Unit 6 - Oxidation - Reduction

Aug 27-7:52 AM

I) Oxidation and Reduction

http://www.calgaryacademy.com/ICT/rr/redox1.html

What is a redox reaction?

What is oxidation?

What types of elements tend to oxidize?

What is reduction?

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I) Oxidation and Reduction

http://www.calgaryacademy.com/ICT/rr/redox1.html

What is a redox reaction?

A reaction that involves a transfer of electrons from one substance to another.

What is oxidation?

a loss of electrons

What types of elements tend to oxidize?

metals, as they give away electron(s) to become cations

What is reduction?

a gain of electrons

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What is a good way to remember oxidation and reduction?

In order for a redox reaction to occur, there must be an oxidation AND a reduction (ie. one substance has to first 'give-up' electrons in order to another substance to gain them).

http://cwx.prenhall.com/petrucci/medialib/media_portfolio/text_images/022_REDOXREACTS1.MOV

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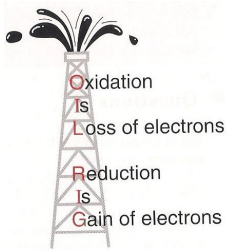
What is a good way to remember oxidation and reduction? OIL RIG

O: oxidation
I: is
L: loss

R: reduction
I: is
G: gain



In order for a redox reaction to occur, there must be an oxidation AND a reduction (ie. one substance has to first 'give-up' electrons in order to another substance to gain them).

http://cwx.prenhall.com/petrucci/medialib/media_portfolio/text_images/022_REDOXREACTS1.MOV

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Single replacement reactions are redox reactions. Take, for example, the reaction of copper metal in silver nitrate solution: http://cwx.prenhall.com/petrucci/medialib/media_portfolio/text_images/024_SILVERCRYSTA.MOV

$Cu_{(s)} + AgNO_{3(aq)} \longrightarrow$

Copper starts as a metal (in atomic form - neutral) and becomes an ion dissolved in solution with charge of _____. Did copper donate or receive electrons? _____ How many? _____ Therefore, Cu underwent _____ and the half-reaction is:

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Single replacement reactions are redox reactions. Take, for example, the reaction of copper metal in silver nitrate solution: http://cwx.prenhall.com/petrucci/medialib/media_portfolio/text_images/024_SILVERCRYSTA.MOV

$Cu_{(s)} + 2AgNO_{3(aq)} \longrightarrow 2Ag_{(s)} + Cu(NO_3)_{(aq)}$

Copper starts as a metal (in atomic form - neutral) and becomes an ion dissolved in solution with charge of __+2__. Did copper donate or receive electrons? __donated__ How many? __2__ Therefore, Cu underwent __oxidation__ and the half-reaction is:

$Cu \longrightarrow Cu^{2+} + 2e$

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Silver, as a reactant, is a dissolved ion with a charge of _____.
What is silver as a product? _____
Did silver donate or accept electrons? _____
Therefore, _____ underwent _____
and the half-reaction is:

Nitrate, NO_3 , started as a dissolved ion and didn't change, so it didn't take part in the redox reaction (it is a spectator ion).

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Silver, as a reactant, is a dissolved ion with a charge of __+1__.
What is silver as a product? __Ag metal (neutral)__
Did silver donate or accept electrons? __accept__
Therefore, __Ag^+__ underwent __reduction__
and the half-reaction is:

$Ag^+ + e \longrightarrow Ag$

Nitrate, NO_3 , started as a dissolved ion and didn't change, so it didn't take part in the redox reaction (it is a spectator ion).

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Overall Reaction:
$Cu_{(s)} + 2AgNO_{3(aq)} \longrightarrow 2Ag_{(s)} + Cu(NO_3)_{2(aq)}$

Oxidation Half-Reaction:

Reduction Half-Reaction:

Balancing Electrons:

Overall NET Redox Reaction:

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Overall Reaction:
$Cu_{(s)} + 2AgNO_{3(aq)} \longrightarrow 2Ag_{(s)} + Cu(NO_3)_{2(aq)}$

Oxidation Half-Reaction: $Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$

Reduction Half-Reaction: $Ag^+_{(aq)} + e \longrightarrow Ag_{(s)}$

Balancing Electrons: $Cu \longrightarrow Cu^{2+} + 2e$
$2Ag^+ + 2e \longrightarrow 2Ag$

Overall NET Redox Reaction:
$Cu_{(s)} + 2Ag^+_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

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For every copper atom that oxidized, two electrons were released, which are enough to _____
(hence the 2:1 stoichiometry).

Notice that the reactant that was in metal form, copper, oxidized to become a cation, Cu^{2+}, as metals tend to do. Ag^+, a metal cation, reduced to become a metal (the opposite process - a reduction). http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/redox/home.html
A half-reaction specifies the oxidation or reduction that occurred, whereas the net redox reaction is the combination of both and shows what and how the substances changed.

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For every copper atom that oxidized, two electrons were released, which are enough to reduce two Ag^+ ions - as can be seen in the balanced reaction (hence the 2:1 stoichiometry).

Notice that the reactant that was in metal form, copper, oxidized to become a cation, Cu^{2+}, as metals tend to do. Ag^+, a metal cation, reduced to become a metal (the opposite process - a reduction). http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/redox/home.html
A half-reaction specifies the oxidation or reduction that occurred, whereas the net redox reaction is the combination of both and shows what and how the substances changed.

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Silverware (silver metal) reacts with H_2S that is present in trace amounts in air to produce Ag_2S (tarnish). The cleaning of silverware is a redox reaction:

$2Al + 3Ag_2S \longrightarrow 6Ag + Al_2S_3$

*This reaction requires heat to attain E_a and water to act as an electron transfer medium. Thus, the silverware is placed into a pan of water that is lined with aluminum foil, and then heated in the oven.

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Oxidation half-reaction:

*Remember, oxidation is a loss of electrons, so electrons will be a product

Reduction half-reaction:

*Reduction is a gain of electrons, so electrons will be a reactant

Balance electrons to create the net redox reaction:

Notice that the electrons are not part of the net reaction, as they have been transferred from one substance to the other.

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Oxidation half-reaction: $Al_{(s)} \longrightarrow Al^{3+}_{(aq)} + 3e^-$

*Remember, oxidation is a loss of electrons, so electrons will be a product

Reduction half-reaction: $Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}$

*Reduction is a gain of electrons, so electrons will be a reactant

Balance electrons to create the net redox reaction:

$Al \longrightarrow Al^{3+} + 3e^-$
$3Ag^+ + 3e^- \longrightarrow 3Ag$
NET: $Al + 3Ag^+ \longrightarrow Al^{3+} + 3Ag$

Notice that the electrons are not part of the net reaction, as they have been transferred from one substance to the other.

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II) Oxidizing Agents and Reducing Agents

An **oxidizing agent** is a substance that _____ another substance. Therefore, the oxidizing agent itself undergoes _____.
What was the oxidizing agent in the previous example? _____ http://www.calgaryacademy.com/ICT/rr/redox3.html

A **reducing agent** is a substance that _____ another substance. Therefore, the reducing agent itself undergoes _____.
What was the reducing agent in the previous example? _____ http://www.calgaryacademy.com/ICT/rr/redox2.html

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II) Oxidizing Agents and Reducing Agents

An **oxidizing agent** is a substance that oxidizes another substance. Therefore, the oxidizing agent itself undergoes reduction.
What was the oxidizing agent in the previous example? Ag^+ http://www.calgaryacademy.com/ICT/rr/redox3.html

A **reducing agent** is a substance that reduces another substance. Therefore, the reducing agent itself undergoes oxidation.
What was the reducing agent in the previous example? Al http://www.calgaryacademy.com/ICT/rr/redox2.html

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Practice Questions: For each of the following reactants, give the
a) oxidation half-reaction (b) reduction half-reaction (c) balanced redox reaction (d) oxidizing agent (e) reducing agent
1) Zn and Cu^{2+}

http://ibchem.com/IB/ibfiles/redox/red_swf/zn%2Bcu.swf

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Practice Questions: For each of the following reactants, give the
a) oxidation half-reaction (b) reduction half-reaction (c) balanced redox reaction (d) oxidizing agent (e) reducing agent
1) Zn and Cu^{2+}

a) $Zn \longrightarrow Zn^{2+} + 2e^-$

b) $Cu^{2+} + 2e^- \longrightarrow Cu$

c) $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

d) Cu^{2+}

e) Zn

http://ibchem.com/IB/ibfiles/redox/red_swf/zn%2Bcu.swf

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2) Li and Fe^{3+}

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2) Li and Fe^{3+}

a) $Li \longrightarrow Li^+ + e$

b) $Fe^{3+} + 3e \longrightarrow Fe$

c) $3Li + Fe^{3+} \longrightarrow 3Li^+ + Fe$

d) Fe^{3+}

e) Li

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3) Br_2 and Cr to produce $2Br^-$ and Cr^{3+}

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3) Br_2 and Cr to produce $2Br^-$ and Cr^{3+}

a) $Cr \longrightarrow Cr^{3+} + 3e^-$

b) $Br_2 + 2e \longrightarrow 2Br^-$

c) $2Cr + 3Br_2 \longrightarrow 2Cr^{3+} + 6Br^-$

d) Br_2

e) Cr

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Assignment 1

Read Hebden p.190 (start at 'Definitions') & 191 and do p.192 #1 & 2

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Assignment 1

Read Hebden p.190 (start at 'Definitions') & 191 and do p.192 #1 & 2

answers in the back of Hebden

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III) Oxidation Numbers

An oxidation number is the real (for atoms, ions and ionic compounds) or apparent (for covalent compounds) charge that a particle possesses. It is very similar to 'combining capacity'.

What are the oxidation numbers for the ionic compound NaCl?

What are the oxidation numbers for the covalent compound H_2O?

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III) Oxidation Numbers

An oxidation number is the real (for atoms, ions and ionic compounds) or apparent (for covalent compounds) charge that a particle possesses. It is very similar to 'combining capacity'.

What are the oxidation numbers for the ionic compound NaCl?

+1 -1

What are the oxidation numbers for the covalent compound H_2O?

+1 -2

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Guidelines for Assigning Oxidation Numbers

When assigning oxidation numbers, the guidelines must be followed IN ORDER.

The sum of the positive and negative charges must equal the overall charge of the substance.

A) For neutral ATOMS, the oxidation number is _____.

Examples: Al P_4 N_2 S_8

B) For ions, the oxidation number equals the charge.

Examples: Mn^{2+} Br S^{2-} Cu^{+} Al^{3+}

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Guidelines for Assigning Oxidation Numbers

When assigning oxidation numbers, the guidelines must be followed IN ORDER.

The sum of the positive and negative charges must equal the overall charge of the substance.

A) For neutral ATOMS, the oxidation number is zero.

Examples: Al P_4 N_2 S_8

0 0 0 0

B) For ions, the oxidation number equals the charge.

Examples: Mn^{2+} Br S^{2-} Cu^{+} Al^{3+}

+2 -1 -2 +1 +3

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C) Polyatomic ions can be assigned a total oxidation number, which would be its ionic charge, or each element in the ion can be assigned its own oxidation number (Part D will explain how this is done).
Find the TOTAL oxidation number:

SO_4^{2-} OH PO_4^{3-}

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C) Polyatomic ions can be assigned a total oxidation number, which would be its ionic charge, or each element in the ion can be assigned its own oxidation number (Part D will explain how this is done).
Find the TOTAL oxidation number:

SO_4^{2-} OH PO_4^{3-}

2- 1- 3-

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D) To assign individual elements an oxidation number in a compound or polyatomic ion, perform the following steps in order:
1) alkali metals are ___
2) alkaline earth metals are ___
3) Other metals with only one possible oxidation number (combining capacity)
4) oxygen is ___
5) hydrogen is +1 (unless it's part of a metal hydride in which case it's -1)
Example: CaH_2: Ca is ___, so each H is ___.
6) halogens are -1 (but can also be +1, +3, or +5 in certain compounds).
7) Lastly, assign any 'hard to predict' atoms that are left over.

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D) To assign individual elements an oxidation number in a compound or polyatomic ion, perform the following steps in order:
1) alkali metals are +1
2) alkaline earth metals are +2
3) Other metals with only one possible oxidation number (combining capacity)
4) oxygen is -2
5) hydrogen is +1 (unless it's part of a metal hydride in which case it's -1)
Example: CaH_2: Ca is +2, so each H is -1.
6) halogens are -1 (but can also be +1, +3, or +5 in certain compounds).
7) Lastly, assign any 'hard to predict' atoms that are left over.

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*Since oxidation numbers are apparent for atoms in covalent compounds, sometimes an atom may have an oxidation number that's a fraction.

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Practice Questions: Determine the oxidation numbers of each element.

a) MnO_4 b) SO_2 c) SO_4^{2-} d) PCl_5

e) Cu_2O f) $HBiO_3$ g) OCl h) NH_2OH

i) NO_3 j) I_2 k) Cr^{3+} l) Cr

m) N_2O_5 n) $H_4P_2O_7$

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Practice Questions: Determine the oxidation numbers of each element.

a) MnO_4 +7 -2

b) SO_2 +4 -2

c) SO_4^{2-} +6 -2

d) PCl_5 +5 -1

e) Cu_2O +1 -2

f) $HBiO_3$ +1+5-2

g) OCl -2 +1

h) NH_2OH -1+1-2+1

i) NO_3 +5 -2

j) I_2 0

k) Cr^{3+} +3

l) Cr 0

m) N_2O_5 +5 -2

n) $H_4P_2O_7$ +1+5-2

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How are oxidation numbers helpful?

When an atom undergoes oxidation during a redox reaction, its oxidation number will _____ as it changes from reactant to product (it loses electrons, therefore it becomes more positive).

When an atom undergoes reduction during a redox reaction, its oxidation number will _____ as it changes from reactant to product (it gains electrons, therefore it becomes more negative).

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How are oxidation numbers helpful?

When an atom undergoes oxidation during a redox reaction, its oxidation number will increase as it changes from reactant to product (it loses electrons, therefore it becomes more positive).

When an atom undergoes reduction during a redox reaction, its oxidation number will decrease as it changes from reactant to product (it gains electrons, therefore it becomes more negative).

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Assigning oxidation numbers to reactants and products in a chemical reaction is helpful in determining whether the reaction is, in fact, a redox reaction, and if so, which reactant is oxidizing and which reactant is reducing.

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Practice Questions

1) For the following half-reactions (with electrons not present), how does the oxidation number of chromium change? Is the chromium oxidizing or reducing?

$Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$

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Practice Questions

1) For the following half-reactions (with electrons not present), how does the oxidation number of chromium change? Is the chromium oxidizing or reducing?

$Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$

+6 +2 +1 +3 +1 -2

+6 → +3; decreases by 3, therefore Cr from $Cr_2O_7^{2-}$ is reducing

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2) What happens to the oxidation number for nitrogen in the unbalanced half-reaction? Is nitrogen oxidizing or reducing?

$NO_2 \longrightarrow N_2O_3$

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2) What happens to the oxidation number for nitrogen in the unbalanced half-reaction? Is nitrogen oxidizing or reducing?

