

Chapter 20
Electrochemistry

- Electrochemistry deals with the relationships between electricity and chemical reactions.
- Oxidation-reduction (redox) reactions were introduced in Chapter 4
- Can be simple displacement reactions:

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

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

20.1 Oxidation-Reduction Reactions

- Redox reactions can also be more complex, with structural and composition changes as well as an exchange of electrons.
- $5\text{VO}^{2+}(\text{aq}) + \text{MnO}_4^-(\text{aq}) + \text{H}_2\text{O(l)} \rightarrow 5\text{VO}_2^+(\text{aq}) + \text{Mn}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq})$
- $5\text{Fe}^{2+}(\text{aq}) + \text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) \rightarrow 5\text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)}$

Oxidation and Reduction

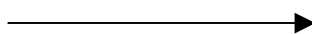
You can't have one without the other!

Oxidation

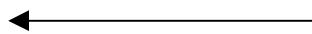
Decrease in number of electrons

(loss of electrons)

Increase in oxidation number



Oxidation number: -3 -2 -1 0 +1 +2 +3



Increase in number of electrons

(gain of electrons)

Decrease in oxidation number

Reduction

Half-Reactions

- Oxidation-reduction reactions can be written as separate oxidation and reduction reactions, called *half-reactions*.
- Oxidation:

$$\text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O(l)} \rightarrow \text{VO}_2^+(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^-$$
- VO^{2+} is called the *reducing agent* (or *reductant*), because it causes the reduction of another substance; the reducing agent is oxidized in the process
- Reduction:

$$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)}$$
- MnO_4^- is called the *oxidizing agent* (or *oxidant*), because it causes the oxidation of another substance; the oxidizing agent is reduced in the process
- Oxidation-reduction can be considered to be the competition between two substances for electrons. The one with the greater attraction for additional electrons becomes the oxidizing agent; the one with the least is the reducing agent.

Spontaneous Oxidation-Reduction Reactions

- If we place a metal in a solution of another metal ion, sometimes we get metal deposition, sometimes not. How do we decide?

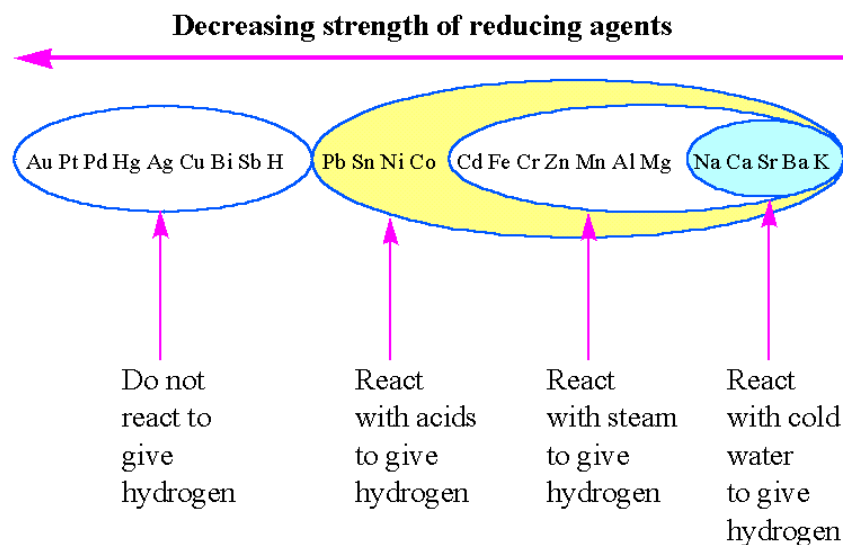
$$\text{Zn} + \text{Sn}^{2+} \rightarrow \text{Zn}^{2+} + \text{Sn}$$

$$\text{Sn} + \text{Zn}^{2+} \rightarrow \text{no rxn}$$
- How do we know which substances will act as oxidizing agents or as reducing agents?
- An activity series gives the relative activity of substances as oxidizing or reducing agents. Review from Chapter 4.
- Determine an activity series in several ways:
 - activity in displacing H_2 from water
 - activity in displacing metals from metal ion solutions (more active metal displaces a less active metal from solution)
- generation of an electrochemical potential or voltage

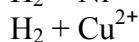
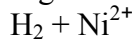
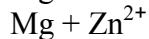
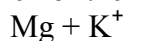
Spontaneous Redox

- Na displaces H_2 from H_2O
- Zn displaces Pb from a solution of Pb^{2+}
- $\text{Pb} + 2\text{Ag}^+ \rightarrow \text{Pb}^{2+} + 2\text{Ag}$
- $\text{Ag} + \text{Pb}^{2+} \rightarrow \text{no rxn}$
- The reverse reactions can be made to occur with electric current, but they are not spontaneous.
- General rule, using an activity series:

stronger	stronger	weaker	weaker
oxidizing	+ reducing	→ reducing	+ oxidizing
agent	agent	agent	agent
- Depending on the relative strengths, the reaction can go to completion, or reach a state of equilibrium.

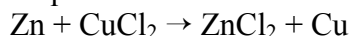
Activity Series

- Predict which of the following combinations will undergo an oxidation-reduction reaction:

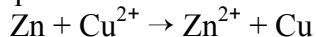


20.2 Balancing Oxidation-Reduction Equations

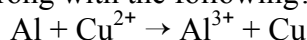
- Some redox equations can be balanced by inspection, just like other types of reactions.



- Net ionic equation:



- Need to balance atoms and charge
- What is wrong with the following?



