

Chapter 17 Electrochemistry



Tesla Motors

- 85 kWh battery rated to deliver 320 miles (265 by EPA)
- Contains 7,104 lithium-ion battery cells in 16 modules wired in series.



Chemistry: OpenStax

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Electrochemistry

- Electrochemistry is the study of batteries and the conversion between chemical and electrical energy.
- Based on redox (oxidation-reduction) reactions in which one substance gains electrons and another loses electrons.
 - These two processes MUST happen together.
- Single replacement, combustion, combination and decomposition rxns are all examples of redox reactions.

Oxidation Numbers

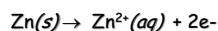
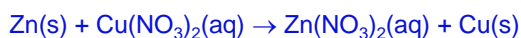
- Free elements: ox # = 0 ($H_2(g)$, $Hg(l)$, etc.)
- Ions in binary ionic compounds: ox # = charge.
 - Ex. For Al_2S_3 , ox # for Al = +3, ox # for S = -2.
- H: ox # = +1, except when H is with alkali metals it is -1 (LiH, NaH, etc.)
- O: ox # = -2, except in peroxides it is -1 (H_2O_2 , K_2O_2)
- The sum of the oxidation numbers of all the atoms in a molecule = 0
- For polyatomic ions, the sum of the oxidation numbers must equal the charge on the ion.

Ox # examples

- Find ox numbers for all the atoms in HCO_3^-
 - H: ox # = +1 O: ox # = -2
- $1 \times H + 1 \times C + 3 \times O = -1$
- $+1 + C + 3 \times (-2) = -1$
 - Ox # for C = +4
- What is the ox # for Cr in $Cr_2O_7^{2-}$?
 - $2Cr + 7O = -2$
 - $2Cr + 7(-2) = -2$
 - Cr = +6

Redox: LEO the lion goes GER

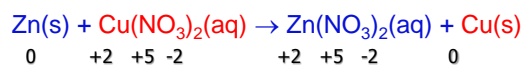
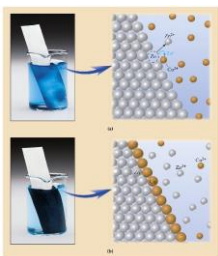
- Oxidation:** loss of electrons; ox # \uparrow , more +
- Reduction:** gain of electrons; ox # \downarrow , more -
- oxidizing agent:** substance that is reduced; it caused oxidization of other substance.
- reducing agent:** substance that is oxidized; it caused reduction of other substance.
- $Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$
- Identify atom oxidized, atom reduced, oxidizing agent and reducing agent.



Zn(s) dissolves



Cu plates on electrode



- Zn 0 to +2 ox # ↑; Zn becomes more +
- Cu +2 to 0 ox # ↓; Cu²⁺ becomes more -

- What is reduced: Cu²⁺ in Cu(NO₃)₂
- What is oxidizing agent: Cu(NO₃)₂
- What is reducing agent: Zn(s)
- What is oxidized: Zn(s)

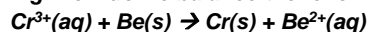
These are always reactants in an overall cell reaction!

Redox Reactions

- What is oxidized/reduced, give the specific atom.
- Agents: give the whole substance as an answer.
- $\text{MnO}_4^- (\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + 5\text{Fe}^{3+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$
- For this reaction:
 - What is the ox # for Mn in MnO₄⁻: +7
 - What is oxidized? Fe²⁺
 - What is reduced? Mn in MnO₄⁻
 - What is the oxidizing agent? MnO₄⁻
 - What is the reducing agent? Fe²⁺

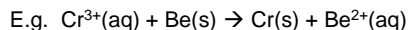
Balancing Simple Redox Reactions

E.g. How do we balance the following reaction?

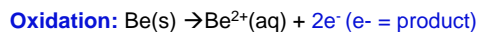


- 1) Break the redox rxn into two **half reactions**, the **oxidation half reaction** and **reduction half reaction**.
- 2) Balance charges by adding electrons (e⁻) to the more + side for each half reaction.
- 3) Balance electrons by multiplying each half reaction by an integer so that the # electrons gained = # electrons lost.