$NO_2 \longrightarrow N_2O_3$

+4 -2 +3 -2

+4 ⟶ +3; decrease by 1, therefore N from NO_2 is reducing

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3) Use oxidation numbers to identify which substance is oxidized and which is reduced in the following redox reactions.

a) $3Cu + 2NO_3^- + 8H^+ \longrightarrow 3Cu^{2+} + 2NO + 4H_2O$

b) $I_2 + 5HOBr + H_2O \longrightarrow 2IO_3^- + 5Br^- + 7H^+$

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3) Use oxidation numbers to identify which substance is oxidized and which is reduced in the following redox reactions.

a) $3Cu + 2NO_3^- + 8H^+ \longrightarrow 3Cu^{2+} + 2NO + 4H_2O$

0 +5 -2 +1 +2 +2-2 +1 -2

Cu is oxidizing; N from NO_3^- is reducing

b) $I_2 + 5HOBr + H_2O \longrightarrow 2IO_3^- + 5Br^- + 7H^+$

0 +1-2+1 +1 -2 +5 -2 -1 +1

I_2 is oxidizing; Br from HOBr is reducing

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4) Use oxidation numbers to determine whether the following reactions are redox reactions.

a) $2H_2O \longrightarrow 2H_2 + O_2$

b) $2AgCl + BaSO_4 \longrightarrow Ag_2SO_4 + BaCl_2$

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4) Use oxidation numbers to determine whether the following reactions are redox reactions.

a) $2H_2O \longrightarrow 2H_2 + O_2$

+1 -2 0 0

Redox

b) $2AgCl + BaSO_4 \longrightarrow Ag_2SO_4 + BaCl_2$

+1 -1 +2 -2 +1 -2 +2 -1

Not Redox

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Assignment 2

Read Hebden p.193-194

1) Do Hebden p.194 #3-6

2) For each unbalanced reaction, what is being oxidized and what is being reduced?

a) $S^{2-} + ClO_3^- \longrightarrow Cl + S$

b) $Cl_2 + SO_2 \longrightarrow 2Cl^- + SO_4^{2-}$

c) $Mn^{2+} + HBiO_3 \longrightarrow Bi^{3+} + MnO_4^-$

d) $FeS + NO_3^- \longrightarrow NO + SO_4^{2-} + Fe^{3+}$

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Assignment 2

Read Hebden p.193-194

1) Do Hebden p.194 #3-6

2) For each unbalanced reaction, what is being oxidized and what is being reduced?

a) $S^{2-} + ClO_3^- \longrightarrow Cl + S$ S^{2-} oxidized, Cl from ClO_3^- reduced

b) $Cl_2 + SO_2 \longrightarrow 2Cl^- + SO_4^{2-}$

S from SO_2 oxidized, Cl_2 reduced

c) $Mn^{2+} + HBiO_3 \longrightarrow Bi^{3+} + MnO_4^-$

Mn^{2+} oxidized, Bi from $HBiO_3$ reduced

d) $FeS + NO_3^- \longrightarrow NO + SO_4^{2-} + Fe^{3+}$

Fe from FeS oxidized, N from NO_3^- reduced

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3) Consider the following reaction:

$Zn_{(s)} + 2H^+_{(aq)} \quad Zn^{2+}_{(aq)} + H_{2(g)}$

The species being oxidized is:

A. H_2

B. H^+

C. Zn

D. Zn^{2+}

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3) Consider the following reaction:

$Zn_{(s)} + 2H^+_{(aq)} \quad Zn^{2+}_{(aq)} + H_{2(g)}$

The species being oxidized is:

A. H_2

B. H^+

C. Zn

D. Zn^{2+}

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4) When SO_4^{2-} reacts to form $S_2O_6^{2-}$, the sulphur atoms

A. lose electrons and are reduced

B. gain electrons and are reduced

C. lose electrons and are oxidized

D. gain electrons and are oxidized

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4) When SO_4^{2-} reacts to form $S_2O_6^{2-}$, the sulphur atoms

A. lose electrons and are reduced

B. gain electrons and are reduced

C. lose electrons and are oxidized

D. gain electrons and are oxidized

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5) In a reaction, the oxidation number of Cr decreases by 3. This indicates that Cr is

A. reduced

B. oxidized

C. neutralized

D. a reducing agent

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5) In a reaction, the oxidation number of Cr decreases by 3. This indicates that Cr is

A. reduced

B. oxidized

C. neutralized

D. a reducing agent

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6) Consider the following redox reaction:

$C_2H_5OH + 2Cr_2O_7^{2-} + 16H^+ \longrightarrow 2CO_2 + 4Cr^{3+} + 11H_2O$

Each carbon atom loses

A. 2 electrons

B. 4 electrons

C. 6 electrons

D. 12 electrons

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6) Consider the following redox reaction:

$C_2H_5OH + 2Cr_2O_7^{2-} + 16H^+ \longrightarrow 2CO_2 + 4Cr^{3+} + 11H_2O$

Each carbon atom loses

A. 2 electrons

B. 4 electrons

C. 6 electrons

D. 12 electrons

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IV) Table of Standard Reduction Potentials

Go to the last page of the data booklet and investigate the redox table.

What is listed on the left side? _____
Another description for the substances on the left are _____.

They are listed from top (strongest) to bottom (weakest). Notice that when you read the reactions from left to right, each is a reduction.

Notice that many of the strongest oxidizing agents are halogens and oxyanions.

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IV) Table of Standard Reduction Potentials

Go to the last page of the data booklet and investigate the redox table.

What is listed on the left side? oxidizing agents
Another description for the substances on the left are things that reduce.

They are listed from top (strongest) to bottom (weakest). Notice that when you read the reactions from left to right, each is a reduction.

Notice that many of the strongest oxidizing agents are halogens and oxyanions.

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What is on the right side? _____
Another description for the substances on the right side are _____.

These are listed from strongest (BOTTOM) to weakest (TOP), like the bases on the acid-base table.
Notice that the oxidation half-reactions are from right to left on the table, as this shows electron(s) being given away.
What types of substance are the strongest reducing agents?

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What is on the right side? reducing agents
Another description for the substances on the right side are things that oxidize.

These are listed from strongest (BOTTOM) to weakest (TOP), like the bases on the acid-base table.
Notice that the oxidation half-reactions are from right to left on the table, as this shows electron(s) being given away.
What types of substance are the strongest reducing agents? metals (like to oxidize to become cations

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The double arrow does not mean the reactions are at equilibrium; it means that the half-reactions can occur in either direction depending on the substances present in a reaction. Once you know the direction to use for a specific half-reaction (depending on whether it's an oxidation or reduction), use only a one way arrow.

Why do some substances such as Cu^+, H_2O_2, and Fe^{2+} appear on both sides of the table?

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The double arrow does not mean the reactions are at equilibrium; it means that the half-reactions can occur in either direction depending on the substances present in a reaction. Once you know the direction to use for a specific half-reaction (depending on whether it's an oxidation or reduction), use only a one way arrow.

Why do some substances such as Cu^+, H_2O_2, and Fe^{2+} appear on both sides of the table?

they can oxidize in the presence of a stronger oxidizing agent, and reduce in the presence of a stronger reducing agent

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Some metals have more than one common oxidation number, and therefore will have multiple half-reactions on the table (such as iron and copper). Locate all of the copper half-reactions on the table.

Be sure you use the correct half-reaction when dealing with these metals.

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Some metals have more than one common oxidation number, and therefore will have multiple half-reactions on the table (such as iron and copper). Locate all of the copper half-reactions on the table.

$Cu^{2+} + 2e^- \longrightarrow Cu$
$Cu^{2+} + e^- \longrightarrow Cu^+$
$Cu^+ + e^- \longrightarrow Cu$

Be sure you use the correct half-reaction when dealing with these metals.

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Some half-reactions require **acidic** conditions (meaning H^+ must be present) and some require **basic** conditions (meaning OH must be present).

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V) Predicting Spontaneity

How might you know whether a redox reaction is spontaneous (_____) or not?
Remember, in order for a redox reaction to occur, there must be an oxidation AND a reduction occurring simultaneously.

The first thing to do is figure out what substance is reducing and what substance is oxidizing. Then jot down the half-reactions for each.

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V) Predicting Spontaneity

How might you know whether a redox reaction is spontaneous (happens on its own) or not?
Remember, in order for a redox reaction to occur, there must be an oxidation AND a reduction occurring simultaneously.

The first thing to do is figure out what substance is reducing and what substance is oxidizing. Then jot down the half-reactions for each.

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1) If the reduction half-reaction is **higher** on the table than the oxidation half-reaction, the reaction is SPONTANEOUS.

2) If the reduction half-reaction is **lower** on the table than the oxidation half-reaction (or if it's on the same line), the reaction is _____.

3) No redox reaction can occur if you have substances that only reduce or substances that only oxidize (both on the same side of the table).

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1) If the reduction half-reaction is **higher** on the table than the oxidation half-reaction, the reaction is SPONTANEOUS.

2) If the reduction half-reaction is **lower** on the table than the oxidation half-reaction (or if it's on the same line), the reaction is non-spontaneous.

3) No redox reaction can occur if you have substances that only reduce or substances that only oxidize (both on the same side of the table).

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Practice Questions

1) Are the following reactions spontaneous?

a) $CuSO_4 + Zn \longrightarrow Cu + ZnSO_4$

http://ibchem.com/IB/ibfiles/redox/red_swf/zn%2Bcu.swf

b) $Zn + Ag_2S \longrightarrow 2Ag + S^{2-} + Zn^{2+}$

c) $S + NO_3 + 6H^+ \longrightarrow NO + H_2S + 2H_2O$

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Practice Questions

1) Are the following reactions spontaneous?

a) $CuSO_4 + Zn \longrightarrow Cu + ZnSO_4$

reduc: $Cu^{2+} + 2e \longrightarrow Cu$
oxid: $Zn \longrightarrow Zn^{2+} + 2e^-$

reduction higher on table than oxidation, therefore SPONTANEOUS

http://ibchem.com/IB/ibfiles/redox/red_swf/zn%2Bcu.swf

b) $Zn + Ag_2S \longrightarrow 2Ag + S^{2-} + Zn^{2+}$

reduc: $Ag_2S + 2e^- \longrightarrow 2Ag + S^{2-}$
oxid: $Zn \longrightarrow Zn^{2+} + 2e^-$

reduction higher than oxidation, therefore SPONTANEOUS

c) $S + NO_3 + 6H^+ \longrightarrow NO + H_2S + 2H_2O$

reduc: $NO_3^- + 4H^+ + 3e^- \longrightarrow NO + 2H_2O$
$S + 2H^+ + 2e^- \longrightarrow H_2S$

NO REACTION POSSIBLE

oxid: NONE

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2) Which metal can be oxidized by I_2 but not by an acidic solution of SO_4^{2-}?

3) Which metal can act as a reducing agent for Sn^{2+} but not for Co^{2+}?

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2) Which metal can be oxidized by I_2 but not by an acidic solution of SO_4^{2-}?

looking for a metal that can be oxidized, therefore the answer is on the RIGHT side of the table below I_2 but above SO_4^{2-} with H^+

Cu

3) Which metal can act as a reducing agent for Sn^{2+} but not for Co^{2+}?

looking for a metal that will act as a reducing agent, therefore the answer is on the RIGHT side of the table below Sn^{2+} but above Co^{2+}

Ni

Aug 27-11:56 AM

Assignment 3

1) Which of the following reactions occur spontaneously?

a) $Cr^{3+} + Ba \Rightarrow Ba^{2+} + Cr$

b) $I_2 + 2Cl^- \Rightarrow Cl_2 + 2I^-$

Aug 27-11:59 AM

Assignment 3

1) Which of the following reactions occur spontaneously?

a) $Cr^{3+} + Ba \Rightarrow Ba^{2+} + Cr$

spontaneous

b) $I_2 + 2Cl^- \Rightarrow Cl_2 + 2I^-$

non-spontaneous

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c) $2NO_3^- + 8H^+ + 3Ni \Rightarrow 2NO + 4H_2O + 3Ni^{2+}$

d) $2Al^{3+} + 3Ca^{2+} \Rightarrow 2Al + 3Ca$

Aug 27-12:01 PM

c) $2NO_3^- + 8H^+ + 3Ni \Rightarrow 2NO + 4H_2O + 3Ni^{2+}$

spontaneous

d) $2Al^{3+} + 3Ca^{2+} \Rightarrow 2Al + 3Ca$

no redox reaction possible

Aug 27-12:01 PM

2) What products are formed when NO_3^- in acidic solution is reduced by $Fe_{(s)}$

3) Read Heben p.195-199 and do p.199 #8ace, 9ac, 10ac, 11aeg, 12

Aug 27-12:02 PM

2) What products are formed when NO_3^- in acidic solution is reduced by $Fe_{(s)}$

NO, H_2O, Fe^{2+}

3) Read Heben p.195-199 and do p.199 #8ace, 9ac, 10ac, 11aeg, 12

answers in the back of Hebden

Aug 27-12:02 PM

VI) Redox Logic Problems

Strategy for solving redox logic problems:
i) separate all substances into oxidizing agents & reducing agents
ii) write the half-reactions
iii) use the given information to order the half-reactions into a mini table
iv) solve the problem using the mini table

Aug 28-7:44 AM

Example:
Given the following information:

$Z + K \longrightarrow Z^- + K$

$K + M \longrightarrow$ no spontaneous reaction

List the reducing agents from strongest to weakest (M is a non-metal and becomes M when reduced).

Aug 28-7:45 AM

Example:
Given the following information:

$Z + K \longrightarrow Z^- + K$

$K + M \longrightarrow$ no spontaneous reaction

List the reducing agents from strongest to weakest (M is a non-metal and becomes M when reduced).

i) O.A. R.A.
K K^-
Z Z^-
M M^-

ii) Build half reactions
O.A. R.A.
$K + e \rightleftharpoons K^-$
$Z + e^- \rightleftharpoons Z^-$
$M + e^- \rightleftharpoons M^-$

iii) order the half-reactions using the rxn info above:
- in the first rxn, Z is reducing and the rxn is spontaneous, so Z rxn higher than K rxn
- in second rxn, M is reducing but rxn is non-spontaneous, so oxidation (K) is higher than reduction (M)....K reaction higher than M rxn

O.A. R.A.
$Z + e^- \rightleftharpoons Z^-$
$K + e \rightleftharpoons K^-$
$M + e^- \rightleftharpoons M^-$

iv) R.A. from strongest to weakest:
M^-, K^-, Z^-

Aug 28-7:45 AM

Example:
Given the following information:

$A^- + L^+ \longrightarrow$ no spontaneous reaction

$A + P \longrightarrow A + P^-$

List the oxidizing agents from strongest to weakest (L is a metal and becomes L^+ when oxidized).

Aug 28-7:56 AM

Example:
Given the following information:

$A^- + L^+ \longrightarrow$ no spontaneous reaction

$A + P \longrightarrow A + P^-$

List the oxidizing agents from strongest to weakest (L is a metal and becomes L^+ when oxidized).

i) O.A. R.A.
A A^-
P P^-
L^+ L

ii) Build half reactions
O.A. R.A.
$A + e \rightleftharpoons A$
$P + e^- \rightleftharpoons P^-$
$L^+ + e \rightleftharpoons L$

iii) order the half-reactions using the rxn info above:
- in the first rxn, L^+ is reducing and the rxn is non-spontaneous, so A rxn higher than L rxn
- in second rxn, P is reducing and rxn is spontaneous, so P rxn higher than A rxn

O.A. R.A.
$P + e^- \rightleftharpoons P$
$A + e \rightleftharpoons A^-$
$L^+ + e^- \rightleftharpoons L$

iv) O.A. from strongest to weakest:
P, A, L^+

Aug 28-7:56 AM

Assignment 4

1) Do Hebden p.200 #14-18

Aug 28-8:01 AM

Assignment 4

1) Do Hebden p.200 #14-18

answers in back of Hebden

Aug 28-8:01 AM

Assignment 5

1) Consider the following oxidation-reduction reaction:

$2Mn^{2+} + 2IO_3^- + 2H_2O \longrightarrow 2MnO_4^- + I_2 + 4H^+$

The reducing agent is

A. I_2
B. IO_3^-
C. H_2O
D. Mn^{2+}

Aug 28-8:09 AM

Assignment 5

1) Consider the following oxidation-reduction reaction:

$2Mn^{2+} + 2IO_3^- + 2H_2O \longrightarrow 2MnO_4^- + I_2 + 4H^+$

The reducing agent is

A. I_2
B. IO_3^-
C. H_2O
D. Mn^{2+}

Aug 28-8:09 AM

2) Consider the reaction:
$Cu^{2+}_{(aq)} + Ni_{(s)} \longrightarrow Ni^{2+}_{(aq)} + Cu_{(s)}$
This reaction will proceed spontaneously because Cu^{2+}

A. is more easily oxidized than Ni^{2+}
B. is a weaker reducing agent than Ni^{2+}
C. is a stronger reducing agent than Ni^{2+}
D. gains electrons more readily than Ni^{2+}

Aug 28-8:11 AM

2) Consider the reaction:
$Cu^{2+}_{(aq)} + Ni_{(s)} \longrightarrow Ni^{2+}_{(aq)} + Cu_{(s)}$
This reaction will proceed spontaneously because Cu^{2+}

A. is more easily oxidized than Ni^{2+}
B. is a weaker reducing agent than Ni^{2+}
C. is a stronger reducing agent than Ni^{2+}
D. gains electrons more readily than Ni^{2+}

Aug 28-8:11 AM

3) A piece of zinc metal is dropped into a solution of $FeCl_2$. The result of this procedure is

A. no reaction
B. the zinc is oxidized by Cl_2
C. the iron is oxidized by Zn^{2+}
D. the zinc is oxidized by Fe^{2+}

Aug 28-8:13 AM

3) A piece of zinc metal is dropped into a solution of $FeCl_2$. The result of this procedure is

