		Hint	Oxidation Numbers for each Element							
a.	$SnCl_4$	*Rule 2*	Sn	+4	Cl	-1				
b.	Ca_3P_2	*Rule 2*	Ca	+2	P	-3				
c.	SnO	*Rules 4, 5*	Sn	+2	O	-2				
d.	Ag_2S	*Rule 2*	Ag	+1	S	-2				
e.	HI	*Rule 3, 5*	H	+1	I	-1				
f.	N_2H_4	*Rule 3, 5*	N	-2	H	+1				
g.	Al_2O_3	*Rule 4, 5*	Al	+3	O	-2				
h.	S_8	*Rule 1*	S	0						
i.	HNO_2		H	+1	N	+3	O	-2		
j.	O_2		O	0						
k.	H_3O^+	*Rules 3, 4, 6*	H	+1	O	-2				
l.	ClO_3^-	*Rules 4, 6*	Cl	+5	O	-2				
m.	$S_2O_3^{2-}$		S	+2	O	-2				
n.	$KMnO_4$		K	+1	Mn	7	O	-2		
o.	$(NH_4)_2SO_4$		N	-3	H	+1	S	+6	O	-2

2. Determine the oxidation number of carbon in each of the following compounds:

a. methane, CH_4

C = -4

b. formaldehyde, CH_2O +2 -2 +1

C = 0

c. carbon monoxide, CO

C = +2

d. carbon dioxide, CO_2

C = +4

3. When elemental iron is made from Fe_2O_3, it iron oxidized or reduced?

$Fe_2O_3 \rightarrow \overset{0}{Fe}$ $+3 \rightarrow 0$

$2Fe + (-2)3 = 0$ ∴ ↓ oxidation # = reduced

$2Fe + (-6) = 0$ ∴ $Fe = 3$

4. Determine which of the following processes are oxidations and which are reductions:

a. Co^{2+} becomes Co gains e- = reduced ↓oxidation #

b. $2I^-$ becomes I_2.

-1 0 loses e^- ∴ oxidized

c. Fe^{3+} becomes Fe^{2+} ↓ oxidation # ∴ gains e- ∴ reduced

d. Sn^{2+} becomes Sn^{4+} ↑ oxidation # ∴ loses e- ∴ oxidized

5. Determine if each of the following changes is an oxidation, a reduction, or neither:

a. $SO_3^{2-} \rightarrow SO_4^{2-}$

$+4 \rightarrow +6$

∴ oxidized

$SO_3^{2-} \rightarrow S + 3(-2) = 2-$

$S + (-6) = -2$

$S = +4$

$SO_4^{2-} \rightarrow S + 4(-2) = -2$

$S + (-8)$

$S = +6$

b. $CaO \rightarrow Ca$

$2+$ 0

∴ gained electrons = reduction

c. $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$

$Cr + (-8) = -2$

$Cr = +6$

$2Cr + (-14) = -2$

$2Cr = +12$

$Cr = +6$

∴ neither

d. $2I^- \rightarrow I_2$

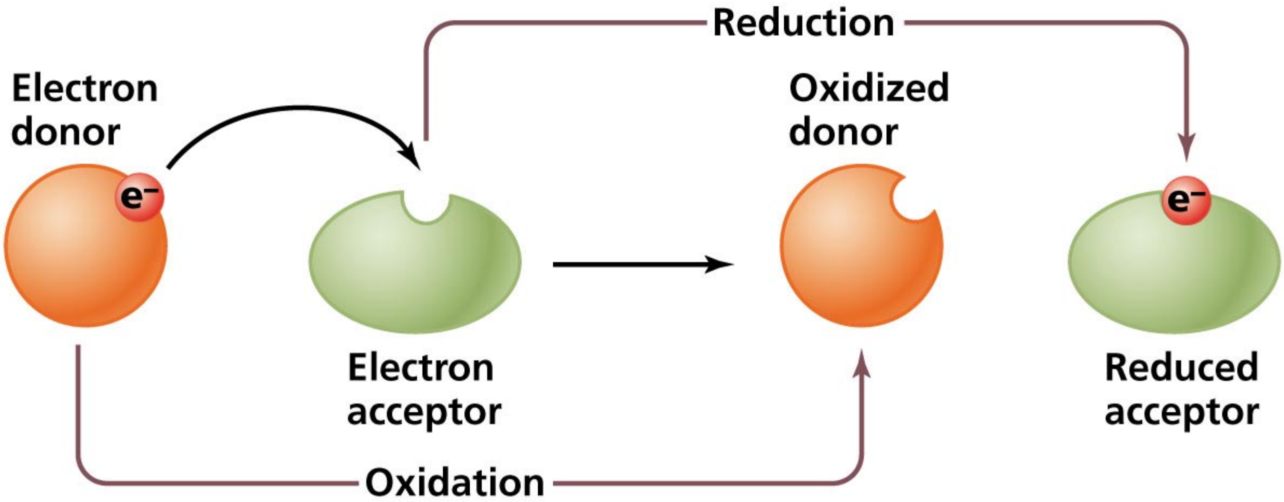
-1 0 ∴ oxidation

e. $IO_3^{1-} \rightarrow \overset{0}{I_2}$

$I + (-6) = -1$

$I = +5$

$+5 \rightarrow 0$ ∴ gain electrons ∴ reduced



Unit 7 - Redox Reactions and Electrochemistry

7.1 - Introduction to Redox Reactions

pages 635-642 in Matter and Change

Redox is short for 'oxidation and reduction,' which are two complimentary types of chemical reactions.

- **Oxidation** originally referred to a reaction where substances combine with oxygen. These types of reactions are commonly called **combustion** or **corrosion**. For example, burning a log or rusting iron.
- **Reduction** originally referred to converting metal ores into pure metals by reducing the mass of the ore.

These two terms have broader meanings today.

The key to all redox reactions is that there is an **exchange of electrons.** That is, when one substance loses electrons, another one gains them.

For example,

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

Since we have aqueous substances, the ions are dissociated... An ionic equation shows us
that:

$Cu_{(s)} + 2Ag^{+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)} + 2Ag_{(s)}$

Now we can remove the spectator ions and observe the net ionic equation.

$Cu_{(s)} + 2Ag^{+}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Ag_{(s)}$

7.1 - Introduction to Redox Reactions

$Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

What do we see now?

1. Copper began as a neutral atom, but became an ion with a +2 charge. An atom becomes positive by **losing** electrons:

   $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^{-}$

   Here, we say that copper has been **oxidized** because it has **lost electrons** (the electrons appear on the product side of the equation).

2. Silver was converted from an ion with a +1 charge to a neutral atom. Therefore, it must have gained an electron:

   $Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$

   We say silver has **gained** electrons and has been **reduced** (electrons appear on the reactant side of the equation).

Now, looking back at the net ionic equation, we can see the electrons gained by silver came from copper.

Note that a substance can **not** lose electrons if there is not another substance to take them. That is **oxidation cannot occur without reduction**.