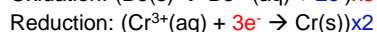
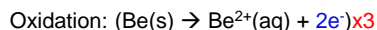
Example Balancing Redox Rxn



- Write the Half reactions



- Balance the electrons



- Add reactions and cancel common terms:



Types of Cells

- **Galvanic cell (also called voltaic cell)**

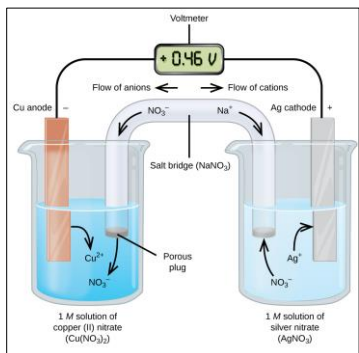
Electrochemical cell in which a spontaneous reaction generates electricity. (*Electric energy is produced.*)

- **Electrolytic cell**

Electrochemical cell in which an electrical current is used to drive a nonspontaneous reaction. (*Electric energy is consumed.*)

Galvanic Cell

Figure 17.4:
A Galvanic cell
provides energy = + V.



[Voltaic Cell: movie](#)
[Electron motion in a cell](#)

Galvanic Cells

Oxidation occurs at the **anode** - both vowels

- mass of anode decreases - it is dissolving as metal atoms lose electrons to form ions in solution

Reduction occurs at the **cathode** - both consonants

- mass of cathode increases as metal ions are reduced to form atoms that plate onto the cathode.

External Circuit - electrons flow from the anode to the cathode via an external wire.

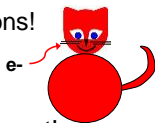
Salt bridge - soluble salt solution in a bridge that connects the two half cells; ions flow through the bridge to complete the electrical circuit.

[Galvanic cell Youtube Videos:](#)

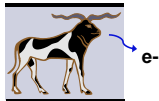
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http://www.youtube.com/watch?feature=player_detail&page&v=C26pH8KC_Wk

Dr. Lisa's cell Mnemonic

- Fat red cat eats electrons!



- Anorexic ox spits them out!



Cathode is reduced (gains e⁻) & mass↑ (plating)
Anode is oxidized (loses e⁻) & mass↓ (dissolves)

What happens in Salt Bridge?

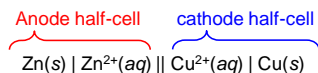
- Migration of ions maintains charge neutrality in both compartments:
- Anions move into the anode where excess + charge builds up as metal cations are formed by oxidation.
- Cations move into the cathode where excess (-) charge builds up as the metal cations are reduced to form neutral metal atoms.

Shorthand Notation for Galvanic Cells

Anode half-reaction: $\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^{-}$

Cathode half-reaction: $\text{Cu}^{2+}(aq) + 2e^{-} \rightarrow \text{Cu}(s)$

Overall cell reaction: $\text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s)$



| = Phase boundary between components in same cell

|| = Salt bridge between two half cells

- electrodes always on 2 ends: anode on left; cathode on right
- Inert electrodes (Pt and graphite) are used for gas phase and aqueous reactions. They allow e⁻ transfer to occur but don't react with cell components.

Electromotive Force

$\text{emf} = E_{\text{cell}} = \text{cell voltage:}$

Electromotive force is the cell potential measured in volts. This is the driving force that pushes electrons away from the anode and towards the cathode.

Joules = Coulombs x Volts

E_{cell} is measured in volts: $V = J/C$

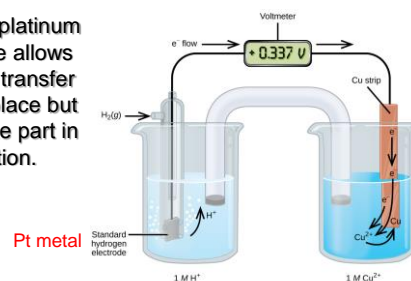
coulomb - the quantity of charge that passes a point in 1 sec when a current of 1 ampere flows.

