Electrochemistry
CHEM 405 & 505

INTRODUCTION TO ELECTROCHEMISTRY

Definition

- Studies the relationship between electrical energy and chemical reactions.

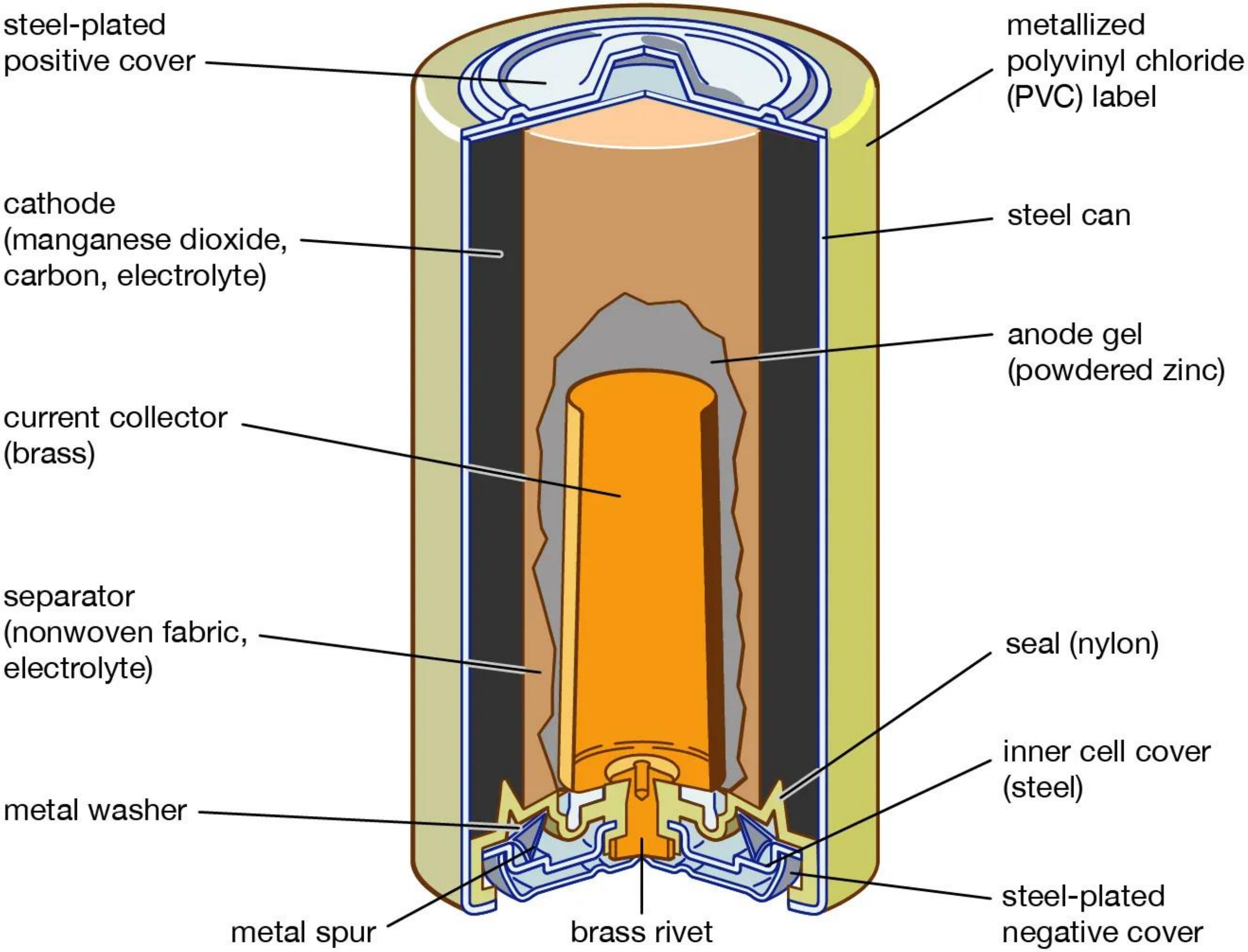
- involves processes where chemical reactions produce either electrical energy or are driven by an electric current.

INTRODUCTION TO ELECTROCHEMISTRY

- Key concepts in electrochemistry include redox (reduction-oxidation) reactions, the movement of electrons, and the role of ions in conducting electricity in solutions.

- This field is essential in developing batteries, fuel cells, electrolysis, corrosion, and various industrial processes.

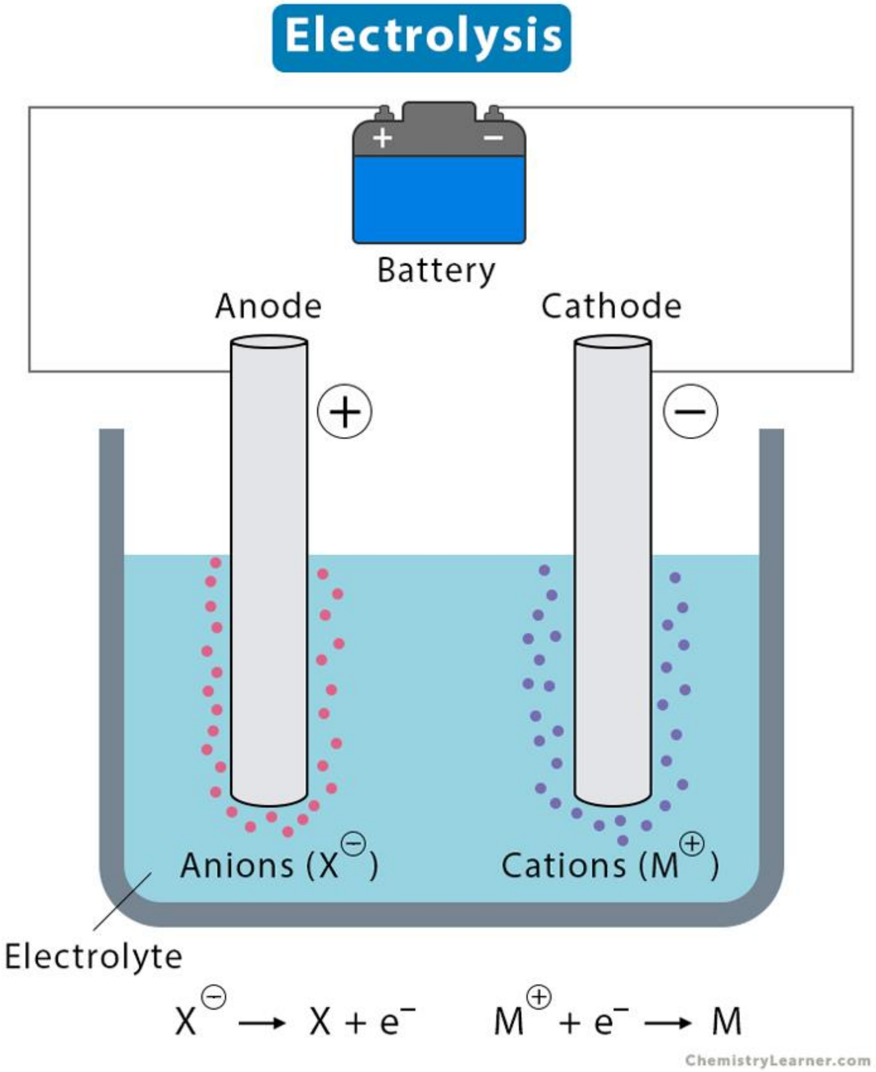
Certain chemical reactions can create electricity. (e.g., batteries).



Cutaway view of an alkaline–manganese dioxide power cell.

https://www.britannica.com/technology/battery-electronics#/media/1/56126/19439

Electricity can make specific chemical reactions happen that wouldn’t happen otherwise. (e.g., electrolysis).



What is ELECTRICITY?

Movement of electrons

$$e^- \rightarrow e^- \rightarrow e^- \rightarrow e^- \rightarrow e^- \rightarrow e^-$$

THROUGH SOMETHING

.

.

.

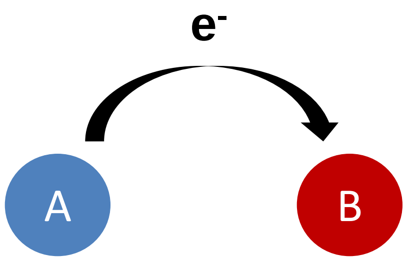
e.g., WIRE

LIGHTBULB

BATTERY

IF Certain chemical reactions can create electricity, then there is a movement of electrons between atoms

Oxidation-Reduction Reactions



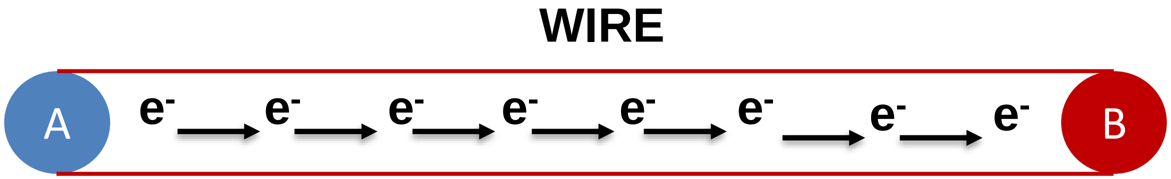
Electrons move between atoms.

Movement of electrons creates electricity

$e^- \rightarrow e^- \rightarrow e^- \rightarrow e^- \rightarrow e^- \rightarrow e^-$

IF certain chemical reactions can create electricity, then there is a movement of electrons between atoms

Oxidation-Reduction
Reactions



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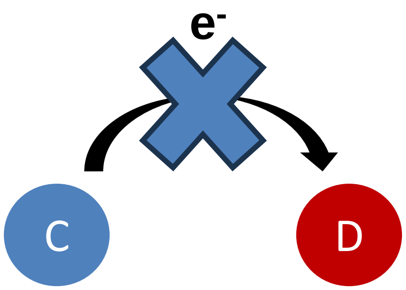
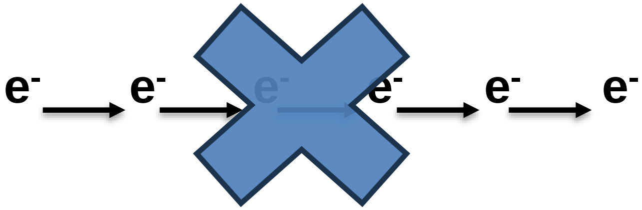
Oxidation-Reduction Reactions



Electrons DON’T move between atoms.



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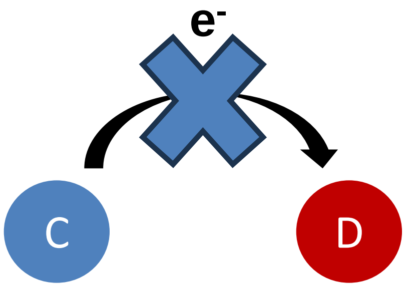
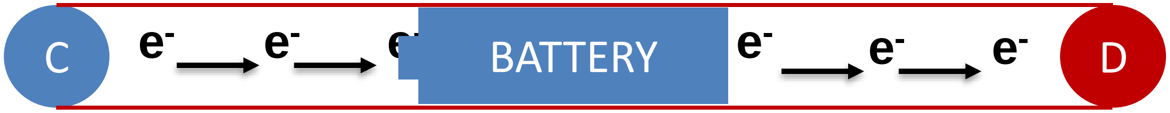
Oxidation-Reduction Reactions

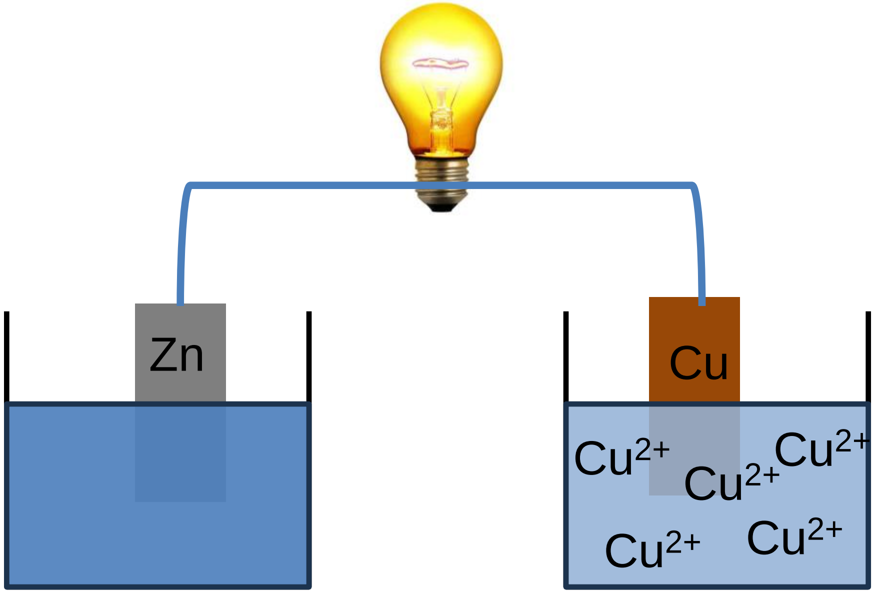




PULL & PUSH

How certain chemical reactions can create electricity?

For example:



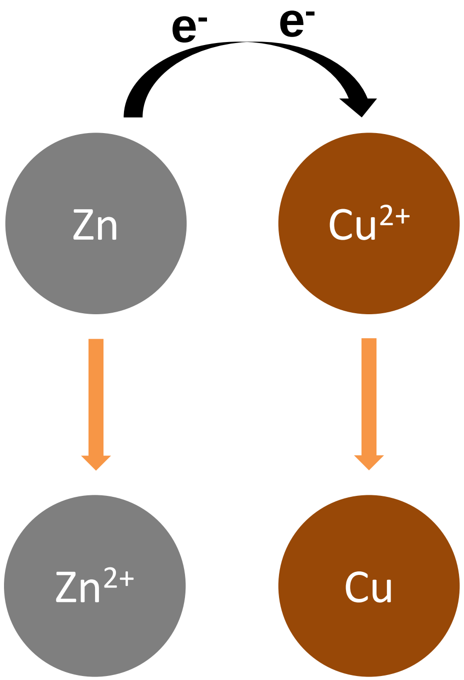
Galvanic or Voltaic Cell – creates electricity using a chemical reaction

How certain chemical reactions can create electricity?

