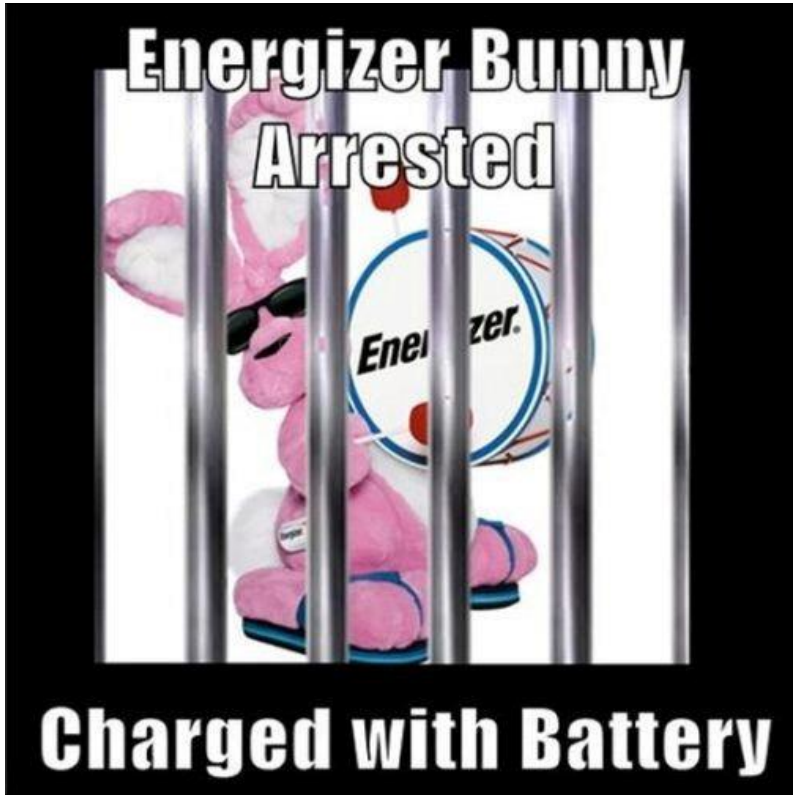
Electrochemical Cells Intro



Outcomes:

- Outline the historical development of voltaic (galvanic) cells.
- Explain the operation of a voltaic cell at the visual, particulate and symbolic levels.

Vocabulary:

Alessandro Volta

- Made the first **BATTERY** using a pile of **METAL** **DISKS** separated by **LEATHER** saturated in **ACID**.

Electrochemical Cell

- A device that converts **CHEMICAL** energy into **ELECTRICAL** energy.

Electrical Current

- The **FLOW** of **ELECTRONS**.

Electrodes

- Conducting **RODS** where electrons flow in and out of.

Voltage (V)

- A unit of ***cell*** or **ELECTRICAL** **POTENTIAL** ***(E˚)***
- The ability of a cell to do **WORK**. Is the **FORCE** or **PRESSURE** that flowing electrons have.

Half-Cell

- The **CONTAINER** in which either the oxidation or reduction reaction occurs.

Electrochemical Cells:

Recall that redox reactions involve a transfer of electrons…

Example:

When a copper strip is placed in an $AgNO_3$ solution, the following reaction takes place:

$$Cu_{(s)} + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag_{(s)}$$

0 +1 +2 0

The copper is "**EATEN AWAY**" and a layer of silver is **FORMED**.

