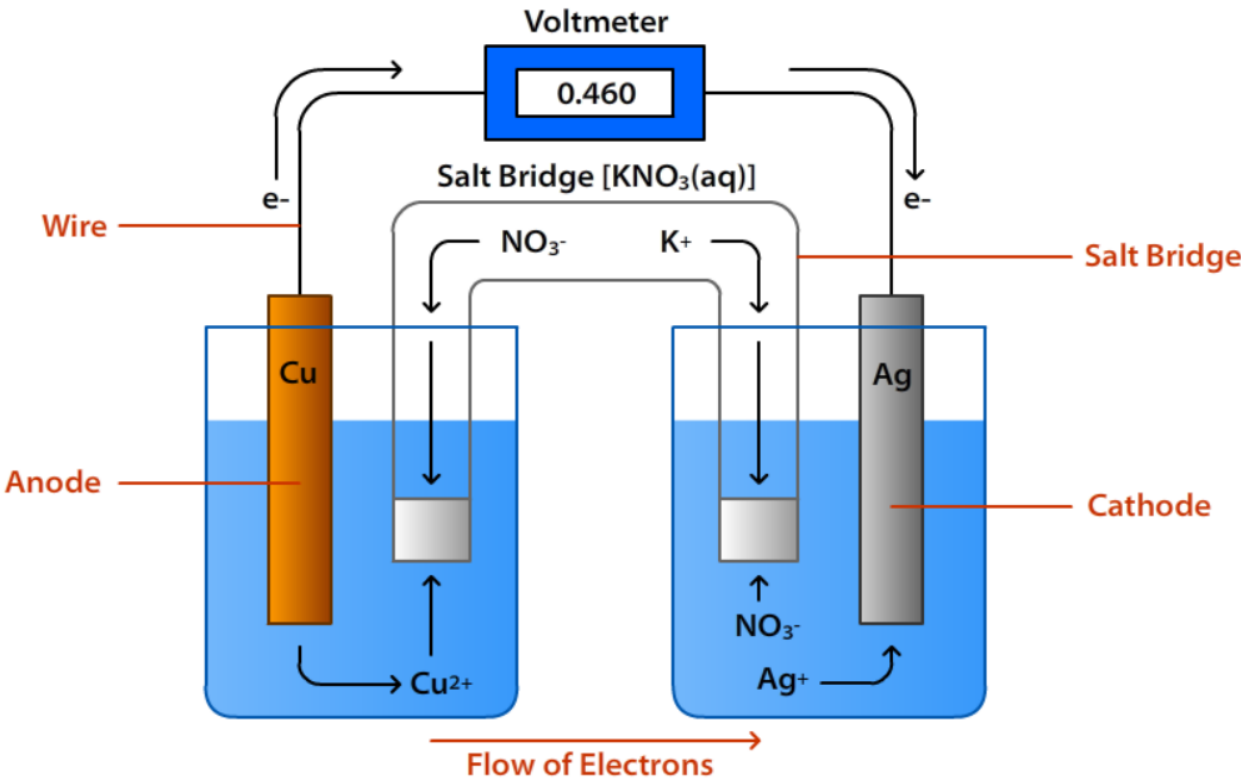
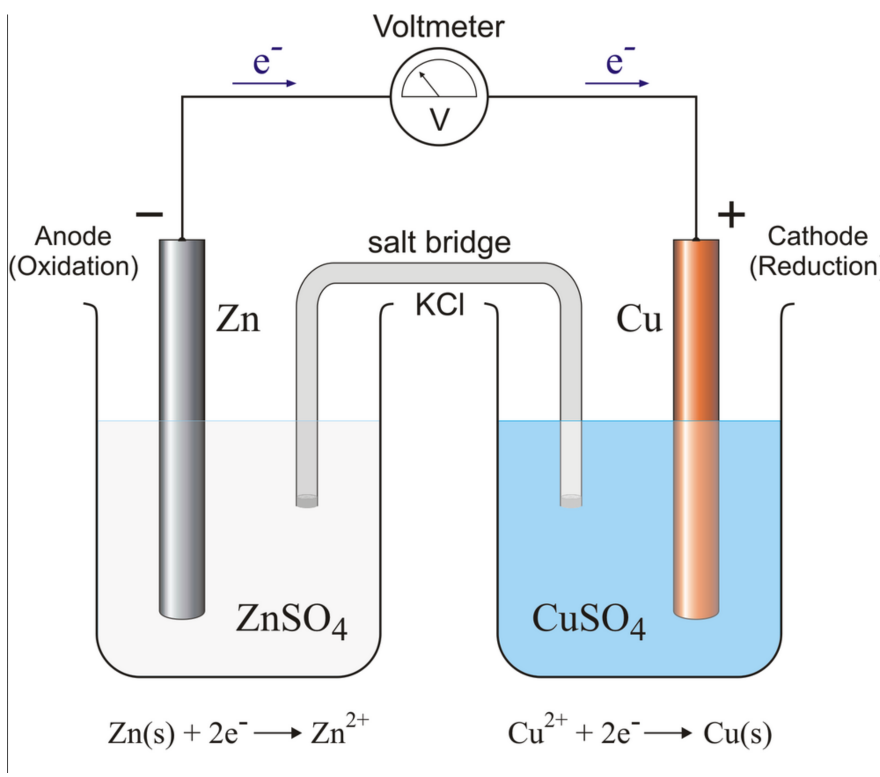
Galvanic Cells