A. no reaction
B. the zinc is oxidized by Cl_2
C. the iron is oxidized by Zn^{2+}
D. the zinc is oxidized by Fe^{2+}

Aug 28-8:13 AM

4) Consider the redox reaction:
$2BrO_3^- + 10Cl^- + 12H^+ \longrightarrow Br_2 + 5Cl_2 + 6H_2O$
The oxidation half-reaction involved in this reaction is

A. $2Cl^- \longrightarrow Cl_2 + 2e^-$
B. $2H^+ \longrightarrow H_2 + 2e^-$
C. $2BrO_3^- + 12H^+ + 10e^- \longrightarrow Br_2 + 6H_2O$
D. $2BrO_3^- + 12H^+ \longrightarrow Br_2 + 6H_2O + 10e^-$

Aug 28-8:14 AM

4) Consider the redox reaction:
$2BrO_3^- + 10Cl^- + 12H^+ \longrightarrow Br_2 + 5Cl_2 + 6H_2O$
The oxidation half-reaction involved in this reaction is

A. $2Cl^- \longrightarrow Cl_2 + 2e^-$
B. $2H^+ \longrightarrow H_2 + 2e^-$
C. $2BrO_3^- + 12H^+ + 10e^- \longrightarrow Br_2 + 6H_2O$
D. $2BrO_3^- + 12H^+ \longrightarrow Br_2 + 6H_2O + 10e^-$

Aug 28-8:14 AM

5) Oxidation is the process involving a

A. gain of electrons
B. loss of protons
C. loss of electrons
D. gain of electrons

Aug 28-8:18 AM

5) Oxidation is the process involving a

A. gain of electrons
B. loss of protons
C. loss of electrons
D. gain of electrons

Aug 28-8:18 AM

6) In a reaction, the oxidation number of S changes from 6 to 4. This information tells us that sulphur has acted as

A. a reducing agent, losing 2e
B. a reducing agent, gaining 2e
C. an oxidizing agent, losing 2e
D. an oxidizing agent, gaining 2e

Aug 28-8:19 AM

6) In a reaction, the oxidation number of S changes from 6 to 4. This information tells us that sulphur has acted as

A. a reducing agent, losing 2e
B. a reducing agent, gaining 2e
C. an oxidizing agent, losing 2e
D. an oxidizing agent, gaining 2e

Aug 28-8:19 AM

7) Consider the reaction:
$2H_2O + Al + MnO_4^- \longrightarrow Al(OH)_4^- + MnO_2$
The substance which undergoes reduction is

A. Al
B. H_2O
C. MnO_4^-
D. $Al(OH)_4^-$

Aug 28-8:21 AM

7) Consider the reaction:
$2H_2O + Al + MnO_4^- \longrightarrow Al(OH)_4^- + MnO_2$
The substance which undergoes reduction is

A. Al
B. H_2O
C. MnO_4^-
D. $Al(OH)_4^-$

Aug 28-8:21 AM

8) Consider the reaction: $2H^+ + 2e^- \longrightarrow H_2$
The reaction represents

A. oxidation
B. reduction
C. electrolysis
D. replacement

Aug 28-8:22 AM

8) Consider the reaction: $2H^+ + 2e^- \longrightarrow H_2$
The reaction represents

A. oxidation
B. reduction
C. electrolysis
D. replacement

Aug 28-8:22 AM

9) Which of the following is non-spontaneous?

A. $Pb(NO_3)_2 + Ni \longrightarrow Pb + Ni(NO_3)_2$
B. $2AgNO_3 + Ni \longrightarrow 2Ag + Ni(NO_3)_2$
C. $Co(NO_3)_2 + Ni \longrightarrow Co + Ni(NO_3)_2$
D. $2Au(NO_3)_3 + 3Ni \longrightarrow 2Au + 3Ni(NO_3)_2$

Aug 28-8:24 AM

9) Which of the following is non-spontaneous?

A. $Pb(NO_3)_2 + Ni \longrightarrow Pb + Ni(NO_3)_2$
B. $2AgNO_3 + Ni \longrightarrow 2Ag + Ni(NO_3)_2$
C. $Co(NO_3)_2 + Ni \longrightarrow Co + Ni(NO_3)_2$
D. $2Au(NO_3)_3 + 3Ni \longrightarrow 2Au + 3Ni(NO_3)_2$

Aug 28-8:24 AM

10) Which one of the following is the strongest reducing agent?

A. Al
B. Ag
C. Ag^+
D. Al^{3+}

Aug 28-8:26 AM

10) Which one of the following is the strongest reducing agent?

A. Al
B. Ag
C. Ag^+
D. Al^{3+}

Aug 28-8:26 AM

11) The oxidation number of an element decreases during a reaction. This implies that atoms of that element

A. lost electrons
B. were oxidized
C. gained electrons
D. acted as a reducing agent

Aug 28-8:27 AM

11) The oxidation number of an element decreases during a reaction. This implies that atoms of that element

A. lost electrons
B. were oxidized
C. gained electrons
D. acted as a reducing agent

Aug 28-8:27 AM

12) Which of the following is the strongest oxidizing agent?

A. Cu^{2+}
B. Pb^{2+}
C. Ni^{2+}
D. Sn^{2+}

Aug 28-8:28 AM

12) Which of the following is the strongest oxidizing agent?

A. Cu^{2+}
B. Pb^{2+}
C. Ni^{2+}
D. Sn^{2+}

Aug 28-8:28 AM

13) Metallic platinum reacts spontaneously with $Au^{3+}_{(aq)}$, but does not react with $Ag^{+}_{(aq)}$. The metals, in order of increasing strength as reducing agents, are

A. Ag, Pt, Au
B. Pt, Au, Ag
C. Au, Ag, Pt
D. Au, Pt, Ag

Aug 28-8:29 AM

13) Metallic platinum reacts spontaneously with $Au^{3+}_{(aq)}$, but does not react with $Ag^{+}_{(aq)}$. The metals, in order of increasing strength as reducing agents, are

A. Ag, Pt, Au
B. Pt, Au, Ag
C. Au, Ag, Pt
D. Au, Pt, Ag

Aug 28-8:29 AM

14) Which of the following pairs of ions will react spontaneously in solution?

A. Cu^{2+} and Fe^{2+}
B. Pb^{2+} and Sn^{2+}
C. Co^{2+} and Cr^{2+}
D. Mn^{2+} and Cr^{2+}

Aug 28-8:30 AM

14) Which of the following pairs of ions will react spontaneously in solution?

A. Cu^{2+} and Fe^{2+}
B. Pb^{2+} and Sn^{2+}
C. Co^{2+} and Cr^{2+}
D. Mn^{2+} and Cr^{2+}

Aug 28-8:30 AM

15) When NO_2 reacts to form N_2O_4, the oxidation number of nitrogen

A. increases by 2
B. increases by 4
C. increases by 8
D. does not change

Aug 28-8:32 AM

15) When NO_2 reacts to form N_2O_4, the oxidation number of nitrogen

A. increases by 2
B. increases by 4
C. increases by 8
D. does not change

Aug 28-8:32 AM

VII) Writing Balanced Equations for Redox Reactions Using the Standard Reduction Table

Steps:
1) Find the appropriate reduction and oxidation half-reactions from the table, and write them down, on above the other.
2) Balance electrons.
3) Cancel where appropriate and write the balanced equation. Electrons should cancel and not be written in the overall redox reaction (electrons should only be present in half-reactions).

Aug 28-8:33 AM

Practice Questions

Write a balanced reaction for the following reactants.
1. MnO_4 (acidic) and H_2SO_3

Aug 28-8:36 AM

Practice Questions

Write a balanced reaction for the following reactants.

1. MnO_4 (acidic) and H_2SO_3

reduction: $(MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O)$ x 2

oxidation: $(H_2SO_3 + H_2O \longrightarrow SO_4^{2-} + 4H^+ + 2e^-)$ x 5

$$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$$
$$5H_2SO_3 + 5H_2O \longrightarrow 5SO_4^{2-} + 20H^+ + 10e^-$$

$$2MnO_4 + 5H_2SO_3 \longrightarrow 2Mn^{2+} + 3H_2O + 5SO_4^{2-} + 4H^+$$

Aug 28-8:36 AM

2) Cu and NO_3^- (acidic) to produce Cu^{2+} and NO

Aug 28-8:42 AM

2) Cu and NO_3^- (acidic) to produce Cu^{2+} and NO

R: $(NO_3 + 4H^+ + 3e \longrightarrow NO + 2H_2O)$ x 2

O: $(Cu \longrightarrow Cu^{2+} + 2e^-)$ x 3

$$2NO_3^- + 8H^+ + 6e^- \longrightarrow 2NO + 4H_2O$$
$$3Cu \longrightarrow 3Cu^{2+} + 6e^-$$

$$2NO_3^- + 8H^+ + 3Cu \longrightarrow 2NO + 4H_2O + 3Cu^{2+}$$

Aug 28-8:42 AM

Assignment 6

Write a balanced equation for each of the following.

a) H_2O_2 (acidic) and $N_2O_{4(aq)}$

Aug 28-8:47 AM

Assignment 6

Write a balanced equation for each of the following.

a) H_2O_2 (acidic) and $N_2O_{4(aq)}$

$$H_2O_2 + 2H^+ + 2e \longrightarrow 2H_2O$$
$$N_2O_4 + 2H_2O \longrightarrow 2NO_3 + 4H^+ + 2e^-$$

$$N_2O_4 + H_2O_2 \longrightarrow 2NO_3^- + 2H^+$$

Aug 28-8:47 AM

b) Ag^+ and $H_2S_{(g)}$

Aug 28-8:53 AM

b) Ag^+ and $H_2S_{(g)}$

$(Ag^+ + e^- \longrightarrow Ag) \times 2$
$H_2S \longrightarrow S + 2H^+ + 2e^-$

$2Ag^+ + 2e^- \longrightarrow 2Ag$
$H_2S \longrightarrow S + 2H^+ + 2e^-$

$2Ag^+ + H_2S \longrightarrow 2Ag + S + 2H^+$

Aug 28-8:53 AM

c) IO_3^- (acidic) and H_2O_2

Aug 28-8:56 AM

c) IO_3^- (acidic) and H_2O_2

$(IO_3^- + 6H^+ + 5e^- \longrightarrow 1/2I_2 + 3H_2O) \times 2$
$(H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-) \times 5$

$2IO_3^- + 12H^+ + 10e^- \longrightarrow I_2 + 6H_2O$
$5H_2O_2 \longrightarrow 5O_2 + 10H^+ + 10e$

$2IO_3 + 2H^+ + 5H_2O_2 \longrightarrow I_2 + 6H_2O + 5O_2$

Aug 28-8:56 AM

d) H_3PO_4 (acidic) and NO

Aug 28-9:01 AM

d) H_3PO_4 (acidic) and NO

$(H_3PO_4 + 2H^+ + 2e \longrightarrow H_3PO_3 + H_2O) \times 3$
$(NO + 2H_2O \longrightarrow NO_3^- + 4H^+ + 3e^-) \times 2$

$3H_3PO_4 + 6H^+ + 6e^- \longrightarrow 3H_3PO_3 + 3H_2O$
$2NO + 4H_2O \longrightarrow 2NO_3^- + 8H^+ + 6e^-$

$3H_3PO_4 + 2NO + H_2O \longrightarrow 3H_3PO_3 + 2NO_3^- + 2H^+$

Aug 28-9:01 AM

VIII) Balancing Half-Reactions Without the Use of the Reduction Table

Using the following guidelines, it is possible to build and balance half-reactions that are not on the redox table starting only with a skeleton half-reaction.

Guidelines: **'MAJOR HYDROXIDE'**
'MAJOR OH^-'

Aug 28-9:06 AM

1) **MAJOR**: Balance all major elements (all elements except O and H).
2) **'O'**: Balance oxygen by adding H_2O molecules to the applicable side.
3) **'H'**: Balance hydrogen by adding H^+ ions to the applicable side
4) $^-$ **:** Balance the charge by adding electrons to the applicable side.
*5) If necessary: If the half-reaction is **basic**, you must use the equation $H_2O \rightleftharpoons H^+ + OH^-$ (can also be written the other way around) to cancel protons (which are acidic) from the half-reaction and end up with OH ions (which are basic).

Aug 28-9:08 AM

Practice Questions (more examples on Hebden p.201-203)

1) Balance the half-reaction whereby NO is reduced to N_2O in acidic solution.
*Skeleton half-reaction: $NO \longrightarrow N_2O$

Aug 28-9:11 AM

Practice Questions (more examples on Hebden p.201-203)

1) Balance the half-reaction whereby NO is reduced to N_2O in acidic solution.
*Skeleton half-reaction: $NO \longrightarrow N_2O$

$NO \longrightarrow N_2O$
MAJOR $2NO \longrightarrow N_2O$
OXYGEN $2NO \longrightarrow N_2O + H_2O$
HYDROG $2H^+ + 2NO \longrightarrow N_2O + H_2O$
CHARGE $2e + 2H^+ + 2NO \longrightarrow N_2O + H_2O$

Aug 28-9:11 AM

2) Balance the half-reaction whereby Mn^{2+} is oxidized to MnO_2 in acidic solution

$Mn^{2+} \longrightarrow MnO_2$

Aug 28-9:15 AM

2) Balance the half-reaction whereby Mn^{2+} is oxidized to MnO_2 in acidic solution

$Mn^{2+} \longrightarrow MnO_2$

$2H_2O + Mn^{2+} \longrightarrow MnO_2 + 4H^+ + 2e$

Aug 28-9:15 AM

3) Balance the half-reaction whereby HO_2 is oxidized to O_2 in **basic** solution.

$HO_2 \longrightarrow O_2$

Aug 28-9:17 AM

3) Balance the half-reaction whereby HO_2 is oxidized to O_2 in **basic** solution.

$HO_2 \longrightarrow O_2$

acidic: $HO_2 \longrightarrow O_2 + H^+ + 2e$

$H^+ + OH^- \longrightarrow H_2O$

basic: $HO_2 + OH^- \longrightarrow O_2 + H_2O + 2e^-$

Aug 28-9:17 AM

4) Balance the following half-reaction in **basic** solution. $Cu_2O \longrightarrow Cu(OH)_2$

Aug 28-9:20 AM

4) Balance the following half-reaction in **basic** solution. $Cu_2O \longrightarrow Cu(OH)_2$

acidic: $3H_2O + Cu_2O \longrightarrow 2Cu(OH)_2 + 2H^+ + 2e^-$

$2H^+ + 2OH^- \longrightarrow 2H_2O$

basic: $H_2O + Cu_2O + 2OH^- \longrightarrow 2Cu(OH)_2 + 2e^-$

Aug 28-9:20 AM

Assignment 7

1) Do Hebden p.203 #19a-m

Aug 28-9:23 AM

Assignment 7

1) Do Hebden p.203 #19a-m

answers in the back of Hebden

Aug 28-9:23 AM

IX) Balancing Full Redox Reactions not on the Redox Data Table

Steps:

1) Figure out which reactant substances match up with which product substances to build skeleton half-reactions. Make sure all 'major' elements are present in a half-reaction from the start.

2) Balance each half-reaction using **MAJOR OH** guidelines. However, do not convert to **basic** (if asked) until after step 4 of this list.

3) Write newly constructed half-reactions one on top of the other and balance electrons.

Aug 28-9:24 AM

4) Put the half-reactions together and cancel where necessary (electrons should always cancel out).
*5. If **necessary**: Convert to **basic** conditions in same manner as previously. If there are no H^+ ions left to convert to basic, then no conversion is possible or necessary.

Aug 28-9:27 AM

Practice Questions

Balance each of the following redox reactions.

1) $H_2PO_2^- + CNO \longrightarrow CN + HPO_3^-$ (acidic)

Aug 28-9:28 AM

Practice Questions

Balance each of the following redox reactions.

