

Thursday Feb. 25, 2016

Objective: The student will be able to (1) balance REDOX reactions using the half-cell method, (2) describe the components of a voltaic cell, (3) determine the electrical output of a voltaic cell.

Homework: Electrochemistry Intro. PS and E-chem PS #1

Warm - Up:

Provide balanced chemical reactions for the following reactions:

1. sulfur dioxide gas is bubbled into cold water
2. aqueous solutions of silver nitrate and 6.0 M ammonia are mixed.
3. Potassium chromate is added to an acidified solution of potassium iodide.

Real-World Applications

We are surrounded by an amazing array of portable electronic gadgets (cell phones, portable music players, laptops, and gaming devices.)

What would our world look like without the batteries which power these devices?

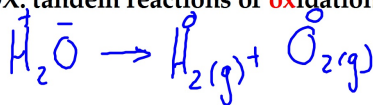
Current research attempts to find batteries with more power, faster charging ability, lighter weight, or cheaper price.



Electrochemical reactions happen here!

Introduction to Redox Terms

REDOX: tandem reactions of oxidation and reduction.



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.

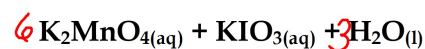
Hydrogen ion **gains** electrons. Hydrogen ion is the oxidizing agent and becomes **reduced**. The oxidation number of H **decreases** from +1 to 0.



OA / RA, OILRIG, show reactions, SYN of H₂O as non-example

Balancing REDOX Reactions

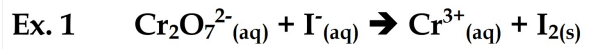
Let's try to balance this bad boy!



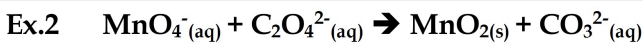
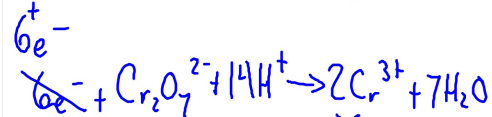
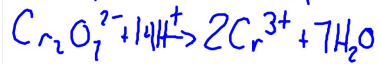
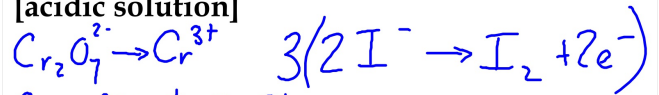
Balancing Redox Reactions Half-Cell Method

Steps:

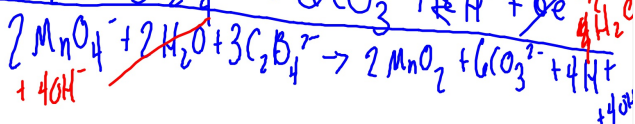
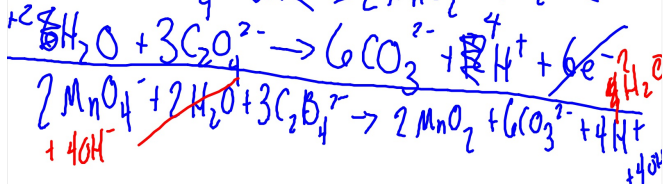
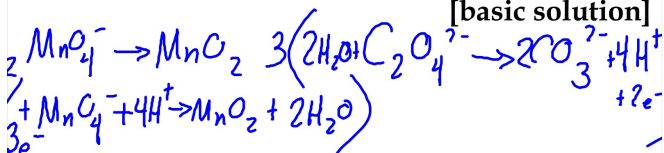
1. Divide the skeleton reaction into two half reactions.
2. Balance the atoms and charges:
 - atoms: $X \rightarrow O \rightarrow H$
 - charges: add electrons to balance charge
3. multiply the electrons so they are equal
4. add half reactions and include SOM
- [4b.] add OH^- (basic solution only)
5. Sight inspect (as a double check)



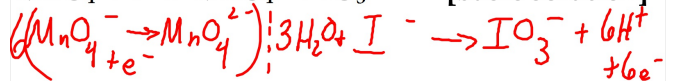
[acidic solution]



[basic solution]



Practice 1



galvanic **Overview of Electrochemical Cells**

VOLTAIC CELL	ELECTROLYTIC CELL
Energy is released from a spontaneous redox reaction.	Energy is absorbed to drive a nonspontaneous redox reaction.
System does work on surroundings.	Surroundings (power supply) do work on system (cell).
Oxidation half-reaction $X \rightarrow X^+ + e^-$	Oxidation half-reaction $A^- \rightarrow A + e^-$
Reduction half-reaction $e^- + Y^+ \rightarrow Y$	Reduction half-reaction $e^- + B^+ \rightarrow B$
Overall (cell) reaction $X + Y^+ \rightarrow X^+ + Y; \Delta G < 0$	Overall (cell) reaction $A^- + B^+ \rightarrow A + B; \Delta G > 0$

