

These Notes are to **SUPPLEMENT** the Text, They do **NOT** Replace reading the Text Book Material. Additional material that is in the Text Book will be on your tests! To get the most information, **READ THE CHAPTER** prior to the Lecture, bring in **these lecture notes and make comments on these notes.** **These notes alone are NOT enough to pass any test!**

**NOTE THESE ARE DRAFT LECTURE NOTES!**

**Electron Transfer or Redox** – Oxidation-Reduction involve a transfer of electrons  
How to spot a Redox reactions, see if there is a “Pure” Element on one side of the Reaction Equation

**Oxidation:** Loss of electrons – or increase in oxidation number

**Reduction:** Gain of electrons – or decrease in oxidation number

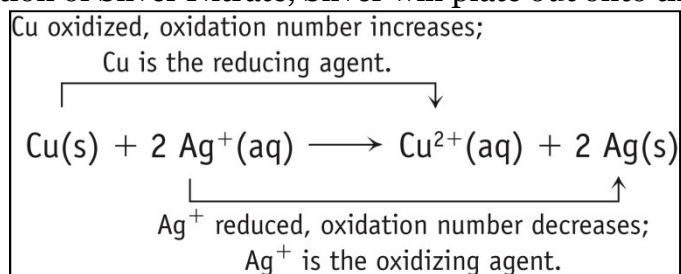
**Oxidation Agent:** Accepts electrons, this species is reduced, oxidation number decreases

**Reducing Agent:** Donates electrons, this species is oxidized, oxidation number increases

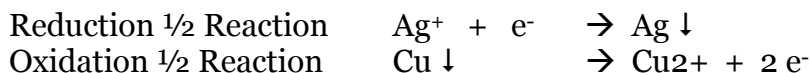
### Oxidation Number

1. A pure atom has an oxidation number of zero
2. Monatomic ions have an oxidation number equal to their ion charge
3. When combined with another element, Fluorine has an oxidation number of -1
4. Oxygen has an oxidation number of -2 in most compounds
  - a. Oxygen combined with Fluorine, oxygen has a positive oxidation number
  - b. Oxygen's oxidation number in peroxides, H<sub>2</sub>O<sub>2</sub>, is -1 and in superoxides, KO<sub>2</sub> is -1/2
5. Cl and Br have oxidation number of -1 in compounds, except when combined with oxygen or fluorine.
6. The oxidation number for H is +1 in most compounds, in hydrides NaH, it is -1.
7. The sum of all of the oxidation numbers for the atoms in a neutral compound must be zero.

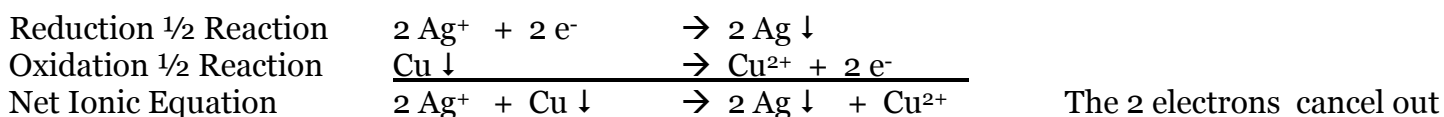
Copper in an aqueous solution of Silver Nitrate, Silver will plate out onto the Copper.

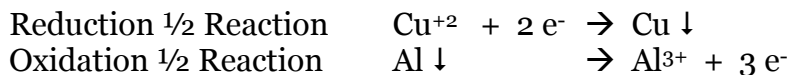


### Balancing of Redox Reactions using half reactions:

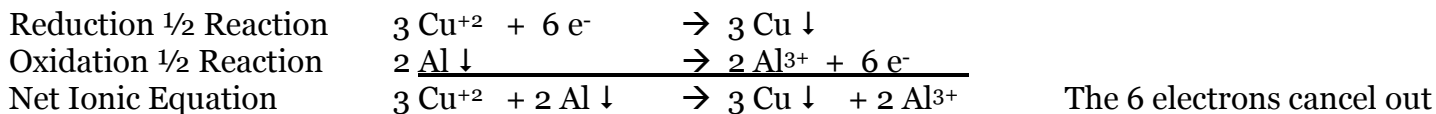


Notice the Silver uses one electron and the copper generates 2 electrons. Therefore we need to multiply the Silver 1/2 reaction by 2:

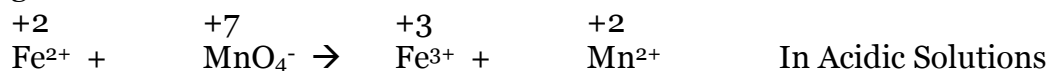


**IE 20.1** Balance Redox of Aluminum and Copper <sup>2+</sup>

Copper uses 2 electrons, Aluminum generates 3 electrons, so cross multiply 3 Cu<sup>+2</sup> and 2 Al ↓

**Balancing Redox Equations in Acid Solutions** Fe<sup>2+</sup> + MnO<sub>4</sub><sup>-</sup> → Fe<sup>3+</sup> + Mn<sup>2+</sup>

1. Assign an oxidation number to each atom



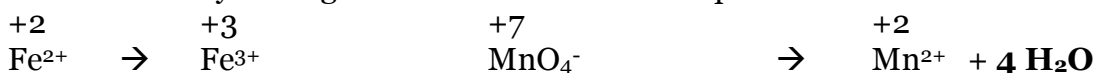
2. Split the reaction to two half reactions:



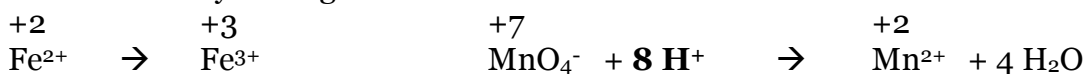
3. Balance the half reactions

a. Balance all atoms except O and H

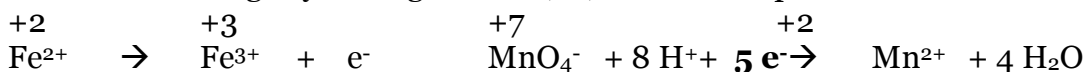
b. Balance O atoms by adding H<sub>2</sub>O's to one side of the equation



