

V. ELECTROCHEMISTRY

V.1 INTRODUCTION TO OXIDATION AND REDUCTION

Key Question: How are electrons gained or lost?

ELECTROCHEMISTRY is the study of oxidation and reduction reactions in which chemical species lose or gain electrons. This occurs due to differences in electronegativity between two chemical species.

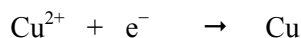
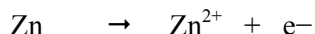
The Nature of Redox

OXIDATION is a process by which _____

REDUCTION is a process by which _____

Consider the reaction: $\text{Cu}^{2+}_{(\text{aq})} + \text{Zn}_{(\text{s})} \rightarrow \text{Cu}_{(\text{s})} + \text{Zn}^{2+}_{(\text{aq})}$

There are two half-reaction processes:



Characteristics of OXIDATION	Characteristics of REDUCTION

Both oxidation and reduction processes occur together. Each half of the full redox reaction is a _____

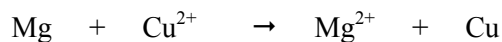
Oxidizing and Reducing Agents

In the example above, Cu^{2+} ions help the Al atoms get oxidized. Cu^{2+} ions = _____

Al metal atoms help the Cu^{2+} ions reduce. Al atoms = _____

Example 1: Identify the RA and OA in the reaction:

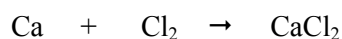
Step 1: RA does the reduction. It is oxidized itself.



OA does the oxidation. It is reduced itself.

Example 2: Identify which substance is oxidized or reduced and the RA and OA in the following reactions:

Step 1: Dissociate salt.



Step 2: RA does the

reduction. It is oxidized.

OA does the oxidation. It is reduced.

Summary of Redox Theory	
1.	A redox reaction is a chemical reaction in which electrons are transferred between chemical entities.
2.	The number of electrons gained by one species equals the number of electrons lost by another.
3.	Reduction is a process in which electrons are gained. (GER)
4.	Oxidation is a process in which electrons are lost. (LEO)
5.	A reducing agent promotes reduction by donating (losing) electrons in a redox reaction.
6.	An oxidizing agent promotes oxidation by removing (gaining) electrons in a redox reaction.

V.3 DETERMINING THE SPONTANEITY OF REDOX REACTIONS

Predicting Redox Using Strength of Oxidizing and Reducing Agents

We can predict if a reaction is: a) spontaneous as written or

b) the reverse reaction is preferred (i.e. forward reaction is non-spontaneous).

Using the Standard Reduction Table

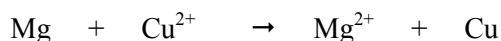
A reaction will be spontaneous if the _____ is above the _____ on the Standard Reduction Table:

Example 3: Predict the spontaneity of the reaction:

Step 1: Identify SOA & SRA.

SOA = highest on left side

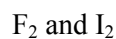
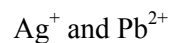
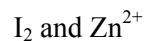
SRA = lowest on right side



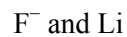
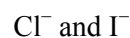
Step 2: SOA higher than SRA = spontaneous. SOA lower than SRA = nonspontaneous.

Example 4: Identify the SOA and SRA from the possible reacting species. Predict whether the reaction is spontaneous.

a. Identify the stronger oxidizing agent in each pair below:



b. Identify the stronger reducing agent in each pair below:



c. Will I_2 react spontaneously with Zn to form I^- and Zn^{2+} ?

d. Will Zn react with the Al^{3+} in an $Al(NO_3)_3$ solution?

Do Hebden Set 36 #1-2, for 7, 8, 11, 12 do (a, c, e, g), 9, 10, 13, 15, 17 (p. 192-200)

V.4 WRITING HALF-REACTION EQUATIONS

IF the balanced half reaction is NOT on the Reduction Table, it is possible to write it. How you balance these depends whether it is in ACIDIC SOLUTION or BASIC SOLUTION.

Method

The officer's names was **MAJOR HYDROXIDE = MAJOR OH⁻**

1. Balance MAJOR species
2. Balance O using H_2O
3. Balance H using H^+
4. Balance charge using e^-

Acidic Solution

Example 5: Write a balanced half-reaction for $\text{N}_2\text{O}_4 \rightarrow \text{NO}_3^-$ in an acid solution.

Step 1: Write unbalanced equation.

Step 2: Balance all atoms except O and H.