Standard Cell Potential, E°

- Gases at 1 atm, Solutions at 1 M, Temperature at 298 K (25°C)
- Standard potential for any galvanic cell is the sum of the half-cell potentials.
 - $E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$
- All cell potentials are compared to hydrogen (**SHE: standard hydrogen electrode**)
- SHE consists of Pt electrode in contact with 1 M H^+ solution and H_2 gas at 1 atm pressure.
- $2\text{H}^+ + 2e^- \rightarrow \text{H}_2(\text{g}) \quad E^\circ_{\text{red}} = 0 \text{ V}$

Galvanic Cells – SHE

An inert platinum electrode allows electron transfer to take place but don't take part in the reaction.



Reduction Potentials

- We can use the table of Standard Reduction Potentials for reduction half reactions to determine the cell potential of a galvanic cell.
- Table 17.1: All potentials are listed as **reduction** potentials.
- **Oxidation potentials:** reverse the reaction and $E^\circ_{\text{ox}}(\text{anode}) = -E^\circ_{\text{red}}(\text{cathode})$

Calculating Cell potentials

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}}(\text{cathode}) + E^\circ_{\text{ox}}(\text{anode})$$

- For the oxidation reaction at the anode, make sure you reverse the reaction and change the sign for E°_{ox} , then add the reduction and oxidation potentials.
- E° is intensive; it does not depend on # of moles involved. Don't multiply E° by a factor if coefficients of reaction are changed.

Standard Reduction Potentials

Standard Reduction Potentials at 25°C (See **Table 17.2** – partial table). Full list is in **Appendix L**.

Half-Reaction	E°_{red} (V)
$\text{F}_2(\text{g}) + 2e^- \rightarrow 2\text{F}^-(\text{aq})$	2.866
$\text{Br}_2(\text{l}) + 2e^- \rightarrow 2\text{Br}^-(\text{aq})$	1.0873
$\text{Ag}^+(\text{aq}) + e^- \rightarrow \text{Ag}(\text{s})$	0.7996
$\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})$	0.337
$2\text{H}^+(\text{aq}) + 2e^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Pb}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Pb}(\text{s})$	-0.126
$\text{Cd}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cd}(\text{s})$	-0.4030
$\text{Zn}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Zn}(\text{s})$	-0.7618
$\text{Li}^+(\text{aq}) + e^- \rightarrow \text{Li}(\text{s})$	-3.04

Increasing strength as oxidizing agents ↑ (left side)
 ↓ (right side) Increasing strength as reducing agents

Strength of Oxidizing Agents

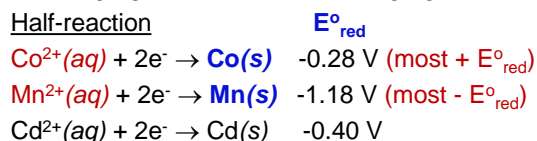
- $\text{F}_2(\text{g}) + 2e^- \rightarrow 2\text{F}^-(\text{aq}) \quad E^\circ_{\text{red}} = 2.866 \text{ V}$
- F_2 has the most positive E°_{red} value.
- F_2 is the easiest to reduce. (It wants e- the most!)
- Thus F_2 is strongest oxidizing agent.
- Active nonmetals are good oxidizing agents. (ox agent = reduced = gain e-)
- As $E^\circ_{\text{red}} \uparrow$, strength ox agent \uparrow

Strength of Reducing Agents

- $\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li}(\text{s}) \quad E^\circ_{\text{red}} = -3.04 \text{ V}$
- Li has the most negative E°_{red} value, so it is the easiest to oxidize.
- Li(s) is strongest reducing agent.
- Li(s) wants to lose electrons, so reverse reaction occurs:
- $\text{Li}(\text{s}) \rightarrow \text{Li}^+(\text{aq}) + \text{e}^- \quad E^\circ_{\text{ox}} = +3.04 \text{ V}$
- Active metals are good reducing agents. (red agent = oxidation = lose e^-).

Cell Potentials

Based on their E°_{red} values, determine the best oxidizing agent, best reducing agent, worst oxidizing agent, and worst reducing agent.



Best/worst agents

For reduction half rxns, oxidizing agents are reactants and reducing agents are products.

- Co^{2+} is the easiest to reduce = best oxidizing agent (most + E°_{red} value).
- Mn^{2+} is the hardest to reduce = worst oxidizing agent. (most - E°_{red} value).
- $\text{Mn}(\text{s})$ is the easiest to oxidize = best reducing agent. (most + E°_{ox} value).
- $\text{Co}(\text{s})$ is the hardest to oxidize = worst reducing agent. (most - E°_{red} value).