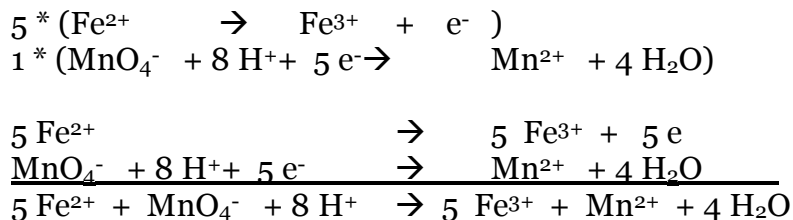
c. Balance H atoms by adding H<sup>+</sup> to one side of the reaction



d. Balance electric charge by adding electrons, e<sup>-</sup>, to the more positive side



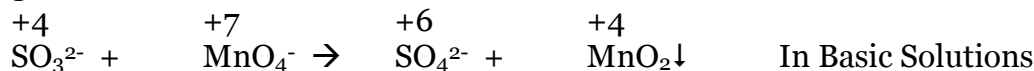
4. Combine each half reaction. Multiply each half reaction by a factor such that when the half reactions are combined, the electrons cancel. Hint: multiply the 1<sup>st</sup> half reaction by the number of electrons in the 2<sup>nd</sup> and multiply the 2<sup>nd</sup> half reaction by the number of electrons in the first.

**IE 20.2** Balance the following in Acidic solution: VO<sub>2</sub><sup>+</sup> + Zn → VO<sup>2+</sup> + Zn<sup>+</sup>

## Balancing Redox Equations in Basic Solutions



1. Assign an oxidation number to each atom



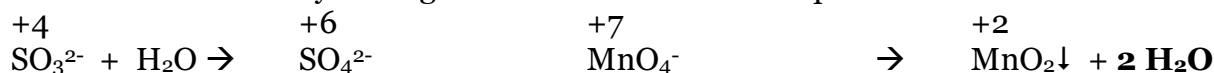
2. Split the reaction to two half reactions:



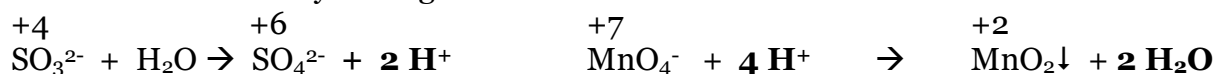
3. Balance the half reactions

a. Balance all atoms except O and H

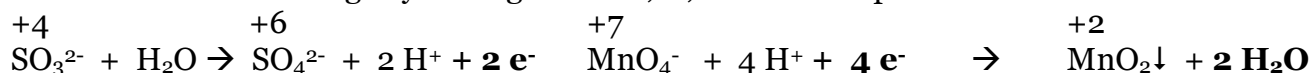
b. Balance O atoms by adding H<sub>2</sub>O's to one side of the equation



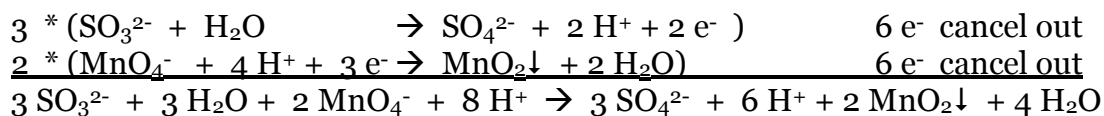
c. Balance H atoms by adding H<sup>+</sup> to one side of the reaction



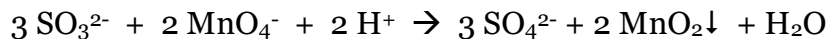
d. Balance electric charge by adding electrons, e<sup>-</sup>, to the more positive side



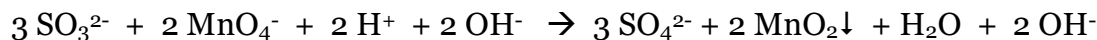
4. Combine each half reaction. Multiply each half reaction by a factor such that when the half reactions are combined, the electrons cancel. Hint: multiply the 1<sup>st</sup> half reaction by the number of electrons in the 2<sup>nd</sup> and multiply the 2<sup>nd</sup> half reaction by the number of electrons in the first.



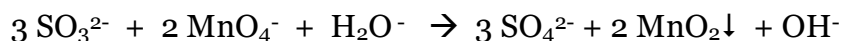
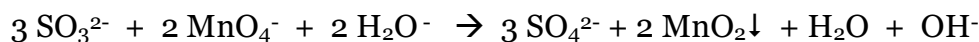
Cancel some water and H<sup>+</sup>