Balancing Equations

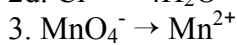
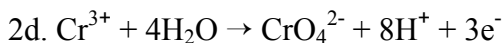
- $5\text{Cr}^{3+} + 3\text{MnO}_4^- + 8\text{H}_2\text{O} \rightarrow 5\text{CrO}_4^{2-} + 3\text{Mn}^{2+} + 16\text{H}^+$
- How do we balance an equation as complex as this?
- Two methods:
 - half-reaction method
 - oxidation number change method
- Will focus on the half-reaction method since it is useful in more circumstances.

Half-Reaction Method

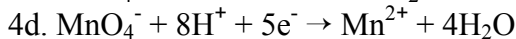
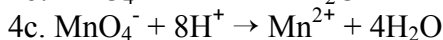
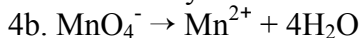
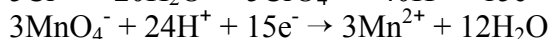
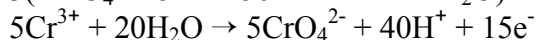
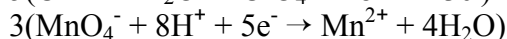
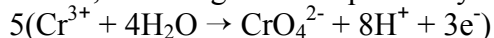
- Write an unbalanced half-reaction for either oxidation or reduction.
- Balance the half-reaction:
 - Balance all atoms other than H and O.
 - Balance O by adding H_2O to the equation.
 - Balance H by adding H^+ to the equation.
 - Balance ionic charges by adding electrons to the equation.
 - If in basic solution, add OH^- to each side of the equation to neutralize all the H^+ .
 - Cancel any H_2O occurring on both sides.
- Write an unbalanced half-reaction for the other process.
- Balance by the same procedure.
- Equalize the number of electrons lost and gained by multiplying each coefficient in each half-reaction by the appropriate constant.
- Add the two half-reactions and cancel equal amounts of anything occurring on both sides.
- Make a final check of atom and charge balances.

Balancing Equations

- $\text{Cr}^{3+} + \text{MnO}_4^- \rightarrow \text{CrO}_4^{2-} + \text{Mn}^{2+}$
- How do we balance this equation in acidic solution?
 - $\text{Cr}^{3+} \rightarrow \text{CrO}_4^{2-}$
 - Cr already balanced
 - $\text{Cr}^{3+} + 4\text{H}_2\text{O} \rightarrow \text{CrO}_4^{2-}$
 - $\text{Cr}^{3+} + 4\text{H}_2\text{O} \rightarrow \text{CrO}_4^{2-} + 8\text{H}^+$



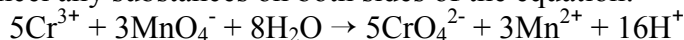
4a. Mn already balanced

5. 3 e⁻ lost, but 5 e⁻ gained. Equalize by multiplying half-reactions:

6. Add the equations:



Cancel any substances on both sides of the equation:



7. Check atom and charge balance:

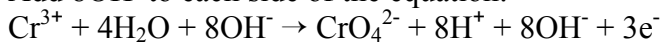
5 Cr, 3 Mn, 20 O, 16 H, 12 + charges on each side

Balancing Equations in Basic Solution

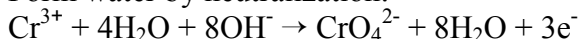
- Neutralize any H⁺ with OH⁻ to form H₂O and add the same number of OH⁻ to the other side of the equation.



- Add 8OH⁻ to each side of the equation:



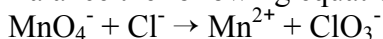
- Form water by neutralization:



- Cancel any water occurring on both sides:



- Balance the following equation in acidic solution:



20.3 Voltaic Cells

- Voltaic (or galvanic) cells: spontaneous redox → electricity (or electrical work)
- Electrolysis cells: electricity → redox
- To produce electricity, we must direct the electron flow through an external circuit. We cannot have direct redox.
- Daniell cell: Zn - Cu

Voltaic Cells

- To produce electricity, we need:
 - Isolated half-reactions, using half-cells
 - Conductive solids (electrodes) connected by external circuits
 - May consist of a reactant/product or be an inert substance such as platinum or graphite
 - Anode: oxidation half-reaction

- Cathode: reduction half-reaction
- Externally, the anode has the negative charge; internally, it has a positive charge)
- Anions flow towards the anode; cations move away from it and towards the cathode.
- Two half-cells must be connected to pass ions
- Salt bridge or porous glass

20.4 Cell EMF

- If the half-reactions are carried out separately (but coupled), we find they generate an electrical current characterized by a voltage (or electromotive force or electrical potential). This is the force pushing electrons through the circuit.
- The voltage produced by a voltaic cell is called the cell potential, E°_{cell} (also the reaction potential, E°_{rxn} , when the half-reactions are not separated)
- The voltage is called the standard electromotive force (emf), E° , under standard conditions.

Electromotive Force

- $E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$
- Values are determined relative to the standard hydrogen electrode, with $E^\circ = 0 \text{ V}$
- $E^\circ_{\text{red}} = \text{emf}$ for the reduction half-rxn
- $E^\circ_{\text{ox}} = \text{emf}$ for the oxidation half-rxn
- $E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$
- We can only measure E°_{cell} , so we must define a standard reference to get E°_{ox} and E°_{red} .