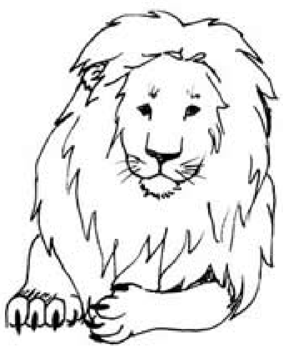
This exchange of electrons is what defines an **oxidation-reduction reaction (redox).**

Here is a mnemonic:

LEO

the Lion Says

GER



LEO

Loss of Electrons is Oxidation

GER

Gain of Electrons is Reduction

Oxidation Numbers

Since redox reactions involve exchanges of electrons between substances, there should be an ordered system to keep track of what gains, what loses,and how many electrons are involved. This record keeping system is called **oxidation numbers**.

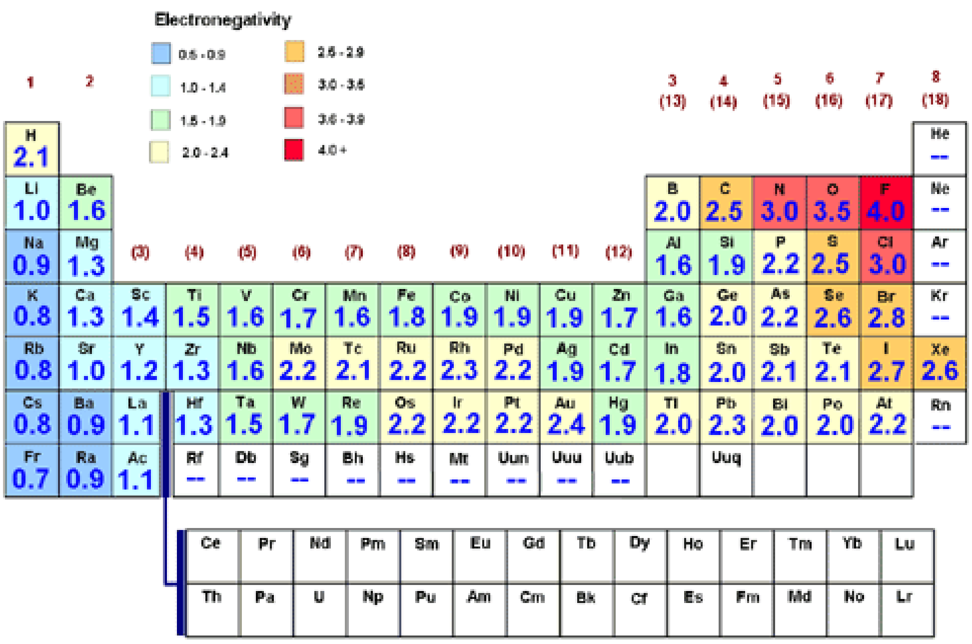
Oxidation numbers are easiest to understand through ionic compounds because we already know there is an exchange of electrons.

However, electronegativity can help us understand oxidation numbers in covalent compounds. REMINDER: **Electronegativity** is the ability of a bonded atom to attract electrons towards itself. *(Table 6))*

-metals have a relatively low electronegativity

-non-metals have a high electronegativity

When two atoms are bonded together by a covalent bond, there is a tug ofwar for the electrons they are sharing due to differences in each atom's electronegativity. The more electronegative atom will pull harder



7.1 - Introduction to Redox Reactions

Consider the water molecule. Oxygen has a higher electronegativity than hydrogen. Therefore, the oxygen pulls the shared electrons closer to itself.

This close proximity of oxygen and the electrons makes oxygen partially negative.

Likewise, the part of the hydrogen that has lost the electron has become partially positive, while the side closest to oxygen has become partially negative.

We use the δ^+ to show the parts of the molecule that are partially positive and the δ^- to show partially negative parts.

We can think of each H 'losing' an electron and the O as 'gaining' two electrons (even though they are technically SHARING the electrons).

Charges given to atoms in this way are called **oxidation numbers**.

For water, hydrogen has an oxidation number of +1 because each has 'lost' an electron. Oxygen has an oxidation number of -2 because it has 'gained' two electrons.

Ex) What are the oxidation numbers for H_2O_2?

Helpful Hint: There are **common oxidation numbers for elements listed in *Table 17*

7.1 - Introduction to Redox Reactions

We will not need to draw Lewis structures often for redox reactions. However, knowing oxidation numbers will be key to our understanding of how these reactions work.

Use the **oxidation number rules** sheet when dealing with redox reactions *(Table 19)*.

- oxidation number of element in stable state = 0
- sum of oxidation numbers for polyatomic compound is 0 unless otherwise specified with accompanying charge

Note that oxidation numbers refer to **each atom** in a molecule or compound, and not the **total** atoms for a given element.

Ex) What is the oxidation number of each element in:

a) $SnCl_4$

b) HI

c) $Mg(NO_2)_2$

d) MgH_2

An increase in the oxidation number of an atom/ion represents oxidation.

A decrease in the oxidation number of an atom/ion represents reduction.

ex) Is Nitrogen oxidized or reduced when NH_3 --> N? How do you know?

7.1 Introduction to Redox Reactions Assignment

		Oxidation Numbers for each Element			
a.	$SnCl_4$	Sn ______	Cl ______		
b.	Ca_3P_2	Ca ______	P ______		
c.	SnO	Sn ______	O ______		
d.	Ag_2S	Ag ______	S ______		
e.	HI	H ______	I ______		
f.	N_2H_4	N ______	H ______		
g.	Al_2O_3	Al ______	O ______		
h.	S_8	S ______			
i.	HNO_2	H ______	N ______	O ______	
j.	O_2	O ______			
k.	H_3O^+	H ______	O ______		
l.	ClO_3^-	Cl ______	O ______		
m.	$S_2O_3^{2-}$	S ______	O ______		
n.	$KMnO_4$	K ______	Mn ______	O ______	
o.	$(NH_4)_2SO_4$	N ______	H ______	S ______	O ______

2. Determine the oxidation number of carbon in each of the following compounds:

   a. methane, CH_4

   b. formaldehyde, CH_2O

   c. carbon monoxide, CO

   d. carbon dioxide, CO_2

3. When elemental iron is made from Fe_2O_3, it iron oxidized or reduced?

4. Determine which of the following processes are oxidations and which are reductions:

a. Co^{2+} becomes Co

b. $2I^-$ becomes I_2.

c. Fe^{3+} becomes Fe^{2+}

d. Sn^{2+} becomes Sn^{4+}

5. Determine if each of the following changes is an oxidation, a reduction, or neither:

a. $SO_3^{2-} \rightarrow SO_4^{2-}$

b. $CaO \rightarrow Ca$

c. $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$

d. $2I^- \rightarrow I_2$

e. $IO_3^{1-} \rightarrow I_2$

7.1 Virtual Redox Lab

Background

Oxidation and reduction reactions have been known for millennia but were not understood until the 17th century. The terms come from metallurgy. Most metals do not naturally exist in their metallic forms (except gold and silver), but were extracted from rocks and minerals. As such the ores were "reduced" to a small amount of metal from a large amount of ore. It was noted that the metals would react with oxygen and form a new substance and hence were oxidized. We now understand that redox (oxidation-reduction) reactions involve the transfer of electrons. Consider for instance, the reaction between copper ions ($Cu^{2+}_{(aq)}$) and zinc metal ($Zn_{(s)}$) react according to the chemical reaction:

$$Cu^{+2}_{(aq)} + Zn_{(s)} \Rightarrow Cu_{(s)} + Zn^{+2}_{(aq)}$$

Electrons were exchanged in this reaction, making it a redox reaction. To make the electron exchange more apparent, we can break this reaction into "half reactions".