5. Add the number of OH<sup>-</sup> to each side as there are H<sup>+</sup> on one side

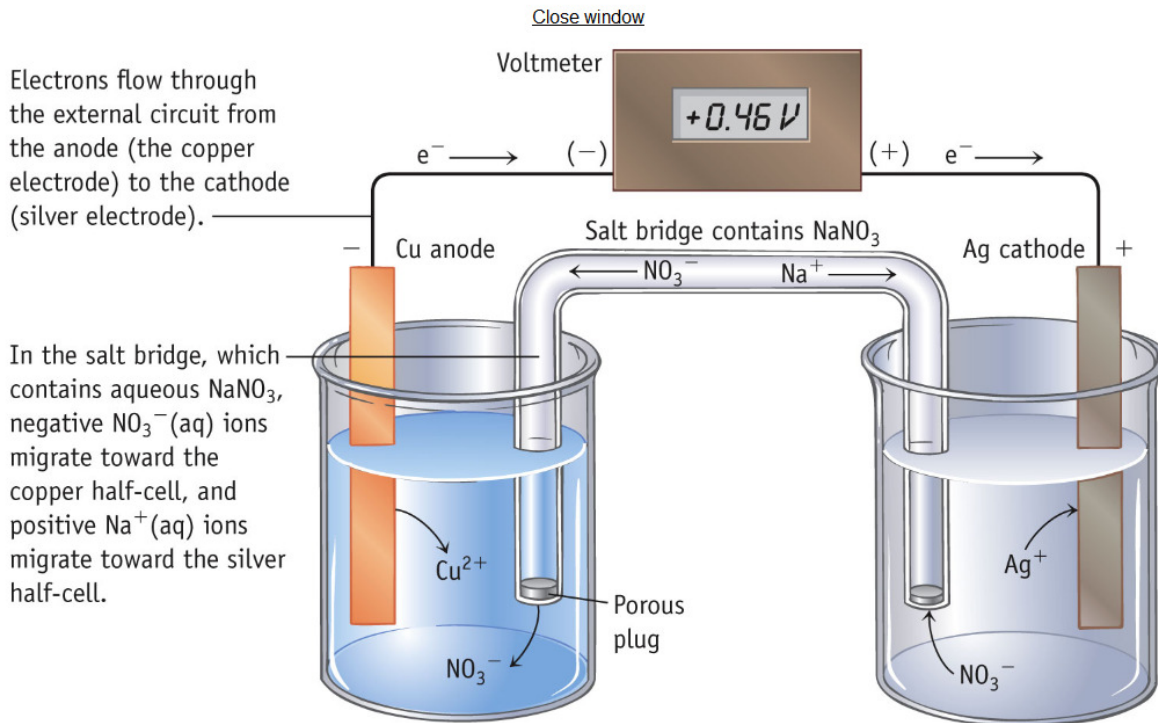


6. Combine 2 H<sup>+</sup> + OH<sup>-</sup> → H<sub>2</sub>O and cancel waters



## 20.2 Simple Voltaic Cells

To save this image, right-click (Windows) or Ctrl-click (Macintosh).



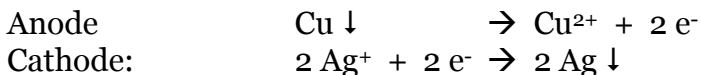
On Left: Copper half cell is copper metal and  $\text{Cu}^{2+}$  solution  
 On Right: Silver half cell is silver metal and  $\text{Ag}^+$  solution

**Salt Bridge:** cell connection that allows cations and anions to move between the two cells, but does not allow the solutions to mix

**Anode (-):** Electrode where oxidation occurs, increase in oxidation number, metal to ion  
 The Copper electrode

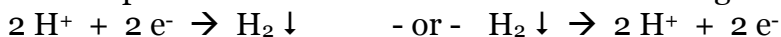
**Cathode (+):** Electrode where reduction occurs, decrease in oxidation number, ion to metal ( plates out)  
 The Silver electrode

**Electrons flow** from the Anode to the Cathode via an external circuit



**Example 20.4** Students do this one  $\text{Fe} \downarrow + \text{Cu}^{2+} \rightarrow \text{Cu} \downarrow + \text{Fe}^{2+}$

**Voltaic Cells with Inert Electrodes** The **Hydrogen Electrode** – this is the standard reference  
 Bubble  $\text{H}_2$  gas over a platinum electrode in a solution containing  $\text{H}^+$



An Iron / Hydrogen cell  $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$  and  $\text{H}_2 \downarrow \rightarrow 2 \text{H}^+ + 2 \text{e}^-$



**Electrochemical Cell Notations**  $\text{Cu} \downarrow | \text{Cu}^{2+} (\text{aq } 1.0 \text{ M}) || \text{Ag}^+ (\text{aq } 1.0 \text{ M}) | \text{Ag} \downarrow$   
Anode on the left Cathode on the right

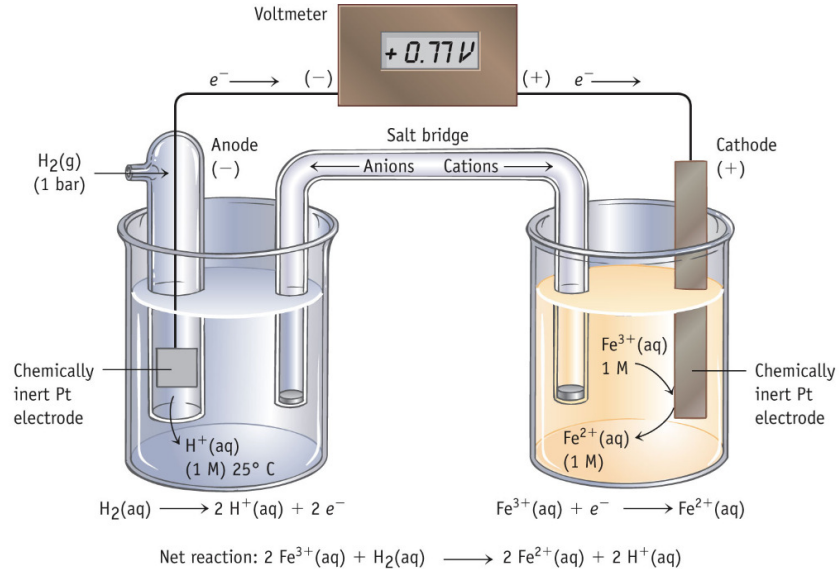
| is a phase boundary such as between the metal and the ions

|| is the salt bridge

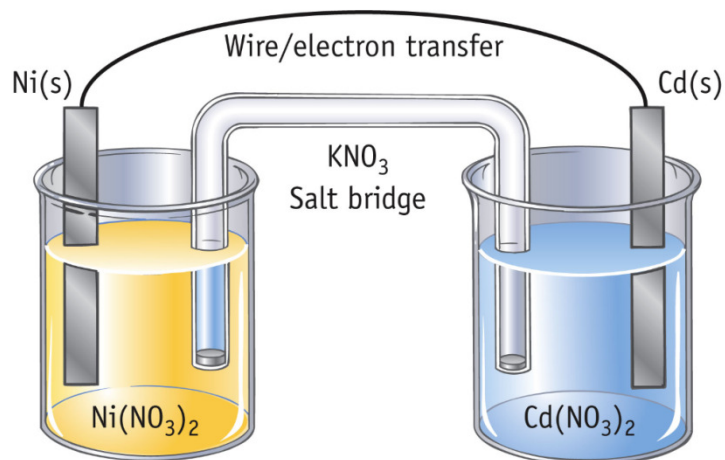
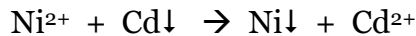
Note: The concentration of ionic solutions is given