Cell Potentials

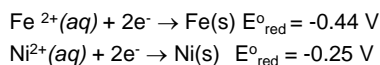
- Positive E°_{cell} means the reaction is product-favored and spontaneous.
 - A negative E°_{cell} means the reaction won't happen in the forward direction.
- Therefore, we want two half reactions that yield the most positive E°_{cell} value.
- Assign rxn with more + E°_{red} as cathode!
- $E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$

Galvanic Cell Potentials

If we were to make a Galvanic cell from the following metals, which would act as the anode and which as the cathode? Refer to the standard reduction potential table.

- Write the half-cell reactions and the overall balanced reaction for each cell.
- Calculate E°_{cell} for the Fe(s) - Ni(s) cell

Example: Fe(s) + Ni(s)



- Assign more + one (Ni) as cathode:
Cathode: $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s}) \quad E^\circ_{\text{red}} = -0.25 \text{ V}$
- For anode: reverse reaction and change sign
Anode: $\text{Fe}(\text{s}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \quad E^\circ_{\text{ox}} = +0.44 \text{ V}$
- Overall: $\text{Ni}^{2+}(\text{aq}) + \text{Fe}(\text{s}) \rightarrow \text{Ni}(\text{s}) + \text{Fe}^{2+}(\text{aq})$
 $E^\circ_{\text{cell}} = -0.25 + 0.44 = 0.19 \text{ V}$

Cell Potentials

Use these values for problem 5 on worksheet.

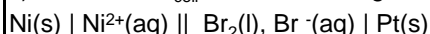
- $\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}(\text{s})$ $E^\circ_{\text{red}} = 1.50 \text{ V}$
- $\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$ $E^\circ_{\text{red}} = 1.07 \text{ V}$
- $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$ $E^\circ_{\text{red}} = -0.13 \text{ V}$
- $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}(\text{s})$ $E^\circ_{\text{red}} = -0.25 \text{ V}$

Cell Potentials in Reactions

5. b) What will E°_{cell} be if Pb^{2+} reacts with Br^- ?

- $E^\circ_{\text{red}} \text{ Pt}(\text{s}) = -0.13 \text{ V}$ $E^\circ_{\text{ox}} \text{ Br}(\text{aq}) = -1.07 \text{ V}$
- Sign changed for bromide since it is oxidized!
- $E^\circ_{\text{cell}} = -1.07 \text{ V} + -0.13 \text{ V} = -1.20 \text{ V}$

c) Calculate E°_{cell} for the following cell



Ni = ox bromine = red

- $E^\circ_{\text{cell}} = 0.25 \text{ V} + 1.07 \text{ V} = 1.32 \text{ V}$
- Reaction for part c is spontaneous (+ E)

Free Energy and Emf

$$\Delta G = -nFE$$

n = number of moles of e⁻ transferred

E = Emf of cell

F = Faraday's constant

$$1F = \frac{96500 \text{ Coulombs}}{1 \text{ mol e}^-} = \frac{96500 \text{ J}}{\text{V} \cdot \text{mol e}^-}$$

spontaneous reaction: $\Delta G < 0$ and $E > 0$

nonspontaneous reaction: $\Delta G > 0$ and $E < 0$

At equilibrium: $\Delta G = 0$ and $E = 0$

Calculate ΔG° from E°

Calculate ΔG° for a reaction between Ni(s) with Au^{3+} .