$Zn_{(s)} \Rightarrow Zn^{+2}_{(aq)} + 2\,e^-$ (zinc metal gives up electrons; hence zinc is oxidized)

$Cu^{+2}_{(aq)} + 2\,e^- \Rightarrow Cu_{(s)}$ (copper ion gains electrons; hence Cu^{2+} is reduced)

Another way of looking at the above reaction is to consider what the Cu^{2+} ion is doing to the Zn. Cu^{2+} is causing the Zn to be oxidized, so Cu^{2+} is acting as the oxidizing agent. Conversely, Zn is causing Cu^{2+} to be reduced, so Zn is a reducing agent. Reactions such as that between Zn(s) and $Cu^{2+}_{(aq)}$ only go in one direction. In other words, we **will not** see the following reaction occur:

$$Cu_{(s)} + Zn^{+2}_{(aq)} \rightarrow Cu^{+2}_{(aq)} + Zn_{(s)}$$

In other words, Zn is able to reduce Cu^{2+} but Cu is not able to reduce Zn^{2+}. We can summarize this by saying that Zn is a stronger reducing agent that Cu.

Purpose

The purpose of this lab is to order Cu, Mg, Zn and Pb from strongest to weakest reducing agent.

Procedure

All of the materials and equipment you need to carry out the lab can be found at the website below:

http://chemcollective.org/vlab/106

The stockroom contains solutions of Cu^{2+}, Mg^{2+}, Zn^{2+} and Pb^{2+} ions and the corresponding metals (Cu, Mg, Zn, Pb). Your first task is to order Cu, Mg, Zn and Pb from stronger to weakest reducing agent. To do this, you need to obtain a beaker to mix a solid metal with a known solution.

For example, if I wanted to test whether Zn could reduce Cu^{2+}, I would take 10 g of solid zinc and place it in a beaker with 100mL of 0.10M $Cu(NO_3)_2$. Using the concentration of ions (seen on the right side of the webpage), I see the following information:

Species	Molarity
H^+	1.015e-7
OH^-	1.015e-7
Cu^{2+}	~0
Zn^{2+}	1.000e-1
NO_3^-	2.000e-1

You can see that I now have Zn^{2+} ions in my solution, but no Cu^{2+} ions. From this data, I can conclude that Cu^{2+} was reduced (as there are no Cu^{2+} ions present) and Zn^{2+} was oxidized (as there are now Zn^{2+} ions present and I didn't add any to my initial solution). =

By that same logic, if I wanted to test whether Cu could reduce Zn^{2+}, I would take 10 g of solid copper and place it in a beaker with 100mL of 0.10M $Zn(NO_3)_2$. Using the concentration of ions (seen on the right side of the webpage), I see the following information:

Species	Molarity
H^+	1.005e-7
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You can see that I do not have Cu^{2+} ions in my solution, but still have Zn^{2+} ions. From this data, I can conclude that Zn^{2+} was not reduced and Cu was not oxidized (as there are not any Cu^{2+} ions present). Thus, Zn is a stronger reducing agent than Cu.

Your task will be to mix all of the solids with all of the solutions to determine if a redox reaction happens. If a reaction happens, write reaction in the data table below. If no reaction occurs, write no reaction (see examples between the copper and zinc). You will then use this data to order Cu, Mg, Zn, and Pb from strongest to weakest reducing agent.

Data:

	$Cu(NO_3)_{2(aq)}$	$Mg(NO_3)_{2(aq)}$	$Zn(NO_3)_{2(aq)}$	$Pb(NO_3)_{2(aq)}$
$Cu_{(s)}$	N/A	No reaction	No reaction	No reaction
$Mg_{(s)}$	reaction	N/A	reaction	reaction
$Zn_{(s)}$	reaction	No reaction	N/A	reaction
$Pb_{(s)}$	reaction	No reaction	No reaction	N/A

Analysis:

Based on the data you collected above, order Cu, Mg, Zn and Pb from strongest to weakest reducing agent.

(strongest) Mg > Zn > Pb > Cu (weakest)

Appendix

A few things that might be helpful to know when using the virtual lab software:

1. Click on this button  for any glassware you need (beaker, flasks, burets, etc.).
2. Click on this button  for any equipment you need (Bunsen burner, etc.).
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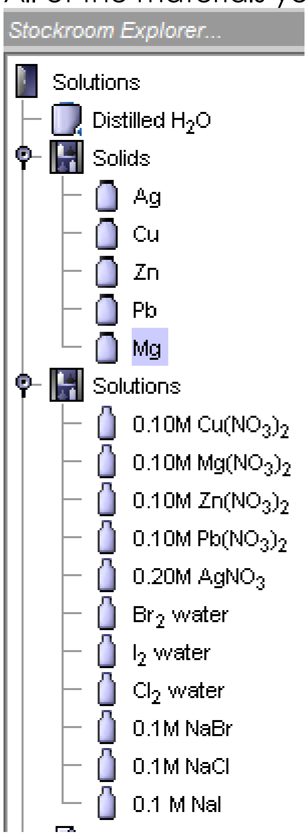

4. To add material to glassware, place the material on top of the glassware until you see a green (+) sign. Then, type in the amount you want to add to the glassware into the transfer amount and click pour

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Your task will be to mix all of the solids with all of the solutions to determine if a redox reaction happens. If a reaction happens, write reaction in the data table below. If no reaction occurs, write no reaction (see examples between the copper and zinc). You will then use this data to order Cu, Mg, Zn, and Pb from strongest to weakest reducing agent.

Data:

	$Cu(NO_3)_{2(aq)}$	$Mg(NO_3)_{2(aq)}$	$Zn(NO_3)_{2(aq)}$	$Pb(NO_3)_{2(aq)}$
$Cu_{(s)}$	N/A		No reaction	
$Mg_{(s)}$		N/A		
$Zn_{(s)}$	reaction		N/A	
$Pb_{(s)}$				N/A

Analysis:

Based on the data you collected above, order Cu, Mg, Zn and Pb from strongest to weakest reducing agent.

