13.3 - Voltaic Cells and Batteries

Electrochemistry

Redox processes convert chemical energy into electrical energy: transferred electrons do work.

Zn/Cu Voltaic cell demo.

Together Forever

1. Example Reaction

Consider this redox reaction:

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$$

What are the half-reactions?

$Zn \rightarrow Zn^{2+} + 2\ e^-$ - oxidation

$Cu^{2+} + 2\ e^- \rightarrow Cu$ - reduction

Electrochemical Cells

Half-reactions are separated in electrochemical cells to harness electrical potential.

The first one was made by Alessandro Volta in 1800.

A cell is two half-cells with an electrolyte, with a different metal electrode in each.

Connecting these is a wire through which electrons flow.

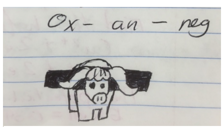
Connecting the electrolyte containers is a salt bridge, through which ions flow.

Details

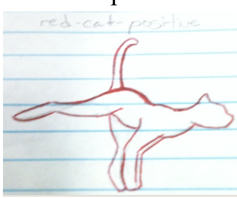
Anode: negative electrode: oxidized.

Cathode: positive electrode: reduced.

Mnemonics: ox-an-neg!



red-cat-positive!

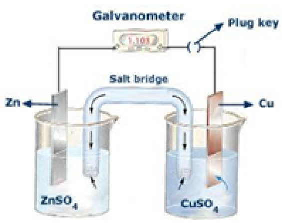


More Details

Salt bridge allows ions to move.

As electrons leave the anode through the wire, negative ions enter that cell to maintain constant charge.

As electrons arrive at the cathode through the wire, positive ions enter that cell.



Reduction Potential

Def.: tendency of substances to gain electrons.
Measured in volts (V) (Resources 9).

When comparing two half-reactions, the one with lower voltage is oxidized.

Voltmeter

Standard Cell Notation

Voltaic cells follow this format:

Reactant | Product ‖ Reactant | Product

Oxidation Side ‖ Reduction Side

The copper/zinc cell is written:

$\boxed{Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu}$

Cell Notation Examples

Write standard cell notation for the following:

2. $La^{3+} + H_2 \longrightarrow La + 2\,H^+$

La^{3+} Potential = -2.37 V Hydrogen = 0 V
Lanthanum is oxidized: $\boxed{La|La^{3+}||2H^+|H_2}$

3. $Au^{3+} + Al \longrightarrow Au + Al^{3+}$

Au Potential = 1.50 V Aluminum = -1.66
Aluminum is oxidized: $\boxed{Al|Al^{3+}||Au^{3+}|Au}$

Calculating Cell Potential

In two half-reactions (oxidation & reduction) of a cell, total cell voltage is the sum:

Equation:

$$\boxed{E^0_{cell} = E^0_{reduction} - E^0_{oxidation}}$$

voltage of reduced chemical

voltage of oxidized chemical

The one with lower potential is oxidized.

4. Voltage Example

$Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$

What's the potential for this cell?

Find the half-cell reactions (Resource P. 9):

Copper: $Cu^{2+} + 2\,e^- \rightarrow Cu$ = + 0.34 V
Zinc: $Zn^{2+} + 2\,e^- \rightarrow Zn$ = - 0.76 V

The potential is:

$$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$$

E^0_{cell} = + 0.34 V – (-0.76 V) = 1.10 V

5. Potential Example

What is the potential of this cell?

$Fe \mid Fe^{2+} \parallel I_2 \mid 2I^-$

Find reduction potentials:

$Fe^{2+} + 2\,e^- \rightarrow Fe$ = - 0.44 V
$I_2 + 2\,e^- \rightarrow 2\,I^-$ = + 0.54 V

Iron has the lower cell potential, so is oxidized.

$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$
E^0_{cell} = + 0.54 V – (-0.44 V) = + 0.98 V

Chem Unit 13.3 Notes - Voltaic Cells & Batteries

Batteries

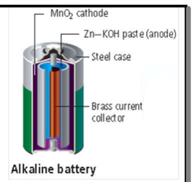
Def: One or more voltaic cells in a single package that generates electric current.

Demo: 6 V & 9 V cut-away batteries.

Primary: reactions not easily reversed – disposable.

Secondary: Reactions are reversible - rechargeable.

6. Dry Cell Batteries



Electrolyte is a moist paste.

A. Alkaline - single use - reaction:

$Zn + H_2O + MnO_2 \rightarrow Mn(OH)_2 + ZnO$

B. NiCad - rechargeable - reaction:

$Cd + 2H_2O + 2NiO(OH) \rightarrow 2Ni(OH)_2 + Cd(OH)_2$

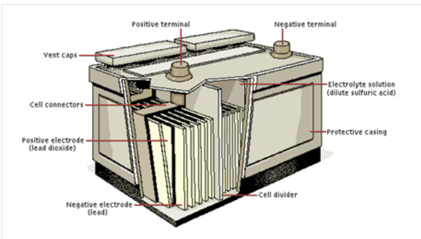
Cadmium metal makes these toxic to the environment.

7. Lead-Acid Battery

Vehicle batteries contain lead and sulfuric acid:

$$Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$$

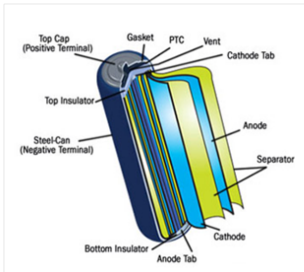
Lead makes these toxic to the environment.



Lithium Ion (primary or secondary)

Lithium batteries have a high reduction potential and a long discharge life span.

Lightweight: has high energy density: energy/mass.



Hydrogen Fuel Cells

H_2 and O_2 can explode violently, but make clean electricity when mixed right.

The first fuel cell was made in 1839 - an old invention.

They use a proton exchange membrane - not an electrolyte.

8. Hydrogen Fuel Cells

Oxidation half-reaction:

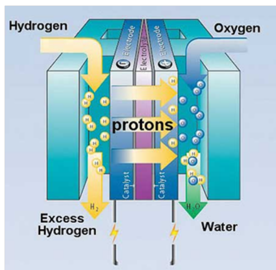
$2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

Reduction half-reaction:

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

The complete reaction:

$2H_2 + O_2 \rightarrow 2H_2O$



Homework

13.3 Problems.
Due: Next Class.