VOLTA'S SET-UP OF AN ELECTRIC CELL





DEFINITIONS

Term	Definition
Electrochemistry	Reactions that produce (or is caused by) a flow of e^-
Electric Cell	A device that continuously converts chemical energy into electrical energy
Electrode	A solid electrical conductor
Electrolyte	An aqueous electrical conductor
Voltage/Volt	Potential for e^- to go through a circuit 1 V → 1 J/Coulomb Volt → The SI unit for electrical potential difference
Current/Ampere/Coulomb	Current → Electrical current is a measure of the amount of electrical charge transferred per unit time. It represents the flow of electrons through a conductive material. 1 A → 1 Coulomb/second Ampere → The SI unit for electrical current Coulomb → The SI unit for electrical charge
Half-Cell	An electrode and an electrolyte forming half a complete cell
Galvanic Cell	An arrangement of 2 half-cells that can produce electricity spontaneously
Cathode/Anode	Cathode (+) → electrode where reduction occurs Anode (-) → electrode where oxidation occurs
Standard Reduction Potential	ΔE^o→ represents the ability of a standard half-cell to attract electrons in a reduction half-reaction
Salt Bridge	A "U"-shaped tube that contains an inert (unreactive) aqueous electrolyte
Volta's Cell	Made up of Copper (Cu) metal, paper soaked in salt solution, and Zinc (Zn) metal

EXPERIMENT

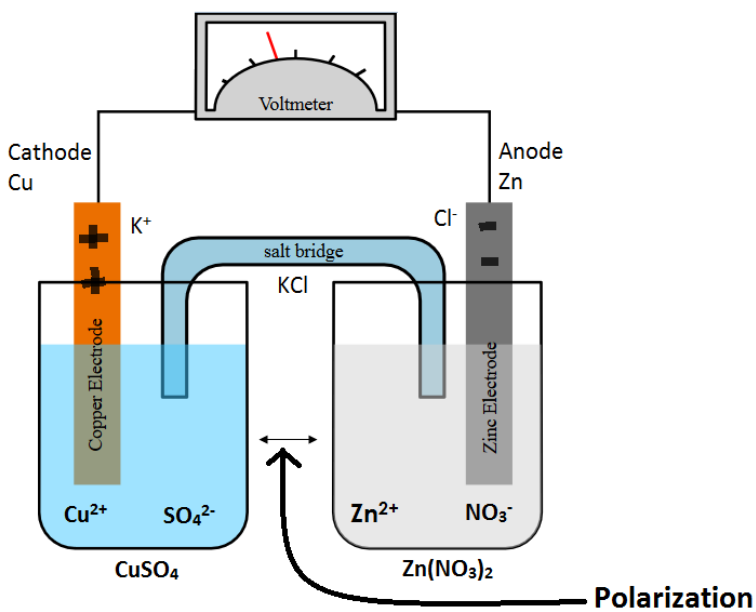
Measured in Volts (V):

	Zn	Cu	Al	Ni	C	Sn
Zn						
Cu	0.6					
Al	0	0.1				
Ni	0	0.09	0.08			
C	0	0.05	0.10	0.02		
Sn	0.1	0.3	0	0.24	0.38	

The components of all electric cells are:

- 2 electrodes
- 1 electrolyte

	Observation	**Interpretation**
Cu Electrode	Buildup of Cu	Reduced as $Cu^{2+} + 2e^- \rightarrow Cu$
Zn Electrode	White residue →	Due to Sulfate
	Getting smaller	Oxidized as $Zn \rightarrow Zn^{2+} + 2\,e^-$
U-tube	Blue (Cu^{2+} ions)	Cu^{2+} ions in the tube
Zn Solution	Limey Green Solution	Presence of SO_4^{2-} ions
Cu Solution	Blue solution	Cu^{2+} ions present within the solution
Voltmeter Reading	0.45 V	Spontaneous Reaction → Redox reaction



Cathode

- $Cu^{2+} + 2e^- \rightarrow Cu$
- $E^o = +0.34\ V$

Anode

- $Zn \rightarrow Zn^{2+} + 2e^-$
- $E^o = -0.76\ V$

Final Equation

$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$

$\Delta E^o = +0.34V - (-0.76V)$

$\Delta E^o = +1.10V$

There are 3 types of Galvanic Cells:

1. 2 Metal Electrodes
2. Metal + Inert
3. 2 Inert Electrodes

TWO METAL ELECTRODES

Using the Standard Reduction Potential Table to predict the anodes and cathodes for the following:

Combination	Anode (Lowest Voltage)	Cathode (Highest Voltage)	Half-Reactions
Ni + Sn	Ni → -0.25V	Sn → -0.16V	$Ni \rightarrow Ni^{2+} + 2\,e^- \;\; E^o = +0.25V$ $Sn^{2+} + 2\,e^- \rightarrow Sn \;\; E^o = -0.16V$ Final Equation: $Ni + Sn^{2+} \rightarrow Ni^{2+} + Sn \;\; \Delta E^o = 0.09V$
Al + Zn	Al → -1.66V	Zn → -0.76V	$2 \cdot (Al \rightarrow Al^{3+} + 3\,e^-) \;\; E^o = +1.66V$ $3 \cdot (Zn^{2+} + 2\,e^- \rightarrow Zn) \;\; E^o = -0.76V$ Final Equation: $2\,Al + 3\,Zn^{2+} \rightarrow 2\,Al^{3+} + 3\,Zn \;\; \Delta E^o = 0.90V$
Sn + Al	Al → -1.66V	Sn → -0.16	$2 \cdot (Al \rightarrow Al^{3+} + 3\,e^-) \;\; E^o = +1.66V$ $3 \cdot (Sn^{2+} + 2\,e^- \rightarrow Sn) \;\; E^o = -0.16V$ Final Equation: $2\,Al + 3\,Sn^{2+} \rightarrow 2\,Al^{3+} + 3\,Sn \;\; \Delta E^o = 1.50V$
Cu + Al	Al → -1.66V	Cu → +0.34V	$2 \cdot (Al \rightarrow Al^{3+} + 3\,e^-) \;\; E^o = +1.66V$ $3 \cdot (Cu^{2+} + 2\,e^- \rightarrow Cu) \;\; E^o = +0.34V$ Final Equation: $2\,Al + 3\,Cu^{2+} \rightarrow 2\,Al^{3+} + 3\,Cu \;\; \Delta E^o = 2.00V$
Ni + Zn	Zn → -0.76V	Ni → -0.25V	$Zn^{2+} + 2\,e^- \rightarrow Zn \;\; E^o = +0.76V$ $Ni \rightarrow Ni^{2+} + 2\,e^- \;\; E^o = -0.25V$ Final Equation: $Ni + Zn^{2+} \rightarrow Ni^{2+} + Zn \;\; \Delta E^o = 0.51V$

METAL + INERT ELECTRODES

Terms	Definition (What the term means for this type of cell)
Inert	Non-reactive (Pt, Carbon)
Anode	Metal Electrode