Appendix 1: Helpful Hints

A few things that might be helpful to know when using the virtual lab software:

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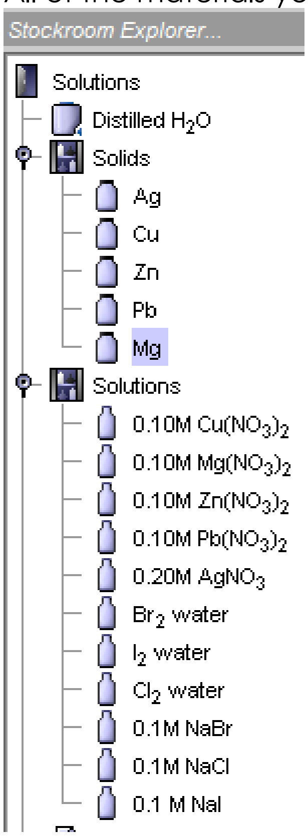

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7.2 Balancing Redox Reactions with Half Reactions Assignment

1. For each of these reactions, determine whether or not it is a redox reaction. If any are, identify oxidizing and reducing agents in those reactions.

a. $CaBr_2 + Pb(NO_3)_2 \rightarrow PbBr_2 + Ca(NO_3)_2$

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

NOT REDOX

element	Initial Ox. No		Final Ox. No.	e⁻ gained or lost	Oxidized or reduced	Agent
		→				
		→				

b. $P_4 + 5O_2 \rightarrow P_4O_{10}$

0 0 4(P)+(-2)10 = 0

4P - 20 = 0

P = 5

element	Initial Ox. No		Final Ox. No.	e⁻ gained or lost	Oxidized or reduced	Agent
P	0	→	5	5 lost	OX	reducing
O	0	→	-2	2 gained	RED	oxidizing

c. $SnCl_2 + 2\,FeCl_3 \rightarrow 2\,FeCl_2 + SnCl_4$

+2 -1 +3 -1 +2 -1 +4 -1

element	Initial Ox. No		Final Ox. No.	e⁻ gained or lost	Oxidized or reduced	Agent
Sn	+2	→	+4	2 lost	OX	Reducing
Fe	+3	→	+2	1 gained	RED	Oxidizing

2. Break each equation into two half-reactions. Identify each half-reaction as oxidation or reduction.

a. $Cu + 2H^+ \rightarrow Cu^{2+} + H_2$

$Cu \rightarrow Cu^{2+} + 2e^-$ oxidation

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

b. $2Al + 3S \rightarrow Al_2S_3$

$+3\ -2$

$2Al \rightarrow 2Al^{3+} + 6e^-$ oxidation

$6e^- + 3S \rightarrow 3S^{2-}$ reduction

3. Balance the following equations using the half-reaction method. Identify what is reduced and what is the reducing agent.

a. $\overset{0}{Na} + \overset{0}{Br_2} \rightarrow \overset{+1}{Na}\overset{-1}{Br}$

$2Na \rightarrow 2Na^+ + \cancel{2e^-}$ red agent

$\cancel{2e^-} + Br_2 \rightarrow 2Br^-$ reduced

$2Na + Br_2 \rightarrow 2NaBr$

b. $Zn + S \rightarrow \overset{2+}{Zn}\overset{2-}{S}$

$Zn \rightarrow Zn^{2+} + \cancel{2e^-}$ red agent

$\cancel{2e^-} + S \rightarrow S^{2-}$ red

$Zn + S \rightarrow ZnS$

c. $Au^{3+}{}_{(aq)} + Cd_{(s)} \rightarrow Au_{(s)} + Cd^{2+}{}_{(aq)}$

$2(Au^{3+} + 3e^- \rightarrow Au)$

$3(Cd \rightarrow Cd^{2+} + 2e^-)$

$2Au^{3+} + \cancel{6e^-} \rightarrow 2Au$ red

$3Cd \rightarrow 3Cd^{2+} + \cancel{6e^-}$ reducing agent

$2Au^{3+}{}_{(aq)} + 3Cd_{(s)} \rightarrow 2Au_{(s)}\ 3Cd^{2+}{}_{(aq)}$

4. Write a balanced equation for each of the following half-reactions, and state whether it represents oxidation or reduction.

a. $HClO_2 \rightarrow Cl^-$ (acidic)

$4e^- + 3H^+ + HClO_2 \rightarrow Cl^- + 2H_2O$

Being that electrons are being gained, this is a reduction half reaction.

b. $Cr(OH)_3 \rightarrow CrO_4^{2-}$ (basic)

$5OH^- + H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+ + 5OH^-$

$5OH^- + \cancel{H_2O} + Cr(OH)_3 \rightarrow CrO_4^{2-} + \overset{4}{\cancel{5}}H_2O$

$5OH^- + Cr(OH)_3 \rightarrow CrO_4^{2-} + 4H_2O + 3e^-$

Being that electrons are being lost, this is an oxidation half reaction.

c. $H_2GeO_3 \rightarrow Ge$ (acidic)

$4e^- + 4H^+ + H_2GeO_3 \rightarrow Ge + 3H_2O$

Being that electrons are being gained, this is a reduction half reaction.

d. $SbO_2^- \rightarrow Sb$ (basic)

$4OH^- + 4H^+ + SbO_2^- \rightarrow Sb + 2H_2O + 4OH^-$

$\overset{2}{\cancel{4}}H_2O + SbO_2^- \rightarrow Sb + \cancel{2H_2O} + 4OH^-$

$2H_2O + SbO_2^- + 3e^- \rightarrow Sb + 4OH^-$

Being that electrons are being gained, this is a reduction half reaction.

5. $Cu + NO_3^- \rightarrow Cu^{2+} + NO$

$(Cu \rightarrow Cu^{2+} + 2e^-)3$

$(3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O)2$

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

7.2 - Balancing Redox Reactions Using Half Reactions

pages 650-653 in Matter and Change

Reduction

Oxidant + e^- → Product

(Gain of Electrons) (Oxidation Number Decreases)

Oxidation

Reductant → Product + e^-

(Loss of Electrons) (Oxidation Number Increases)

We can use oxidation numbers to identify which reactions are redox, which element is gaining electrons, and which is losing electrons.

For example,

$2Na + Cl_2 \rightarrow 2NaCl$

It is useful to write in the oxidation of every element in every compound above the element in the equation.

Remember that the balancing coefficients in the chemical equations **do not** affect the oxidation numbers.

A chart is a useful way of organizing the changes in oxidation number for each element:

element	Initial Ox. #	Final Ox. #	Change in e^-	Oxidized or Reduced?

By looking at the table, we see:

i. oxidation numbers **did** change, so it is a redox reaction.

ii. Na increased its oxidation number from 0 to +1. Therefore, it has **lost**electrons. (**LEO**)

iii. Cl decreased its oxidation number from 0 to -1. Therefore it has **gained** electrons. (**GER**)

So, an **increase** in oxidation number indicates **oxidation**, while a **decrease** in oxidation number indicates **reduction**.

Ex 1) Consider the following reaction:

$2Mg + O_2 \rightarrow 2MgO$

Summarize the changes in oxidation number for each element, determine how many electrons has been transferred per atom, and identify what has been oxidized and what has been reduced.

element	Initial Ox. #	Final Ox. #	Change in e^-	Oxidized or Reduced?