Step 3: Balance O by adding H_2O (on side with less O's)

Step 4: Balance H by adding H^+ (on side with less H's)

Step 5: Balance charge by adding e^- (on side with more "+" charge)

Check.

*You do not have to show each step separately.

Basic Solution

Balance as for an acidic solution. One more step: Add OH^- to both sides of the equation to neutralize H^+ .

Example 6: Write a balanced half-reaction for $\text{N}_2\text{O}_4 \rightarrow \text{NO}_3^-$ in a basic solution.

Step 1 - 5: As above.

Step 6: Add OH^- to both sides to neutralize H^+ .

Step 7: Cancel and Check.

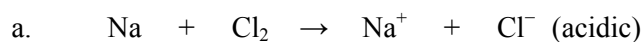
Do Hebden Set 37 #19 (a, c, e, g, i, l, m) (p. 203)

V.5 BALANCING FULL REDOX REACTIONS: USING HALF-REACTIONS METHOD

Method

1. Break up reaction into _____
2. **MULTIPLY** each half reaction by an integer so as to make _____
3. **ADD** the half reactions and **CANCEL**.
4. If basic, then neutralize H^+ with OH^- and simplify.

Example 7: Balance using half-reactions method:

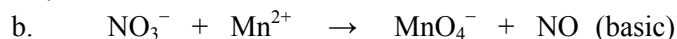


Step 1: Find or write half-reactions.

Step 2: Multiply to make e^- equal on both sides.

Step 3: Add half-reactions and cancel.

Check (charges are balanced).



Step 1: Find or write half-reactions.

*When looking up half reactions, don't worry about the H_2O , H^+ or OH^- in the equations. Just match up the major species.

Step 2: Multiply to make e^- equal on both sides.

Step 3: Add half-reactions and cancel.

Step 4: Neutralize H^+ with OH^- and cancel.

Check (charges balanced?)

Do Hebden Set 38 #24 (a, b, g, h, i, k, o, r, t, v) (p. 207)

V.2 OXIDATION NUMBERS

Oxidation States

Oxidation numbers (O.N.) are imaginary electric charges of particles if the particles were _____

Common Oxidation Numbers

Chemical Species	Examples	Oxidation Number
all atoms in ELEMENTS (monatomic and diatomic)		
IONIC COMPOUNDS: Alkali (Group 1)		

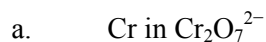
Halogens		
HYDROGEN in Compounds		
Metallic hydrides		
OXYGEN in Compounds		
Peroxides (O_2^{2-})		
all MONATOMIC IONS		

Many chemical elements can have many different oxidation states and, therefore, many different oxidation numbers. For example, some of sulphur's oxidation states are -2 , $+2$, $+4$, $+6$. A system has been developed to determining the oxidation numbers of these kinds of elements.

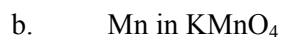
Rules for Determining Oxidation Numbers

1. Assign _____
2. The sum of the oxidation numbers equals the charge of the molecule or ion.
 - a) The ON for a neutral compound =
 - b) The ON for a polyatomic ion =
3. Any unknown oxidation number is determined algebraically from the sum of the known oxidation numbers

Example 8: Find the oxidation number of



Assign common O.N. Then
let the unknown O.N. = x .
Solve algebraically.



Assign common O.N. Then
let the unknown O.N. = x .
Solve algebraically.

Changes in Oxidation Numbers

These numbers are useful in deciding whether oxidation or reduction has occurred, and for balancing redox

equations.

If the oxidation number **increases** = _____

If the oxidation number **decreases** = _____

Example 9: Find the Δ ON for each species and identify each as an oxidation or reduction.



ON inc. = oxidation.

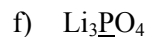
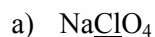
ON dec. = reduction



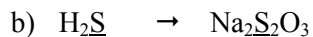
ON inc. = oxidation.

ON dec. = reduction

Example 10: Find the oxidation number of the underlined element.



Example 11: Given the reaction, use oxidation numbers to determine whether the element has been reduced or oxidized.



Do Hebden Set 39 #3 (a - i, n, s), 4, 5, 20, 22, 23 (p. 194-204)

V.6 BALANCING FULL REDOX EQUATIONS – USING OXIDATION NUMBERS

Method

1. Determine and _____ and find _____
2. Make _____. (MULTIPLY Δ ON to get a common multiple \rightarrow this becomes the #of atoms of each species).
3. Balance any other _____ elements (not H or O) that are not yet balanced.
4. Add H_2O to balance
5. Add H^+ to balance
6. If basic then add OH^- to both sides to neutralize any H^+ . (Combine H^+ and OH^- to form H_2O).
7. Simplify/cancel and check.