In this reaction, Cu is **OXIDIZED** and Ag is **REDUCED**:

Oxidation: $Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$

Reduction: $(Ag^+ + 1e^- \rightarrow Ag_{(s)})\ x2$

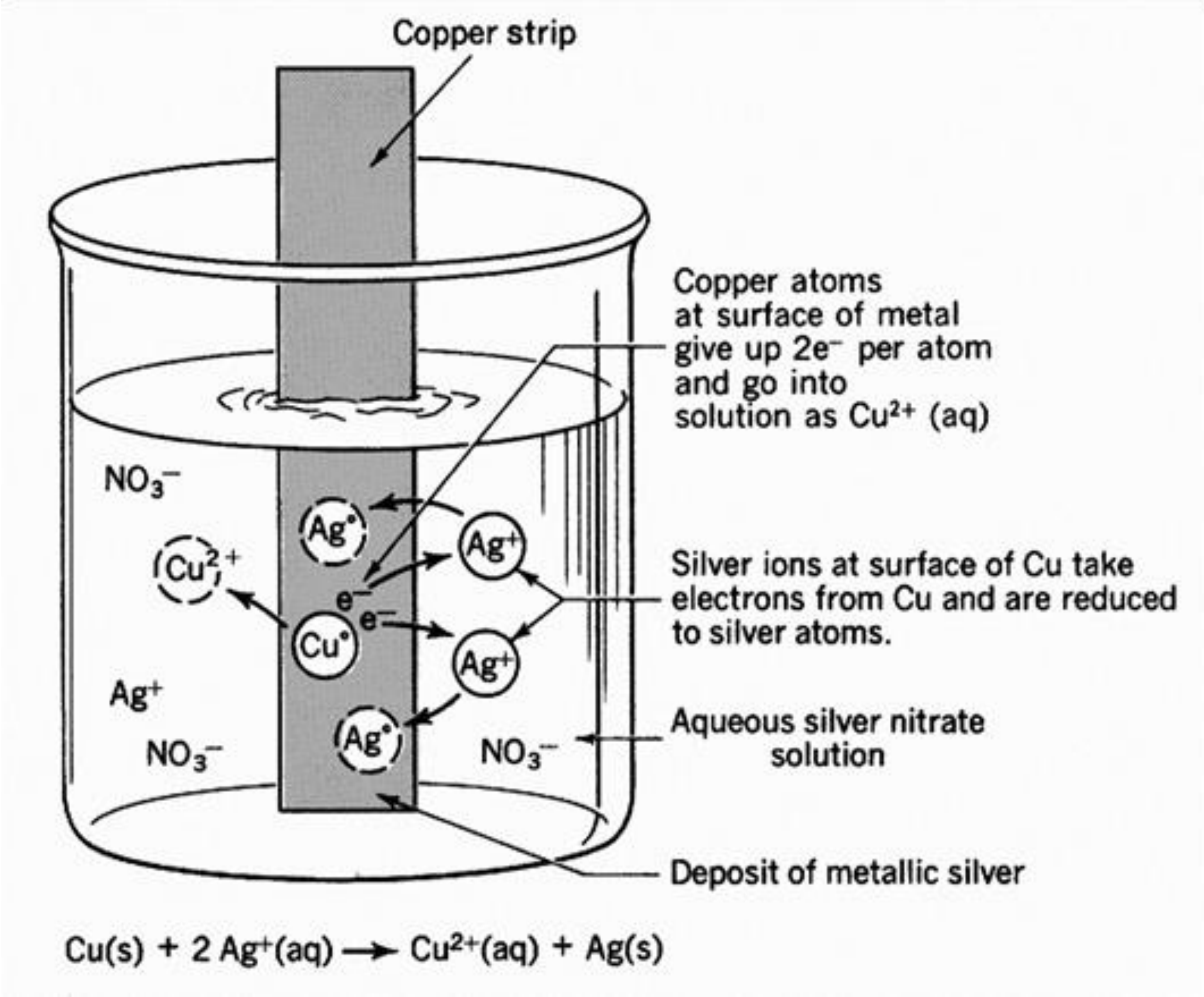
Net Ionic: $Cu_{(s)} + 2Ag^+ \rightarrow Cu^{2+} + 2Ag_{(s)}$

Here, the electron transfer is direct.