Recall that oxidation cannot occur without reduction (and vice versa). Therefore, if one substance cannot accept the electrons a substance is giving away, they cannot be given away in the first place. That is, one **allows** for the other to occur.

- We call the substance that is oxidized the **reducing agent** because it allows another element to be reduced.
- We call the substance that is being reduced the **oxidizing agent** because it allows another element to be oxidized.

So, in our example above, since Mg was ____________ and it is the ________________.
Likewise, O is the _________________ because it was ________________.

Ex 2) In the chemical reaction

$N_2 + 3H_2 \rightarrow 2NH_3$

summarize the changes in oxidation numbers, determine the number of electrons transferred per atom, identify what has been oxidized or reduced, and identify the oxidizing and reducing agents.

element	Initial Ox. #	Final Ox. #	Change in e^-	Oxidized or Reduced?	Type of Agent?

**If there is no transfer of electrons, the reaction is not a REDOX reaction. Charges of atoms will remain the same in a situation like this.

7.2 - Balancing Redox Reactions with Half Reactions

Now that we can recognize redox reactions, we need to know how to balance them.

Balancing redox reactions can be tricky; therefore, there are two different methods we will look at: Half Reaction and Oxidation Number Methods

Oxidation and reduction always occur together. Each part of the reaction is called a "half-reaction" *(Table 18 and 19)*

Steps:

1. Break the reaction into **two** half reactions (an oxidation and a reduction) and remove any spectator ions from the equations (or use Table 18 and 19 to help find half reactions)

   Note - oxidation numbers may help decide what is being oxidized and what is being reduced.

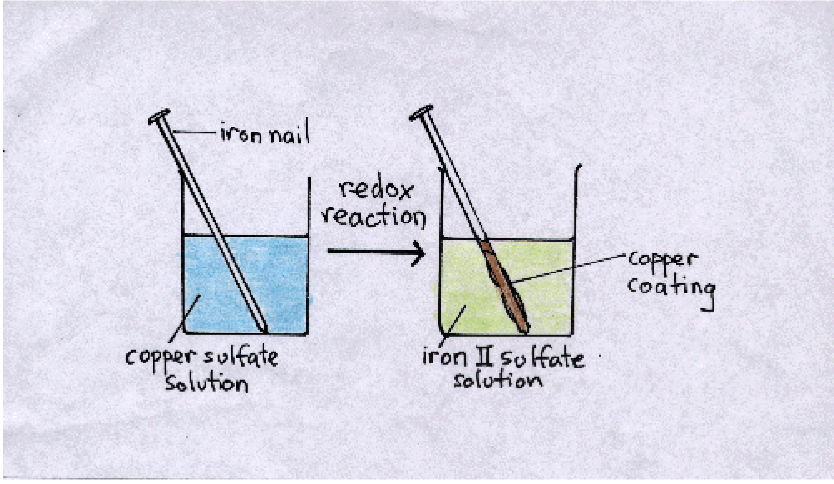
2. Balance each half-reaction separately, first by number of atoms, then by charge by adding electrons to the appropriate side of the equation.

3. Compare the number of electrons in each equation. They must be equal. If they are not equal, make them equal by multiplying everything in **one** half-reaction by a coefficient.

4. Add the two equations together and replace all spectator ions in the correct spot. Electrons do not belong here (they should cancel). Furthermore, all compounds previously broken up are written as bonded together again.

Ex) Balance the following using the half reaction method:

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



Ex) Balance the following using half reactions:

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + Fe_2(SO_4)_{3\ (aq)}$

•

Balancing Half Reactions in **Acidic** Solutions

When you are told that a half reaction is in an acidic solution we need to understand that the half reaction itself does not show all substances present.

Since a half reaction only explains how oxidation numbers change for a particular substance, we may need to add $\underline{H_2O}$ (since we are in a solution) and $\underline{H^+}$ (since the solution is acidic) to the equation.

Ex) Write the balanced equation for the half reaction in which nitric oxide (NO) is reduced to nitrous oxide (N_2O) in an acidic solution.

The Steps:

1. Balance all elements for mass **except** for hydrogen and oxygen.

2. Add H_2O to one side of the reaction to balance the oxygens first.

3. You will probably have a discrepancy with the number of hydrogens, so add a number of H^+ to the opposite side as H_2O to balance them out.

4. Now, add electrons to one of the sides so that each side of the reaction is electrically the same.

This is now the balanced chemical equation for this HALF-REACTION; Note, that since electrons are ______________, this process is _____________

Balancing Half Reactions in **Basic** Solutions

We will use a similar process, realizing this time that water and **OH^-** must be present.

Ex) Write the balanced equation for the oxidation half reaction in which Cl_2 is oxidized to ClO_3^- in basic solution.

The Steps:

1. Balance all elements for mass except hydrogen and oxygen.

2. Add H_2O to balance for oxygens.

3. Add H^+ to balance for hydrogens.

4. Add the same number of OH^- as H^+ ions in previous step to **both** sides of the equation.

5. Cancel off as many waters as you can (you should now have H_2O on both sides of the equation).

6. Add in electrons to make both sides equal in charge.

This is now the balanced chemical equation for this HALF-REACTION; Note, that sinc electrons are __________, _____________ has occurred in this process

7.2 Balancing Redox Reactions with Half Reactions Assignment

1. For each of these reactions, determine whether or not it is a redox reaction. If any are, identify oxidizing and reducing agents in those reactions.

a. $CaBr_2 + Pb(NO_3)_2 \rightarrow PbBr_2 + Ca(NO_3)_2$

element	Initial Ox. No		Final Ox. No.	e^- gained or lost	Oxidized or reduced	Agent
		→				
		→				

b. $P_4 + 5O_2 \rightarrow P_4O_{10}$

element	Initial Ox. No		Final Ox. No.	e^- gained or lost	Oxidized or reduced	Agent
		→				
		→				

c. $SnCl_2 + 2\ FeCl_3 \rightarrow 2\ FeCl_2 + SnCl_4$

element	Initial Ox. No		Final Ox. No.	e^- gained or lost	Oxidized or reduced	Agent
		→				
		→				

2. Break each equation into two half-reactions. Identify each half-reaction as oxidation or reduction.

a. $Cu + 2\,H^+ \rightarrow Cu^{2+} + H_2$

b. $2\,Al + 3\,S \rightarrow Al_2S_3$

3. Balance the following equations using the half-reaction method. Identify what is reduced and what is the reducing agent.

a. $Na + Br_2 \rightarrow NaBr$

b. $Zn + S \rightarrow ZnS$

c. $Au^{3+}_{(aq)} + Cd_{(s)} \rightarrow Au_{(s)} + Cd^{2+}_{(aq)}$

4. Write a balanced equation for each of the following half-reactions, and state whether it represents oxidation or reduction.

a. $HClO_2 \rightarrow Cl^-$ (acidic)

b. $Cr(OH)_3 \rightarrow CrO_4^{2-}$ (basic)

c. $H_2GeO_3 \rightarrow Ge$ (acidic)

d. $SbO_2^- \rightarrow Sb$ (basic)

5. CHALLENGE: By using **balanced** half-reaction equations from your tables work out **overall** redox equation for the reaction below in an acidic solution.

a. $Cu + NO^-{}_3 \rightarrow Cu^{2+} + NO$

7.3 – Balancing Equations with Oxidation Numbers Assignment

1. Balance the following redox reactions using the oxidation number method.

a. $\overbrace{SnCl_2 + HgCl_2 \rightarrow SnCl_4}^{ox} + HgCl$

+2 -1 +2 -1 +4 -1 +1 -1

red

Sn : +2 → +4 : +2 change

Hg : +2 → +1 : -1 change.