Example 12: Balance the following reaction using oxidation number method.

Write O.N. and find Δ O.N.

Δ O.N. = + = reduction

Δ O.N. = - = oxidation

MAJOR OH (don't add e⁻)



Example 13: Balance the redox reactions using oxidation numbers. Use a separate sheet of paper.

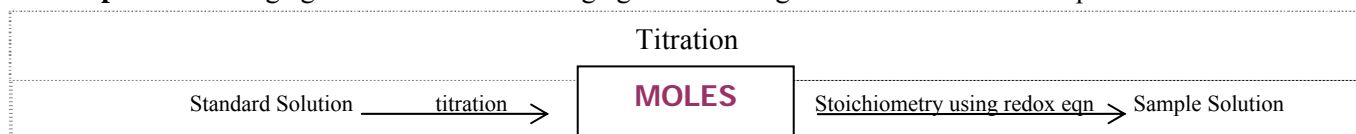
- a) $\text{OsO}_4 + \text{C}_2\text{H}_5\text{OH} \rightarrow \text{OsO}_3^- + \text{CH}_3\text{COOH}$ (acid solution)
- b) $\text{Al}_{(l)} + \text{MnO}_{(s)} \rightarrow \text{Al}_2\text{O}_{3(s)} + \text{Mn}_{(l)}$ (acid solution)
- c) $\text{I}_2\text{O}_{5(s)} + \text{CO}_{(g)} \rightarrow \text{I}_{2(s)} + \text{CO}_{2(g)}$ (acid solution)
- d) $\text{ClO}_3^- \rightarrow \text{Cl}_2 + \text{ClO}_4^-$ (basic solution)

Do Hebden Set 40 #25 (a – f, l, m, n) (p. 210)

V.7 REDOX TITRATION & QUANTITATIVE ASPECTS OF REDOX REACTIONS

Like acids and bases, redox reaction can also be titrated -- the oxidizing agent can be titrated against the reducing agent until an end point is reached. The purpose would be to determine the concentration of a solution.

Principle: Oxidizing agent reacts with reducing agent according to balanced REDOX equation.



Example 14: A solution containing I_2 is titrated to the end point using 18.0 mL of 0.13 M $\text{Na}_2\text{S}_2\text{O}_3$ solution to form I^- and $\text{S}_4\text{O}_6^{2-}$. How many moles of I_2 were present? [Answer: 1.2×10^{-3} mol]

Step 1: Identify species present.

If the reactants are not obvious or given from the question, then SOA and SRA present will react.

Step 2: Write balanced redox reaction.

Step 3: Calculate moles of given species.

Step 4: Use mole ratio to calculate moles of unknown at equivalence point.

Do Hebden Set 41 #26, 27, 29, 30, 32 (p. 213-214)

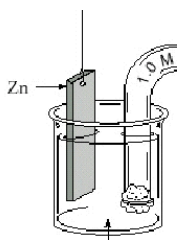
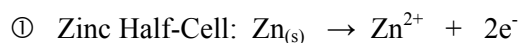
ELECTROCHEMICAL AND ELECTROLYTIC CELLS

Key Question: How are electrochemical principles applied in electrochemical and electrolytic cells?

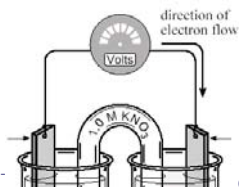
V.8 ELECTROCHEMICAL CELLS

Consider the Zn and Cu electrochemical cell:

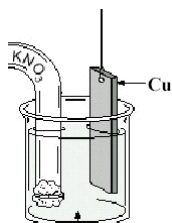
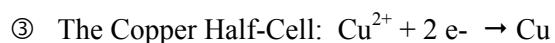
ANODE - oxidation



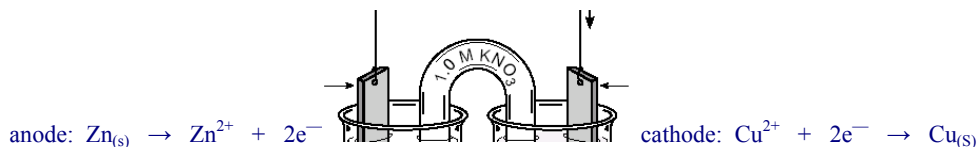
② **FLOW OF ELECTRONS**



CATHODE - reduction



④ SALT BRIDGE

**Diagramming Electrochemical Cells**

Draw a standard electrochemical cell using Zn and Cu half-cells. Indicate electron flow, electrolytes, half reactions, overall reaction, cell voltage, the anode, the cathode, the positive and negative electrode:

Summary

- REDUCTION occurs at the _____ and it _____ mass, and _____ ions migrate to _____ ode. It is connected to _____ tive terminal of the voltmeter.
- OXIDATION occurs at the _____ and it _____ mass, and _____ ions migrate to _____ ode. It is connected to _____ tive terminal of the voltmeter.

Electrons travel through wires towards the _____.

Example 16: Describe a standard electrochemical cell using Fe and Mg half-cells. Sketch and label: electron flow, electrolytes, half reactions, overall reaction, the anode, the cathode, the positive and negative electrode.



Steps:

1. One half cell must contain Fe electrode dipped into a solution containing Fe^{2+} .
2. One half cell must contain Mg electrode dipped into a solution containing Mg^{2+} .

3. Since Mg is being oxidized, it is the _____.
4. Since Fe is being reduced it is the _____. **REACTION HIGHER ON REDUCTION TABLE WILL BE CATHODE**
5. Anions move toward the _____. Cations move toward the _____.
6. e^- travel from _____ to the _____.

Do Hebden Set 42 #34–35 (p. 217)

V.9 STANDARD REDUCTION POTENTIALS

VOLTAGE =

REDUCTION POTENTIAL =

Half-Cell Potentials

CELL POTENTIAL E_{cell} is the maximum voltage of an electrochemical cell. (strength of attracting e^-)

STANDARD CELL POTENTIAL E°_{cell} is the maximum voltage when:

The purpose of E°_{cell} is to:

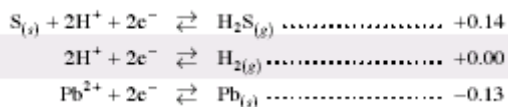
- 1) Calculate voltage of a particular electrochemical cell.
- 2) Predict whether a particular redox reaction will occur (=spontaneous)

E°_{cell} is calculated from the _____ between the half reactions.

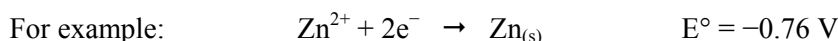
Table of Standard Reduction Potentials

Oxidizing Agents	Reducing Agents	E° (Volts)
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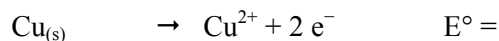
STANDARD REFERENCE HALF-CELL



All other half-cells are relative compared to this half-cell. The higher the species on the reduction potential table, the greater its attraction for electrons, the greater its reduction potential and the stronger it acts as an oxidizing agent.



Each reduction half reaction can be written as an oxidation.



Calculating E°_{cell} for an Electrochemical Cell

Example 17: What is the E°_{cell} for the electrochemical cell for:



Step 1: Determine SOA, SRA

Step 2: Write the cathode and anode reaction with their E° values.

CATHODE IS HIGHER ON
REDUCTION TABLE.

Remember reverse oxidation reaction AND reverse the sign of the E° value.

cathode: $E^\circ =$

anode: $E^\circ =$

Step 3: Multiply half reactions to make e- cancel. DO NOT
MULTIPLY E° VALUES.

Step 4: Add the reaction and E° values to get E°_{cell}.

overall: $E^\circ_{\text{cell}} =$

V.10 PREDICTING SPONTANEITY OF REACTION USING CELL POTENTIALS

When an oxidizing and reducing agent are mixed there are two possible outcomes:

SPONTANEOUS: reactants will _____ in a redox reaction.

NON-SPONTANEOUS: reactants will _____

Method 1

Overall E°_{cell} is (+) =

Overall E°_{cell} is (—) =

Example 18: Will salt in seawater (ignore water) react with an iron pail that it is stored in? [Answer: non-spontaneous, $E^\circ_{\text{cell}} = -2.26\text{V}$]

Step 1: Determine SOA, SRA

Step 2: Write the cathode and anode reaction with their E° values. CATHODE IS HIGHER ON RED. TABLE.

Step 3: Multiply half reactions to make e- cancel. DO NOT MULTIPLY E° VALUES.