1) $H_2PO_2^- + CNO \longrightarrow CN + HPO_3^-$ (acidic)

$H_2PO_2^- \longrightarrow HPO_3^-$

$(H_2O + H_2PO_2^- \longrightarrow HPO_3 + 3H^+ + 3e^-) \times 2$

$CNO \longrightarrow CN$

$(2H^+ + 2e + CNO \longrightarrow CN + H_2O) \times 3$

$2H_2O + 2H_2PO_2^- \longrightarrow 2HPO_3 + 6H^+ + 6e^-$

$6H^+ + 6e^- + 3CNO \longrightarrow 3CN^- + 3H_2O$

$2H_2PO_2^- + 3CNO^- \longrightarrow 2HPO_3^- + 3CN^- + H_2O$

Aug 28-9:28 AM

Disproportionation Reaction: a redox reaction in which the same substance is both reduced AND oxidized to make two different products.
2) Balance the following disproportionation reaction under basic conditions:

$$P_4 \longrightarrow H_2PO_2^- + PH_3$$

Aug 28-4:38 PM

Disproportionation Reaction: a redox reaction in which the same substance is both reduced AND oxidized to make two different products.
2) Balance the following disproportionation reaction under basic conditions:

$$P_4 \longrightarrow H_2PO_2^- + PH_3$$

$P_4 \longrightarrow H_2PO_2^-$

$(8H_2O + P_4 \longrightarrow 4H_2PO_2^- + 8H^+ + 4e^-) \times 3$

$P_4 \longrightarrow PH_3$

$12e^- + 12H^+ + P_4 \longrightarrow 4PH_3$

$24H_2O + 3P_4 \longrightarrow 12H_2PO_2 + 24H^+ + 12e^-$

$12e^- + 12H^+ + P_4 \longrightarrow 4PH_3$

$24H_2O + 4P_4 \longrightarrow 12H_2PO_2 + 12H^+ + 4PH_3$

lowest terms: $6H_2O + P_4 \longrightarrow 3H_2PO_2 + 3H^+ + PH_3$

$3H^+ + 3OH^- \longrightarrow 3H_2O$

$3H_2O + P_4 + 3OH \longrightarrow 3H_2PO_2 + PH_3$

Aug 28-4:38 PM

3) $H_2O_2 + SCN^- \longrightarrow NH_4^+ + H_2O + HCO_3^- + HSO_4^-$ (acidic)

Aug 28-4:45 PM

3) $H_2O_2 + SCN^- \longrightarrow NH_4^+ + H_2O + HCO_3^- + HSO_4^-$ (acidic)

$SCN^- \longrightarrow NH_4^+ + HCO_3^- + HSO_4^-$

$7H_2O + SCN \longrightarrow NH_4^+ + HCO_3^- + HSO_4^- + 8H^+ + 8e^-$

$H_2O_2 \longrightarrow H_2O$

$(2H^+ + 2e^- + H_2O_2 \longrightarrow 2H_2O) \times 4$

$8H^+ + 8e^- + 4H_2O_2 \longrightarrow 8H_2O$

$7H_2O + SCN^- \longrightarrow NH_4^+ + HCO_3^- + HSO_4^- + 8H^+ + 8e^-$

$4H_2O_2 + SCN \longrightarrow NH_4^+ + HCO_3^- + HSO_4^- + H_2O$

Aug 28-4:45 PM

Assignment 8

1) Do Hebden p.207 #24adijqw

Aug 28-4:50 PM

Assignment 8

1) Do Hebden p.207 #24adijqw

answers in the back of Hebden

Aug 28-4:50 PM

X) Redox Titrations

Similar to acid-base titrations, redox titrations are also useful for determining the unknown concentration of a solution. However, the titration reaction is a redox reaction rather than an acid-base reaction.

Aug 28-4:51 PM

A common solution used for titrations is acidified aqueous potassium permanganate, $KMnO_{4(aq)}$ (actually $K^+_{(aq)}$ and $MnO_4^-{}_{(aq)}$), because permanganate (MnO_4^-) is a strong oxidizing agent (top left on the table) when in the presence of an acid, and there is a built in colour change as MnO_4 is purple and reacts to become colourless $Mn^{2+}_{(aq)}$:

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

purple → colourless

Aug 28-4:53 PM

Purple MnO_4^- solution (from the buret) will spontaneously oxidize most reducing agents of unknown concentration (in the flask). How do you know this?

As purple MnO_4^- ions drop into the flask, they will react with the reducing agent to produce Mn^{2+} ions, so the mixture in the flask will remain colourless. Once all of the reducing agent from the flask has reacted with MnO_4 , the equivalence point has been reached, so the next drop of MnO_4 solution added will ...

Aug 29-7:51 AM

Purple MnO_4^- solution (from the buret) will spontaneously oxidize most reducing agents of unknown concentration (in the flask). How do you know this? b/c it's high on the top left of the redox table

As purple MnO_4^- ions drop into the flask, they will react with the reducing agent to produce Mn^{2+} ions, so the mixture in the flask will remain colourless. Once all of the reducing agent from the flask has reacted with MnO_4 , the equivalence point has been reached, so the next drop of MnO_4 solution added will ...

turn the solution the flask purple, as MnO_4^- no longer has anything to react with. This is the endpoint of the titration.

Aug 29-7:51 AM

Practice Question

When 25.00mL of Cr^{3+} solution is titrated with 0.300M $KMnO_4$ solution, the titration takes 28.32mL of $KMnO_4$ solution to reach the endpoint, according to the following equation:

$6MnO_4 + 10Cr^{3+} + 11H_2O \longrightarrow 6Mn^{2+} + 5Cr_2O_7^{2-} + 22H^+$

Calculate the original $[Cr^{3+}]$.

Aug 29-7:55 AM

Practice Question

When 25.00mL of Cr^{3+} solution is titrated with 0.300M $KMnO_4$ solution, the titration takes 28.32mL of $KMnO_4$ solution to reach the endpoint, according to the following equation:

$6MnO_4 + 10Cr^{3+} + 11H_2O \longrightarrow 6Mn^{2+} + 5Cr_2O_7^{2-} + 22H^+$

Calculate the original $[Cr^{3+}]$.

$KMnO_4 \longrightarrow K^+ + MnO_4^-$

0.300M 0.300M

moles MnO_4 = (0.300M)(0.02832L) = 0.008496mol

$$0.008496\text{mol} \times \frac{10\text{mol } Cr^{3+}}{6\text{mol } MnO_4^-} = 0.01416\text{mol } Cr^{3+}$$

$$[Cr^{3+}] = \frac{0.01416\text{mol}}{0.02500\text{L}} = 0.566\text{M}$$

Aug 29-7:55 AM

The dichromate ion ($Cr_2O_7^{2-}$), which creates an orange solution, can be used as an oxidizing agent for many redox titrations if in an acidic environment, as it's high on the left side of the table, and it reacts to produce the Cr^{3+} ion, which forms a green solution.

Aug 29-8:01 AM

Assignment 9

1) Do Hebden p.213-214 #26, 28, 29

Aug 29-8:03 AM

Assignment 9

1) Do Hebden p.213-214 #26, 28, 29

answers in the back of Hebden

Aug 29-8:03 AM

XI) The Electrochemical Cell

What is another name for an electrochemical cell?

What is the point of an electrochemical cell?

Aug 29-8:03 AM

XI) The Electrochemical Cell

What is another name for an electrochemical cell? battery

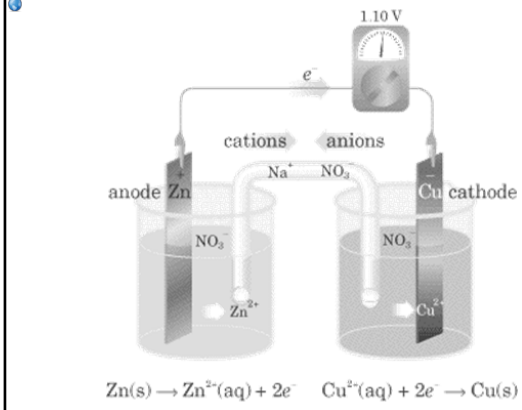
What is the point of an electrochemical cell?

to get electrons moving through an external circuit, thereby creating electricity

Aug 29-8:03 AM

How does an electrochemical cell work?

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/CuZncell.html



Aug 29-8:06 AM

AN OX CARED

Aug 29-8:08 AM

AN OX CARED

anode oxidation cathode reduction

Aug 29-8:08 AM

Description of the Cu/Zn Electrochemical Cell

Oxidation takes place at the anode, as Zn metal atoms lose two electrons to become Zn^{2+} cations ($Zn \longrightarrow Zn^{2+} + 2e^-$) which then dissolve in the $Zn(NO_3)_{2(aq)}$ solution. The electrons travel up the Zn electrode and through the wire to the cathode Cu electrode where they are used to reduce Cu^{2+} cations to Cu metal atoms ($Cu^{2+} + 2e \longrightarrow Cu$). The electrons travel from the oxidation to the reduction half-cell because they feel the 'pull' for electrons from the Cu^{2+} ions (similar to how we feel the 'pull' from gravity). The Cu^{2+} ions are attracted to the cathode (Cu metal strip) by the electrons that are arriving through the wire.

Aug 29-8:09 AM

The anode Zn electrode loses mass (as Zn atoms make Zn^{2+} ions that dissolve in solution) and the cathode Cu electrode gains mass (as Cu^{2+} ions make Cu atoms). The point of the whole process is to get electrons moving through the wire. Then you can hook up and electrically power a device such as a light bulb.
To balance the positive charge build-up in the solution at the Zn anode (due to the production of Zn^{2+}), NO_3 anions from the salt bridge migrate into the anode solution and Zn^{2+} cations migrate out of the anode solution and into the salt bridge.

Aug 29-8:13 AM

To avoid negative charge build-up at the Cu cathode (due to loss of Cu^{2+}), Na^+ cations from the salt bridge migrate into the cathode solution and NO_3 anions migrate out of the cathode solution into the salt bridge. If there is no salt bridge, the cell would cease to operate due to the resistance by build-up of like charges.

http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/galvan5.swf

Aug 29-8:17 AM

Key Terms

Electrode:

Anode:

Cathode:

Wire:

Voltmeter:

Salt Bridge:

Aug 29-8:19 AM

Key Terms

Electrode: a conductor where a half-reaction occurs

Anode: The electrode which is the site of oxidation. Electrons are released here and travel up the electrode into the wire.

Cathode: The electrode which is the site of reduction. Electrons travel to the cathode where they are used in a reduction half-reaction.

Wire: Electrons travel through the wire from anode to cathode. The electricity is used to do work (power a device).

Voltmeter: used to measure the voltage of the cell

Salt Bridge: Made up of spectator ions (ex. KNO_3, Na_2SO_4). Ions from the salt bridge help to maintain neutrality in both half-cells.

Aug 29-8:19 AM

If the half-cells were not separated, what would happen?

Aug 29-8:23 AM

If the half-cells were not separated, what would happen?

The redox reaction between Zn and Cu^{2+} would occur directly in the solution (Cu^{2+} ions would be right beside Zn atoms, taking electrons) and there would be no electron flow in the wire. Thus, electrical energy from the cell could not be harnessed to do work.

Aug 29-8:23 AM

Determining which redox reaction is occurring in an electrochemical cell:

Electrochemical cells are **spontaneous** (as opposed to electrolytic cells), therefore the reduction half-reaction is _____ on the table than the oxidation half-reaction. To determine which electrode is the cathode (site of reduction), look for the substance in either of the half-cells that is best at reducing (strongest oxidizing agent).

Aug 29-8:25 AM

Determining which redox reaction is occurring in an electrochemical cell:

Electrochemical cells are **spontaneous** (as opposed to electrolytic cells), therefore the reduction half-reaction is higher on the table than the oxidation half-reaction. To determine which electrode is the cathode (site of reduction), look for the substance in either of half-cells that is best at reducing (highest on the left).

Aug 29-8:25 AM

Example: In an Mg/Pb electrochemical cell, which is the cathode? The anode? What half-reactions are occurring?

Aug 29-8:28 AM

Example: In an Mg/Pb electrochemical cell, which is the cathode? The anode? What half-reactions are occurring?

The Pb half-reaction is higher on the table than the Mg half-reaction, so Pb is the cathode and Mg is the anode.

At Cathode: $Pb^{2+} + 2e^- \longrightarrow Pb$

At Anode: $Mg \longrightarrow Mg^{2+} + 2e^-$

Aug 29-8:28 AM

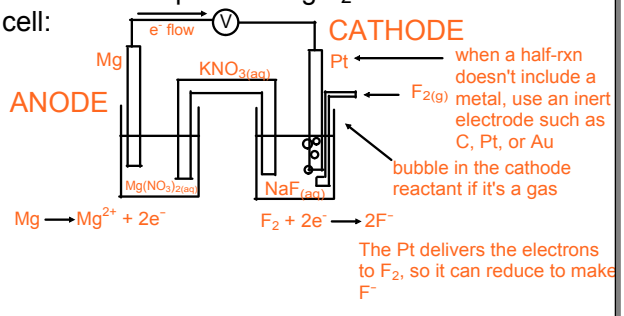
Half-Reactions

Any two half-reactions from the redox table can be put together to create an electrochemical cell. However, not all half-reactions include a metal. How can these reactions be set up in a half-cell? Sketch and explain an Mg/F_2 electrochemical cell:

Aug 29-8:29 AM

Half-Reactions

Any two half-reactions from the redox table can be put together to create an electrochemical cell. However, not all half-reactions include a metal. How can these reactions be set up in a half-cell? Sketch and explain an Mg/F_2 electrochemical cell:



Aug 29-8:29 AM

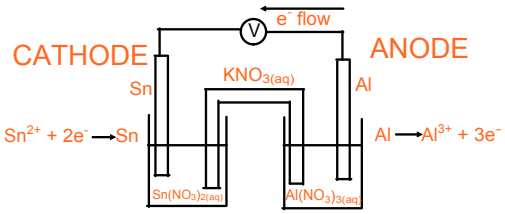
Practice Questions

1) Draw and label all parts of a Sn/Al electrochemical cell.

Aug 29-8:42 AM

Practice Questions

1) Draw and label all parts of a Sn/Al electrochemical cell.



Aug 29-8:42 AM

a) Identify a suitable electrolyte for the 'Al' half cell.
b) In which direction will electrons flow in the wire?
c) Which electrode will lose mass?
d) Toward which half-cell will the K^+ in the salt bridge migrate?
e) Write the half-reaction occurring at the anode.

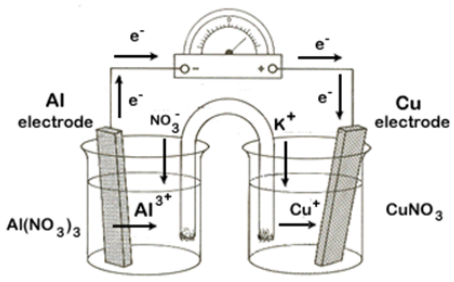
f) Identify the cathode.
g) What happens to the [Al^{3+}] in the aluminum half-cell?
h) Write the net redox equation.

Aug 29-8:45 AM

a) Identify a suitable electrolyte for the 'Al' half cell. $Al(NO_3)_{3(aq)}$
b) In which direction will electrons flow in the wire? Al to Sn (anode to cathode)
c) Which electrode will lose mass? Al (the anode)
d) Toward which half-cell will the K^+ in the salt bridge migrate? toward Sn (the cathode)
e) Write the half-reaction occurring at the anode.
$Al \longrightarrow Al^{3+} + 3e^-$
f) Identify the cathode. Sn
g) What happens to the [Al^{3+}] in the aluminum half-cell? it increases
h) Write the net redox equation.
$3Sn^{2+} + 2Al \longrightarrow 3Sn + 2Al^{3+}$

Aug 29-8:45 AM

2) Examine the following diagram of an electrochemical cell.



Aug 29-8:50 AM

a) Identify the cathode.
b) What happens to the [NO_3^-] in the Al half-cell as the cell is operating?

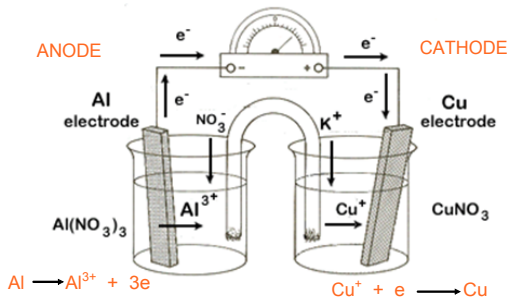
c) What happens to the [Cu^+] in the copper half-cell as the cell is operating?

d) What is the role of K^+ from the salt bridge?

e) Write the net redox reaction.

Aug 29-8:53 AM

2) Examine the following diagram of an electrochemical cell.



Aug 29-8:50 AM

a) Identify the cathode. Cu
b) What happens to the $[NO_3^-]$ in the Al half-cell as the cell is operating?
$[NO_3^-]$ is increasing to balance the inc in Al^{3+}
c) What happens to the $[Cu^+]$ in the copper half-cell as the cell is operating?
$[Cu^+]$ decreasing as it's reducing to form Cu
d) What is the role of K^+ from the salt bridge?
K^+ moves to the cathode (Cu half-cell) to replenish the reacted Cu^+ ions
e) Write the net redox reaction.
$Al + 3Cu^+ \longrightarrow Al^{3+} + 3Cu$

Aug 29-8:53 AM

Assignment 10

1) Draw an electrochemical cell with Mn and Pb electrodes and solutions of $MnCl_2$ and $Pb(NO_3)_2$ with a KNO_3 salt bridge.
a) Identify the anode.
b) Write the oxidation half-reaction.