<u>Electrochemical Cells:</u>

It is possible to **<u>SEPARATE</u>** these half reactions so we can create an **<u>ELECTRICAL</u>** **<u>CURRENT</u>**.

→ Called a **<u>VOLTAIC</u>** ***(<u>ELECTROCHEMICAL</u>)*** **<u>CELL</u>*.***



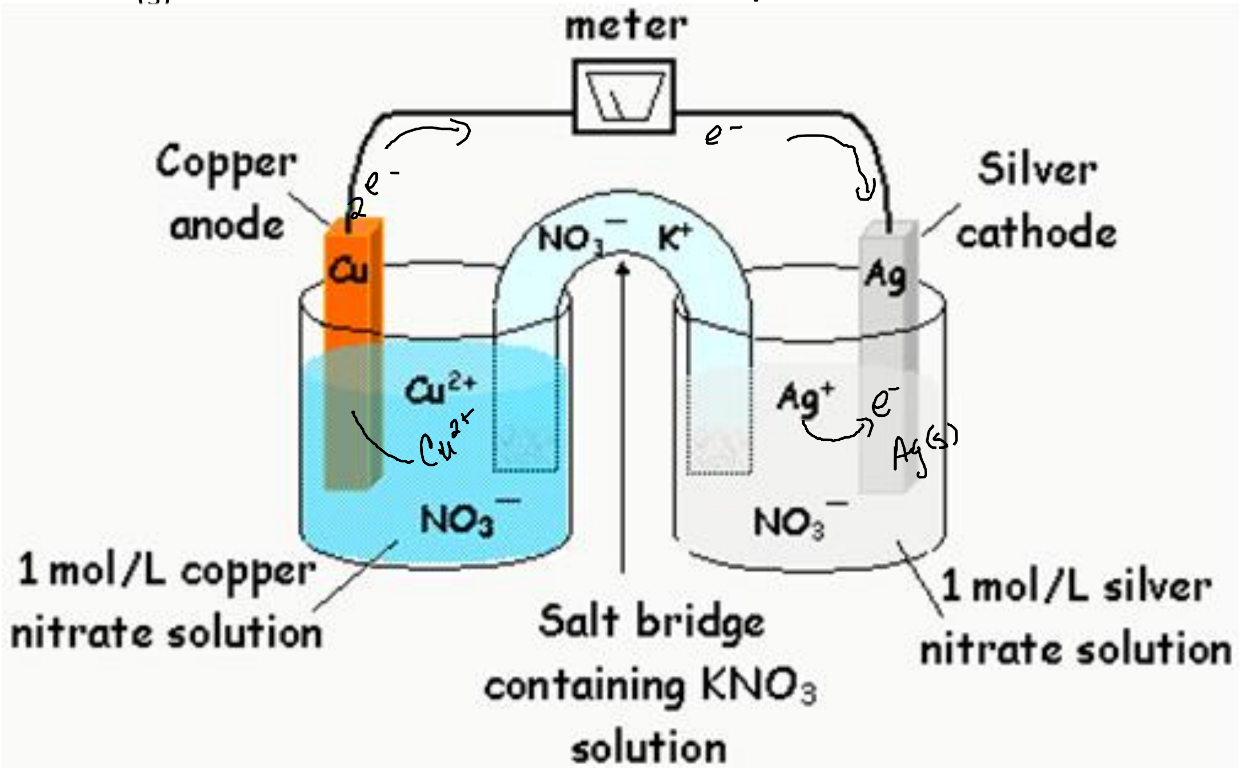
Electrochemical Cells:

If we take the copper & silver redox reaction from before, but **SEPARATE** the half reactions into two **HALF**-**CELLS**, we will get a **VOLTAIC** (electrochemical) cell