Step 4: Add the reaction and E° values to get E°_{cell} . (+) = spont, (-) = non-spont

cathode:	E°	=
<u>anode:</u>	E°	=

overall:	E°_{cell}	=
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Method 2

If OA is higher than RA =

Example 19: Which of Fe^{3+} , Al^{3+} , Cl_2 or Au^{3+} will best oxidize $\text{Pb}_{(s)}$? [Answer: Au^{3+}]

Step 1: Determine the preferred reactants (SOA, SRA).

Step 2: SOA higher than SRA = spontaneous.

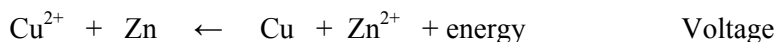
Do Hebden Set 43 #36 (a, c, e, h), 37, 38, 40 (p. 224-225)

V.11 APPLICATIONS OF ELECTROCHEMISTRY

When cells are first constructed, they are _____. All the voltages calculated are _____. All electrochemical cells are exothermic (strong tendency to form products).

Consider: $\text{Cu}^{2+} + \text{Zn} \rightarrow \text{Cu} + \text{Zn}^{2+} + \text{energy}$

INITIALLY: As the cell operates, $[\text{Cu}^{2+}]$ _____ creases. And $[\text{Zn}^{2+}]$ _____ creases. Both these changes pushes the reaction to shift _____, so



Eventually, these tendencies will be _____. At this point, the cell has reached _____.

Factors affecting Electrochemical Cells:

Concentration

A change in concentration means concentration $\neq 1 \text{ M}$, not standard conditions, remove “ $^\circ$ ” from E°_{cell}

Apply Le Chatelier’s Principle:

Consider: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}_{(\text{s})}$ $E^\circ_{\text{cell}} = +0.80 \text{ V}$ if $[\text{Ag}^+] = 1 \text{ M}$

$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}_{(\text{s})}$ E_{cell} if $[\text{Ag}^+] = 0.5 \text{ M}$

Lower $[\text{Ag}^+]$, LCP: shift _____, reduction favoured _____, so _____ E_{cell} value.

$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}_{(\text{s})}$ E°_{cell} if $[\text{Ag}^+] = 1.5 \text{ M}$

Higher $[\text{Ag}^+]$, LCP: shift _____, reduction favoured _____, so _____ E_{cell} value.

Surface Area

What if we used a larger electrode? Will the E°_{cell} be larger, smaller, or the same?

The electrode is solid. Changing the amount of solid will not affect its concentration. If concentration is not affected, neither is the redox equilibrium.

Note: RATE at which cell works (reaction rate) will change.

Multiple Electrodes

What if we have multiple anodes and cathodes (mixture of substances)? Which electrode will the reaction proceed?

The other electrodes will be spectators.

Do Hebden Set 44 #47–48 (p. 228)

Breathalyser

Read about Breathalyser (p. 228 – 229) in Hebden and Do #49.

After drinking, breath contains ethanol C_2H_5OH . Acidified dichromate (at $E^\circ = 1.23$ on the reduction table) will oxidize alcohol. The unbalanced formula equation is:



yellow (orange at
higher concentrations)

$Cr_2(SO_4)_3$ is green

Exhaled air is mixed with standardized acidified dichromate. Put in a spectrophotometer set at the wavelength of green light. More alcohol produces more green $Cr_2(SO_4)_3$. Machine is calibrated with known concentration samples of alcohol to ensure accuracy.

BATTERIES

Read about the different types of batteries (p. 230 – 233) in Hebden and complete the notes by making notes on the anode and cathode, electrolyte, and oxidation and reduction half reactions.

Zinc-carbon battery

<p>Label the cathode and anode. (p. 231)</p> <p style="text-align: center;">not rechargeable</p>	<ul style="list-style-type: none"> ▪ Carbon cathode and zinc casing acts as anode ▪ NH_4Cl/MnO_2 paste inside ▪ Carbon cathode is inert, the simple cathode reaction is: ▪ Zinc anode reaction is:
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Why do batteries die?

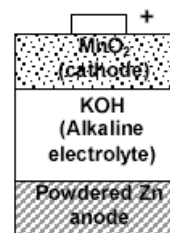
E°_{cell} is initial cell voltage. Overtime, the redox reaction reaches equilibrium. Equilibrium is when no more reactant is left (no Zn) or when the electrode gets coated with a by-product (cannot reach electrode). $E^\circ_{cell} = 0$ V. The battery dies when you run out of zinc or when the electrode gets covered with $Zn(NH_4)_4^{2+}$.