For example:

Zn weakly pulls electrons

Zn **loses** electrons
It is **oxidized**



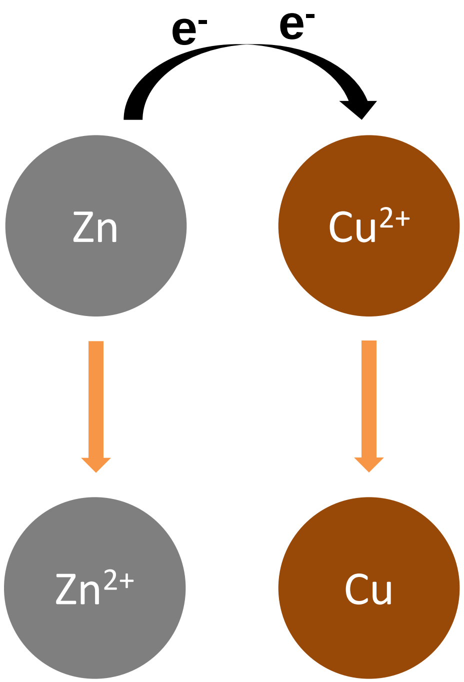
Cu^{2+} strongly pulls electrons

Cu^{2+} **gains** electrons
It is **reduced**

Oxidation - Loss of electrons
Reduction - Gain of electrons

How certain chemical reactions can create electricity?

For example:



Zn weakly pulls electrons

Zn **loses** electrons
It is **oxidized**

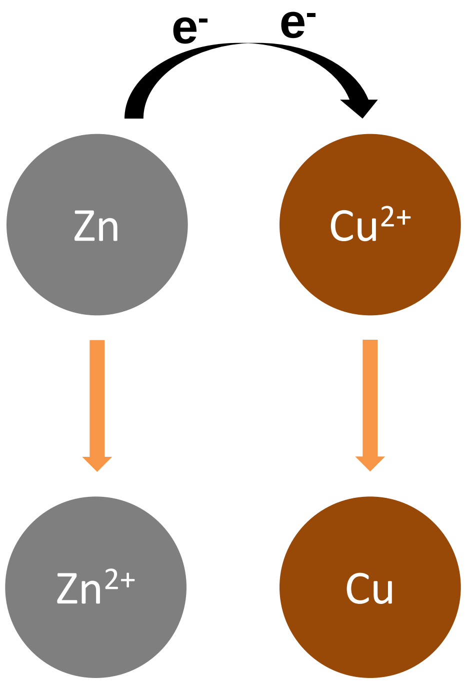
Cu^{2+} strongly pulls electrons

Cu^{2+} **gains** electrons
It is **reduced**

This process happens on its own.
It is spontaneous.

How certain chemical reactions can create electricity?

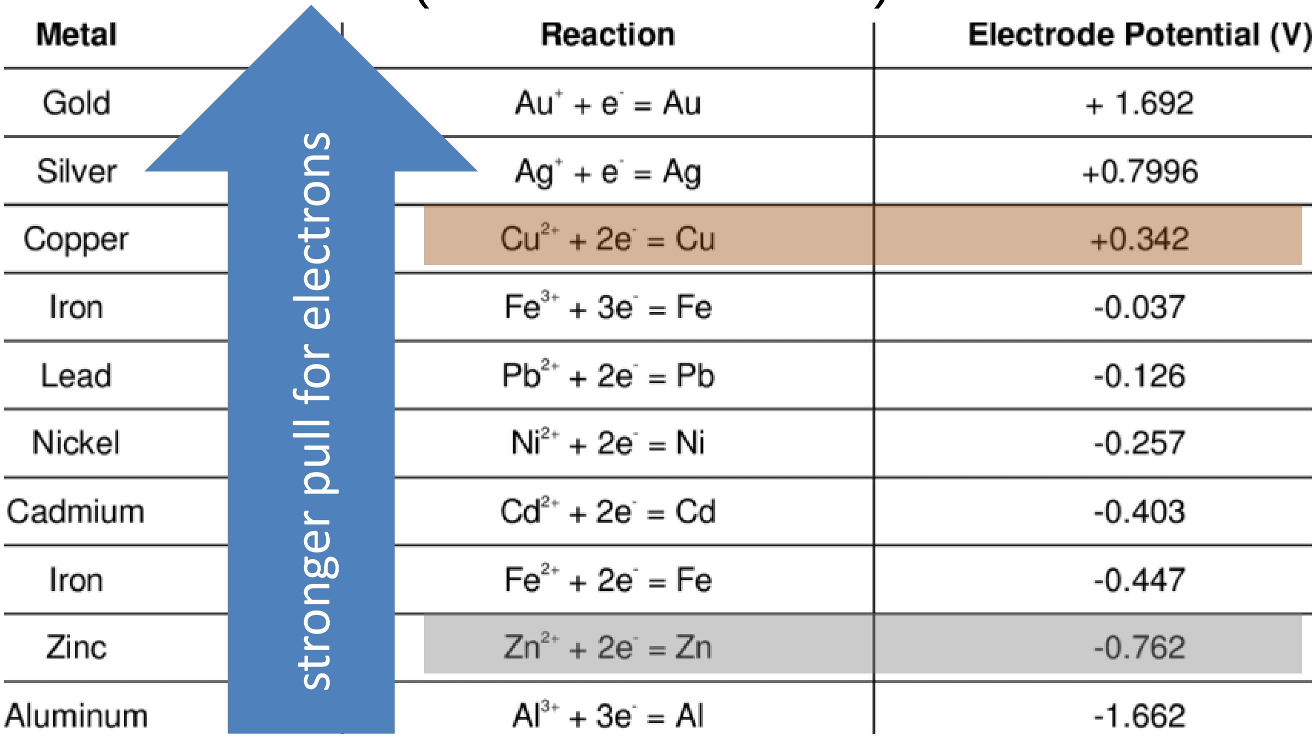
How do we ensure that electrons spontaneously move from Zn to Cu^{2+}?



It is spontaneous.

Standard reduction potentials

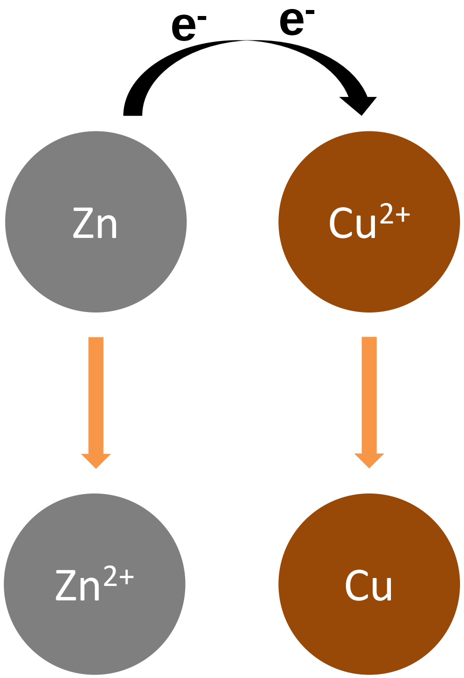
(relative to the SHE)



Metal	Reaction	Electrode Potential (V)
Gold	$Au^{+} + e^{-} = Au$	+ 1.692
Silver	$Ag^{+} + e^{-} = Ag$	+0.7996
Copper	$Cu^{2+} + 2e^{-} = Cu$	+0.342
Iron	$Fe^{3+} + 3e^{-} = Fe$	-0.037
Lead	$Pb^{2+} + 2e^{-} = Pb$	-0.126
Nickel	$Ni^{2+} + 2e^{-} = Ni$	-0.257
Cadmium	$Cd^{2+} + 2e^{-} = Cd$	-0.403
Iron	$Fe^{2+} + 2e^{-} = Fe$	-0.447
Zinc	$Zn^{2+} + 2e^{-} = Zn$	-0.762
Aluminum	$Al^{3+} + 3e^{-} = Al$	-1.662

Table: Credit to Thesis “Environmental Factors Governing Corrosion Rates on Overhead Transmission Structures, Hardware and Conductors”, December 2011 Advisor: Fabio Bologna, Murray B. Mcbride, Neal Scott Murray

How certain chemical reactions can create electricity?



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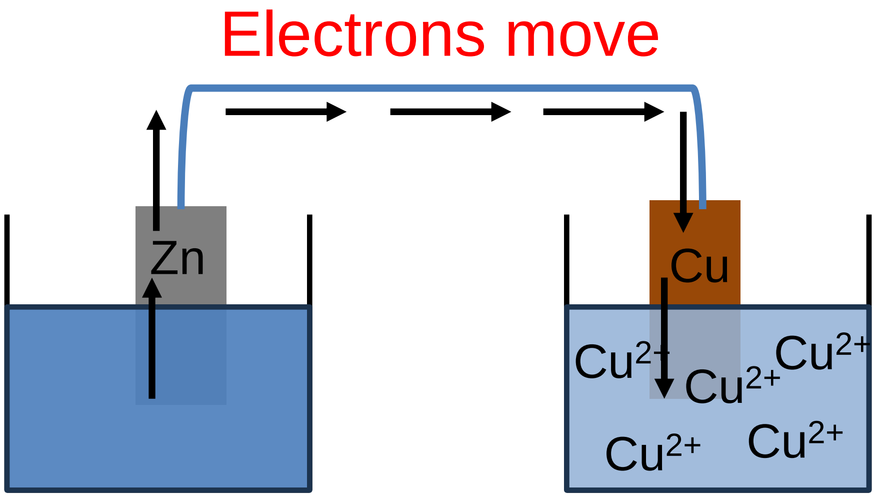
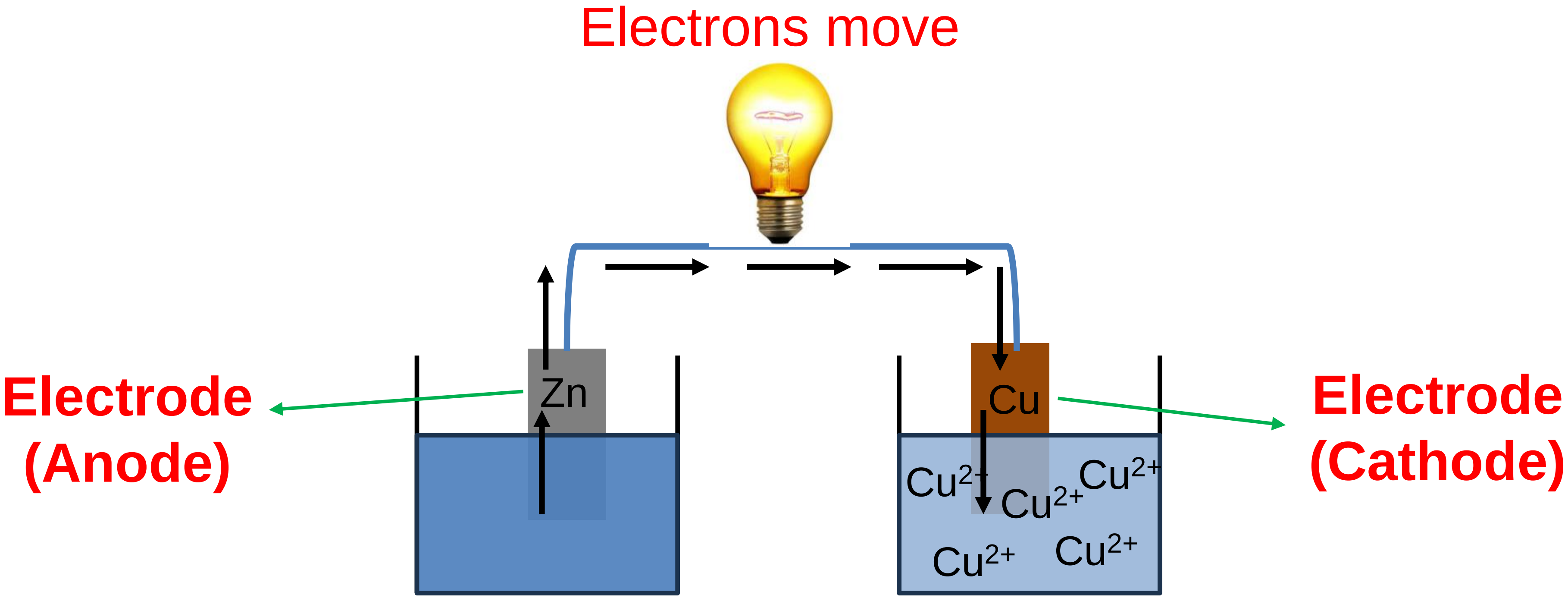


Galvanic or Voltaic Cell – creates electricity using a chemical reaction

How certain chemical reactions can create electricity?



Galvanic or Voltaic Cell –
creates electricity using a
chemical reaction

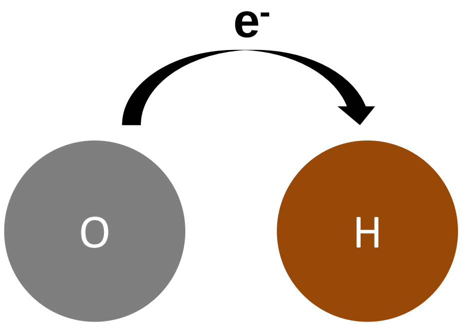
Anode where **Oxidation** happens - Loss of electrons (**An Ox**)
Cathode where **Reduction** - Gain of electrons (**Red Cat**)

Electricity can make specific chemical reactions happen that wouldn't happen otherwise. (e.g., electrolysis).