**Hydrogen / Iron cell**

$\text{Pt} | \text{H}_2 (\text{P} = 1 \text{ bar}) | \text{H}^+ (\text{aq } 1.0 \text{ M}) || \text{Fe}^{3+} (\text{aq } 1.0 \text{ M}), \text{Fe}^{2+} (\text{aq } 1.0 \text{ M}) | \text{Pt}$   
Hydrogen Electrode Ferric / Ferrous



**Review Check 20.2**



Write out balanced half reactions and label Anode, Cathode, Electron Flow

**20.3 Commercial Voltaic Cells**

The previously discussed Voltaic Cells do not make good Commercial Cells

- 1. They are not robust
- 2. Voltage varies with the concentration of the ions
- 3. Current flow is low

A good Commercial Cell

- 1. Have a large mass of reactants to produce current over a long period
- 2. Would be nice to be Re-Chargeable

**Primary Battery:** cannot be recharged

**Secondary or Rechargeable:** the reactions can be reversed to recharge the cell

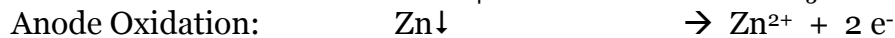
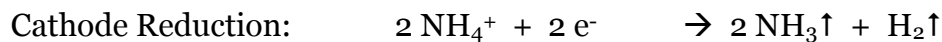
**Common Dry Cell or Zinc / Graphite cell**

Voltage: 1.5

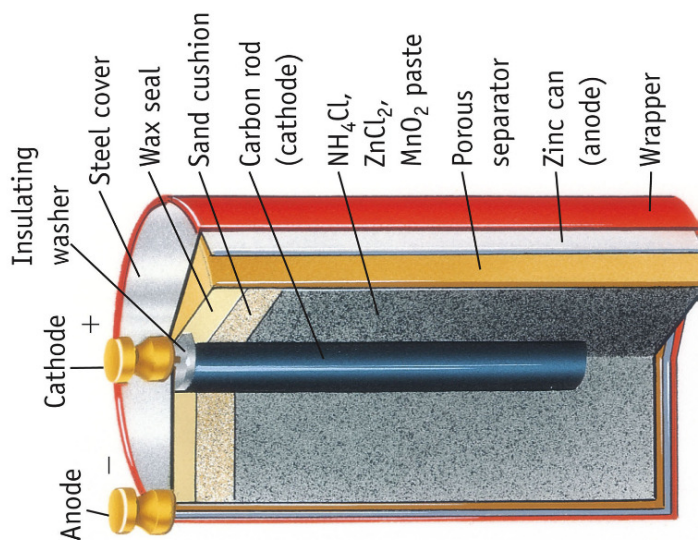
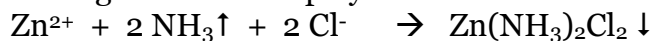
Pros: Inexpensive

Cons: Can leak with use as Zn is used up, generates gas's, voltage drops,

Contains a moist paste of  $\text{NH}_4\text{Cl}$ ,  $\text{ZnCl}_2$  and  $\text{MnO}_2$ , Zinc metal can and graphite rod  
The Zinc outer can is used up and the battery can leak:



The cathode gases are use up by:

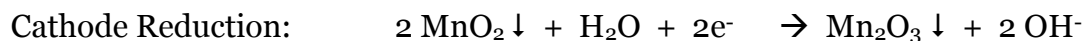


**Alkaline Batteries** – last 50% longer then dry cells,

Voltage: 1.54 v

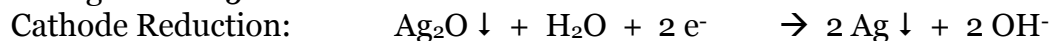
Pros: Voltage does not decline with use, can generate high current

Cons:



**Silver Oxide Battery**

Voltage: 1.5 v



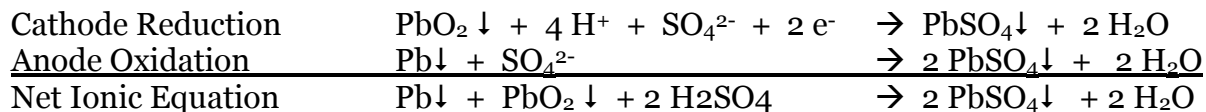
**Zinc-Oxygen Battery**

Voltage: 1.15 – 1.35 v



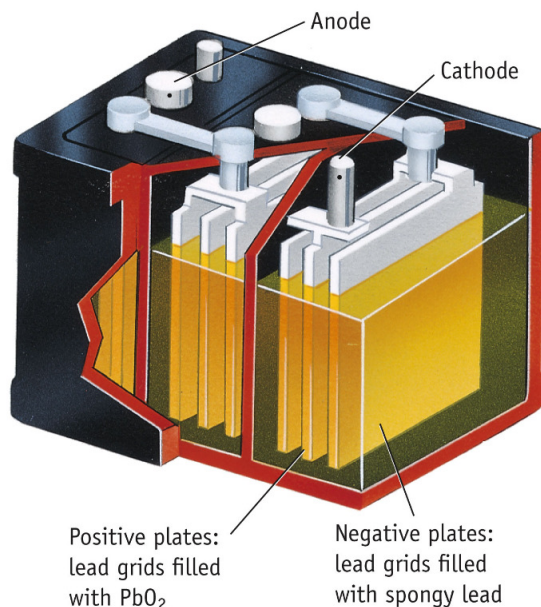
## Rechargeable Batteries

**Lead Storage.** Cathode is Lead metal / Lead (IV) Oxide, the anode is Lead metal.



Students should be able to prove the Net Ionic Equation

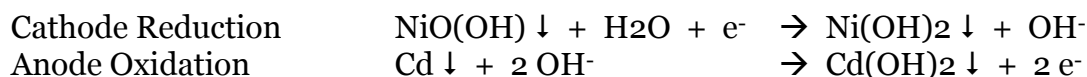
This is the current generating equation, to charge the battery, reverse the reactions