- Cathode: $2 \times (\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}(\text{s}))$ $E^\circ_{\text{red}} = 1.50 \text{ V}$
 - Anode: $3 \times (\text{Ni}(\text{s}) \rightarrow \text{Ni}^{2+} + 2\text{e}^-)$ $E^\circ_{\text{ox}} = -(-0.25 \text{ V})$
- $$2\text{Au}^{3+} + 3\text{Ni}(\text{s}) \rightarrow 2\text{Au}(\text{s}) + 3\text{Ni}^{2+} \quad E^\circ_{\text{cell}} = 1.75 \text{ V}$$
- n = 6 mol e⁻ transferred from anode to cathode

$$\Delta G^\circ = -(6 \text{ mol e}^-) \left(\frac{96500 \text{ J}}{\text{V} \cdot \text{mol e}^-} \right) (1.75 \text{ V})$$

$$= -1.01 \times 10^6 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = -1.01 \times 10^3 \text{ kJ}$$

*note: the units for n may be omitted, and ΔG° may be expressed in kJ/mol

Standard emf & K

- At equilibrium, $E = 0$ and $Q = K$ (plug these conditions into Nernst Equation)

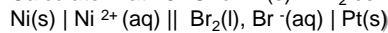
$$0 = E^\circ - \frac{RT}{nF} \ln K \quad \Rightarrow \quad E^\circ = \frac{RT}{nF} \ln K$$

$$R = 8.314 \frac{\text{J}}{\text{K} \cdot \text{mol}}, \quad F = 96500 \frac{\text{J}}{\text{V} \cdot \text{mol e}^-}$$

- At 298 K, $E^\circ = \frac{0.0592 \text{ V}}{n} \log K$

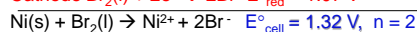
Calculating K from E°

Calculate K at 25 °C for Ni(s) + Br_2 cell (#4c):



Anode $\text{Ni}(\text{s}) \rightarrow \text{Ni}^{2+} + 2\text{e}^-$ $E^\circ_{\text{ox}} = +0.25 \text{ V}$

Cathode $\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-$ $E^\circ_{\text{red}} = 1.07 \text{ V}$



From rearranging the equations relating E° to K

$$K = 10^{\left(\frac{nE^\circ}{0.0592 \text{ V}} \right)} = 10^{\left(\frac{2 \times 1.32 \text{ V}}{0.0592 \text{ V}} \right)} = 3.93 \times 10^{44}$$

or

$$K = e^{\left(\frac{nFE^\circ}{RT} \right)} = e^{\left(\frac{2 \times 96500 \text{ J}}{8.314 \text{ J}} \times \frac{-1.32 \text{ V}}{\text{K} \cdot \text{mol}} \times 298 \text{ K} \right)} = 4.54 \times 10^{44}$$

Cell Potentials Summary

- Positive E°_{cell} value
 - Provides energy
 - ΔG° is negative
 - K is large
 - Reaction is product-favored
- Negative E°_{cell} value
 - Consumes energy
 - ΔG° is positive
 - K is small
 - Reaction is reactant-favored

The Nernst Equation: Calculate emf for nonstandard conditions

Recall: $\Delta G = \Delta G^\circ + RT \ln Q$

- Plugging in $\Delta G = -nFE$ and $\Delta G^\circ = -nFE^\circ$
- $$-nFE = -nFE^\circ + RT \ln Q$$

Nernst Equation:
$$E = E^\circ - \frac{RT}{nF} \ln Q$$

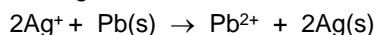
$$E = E^\circ - \frac{2.303RT}{nF} \log Q \quad R = 8.314 \frac{\text{J}}{\text{K} \cdot \text{mol}}, \quad F = 96500 \frac{\text{J}}{\text{V} \cdot \text{mole}}$$

At 298 K:
$$E = E^\circ - \frac{0.0592\text{V}}{n} \log Q$$

- Both []'s & P's can be plugged into Q; V = volts (unit)

Pb and Ag⁺ cell (worksheet)

#8. A galvanic cell utilizes the reaction:



- Write the half reactions and calculate E°_{cell} .
- Write the short-hand notation for this cell.

a) Look up reduction potentials

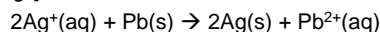


$$E^\circ_{\text{cell}} = 0.80 \text{ V} + 0.13 \text{ V} = 0.93 \text{ V}$$

- $\text{Pb(s)}|\text{Pb}^{2+}(\text{aq})||\text{Ag}^+(\text{aq})|\text{Ag(s)}$

Nernst Equation: Pb/Ag⁺ cell

8c) Calculate the cell potential if $[\text{Pb}^{2+}] = 0.88 \text{ M}$ and $[\text{Ag}^+] = 0.14 \text{ M}$.