Oxygen **loses** electrons
It is **oxidized**

$$\overset{+1}{H_2}\overset{-2}{O} \longrightarrow \overset{0}{H_2} + ½\ \overset{0}{O_2}$$

Hydrogen **gains** electrons
It is **reduced**



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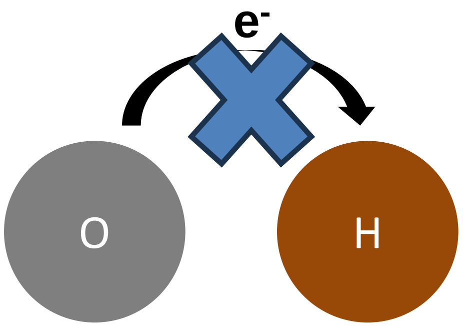

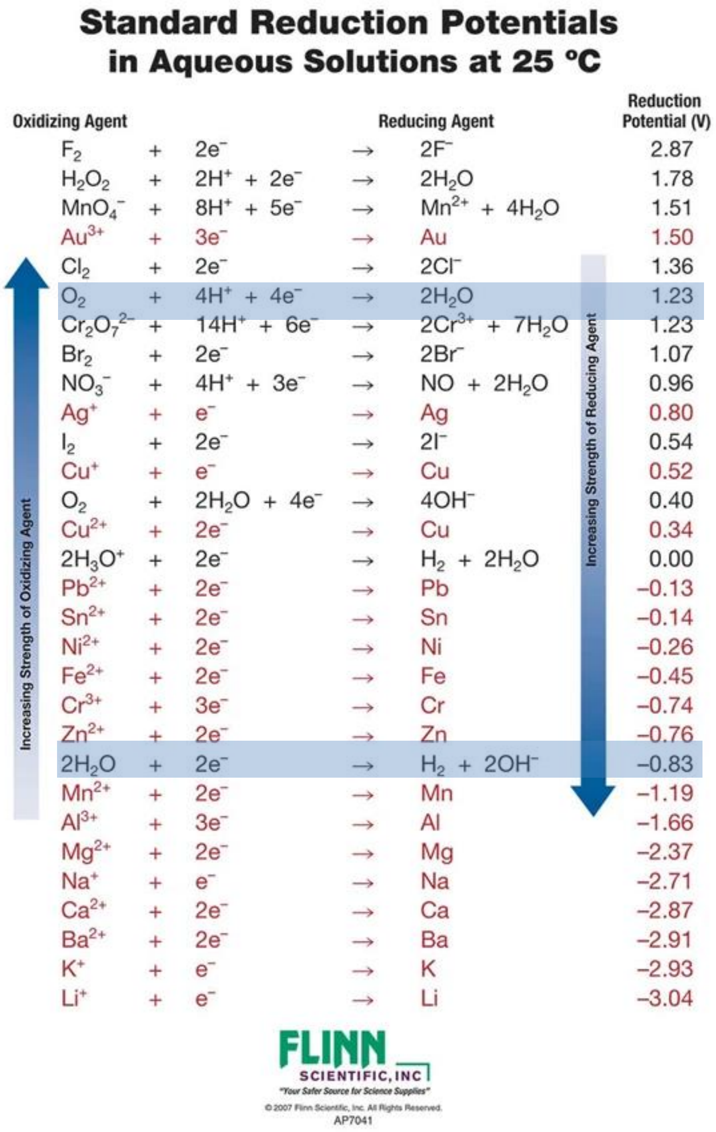

Standard Reduction Potentials in Aqueous Solutions at 25 °C

Oxidizing Agent				Reducing Agent	Reduction Potential (V)
F_2	+	$2e^-$	→	$2F^-$	2.87
H_2O_2	+	$2H^+ + 2e^-$	→	$2H_2O$	1.78
MnO_4^-	+	$8H^+ + 5e^-$	→	$Mn^{2+} + 4H_2O$	1.51
Au^{3+}	+	$3e^-$	→	Au	1.50
Cl_2	+	$2e^-$	→	$2Cl^-$	1.36
O_2	+	$4H^+ + 4e^-$	→	$2H_2O$	1.23
$Cr_2O_7^{2-}$	+	$14H^+ + 6e^-$	→	$2Cr^{3+} + 7H_2O$	1.23
Br_2	+	$2e^-$	→	$2Br^-$	1.07
NO_3^-	+	$4H^+ + 3e^-$	→	$NO + 2H_2O$	0.96
Ag^+	+	e^-	→	Ag	0.80
I_2	+	$2e^-$	→	$2I^-$	0.54
Cu^+	+	e^-	→	Cu	0.52
O_2	+	$2H_2O + 4e^-$	→	$4OH^-$	0.40
Cu^{2+}	+	$2e^-$	→	Cu	0.34
$2H_3O^+$	+	$2e^-$	→	$H_2 + 2H_2O$	0.00
Pb^{2+}	+	$2e^-$	→	Pb	–0.13
Sn^{2+}	+	$2e^-$	→	Sn	–0.14
Ni^{2+}	+	$2e^-$	→	Ni	–0.26
Fe^{2+}	+	$2e^-$	→	Fe	–0.45
Cr^{3+}	+	$3e^-$	→	Cr	–0.74
Zn^{2+}	+	$2e^-$	→	Zn	–0.76
$2H_2O$	+	$2e^-$	→	$H_2 + 2OH^-$	–0.83
Mn^{2+}	+	$2e^-$	→	Mn	–1.19
Al^{3+}	+	$3e^-$	→	Al	–1.66
Mg^{2+}	+	$2e^-$	→	Mg	–2.37
Na^+	+	e^-	→	Na	–2.71
Ca^{2+}	+	$2e^-$	→	Ca	–2.87
Ba^{2+}	+	$2e^-$	→	Ba	–2.91
K^+	+	e^-	→	K	–2.93
Li^+	+	e^-	→	Li	–3.04

Increasing Strength of Oxidizing Agent

Increasing Strength of Reducing Agent

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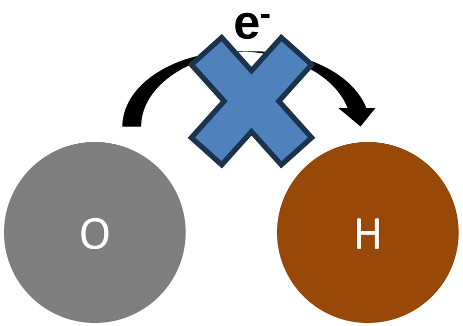

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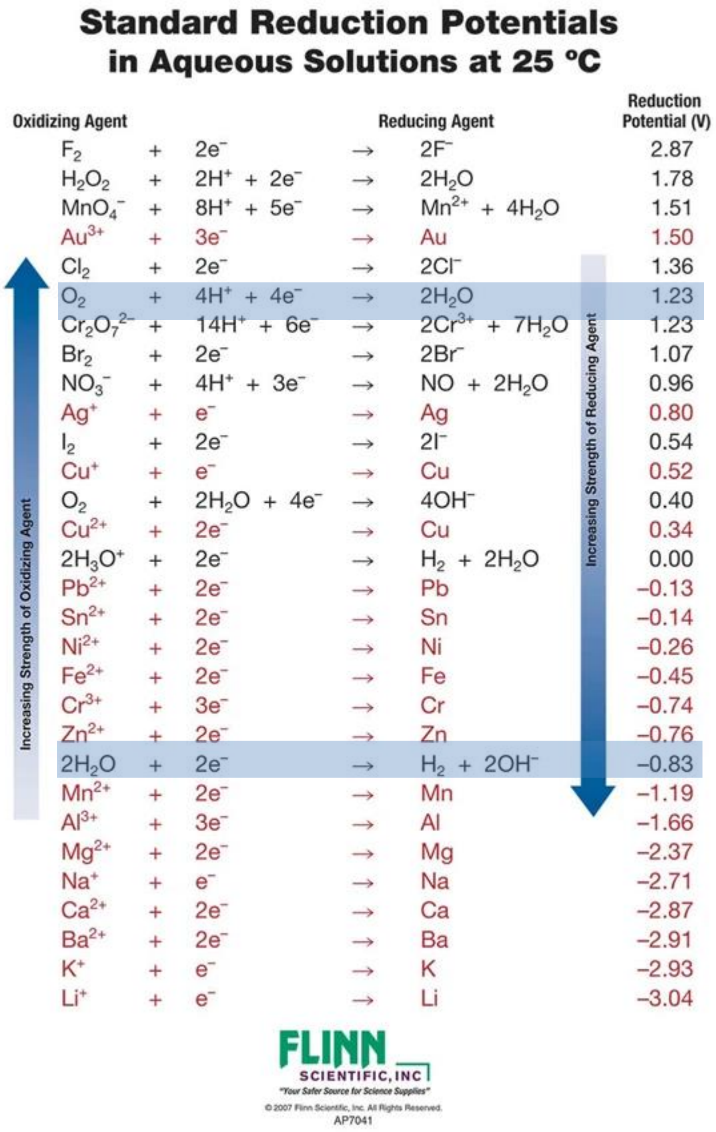
It is **NOT spontaneous**.



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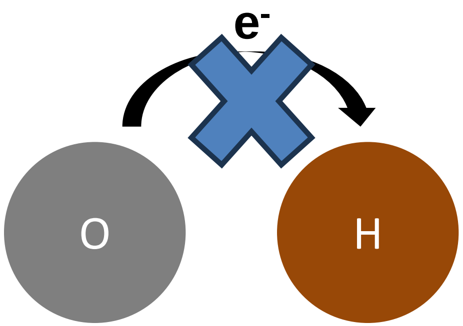

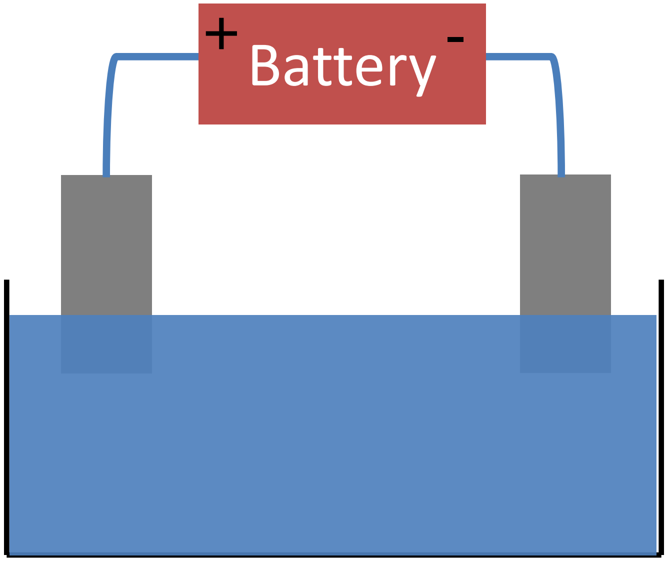

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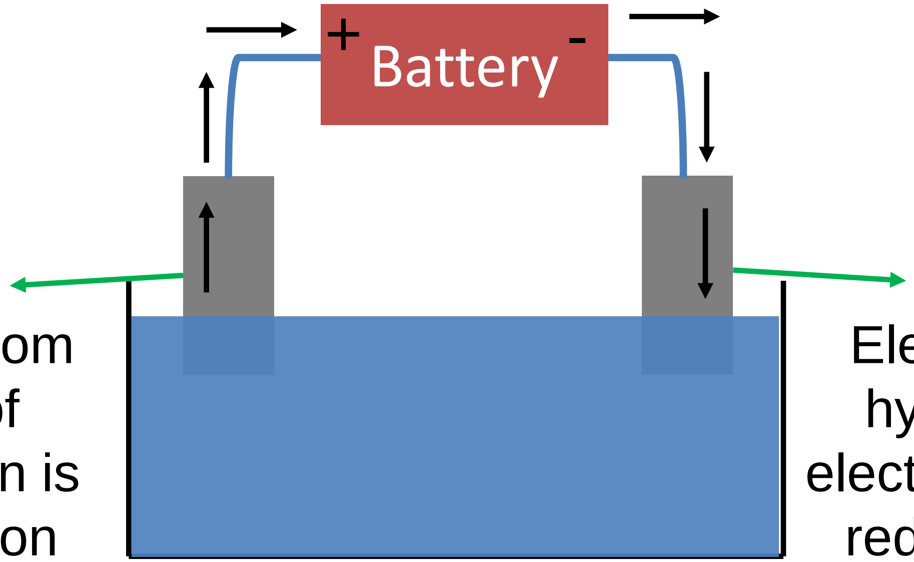


Anode
Electrons pulled from oxygen - Loss of electrons – Oxygen is oxidized - Oxidation

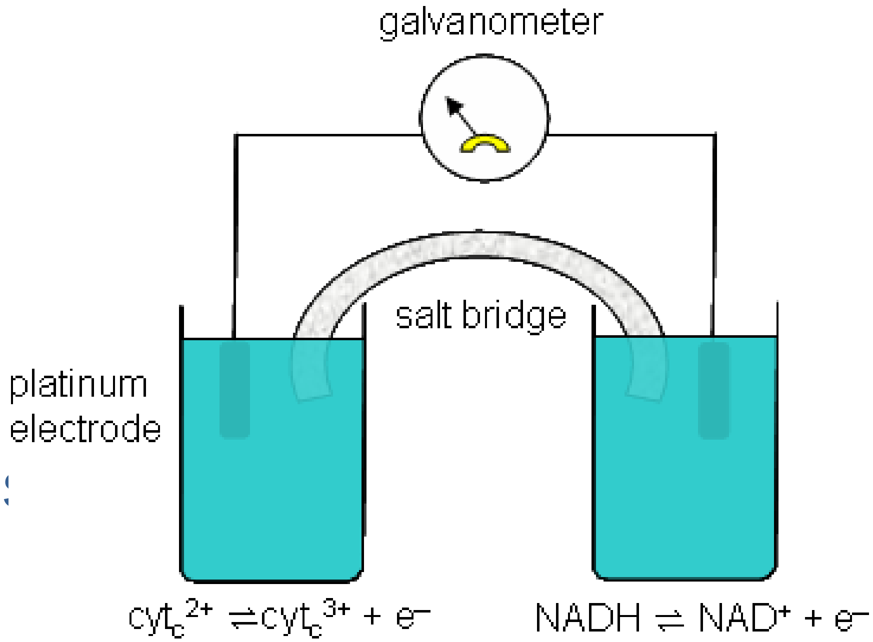
Cathode
Electrons pushed to hydrogen - Gain of electrons – Hydrogen is reduced - Reduction

Anode where **Oxidation** happens - Loss of electrons (**An Ox**)
Cathode where **Reduction** - Gain of electrons (**Red Cat**)

FUNDAMENTALS OF ELECTROCHEMISTRY

PART I :

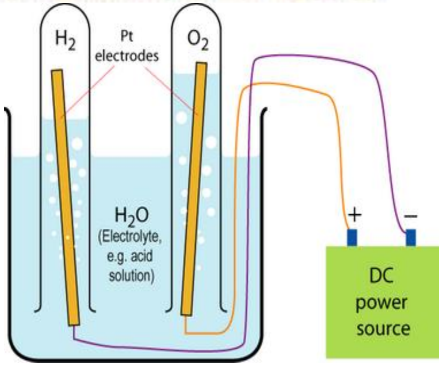
- Terminology
- History
- Electrochemical Cells
- Standard Half Cell Potentials
- Standard Cell Potentials