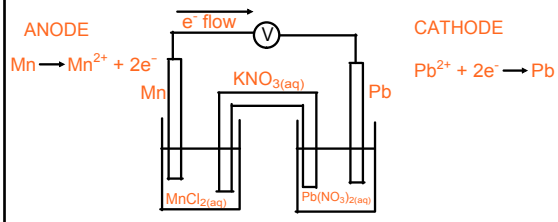
c) Write the reaction occurring at the cathode.

d) Toward which electrode do the K^+ ions migrate?
e) Toward which electrode do the electrons travel in the wire?
f) What happens to the mass of the Pb electrode?

Aug 29-8:59 AM

Assignment 10

1) Draw an electrochemical cell with Mn and Pb electrodes and solutions of $MnCl_2$ and $Pb(NO_3)_2$ with a KNO_3 salt bridge.



Aug 29-8:59 AM

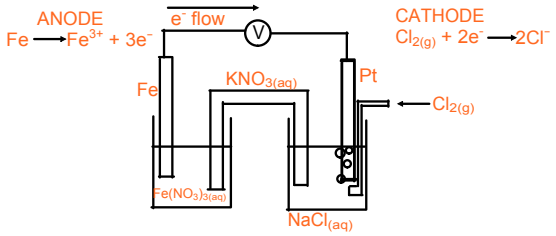
a) Identify the anode. Mn
b) Write the oxidation half-reaction.
$Mn \longrightarrow Mn^{2+} + 2e^-$
c) Write the reaction occurring at the cathode.
$Pb^{2+} + 2e^- \longrightarrow Pb$
d) Toward which electrode do the K^+ ions migrate? Pb (cathode)
e) Toward which electrode do the electrons travel in the wire? Mn to Pb (anode to cathode)
f) What happens to the mass of the Pb electrode?
increase

Aug 29-8:59 AM

2) Draw an electrochemical cell with one half-cell having an iron electrode and iron III nitrate solution and the other half-cell having an inert platinum electrode with $Cl_{2(g)}$ being bubbled in. The electrolyte in that half-cell is aqueous sodium chloride. Use KNO_3 in the salt bridge.

Aug 29-9:06 AM

2) Draw an electrochemical cell with one half-cell having an iron electrode and iron III nitrate solution and the other half-cell having an inert platinum electrode with $Cl_{2(g)}$ being bubbled in. The electrolyte in that half-cell is aqueous sodium chloride. Use KNO_3 in the salt bridge.



Aug 29-9:06 AM

a) Identify the cathode.
b) Write the half-reaction occurring in the iron half-cell.
c) What happens to the [NO_3] in the iron half-cell?
d) Write the half-reaction occurring in the Cl_2 half-cell.
e) What happens to the extra Cl^- ions being produced at the cathode?

f) What is the function of the platinum electrode at the cathode?
g) Write the net redox reaction.

Aug 29-9:13 AM

a) Identify the cathode. Pt, Au, or C
b) Write the half-reaction occurring in the iron half-cell. $Fe \longrightarrow Fe^{3+} + 3e^-$
c) What happens to the [NO_3] in the iron half-cell? increase
d) Write the half-reaction occurring in the Cl_2 half-cell. $Cl_2 + 2e^- \longrightarrow 2Cl^-$
e) What happens to the extra Cl^- ions being produced at the cathode?
they migrate into the salt bridge
f) What is the function of the platinum electrode at the cathode? transfer electrons to Cl_2
g) Write the net redox reaction.
$3Cl_2 + 2Fe \longrightarrow 6Cl^- + 2Fe^{3+}$

Aug 29-9:13 AM

3) Read Hebden p.215-217 and do #34, 35

Aug 29-9:18 AM

3) Read Hebden p.215-217 and do #34, 35
answers in the back of Hebden

Aug 29-9:18 AM

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_18&folder=zinc_copper_cell

XII) Standard Potentials

Cell potential, measured in volts (V), is the 'pull' on electrons. Potential, or voltage, acts on electrons just like 'pressure' acts on water, or like gravity acts on objects. In an electrochemical cell, the substance that wants to reduce pulls the electrons through the wire, and this movement of electrons is what is used to do work, such as power a motor. One volt is defined as 1 joule of work per coulomb of charge transferred, so voltage gives an indication of how much work the electrochemical cell can do. The higher the voltage, the stronger the battery.

Aug 29-9:19 AM

The magnitude of the voltage in an electrochemical cell depends on its half-reactions. If a substance that is very good at giving up electrons oxidizes (low on the right side of the table), and a substance that is very good at reducing gains electrons (high on the left side), the cell will have a large voltage (large gap between half-reactions on the table). The smaller the gap between the reduction and oxidation half-reactions, the smaller the voltage.

http://www.absorblearning.com/media/attachment.action?quick=uo&att=2198

Aug 29-9:26 AM

Find the cell voltage of a lithium/fluorine cell:
When calculating cell voltage, switch the sign of the oxidation voltage

Coefficients don't affect voltage as it's joule per coulomb of charge

Aug 29-9:27 AM

Find the cell voltage of a lithium/fluorine cell:
When calculating cell voltage, switch the sign of the oxidation voltage

$F_2 + 2e^- \longrightarrow 2F$ +2.87V

$Li \longrightarrow Li^+ + e$ +3.04V

E° cell = +5.91V

Coefficients don't affect voltage as it's joule per coulomb of charge

Aug 29-9:27 AM

The E° values on the table are at 'standard state', meaning 25°C, 1atm pressure for gases (room pressure), 1M concentrations for all solutions. If any of these conditions are altered, the voltage is just an E value as it's no longer at standard state.

Aug 29-9:32 AM

Notice that each half-reaction is assigned a certai voltage on the redox table. It is impossible to set up just a half-cell (without another half-cell) and get a voltage reading. You can only get a voltage from a complete cell with oxidation and reduction half-reactions occurring simultaneously. So where do these half-reaction voltages on the table come from?

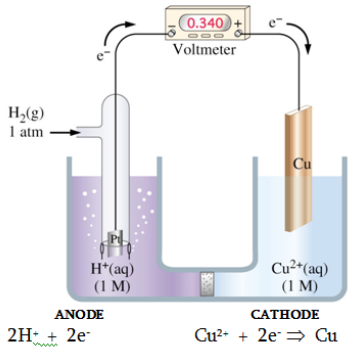
The hydrogen half-cell at standard state was used as a reference half-cell, and arbitrarily assigned a voltage of 0.00V. $2H^+ + 2e \longrightarrow H_2$

$E^\circ = 0.00V$

Every other half-cell was hooked up to the hydrogen half-cell and the voltage was recorded, and became the half-cell voltage for that respective half-reaction.

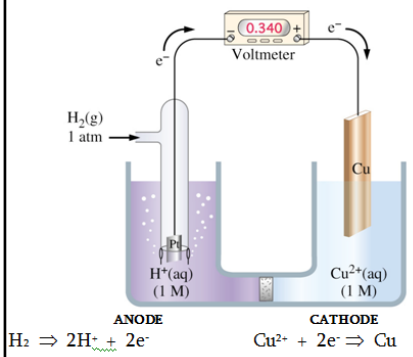
Aug 29-10:46 AM

Here is an example of the hydrogen half-cell and a copper half-cell:



Aug 29-10:51 AM

Here is an example of the hydrogen half-cell and a copper half-cell:



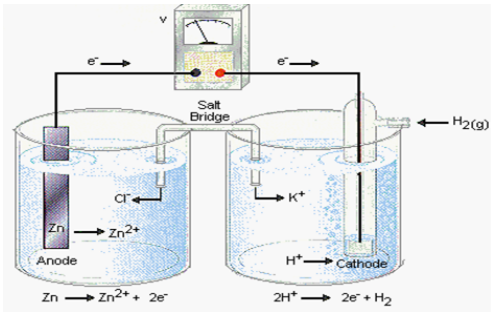
$H_2 \rightarrow 2H^+ + 2e^-$ 0.00V
$Cu^{2+} + 2e^- \rightarrow Cu$
E^o_{cell} = +0.34V

Therefore, the $Cu^{2+} + 2e^- \rightarrow Cu$ reduction is given a voltage of 0.34V (see table)

$H_2 \Rightarrow 2H^+ + 2e^-$ $\quad$ $Cu^{2+} + 2e^- \Rightarrow Cu$

Aug 29-10:51 AM

What about a hydrogen half-cell with a zinc half-cell?



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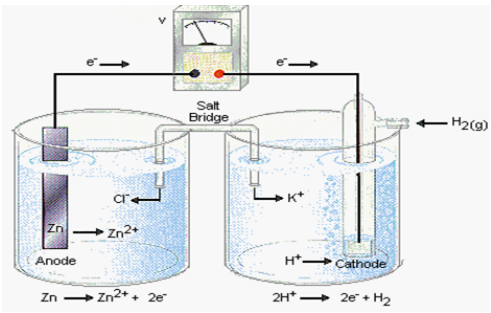
Aug 29-10:55 AM

What about a hydrogen half-cell with a zinc half-cell?



$2H^+ + 2e^- \rightarrow H_2$ 0.00V
$Zn \rightarrow Zn^{2+} + 2e^-$
E^o_{cell} = 0.76V

Therefore, the oxidation half-reaction $Zn \rightarrow Zn^{2+} + 2e^-$ is 0.76V, so the reduction potential $Zn^{2+} + 2e \rightarrow Zn$ is -0.76V (see table)

http://www.absorblearning.com/media/attachment.action?quick=uo&att=2198

Aug 29-10:55 AM

Any other half-reaction could have been chosen as the reference half-cell and been assigned 0.00V, but the hydrogen half-cell was used.

The voltages given on the table are for **reduction** half-reactions. To get **oxidation** half-reaction voltages, **reverse** the sign on the voltage.
reduction of Cu^{2+}: $Cu^{2+} + 2e^- \longrightarrow Cu$ $E^o = 0.34V$
oxidation of Cu : $Cu \longrightarrow Cu^{2+} + 2e^-$ $E^o = -0.34V$

Aug 29-10:59 AM

To predict the cell voltage of two half-cells, simply add the two half-reaction voltages together:
Example: $Br_{2(l)} + 2e \longrightarrow 2Br$
$Pb_{(s)} \longrightarrow Pb^{2+} + 2e^-$
E^o_{cell} =

A **spontaneous** reaction has a **positive** E^o_{cell} and it's an electro**chemical** cell.

A **non-spontaneous** reaction has a **negative** E^o_{cell} or a **zero** E^o_{cell} and is an electro**lytic** cell (it won't occur unless supplied with voltage).

Aug 29-11:02 AM

To predict the cell voltage of two half-cells, simply add the two half-reaction voltages together:
Example: $Br_{2(l)} + 2e \longrightarrow 2Br$ +1.09V
$Pb_{(s)} \longrightarrow Pb^{2+} + 2e^-$ +0.13V
E^o_{cell} = +1.22V

A **spontaneous** reaction has a **positive** E^o_{cell} an its an electro**chemical** cell.

A **non-spontaneous** reaction has a **negative** E^o_{cell} or a **zero** E^o_{cell} and is an electro**lytic** cell (it won't occur unless supplied with voltage).

Aug 29-11:02 AM

Practice Questions

1) Calculate the standard potential of the cell:

$Cu^{2+} + Co \longrightarrow Cu + Co^{2+}$

Aug 29-4:32 PM

Practice Questions

1) Calculate the standard potential of the cell:

$Cu^{2+} + Co \longrightarrow Cu + Co^{2+}$

$Cu^{2+} + 2e \longrightarrow Cu \quad +0.34V$

$Co \longrightarrow Co^{2+} + 2e^{-} \quad +0.28V$

$E^{o}_{cell} = +0.62V$

Aug 29-4:32 PM

2) Calculate the potential of the cell:

$3Zn^{2+} + 2Al \longrightarrow 3Zn + 2Al^{3+}$

**Remember: Coefficients don't affect voltage*

Aug 29-4:34 PM

2) Calculate the potential of the cell:

$3Zn^{2+} + 2Al \longrightarrow 3Zn + 2Al^{3+}$

**Remember: Coefficients don't affect voltage*

$Zn^{2+} + 2e \longrightarrow Zn \quad -0.76V$

$Al \longrightarrow Al^{3+} + 3e^{-} \quad +1.66V$

$E^{o}_{cell} = +0.90V$

Aug 29-4:34 PM

3) Write the complete redox equation and calculate the E^{o}_{cell} of a Pb/Cr electrochemical cell.

Aug 29-4:38 PM

3) Write the complete redox equation and calculate the E^{o}_{cell} of a Pb/Cr electrochemical cell.

spontaneous cell - Pb is higher so that is reduction

$Pb^{2+} + 2e \longrightarrow Pb \quad -0.13V$

$Cr \longrightarrow Cr^{3+} + 3e^{-} \quad +0.74V$

$E^{o}_{cell} = +0.61V$

$3Pb^{2+} + 2Cr \longrightarrow 3Pb + 2Cr^{3+}$

Aug 29-4:38 PM

Cells at Equilibrium

The E^{o}_{cell} of a cell is the voltage at standard state, so solution concentrations are 1.0M. As the cell operates, reactants are being used up, so their concentrations are continually decreasing, thus the voltage of the cell starts to decrease.
When all reactants are used up, the cell will cease to operate (the battery is dead), and the cell voltage will be zero. At this point, the cell is said to be at equilibrium.
Therefore, the voltage of a cell **at equilibrium** is 0 (a dead battery).

Aug 29-4:41 PM

Assignment 11

1) Do Hebden p.224-225 #36abcdef, 37
p.226 #46
(36f is a disproportionation reaction)

Aug 29-4:44 PM

Assignment 11

1) Do Hebden p.224-225 #36abcdef, 37
p.226 #46
(36f is a disproportionation reaction)

answers in the back of Hebden

Aug 29-4:44 PM

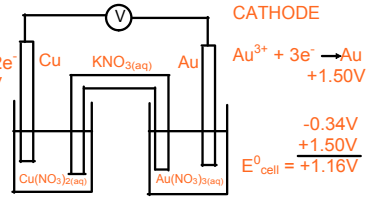
XIII) Cell Potential Practice Questions

1) Draw and label an electrochemical cell using a copper anode and having an E^{o}_{cell} greater than 1.0V (The product of the oxidation is Cu^{2+}).

Aug 29-4:56 PM

XIII) Cell Potential Practice Questions

1) Draw and label an electrochemical cell using a copper anode and having an E^{o}_{cell} greater than 1.0V (The product of the oxidation is Cu^{2+}).



$Cu \rightarrow Cu^{2+} + 2e^{-}$ $E^{o} = -0.34V$
so for $E^{o}_{cell} > 1.0V$, the reduction half-reaction must be greater than 1.34V. Thus, reduction from Cl_2 up on the table will be sufficient

-0.34V
+1.50V
E^{0}_{cell} = +1.16V

Aug 29-4:56 PM

2) If the E^{o} fro $Ni^{2+} + 2e^{-} \longrightarrow Ni$ were set at zero volts, what would be the E^{o} for $Cu^{2+} + 2e^{-} \rightarrow Cu$?

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Aug 29-5:04 PM

2) If the E^o fro $Ni^{2+} + 2e^- \longrightarrow Ni$ were set at zero volts, what would be the E^o for $Cu^{2+} + 2e^- \longrightarrow Cu$?

http://www.absorblearning.com/media/attachment.action?quick=uo&att=2198

$Ni^{2+} + 2e^- \longrightarrow Ni$ is -0.26V, so if it was set at 0.00V, that is 0.26V higher

Therefore, everything would be 0.26V higher. Currently, $Cu^{2+} + 2e \longrightarrow Cu$ is currently 0.34V, so it would be 0.34V + 0.26V = 0.60V

Aug 29-5:04 PM

3) What will happen to an aluminum spoon if it is used to stir a solution of $Fe(NO_3)_2$? Use E^o values to support your answer.

Aug 29-5:08 PM

3) What will happen to an aluminum spoon if it is used to stir a solution of $Fe(NO_3)_2$? Use E^o values to support your answer.

This is a possible single replacement reaction. But is it spontaneous?