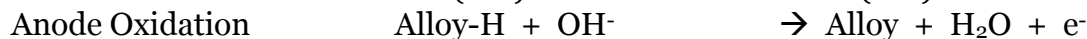
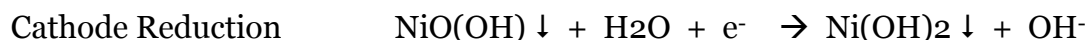
## Nickel-Cadmium Battery

Pros: Produce a constant voltage, lightweight, rechargeable

Cons: Cost is high, Cadmium is toxic



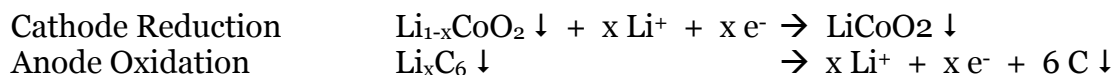
## Nickel-Metal Hydride



Alloy contains Lanthanum, Cerium or Neodymium

## Lithium Battery

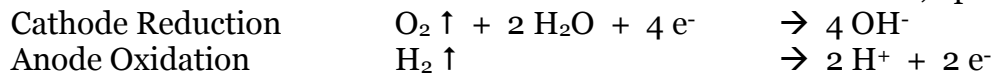
Pros: Can be made into any shape, high current, many cycles, light weight

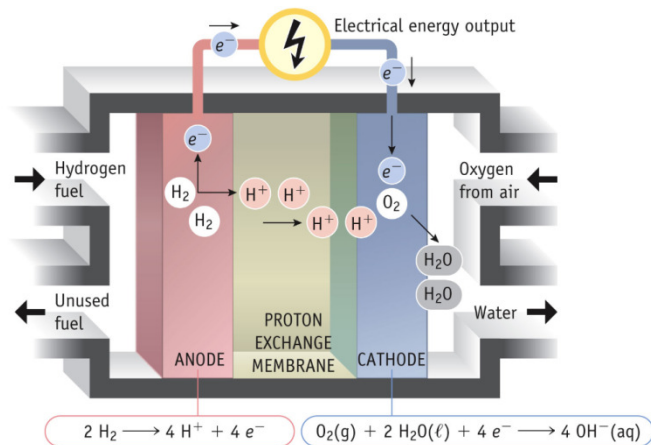


## Fuel Cell

Pros: Reactants can be supplied continuously 40-60% efficiency

Cons: Reactants are hard to distribute via a "Gas Station", operating temp 70-80 °C





## 20.4 Standard Electrochemical Potentials

Electromotive Force or emf is measured in Volts.

One Volt is the potential difference needed to impart one joule of energy to an electric charge of one coulomb  $J = V C$

One coulomb is the quantity of current when 1 ampere flows for 1 sec  $C = A \text{ Time}$

### Standard Potentials

1. Reactants and products are present in their standard states
2. Aqueous solutions are 1.0 M
3. Gases are at a pressure of 1.0 Bar

**Standard Cell Potential**  $E^\circ_{\text{cell}}$  is at 25 °C Are referenced to the Standard Hydrogen Electrode  
Which way do the electrons flow? They always flow from the anode (-) to the cathode (+)

### Table 20.1 Standard Reduction Potentials in Aqueous Solutions at 25 °C

1.  $\frac{1}{2}$  Reactions are written as "Oxidized +  $e^- \rightarrow$  Reduced", species on left is oxidizing agent, on right reducing agent. All potentials are for Reduction Reactions, decrease in Oxidation Number, potentials are in Volts vs the Standard Hydrogen Electrode (SHE) and are called Standard Reduction Potentials.
2. The more Positive value of  $E^\circ$ , the better the oxidizing ability.
3. The more Negative value for  $E^\circ$ , the less likely the  $\frac{1}{2}$  reaction will occur.
4. The reaction of any substance on the left (an oxidizing agent) with a substance LOWER than it on the right (a reducing agent) is product favored at equilibrium (it will occur). Product favored reactions will always involve a reducing agent that is "Southeast" of the oxidizing agent

Reduction Half-Reaction	
$I_2(s) + 2 e^- \rightarrow 2 I^-(aq)$	The northwest-southeast rule: The reducing agent always lies to the southeast of the oxidizing agent in a product-favored reaction.
$Cu^{2+}(aq) + 2 e^- \rightarrow Cu(s)$	
$2 H^+(aq) + 2 e^- \rightarrow H_2(g)$	
$Fe^{2+}(aq) + 2 e^- \rightarrow Fe(s)$	
$Zn^{2+}(aq) + 2 e^- \rightarrow Zn(s)$	

5. The algebraic sign of the  $\frac{1}{2}$  reaction reduction potential is compared to the SHE.
6. Electrochemical potentials depend on the nature of the reactants and products and their concentrations, not on the quantities of material used.

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$  When  $E^\circ_{\text{cell}} = \text{positive}$ , the reaction as written is predicted to be **Product-Favored**, If **negative**, it is **Reactant Favored**.