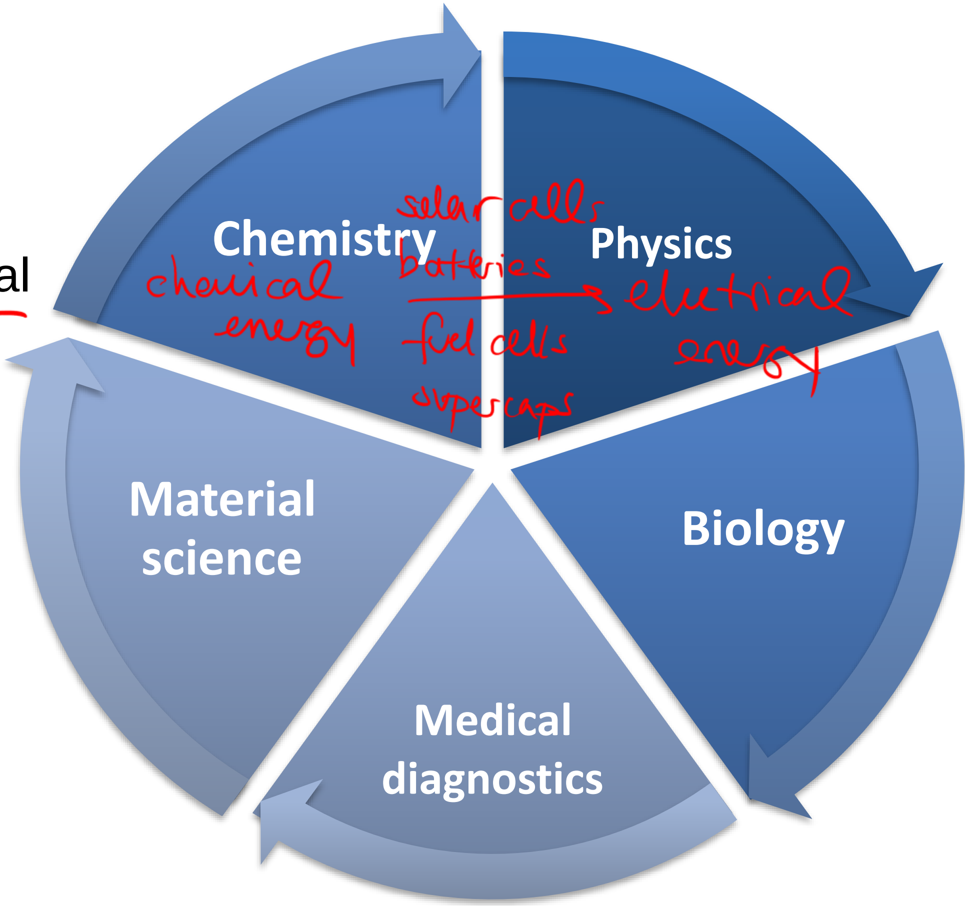


Electrochemistry

Electrochemistry

- electron transfer process
- properties of electrolytes
- processes occurring at electrodes
- interchange of chemical and electrical energy





Chemistry

Physics

Biology

Medical diagnostics

Material science

solar cells

batteries

chemical energy → electrical energy

fuel cells

supercaps

Peter Atkins " I think we ought to expand our view of electrochemistry is about. It is not just about electrode potential and electrolysis, it is half of modern inorganic chemistry "

History

- In 1600, **Sir William Gilbert** coined the word *electric* from Greek word for amber (electron) to describe the substances that acquired the power to attract light objects when rubbed.

- In 1663, **Otto von Guericke** created the first electric generator, which produced static electricity by applying friction in the machine.

- In 1747, **Benjamin Franklin** proposed ‘’ one-fluid theory of electricity’’.

- In 1781, **Charles-Augustin de Coulomb** developed the law of electrostatic attraction

- In the late 1700s, **Luigi Galvani** marked the birth of electrochemistry by establishing a bridge between chemical reactions and electricity on his essay *"Commentary on the Effect of Electricity on Muscular Motions"* where he proposed *animal electricity*.

- In 1800s, **Alessandro Volta** referred the electricity generated as "metallic electricity" and decided that the muscle, by contracting when touched by metal, resembled the action of an electroscope. He invented the first practical battery – the Voltaic Pile.

- In 1800, **William Nicholson** and **Johann Ritter** decomposed water by electrolysis.

History

- In 1827, **Georg Ohm** expressed his law in this famous book *"The Galvanic Circuit Investigated Mathematically "* in which he gave his complete theory of electricity

- In 1832, **Michael Faraday**'s experiments on electrochemistry led him to state his two laws of electrolysis.
  - (1) the amount of chemical change produced by current at an electrode-electrolyte boundary is proportional to the quantity of electricity used and (2) the amounts of chemical changes produced by the same quantity of electricity in different substances are proportional to their equivalent weights.
- In 1839, **Sir William Grove** produced the first fuel cell.

- In 1894, **Friedrich Ostwald** concluded important studies of the electrical conductivity and electrolytic dissociation of organic acids.

- In 1888, **Hermann Nernst** developed the theory of the electromotive force of the voltaic cell

- In 1909, **Robert Andrews Millikan** began a series of experiments to determine the electric charge carried by a single electron.

- In 1923, **Johannes Nicolaus Brønsted** and **Thomas Martin Lowry** published about how acids and bases behave, using an electrochemical basis.

‘ *Galvani ! Stop playing the cell and please eat the frog legs*’.

Electrical Charge and Current

Electrical charge , symbolized *q*, is the basic property of matter carried by some elementary particles.

Metal	Electrolyte	Plasma
• Valence electrons	• Ions	• Electrons &Cations

Electrical current is the rate of charge flow passing through a given point in an electric circuit, measured in **Coulombs/second**, named **Amperes**. Current can be introduced through electrodes.

Electrical Units

- Coulomb (C)
  - Charge on 6.25 x 10^{18} electrons
  - Units: charge **q** in coulombs (**C**), current **I** in amps (**A**), time **t** in seconds (s), **q=It**

- Amperes (A)
  - Electric current,   A=1C/sec

- Volt (V)
  - Potential driving current flow, V= 1 J/C
- Ohm's law
  - **ε = IR**, **ε** = potential, **I** =current, and **R**=resistance

Electrical Units

	symbol	*unit*	*relationships*
Charge	**q**	**Coulomb (C)**	**$q = I \times t$ (A.s)**
Current	**I**	**Ampere (A)**	**$I = q/t$ (t in s)**
Potential	**ε**	**Volt (V)**	**$\varepsilon = I \times R$**
Power	**P**	**Watt (W)**	**$P = \varepsilon \times I$**
Energy	**E**	**Joule (J)**	**$Pt = \varepsilon \times I \times t = \varepsilon \times q$**
Resistance	**R**	**Ohm (Ω)**	**$R = \varepsilon/I$**

Question: A current of 15A flows through a fuse for 0.05s before it melts. How much charge will have passed through it?

q = I × t (Unit in Coulombs)

q = 15 A × 0.05 s

q = 7.5 C

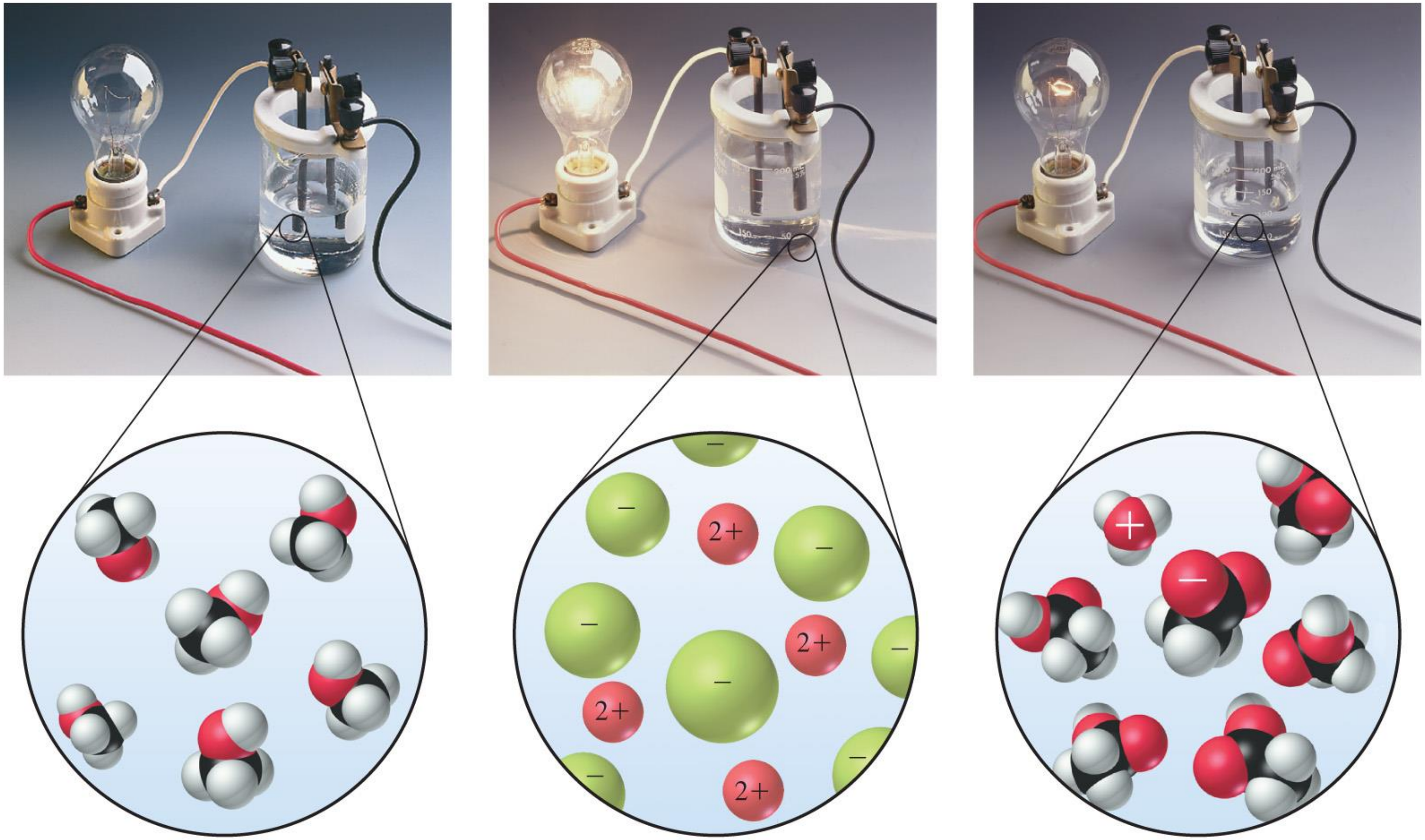
Question: What is the current in a resistor when 0.56C flows through the resistor in 800s?

I = q/t

I = 0.56 C / 800 s

q = 7E-4 A

Electrolytes

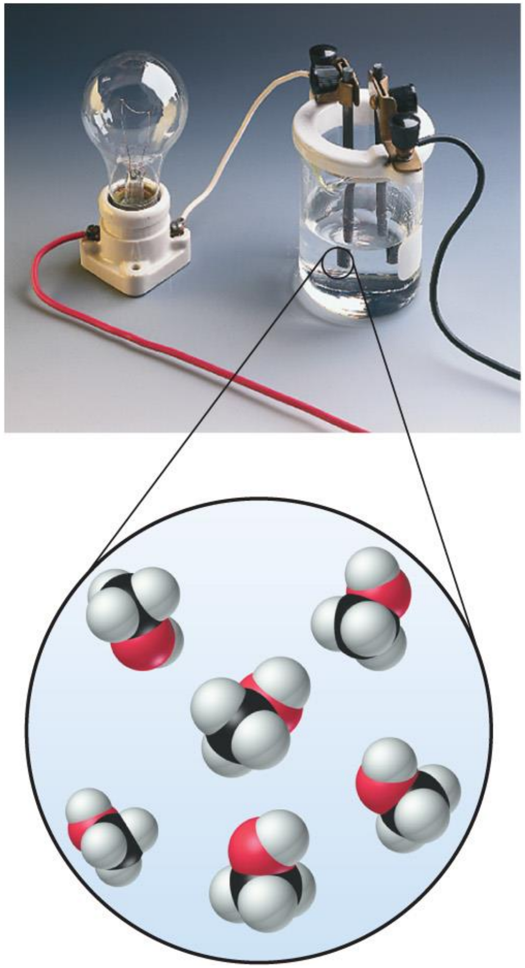


Non-electrolyte

Strong electrolyte

Weak electrolyte

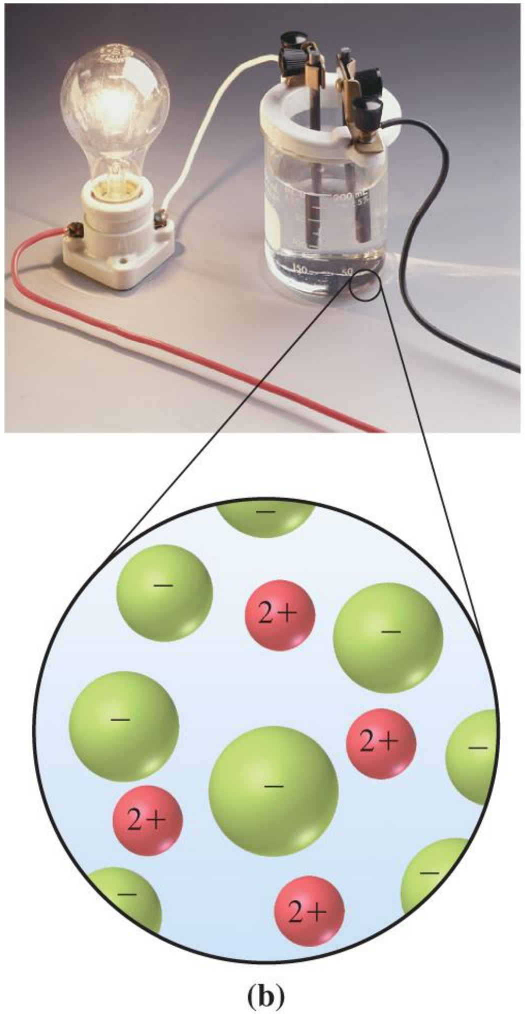
Electrolytes



Non-electrolyte

Sugar or ethanol dissolve in water, BUT they do not produce **ions**

Electrolytes

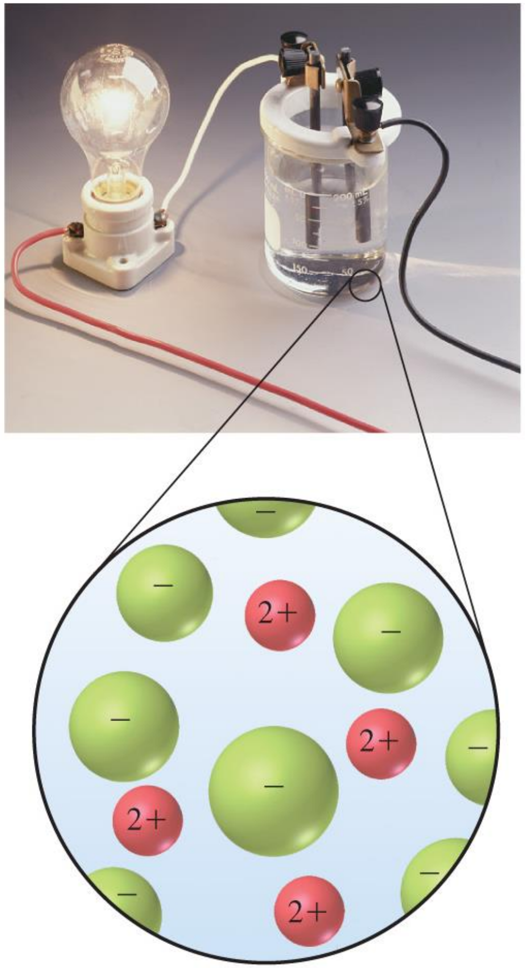


Strong electrolyte

Electrolytes are ionic compounds which dissociate into ions when dissolved in a solvent and they conduct a current (conductivity).