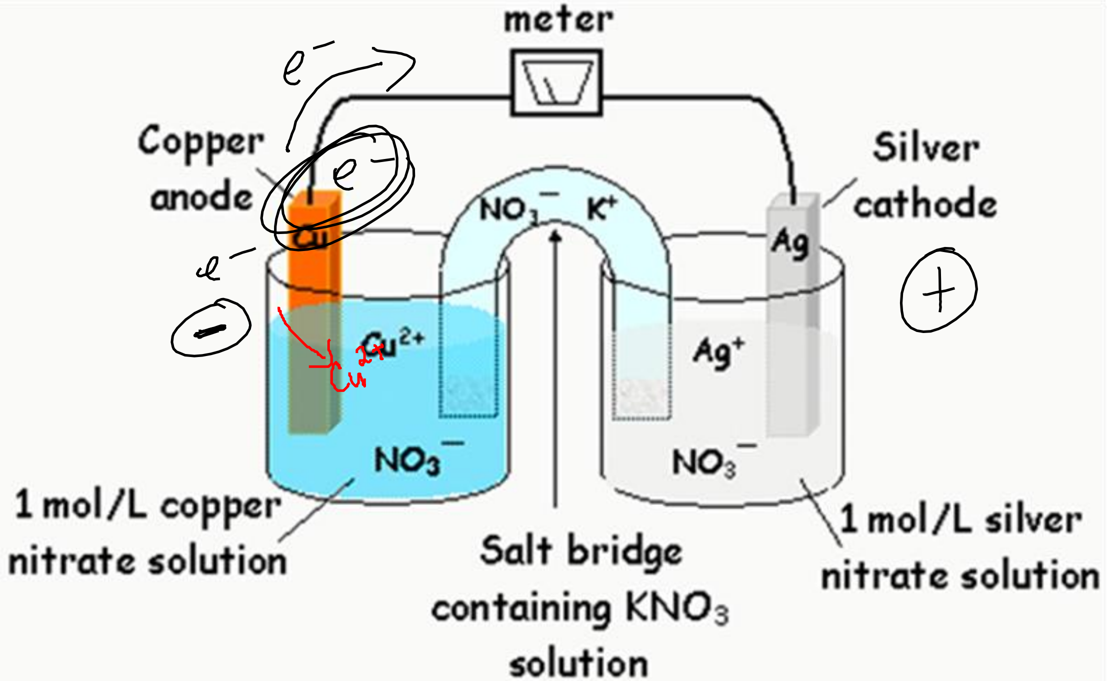
The spontaneous redox reaction is:

$$Cu_{(s)} + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag_{(s)}$$

Here, Ag^{+} is **REDUCED** and $Cu_{(s)}$ is **OXIDIZED**. We can set up the cell as follows:



Electrochemical Cells:



At the **ANODE:** (LEO)

- The **OXIDATION** half-reaction occurs.

$$Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$$

- Electrons are **PRODUCED**, so it is the **NEGATIVE** electrode.

At the **CATHODE:**

- The **REDUCTION** half-reaction occurs.