Cathode	Reduction of ions

Ex. Carbon + Metal Electrode

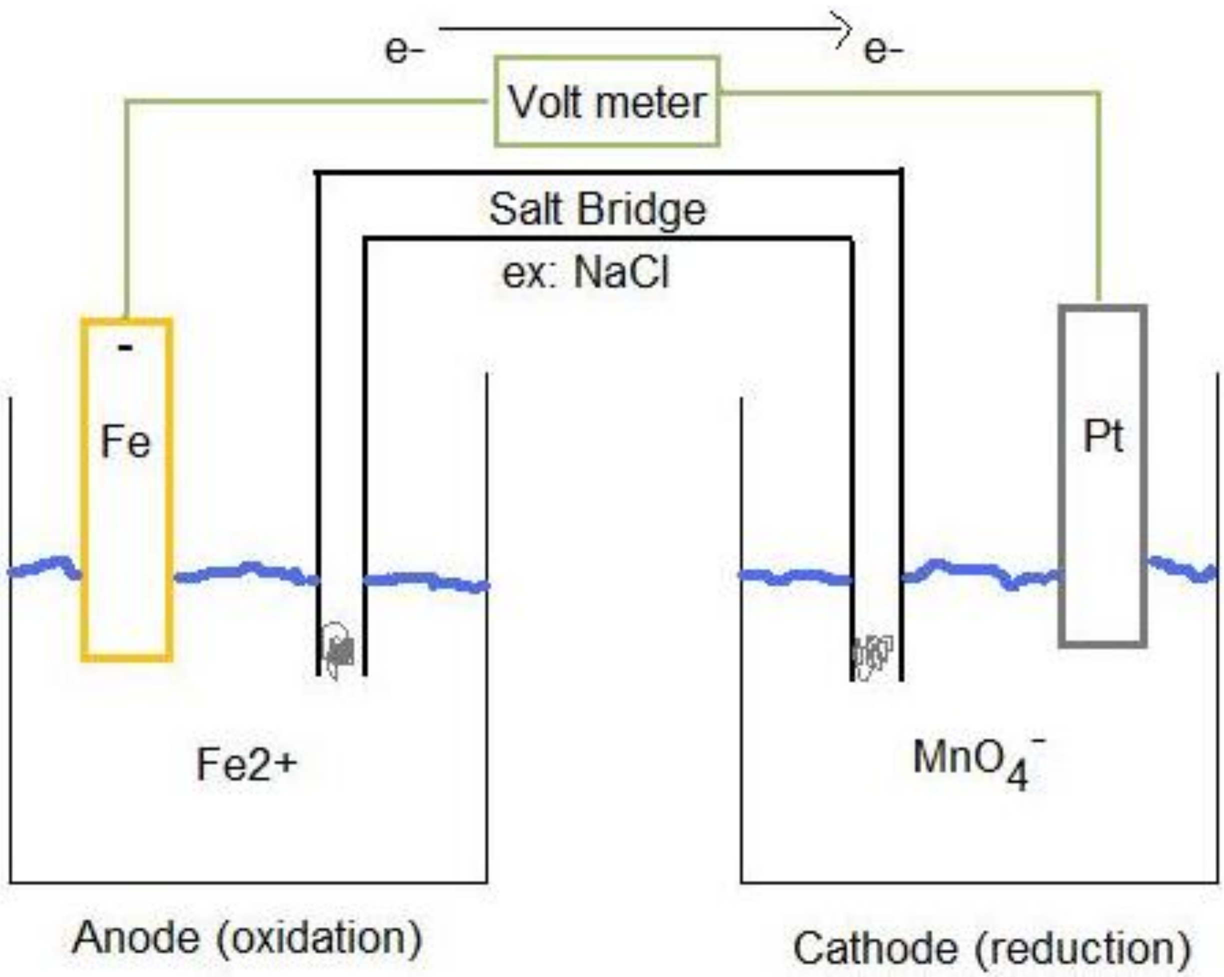
What reaction occurred in the solution ($NaNO_3$)?

$$NO_3^- + 3H^+ + 2e^- \rightarrow HNO_2 + H_2O \;\; E^o = +0.94V$$

An inert electrode is a metal submerged in an aqueous solution of ion compounds that transfers electrons rather than exchanging ions with the aqueous solution. It does not participate or interfere in the chemical reaction but serves as a source of electrons. Platinum is usually the metal used as an inert electrode. An active electrode is an electrode that can be oxidized or reduced in half reaction.

Anode: $Fe^{2+}{}_{(aq)} \rightarrow Fe^{3+}{}_{(s)} + e^-$

Cathode: $MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5e^- \rightarrow Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$



TWO INERT ELECTRODES

Inert → non-reactive

Galvanic Cells are characterized by their ability to induce spontaneous reactions. However, the primary difference between the various types of Galvanic Cells is simply the type of electrode used. Specifically, the primary element found within the electrode.