E° Values

- Reference for E° is the standard hydrogen electrode, using the reaction:
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad E^\circ = 0$
 1 M 1 atm (std conditions)

Cell Potential

- Other values are then measured from E°_{cell} with the standard hydrogen electrode or with other known half-cells.
 - $\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^- \quad E^\circ = 0.22 \text{ V}$
 - $\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^- \quad E^\circ = 0.2802 \text{ V}$
- Half-cells such as these are used as reference electrodes. The Ag/AgCl electrode, along with a glass electrode, is used in a pH meter.

E° Values

- Then get other half-reaction potentials from measured E°_{cell} values.
 - $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2 \quad E^\circ_{\text{cell}} = 0.76 \text{ V}$
 - $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^- \quad E^\circ_{\text{ox}} = 0.76 \text{ V}$ since $E^\circ_{\text{red}} = 0.00 \text{ V}$

Reference to Hydrogen Electrode

- The E° of the hydrogen electrode is defined as 0.00 V. What would be the values of E° for other half-reactions if this were defined as 1.00 V?
- The E° of the hydrogen electrode is defined as 0.00 V. What would be the values of E° for other half-reactions if this were defined as 1.00 V?

- What would be the E° values of overall redox reactions if the reference were defined as 1.00 V?
- What would be the E° values of overall redox reactions if the reference were defined as 1.00 V?

E° Values

- If we reverse a half-reaction, we change the sign of E° .

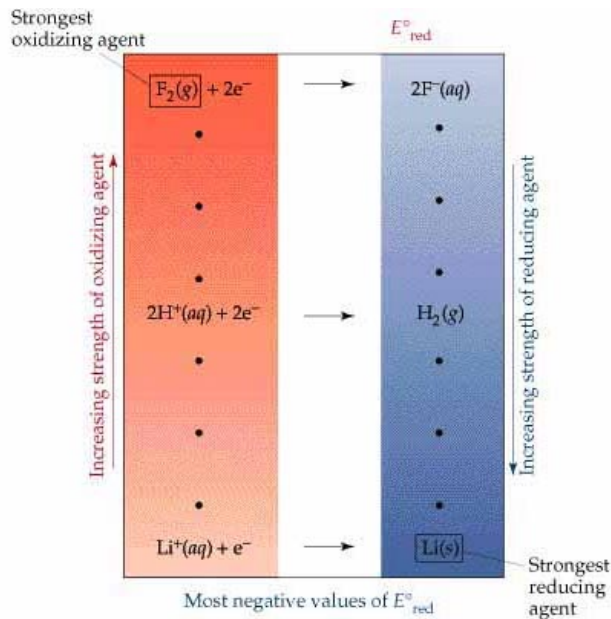
$$\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn} \quad E^\circ_{\text{red}} = -0.76 \text{ V}$$
- Measure a value of $E^\circ_{\text{cell}} = 0.63 \text{ V}$ for the following reaction.

$$\text{Zn} + \text{Pb}^{2+} \rightarrow \text{Zn}^{2+} + \text{Pb}$$
- What is the E°_{red} of

$$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb} \text{ ?}$$
- $$\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^- \quad E^\circ_{\text{ox}} = 0.76 \text{ V}$$
- $$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb} \quad E^\circ_{\text{red}} = ?$$
- $$\text{Zn} + \text{Pb}^{2+} \rightarrow \text{Zn}^{2+} + \text{Pb} \quad E^\circ_{\text{cell}} = 0.63 \text{ V}$$
- $$E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$$
- $$0.63 \text{ V} = E^\circ_{\text{red}} + 0.76 \text{ V}$$
- $$E^\circ_{\text{red}} = 0.63 \text{ V} - 0.76 \text{ V} = -0.13 \text{ V}$$
- Values determined in this way are listed in Table 20.1 and Appendix E.

Reduction Potentials

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



20.5 Spontaneity of Redox Reactions

- A redox reaction is spontaneous if $E^\circ_{\text{rxn}} > 0$.
- To determine spontaneity, add the two half-reactions and their E° values to see if E°_{rxn} has a positive value.
- What will happen if we place a piece of Zn and a piece of Cu in a solution that contains a mixture of Zn^{2+} and Cu^{2+} ?

- Two possibilities:
 - $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$
 - $\text{Cu} + \text{Zn}^{2+} \rightarrow \text{Cu}^{2+} + \text{Zn}$

Spontaneous Redox

- Two possible reduction half-reactions:

$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	$E^\circ_{\text{red}} = -0.76 \text{ V}$
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	$E^\circ_{\text{red}} = 0.34 \text{ V}$
- Two possible oxidation half-reactions:

$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$	$E^\circ_{\text{ox}} = 0.76 \text{ V}$
$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$	$E^\circ_{\text{ox}} = -0.34 \text{ V}$
- Two ways to combine them:

$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	$E^\circ_{\text{red}} = -0.76 \text{ V}$
$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$	$E^\circ_{\text{ox}} = -0.34 \text{ V}$
$\text{Cu} + \text{Zn}^{2+} \rightarrow \text{Cu}^{2+} + \text{Zn}$	$E^\circ_{\text{rxn}} = -1.10 \text{ V}$
- Two ways to combine them:

$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	$E^\circ_{\text{red}} = 0.34 \text{ V}$
$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$	$E^\circ_{\text{ox}} = 0.76 \text{ V}$
$\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$	$E^\circ_{\text{rxn}} = 1.10 \text{ V}$
- The combination with $E^\circ_{\text{rxn}} > 0$ is spontaneous.