The further apart the Oxidizing and Reducing Agents, the greater is  $E^\circ_{\text{cell}}$



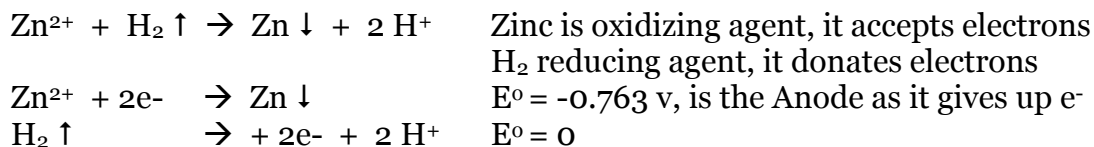
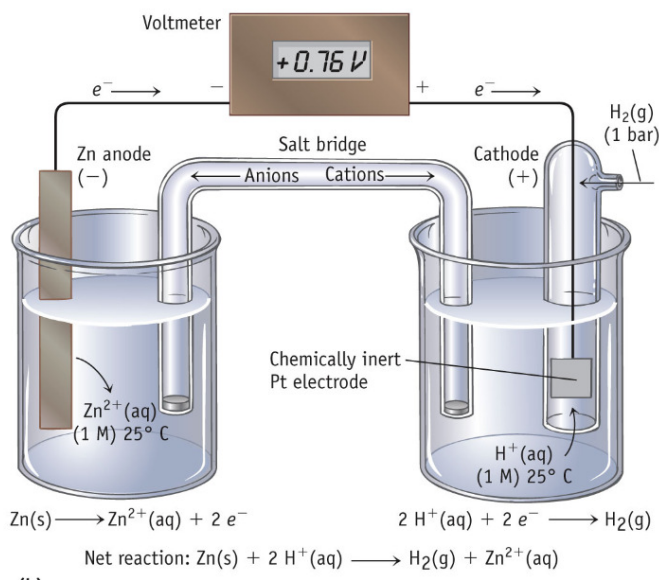
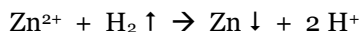
**Table 20.1** Standard Reduction Potentials In Aqueous Solution at 25 °C\*

Reduction Half-Reaction	$E^\circ$ (V)
$F_2(g) + 2 e^- \rightarrow 2 F^-(aq)$	+2.87
$H_2O_2(aq) + 2 H^+(aq) + 2 e^- \rightarrow 2 H_2O(l)$	+1.77
$PbO_2(s) + SO_4^{2-}(aq) + 4 H^+(aq) + 2 e^- \rightarrow PbSO_4(s) + 2 H_2O(l)$	+1.685
$MnO_4^-(aq) + 8 H^+(aq) + 5 e^- \rightarrow Mn^{2+}(aq) + 4 H_2O(l)$	+1.51
$Au^{3+}(aq) + 3 e^- \rightarrow Au(s)$	+1.50
$Cl_2(g) + 2 e^- \rightarrow 2 Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}(aq) + 14 H^+(aq) + 6 e^- \rightarrow 2 Cr^{3+}(aq) + 7 H_2O(l)$	+1.33
$O_2(g) + 4 H^+(aq) + 4 e^- \rightarrow 2 H_2O(l)$	+1.229
$Br_2(l) + 2 e^- \rightarrow 2 Br^-(aq)$	+1.08
$NO_3^-(aq) + 4 H^+(aq) + 3 e^- \rightarrow NO(g) + 2 H_2O(l)$	+0.96
$OCl^-(aq) + H_2O(l) + 2 e^- \rightarrow Cl^-(aq) + 2 OH^-(aq)$	+0.89
$Hg^{2+}(aq) + 2 e^- \rightarrow Hg(l)$	+0.855
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.799
$Hg_2^{2+}(aq) + 2 e^- \rightarrow 2 Hg(l)$	+0.789
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.771
$I_2(s) + 2 e^- \rightarrow 2 I^-(aq)$	+0.535
$O_2(g) + 2 H_2O(l) + 4 e^- \rightarrow 4 OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2 e^- \rightarrow Cu(s)$	+0.337
$Sn^{4+}(aq) + 2 e^- \rightarrow Sn^{2+}(aq)$	+0.15
$2 H^+(aq) + 2 e^- \rightarrow H_2(g)$	0.00
$Sn^{2+}(aq) + 2 e^- \rightarrow Sn(s)$	-0.14
$Ni^{2+}(aq) + 2 e^- \rightarrow Ni(s)$	-0.25
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$	-0.255
$PbSO_4(s) + 2 e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$	-0.356
$Cd^{2+}(aq) + 2 e^- \rightarrow Cd(s)$	-0.40
$Fe^{2+}(aq) + 2 e^- \rightarrow Fe(s)$	-0.44
$Zn^{2+}(aq) + 2 e^- \rightarrow Zn(s)$	-0.763
$2 H_2O(l) + 2 e^- \rightarrow H_2(g) + 2 OH^-(aq)$	-0.8277
$Al^{3+}(aq) + 3 e^- \rightarrow Al(s)$	-1.66
$Mg^{2+}(aq) + 2 e^- \rightarrow Mg(s)$	-2.37
$Na^+(aq) + e^- \rightarrow Na(s)$	-2.714
$K^+(aq) + e^- \rightarrow K(s)$	-2.925
$Li^+(aq) + e^- \rightarrow Li(s)$	-3.045

Increasing strength of oxidizing agents

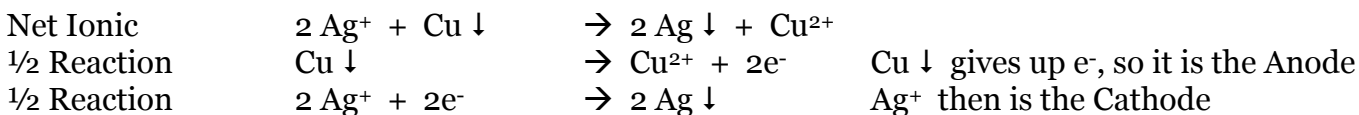
Increasing strength of reducing agents

Will this reaction occur as written:

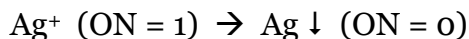


$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}} = -0.763 - 0 = -0.763$  This reaction will not occur as written!