Alkaline Battery (Duracell, Energizer)

- Same set up as the zinc-carbon battery except:
- NH_4Cl/MnO_2 paste is replaced by KOH base (alkaline)

What are the advantages of the alkaline cell? (p. 232)

1. It lasts longer because there is no $Zn(NH_4)_4^{2+}$ build up at electrodes.



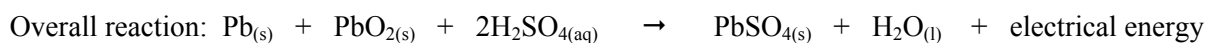
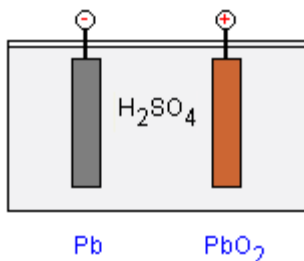
2. _____

“Car Battery” (Lead-Acid Battery)

Complete the notes: (p. 230)

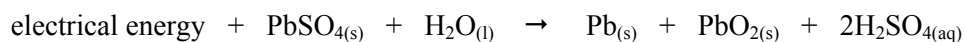
Anode reaction:

Cathode reaction:



Why can car batteries be recharged?

The battery dies because of $\text{PbSO}_{4(s)}$ build up on the electrodes. “Boosting” the car, sends electricity through the battery reversing the reaction.



(from alternator or generator)

Why do car batteries eventually die? (i.e. Why does the cell deteriorate?)

Fuel Cells

Complete the notes: (p. 233)

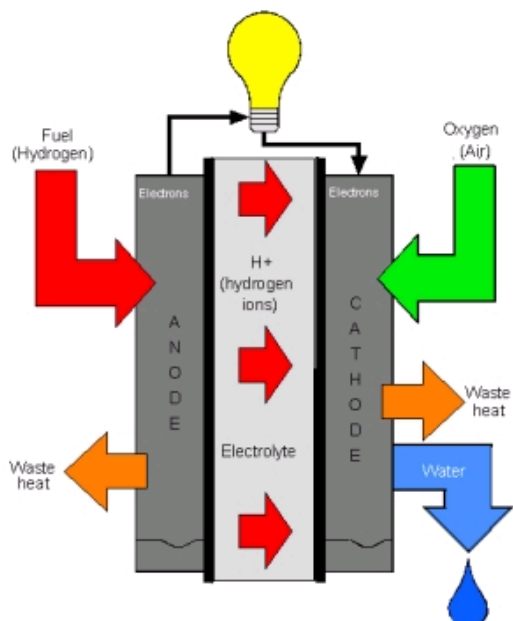
What is a fuel cell?

What are the advantages of fuel cells?

What are the disadvantages of fuel cells?

- electrodes corrode quickly and require constant maintenance.

- expensive
- need large cell to produce sufficient electrical energy.



The most common type of fuel cell is the **HYDROGEN-OXYGEN FUEL CELL.**

The anode reaction is:

The cathode reaction is:

The overall reaction is:

V.12 CORROSION

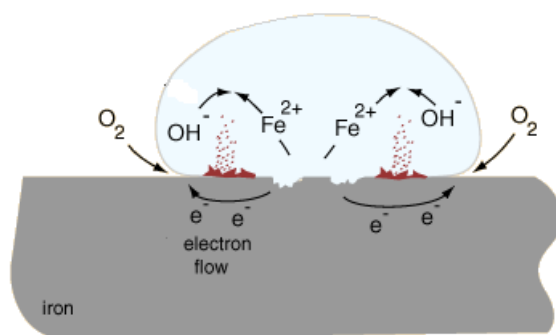
CORROSION = _____

“Rusting” is oxidation of iron

Describe the conditions necessary for corrosion to occur: _____

Rusting is accelerated in the presence of acidic solutions, electrolytes, mechanical stresses, and contact with less active metals.

Analyze the process of metal corrosion using electrochemistry.



1. The “electrochemical cell” cell is a drop of water.
2. Anode is the centre of the drop. Cathode is the edge of the drop.

3. Fe is oxidized to Fe^{2+} (anode). O_2 picks up e^- from the Fe and is reduced to OH^- (cathode).
4. The Fe^{2+} reacts with the OH^- to make $\text{Fe}(\text{OH})_{2(s)}$
5. The $\text{Fe}(\text{OH})_2$ is then oxidized $\rightarrow \text{Fe}(\text{OH})_3 \rightarrow \text{Fe}_2\text{O}_3(s) = \text{RUST}$.