Hg : Sn

1 : 2

$1SnCl_2 + 2HgCl_2 \rightarrow 1SnCl_4 + 2HgCl$

b. $HNO_3 + \overbrace{H_2S \rightarrow NO + S}^{ox} + H_2O$

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

red

N : +5 → +2 : -3 change

S : -2 → 0 : +2 change

N : S

2 3

$2HNO_3 + 3H_2S \rightarrow 2NO + 3S + 4H_2O$

c. $NaClO + \overbrace{H_2S \rightarrow NaCl + H_2SO_4}^{ox}$

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

red

S : -2 → +6 ⇒ +8 change

Cl : +1 → -1 ⇒ -2 change

Cl : S

2 8

$8NaClO + 2H_2S \rightarrow 8NaCl + 2H_2SO_4$

red

d. $MnO_4^- + H^+ + Cl^- \rightarrow Mn^{2+} + Cl_2 + H_2O$

+7 -2 +1 -1 2+ 0 +1 -2

ox

Mn : +7 → +2 ⇒ -5 change

Cl : -1 → 0 ⇒ +1 change

Mn : Cl

1 5

$MnO_4^- + 8H^+ + 5Cl^- \rightarrow Mn^{2+} + \frac{5}{2}Cl_2 + 4H_2O$

x2

$2MnO_4^- + 16H^+ + 10Cl^- \rightarrow 2Mn^{2+} + 5Cl_2 + 8H_2O$

2. Balance each of the following redox reactions in **acidic** solutions using both methods:

a) $ClO_4^- + Br^- \rightarrow Cl^- + BrO_3^-$

+7 -2 -1 -1 +5 -2

i) Oxidation Numbers:

Cl: +7 → -1 ⇒ -8 change

Br: -1 → +5 ⇒ +6 change

Br : Cl

8 : 6

$6ClO_4^- + 8Br^- \rightarrow 6Cl^- + 8BrO_3^-$

ii) Half Reactions

$(8e^- + 8H^+ + ClO_4^- \rightarrow Cl^- + 4H_2O)6$

$(3H_2O + Br^- \rightarrow BrO_3^- + 6H^+ + 6e^-)8$

$\cancel{48e^-} + \cancel{48H^+} + \cancel{24H_2O} + 6ClO_4^- + 8Br^- \rightarrow 6Cl^- + \cancel{24H_2O} + 8BrO_3^- + \cancel{48H^+} + \cancel{48e^-}$

$\boxed{6ClO_4^- + 8Br^- \rightarrow 6Cl^- + 8BrO_3^-}$

a) $HNO_3 + Cu \rightarrow NO_2 + Cu^{2+}$

+1 +5 -2 0 +4 -2 2+

i) Oxidation Numbers:

N: +5 → +4 : -1 change

Cu: 0 → 2+ : +2 change

Cu : N

1 2

$2H^+ + 2HNO_3 + 1Cu \rightarrow 2NO_2 + 1Cu^{2+} + 2H_2O$

ii) Half Reactions

$(1e^- + H^+ + HNO_3 \rightarrow NO_2 + H_2O)2$

$\cancel{2e^-} + 2H^+ + 2HNO_3 \rightarrow 2NO_2 + 2H_2O$

$Cu \rightarrow Cu^{2+} + \cancel{2e^-}$

$2H^+ + 2HNO_3 + Cu \rightarrow 2NO_2 + Cu^{2+} + 2H_2O.$

3. Balance the following redox reactions in **basic solutions** using both methods:

a) $ClO_3^- + MnO_2 \rightarrow Cl^- + MnO_4^-$

$+5\ -2 \quad +4\ -2 \quad -1 \quad +7\ -2$

i) Oxidation Numbers

Mn: $+4 \rightarrow +7 \rightarrow +3$ change

Cl: $+5 \rightarrow -1 \rightarrow -6$ change

Mn : Cl

6 : 3

$6OH^- + \cancel{3H_2O} + 3ClO_3^- + 6MnO_2 \rightarrow 3Cl^- + 6MnO_4^- + \underbrace{6H^+ + 6OH^-}_{6H_2O}$

$6OH^- + 3ClO_3^- + 6MnO_2 \rightarrow 3Cl^- + 6MnO_4^- + 3H_2O$

ii) Half Reactions

$6e^- + \cancel{6OH^-} + \cancel{6H^+}\ (\cancel{6}H_2O\ 3) + ClO_3^- \rightarrow Cl^- + \cancel{3H_2O} + 6OH^-$

$(4OH^- + \cancel{2H_2O} + MnO_2 \rightarrow MnO_4^- + \underbrace{\cancel{4H^+ + 4OH^-}}_{\cancel{4}H_2O\ 2} + 3e^-)\ 2$

$\cancel{6e^-} + \cancel{3H_2O} + ClO_3^- + \cancel{8}OH^-\ 2 + 2MnO_2 \rightarrow Cl^- + 2MnO_4^- + \cancel{4}H_2O\ 1 + \cancel{6OH^-} + \cancel{6e^-}$

$ClO_3^- + 2OH^- + 2MnO_2 \rightarrow Cl^- + 2MnO_4^- + H_2O$

b) $ReO_4^- + IO^- \rightarrow IO_3^- + Re$

$+7\ -2 \quad +1\ -2 \quad +5\ -2 \quad 0$

i) Oxidation Numbers

Re: $+7 \rightarrow 0$: -7 change

I: $+1 \rightarrow +5$: $+4$ change

Re : I

4 : 7

$+4OH^- \quad 4H^+ + 4ReO_4^- + 7IO^- \rightarrow 7IO_3^- + 4Re + \cancel{2H_2O} \quad +4OH^-$

$\cancel{4}H_2O\ 2$

$2H_2O + 4ReO_4^- + 7IO^- \rightarrow 7IO_3^- + 4Re + 4OH^-$

ii) Half Reactions

$7e^- + \overbrace{8OH^- + 8H^+}^{\cancel{8}H_2O\ 4} + ReO_4^- \rightarrow Re + \cancel{4H_2O} + 8OH^-\ )\ 4$

