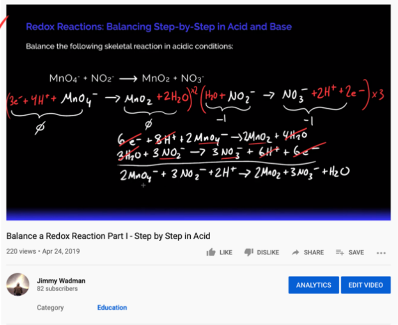
CH 305 Unit 2

THE STANDARD CELL

Balancing Redox Reactions



Acidic Conditions: Follow these steps to balance redox reaction in acidic solutions

1. Identify the species for the oxidation and reduction and write them as two separate half reactions (they will not be balanced at this point).
2. For each half-reaction, balance all elements except for hydrogen and oxygen.
3. Next, balance the oxygens in the reaction by adding H_2O.
4. Now finish by balancing the hydrogens using H^+.
5. Now total the charge on each side of the reaction and balance the charge using electrons (e^-).
6. Make the electron count match in the reduction and the oxidation by multiplying one or both of the balanced half-reactions by whole numbers to equalize the number of the electrons.
7. Now combine the two reactions to make your overall redox reaction. Cancel any possible species that are on both sides of the equation (H_2O and/or H^+).
8. Now do a final overall check by making sure that all the elements and charge are balanced on each side of the reaction.

stoichiometry independent

Electrochemical Cell Definitions

1. **Anode**: the site of oxidation (An Ox) ; gives electrons to the wire
2. **Cathode**: the site of reduction (Red Cat) ; takes in electrons from the wire

} solid

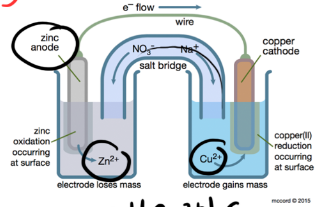
3. **Voltage**: the difference in potential per unit charge (J/C or V); a measurement of the "pulling power" on the electrons

battery

4. **Voltaic Cell:** an electrochemical cell with a positive standard cell potential ; the redox reaction of the cell occurs without an external power source (spontaneous) ; $\varepsilon^\circ_{cathode} > \varepsilon^\circ_{anode}$ (reduction potentials)
5. **Electrolytic Cell:** an electrochemical cell with a negative standard cell potential ; the redox reaction of the cell relies on an external power source (non-spontaneous) $\varepsilon^\circ_{cathode} < \varepsilon^\circ_{anode}$ (reduction potentials)
6. Shorthand Notation:

anode|anodic solution ||cathodic solution | cathode

R | P || R | P



$Zn|Zn^{2+}||Cu^{2+}|Cu$

Electrochemical Potential

Electrical Cell Potential (ε°_{cell}): the voltage associated with the redox reaction occurring in an electrochemical cell

$$\varepsilon^\circ_{cell} = \varepsilon^\circ_{cathode} - \varepsilon^\circ_{anode}$$

in this equation, both ε° values are reduction potentials read from a table

$F_2 + 2e^- \rightleftharpoons 2F^-$ → cathode

$Li \rightleftharpoons Li^+ + e^-$ → anode

$5.92 = 2.87 - (-3.05)$

strongest oxidizing agent

Half Reaction					potential
F_2	+	$2e^-$	⇌	$2F^-$	+2.87 V
Pb^{4+}	+	$2e^-$	⇌	Pb^{2+}	+1.67 V
Cl_2	+	$2e^-$	⇌	$2Cl^-$	+1.36 V
Ag^+	+	$1e^-$	⇌	Ag	+0.80 V
Fe^{3+}	+	$1e^-$	⇌	Fe^{2+}	+0.77 V
Cu^{2+}	+	$2e^-$	⇌	Cu	+0.34 V
$2H^+$	+	$2e^-$	⇌	H_2	0.00 V
Fe^{3+}	+	$3e^-$	⇌	Fe	−0.04 V
Pb^{2+}	+	$2e^-$	⇌	Pb	−0.13 V
Fe^{2+}	+	$2e^-$	⇌	Fe	−0.44 V
Zn^{2+}	+	$2e^-$	⇌	Zn	−0.76 V
Al^{3+}	+	$3e^-$	⇌	Al	−1.66 V
Mg^{2+}	+	$2e^-$	⇌	Mg	−2.36 V
Li^+	+	$1e^-$	⇌	Li	−3.05 V

increasing strength as an oxidizing agent

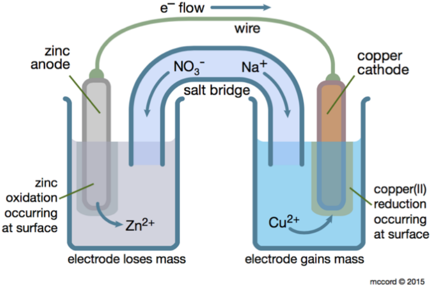
increasing strength as an reducing agent

strongest reducing agent

The Electrochemical Cell Checklist

Make sure you can identify the following for a voltaic or electrolytic cell:

- ❑Cathode (including charge and half-reaction)
- ❑Anode (including charge and half-reaction)
- ❑Salt Bridge and flow of ions
- ❑Direction of electron flow and current flow
- ❑What is driving the push/pull of electrons
- ❑Spontaneity of the redox reaction
- ❑Voltmeter or power supply



$Cu^{2+} + 2e^- \rightarrow Cu$ (0.34V) ✓

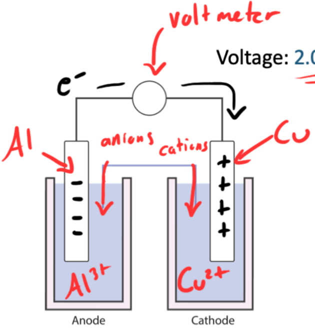
$Al^{3+} + 3e^- \rightarrow Al$ (-1.66V)

The Voltaic Cell

A voltaic cell produces a positive voltage through a spontaneous redox reaction

$$\varepsilon^\circ_{cell} = \varepsilon^\circ_{cathode} - \varepsilon^\circ_{anode}$$

Voltage: 2.00 V = (0.34 V) – (-1.66 V) ; Al | Al^{3+} || Cu^{2+} | Cu



red, cathode : $Cu^{2+} + 2e^- \rightarrow Cu$

ox, anode : $Al \rightarrow Al^{3+} + 3e^-$

Q: touch the electrode that is growing throughout the reaction?

$$Cu^{2+} + 2e^- \rightarrow Cu\ (0.34V)$$

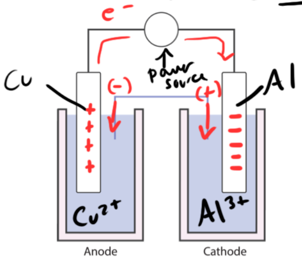
$$Al^{3+} + 3e^- \rightarrow Al\ (-1.66V)$$

The Electrolytic Cell

An electrolytic cell requires a power source to run a non-spontaneous redox reaction

$$\varepsilon^\circ_{cell} = \varepsilon^\circ_{cathode} - \varepsilon^\circ_{anode}$$

Electrolytic Voltage: -2.00 V = (-1.66 V) – (0.34 V) ; Cu | Cu^{2+} || Al^{3+} | Al



red, cathode: $Al^{3+} + 3e^- \rightarrow Al$

ox, anode: $Cu \rightarrow Cu^{2+} + 2e^-$

Shorthand Notation Overview

| → phase change

, → rxn without phase change

anode |anodic solution ||cathodic solution | cathode

R P R P

Al| Al^{3+} ||Cu^{2+} | Cu

s aq aq s

inert electrode

(g) (aq)

Pt | H_2 | H^+ || Cu^{2+} | Cu

$H_2 \rightarrow 2H^+ + 2e^-$

Al| Al^{3+} || Cr^{3+} , Cr^{2+} | C (graphite)

aq aq

inert electrode

The Electrochemical Cell Summary

	Voltaic cells	**Electrolytic Cells**
Spontaneous?	Yes	No
potential, E	positive (+)	negative (–)
push/pull of electrons	Anode to cathode via spontaneous chemical reaction, supplying a voltage	Anode to cathode, requiring an external power source
anode	negative (–) electrons don't want to be here so…	positive (+) electrons want to be here but…
cathode	positive (+) they spontaneously go here (where they want to be)	negative (–) they are pushed to go here (where they don't want to be)

In **all** electrochemical cells, the electrons travel from the site of oxidation (anode) to the site of reduction (cathode).

The main difference is that voltaic cells are spontaneous cells, where the redox reaction drives the current. In an electrolytic cell, the redox reaction is non-spontaneous. Therefore, the push/pull of current is driven by an external power source.