$$Ag^+ + 1e^- \rightarrow Ag_{(s)}$$

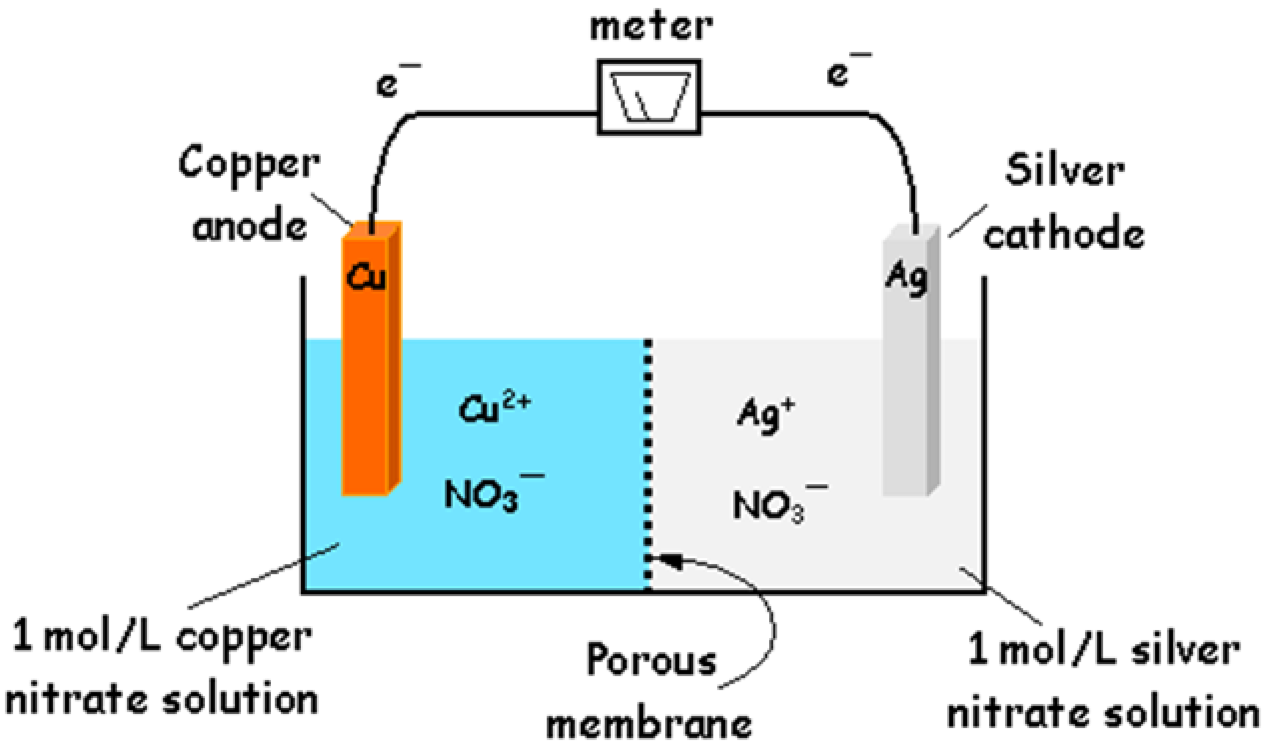
- Electrons are **CONSUMED**, so this is the **POSITIVE** electrode.

Electrochemical Cell Part 1

Electrochemical Cells:

The **Salt Bridge:**

- Joins the two **HALF**-**CELLS**, allowing movement of **IONS** between the **SOLUTIONS**.
- Can also be made of a **POROUS** **MEMBRANE**:



How a Cell Works:

Zn(s) / Zn(NO3)2 // Cu(NO3)2 / Cu(s)

In ALL Electrochemical Cells:

(−)

1. **OXIDATION** occurs at the **ANODE** (both start with vowels).

2. **REDUCTION** occurs at the **CATHODE**.

3. ***Electrons travel from the ANODE to the CATHODE***. In our cell, the $Cu_{(s)}$ **PRODUCES** electrons which travel to the **CATHODE**.

4. The **CATHODE GAINS MASS**. In our example, the Ag^{+} accepts electrons, and produces more $Ag_{(s)}$.

5. The **ANODE LOSES MASS**. In our example, when the **COPPER** loses electrons it becomes Cu^{2+} ions that **DISSOLVE** in solution.

6. We use shorthand notation to describe this cell:

Anode Cathode

$$Cu_{(s)} / Cu(NO_3)_2 // AgNO_3 / Ag_{(s)}$$

- The **//** represents the **SALT BRIDGE**.