Predicting Reactions

- Vanadium Reduction Potentials:

$\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$	$E^\circ_{\text{red}} = +1.00 \text{ V}$
$\text{VO}^{2+} + 2\text{H}^+ + \text{e}^- \rightarrow \text{V}^{3+} + \text{H}_2\text{O}$	$E^\circ_{\text{red}} = +0.36 \text{ V}$
$\text{V}^{3+} + \text{e}^- \rightarrow \text{V}^{2+}$	$E^\circ_{\text{red}} = -0.26 \text{ V}$
$\text{V}^{2+} + 2\text{e}^- \rightarrow \text{V}$	$E^\circ_{\text{red}} = -1.20 \text{ V}$
- Chromium Reduction Potentials:

$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	$E^\circ_{\text{red}} = +1.33 \text{ V}$
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Cr}^{2+}(\text{aq})$	$E^\circ_{\text{red}} = -0.41 \text{ V}$
$\text{Cr}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cr}(\text{s})$	$E^\circ_{\text{red}} = -0.91 \text{ V}$
- Manganese Reduction Potentials:

$\text{MnO}_4^-(\text{aq}) + \text{e}^- \rightarrow \text{MnO}_4^{2-}(\text{aq})$	$E^\circ_{\text{red}} = +0.56 \text{ V}$
$\text{MnO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	$E^\circ_{\text{red}} = +2.26 \text{ V}$
$\text{MnO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{Mn}^{3+}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$	$E^\circ_{\text{red}} = +0.95 \text{ V}$
$\text{Mn}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq})$	$E^\circ_{\text{red}} = +1.51 \text{ V}$
$\text{Mn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mn}(\text{s})$	$E^\circ_{\text{red}} = -1.18 \text{ V}$

Predicting Reactions

- Add excess Cr^{2+} to VO_2^+ . What is the product?

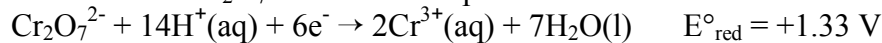
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Cr}^{2+}(\text{aq})$	$E^\circ_{\text{red}} = -0.41 \text{ V}$
$\text{Cr}^{2+}(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{e}^-$	$E^\circ_{\text{ox}} = 0.41 \text{ V}$
$\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$	$E^\circ_{\text{red}} = +1.00 \text{ V}$
$\text{VO}^{2+} + 2\text{H}^+ + \text{e}^- \rightarrow \text{V}^{3+} + \text{H}_2\text{O}$	$E^\circ_{\text{red}} = +0.36 \text{ V}$



- Product is V^{2+} , because successive $E^{\circ}_{\text{rxn}} = 1.41, 0.77, 0.15, -0.79 \text{ V}$

Predicting Reactions

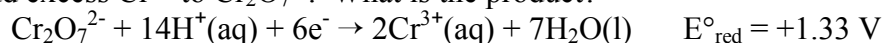
- Add excess Cr to $\text{Cr}_2\text{O}_7^{2-}$. What is the product?



- Product is Cr^{2+} ; the successive values of E°_{rxn} are 2.24 and 0.50 V.

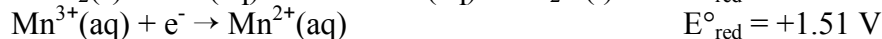
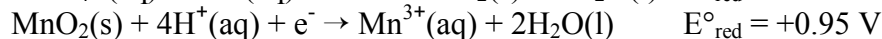
Predicting Reactions

- Add excess Cr^{2+} to $\text{Cr}_2\text{O}_7^{2-}$. What is the product?



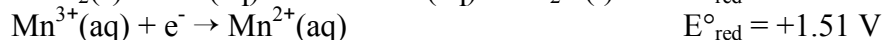
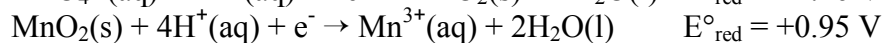
Predicting Reactions

- Add excess V^{2+} to MnO_4^{-} . What is the product?



Predicting Reactions

- Add Mn^{2+} to MnO_4^{-} . What is the product?



- Predict no reaction, but reaction actually occurs to form Mn^{3+} or MnO_2 . We will deal with this discrepancy in the next section. It arises from the fact that there are more half-reactions that could be considered, which arise from combinations of these half-reactions.

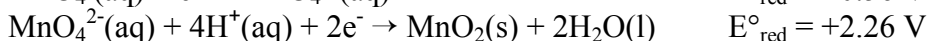
Stability in Aqueous Systems

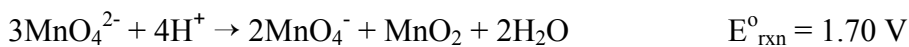
- Disproportionation Reactions

A substance reacts with itself to form new substances with higher and lower oxidation numbers.

- No examples with V or Cr.

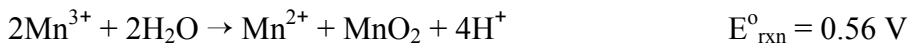
- MnO_4^{2-} will disproportionate:





Stability in Aqueous Systems

- Mn^{3+} will disproportionate:
 $\text{MnO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{Mn}^{3+}(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \quad E^\circ_{\text{red}} = +0.95 \text{ V}$
 $\text{Mn}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) \quad E^\circ_{\text{red}} = +1.51 \text{ V}$