How can we stop corrosion?

Method	Description
1. Protect the Metal Surface	
2. CATHODIC PROTECTION	$\begin{array}{l} \text{Fe}^{2+} + 2e^- \rightleftharpoons \text{Fe}_{(s)} \dots\dots\dots -0.45 \\ \text{Ag}_2\text{S}_{(s)} + 2e^- \rightleftharpoons 2\text{Ag}_{(s)} + \text{S}^{2-} \dots\dots\dots -0.69 \\ \text{Cr}^{3+} + 3e^- \rightleftharpoons \text{Cr}_{(s)} \dots\dots\dots -0.74 \\ \text{Zn}^{2+} + 2e^- \rightleftharpoons \text{Zn}_{(s)} \dots\dots\dots -0.76 \\ \text{Te}_{(s)} + 2\text{H}^+ + 2e^- \rightleftharpoons \text{H}_2\text{Te} \dots\dots\dots -0.79 \\ 2\text{H}_2\text{O} + 2e^- \rightleftharpoons \text{H}_{2(g)} + 2\text{OH}^- \dots\dots\dots -0.83 \\ \text{Mn}^{2+} + 2e^- \rightleftharpoons \text{Mn}_{(s)} \dots\dots\dots -1.19 \\ \text{Al}^{3+} + 3e^- \rightleftharpoons \text{Al}_{(s)} \dots\dots\dots -1.66 \\ \text{Mg}^{2+} + 2e^- \rightleftharpoons \text{Mg}_{(s)} \dots\dots\dots -2.37 \end{array}$

Do Hebden Set 45 #52-63 (p. 229-236)

V.13 ELECTROLYSIS

ELECTROCHEMICAL CELL - spontaneous redox reaction produces electricity

ELECTROLYTIC CELL - add electricity to non-spontaneous reactions to make them go

Electrolysis of Water

Water will NOT spontaneously reduce to H_2 :



Water will NOT spontaneously oxidize to O_2 :



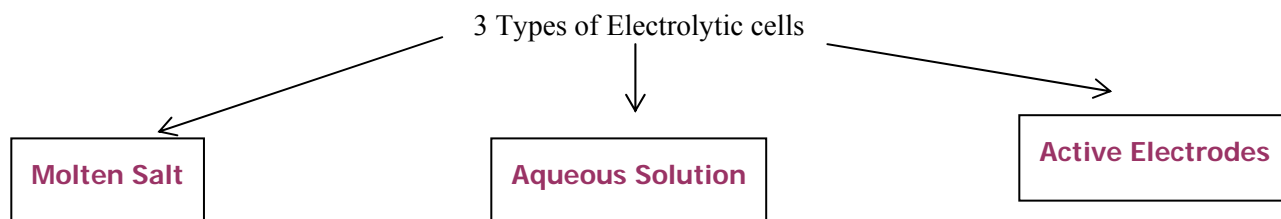
But if we add electricity (e^-) we can start the reduction process:

Half-cell **OVERPOTENTIAL**: In practice, these half-cells are not at standard state or there is resistance in the wire for example. Therefore, a half-cell might _____ (aka: OVER POTENTIAL) to actually undergo oxidation-reduction.

Electrolysis of water's overpotential is shown on the Table of Standard Reduction Potentials.

$\text{ClO}_4^- + 8\text{H}^+ + 8\text{e}^- \rightleftharpoons \text{Cl}^- + 4\text{H}_2\text{O}$	+1.39	Overpotential Effect
$\text{Cl}_{2(g)} + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-$	+1.36	
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+1.23	
$\frac{1}{2}\text{O}_{2(g)} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}$	+1.23	
$\text{MnO}_{2(s)} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 2\text{H}_2\text{O}$	+1.22	
$\text{IO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightleftharpoons \frac{1}{2}\text{I}_{2(s)} + 3\text{H}_2\text{O}$	+1.20	
$\text{Br}_{2(l)} + 2\text{e}^- \rightleftharpoons 2\text{Br}^-$	+1.09	
$\text{AuCl}_4^- + 3\text{e}^- \rightleftharpoons \text{Au}_{(s)} + 4\text{Cl}^-$	+1.00	
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{NO}_{(g)} + 2\text{H}_2\text{O}$	+0.96	
$\text{Hg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Hg}_{(l)}$	+0.85	
$\frac{1}{2}\text{O}_{2(g)} + 2\text{H}^+(10^{-7}\text{M}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}$	+0.82	
$2\text{NO}_3^- + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{N}_2\text{O}_4 + 2\text{H}_2\text{O}$	+0.80	
$\text{Cr}^{3+} + \text{e}^- \rightleftharpoons \text{Cr}^{2+}$	-0.41	
$2\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{H}_2 + 2\text{OH}^-(10^{-7}\text{M})$	-0.41	
$\text{Fe}^{2+} + 2\text{e}^- \rightleftharpoons \text{Fe}_{(s)}$	-0.45	
$\text{Ag}_2\text{S}_{(s)} + 2\text{e}^- \rightleftharpoons 2\text{Ag}_{(s)} + \text{S}^{2-}$	-0.69	
$\text{Cr}^{3+} + 3\text{e}^- \rightleftharpoons \text{Cr}_{(s)}$	-0.74	
$\text{Zn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Zn}_{(s)}$	-0.76	
$\text{Te}_{(s)} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{Te}$	-0.79	