Steps

1. Find 2 equations in the Standard Reduction Potential Table that contain all of the ions
2. Determine the anode and cathode
3. Reverse anode reaction and sign on the voltage
4. Balance the electrons
5. Add the equations

Example 1

Consider the following combination of ions:

BrO_3^- H^+ Br^- H_2O $S_4O_6^{2-}$ $S_2O_3^{2-}$

1. "*Find 2 equations in the Standard Reduction Potential Table that contain all of the ions*"

In this example, the two equations that contain all of the ions are:

- $BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O \quad E^o = +1.44$
- $S_4O_6^{2-} + 2e^- \rightarrow 2\, S_2O_3^{2-} \quad E^o = +0.169$

2. "*Determine the anode and cathode*"

Remember that the anode of a galvanic cell is the ion of the equation with the lowest voltage. Anode of a galvanic cell is where reduction occurs, thus it has to be lower on the Standard **Reduction** Potential table.

- Anode: $S_4O_6^{2-} \quad E^o = +0.169$
- Cathode: $BrO_3^- \quad E^o = +1.44$

3. "*Reverse anode reaction and sign on the voltage*"

To reverse a reaction, the products of original anode equation become the reactants. Due to this, the sign on the voltage is reversed as well.

Original Equation:

$S_4O_6^{2-} + 2e^- \rightarrow 2\, S_2O_3^{2-} \quad E^o = +0.169$

Equation upon reversal:

$2\, S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^- \quad E^o = -0.169$

4. "*Balance electrons*"

Let's look at our 2 equations again.

$BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O \quad E^o = +1.44$

$2\, S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^- \quad E^o = -0.169$

There are $6e^-$ reacting, but only $2e^-$ being produced. Therefore, we must multiply the anode reaction by a factor of 3 ($Numeric\ Factor = \frac{Number\ of\ electrons\ reacting}{Number\ of\ electrons\ produced} = \frac{6}{2} = 3$) in order to balance the electrons. Remember **NOT** to multiply the voltage.

$3 \cdot [2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-] \quad E^o = -0.169$

$= 6\ S_2O_3^{2-} \rightarrow 3\ S_4O_6^{2-} + 6e^- \quad E^o = -0.169$

After balancing:

$BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O \quad E^o = +1.44$

$6\ S_2O_3^{2-} \rightarrow 3\ S_4O_6^{2-} + 6e^- \quad E^o = -0.169$

5. *"Add the equations"*

$(BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O \quad E^o = +1.44) \quad + \quad (6\ S_2O_3^{2-} \rightarrow 3\ S_4O_6^{2-} + 6e^- \quad E^o = -0.169)$

$BrO_3^- + 6H^+ + 6\ S_2O_3^{2-} + \cancel{6e^-} \rightarrow Br^- + 3H_2O + 3\ S_4O_6^{2-} + \cancel{6e^-} \quad \Delta E^o = +1.44 - 0.169 = +1.271$

Final Equation:

$BrO_3^- + 6H^+ + 6\ S_2O_3^{2-} \rightarrow Br^- + 3H_2O + 3\ S_4O_6^{2-} \quad \Delta E^o = +1.271$

Example 2

$MnO_4^- \quad H^+ \quad Mn^{2+} \quad H_2SO_3 \quad SO_4^{2-} \quad H_2O$

Cathode → $\boldsymbol{E^o = +1.49} \quad MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

Anode → $\boldsymbol{E^o = +0.17} \quad SO_4^{2-} + 4\ H^+ + 2\ e^- \rightarrow H_2SO_3 + H_2O$

After Reversal → $\boldsymbol{E^o = -0.17} \quad H_2SO_3 + H_2O \rightarrow SO_4^{2-} + 4\ H^+ + 2\ e^-$

Balancing e^- → $2 \cdot [MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O]$
$5 \cdot [H_2SO_3 + H_2O \rightarrow SO_4^{2-} + 4\ H^+ + 2\ e^-]$

ΔE^o → $\Delta E^o = +1.49 - 0.17 = +1.32$

Add Equations → $\mathbf{2}\ MnO_4^- + 5\ H_2SO_3 + 5\ H_2O + 16\ H^+ \rightarrow 2\ Mn^{2+} + 8\ H_2O + 20\ H^+ + 5\ SO_4^{2-}$

Final Equation → $\mathbf{2}\ MnO_4^- + 5\ H_2SO_3 + 5\ H_2O \rightarrow 2\ Mn^{2+} + 8\ H_2O + 4\ H^+ + 5\ SO_4^{2-}\ \Delta E^o = +1.32$