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 circuit1 V 🡪 1 J/CoulombVolt 🡪 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/secondAmpere 🡪 The SI unit for electrical currentCoulomb 🡪 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 occursAnode (-) 🡪 electrode where oxidation occurs |
| Standard Reduction Potential | $∆E^{o}$🡪 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

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| --- |
| **Cathode*** Cu2+ + 2e- 🡪 Cu
* $E^{o}=+0.34 V$
 |
| **Anode*** Zn 🡪 Zn2+ + 2e-
* $E^{o}=-0.76 V$
 |
| **Final Equation**Cu2+ + Zn 🡪 Cu + Zn2+ $$∆E^{o}=+0.34V-\left(-0.76V\right)$$$$∆E^{o}=+1.10V$$ |

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| --- | --- | --- |
|  | Observation | Interpretation |
| Cu Electrode | Buildup of Cu | Reduced as $Cu^{2+}+2e ^{-}\rightarrow Cu$ |
| Zn Electrode | White residueGetting smaller | Due to SulfateOxidized as $ Zn\rightarrow Zn^{2+}+2 e^{-}$ |
| U-tube | Blue (Cu2+ ions) | Cu2+ ions in the tube  |
| Zn Solution | Limey Green Solution | Presence of SO42- ions |
| Cu Solution | Blue solution | Cu2+ ions present within the solution |
| Voltmeter Reading | 0.45 V | Spontaneous Reaction 🡪Redox reaction |



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:$$Ni+Sn^{2+}\rightarrow Ni^{2+}+Sn ∆E^{o}=0.09V$$ |
| Al + Zn | Al 🡪 -1.66V | Zn 🡪 -0.76V | $2∙\left(Al\rightarrow Al^{3+}+3 e^{-}\right) E^{o}=+1.66V$ $3∙\left(Zn^{2+}+2 e^{-}\rightarrow Zn\right) E^{o}=-0.76V$ Final Equation:$$2 Al+3 Zn^{2+}\rightarrow 2 Al^{3+}+3 Zn ∆E^{o}=0.90V$$ |
| Sn + Al | Al 🡪 -1.66V | Sn 🡪 -0.16 | $2∙\left(Al\rightarrow Al^{3+}+3 e^{-}\right) E^{o}=+1.66V$ $$3∙\left(Sn^{2+}+2 e^{-}\rightarrow Sn\right) E^{o}=-0.16V$$Final Equation:$$2 Al+3 Sn^{2+}\rightarrow 2 Al^{3+}+3 Sn ∆E^{o}=1.50V$$ |
| Cu + Al | Al 🡪 -1.66V | Cu 🡪 +0.34V | $2∙\left(Al\rightarrow Al^{3+}+3 e^{-}\right) E^{o}=+1.66V$ $$3∙\left(Cu^{2+}+2 e^{-}\rightarrow Cu\right) E^{o}=+0.34V$$Final Equation:$$2 Al+3 Cu^{2+}\rightarrow 2 Al^{3+}+3 Cu ∆E^{o}=2.00V$$ |
| 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:$$Ni+Zn^{2+}\rightarrow Ni^{2+}+Zn ∆E^{o}=0.51V$$ |

# Metal + inert electrodes

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| --- | --- |
| 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 ElectrodeWhat reaction occurred in the solution (NaNO3)?$$NO\_{3}^{-}+3H^{+}+2e^{-}\rightarrow HNO\_{2}+H\_{2}O E^{o}=+0.94V$$ |

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:** Fe2+ (aq) → Fe3+(s) + e-

**Cathode:** MnO4-(aq) + 8H+ (aq) + 5e- → Mn2+ (aq) + 4H2O (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:

$$BrO\_{3}^{-} H^{+} Br^{-} H\_{2}O S\_{4}O\_{6}^{2-} S\_{2}O\_{3}^{2-}$$

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:

* $BrO\_{3}^{-}+6H^{+}+6e^{-} \rightarrow Br^{-}+3H\_{2}O E^{o}=+1.44$
* $S\_{4}O\_{6}^{2-}+2e^{-} \rightarrow 2 S\_{2}O\_{3}^{2-} E^{o}=+0.169$
1. “*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\_{4}O\_{6}^{2-} E^{o}=+0.169 $
* Cathode: $BrO\_{3}^{-} E^{o}=+1.44$
1. “*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:

$$S\_{4}O\_{6}^{2-}+2e^{-} \rightarrow 2 S\_{2}O\_{3}^{2-} E^{o}=+0.169$$

Equation upon reversal:

$$2 S\_{2}O\_{3}^{2- }\rightarrow S\_{4}O\_{6}^{2-}+2e^{-} E^{o}=-0.169$$

1. “*Balance electrons*”