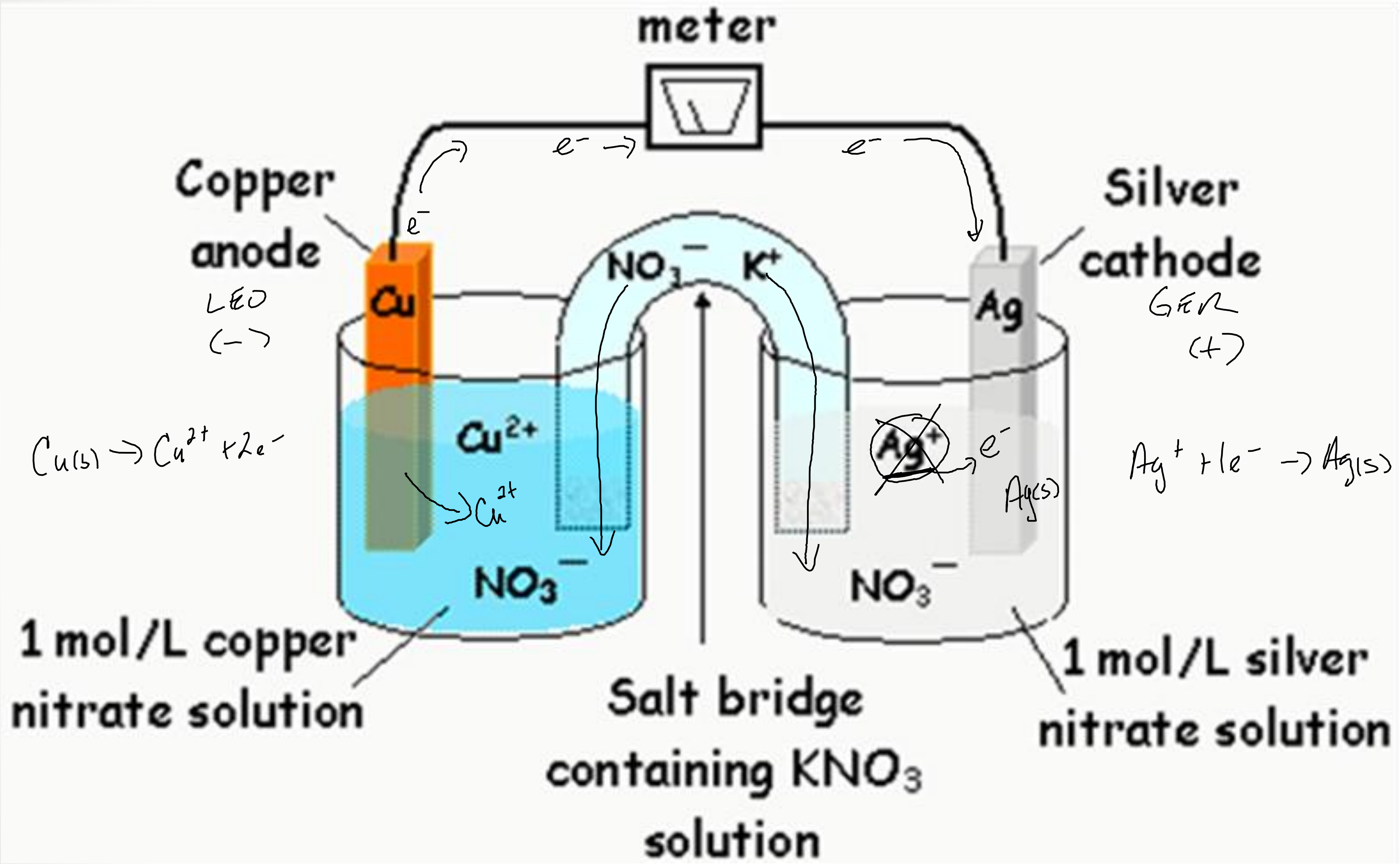
How a Cell Works:

The Salt Bridge:

- Is usually a tube containing an **ELECTROLYTE** that is **NON-REACTIVE** with the cell's **COMPONENTS**.
- It maintains the electric **NEUTRALITY** of the cell by allowing **IONS** to move between the **SOLUTIONS**.