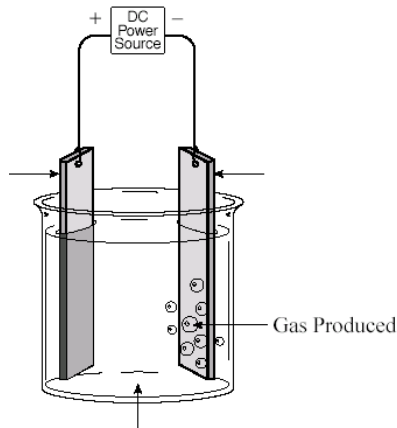
More than the theoretical voltage is required. For water this effect is very profound. Water reactions must be considered to be ranked at the position of the overpotential arrows.



Electrolysis Cell: Molten (Melted) Salt

- Contains _____ (usually Pt or carbon-graphite)
- Requires high temperatures to melt salts (therefore may be difficult to perform)
- Relatively simple to analyze since only one OA and one RA (from the binary salt)

Example 20: Electrolysis of molten NaCl makes Cl_2 and Na. $\text{NaCl}_{(l)} \rightarrow \text{Na}^+_{(l)} + \text{Cl}^-_{(l)}$



Note: E°_{cell} for Electrolytic cells are ALWAYS NEGATIVE

Electrolysis Cell:– Aqueous Solution (Dissolved Binary Salt)

- Contains inert electrodes in _____
- More complicated reactions, but easier in real life because lower temperatures are required.
- Main difference: _____. (Therefore, must analyze H_2O too, might be SOA or SRA)

Example 20: Aqueous NaCl. What is the reaction at the anode and cathode? What is the overall reaction?

Step 1: List ALL species present.

Step 2: Determine the conditions:

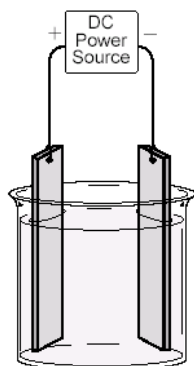
Neutral – Consider the two H_2O reactions

Acidic – Consider the $\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_{(\text{g})}$ reduction reaction

Step 3: Determine SOA and SRA. SOA will react with SRA, the other species will be spectators.

***OVERPOTENTIAL EFFECT:** If Br^- or Cl^- are present, it will be the anode reaction.

Step 4: Write anode, cathode with E° values and find overall reaction and E°_{cell} .



species:

conditions:

cathode:

anode: _____

overall:

Do Hebden Set 46 #64-67, 68 (a, d, h), 69, 70, 72 (p. 238-242)

Active electrode - ELECTROPLATING

= an unwanted or cheap metal is coated with a better metal from the salt solution. (Coated at the CATHODE)

Example 21: Silver Plating

	<p>Cathode =</p> <p>Anode =</p> <p>Solution =</p>
--	---

Note: Symbol for battery is $| |$ (Short end = negative. e^- comes out of negative end and goes to cathode.)

ELECTROREFINING = an impure metal is purified.

Read about **ELECTROREFINING** p. 245 – 246 in Hebden.

The **Impure Metal** is the (anode/cathode) _____

The **Pure metal** is the (anode/cathode) _____

The **Electrolyte** (solution) contains the (anion/cation) _____ of the metal to be purified.

ELECTROWINNING = reduction of ores to produce metals in industry.

ALUMINUM PRODUCTION = purifying Bauxite to get Aluminum

Do Hebden Set 47 #73-77 (p. 244-246)