$2Al + 3Fe(NO_3)_2 \longrightarrow 2Al(NO_3)_3 + 3Fe$

$Al \longrightarrow Al^{3+} + 3e^-$ +1.66V

$Fe^{2+} + 2e^- \longrightarrow Fe$ -0.45V

E^o_{cell} = 1.21V spontaneous

The aluminum spoon will react and 'disappear'.

Aug 29-5:08 PM

4) Four metals were used to set up the following electrochemical cells with their ions J^{2+}, L^+, K^{2+}, and M^{3+}.

ANODE	CATHODE	CELL VOLTAGE
J	L	+0.30V
L	K	+1.60V
L	M	+1.30V

a) Identify the strongest reducing agent.

b) Identify the strongest oxidizing agent.

c) Calculate the voltage of a J/M cell.

Aug 29-5:15 PM

4) Four metals were used to set up the following electrochemical cells with their ions J^{2+}, L^+, K^{2+}, and M^{3+}.

ANODE	CATHODE	CELL VOLTAGE
J	L	+0.30V
L	K	+1.60V
L	M	+1.30V

a) Identify the strongest reducing agent. J

b) Identify the strongest oxidizing agent. K^+

c) Calculate the voltage of a J/M cell.

1.60V

Aug 29-5:15 PM

5) Four half-cells are constructed by placing strips of the four metals W, X, Y, and Z in 1.00M solutions of their ions W^+, X^{2+}, Y^{3+}, and Z^{2+} respectively. Various combinations of these half-cells are connected, giving the following data:

ANODE	CATHODE	E^o_{cell}
W	X	+0.20V
W	Y	+0.36V
Z	W	+0.14V

a) Which metal is the strongest reducing agent? Z

b) Which metal ion is the strongest oxidizing agent? Y^{3+}

c) Write the balanced equation for the cell reaction that would occur when half-cells X and Y are connected. $2Y^{3+} + 3X \longrightarrow 2Y + 3X^{2+}$

Aug 29-5:39 PM

d) Calculate the voltage produced when half-cells X and Y are connected to produce a spontaneous reaction.

Aug 29-5:43 PM

d) Calculate the voltage produced when half-cells X and Y are connected to produce a spontaneous reaction.

Anode	Cathode	Voltage
X	~~W~~	-0.20V
~~W~~	Y	+0.36V
X	Y	+0.16V

Aug 29-5:43 PM

6) Consider the following redox data:
$3V^{2+} + 2Ga \longrightarrow 3V + 2Ga^{3+}$ $E^o = -0.64V$
$3V^{2+} + 2Al \longrightarrow 3V + 2Al^{3+}$ $E^o = +0.46V$
Based on these observations, a student concludes that Ga^{3+} and Al will react spontaneously. Evaluate this conclusion, and support your answer with calculations.

Aug 29-5:45 PM

6) Consider the following redox data:
$3V^{2+} + 2Ga \longrightarrow 3V + 2Ga^{3+}$ $E^o = -0.64V$
$3V^{2+} + 2Al \longrightarrow 3V + 2Al^{3+}$ $E^o = +0.46V$
Based on these observations, a student concludes that Ga^{3+} and Al will react spontaneously. Evaluate this conclusion, and support your answer with calculations.

$3V + 2Ga^{3+} \longrightarrow 3V^{2+} + 2Ga$ $E^o = +0.64V$
$3V^{2+} + 2Al \longrightarrow 3V + 2Al^{3+}$ $E^o = +0.46V$
$Ga^{3+} + Al \longrightarrow Ga + Al^{3+}$ $E^o = 1.10V$
spontaneous

Aug 29-5:45 PM

Assignment 12

1) An electrochemical cell has electrodes of Pb in $Pb(NO_3)_2$ and Cr in $Cr(NO_3)_3$. Calculate E^o_{cell}.

Aug 30-8:13 AM

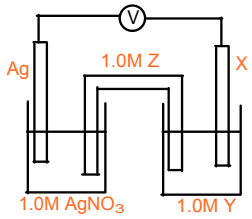
Assignment 12

1) An electrochemical cell has electrodes of Pb in $Pb(NO_3)_2$ and Cr in $Cr(NO_3)_3$. Calculate E^o_{cell}.

$Cr \longrightarrow Cr^{3+} + 3e$ 0.74V
$Pb^{2+} + 2e^- \longrightarrow Pb$ -0.13V
$E^o_{cell} = +0.61V$

Aug 30-8:13 AM

2) Use the following diagram of a cell at 25°C to answer the questions:



a) Identify the metal 'X'.

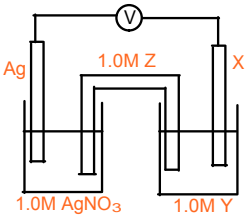
b) Identify a suitable electrolyte 'Y'.

c) Identify a suitable electrolyte 'Z'.

d) Indicate on the diagram the direction of electron flow.

Aug 30-8:16 AM

2) Use the following diagram of a cell at 25°C to answer the questions:



a) Identify the metal 'X'. Pb

b) Identify a suitable electrolyte 'Y'. $Pb(NO_3)_{2(aq)}$

c) Identify a suitable electrolyte 'Z'. $KNO_{3(aq}$

d) Indicate on the diagram the direction of electron flow. Pb to Ag (anode to cathode)

Aug 30-8:16 AM

3) Can $Fe_2(SO_4)_3$ be stored in a container made of nickel? Support with E^o calculations.

Aug 30-8:17 AM

3) Can $Fe_2(SO_4)_3$ be stored in a container made of nickel? Support with E^o calculations.

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$ 0.77V

$Ni \longrightarrow Ni^{2+} + 2e^-$ 0.26V

$E^o_{cell} = +1.03V$

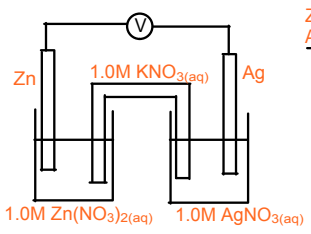
No, as a spontaneous reaction will ensue and the container will break down.

Aug 30-8:17 AM

4) Draw and label the parts of an operating electrochemical cell using a zinc anode that will produce an electric current having a voltage of 1.56V under standard conditions.

Aug 30-8:26 AM

4) Draw and label the parts of an operating electrochemical cell using a zinc anode that will produce an electric current having a voltage of 1.56V under standard conditions.



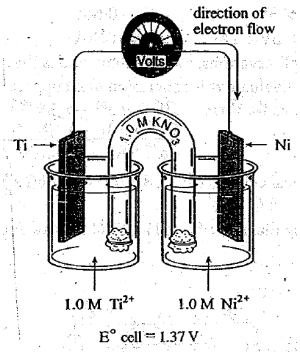
$Zn \longrightarrow Zn^{2+} + 2e^-$ 0.76V

$Ag^+ + e^- \longrightarrow Ag$ 0.80V

$E^o_{cell} = 1.56V$

Aug 30-8:26 AM

5) Consider the following electrochemical cell:

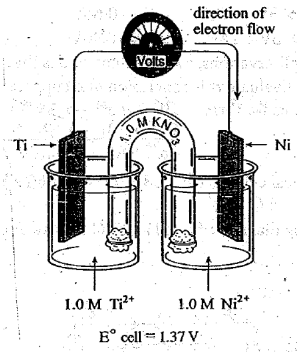


a) Write the equation for half-reaction that occurs at the anode.

b) Calculate the reduction potential of Ti^{2+}.

Aug 30-8:30 AM

5) Consider the following electrochemical cell:



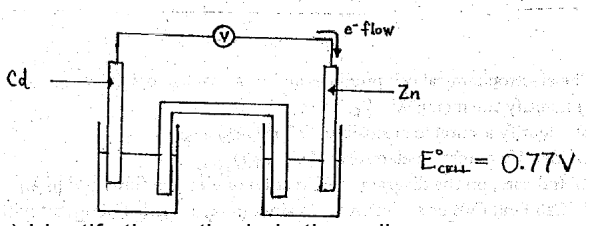
a) Write the equation for half-reaction that occurs at the anode.

$Ti \longrightarrow Ti^{2+} + 2e$

b) Calculate the **reduction** potential of Ti^{2+}. -1.63V

Aug 30-8:30 AM

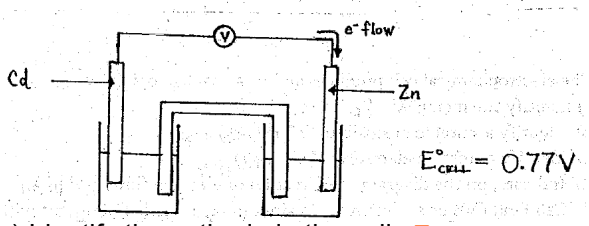
6) Given the following diagram of an electrochemical cell, which is constructed in order to determine the reduction potential of a Cadmium half-cell, answer the questions below.



a) Identify the cathode in the cell.
b) Write the half-reaction occuring at the Cd electrode.
c) Determine the E^o for the half-reaction:
$Cd^{2+} + 2e \longrightarrow Cd$

Aug 30-9:34 AM

6) Given the following diagram of an electrochemical cell, which is constructed in order to determine the reduction potential of a Cadmium half-cell, answer the questions below.



a) Identify the cathode in the cell. Zn
b) Write the half-reaction occuring at the Cd electrode. $Cd \longrightarrow Cd^{2+} + 2e$
c) Determine the E^o for the half-reaction:
$Cd^{2+} + 2e \longrightarrow Cd$ -1.53V

Aug 30-9:34 AM

7) Four metals were used to set up the following electrochemical cells.

ANODE	CATHODE	CELL VOLTAGE
A	B	+1.402V
A	C	+1.230V
Z	A	+0.080V

a) Rank the metals from strongest reducing agent to weakest.
b) Predict the E^O of a Z/B cell.

Aug 30-9:38 AM

7) Four metals were used to set up the following electrochemical cells.

ANODE	CATHODE	CELL VOLTAGE
A	B	+1.402V
A	C	+1.230V
Z	A	+0.080V

a) Rank the metals from strongest reducing agent to weakest. Z, A, C, B
b) Predict the E^O of a Z/B cell. +1.482V

Aug 30-9:38 AM

8) The following reactions occur at 25°C with all substances present in 1.0M concentrations.

$Zn + Pb^{2+} \Rightarrow Zn^{2+} + Pb$

$Ti + Zn^{2+} \Rightarrow Ti^{2+} + Zn$

$2Lu + 3Ti^{2+} \Rightarrow 2Lu^{3+} + 3Ti$

Predict whether each of the following reactions will occur:

a) $Pb + Ti^{2+} \Rightarrow Pb^{2+} + Ti$

b) $2Lu + 3Pb^{2+} \Rightarrow 2Lu^{3+} + 3Pb$

c) $2Lu^{3+} + 3Zn \Rightarrow 3Zn^{2+} + 2Lu$

Aug 30-9:40 AM

8) The following reactions occur at 25°C with all substances present in 1.0M concentrations.

$Zn + Pb^{2+} \Rightarrow Zn^{2+} + Pb$

$Ti + Zn^{2+} \Rightarrow Ti^{2+} + Zn$

$2Lu + 3Ti^{2+} \Rightarrow 2Lu^{3+} + 3Ti$

Predict whether each of the following reactions will occur:

a) $Pb + Ti^{2+} \Rightarrow Pb^{2+} + Ti$ WILL NOT OCCUR

b) $2Lu + 3Pb^{2+} \Rightarrow 2Lu^{3+} + 3Pb$ WILL OCCUR

c) $2Lu^{3+} + 3Zn \Rightarrow 3Zn^{2+} + 2Lu$ WILL NOT OCCUR

Aug 30-9:40 AM

9) The metals Rh, Ti, Cr, and Pd are individually placed in 1.0M solutions of Rh^{2+}, Ti^{2+}, Cr^{2+}, and Pd^{2+}, and the cell voltages of the spontaneous reactions are determined.

ION / METAL	Rh^{2+}	Ti^{2+}	Pd^{2+}	Cr^{2+}
Rh		no reaction	0.35 V	no reaction
Ti	2.23 V		2.58 V	?
Pd	no reaction	no reaction		no reaction
Cr	1.51 V	no reaction	1.86 V	

a) Arrange the metals in order of **increasing** strength as reducing agents.

b) Determine the cell voltage for Ti in a 1.0M solution of Cr^{2+}.

Aug 30-9:43 AM

9) The metals Rh, Ti, Cr, and Pd are individually placed in 1.0M solutions of Rh^{2+}, Ti^{2+}, Cr^{2+}, and Pd^{2+}, and the cell voltages of the spontaneous reactions are determined.

ION / METAL	Rh^{2+}	Ti^{2+}	Pd^{2+}	Cr^{2+}
Rh		no reaction	0.35 V	no reaction
Ti	2.23 V		2.58 V	?
Pd	no reaction	no reaction		no reaction
Cr	1.51 V	no reaction	1.86 V	

a) Arrange the metals in order of **increasing** strength as reducing agents. Pd, Rh, Cr, Ti

b) Determine the cell voltage for Ti in a 1.0M solution of Cr^{2+}. +0.72V

Aug 30-9:43 AM

10) Do Hebden p.225 #38, 40, 41

Aug 30-9:47 AM

10) Do Hebden p.225 #38, 40, 41

answers in the back of Hebden

Aug 30-9:47 AM

XIV) Selecting Preferred Reactions

In many reaction environments, there's more than one substance that can reduce, and/or more than one substance that can oxidize. How can you determine the reaction that actually occurs?

1) When several different reductions can occur,...

2) When several different oxidations can occur, ...

Aug 30-9:48 AM

XIV) Selecting Preferred Reactions

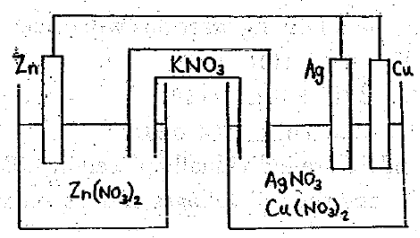
In many reaction environments, there's more than one substance that can reduce, and/or more than one substance that can oxidize. How can you determine the reaction that actually occurs?

1) When several different reductions can occur,...
the one with the highest E^o will occur preferentially (highest on the left of the table).

2) When several different oxidations can occur, ...
the one with the highest E^o will occur preferentially (lowest on the right of the table).

Aug 30-9:48 AM

Write the redox equation that will occur for the following cell:

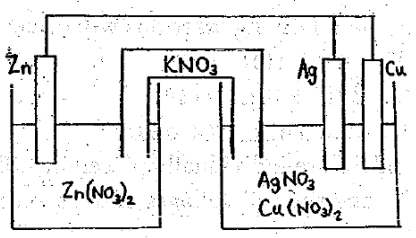


A method to deduce the actual reaction occurring:
i)List all possible <u>oxidizing agents</u> <u>reducing agents</u>

ii) Circle each agent that is the most preferred.

Aug 30-9:51 AM

Write the redox equation that will occur for the following cell:



A method to deduce the actual reaction occurring:
i)List all possible <u>oxidizing agents</u> <u>reducing agents</u>

oxidizing agents		reducing agents	
Zn^{2+}	Cu^{2+}	Zn	Cu
Ag^+	H_2O	Ag	H_2O

ii) Circle each agent that is the most preferred.

Aug 30-9:51 AM

iii) Write the overall reaction.

Aug 30-9:53 AM

iii) Write the overall reaction.

$2Ag^+ + Zn \longrightarrow 2Ag + Zn^{2+}$

Aug 30-9:53 AM

2) A beaker contains an iron nail wrapped with both a piece of copper wire and a piece of magnesium ribbon, immersed in a solution containing $CuSO_{4(aq)}$ and some dissolved $Cl_{2(g)}$. What is the preferred redox reaction?

Aug 30-9:55 AM

2) A beaker contains an iron nail wrapped with both a piece of copper wire and a piece of magnesium ribbon, immersed in a solution containing $CuSO_{4(aq)}$ and some dissolved $Cl_{2(g)}$. What is the preferred redox reaction?

$Cl_2 + Mg \longrightarrow Mg^{2+} + 2Cl^-$

Aug 30-9:55 AM

Assignment 13

Do Hebden p.228 #47

Aug 30-9:57 AM

Assignment 13

Do Hebden p.228 #47

answers in the back of Hebden

Aug 30-9:57 AM

XV) Applied Electrochemistry

Breathalyzer

Police use a breathalyzer to test for alcohol consumption. The breathalyzer consists of a redox reaction that includes a colour change. The intensity of green colour produced is measured to find the amount of alcohol in the blood. The colour change is due to the orange dichromate ion ($Cr_2O_7^{2-}$) reacting to produce the green chromium ion (Cr^{3+}) due to a reaction with ethanol, C_2H_5OH, the alcohol used in beverages.