Stability in Aqueous Systems

- Reaction with Water
 Reduce hydronium ion to release hydrogen gas:
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad E^\circ_{\text{red}} = 0.000 \text{ V}$
- Any substance with $E^\circ_{\text{ox}} > 0$ will reduce H^+ to H_2
- Examples are V, V^{2+} , Cr, Cr^{2+} , Mn
- The ions will react, but tend to react only very slowly. There seems to be a kinetic factor that results in a fast reaction only if $E^\circ_{\text{rxn}} > 0.4\text{-}0.5 \text{ V}$ (called an overvoltage).
- Reaction with Water
 Oxidize water to release oxygen gas:
 $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \quad E^\circ_{\text{ox}} = -1.23 \text{ V}$
- Any $E^\circ_{\text{red}} > 1.23 \text{ V}$ will result in production of O_2 . Generally need $E^\circ_{\text{rxn}} > 0.4\text{-}0.5 \text{ V}$ for fast reaction.
- Examples are $\text{Cr}_2\text{O}_7^{2-}$ (very slow), MnO_4^{2-} (disproportionates faster), Mn^{3+} (disproportionates faster)

Stability in Aqueous Systems

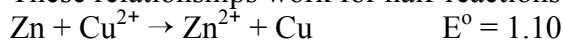
- Oxidation by O_2 in Air
 $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l}) \quad E^\circ_{\text{red}} = 1.23 \text{ V}$
 - Any $E^\circ_{\text{ox}} > -1.23 \text{ V}$ will result in oxidation by air. Many substances fall into this category ($E^\circ_{\text{rxn}} > 0.4\text{-}0.5 \text{ V}$ for fast reaction).
- | | | |
|------------------------------|----------------------|--|
| V | Cr | Mn |
| V^{2+} | Cr^{2+} | not Mn^{2+} |
| V^{3+} | not Cr^{3+} | Mn^{3+} |
| VO^{2+} (very slow) | | not MnO_2 |
| | | MnO_4^{2-} (disproportionates faster) |

EMF and Free Energy Change

- We have seen three criteria for spontaneity:
 - $E^\circ > 0$
 - $\Delta G^\circ < 0 \quad 1 \text{ J} = 1 \text{ coul} \times 1 \text{ V}$
 - $K \gg 1$
- These criteria are related:
 - $\Delta G^\circ = -RT \ln K$
 - $\Delta G^\circ = -nFE^\circ, \Delta G = -nFE$
 where n = number of e^- transferred and F = Faraday constant (charge on 1 mole e^-)
- $1 F = 96,500 \text{ coul/mol } \text{e}^- = 96,500 \text{ J/V mol } \text{e}^-$

Thermodynamics

- These relationships work for half-reactions or complete redox reactions.



$$n = 2$$

$$\Delta G^\circ = -2 \text{ mol } e^- \times 96500 \text{ J/V mol } e^- \times 1.10 \text{ V}$$

$$\Delta G^\circ = -212,300 \text{ J} = -212.3 \text{ kJ}$$

- ΔG° depends on the number of moles, but E° does not

Voltage and Moles

- Note that different size alkaline cells all deliver the same voltage, in spite of different number of moles of reactants.

Thermodynamics

- We can add E°_{ox} to E°_{red} to give E°_{cell} or E°_{rxn} in the same way that we can add half-reactions to give an overall reaction.



$$\Delta G^\circ_{\text{ox}} = -2 \times 96500 \times 0.44 = -84900 \text{ J}$$



$$\Delta G^\circ_{\text{red}} = -2 \times 96500 \times 1.36 = -262500 \text{ J}$$



$$\Delta G^\circ = -2 \times 96500 \times 1.80 = -347400 \text{ J}$$

- $\Delta G^\circ_{\text{rxn}} = \Delta G^\circ_{\text{ox}} + \Delta G^\circ_{\text{red}}$

$$= -84.9 + -262.5 = -347.4 \text{ kJ}$$

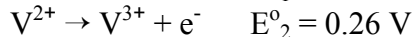
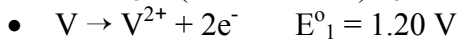
- From Chapter 19, we know that ΔG° values are additive when we add reactions.
- E° s are additive when we add half-reactions to give a complete reaction because the value of n is the same for the half-reactions and the complete reaction.
- E° s are not additive when adding two half-reactions to give a third half-reaction because the value of n is not constant.
- We can add ΔG° under all circumstances:

$$\Delta G^\circ_3 = \Delta G^\circ_1 + \Delta G^\circ_2$$

$$-n_3 F E^\circ_3 = -n_1 F E^\circ_1 - n_2 F E^\circ_2$$

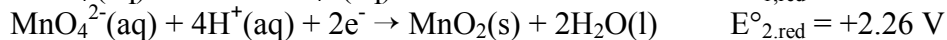
$$n_3 E^\circ_3 = n_1 E^\circ_1 + n_2 E^\circ_2$$

$$E^\circ_3 = (n_1 E^\circ_1 + n_2 E^\circ_2) / n_3$$



$$E^\circ_3 = (2 \times 1.20 + 1 \times 0.26) / 3 = 0.887 \text{ V}$$

- Given the first two half-reactions, what is the value of E°_{red} for the third?



- $E^\circ_{3,\text{red}} = 1.69 \text{ V}$

- This is why we could not make correct predictions for this system earlier.

Thermodynamics

- We can calculate K_{eq} from E° :

$$\Delta G^\circ = -nFE^\circ = -RT \ln K$$

$$E^\circ = (RT/nF) \ln K = 2.303 (RT/nF) \log K$$
- At 25°C, $2.303 RT/F = 0.05916$

$$E^\circ = (0.05916/n) \log K \text{ at } 25^\circ\text{C}$$
- Thus, we can measure E° for a redox reaction and then calculate the equilibrium constant for that reaction.