In our example:

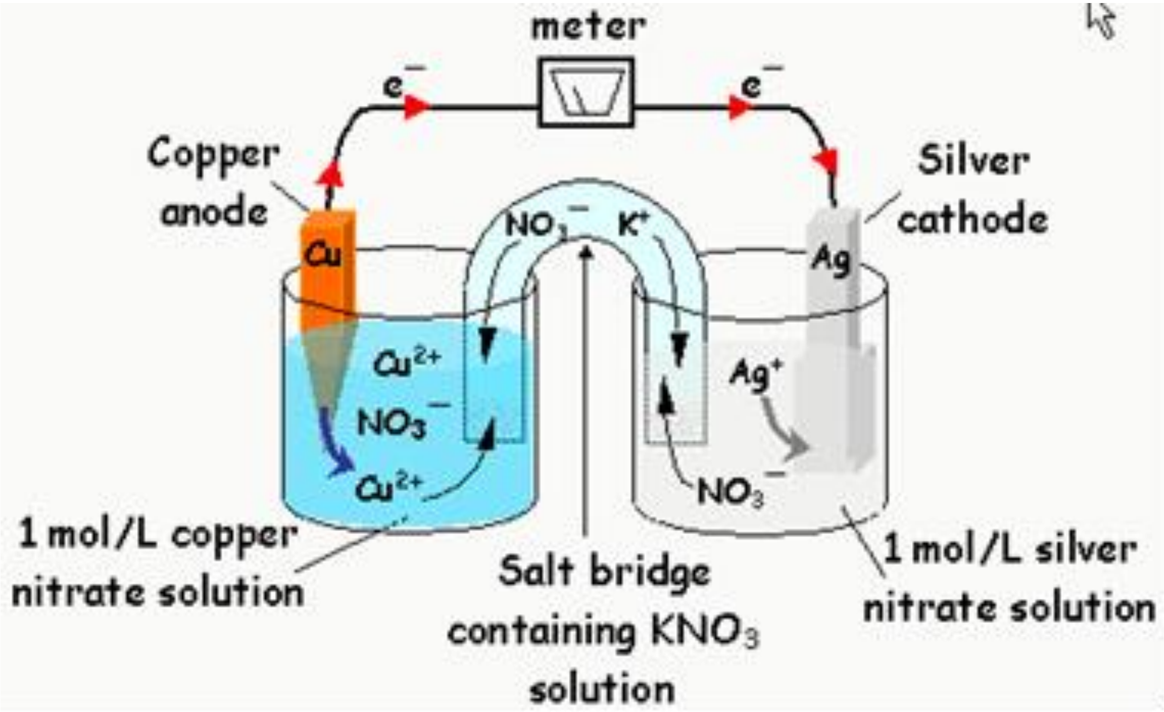
- The anode is **PRODUCING Cu^{2+} ions**, which would eventually create a **NET +VE CHARGE**.
- The cathode is **REMOVING Ag^{+} ions**, which would eventually create a **NET –VE CHARGE** → will **REPEL** electrons and cause cell to **STOP WORKING**.


meter
Copper anode
LEO
(−)
Silver cathode
GER
(+)
Cu
Ag
e^-
NO_3^-
K^+
Cu^{2+}
Ag^+
$Ag_{(s)}$
NO_3^-
NO_3^-
$Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$
$Ag^+ + 1e^- \rightarrow Ag_{(s)}$
1 mol/L copper nitrate solution
Salt bridge containing KNO_3 solution
1 mol/L silver nitrate solution

How a Cell Works:

To prevent charge buildup:

- **CATIONS** move towards the **CATHODE** to replace the lost **Ag^+**
- **ANIONS** move towards the **ANODE** to balance excess **Cu^{2+}**.



- **ANIONS** and **CATIONS** get their name from the **ELECTRODE** they are **ATTRACTED TO**.