Notation for a Voltaic Cell

SC
25°C, 1 atm 1M

anode (1M) cathode (0.050)

$$\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu(s)}$$

Pt, graphite

- diff between single and double lines
- inclusion of concentrations (1M if not stated)
- diff. between inactive and active electrodes

Notation for a Voltaic Cell

Graphite $\mid \text{I}^-(\text{aq}) \mid \text{I}_2(\text{s}) \parallel \text{MnO}_4^-(\text{aq}), \text{H}^+(\text{aq}), \text{Mn}^{2+}(\text{aq}) \mid \text{graphite}$

inert electrodes are shown

commas are used to show substances in the same phase

Standard State of Redox Reactions

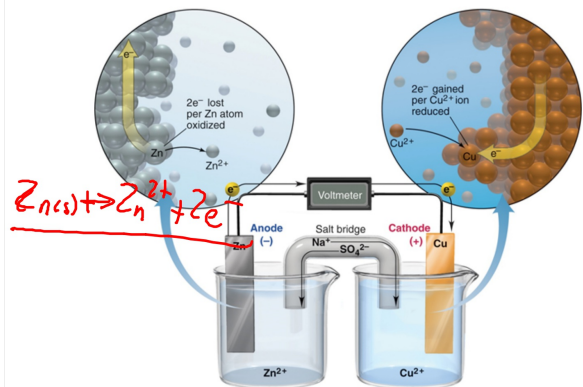
ΔG ΔG°

1M
25°C
1 atm

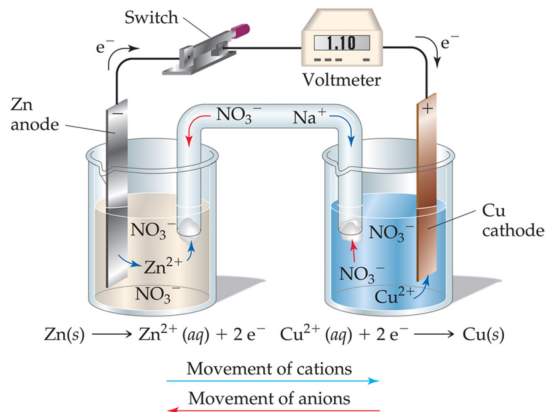
$$\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$$

Electrochemical reactions are applied thermodynamic systems. Discuss standard conditions (SC) again.

Components of a Voltaic Cell



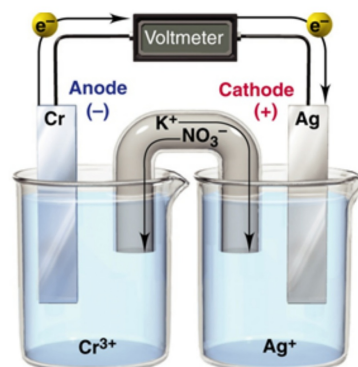
Movement of Charged Particles



Example Voltaic Cell

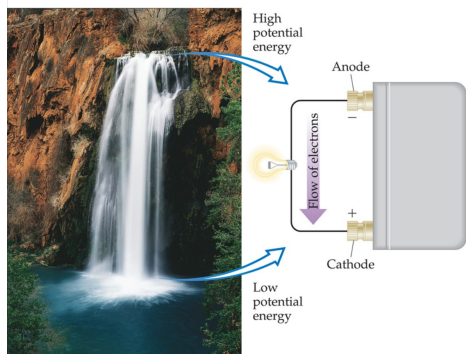
Draw a diagram, show balanced equations, and write the notation for a voltaic cell that consists of one half-cell with a chromium bar in a chromium(III) nitrate solution, another half-cell with a silver bar in a silver nitrate solution, and a potassium nitrate salt bridge. Measurement indicates that the chromium electrode is negative relative to the silver electrode.

Sample Problem Solution



Electromotive Force (emf)

Consider the movement of water in a waterfall:



Cell Potential: Output of a Voltaic Cell

cell potential (E_{cell} , voltage, emf): the difference in electrical potential between the two electrodes.

anode \rightarrow *cathode*
In a voltaic cell, electrons flow from the more negative electrode to the more positive

Sign Conventions:

$E_{\text{cell}} > 0$ spontaneous reaction

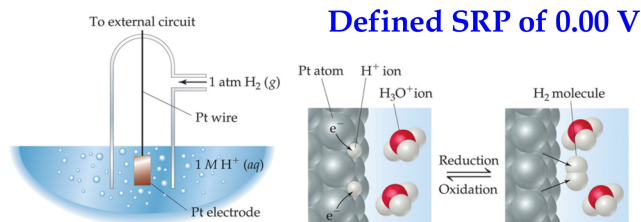
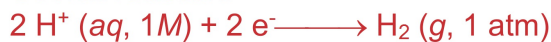
$E_{\text{cell}} = 0$ equilibrium, cell can do no more work

$E_{\text{cell}} < 0$ nonspontaneous reaction

$$1 \text{ V} = 1 \frac{\text{J}}{\text{C}}$$

Standard Hydrogen Electrode (SHE)

Overall reaction:



Defined SRP of 0.00 V

EQ: Why use platinum wire?