Aug 30-9:57 AM

$C_2H_5OH + K_2Cr_2O_7 + H_2SO_4 \longrightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$

ethanol orange green

All reactants are in the breathalyzer in excess, except the ethanol, which comes from one's breath. The amount of Cr^{3+} produced is dependent on the amount of ethanol, therefore the colour change and intensity of green is dependent on the amount of ethanol provided.

How do you know the above equation is a redox equation?

Aug 30-10:00 AM

$C_2H_5OH + K_2Cr_2O_7 + H_2SO_4 \longrightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$
ethanol orange green

All reactants are in the breathalyzer in excess, except the ethanol, which comes from one's breath. The amount of Cr^{3+} produced is dependent on the amount of ethanol, therefore the colour change and intensity of green is dependent on the amount of ethanol provided.

How do you know the above equation is a redox equation?

the oxidation numbers for C and Cr change

Aug 30-10:00 AM

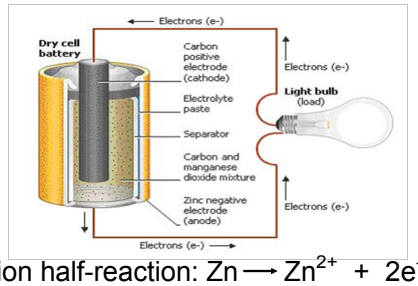
Batteries

For any battery, electrons produced at the anode leave the battery to power a device and return via the electrical circuit to the cathode for reduction.

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/flashlight.html

Aug 30-10:04 AM

Zinc-Carbon Battery (aka Dry Cell)
'Dry' because the ions are in a paste



oxidation half-reaction: $Zn \longrightarrow Zn^{2+} + 2e^-$
reduction half-reaction:
$2MnO_2 + 2NH_4^+ + 2e^- \longrightarrow Mn_2O_3 + 2NH_3 + H_2O$

Aug 30-10:05 AM

The Zn casing around the battery is the anode. The carbon rod down the centre is the inert cathode.

The battery dies when all of the Zn is consumed.

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/ZnCbatteryV8web.html

Advantages: cheap materials
Disadvantages: not rechargeable, short shelf life, voltage inconsistent

Aug 30-10:08 AM

Alkaline Dry Cell (Modified Version of the Zinc/Carbon Dry Cell)

Same as the zinc-carbon battery except the paste is manganese dioxide and potassium hydroxide. The KOH electrolyte gives this battery its name (alkaline = basic). The half-reactions are the same as the zinc-carbon battery except they are under basic conditions. Alkaline batteries are the most common battery today.

Anode: $Zn + 2OH \longrightarrow ZnO + H_2O + 2e$
Cathode: $2MnO_2 + H_2O + 2e \longrightarrow Mn_2O_3 + 2OH^-$

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_18&folder=alkaline_battery

Aug 30-10:10 AM

Advantages:
-more efficient ion transport in basic (alkaline) electrolyte
-more constant voltage than zinc-carbon battery

Disadvantages:
-materials more expensive

Aug 30-10:13 AM

Lead-Acid Storage Battery (Car Battery)

This is the type of battery found in automobiles. It is made up of 6 individual cells connected in series, with each cell producing 2 volts (making a 12 volt battery). Each cell consists of one Pb plate and one PbO_2 plate immersed in Sulphuric Acid. The cathode and anode do not need separate compartments since both Pb and PbO_2 are solids, so they cannot come in direct contact. This battery primarily serves to start the car.

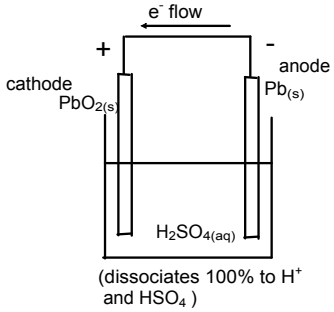
http://www.wainet.ne.jp/~yuasa/flash/EngLead_Storage_Battery.swf

Aug 30-9:51 AM

Discharging half-reactions:

Anode: $Pb_{(s)} + HSO_4^- \longrightarrow PbSO_{4(s)} + H^+ + 2e^-$

Cathode: $PbO_{2(s)} + 3H^+ + HSO_4^- + 2e^- \longrightarrow PbSO_{4(s)} + 2H_2O$



(dissociates 100% to H^+ and HSO_4^-)

Recharging half-reactions would be the reverse of the discharging half-reactions.

Aug 30-11:22 AM

Notice that solid $PbSO_4$ is a product of each half-reaction. It sticks to both electrodes and serves as the reactant when the battery is recharged, which is accomplished by applying a voltage to reverse the half-reactions. The *alternator* does this; it is powered by the motor to convert mechanical energy to chemical energy, which then provides the necessary voltage to the non-spontaneous recharging reaction. Eventually, $PbSO_4$ can 'flake off' of the electrodes, due to bumps, erratic driving, etc. and fall to the bottom of the H_2SO_4 bath. This loss of reactants disallows the battery from fully recharging. Eventually, a new battery is required.

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/PbbatteryV9web.html

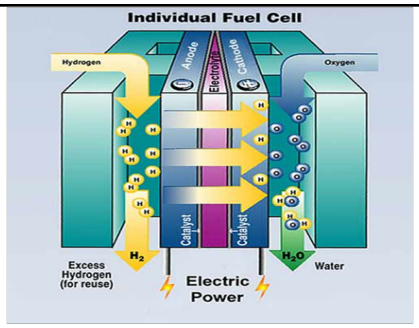
Aug 30-9:45 AM

Fuel Cell

Fuel cells are different from the batteries described above because fuel cells have to be continually supplied with hydrogen and oxygen gas, whereas the other batteries are self contained units.

http://www.wainet.ne.jp/~yuasa/flash/EngFuel_Cell.swf

Aug 30-1:12 PM



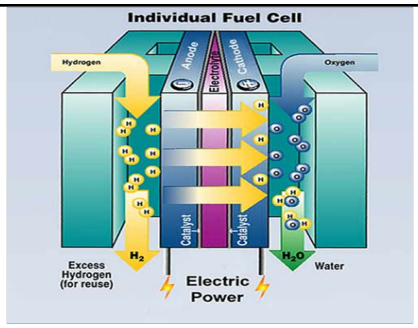
Anode Half-Reaction: $2H_2 \longrightarrow 4H^+ + 4e^-$

Cathode Half-Reaction: $O_2 + 4H^+ + 4e \longrightarrow 2H_2O$

Overall Reaction:

http://www.odec.ca/projects/2006/wong6j2/hydrogenAnimation.html

Aug 30-1:15 PM



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Overall Reaction: $2H_2 + O_2 \longrightarrow 2H_2O$

http://www.odec.ca/projects/2006/wong6j2/hydrogenAnimation.html

Aug 30-1:15 PM

Advantages:
non-polluting, produces drinkable water, efficient

Disadvantages:
expensive, requires a lot of maintenance, dangerous due to compressed gas

http://www.fueleconomy.gov/feg/fuelcell8.swf

BALLARD: http://www.youtube.com/watch?v=Sncuy89bjCc

HONDA: http://automobiles.honda.com/fcx-clarity/

VANCOUVER BUSES: http://www.youtube.com/watch?v=HvP1nSEB-po

Aug 30-1:20 PM

Assignment 14

1) Read Hebden p.228-233 and do #52, 53, 55a, 56

Aug 30-1:21 PM

Assignment 14

1) Read Hebden p.228-233 and do #52, 53, 55a, 56

answers in the back of Hebden

Aug 30-1:21 PM

XVI) Corrosion of Iron

What must be present in order for iron to rust?

Iron corroding is actually a redox reaction. What do you think happens to the Fe metal?

Where do the resulting electrons go?

So, what serves as the anode? The cathode?

Aug 30-1:22 PM

XVI) Corrosion of Iron

What must be present in order for iron to rust?

H_2O and O_2

Iron corroding is actually a redox reaction. What do you think happens to the Fe metal?

it oxidizes to form the ion Fe^{2+}

$Fe \longrightarrow Fe^{2+} + 2e$

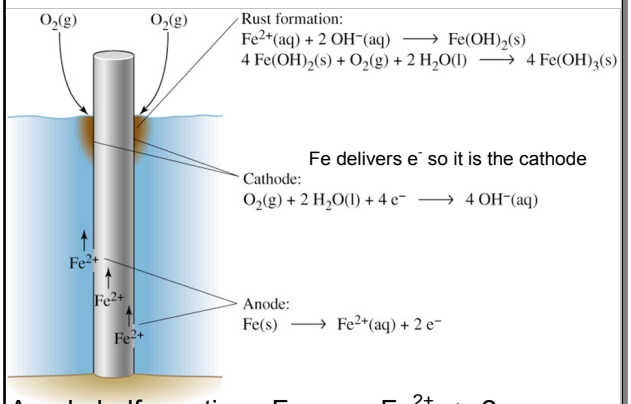
Where do the resulting electrons go?

they are used to reduce O_2 in the presence of H_2O $\quad O_2 + 2H_2O + 4e \longrightarrow 4OH^-$

So, what serves as the anode? The cathode?

Fe serves as the anode and the cathode

Aug 30-1:22 PM



Anode half-reaction: $Fe \longrightarrow Fe^{2+} + 2e$

Cathode half-reaction: $O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$

Aug 30-1:27 PM

The Fe^{2+} created at the anode follows the electrons to the cathode and dissolve in the water present. This helps maintain neutrality as the negatively charged OH is produced at the cathode. In fact, Fe^{2+} combines with the hydroxide ion, producing solid iron (II) hydroxide.

$$Fe^{2+} + 2OH \longrightarrow Fe(OH)_2$$

Aug 30-1:31 PM

The iron (II) hydroxide then reacts with oxygen and water to make $Fe(OH)_3$. Then $Fe(OH)_3$ decomposes to $FeO(OH) + H_2O$, and then two FeO(OH) molecules collide to make $Fe_2O_3 + H_2O$. Rust is actually Fe_2O_3 as well as its hydrated form $Fe_2O_3 \bullet xH_2O$. This accounts for the different colours in rust.
Corrosion cannot occur in dry air or in oxygen depleted water (deep water). Why?

Aug 30-1:33 PM

The iron (II) hydroxide then reacts with oxygen and water to make $Fe(OH)_3$. then $Fe(OH)_3$ decomposes to $FeO(OH) + H_2O$, and then two FeO(OH) molecules collide to make $Fe_2O_3 + H_2O$. Rust is actually Fe_2O_3 as well as its hydrated form $Fe_2O_3 \bullet xH_2O$. This accounts for the different colours in rust.
Corrosion cannot occur in dry air or in oxygen depleted water (deep water). Why?

- dry air has no H_2O vapour
- oxygen depleted water has no O_2

Aug 30-1:33 PM

Protection from Corrosion

There are two protection types: Physical and Electrochemical

Physical Protection

What does physical protection from corrosion mean?

What are some different types of physical protection?

Aug 30-1:37 PM

Protection from Corrosion

There are two protection types: Physical and Electrochemical

Physical Protection

What does physical protection from corrosion mean? there is a physical layer on top of the iron, keeping it from being exposed to O_2 and H_2O

What are some different types of physical protection?

paint
chrome plating
plastic
grease
tin or zinc plating

Aug 30-1:37 PM

Electrochemical Protection

What does electrochemical protection from corrosion mean?

There are two types of electrochemical protection:

1) **Cathodic Protection**

Why do people put zinc strips on the iron hull of boats?

Zinc oxidizes more readily than iron (check your table). Therefore, if zinc is available, it will oxidize to Zn^{2+} before any Fe will oxidize.

Aug 31-8:04 AM

Electrochemical Protection

What does electrochemical protection from corrosion mean? Electrons are provided for O_2 to reduce in the presence of H_2O from a source other than Fe.

There are two types of electrochemical protection:

1) **Cathodic Protection**

Why do people put zinc strips on the iron hull of boats?

Zinc oxidizes more readily than iron (check your table). Therefore, if zinc is available, it will oxidize to Zn^{2+} before any Fe will oxidize.

Aug 31-8:04 AM

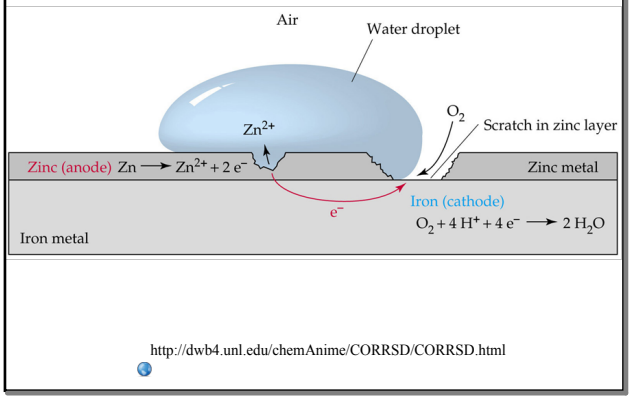
Attaching zinc strips to iron will prevent iron from rusting as the zinc will oxidize first, and the electrons from the zinc will conduct through the iron to the site of reduction (where the water and oxygen is). The zinc is called a **sacrificial anode** (it sacrifices itself for the iron). Zinc provides Fe with electrons and Fe acts as the cathode to conduct the electrons to O_2 and H_2O. The zinc does not have to completely cover the Fe.

Anode Half-Reaction: $Zn \longrightarrow Zn^{2+} + 2e$

Cathode Half-Reaction: $O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$

Aug 31-8:07 AM

The oxidation of Zn will decrease its mass. Regularly replacing the sacrificial anode (Zn strips) will ensure that the Fe will not corrode.



http://dwb4.unl.edu/chemAnime/CORRSD/CORRSD.html

Aug 31-8:10 AM

2) **Supplying an electrical current to the iron**

Some ships will supply a low voltage electrical current (a stream of electrons) to the iron hull. This prevents the iron hull from having to oxidize to supply electrons to oxygen and water, thereby preventing corrosion.

Aug 31-8:11 AM

Assignment 15

1) Could Zn be used as a sacrificial anode for Al? Explain.

2) Would it be smart to use a tin based paint on the bottom of an aluminum boat? Why or why not?

Aug 31-8:13 AM

Assignment 15

1) Could Zn be used as a sacrificial anode for Al? Explain.

No, as Al oxidizes more readily than Zn.

2) Would it be smart to use a tin based paint on the bottom of an aluminum boat? Why or why not?

No, as Al oxidizes more readily than Sn, so the tin would play no active role.

Aug 31-8:13 AM

3) The Statue of Liberty consists of an iron frame covered with copper. Discuss the reason for the rapid corrosion within this structure.

4) Why hasn't the Titanic corroded?

Aug 31-8:15 AM

3) The Statue of Liberty consists of an iron frame covered with copper. Discuss the reason for the rapid corrosion within this structure.

Any scratches in the copper exposes Fe, which then will corrode due to contact with O_2 and H_2O.

4) Why hasn't the Titanic corroded?

There is no dissolved O_2 present at the bottom of the ocean.

Aug 31-8:15 AM

5) To prevent an environmental disaster, how could you stop an underground iron septic tank from rusting through without actually having to dig it up?

6) Why does most car rust start in or around the wheel wells?

Aug 31-8:20 AM

5) To prevent an environmental disaster, how could you stop an underground iron septic tank from rusting through without actually having to dig it up?

Push a metallic wire that's attached to a chunk of zinc through the ground so it's in contact with the tank. Now, the zinc will corrode instead of the Fe tank. Replace the chunk of Zn when necessary.

6) Why does most car rust start in or around the wheel wells?

it is the wettest area of the vehicle

Aug 31-8:20 AM

XVII) Electrolysis

Recall that a spontaneous cell (a _____) is called an _____ cell.

What is an **electrolytic** cell?

How does an electrolytic cell function?

What is **electrolysis**?

Aug 31-8:23 AM

XVII) Electrolysis

Recall that a spontaneous cell (a battery) is called an electrochemical cell.

What is an **electrolytic** cell?

a cell in which a non-spontaneous redox reaction occurs by hooking up a battery to the circuit, which provides electrical energy

How does an electrolytic cell function?

the battery drives electrons from anode to cathode, forcing the non-spontaneous redox reaction to occur

What is **electrolysis**?

A non-spontaneous redox reaction that occurs by supplying a power source such as a battery.

Aug 31-8:23 AM

Why don't electrolytic cell need separate half-cells?

Aug 31-8:27 AM

Why don't electrolytic cell need separate half-cells?

The reactants can be in contact in one cell because the reaction is non-spontaneous. Therefore, it won't occur directly. It will only occur through the external wire and battery.