20.6 Effect of Concentration on Cell EMF

- So far, we have been using standard state conditions, but we don't always have 1 M solutions. We can correct E° to E by using the Nernst equation.

$$\Delta G = \Delta G^\circ + RT \ln Q$$
 But, $\Delta G = -nFE$ and $\Delta G^\circ = -nFE^\circ$, so

$$-nFE = -nFE^\circ + 2.303 RT \log Q$$

$$E = E^\circ - (2.303 RT/nF) \log Q$$
- At 25°C, $E = E^\circ - (0.05916/n) \log Q$

Nernst Equation

- At 25°C, $E = E^\circ - (0.05916/n) \log Q$
- $E = E^\circ$ if $Q = 1$
- When the system reaches equilibrium, $Q = K$, and $E = 0$, because $E^\circ = (0.05916/n) \log K$, and the cell has "run down".
- Consider the Zn/Cu^{2+} reaction if more Cu^{2+} is added to the cell. The voltage becomes greater than 1.10 V.
- What is E of the Zn/Cu^{2+} reaction if $[\text{Cu}^{2+}] = 0.010 \text{ M}$ and $[\text{Zn}^{2+}] = 1.99 \text{ M}$? Note that this corresponds to starting with standard conditions and changing to 99% completion of reaction. $E^\circ = 1.10 \text{ V}$ (with $[\text{Cu}^{2+}] = [\text{Zn}^{2+}] = 1.00 \text{ M}$)
- $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$
 For this reaction, $n = 2$.

$$Q = [\text{Zn}^{2+}]/[\text{Cu}^{2+}]$$

$$E = E^\circ - (0.05916/n) \log Q$$

$$E = 1.10 \text{ V} - (0.05916/2) \log (1.99/0.010)$$

$$E = 1.10 \text{ V} - (0.05916/2) \log 199$$

$$E = 1.10 - 0.068 = 1.03 \text{ V}$$
- For 99.9% reaction (1.999 M Zn^{2+} , 0.001 M Cu^{2+}), $E = 1.10 - 0.098 = 1.00 \text{ V}$
- For 99.99% reaction, 1.9999 M Zn^{2+} , 0.0001 M Cu^{2+}), $E = 1.10 - 0.127 = 0.97 \text{ V}$

Concentration Cells

- We can generate a voltage with a cell that contains the same materials in the cathode and anode compartments, but at different concentrations.

20.7 Batteries (Voltaic Cells)

- Voltaic cells are used as portable electricity sources.
- one unit = cell

- several cells = battery
- Commercial cells:
 - Anode: usually a metal that can be oxidized
 - Cathode: usually an inert conductor surrounded by a substance with a high oxidation number that can be reduced
 - Electrolyte: aqueous solution or moist paste for ion movement
 - Electrolytes: acidic, alkaline, organic liquid, molten salt, solid state
- Primary cells: one use, then discard
- Secondary cells: rechargeable, reusable
- Consider some commercial cells and batteries
 - What are the anode and cathode?
 - What is used as an electrolyte?
 - What half-reactions are used?
 - What voltage is generated?

LeClanche Dry Cell

- Used in flashlight batteries.
- Electrolyte is NH_4Cl
- $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ $E^\circ = 0.763 \text{ V}$
- $2\text{MnO}_2 + 2\text{NH}_4^+ + 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{NH}_4\text{OH} + 2\text{MnO}(\text{OH})$ $E^\circ = 0.5 \text{ V}$
- $\text{Zn} + 2\text{MnO}_2 + 2\text{NH}_4^+ + 2\text{H}_2\text{O} \rightarrow \text{Zn}^{2+} + 2\text{NH}_4\text{OH} + 2\text{MnO}(\text{OH})$ $E^\circ = 1.26 \text{ V}$
- The measured voltage is actually 1.5 V, because concentrations are higher than standard state concentrations.

Alkaline Dry Cell

- Electrolyte is KOH
- $\text{Zn} + 2\text{OH}^- \rightarrow \text{ZnO} + \text{H}_2\text{O} + 2\text{e}^-$
- $2\text{MnO}_2 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{OH}^- + \text{Mn}_2\text{O}_3$
- $\text{Zn} + 2\text{MnO}_2 \rightarrow \text{ZnO} + \text{Mn}_2\text{O}_3$ $E = 1.5 \text{ V}$
- Reactions are similar to the acidic dry cell
- Can be recharged a few times

Lead Storage Battery

- Used in automobiles
- $\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$ $E^\circ = 1.69 \text{ V}$
- $\text{Pb} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4 + 2\text{e}^-$ $E^\circ = 0.35 \text{ V}$
- $\text{Pb} + \text{PbO}_2 + 2\text{SO}_4^{2-} + 4\text{H}^+ \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$ $E^\circ = 2.04 \text{ V}$
- With 6 M H_2SO_4 , the cell produces 2.0 V; the voltage decreases as the cell reactions occur.
- Six cells connected in series makes up the common 12 V car battery.
- Can be recharged:

$$2\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow \text{Pb} + \text{PbO}_2 + 2\text{SO}_4^{2-} + 4\text{H}^+$$
- Must pass 2 V per cell back through the battery to recharge.
- Rapid recharging can electrolyze the acid, producing hydrogen gas, which might explode, or which can dislodge the lead oxide or sulfate, which shortens the battery life.