Standard Reduction Potentials

Potential (V)	Reduction Half-Reaction
+2.87	$\text{F}_2(\text{g}) + 2 \text{e}^- \longrightarrow 2 \text{F}^-(\text{aq})$
+1.51	$\text{MnO}_4^-(\text{aq}) + 8 \text{H}^+(\text{aq}) + 5 \text{e}^- \longrightarrow \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$
+1.36	$\text{Cl}_2(\text{g}) + 2 \text{e}^- \longrightarrow 2 \text{Cl}^-(\text{aq})$
+1.33	$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) + 6 \text{e}^- \longrightarrow 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\text{l})$
+1.23	$\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 \text{e}^- \longrightarrow 2 \text{H}_2\text{O}(\text{l})$
+1.06	$\text{Br}_2(\text{l}) + 2 \text{e}^- \longrightarrow 2 \text{Br}^-(\text{aq})$
+0.96	$\text{NO}_3^-(\text{aq}) + 4 \text{H}^+(\text{aq}) + 3 \text{e}^- \longrightarrow \text{NO}(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$
+0.80	$\text{Ag}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Ag}(\text{s})$
+0.77	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \longrightarrow \text{Fe}^{2+}(\text{aq})$
+0.68	$\text{O}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \longrightarrow \text{H}_2\text{O}_2(\text{aq})$
+0.59	$\text{MnO}_4^-(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) + 3 \text{e}^- \longrightarrow \text{MnO}_2(\text{s}) + 4 \text{OH}^-(\text{aq})$
+0.54	$\text{I}_2(\text{s}) + 2 \text{e}^- \longrightarrow 2 \text{I}^-(\text{aq})$
+0.40	$\text{O}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l}) + 4 \text{e}^- \longrightarrow 4 \text{OH}^-(\text{aq})$
+0.34	$\text{Cu}^{2+}(\text{aq}) + 2 \text{e}^- \longrightarrow \text{Cu}(\text{s})$
0 (defined)	$2 \text{H}^+(\text{aq}) + 2 \text{e}^- \longrightarrow \text{H}_2(\text{g})$
-0.28	$\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^- \longrightarrow \text{Ni}(\text{s})$
-0.44	$\text{Fe}^{2+}(\text{aq}) + 2 \text{e}^- \longrightarrow \text{Fe}(\text{s})$
-0.76	$\text{Zn}^{2+}(\text{aq}) + 2 \text{e}^- \longrightarrow \text{Zn}(\text{s})$
-0.83	$2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^- \longrightarrow \text{H}_2(\text{g}) + 2 \text{OH}^-(\text{aq})$
-1.66	$\text{Al}^{3+}(\text{aq}) + 3 \text{e}^- \longrightarrow \text{Al}(\text{s})$
-2.71	$\text{Na}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Na}(\text{s})$
-3.05	$\text{Li}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Li}(\text{s})$

A word about E°_{cell} calculations

OILRIG

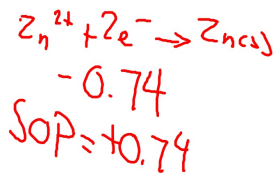
Two standard methods to calculate:

Method 1: (subtraction)

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

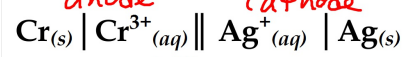
Method 2: (addition)

$$E^\circ_{\text{cell}} = \text{SOP} + \text{SRP}$$

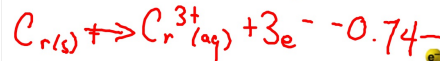


Example (from a voltaic cell previously drawn)

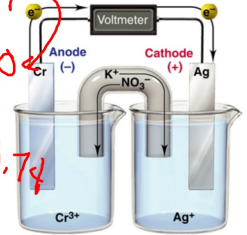
Calculate the E°_{cell} for the voltaic cell:



$$E^\circ_{\text{cell}} = 1.54\text{V}$$



$$E^\circ_{\text{cell}} = E^\circ_{\text{cath}} - E^\circ_{\text{an}} = 0.80 + 0.74$$



Strength of Oxidizers and Reducers

