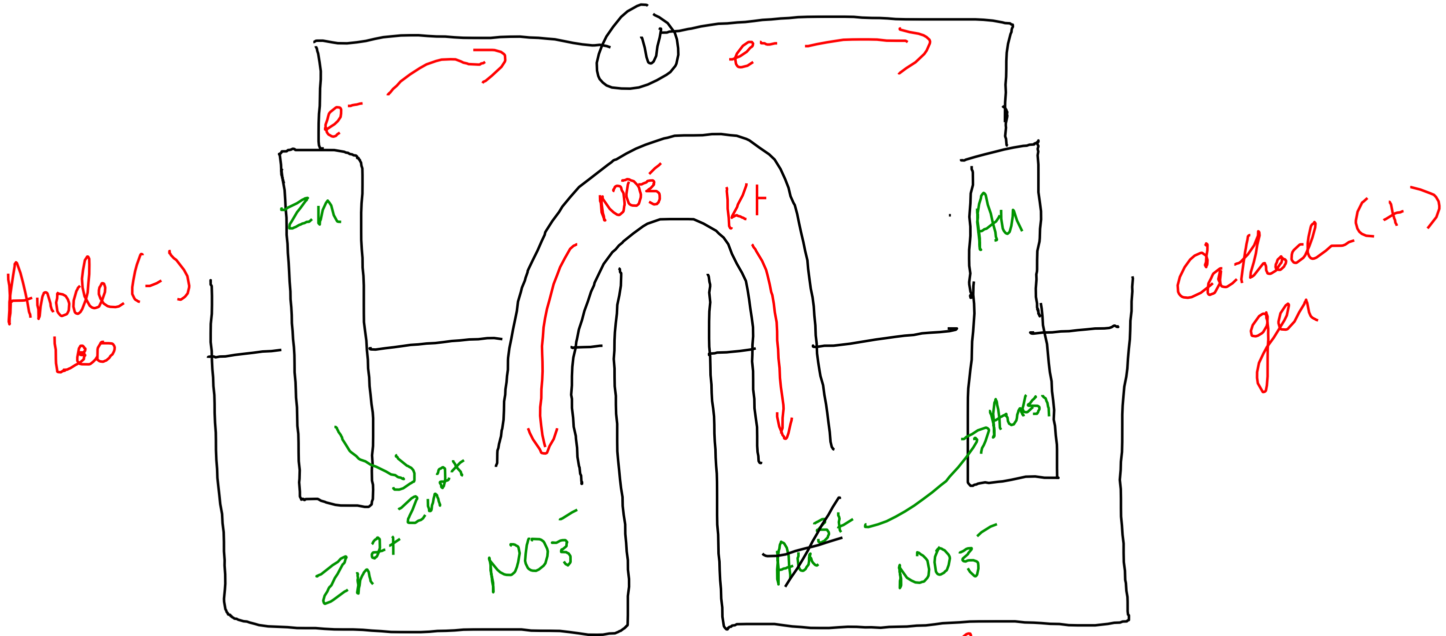
NET: $3Mg_{(s)} + 2Al^{3+} \rightarrow 3Mg^{2+} + 2Al_{(s)}$

$E^\circ_{cell} = +0.71\,V$

Putting it all together:

Try this one…

Draw a zinc-gold cell labelling everything. Write the net reaction and determine the cell potential.



$3Zn_{(s)} \rightarrow 3Zn^{2+} + \cancel{2}6e^-$

$E^\circ_{ox} = +0.76\ V$

$2Au^{3+} + \cancel{3}6e^- \rightarrow 2Au_{(s)}$

$E^\circ_{red} = 1.5\ V$

Net: $3Zn_{(s)} + 2Au^{3+} \rightarrow 3Zn^{2+} + 2Au_{(s)}$

$E^\circ_{cell} = +2.26\ V$