NiCad Cell

- Used in rechargeable power tools.
- Rechargeable because the reaction products stick to the electrodes.
- $\text{Cd} + 2\text{OH}^- \rightarrow \text{Cd}(\text{OH})_2 + 2\text{e}^-$ $E^\circ = 0.81 \text{ V}$
- $\text{NiO}_2 \text{ (on Ni)} + 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Ni}(\text{OH})_2 + 2\text{OH}^-$ $E^\circ = 0.49 \text{ V}$
- $\text{Cd} + \text{NiO}_2 + 2\text{H}_2\text{O} \rightarrow \text{Cd}(\text{OH})_2 + \text{Ni}(\text{OH})_2$ $E^\circ = 1.30 \text{ V}$

Silver Cell

- Used in cameras, watches, hearing aids
- Electrolyte is 20-40% KOH
- $\text{Zn} + 2\text{OH}^- \rightarrow \text{ZnO} + \text{H}_2\text{O} + 2\text{e}^-$ $E^\circ = 1.25 \text{ V}$
- $2\text{AgO} + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Ag}_2\text{O} + 2\text{OH}^-$ $E^\circ = 0.61 \text{ V}$
- $\text{Zn} + 2\text{AgO} + \text{H}_2\text{O} \rightarrow \text{Zn}(\text{OH})_2 + \text{Ag}_2\text{O}$ $E^\circ = 1.86 \text{ V}$
- Rechargeable if don't pass this point.
- After the AgO is depleted:
- $\text{Zn} + \text{Ag}_2\text{O} + \text{H}_2\text{O} \rightarrow \text{Zn}(\text{OH})_2 + 2\text{Ag}$ $E^\circ = 1.58 \text{ V}$

Fuel Cell

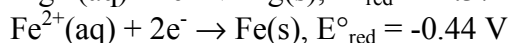
- Used for space travel, and in new hydrogen fuel automobiles
- Reactants are stored external to the cell and introduced to the electrodes as they are needed.
- The reactants are usually gaseous, such as O_2 and H_2 , or O_2 and CH_4 , or O_2 and NH_3 .
- $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$ $E^\circ = 0.40 \text{ V}$
- $2\text{H}_2 + 4\text{OH}^- \rightarrow 4\text{H}_2\text{O} + 4\text{e}^-$ $E^\circ = 0.83 \text{ V}$
- $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ $E^\circ = 1.23 \text{ V}$
- Recent development uses butane

20.8 Corrosion

- Corrosion of Iron
- Since $E^\circ_{\text{red}}(\text{Fe}^{2+}) < E^\circ_{\text{red}}(\text{O}_2)$ iron can be oxidized by oxygen.
- Cathode: $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$.
- Anode: $\text{Fe}(\text{s}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$.
- Dissolved oxygen in water usually causes the oxidation of iron.
- Fe^{2+} initially formed can be further oxidized to Fe^{3+} which forms rust, $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}(\text{s})$.
- Oxidation occurs at the site with the greatest concentration of O_2 .

Preventing the Corrosion of Iron

- Corrosion can be prevented by coating the iron with paint or another metal.
- Galvanized iron is coated with a thin layer of zinc.
- Zinc protects the iron since Zn is the anode and Fe the cathode:
 $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s}), E^\circ_{\text{red}} = -0.76 \text{ V}$
 $\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s}), E^\circ_{\text{red}} = -0.44 \text{ V}$
- With the above standard reduction potentials, Zn is easier to oxidize than Fe.
- To protect underground pipelines, a sacrificial anode is added.
- The water pipe is turned into the cathode and an active metal is used as the anode.
- Often, Mg is used as the sacrificial anode:



- Also used in water heaters, on gasoline storage tanks, and on ship hulls

20.9 Electrolysis

- Since chemical oxidation-reduction involves the transfer of electrons from one substance to another, it should be possible to harness the flow of electrons to produce electricity. We do this with voltaic cells.
- Electricity can also be used to cause non-spontaneous chemical reactions. This process is called electrolysis.

Electrolytic Processes with Metals

- A variety of metals can be prepared by electrolysis, if a cheap source of electricity is available. In addition, some metals* are purified by electrolysis.

aluminum	cadmium
calcium	copper*
gold*	lead*
magnesium	sodium
zinc	

Purification of Copper

- Recovered from its ores by chemical reduction.
- Purified by electrolysis.
- Recover impurities:
 - Mo (25%)
 - Se (93%)
 - Te (96%)
 - Au (32%)
 - Ag (28%)

Production of Aluminum

- Recovered as $\text{Al}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ from clays and bauxite.
- The oxide is difficult to reduce.
- The metal forms a protective coating of the oxide.
- Al is very active but can be used in air because of the oxide coating.
- The oxide coating can be thickened and colored by anodizing.
- Electrolysis of Al_2O_3 (melts at 2045°C) dissolved in cryolite, Na_3AlF_6 (melts at 1000°C).
- Uses 5% of U.S. electricity production

Electrolysis

- Electrolysis is used for isolating active elements, purifying metals, and electroplating.
- Pure compounds: H_2O , molten salts
- Use inert electrodes in the liquid and pass electricity through the system
- The negative electrode (cathode) attracts cations; reduction occurs.
- The positive electrode (anode) attracts anions; oxidation occurs.

Electrolysis of NaCl

- Cathode:
 $\text{Na}^+(\text{l}) + \text{e}^- \rightarrow \text{Na}(\text{l}) \quad E^\circ = -2.71 \text{ V}$
- Anode:
 $2\text{Cl}^-(\text{l}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^- \quad E^\circ = -1.36 \text{ V}$
- $2\text{Na}^+(\text{l}) + 2\text{Cl}^-(\text{l}) \rightarrow 2\text{Na}(\text{l}) + \text{Cl}_2(\text{g}) \quad E^\circ = -4.08 \text{ V}$
- Must supply at least 4.08 V to electrolyze molten sodium chloride.
- NaCl melts at 804°C, where Na vaporizes and burns.
- Lower the temperature by adding CaCl₂. (Why does this work?)
- Na reacts with Cl₂, even at room temperature.
- Commercial operations use a Downs Cell.
- How does the Downs Cell solve the problem of reaction between Na and Cl₂?