### Calculating Standard Cell Potentials



Silver is the Cathode, at the cathode there is a decrease in oxidation number, it is the Oxidation Agent



Copper is then the Anode and is the Reducing Agent

$E^0 \text{Ag}^+ = +0.799 \text{ v}$                        $E^0 \text{Cu} \downarrow = -0.337 \text{ v}$

$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}} = E^0_{\text{Ag}^+} - E^0_{\text{Cu}_1} = (+0.799) - (-0.337) = +0.462 \text{ v}$

The Positive value for  $E^0_{\text{cell}}$  says this is a Product-Favored Reaction – it will run as written

### Relative Strengths of Oxidizing and Reducing Agents

$E^0, \text{V}$		Reduction Half-Reaction
+1.36	Increasing strength ↑ as oxidizing agents	$\text{Cl}_2(\text{g}) + 2 e^- \rightarrow 2 \text{Cl}^-(\text{aq})$
+0.80		$\text{Ag}^+(\text{aq}) + e^- \rightarrow \text{Ag}(\text{s})$
+0.00		$2 \text{H}^+(\text{aq}) + 2 e^- \rightarrow \text{H}_2(\text{g})$
-0.25		$\text{Ni}^{2+}(\text{aq}) + 2 e^- \rightarrow \text{Ni}(\text{s})$
-0.76		$\text{Zn}^{2+}(\text{aq}) + 2 e^- \rightarrow \text{Zn}(\text{s})$

Upper left of this and Table 20.1 shows  $\text{Cl}_2$  the strongest oxidizing agents, these are easily reduced.

Bottom left shows  $\text{Zn}^{2+}$  which is a poor oxidizing agent or easily reduced.

Upper Right shows  $\text{Cl}^-$ , a poor reducing agent

Lower Right shows  $\text{Zn}$  ↓, which is a good reducing agent.

**Oxidizing Agents:**  $\text{Cl}_2 > \text{Ag}^+ > \text{H}^+ > \text{Ni}^{2+} > \text{Zn}^{2+}$

**Reducing Agents:**  $\text{Zn} > \text{Ni} > \text{H}_2 > \text{Ag} > \text{Cl}^-$   
**Strong** **Weak**

## 20.5 Electrochemical Cells under Nonstandard Conditions

**The Nernst Equation:**  $E = E^\circ - (RT/nF) \ln Q$

E is the cell voltage,  $E^\circ$  the standard cell voltage

R the gas constant 8.314472 J/K mol T is temperature in K

n is the number of moles of electrons transferred

F the **Faraday Constant**  $9.6485338 \times 10^4 \text{ C/mol}$  C = coulomb = 1 ampere for 1 second

Q is the reaction quotient

The Nernst Equation Reduces to:  $E = E^\circ - (0.0257/n) \ln Q$  at 25 °C

### The Math behind the pH Meter

The following Voltaic Cell: 1.00 M  $\text{Cu}/\text{Cu}^{2+}$  vs the SHE

Cathode Reduction:  $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \downarrow$

Anode Oxidation:  $\text{H}_2 \uparrow \rightarrow 2\text{H}^+ + 2e^-$   $\text{H}_2 \uparrow$  gives up  $e^-$ , so it is the Anode

Net Ionic Equation:  $\text{Cu}^{2+} + \text{H}_2 \uparrow \rightarrow \text{Cu} \downarrow + 2\text{H}^+$

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = E^\circ_{\text{Cu}^{2+}} - E^\circ_{\text{H}_2 \uparrow} = (+0.337 \text{ v}) - (0.00\text{v}) = + 0.337 \text{ v}$

$$Q = [\text{H}^+]^2 / [\text{Cu}^{2+}] P_{\text{H}_2}$$

$$E = 0.337\text{v} - (0.257/2) \ln ([\text{H}^+]^2 / [\text{Cu}^{2+}] P_{\text{H}_2})$$

E is measured at 0.490 v

$$0.490 \text{ v} = 0.337 \text{ v} - (0.257/2) \ln ([\text{H}^+]^2 / (1.00)(1.00))$$

$$-11.9 = \ln ([\text{H}^+]^2)$$

$$[\text{H}^+] = 3 \times 10^{-3} \text{ M}$$

$$\text{pH} = 2.6$$

In practice a standard  $\text{Ag}/\text{AgCl}$  or **Calomel Electrode** is used in place of the Copper electrode

### $E^\circ$ and the Equilibrium Constant

As a Voltaic cell produces electric current, the reactant concentration decreases and the cell voltage changes. When  $E_{\text{cell}} = 0$ , there is no further net reaction and the reaction is in equilibrium.

$$E = E^\circ - (0.0257/n) \ln Q = 0 = E^\circ - (0.0257/n) \ln K \quad \text{or}$$

$$\ln K = n E^\circ / 0.0257 \text{ at } 25 \text{ }^\circ\text{C}$$

**IE 20.9** Calculate K for:  $\text{Fe}\downarrow + \text{Cd}^{2+} \rightleftharpoons \text{Fe}^{2+} + \text{Cd}\downarrow$

Cathode Reduction  $\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}\downarrow$   $E^\circ = -0.40 \text{ v}$

Anode Oxidation  $\text{Fe}\downarrow \rightarrow \text{Fe}^{2+} + 2\text{e}^-$   $E^\circ = -0.44 \text{ v}$

Note: Even though the reaction is reversed, the sign of the voltage IS NOT reversed

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = E^\circ_{\text{Cd}^{2+}} - E^\circ_{\text{Fe}} = (-0.40 \text{ v}) - (-0.44 \text{ v}) = +0.04 \text{ V}$$

$$\ln K = n E^\circ / 0.0257 = 2 * 0.04 \text{ V} / 0.0257 = 3.1$$

$$K = 20$$

**20.7 Electrolysis** the use of electrical energy to bring about a chemical change

Electrolysis of molten Sodium Chloride to get Sodium metal and Chlorine gas

Cathode Reduction  $2 \text{Na}^+ + 2 \text{e}^- \rightarrow 2 \text{Na}\downarrow$   $\text{e}^-$  from electrode to chemical species

Anode Oxidation  $2 \text{Cl}^- \rightarrow \text{Cl}_2\uparrow + 2 \text{e}^-$   $\text{e}^-$  from chemical species to electrode

Net Ionic Equation  $2 \text{Na}^+ + 2 \text{Cl}^- \rightarrow 2 \text{Na}\downarrow + \text{Cl}_2\uparrow$

### Electrolysis of Aqueous Solutions

Water is an Electro Active substance, it can electrolyze as well as molten NaCl

Electrolysis of an aqueous solution of NaI yields the following ions:  $\text{Na}^+$ ,  $\text{I}^-$ , and  $\text{H}_2\text{O}$ .