Identifying the Electrodes:

Look at the **REDUCTION** **TABLE**

- All half-rx's are **REVERSIBLE** (can go forward or backward)
- All are written as **REDUCTIONS** (**GER**)
- Their reverse would be **OXIDATIONS** (**LEO**)
- The half-rx with the **GREATER** **POTENTIAL** to be reduced is **LOWER** on the table (**HIGHER REDUCTION POTENTIAL E°**)

So the **LOWER** half-rx is the **CATHODE** (**LIC**)

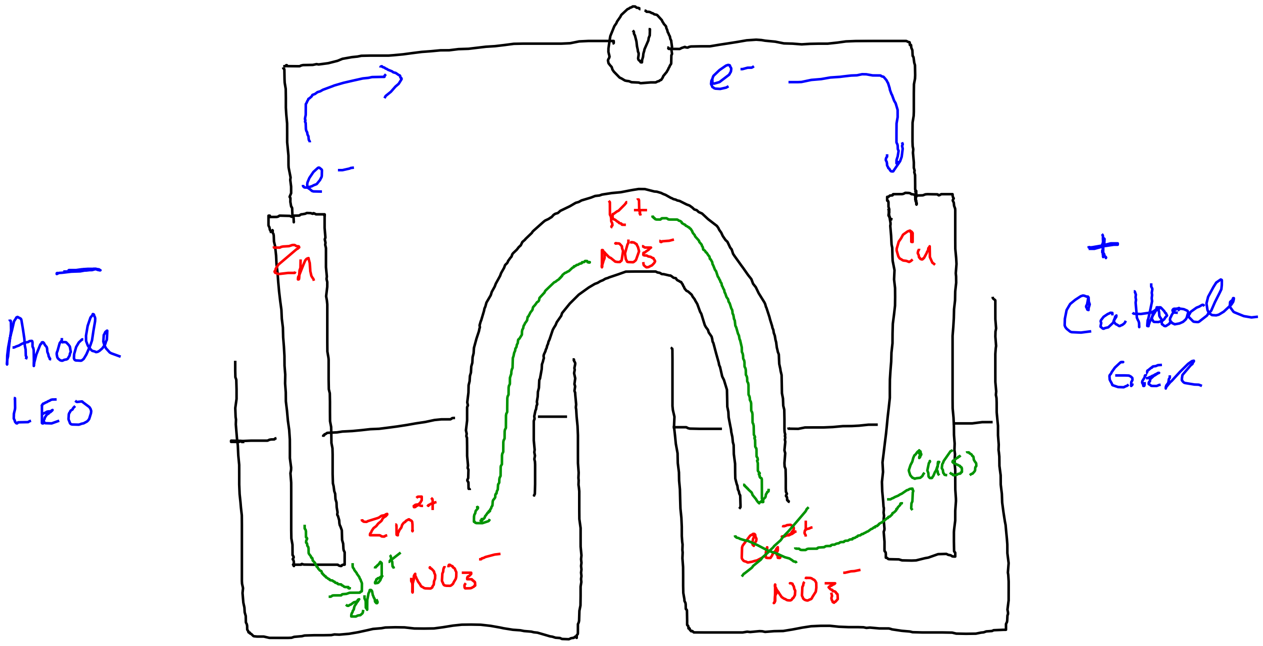
(Notice $Ag^+ + 1e^- \rightarrow Ag$ is **LOWER** than $Cu^{2+} + 2e^- \rightarrow Cu$ so Ag gets to be the **CATHODE**)

Also notice that the **ANODE** reaction Is **REVERSED** (**AIR**)

(Anode rx: $\boldsymbol{Cu \rightarrow Cu^{2+} + 2e}$)

Drawing a Cell:

Draw a zinc-copper electrochemical cell. Show movement of ions, electrons, label the anode and cathode, and identify the half-reactions at each electrode.



$Zn(s) \rightarrow Zn^{2+} + 2e^-$ $Cu^{2+} + 2e^- \rightarrow Cu(s)$