Electrolysis of H₂O

- Cathode:
 $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-$
 $E^\circ = -0.82 \text{ V at pH 14, } -0.41 \text{ V at pH 7, } 0.00 \text{ V at pH 0}$
- Anode:
 $2\text{H}_2\text{O} \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+ + 4\text{e}^-$
 $E^\circ = -0.41 \text{ V at pH 14, } -0.82 \text{ V at pH 7, } -1.23 \text{ V at pH 0}$
- $2\text{H}_2\text{O} \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
 $E^\circ = -1.23 \text{ V at any pH}$
- Must supply at least 1.23 V to electrolyze water. Add an electrolyte to increase electrical conductivity

Electrolysis of Aqueous Solutions

- Products depend on whether it is easier to oxidize or reduce the dissolved ions or water.
- Consider a solution of VCl₃ under standard conditions at pH 7.
- Cathode:
 $\text{V}^{3+} + \text{e}^- \rightarrow \text{V}^{2+} \quad E^\circ = -0.26 \text{ V}$
 $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^- \quad E^\circ = -0.41 \text{ V at pH 7}$
- Because of lower voltage, will reduce V³⁺, not H₂O.
- Anode:
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^- \quad E^\circ = -1.36 \text{ V}$
 $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^- \quad E^\circ = -0.82 \text{ V at pH 7}$
- Because of lower voltage, will oxidize H₂O, not Cl⁻.
- Products are V²⁺ and O₂.
- What are the products of electrolysis of VBr₃ at pH 7?
- Cathode:
 $\text{V}^{3+} + \text{e}^- \rightarrow \text{V}^{2+} \quad E^\circ = -0.26 \text{ V}$
 $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^- \quad E^\circ = -0.41 \text{ V at pH 7}$

- Anode:
 $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$ $E^\circ = -1.07 \text{ V}$
 $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ $E^\circ = -0.82 \text{ V at pH 7}$
- What are the products of electrolysis of VI_3 at pH 7?
- Cathode:
 $\text{V}^{3+} + \text{e}^- \rightarrow \text{V}^{2+}$ $E^\circ = -0.26 \text{ V}$
 $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$ $E^\circ = -0.41 \text{ V at pH 7}$
- Anode:
 $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$ $E^\circ = -0.54 \text{ V}$
 $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ $E^\circ = -0.82 \text{ V at pH 7}$
- What are the products of electrolysis of a mixture of CuCl_2 and VCl_3 at pH 7?
- Cathode:
 $\text{V}^{3+} + \text{e}^- \rightarrow \text{V}^{2+}$ $E^\circ = -0.26 \text{ V}$
 $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ $E^\circ = 0.34 \text{ V}$
 $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$ $E^\circ = -0.41 \text{ V at pH 7}$
- Anode:
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ $E^\circ = -1.36 \text{ V}$
 $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ $E^\circ = -0.82 \text{ V at pH 7}$

Faraday's Law

- Faraday's Law: the mass of product produced by a given amount of current is proportional to the equivalent weight
- Equivalent weight: molar mass/no. e^- transferred
- Equivalent weight is the mass of substance oxidized or reduced by 1 mole of electrons
- We can pass electrons through a series of cells and compare the amount of substance deposited.
- What is the equivalent weight of ... ?

AgNO_3	CuSO_4
AuCl_3	HCl
- Moles deposited:
 - Ag^+ , 1 mole $\text{e}^- \rightarrow 1 \text{ mol Ag}$
 - Cu^{2+} , 1 mole $\text{e}^- \rightarrow 1/2 \text{ mol Cu}$
 - Au^{3+} , 1 mole $\text{e}^- \rightarrow 1/3 \text{ mol Au}$
 - H^+ , 1 mole $\text{e}^- \rightarrow 1/2 \text{ mol H}_2$
- F = charge on 1 mol $\text{e}^- = 96500 \text{ coul/mol}$
- charge = current x time
- 1 coul = 1 A s
- moles $\text{e}^- = \text{charge (coul)} \times 1 \text{ mol}/96500 \text{ coul}$
- moles $\text{e}^- = \text{current (A)} \times \text{time (s)} \times 1 \text{ coul}/1 \text{ A s} \times 1 \text{ mol}/96500 \text{ coul}$
- If we electrolyze molten NaCl with a current of 5000 A for 30 min (or 1800 s), what mass of Na is produced?

- $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$
- moles $\text{e}^- = 5000 \text{ A} \times 1800 \text{ s} \times 1 \text{ mol}/96500 \text{ coul} = 93.26 \text{ mol}$
- moles $\text{Na} = 93.26 \text{ mol } \text{e}^- \times 1 \text{ mol Na}/1 \text{ mol } \text{e}^- = 93.26 \text{ mol}$
- mass $\text{Na} = 93.26 \text{ mol} \times 22.99 \text{ g/mol} = 2144 \text{ g}$
- We can also calculate how much electrical energy it will take for an electrolysis. We will not pursue these calculations.

- moles $\text{e}^- = \text{current (A)} \times \text{time (s)} \times 1 \text{ coul}/1 \text{ A s} \times 1 \text{ mol}/96500 \text{ coul}$
- How long would we have to electrolyze molten NaCl with a current of 3000 A to produce 150 g of Na?