Cathode Reduction  $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}\downarrow$   
 $2 \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{H}_2\uparrow + 2 \text{OH}^-$

Anode Oxidation  $2 \text{I}^- \rightarrow \text{I}_2 + 2 \text{e}^-$   
 $2 \text{H}_2\text{O} \rightarrow \text{O}_2\uparrow + 4 \text{H}^+ + 4 \text{e}^-$

Any Sodium Metal ( $\text{Na}\downarrow$ ) formed would immediately react with water to produce  $\text{NaOH}$  and  $\text{H}_2\uparrow$

Cathode Reduction  $2 \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{H}_2\uparrow + 2 \text{OH}^-$   $E^\circ_{\text{a}} = -0.8277 \text{ V}$

Anode Oxidation  $2 \text{I}^- \rightarrow \text{I}_2 + 2 \text{e}^-$   $E^\circ_{\text{a}} = -0.535 \text{ V}$

Note Text states  $-0.621 \text{ v}$ , See table 20.1 and reverse the value!

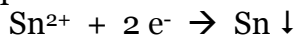
Net Ionic Equation  $2 \text{H}_2\text{O} + 2 \text{I}^- \rightarrow \text{H}_2\uparrow + 2 \text{OH}^- + \text{I}_2$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (-0.8277 \text{ V}) - (-0.535 \text{ V}) = -0.293 \text{ v. The process is NOT spontaneous}$$

### 20.8 Counting Electrons

Electrolysis of aqueous  $\text{AgNO}_3$  requires one mole of electrons to produce one mole of silver as only one electron is transferred  $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}\downarrow$

Electrolysis of aqueous Tin +2 requires 2 moles of electrons as two electrons are transferred:



Electric Charge (Coulombs, C) = Current (amperes, A) \* Time (seconds)

Moles of  $\text{e}^-$  = C/F = Coulombs /  $9.648 \times 10^4 \text{ C/mol e}^-$

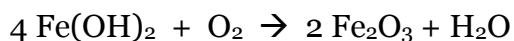
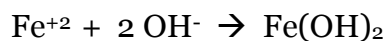
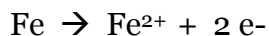
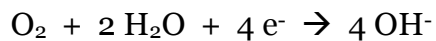
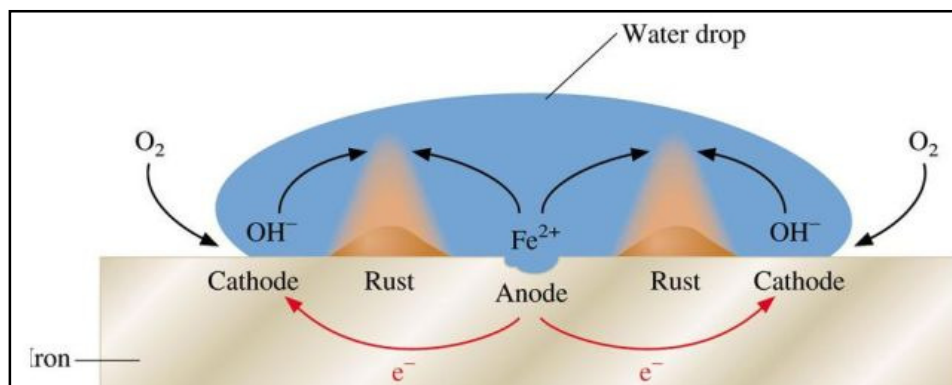
**IE 20.11** 2.40 A is passed through a  $\text{Cu}^{2+}$  aqueous solution for 30.0 min, how many grams of Cu is deposited?

$$C = A \cdot t = 2.40 \text{ A} \cdot 30.0 \text{ min} \cdot 60 \text{ sec/min} = 4.32 \times 10^3 \text{ C}$$

$$\text{Moles of } e^- = (4.32 \times 10^3 \text{ C}) / 96,485 \text{ C/mol } e^- = 4.48 \times 10^{-2} \text{ mol } e^-$$

$$\text{g Cu} = (4.48 \times 10^{-2} \text{ mol } e^-) \cdot (1 \text{ mol Cu} / 2 \text{ mol } e^-) (63.55 \text{ g Cu} / 1 \text{ mol Cu}) = 1.42 \text{ g}$$

**Rusting from a drop of water** – the electrochemical process of rusting iron



## CHEAT SHEET

**Oxidation:** Loss of electrons – or increase in oxidation number

**Reduction:** Gain of electrons – or decrease in oxidation number

**Oxidation Agent:** Accepts electrons, this species is reduced, oxidation number decreases

**Reducing Agent:** Donates electrons, this species is oxidized, oxidation number increases

### Oxidation Number

**Salt Bridge:** cell connection that allows cations and anions to move between the two cells, but does not allow the solutions to mix

**Anode (-):** Electrode where **oxidation** occurs, increase in oxidation number, metal to ion

**Cathode (+) :** Electrode where **reduction** occurs, decrease in oxidation number, ion to metal (plates out)

**Electrons flow from the Anode to the Cathode** via an external circuit

To save this image, right-click (Windows) or Ctrl-click (Macintosh).

[Close window](#)

Electrons flow through the external circuit from the anode (the copper electrode) to the cathode (silver electrode).

In the salt bridge, which contains aqueous  $\text{NaNO}_3$ , negative  $\text{NO}_3^-$  (aq) ions migrate toward the copper half-cell, and positive  $\text{Na}^+$  (aq) ions migrate toward the silver half-cell.