Even insoluble ionic compounds such as $CaCO_3$ are electrolytes because they can conduct a current in the molten (melted) state.

Electrolytes

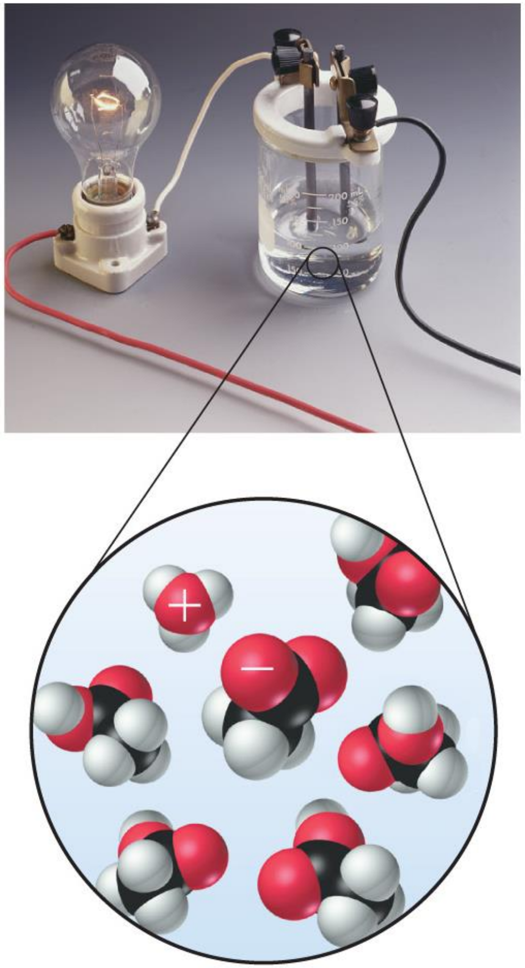


(b)

Strong electrolyte

strong acids	HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$, and H_2SO_4
strong bases	NaOH, KOH, LiOH, $Ba(OH)_2$, and $Ca(OH)_2$
salts	NaCl, KBr, $MgCl_2$

Electrolytes



(c)

Weak electrolyte

weak acids	HF, $HC_2H_3O_2$ (acetic acid), H_2CO_3 (carbonic acid), H_3PO_4 (phosphoric acid)
weak bases	NH_3 (ammonia), C_5H_5N (pyridine), and several more, all containing "N"

Oxidation-Reduction Reactions

Involve the transfer of electrons between different reactants

- Oxidation: formation of oxides
- Reduction: removal of oxygen from the oxide

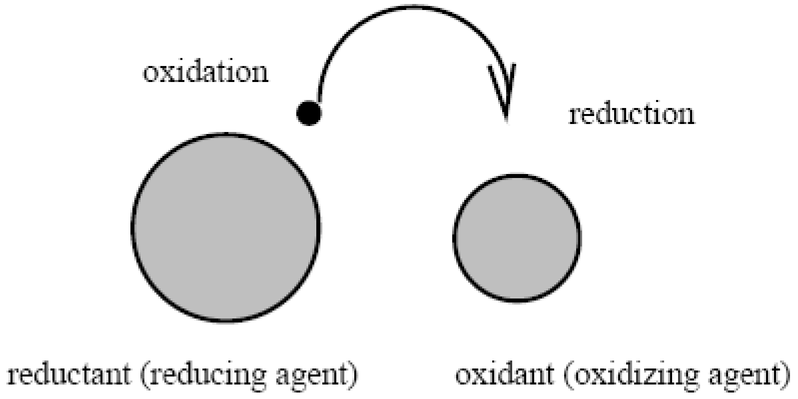
Oxidation - Loss of Electrons
Reduction - Gain of Electrons

Oxidant - Reactant which oxidizes (gets reduced)

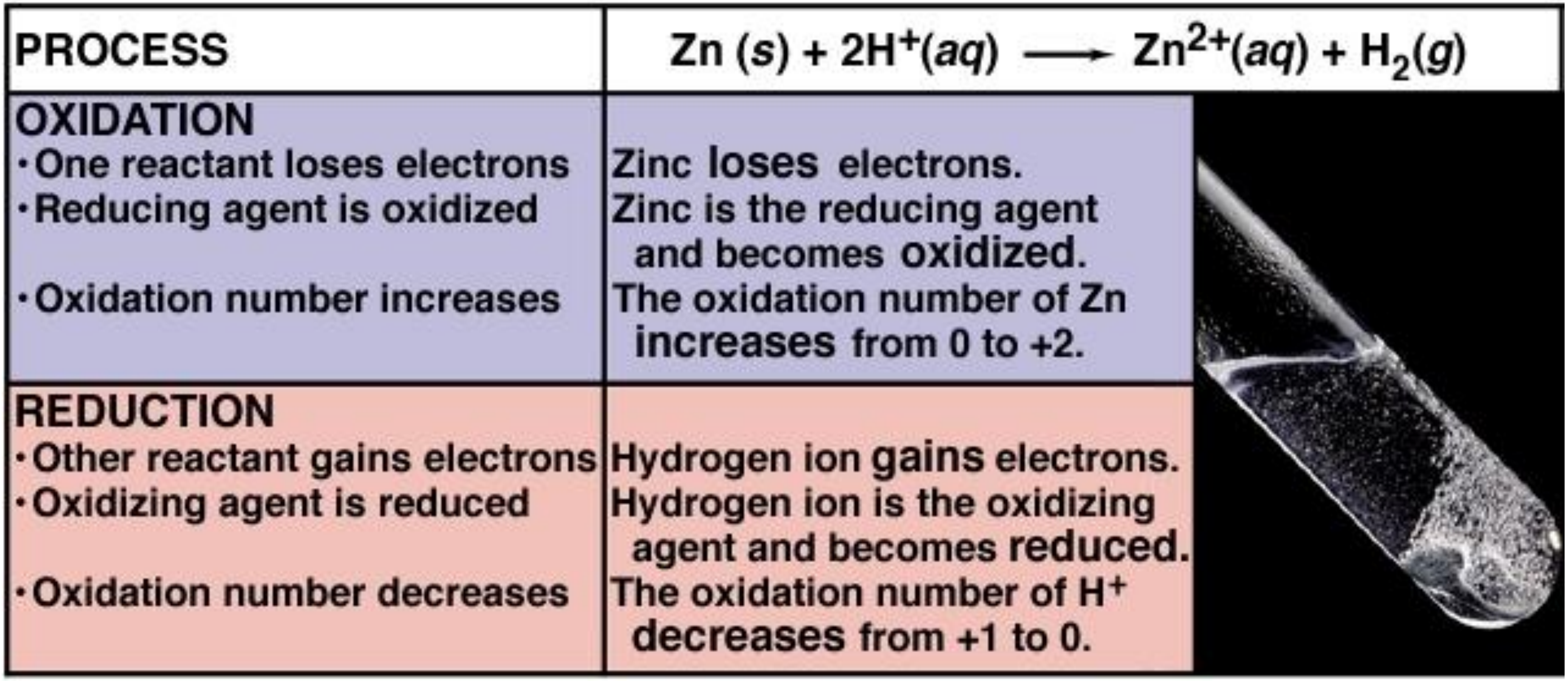
Reductant - Reactant which reduces (gets oxidized)

$$\mathbf{4Fe + 3O_2 \longrightarrow 2Fe_2O_3}$$

$$Fe \longrightarrow Fe^{+3} + 3e^-$$
$$O_2 + 4e^- \longrightarrow 2O^{-2}$$

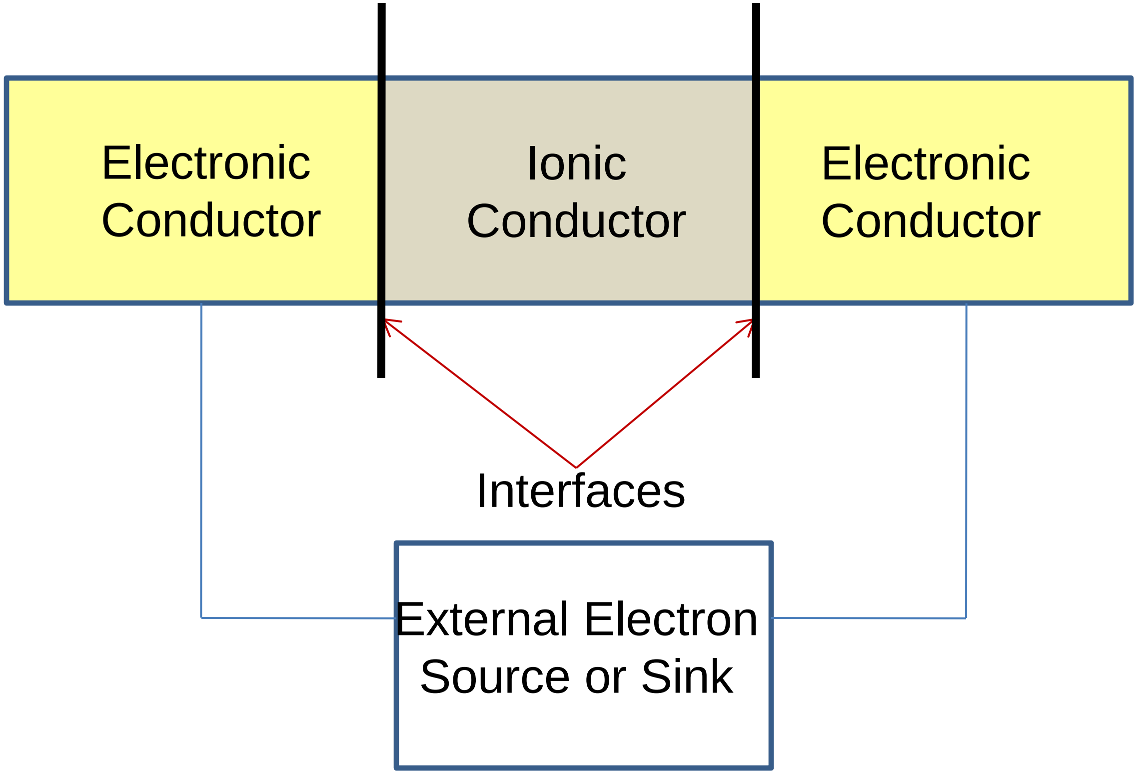


Oxidation-Reduction Reactions



PROCESS	$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$
OXIDATION • One reactant loses electrons • Reducing agent is oxidized • Oxidation number increases	Zinc **loses** electrons. Zinc is the reducing agent and becomes **oxidized.** The oxidation number of Zn **increases** from 0 to +2.
REDUCTION • Other reactant gains electrons • Oxidizing agent is reduced • Oxidation number decreases	Hydrogen ion **gains** electrons. Hydrogen ion is the oxidizing agent and becomes **reduced.** The oxidation number of H^+ **decreases** from +1 to 0.

Electrochemical System



An electrochemical system is not *homogeneous but is heterogeneous*

Electrochemical Cell

- An electrochemical cell is an array consisting of two electrodes, each of which is in contact with an electrolyte solution.

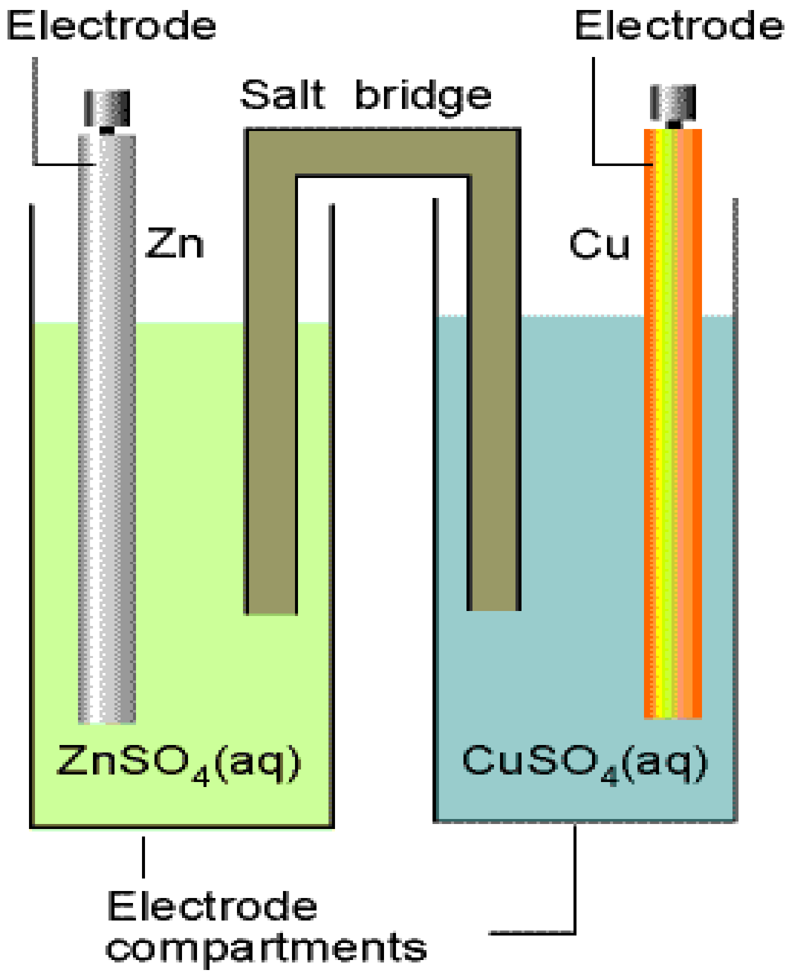
- Electrode is a conductor at the surface of which electron transfer to or from the surrounding solution takes place.