$$E^\circ_{\text{cell}} = 0.93 \text{ V}, \quad n = 2$$

$$Q = (0.88 \text{ M}) / (0.14 \text{ M})^2 = 44.9$$

$$E = E^\circ - (0.0592 \text{ V}/n) \log Q$$

$$E = 0.93 \text{ V} - (0.0592 \text{ V}/2) \times \log 44.9$$

$$E = 0.93 \text{ V} - (0.0296 \text{ V})(1.65) = 0.93 - 0.049 = 0.88 \text{ V}$$

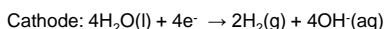
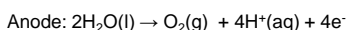
$$E = 0.88 \text{ V} \quad (E \downarrow \text{ because } Q > 1)$$

Do Worksheet problem #9

Electrolysis

Electrolytic Cell: Electrical energy from an external source (outlet or a battery) is used to force a nonspontaneous redox reaction to proceed.

- Water doesn't naturally decompose into H_2 and O_2 . A battery must supply energy to drive this reaction forward.
- $2\text{H}_2\text{O(l)} \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \quad E^\circ_{\text{cell}} = -1.23 \text{ V}$

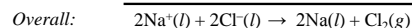
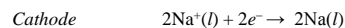
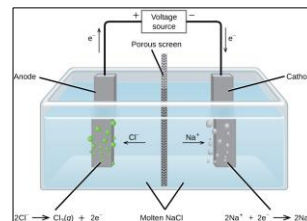


- The electrodes are the still the same for electrolysis: (oxidation at anode, reduction at cathode).

Electrolysis

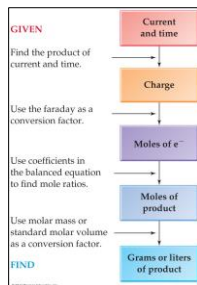
Figure 17.19 Electrolysis of *molten* sodium chloride; $E^\circ_{\text{cell}} = -4 \text{ V}$

The reactions are not spontaneous in the forward direction: Na likes to lose e⁻ and Cl₂ likes to gain e⁻.



Electrolysis Calcs

- Used to find mass or volume of product produced by passing current through cell.
- Current: measured in Amps
 $A = C/s$
- ampere**: unit of electric current; rate of flow of e^- .
- charge = current x time**
- coulombs = amps x seconds**



Electrolysis Calculations

- How many grams of Cu can be collected in 1.00 hour by a current of 1.62 A from a $CuSO_4$ solution?
- Reduction reaction: $Cu^{2+} + 2e^- \rightarrow Cu(s)$
- Calculate coulombs: $Current (C/s) \times time (s) = C$
- $C \rightarrow mol e^- \rightarrow mol solid \rightarrow g solid$
- > **convert time to secs: 1 hr = 3600 secs**

$$3600s \times \frac{1.62C}{s} \times \frac{1 mol e^-}{96,500C} \times \frac{1 mol Cu}{2 mol e^-} \times \frac{63.55g}{mol} = 1.92 g Cu$$

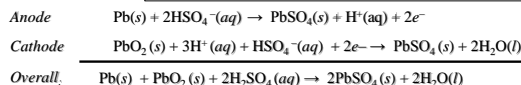
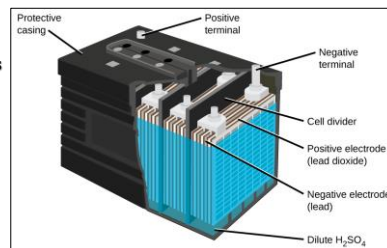
Batteries

- Batteries are examples of Galvanic cells we use in everyday life.
- For a Galvanic cell the voltage keeps dropping as the spontaneous reaction proceeds right.
- Battery is dead when $E_{cell} = 0$ (at equilibrium the rxn no longer proceeds forward; this is minimum pt on free energy curve).
- Bigger batteries only last longer, they don't have more volts. Volts are determined by the chemical reaction that occurs.