$(4OH^- + \cancel{2H_2O} + IO^- \rightarrow IO_3^- + \underbrace{4H^+ + 4OH^-}_{\cancel{4}H_2O\ 2} + 4e^-)\ 7$

$\cancel{28e^-} + \cancel{16}H_2O\ 2 + 4ReO_4^- + \cancel{28OH^-} + 7IO^- \rightarrow 4Re + \cancel{32}OH^-\ 4 + 7IO_3^- + \cancel{14H_2O} + \cancel{28e^-}$

$2H_2O + 4ReO_4^- + 7IO^- \rightarrow 4Re + 4OH^- + 7IO_3^-$

7.3 - Balancing Redox Reactions with Oxidation Numbers

pages 644-649 in Matter and Change

I	II											III	IV	V	VI	VII
H +1																
Li +1	Be +2											B +3	C +4 +2 -4	N +5 +4 +3 +2 +1 -3	O -2	F -1
Na +1	Mg +2											Al +3	Si +4	P +5 +3 -3	S +6 +4 -2	Cl +7 +5 +3 +1 -1
K +1	Ca +2	Sc +3	Ti +4 +3	V +5 +4 +3 +2	Cr +6 +3 +2	Mn +7 +4 +3 +2	Fe +3 +2	Co +3 +2	Ni +2	Cu +2 +1	Zn +2	Ga +3 +1	Ge +4 +2	As +5 +3 -3	Se +6 +4 -2	Br +7 +5 +3 +1 -1
Rb +1	Sr +2									Ag +1	Cd +2	In +3 +1	Sn +4 +2	Sb +5 +3	Te +6 +4 -2	I +7 +5 +3 +1 -1

7.3 - Balancing Redox Reactions with Oxidation Numbers

The second balancing method for Redox reactions is the **oxidation number method.** This is very helpful for difficult equations...

For example: $Cu_{(s)} + Ag^{+}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + Ag_{(s)}$

appears to be balanced, but the **charges** are not balanced.

The total electrical charge on the reactant side must equal the total electrical charge on the product side.

Therefore, like mass (Conservation of Mass), charge is conserved during a chemical reaction.

To balance the equation using oxidation numbers we will follow these steps:

1. Assign oxidation numbers to all atoms in the equation

2. Identify the atoms that are oxidized and the atoms that are reduced

3. Determine the change in the oxidation number for the atoms that are oxidized and for the atoms that are reduced

4. Make the change in oxidation numbers equal in magnitude by adjusting the coefficients of the equation

5. If necessary, use the conventional method (by mass) to balance the remainder of the equation

Ex 1) $Cu_{(s)} + HNO_{3(aq)} \dashrightarrow Cu(NO_3)_{2(aq)} + NO_2(g) + H_2O_{(l)}$

•

7.3 - Balancing Redox Reactions with Oxidation Numbers

Ex 2) $MnO_4^- + Fe^{2+} + H^+ \rightarrow Mn^{2+} + Fe^{3+} + H_2O$

Ex 3) $NH_3 + O_2 \rightarrow NO_2 + H_2O$

7.3 - Balancing Redox Reactions with Oxidation Numbers

Note - if a polyatomic ion stays intact, the oxidation numbers of its elements will not change.

Ex) $K_2Cr_2O_7 + NaI + H_2SO_4 \rightarrow Cr_2(SO_4)_3 + I_2 + H_2O + Na_2SO_4 + K_2SO_4$

•

7.3 - Balancing Redox Reactions with Oxidation Numbers

Overall Redox In Acidic and Basic Solutions

Ex) Balance the following redox reaction in an acidic solution:

$P_4 + IO_3^- \Rightarrow H_2PO_4^- + I^-$

Method 1 - Oxidation Numbers

Now we must balance for mass. Start with oxygen first by adding the H_2O molecule. Finish off by balancing the hydrogens by adding H^+.

Method 2 - Half Reactions

$P_4 + IO_3^- \Rightarrow H_2PO_4^- + I^-$

Break the reaction up into two half reactions.

1. Recall that step 1 is to balance for mass. One reaction will require you to add H_2O and H^+.

2. Balance for charge by adding the correct number of electrons in each half reaction.

3. Finally, get the number of electrons equal in both half reactions by multiplying by a coefficient.

4. Add these two half reactions up to get the overall balanced redox reaction.

•

7.3 - Balancing Redox Reactions with Oxidation Numbers

Redox reactions that are in basic solutions are similar, however we need to deal with adding OH^- as well.

Ex) Balance the equation for the reaction of the permanganate ion with the sulfite ion in a **basic** solution to give manganese dioxide (MnO_2) and the sulfate ion.

Method 1 - Oxidation Numbers

In this method, we need add H_2O on the opposite side to balance for oxygens, use H^+ to balance for charge; then add the equivalent number of OH^- ions to both sides.

Method 2 - Half Reactions

Note that each half reaction will contain oxygen, so you will need to add H_2O to both equations. (Just like in 7.2) The best way to do this is to treat this reaction as if it were in an acidic solution then neutralize the number of H^+ in the final equation with the same number of OH^- on both sides.

7.3 - Balancing Redox Reactions with Oxidation Numbers

7.3 – Balancing Equations with Oxidation Numbers Assignment

1. Balance the following redox reactions using the oxidation number method.

a. $SnCl_2 + HgCl_2 \text{ --> } SnCl_4 + HgCl$

b. $HNO_3 + H_2S \text{ --> } NO + S + H_2O$

c. $NaClO + H_2S \text{ --> } NaCl + H_2SO_4$

d. $MnO_4^- + H^+ + Cl^- \text{ --> } Mn^{2+} + Cl_2 + H_2O$

7.3 - Balancing Redox Reactions with Oxidation Numbers

2. Balance each of the following redox reactions in **acidic** solutions using both methods:

a) $ClO_4^- + Br^- \rightarrow Cl^- + BrO_3^-$

i) Oxidation Numbers:

ii) Half Reactions

a) $HNO_3 + Cu \rightarrow NO_2 + Cu^{2+}$

i) Oxidation Numbers:

ii) Half Reactions

7.3 - Balancing Redox Reactions with Oxidation Numbers

3. Balance the following redox reactions in **basic** solutions using both methods:

a) $ClO_3^- + MnO_2 \rightarrow Cl^- + MnO_4^-$

i) Oxidation Numbers

ii) Half Reactions

b) $ReO_4^- + IO^- \rightarrow IO_3^- + Re$

i) Oxidation Numbers

ii) Half Reactions

7.4 Electrochemistry Worksheet

1. Calculate the voltage produced for each of the electrochemical cells containing the substances below. On each diagram, label:

   i. A cathode and the substance it is made of
   ii. An anode and the substance it is made of
   iii. The correct half reaction under each beaker
   iv. The ions coming off or attaching to each electrode
   v. The direction of the ions in the salt bridge
   vi. The direction of the flow of electrons