A cathode is an electrode where reduction occurs.

An anode is an electrode where oxidation takes place.

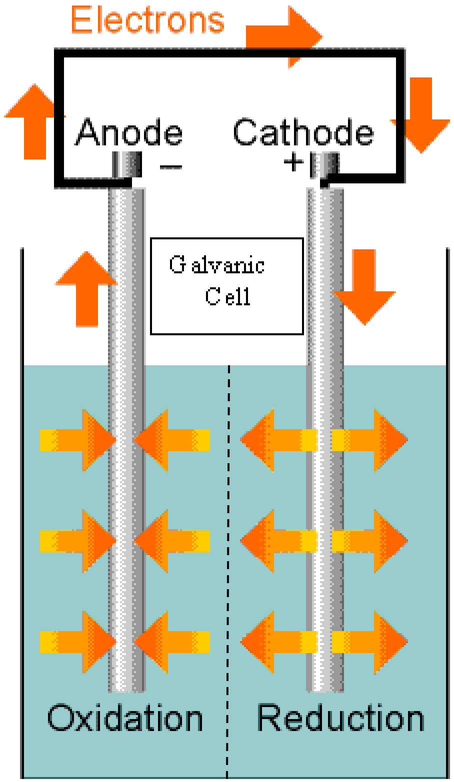
- Electrons carry the charge within the electrodes and the ext. conductor

- Anions and cations are the charge carriers within the cell

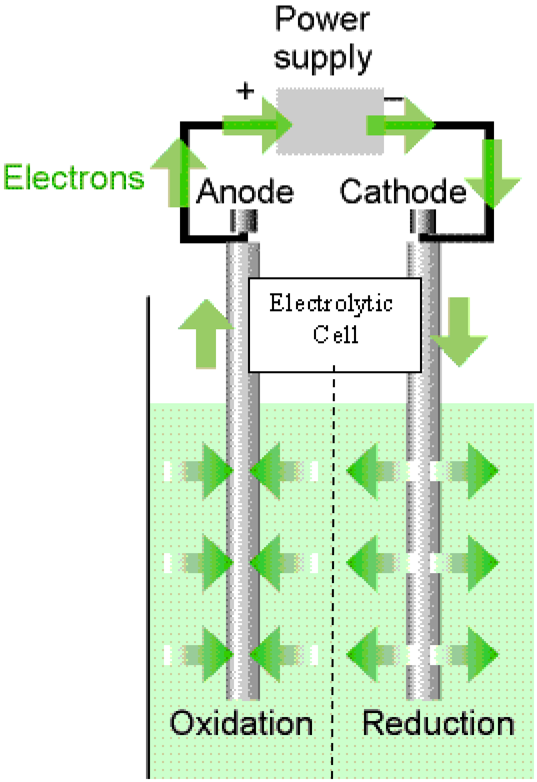


Electron Flow → always from anode to cathode (through the wire)

Electrochemical Cell



Galvanic Cell -- an electrochemical cell in which a spontaneous reaction generates electricity.



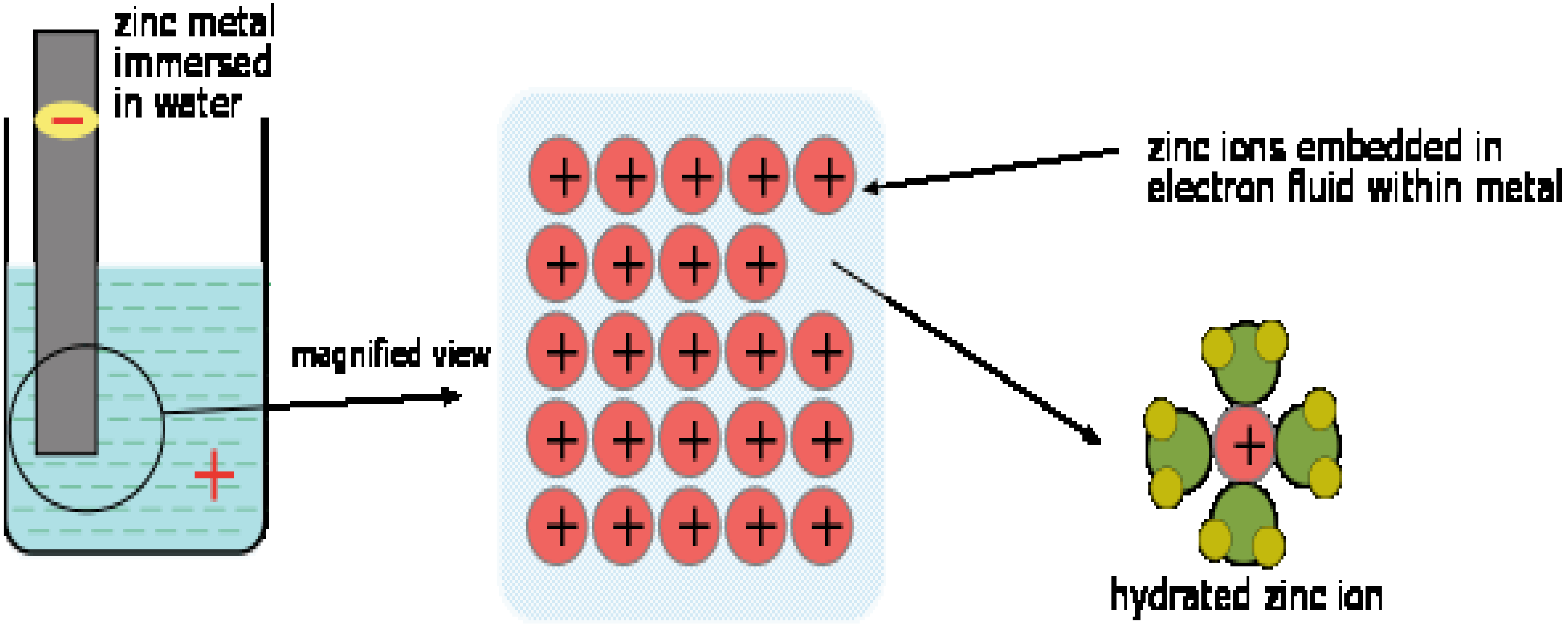
Electrolytic Cell -- an electrochemical cell in which electricity (an external voltage) is used to drive a nonspontaneous reaction.

Electroneutrality

electroneutrality principle states all pure substances carry a net charge of zero (electrolyte solutions have essentially the same number of cations and anions regardless of the value of the electrical potential)

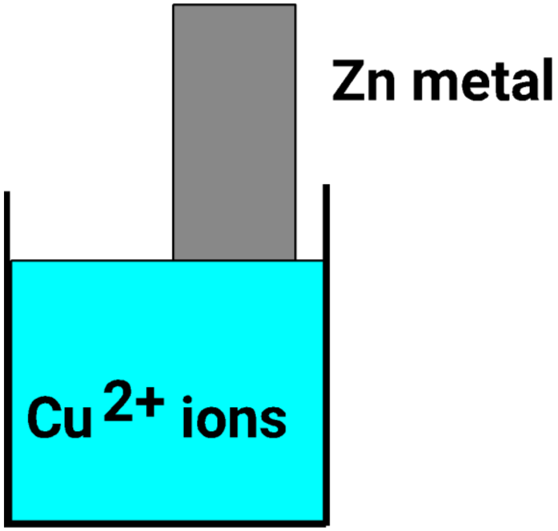
In any solution:

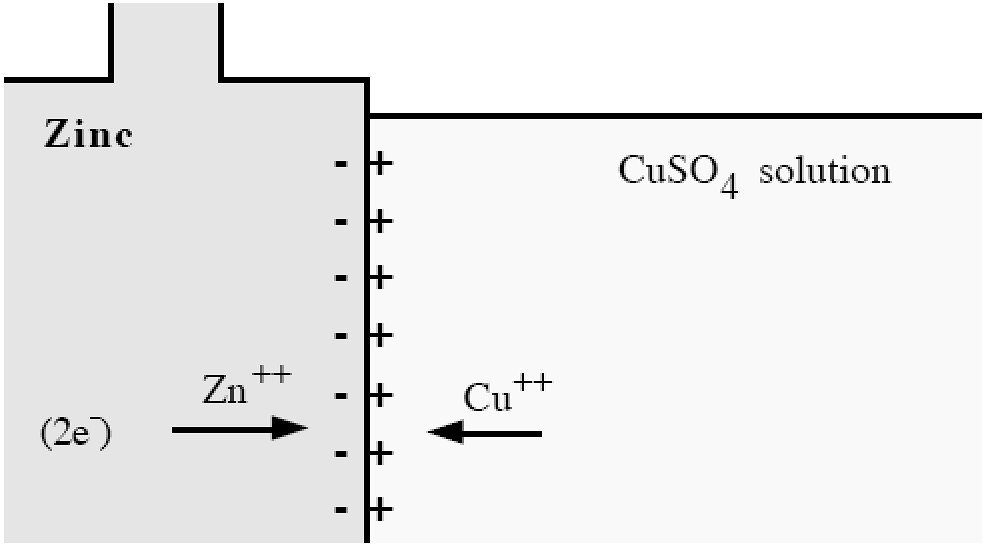
Number of + charges = Number of - charges



Oxidation-Reduction

The reaction is a simple oxidation-reduction process, a transfer of two electrons from the zinc to the copper



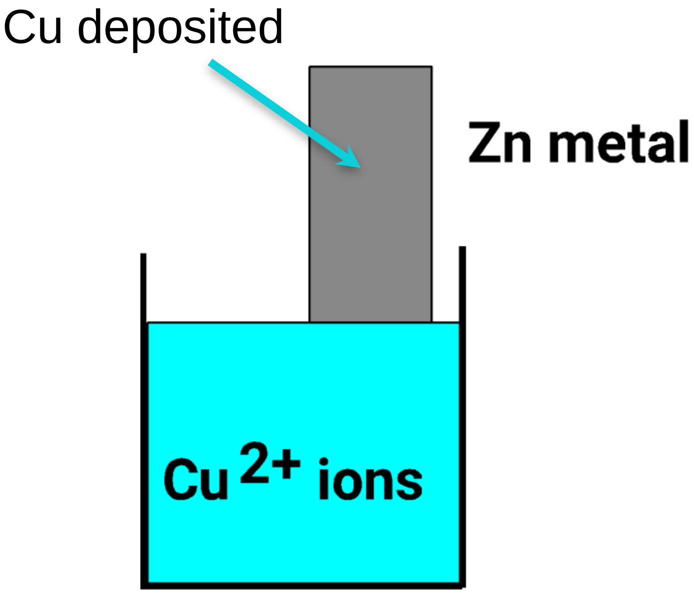


$Zn(s) \longrightarrow Zn^{2+} + 2e^{-}$

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

$$Zn(s) + Cu^{2+} \longrightarrow Zn^{2+} + Cu(s)$$

Oxidation-Reduction



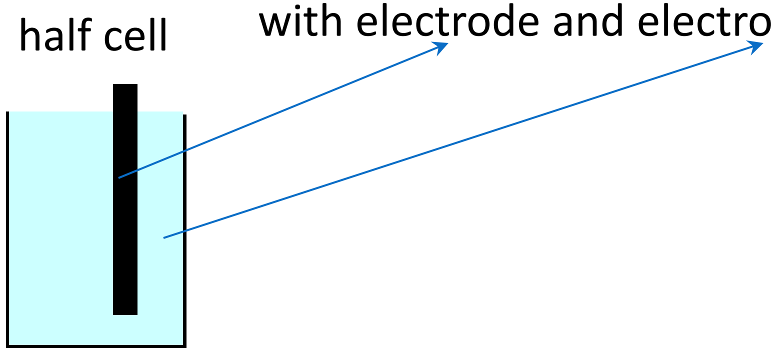
With time, Cu plates out onto Zn metal strip, and Zn strip “disappears.”