Car Batteries

Figure 17.4: Lead Storage (Car) Batteries

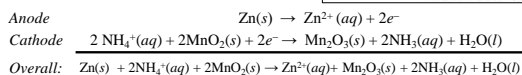
- 6 cells
- 2 V per cell
- 12 V total
- Rechargeable



Dry Cell Batteries

Figure 17.10: Dry Cell (or Laclanche) Batteries: Zinc container (anode), Graphite rod (cathode), 1.5 V

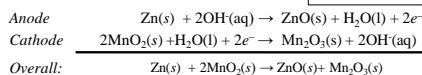
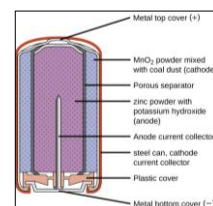
D, C, A, AA, and AAA batteries all have the same Voltage rating. However, larger batteries can deliver more moles of electrons. As the zinc container oxidizes, its contents eventually leak out, so this type of battery should not be left in any electrical device for extended periods.



Alkaline Batteries

Figure 17.11: Alkaline Batteries: Zinc container (anode), Graphite rod (cathode), 1.43 V

Alkaline batteries were designed as direct replacements for zinc-carbon (dry cell) batteries in the 1950's. They can deliver about 3 – 5x the energy of a zinc-carbon dry cell battery of similar size.



9 V Batteries

9 Volt Batteries – 6 x 1.5 V Dry Cell batteries connected in series

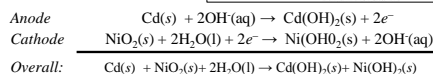
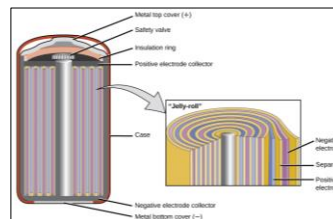


https://upload.wikimedia.org/wikipedia/commons/c/ce/9V_innards_AAAA_removal-1.jpg

NiCd Batteries

NiCd Batteries: Nickel-plated cathode and cadmium-plated anode.

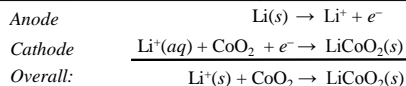
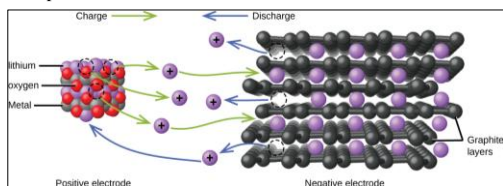
NiCd batteries use a "jelly-roll" design that significantly increases the amount of current the battery can deliver as compared to a similar-sized alkaline battery.



Lithium ion Batteries

Lithium ion (Rechargeable) batteries: charge flows between the electrodes as the lithium ions move between the anode and cathode

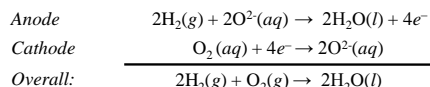
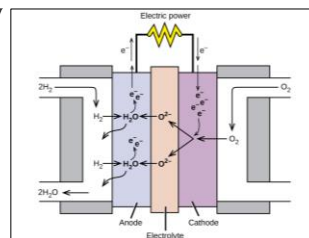
- Cell potential: 3.7 V



Fuel Cells How it works

Fuel Cells; $E^\circ_{\text{cell}} = 1.23 \text{ V}$

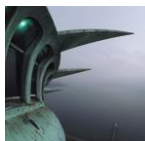
In this hydrogen fuel-cell schematic, oxygen from the air reacts with hydrogen, producing water and electricity.



17.6 Corrosion

The term *corrosion* generally refers to the deterioration of a metal by an electrochemical process (e.g., rusting of iron).

A view from the top of the Statue of Liberty, showing the green patina coating the statue. The patina is formed by corrosion of the copper skin of the statue, which forms a thin layer of an insoluble compound that contains copper(II), sulfate, and hydroxide ions.



(a) The Statue of Liberty is covered with a copper skin, and was originally brown, as shown in this painting. (b) Exposure to the elements has resulted in the formation of the blue-green patina seen today.

Image from Principles of General Chemistry <http://2012books.lardbucket.org/books/principles-of-general-chemistry-v1.0/s23-02-standard-potentials.html> CC-BY-NC-SA 3.0 license