Let’s look at our 2 equations again.

$$BrO\_{3}^{-}+6H^{+}+6e^{-} \rightarrow Br^{-}+3H\_{2}O E^{o}=+1.44$$

$$2 S\_{2}O\_{3}^{2- }\rightarrow S\_{4}O\_{6}^{2-}+2e^{-} E^{o}=-0.169$$

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.

$$3∙[2 S\_{2}O\_{3}^{2- }\rightarrow S\_{4}O\_{6}^{2-}+2e^{-}] E^{o}=-0.169$$

$$=6 S\_{2}O\_{3}^{2- }\rightarrow 3 S\_{4}O\_{6}^{2-}+6e^{-} E^{o}=-0.169$$

**After balancing:**

$$BrO\_{3}^{-}+6H^{+}+6e^{-} \rightarrow Br^{-}+3H\_{2}O E^{o}=+1.44$$

$$6 S\_{2}O\_{3}^{2- }\rightarrow 3 S\_{4}O\_{6}^{2-}+6e^{-} E^{o}=-0.169$$

1. “*Add the equations*”

$$(BrO\_{3}^{-}+6H^{+}+6e^{-} \rightarrow Br^{-}+3H\_{2}O E^{o}=+1.44) + (6 S\_{2}O\_{3}^{2- }\rightarrow 3 S\_{4}O\_{6}^{2-}+6e^{-} E^{o}=-0.169)$$

$$BrO\_{3}^{-}+6H^{+}+ 6 S\_{2}O\_{3}^{2- }+6e^{-}\rightarrow Br^{-}+3H\_{2}O + 3 S\_{4}O\_{6}^{2-}+ 6e^{-} ∆E^{o}=+1.44-0.169=+1.271$$

**Final Equation:**

$$BrO\_{3}^{-}+6H^{+}+ 6 S\_{2}O\_{3}^{2- }\rightarrow Br^{-}+3H\_{2}O + 3 S\_{4}O\_{6}^{2-} ∆E^{o}=+1.271$$

### **Example 2**

$$MnO\_{4}^{-} H^{+} Mn^{2+} H\_{2}SO\_{3} SO\_{4}^{2-} H\_{2}O$$

|  |  |
| --- | --- |
| **Cathode 🡪****Anode 🡪****After Reversal 🡪****Balancing e- 🡪** | $E^{o}=+1.49 MnO\_{4}^{-}+8 H^{+}+5 e^{-} \rightarrow Mn^{2+}+4 H\_{2}O$  |
| $E^{o}=+0.17 SO\_{4}^{2-}+4 H^{+}+2 e^{-} \rightarrow H\_{2}SO\_{3}+H\_{2}O$ $E^{o}=-0.17 H\_{2}SO\_{3}+H\_{2}O \rightarrow SO\_{4}^{2-}+4 H^{+}+2 e^{-} $ $$2∙[MnO\_{4}^{-}+8 H^{+}+5 e^{-} \rightarrow Mn^{2+}+4 H\_{2}O]$$$$5∙[H\_{2}SO\_{3}+H\_{2}O \rightarrow SO\_{4}^{2-}+4 H^{+}+2 e^{-}]$$ |
| $∆E^{o}$🡪 | $$∆E^{o}=+1.49-0.17=+1.32$$ |
| **Add Equations 🡪****Final Equation 🡪** | $$2 MnO\_{4}^{-}+5 H\_{2}SO\_{3}+5 H\_{2}O +16 H^{+}\rightarrow 2 Mn^{2+}+8 H\_{2}O+20 H^{+}+5 SO\_{4}^{2-}$$$$2 MnO\_{4}^{-}+5 H\_{2}SO\_{3}+5 H\_{2}O\rightarrow 2 Mn^{2+}+8 H\_{2}O+4 H^{+}+5 SO\_{4}^{2-} ∆E^{o}=+1.32 $$ |