- Zn is oxidized and is the reducing agent
  $Zn(s) \rightarrow Zn^{2+}(aq) + 2e-$

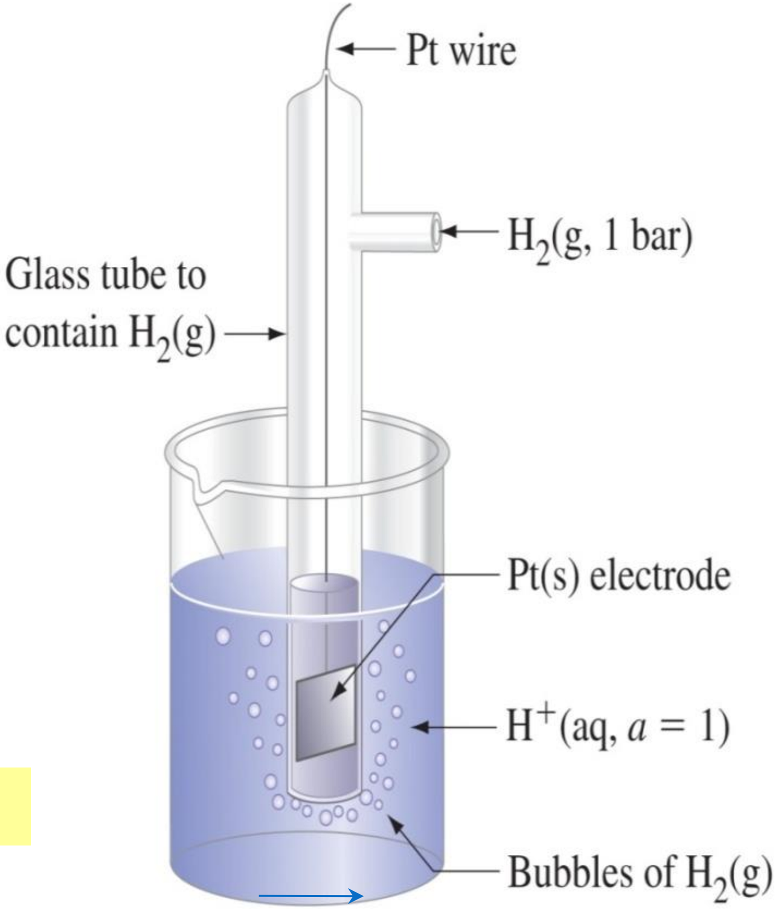
- Cu^{2+} is reduced and is the oxidizing agent
  $Cu^{2+}(aq) + 2e- \rightarrow Cu(s)$

Standard Electrode Potentials

electrode is referred to the entire half cell with electrode and electrolyte



The potential of an individual electrode is difficult to establish



$$2\ H^+(a = 1) + 2\ e^- \longrightarrow H_2(g,\ 1\ bar) \qquad E° = 0\ V$$

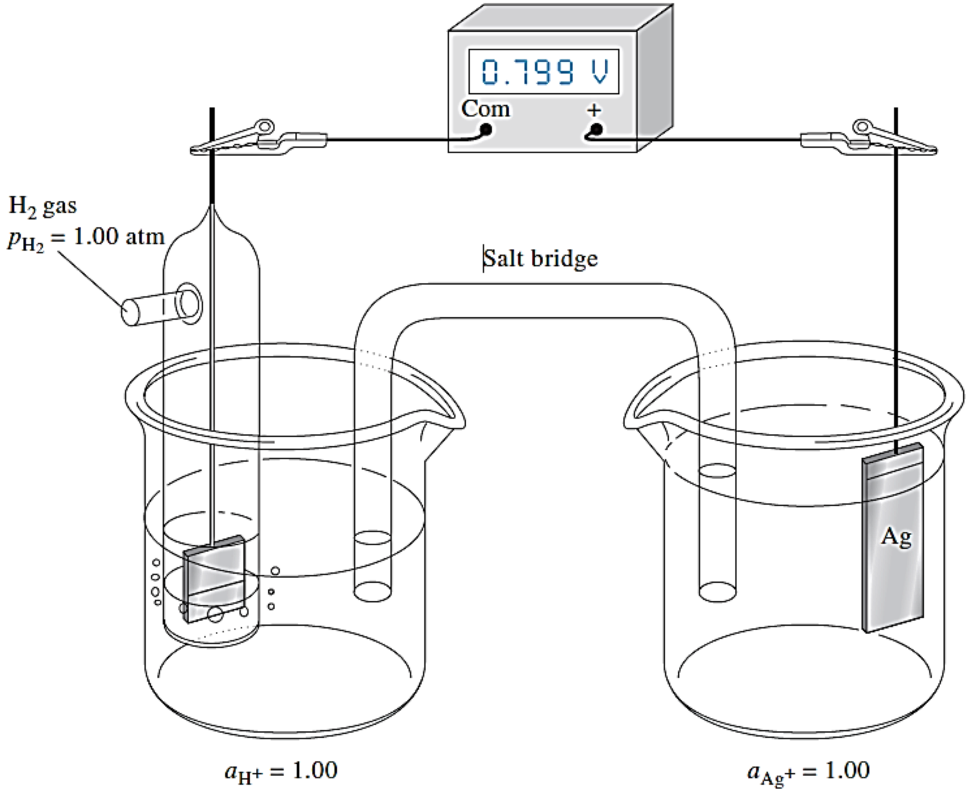
$Pt|H_2(g, 1\ bar)|\ H^+(a = 1)$

$$2H^+ + 2e^- \rightleftharpoons H_2(aq)$$

$$H_2(aq) \rightleftharpoons H_2(g)$$

Standard Electrode Potentials

The standard potential, Eº, of another couple is assigned by constructing a cell in which it is the right-hand electrode and the standard hydrogen electrode is the left hand electrode.



Standard electrode potential, "E°",

- *E°* defined by international agreement. **The tendency for a *reduction* process to occur at an electrode.**
- All ionic species present at a=1
- All gases are at 1 bar (approximately 1 atm).
- At the temperature of 298.15 K (25 °C).

Pt, H_2(*p* =1.00 atm) | H^+(a_{H+}=1.00) || Ag^+ (a_{Ag+}=1.00) | Ag

SHE || Ag^+ (a_{Ag+}=1.00) | Ag

Standard Electrode Potentials

Electrode potentials are always defined for the reduction direction.

Whitten/Davis/Peck, General Chemistry, 6/e
Table 21.2

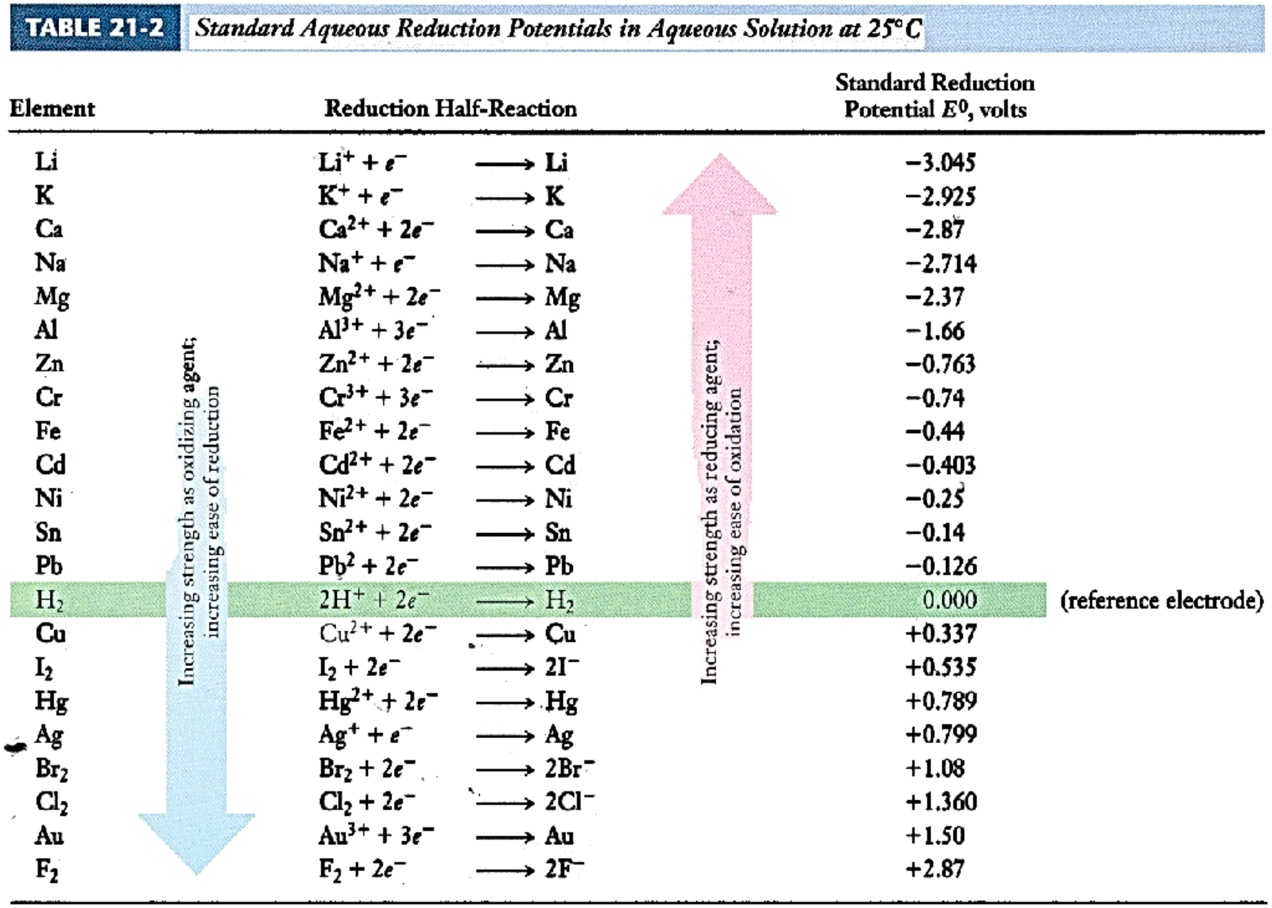


TABLE 21-2 *Standard Aqueous Reduction Potentials in Aqueous Solution at 25°C*

Element	Reduction Half-Reaction	Standard Reduction Potential E^0, volts	
Li	$Li^+ + e^- \longrightarrow Li$	−3.045	
K	$K^+ + e^- \longrightarrow K$	−2.925	
Ca	$Ca^{2+} + 2e^- \longrightarrow Ca$	−2.87	
Na	$Na^+ + e^- \longrightarrow Na$	−2.714	
Mg	$Mg^{2+} + 2e^- \longrightarrow Mg$	−2.37	
Al	$Al^{3+} + 3e^- \longrightarrow Al$	−1.66	
Zn	$Zn^{2+} + 2e^- \longrightarrow Zn$	−0.763	
Cr	$Cr^{3+} + 3e^- \longrightarrow Cr$	−0.74	
Fe	$Fe^{2+} + 2e^- \longrightarrow Fe$	−0.44	
Cd	$Cd^{2+} + 2e^- \longrightarrow Cd$	−0.403	
Ni	$Ni^{2+} + 2e^- \longrightarrow Ni$	−0.25	
Sn	$Sn^{2+} + 2e^- \longrightarrow Sn$	−0.14	
Pb	$Pb^{2} + 2e^- \longrightarrow Pb$	−0.126	
H_2	$2H^+ + 2e^- \longrightarrow H_2$	0.000	(reference electrode)
Cu	$Cu^{2+} + 2e^- \longrightarrow Cu$	+0.337	
I_2	$I_2 + 2e^- \longrightarrow 2I^-$	+0.535	
Hg	$Hg^{2+} + 2e^- \longrightarrow Hg$	+0.789	
Ag	$Ag^+ + e^- \longrightarrow Ag$	+0.799	
Br_2	$Br_2 + 2e^- \longrightarrow 2Br^-$	+1.08	
Cl_2	$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.360	
Au	$Au^{3+} + 3e^- \longrightarrow Au$	+1.50	
F_2	$F_2 + 2e^- \longrightarrow 2F^-$	+2.87	

Increasing strength as oxidizing agent; increasing ease of reduction

Increasing strength as reducing agent; increasing ease of oxidation

Harcourt, Inc.

Standard Electrode Potentials

- Which is the best oxidizing agent:

  O_2, H_2O_2, or Cl_2 ?

H_2O_2 /H_2O	+1.77
Cl_2 /Cl^-	+1.36
O_2 /H_2O	+1.23

This order is based on their standard reduction potentials, where more positive values correspond to stronger oxidising agents.

- Which is the best reducing agent:

  Sn, Hg, or Al ?

Al^{3+} /Al	-1.66
Sn^{2+} /Sn	-0.14
Hg^{2+} /Hg	+0.86

This order is based on their standard reduction potentials, where more negative values correspond to stronger reducing agents.

Standard Electrode Potentials

The following reactions are spontaneous and thus proceed to the right, as written:

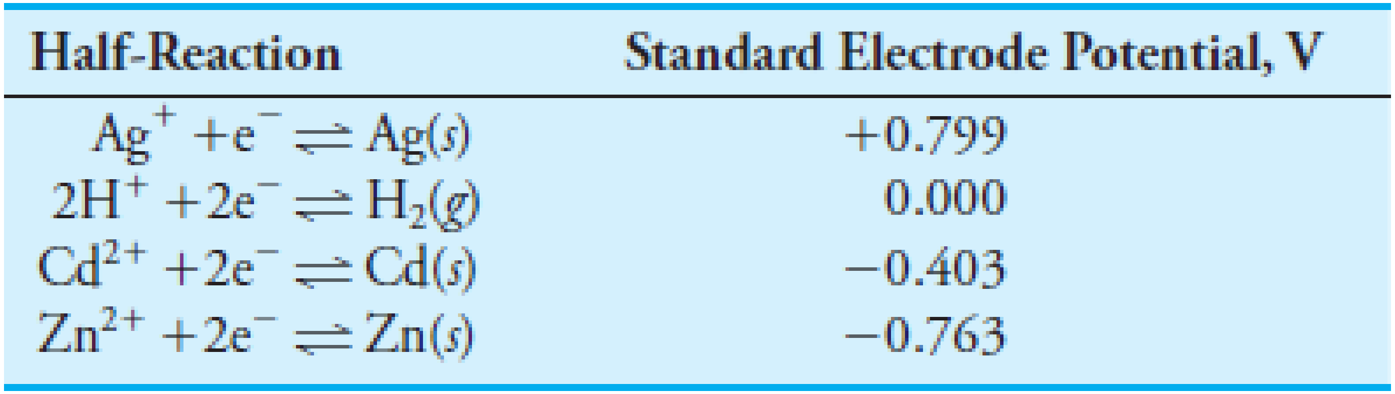
$$2H^+ + Cd(s) \rightleftharpoons H_2 + Cd^{2+}$$