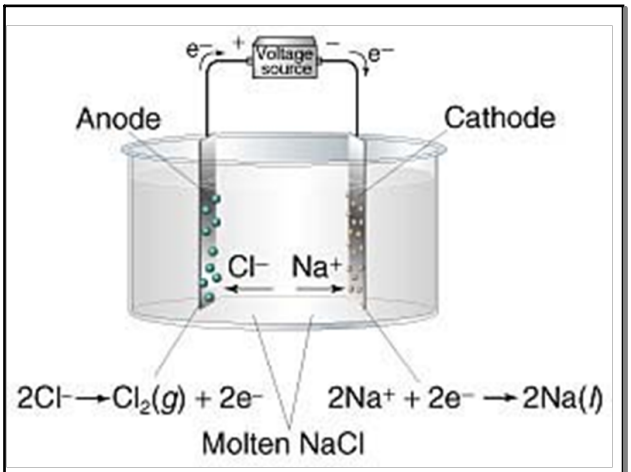
There are two types of electrolytic cells.

Aug 31-8:27 AM

1) Molten Electrolytic Cell - Electrolysis of a Molten (Liquid) Salt

A molten salt is a salt that has been melted into liquid form, so no water is present in the cell. This takes very high temperatures and is expensive.

Aug 31-8:30 AM



Aug 31-8:31 AM

The only substances in the cell available to react are Na^+ and Cl^-. The electrodes are **inert**, meaning they are unreactive.

The inert cathode electrode becomes negatively charged as the battery pumps electrons to it. The negative charge draws Na^+ ions to the cathode to be reduced.

The inert anode electrode becomes positively charged because the battery pulls electrons away from it. The positive charge draws Cl^- ions to the anode to be oxidized.

Aug 31-8:32 AM

Cathode Half-Reaction:

$Na^+ + e^- \longrightarrow Na$ $\quad E^o_{reduction}$ =

Anode Half-Reaction:

$2Cl^- \longrightarrow Cl_2 + 2e^-$ $\quad E^o_{oxidation}$ =

Overall Reaction:

E^o_{cell} =

Slightly more than _____V would have to be supplied by the battery in order for this cell to function.

http://www.absorblearning.com/media/attachment.action?quick=ua&att=2170

Aug 31-8:28 AM

Cathode Half-Reaction:
$Na^+ + e^- \longrightarrow Na$ $E^o_{reduction} = -2.71V$
Anode Half-Reaction:
$2Cl^- \longrightarrow Cl_2 + 2e^-$ $E^o_{oxidation} = -1.36V$

Overall Reaction:
$2Na^+ + 2Cl \longrightarrow 2Na + Cl_2$ $E^o_{cell} = -4.07V$

Slightly more than 4.07 V would have to be supplied by the battery in order for this cell to function.

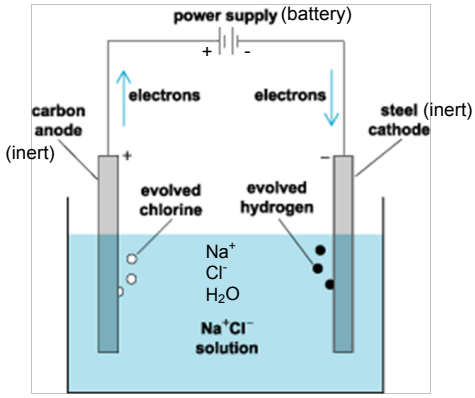
http://www.absorblearning.com/media/attachment.action?quick=ua&att=2170

Aug 31-8:28 AM

Notice on the molten cell that the positive side of the battery (the cathode) is always connected to the anode of the electrolytic cell.
The negative side of the battery (the anode) is always connected to the cathode of the electrolytic cell.
Sometimes, when solving a problem, this is the only hint you're given in order to determine the anode and cathode of the electrolytic cell.

Aug 31-8:39 AM

2) Aqueous Electrolytic Cell - Electrolysis of an Aqueous Salt (a salt solution)



Aug 31-8:41 AM

For an aqueous electrolytic cell, water is present with the salt ions, so it could oxidize or reduce in place of the ions (depends on the half-reactions on the table).
Use the table to determine the half-reactions.
Potential O.A. Potential R.A.

Aug 31-8:43 AM

For an aqueous electrolytic cell, water is present with the salt ions, so it could oxidize or reduce in place of the ions (depends on table placement).
Use the table to determine the half-reactions.
Potential O.A. Potential R.A.

Na^+ H_2O Cl^- H_2O

overpotential effect

Aug 31-8:43 AM

Cathode Half-Reaction:
$E^o =$

Anode Half-Reaction:
$E^o =$

Overall Reaction:
$E^o_{cell} =$

The battery must provide a voltage of slightly more than _____V. What colour would the solution be if phenolphthalein was added? Why?

http://www.absorblearning.com/media/attachment.action?quick=ui&att=2186

Aug 31-8:47 AM

Cathode Half-Reaction:

$2H_2O + 2e \longrightarrow H_2 + 2OH^-$ $E^o = -0.41V$

Anode Half-Reaction:

$2Cl^- \longrightarrow Cl_2 + 2e$ $E^o = -1.36V$

Overall Reaction:

$2H_2O + 2Cl^- \longrightarrow H_2 + 2OH^- + Cl_2$ $E^o_{cell} = -1.77V$

The battery must provide a voltage of slightly more than 1.77 V. What colour would the solution be if phenolphthalein was added? pink

Why? OH^- ions produced, therefore a basic environment

http://www.absorblearning.com/media/attachment.action?quick=ui&att=2186

Aug 31-8:47 AM

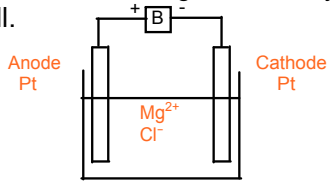
Practice Questions

1) Draw a diagram for, find the half-reactions, overall reaction, and voltage necessary to run a $MgCl_{2(l)}$ cell.

Aug 31-8:53 AM

Practice Questions

1) Draw a diagram for, find the half-reactions, overall reaction, and voltage necessary to run a $MgCl_{2(l)}$ cell.



Cathode: $Mg^{2+} + 2e^- \longrightarrow Mg$ $E^o = -2.37V$

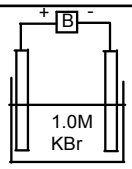
Anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$ $E^o = -1.36V$

Overall: $Mg^{2+} + 2Cl^- \rightarrow Mg + Cl_2$ $E^o_{cell} = -3.73V$

Need a battery with slightly more than 3.73V to make the cell run.

Aug 31-8:53 AM

2) Consider the following cell:



a) Is this cell molten or aqueous?

b) Which side is the anode?

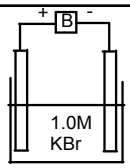
c) What are the anode and cathode composed of?

d) Which substances are competing to reduce? To oxidize?

e) What is the 'overpotential effect'?

Aug 31-9:00 AM

2) Consider the following cell:



a) Is this cell molten or aqueous? aqueous (1.0M)

b) Which side is the anode? left (+ side of battery)

c) What are the anode and cathode composed of?
they are inert: Pt, steel, or C

d) What substance are competing to reduce? To oxidize? Reduce: K^+, H_2O Oxidize: Br^- H_2O

e) What is the 'overpotential effect'?
Water reduces and oxidizes out of order than what's on the table, even though it may have a higher voltage than the substances it 'jumps over'.

Aug 31-9:00 AM

f)Write the half-reactions and overall reaction (with voltages).

g) What observations could you make for the cathode?

Aug 31-9:08 AM

f)Write the half-reactions and overall reaction (with voltage).

Anode: $2Br^- \longrightarrow Br_2 + 2e^-$ $E^o = -1.09V$

Cathode: $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$ $E^o = -0.41V$

Overall: $2Br^- + 2H_2O \longrightarrow Br_2 + H_2 + 2OH$

$E^o_{cell} = -1.50V$

g) What observations could you make for the cathode?

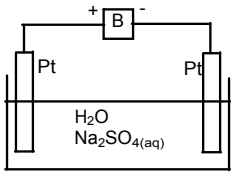
You'd see bubbling due to the H_2 gas formed.

Aug 31-9:08 AM

3) Draw an electrolytic cell capable of electrolyzing water (water is oxidized and reduced) to make H_2 and O_2. Ions must be present in the cell (to maintain cell neutrality at each electrode), so a salt must be part of the cell. Give the half-reactions and the overall reaction (including voltages).

http://educypedia.karadimov.info/library/electrolysis.swf

Aug 31-9:18 AM



*need a salt that will not oxidize or reduce in place of H_2O

Anode: $H_2O \longrightarrow 1/2O_2 + 2H^+ + 2e^-$ $E^o = -0.82V$

Cathode: $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$ $E^o = -0.41V$

Overall: $H_2O \longrightarrow 1/2O_2 + H_2$ $E^o_{cell} = -1.23V$

http://educypedia.karadimov.info/library/electrolysis.swf

Aug 31-9:20 AM

Assignment 16

1) Do Hebden p.238 #64ab & p.242 #65abc, 67, 69, 71

Aug 31-9:26 AM

Assignment 16

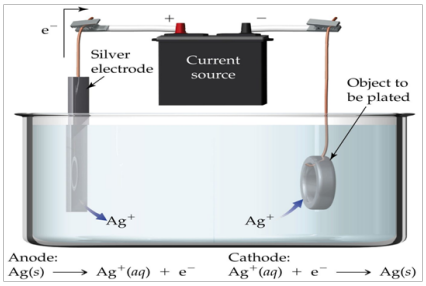
1) Do Hebden p.238 #64ab & p.242 #65abc, 67, 69, 71

answers in the back of Hebden

Aug 31-9:26 AM

Electroplating

Electroplating is an electrolytic process where a metal is plated onto another metal for protection or decoration.



Aug 31-9:27 AM

The metal that will be used for plating is first oxidized at the anode (from metal to ion). It then travels across the electrolyte as a cation (it's attracted to the electrons waiting at the cathode) and is reduced at the cathode (from ion back to metal).
What is the overall reaction of the electroplating cell?

What is the point of electroplating if there is no overall reaction?

Aug 31-9:28 AM

The metal that will be used for plating is first oxidized at the anode (from metal to ion). It then travels across the electrolyte as a cation (it's attracted to the electrons waiting at the cathode) and is reduced at the cathode (from ion back to metal).
What is the overall reaction of the electroplating cell? There is no overall reaction:

$Ag^+ + e^- \longrightarrow Ag$	0.80V
$Ag \longrightarrow Ag^+ + e^-$	-0.80V
cancels out	0.00V

What is the point of electroplating if there is no overall reaction? to move the silver metal from the anode to the object we want to plate it on (the cathode)

Aug 31-9:28 AM

For an electroplating cell, what is the anode and what is the cathode?

What should the electrolyte be composed of?

What voltage must be supplied by the battery in order for the electroplating to occur?

Aug 31-10:33 AM

For an electroplating cell, what is the anode and what is the cathode?
The anode is the plating metal. The cathode is the metal object that is to be plated on.

What should the electrolyte be composed of?
the ion of the plating metal

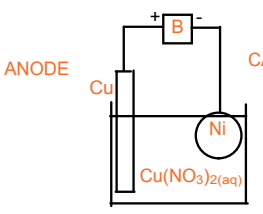
What voltage must be supplied by the battery in order for the electroplating to occur?
anything greater than 0V (any voltage)

Aug 31-10:33 AM

Design an electrolytic cell which will result in a nickel coin being plated with copper. Label the anode, cathode, and electrolyte. Give the half-reactions, the overall reaction, the voltage, and the voltage required for the process to occur.

Aug 31-10:35 AM

Design an electrolytic cell which will result in a nickel coin being plated with copper. Label the anode, cathode, and electrolyte. Give the half-reactions, the overall reaction, the voltage, and the voltage required for the process to occur.



ANODE: $Cu \longrightarrow Cu^{2+} + 2e^-$
CATHODE: $Cu^{2+} + 2e^- \longrightarrow Cu$
OVERALL: none $E^\circ_{cell} = 0.00V$

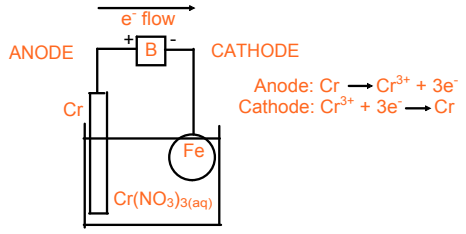
any voltage greater than 0V and this cell will run

Aug 31-10:35 AM

Design an electrolytic cell which could be used to electroplate an iron hood ornament with chromium. Draw the cell and label all parts. Show the direction of electron flow in the wire. Write the half-reactions and identify a suitable electrolyte.

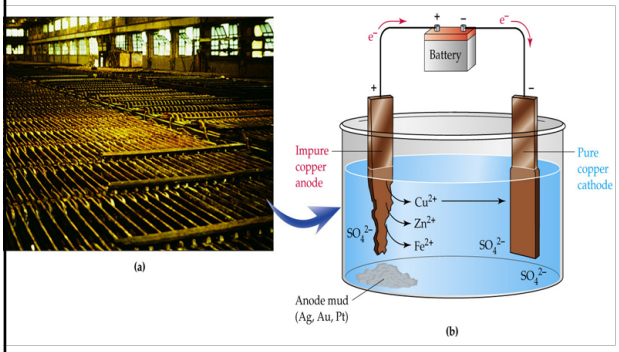
Aug 31-10:41 AM

Design an electrolytic cell which could be used to electroplate an iron hood ornament with chromium. Draw the cell and label all parts. Show the direction of electron flow in the wire. Write the half-reactions and identify a suitable electrolyte.



Aug 31-10:41 AM

Electrorefining
An electrolytic process where impure metal ore becomes pure metal. Below is an example of the electrorefining of copper.



Aug 31-10:44 AM

The impure copper ore anode contains various amounts of zinc, iron, silver, platinum, and gold, but mostly copper. What order will these metals oxidize?

Aug 31-10:45 AM

The impure copper ore anode contains various amounts of zinc, iron, silver, platinum, and gold, but mostly copper. What order will these metals oxidize?

Zn, then Fe, then Cu

Aug 31-10:45 AM

Because copper will always be on the surface of the anode (since the anode is *mostly* copper), silver, platinum, and gold will never get a chance to oxidize and will fall to the bottom of the cell making the 'anode mud'. Therefore, when zinc is on the surface of the ore, it will oxidize. When it isn't, Fe will oxidize. And when neither is on the surface, copper will oxidize.
Therefore, what three cations will exist in the solution?

Which ion out of the three listed is best at reducing?

Aug 31-10:47 AM

Because copper will always be on the surface of the anode (since the anode is *mostly* copper), silver, platinum, and gold will never get a chance to oxidize and will fall to the bottom of the cell making the 'anode mud'. Therefore, when zinc is on the surface of the ore, it will oxidize. When it isn't, Fe will oxidize. And when neither is on the surface, copper will oxidize.
Therefore, what three cations will exist in the solution? Zn^{2+}, Fe^{2+}, Cu^{2+}

Which ion out of the three listed is best at reducing? Cu^{2+}

Aug 31-10:47 AM

Therefore, why will Cu^{2+} be the only ion to reduce onto the pure Cu cathode?

Once all of the copper is reduced, the cell is terminated. The electrolyte is $CuSO_{4(aq)}$, in order to have extra Cu^{2+} ions in solution.

http://www.absorblearning.com/media/attachment.action?quick=uk&att=2190

Aug 31-10:49 AM

Therefore, why will Cu^{2+} be the only ion to reduce onto the pure Cu cathode?
Cu^{2+} will always be reducing as there is such a large amount of it in solution. Zn^{2+} and Fe^{2+} will never get the opportunity to reduce.
Once all of the copper is reduced, the cell is terminated. The electrolyte is $CuSO_{4(aq)}$, in order to have extra Cu^{2+} ions in solution.

http://www.absorblearning.com/media/attachment.action?quick=uk&att=2190

Aug 31-10:49 AM

Assignment 17

1) Do Hebden p.244 #73, 75, 76

2) Read Hebden p.245 & 246 and do #77

http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH7904

Aug 31-10:51 AM

Assignment 17

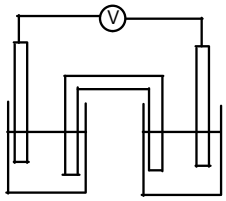
1) Do Hebden p.244 #73, 75, 76

2) Read Hebden p.245 & 246 and do #77

answers in the back of Hebden

http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH7904

Aug 31-10:51 AM



Aug 29-8:35 AM