$$2Ag^+ + H_2(g) \rightleftharpoons 2Ag(s) + 2H^+$$

$$Cd^{2+} + Zn(s) \rightleftharpoons Cd(s) + Zn^{2+}$$

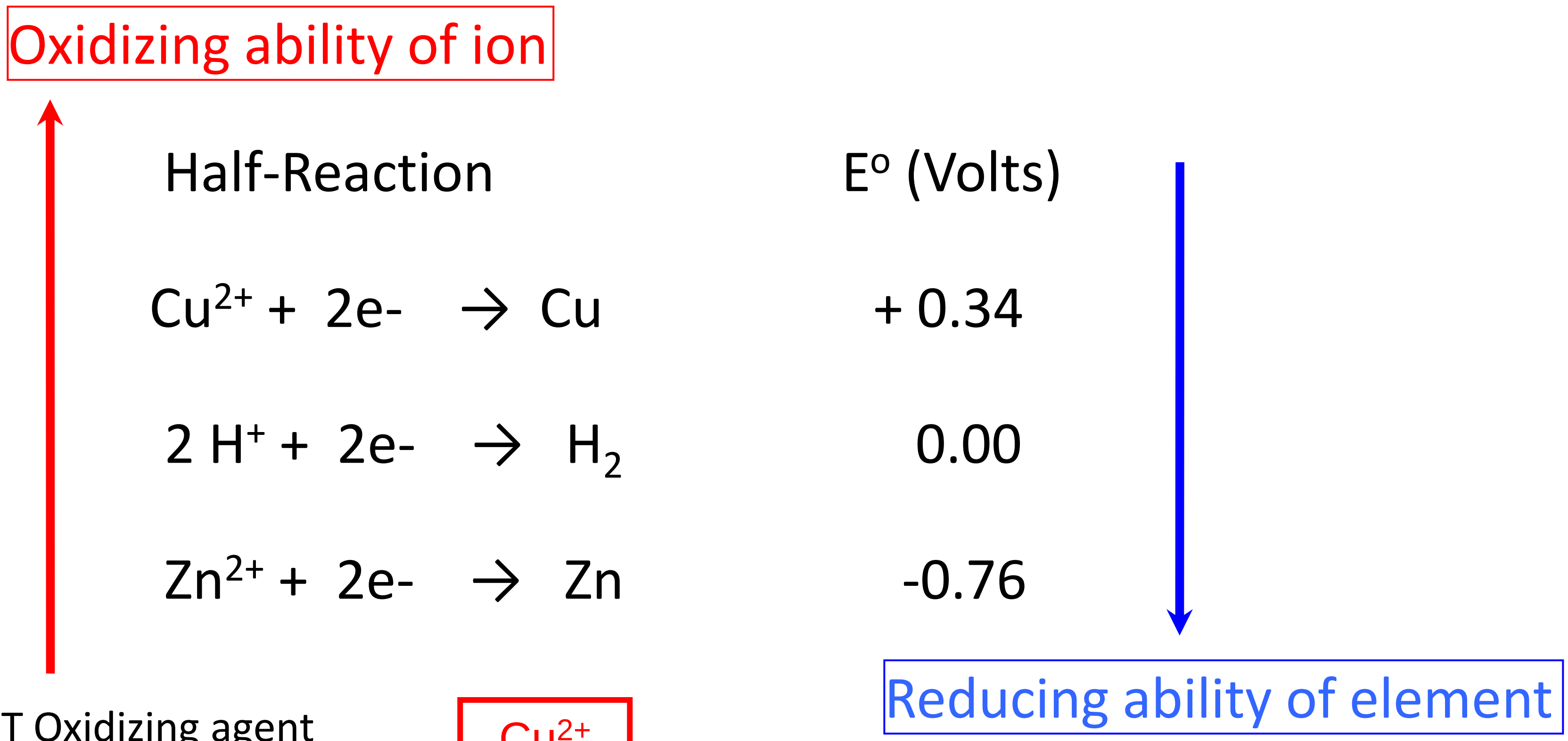
Order of oxidizing strength???

The second reaction establishes that Ag^+ is a more effective electron acceptor than H^+; the first reaction demonstrates that H+ is more effective than Cd^{2+}. Finally, the third equation shows that Cd^{2+} is more effective than Zn^{2+}. Thus, the order of oxidizing strength is **$Ag^+ > H^+ > Cd^{2+} > Zn^{2+}$**



Half-Reaction	Standard Electrode Potential, V
$Ag^+ + e^- \rightleftharpoons Ag(s)$	+0.799
$2H^+ + 2e^- \rightleftharpoons H_2(g)$	0.000
$Cd^{2+} + 2e^- \rightleftharpoons Cd(s)$	−0.403
$Zn^{2+} + 2e^- \rightleftharpoons Zn(s)$	−0.763

Standard Electrode Potentials



Oxidizing ability of ion

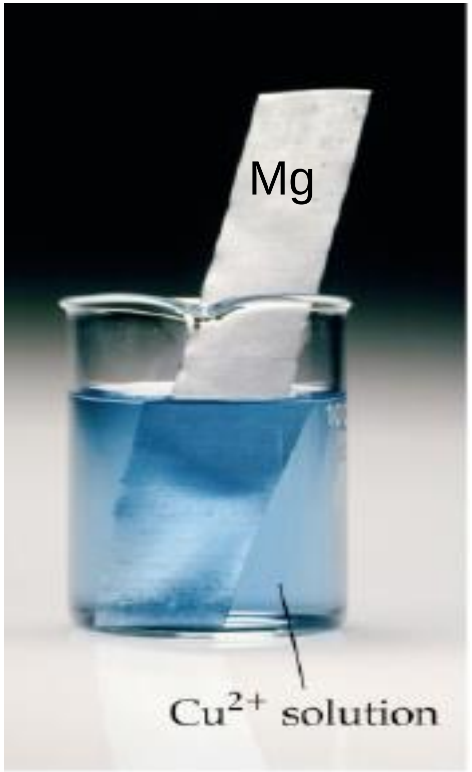
Half-Reaction	E° (Volts)
$Cu^{2+} + 2e^- \rightarrow Cu$	+ 0.34
$2 H^+ + 2e^- \rightarrow H_2$	0.00
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76

Reducing ability of element

BEST Oxidizing agent Cu^{2+}

BEST Reducing agent Zn

Standard Electrode Potentials



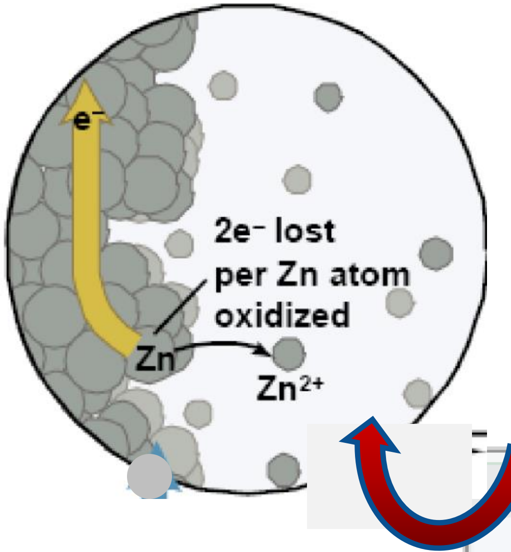
Mg$^{2+/}$Mg(s)	-2.363 V
Cu^{+2}/ Cu(s)	0.337 V

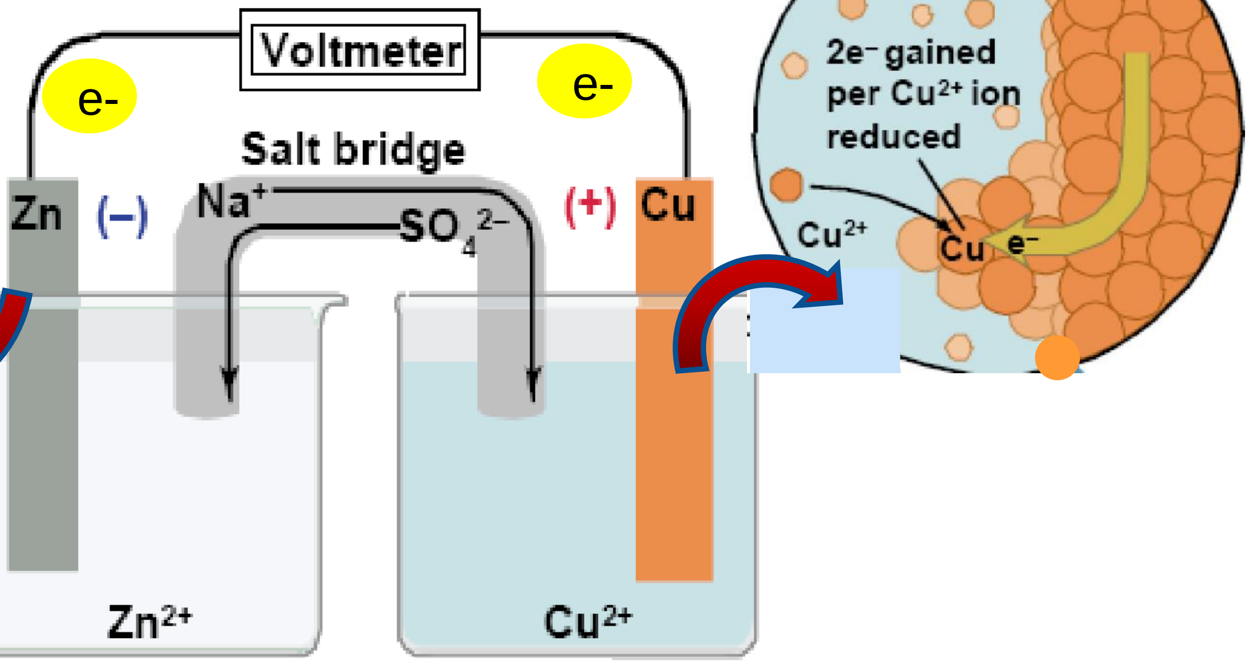
Copper plates on Magnesium

Magnesium dissolves as Mg^{2+}

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$$

Electrochemical Cell





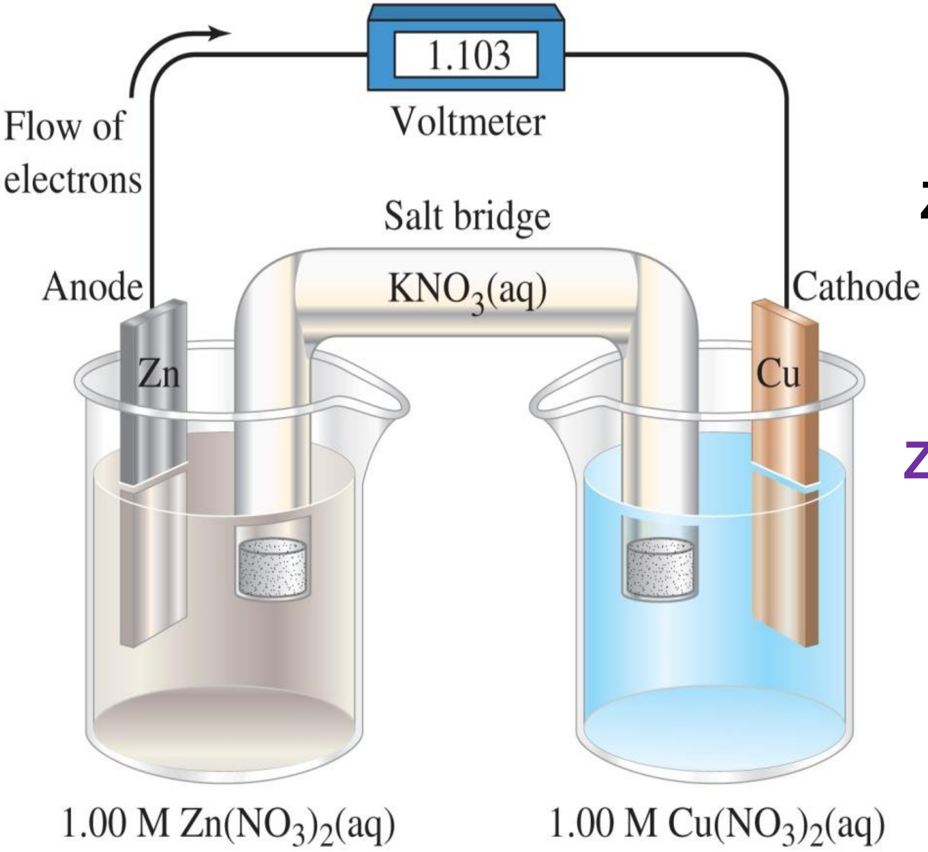
Oxidation half-reaction

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

Reduction half-reaction

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Cell Description Conventions



$$\mathrm{Zn(s)\,|\,Zn^{2+}(aq)\;\;||\;\;Cu^{2+}(aq)\,|\,Cu(s)}$$

$$\mathrm{Zn\,|\,Zn^{2+}(x\ M)\,||\,Cu^{2+}(y\ M)\,|\,Cu}$$

$$\mathrm{Zn\,|\,Zn(NO_3)_2(x\ M)\,||\;\;Cu(NO_3)_2(y\ M)\,|\,Cu}$$

$$\underbrace{\mathrm{Zn(s)\,|\,Zn^{2+}(aq)}}_{\text{anode}}\;\;||\;\;\underbrace{\mathrm{Cu^{2+}(aq)\,|\,Cu(s)}}_{\text{cathode}}$$

Cell Description Conventions

$$Zn \mid Zn^{2+}(1M) \parallel Cu^{2+}(1M) \mid Cu$$

The components of the cell are written in **order**, from **left-hand** and moving across the salt bridge **to the right**. A single **vertical line** indicates a **phase boundary**, such as that between the solid Zn electrode and $Zn^{2+(}$aq). The **double vertical line** represents a **salt bridge**. Spectator ions, like SO_2^{-4}(aq) in the Zn–Cu cell, are usually omitted.

By convention, the **electrode** written to **the left of the salt bridge** is always taken to be **the anode**, and **the associated half-equation is always written as an oxidation**. The **right-hand electrode** is therefore always the **cathode**, and the **half-equation is always written as a reduction**.

$$Ag \mid Ag^{+} \parallel H^{+} \mid H_2 \mid Pt_{(s)}$$

$$2Ag_{(s)} + 2H^{+} \rightarrow 2Ag^{+} + H_{2\,(g)}$$

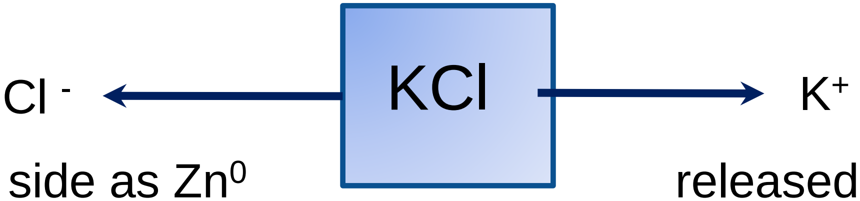
Salt Bridge

- Salt bridge exists to provide the electrical connection between the two reaction vessels while keeping the two reactions separate.

- Allows ion migration in solution but prevents extensive mixing of electrolytes.

- Can be a simple porous disk or a gel saturated with a non-interdering salt like KCl.

$$Zn(s) \mid Zn^{2+}(aq) \quad \| \quad Cu^{2+}(aq) \mid Cu(s)$$



released to Zn side as Zn^0 is converted to Zn^{2+}

released to Cu side as Cu^{+2} is converted to Cu^0

Standard Cell Potential

Standard cell potential (E^o): the potential difference of a cell formed from two *standard* electrodes.

$$E°_{cell} = E°_{cathode} - E°_{anode}$$

E^o is a quantitative measure of the tendency of reactants to proceed to products when all are in their standard states at 25 °C.