-0.41 -0.13

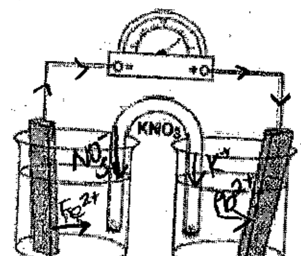
a) ($Fe \mid Fe^{2+}$) and ($Pb \mid Pb^{2+}$)

ox red
anode cathode

Red. $Pb^{2+} + 2e \rightarrow Pb$ -0.13

ox $Fe_{(s)} \longrightarrow Fe^{2+} + 2e$ $+0.44$

$Pb^{2+} + Fe \rightarrow Pb + Fe^{2+}$ $E = 0.31\,V$



anode cathode

red ox

b) ($Cr \mid Cr^{3+}$) and ($Li \mid Li^{+}$)

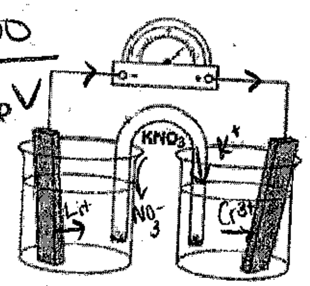
cath. ano.

RED $Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$ -0.74

OX $3(Li_{(s)} \rightarrow Li^+ + e^-)$ $+3.00$

$E = 2.26\,V$

$Cr^{3+}_{(aq)} + 3Li_{(s)} \rightarrow 3Li^+_{(aq)} + Cr_{(s)}$



$3Li \rightarrow 3Li^+ + 3e^-$
ANODE

$Cr^{3+} + 3e^- \rightarrow Cr$
CATHODE

2. Determine the E° for a (Ag+ | Ag) and (Zn | Zn^{2+}) electrochemical cell

RED $(Ag^+ + 1e^- \rightarrow Ag)2$ $+0.80$

OX. $Zn \rightarrow Zn^{2+} + 2e^-$ $+0.76$

$2Ag^+ + Zn \rightarrow 2Ag^+ Zn^{2+}$ $E = 1.56V$

3. Determine E° for a $Fe_{(s)}$ | $Fe^{2+}_{(aq)}$ | | $Cu^{2+}_{(aq)}$ | $Cu_{(s)}$

OX $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$ $+0.44$

RED $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $+0.34$

$Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$ $E = 0.78V$

4. a) An electrochemical cell is created using gold and magnesium half-cells. Write the redox reaction. Determine which half-cell will undergo oxidation and which will undergo reduction, identify which substance is the anode and cathode, and calculate the voltage for the cell. You do not need to draw a diagram of the cell.

RED $2(Au^{3+} + 3e^- \rightarrow Au_{(s)})$ $+1.50$ CATHODE

OX $(Mg_{(s)} \rightarrow Mg^{2+} + 2e^-)3$ $+2.37$ ANODE

$2Au^{3+} + 3Mg_{(s)} \rightarrow 3Mg^{2+} + 2Au$ $E = 3.87V$

b. If the mass of the magnesium electrode ~~changes~~ decreases. by 5.0 g, what will be the change in mass of the gold electrode, and will its mass increase or decrease? (hint – use mass to mass stoichiometry using the redox equation found in part a)

mass will increase

Mg: $n = \frac{wt}{mm} = \frac{5.0g}{24.31g/mol} = 0.206$ mol Mg

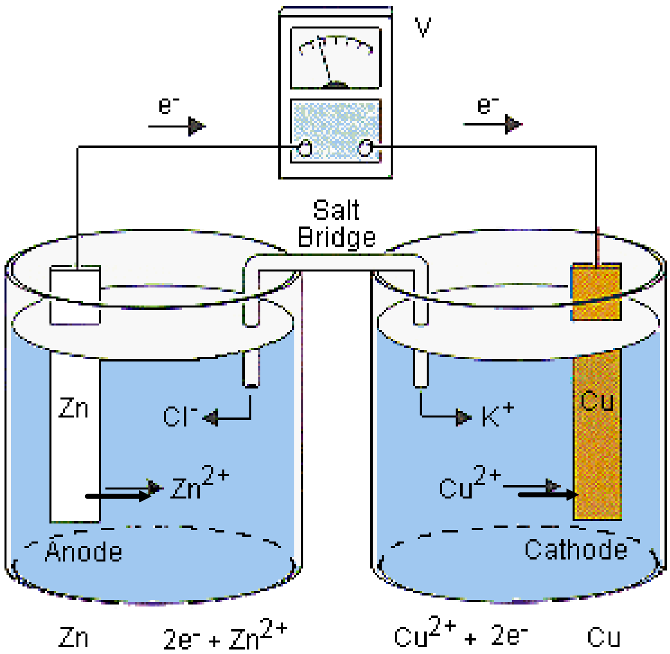
$\frac{3 mol Mg}{0.206 mol} = \frac{2 mol Au}{x}$

$x = 0.137$ mol Au

$wt = n \cdot mm = (0.137 mol)(196.97 g/mol) = 27.0g$ Au

7.4 - Introduction to Electrochemistry and Calculating Voltages

pages 633-642 in Heath



During a redox reaction, electrons are passed from one substance to another.

- We call the flow of electrons **electric current**. This current can be harnessed to do work.

Electrochemistry is the branch of chemistry that deals with the conversion between chemical and electrical energy.

An **electrochemical cell** is the basic unit of a battery. It converts energy from a spontaneous redox reaction into electricity (from the flow of electrons through a metal or ion movement in a solution).

- This transformation from chemical energy to electricity will happen as the electrons are passed to from one substance to another in a redox reaction.

For example, an electrochemical cell can be made by using the following reaction:

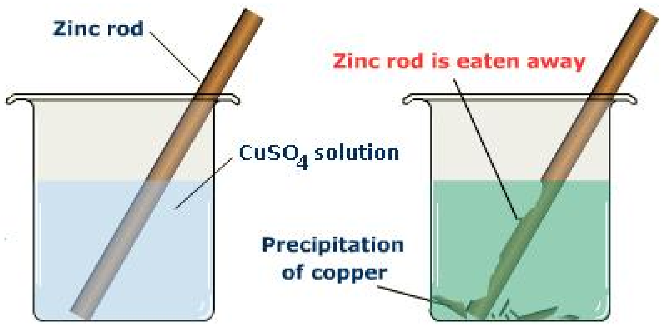
$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

This reaction involves the two half reactions:

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ and $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

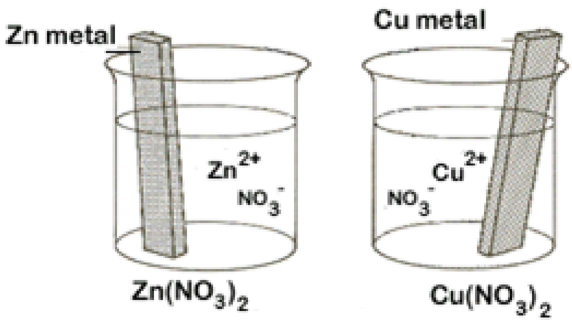
oxidation reduction

To create electricity, we must have our electrons pass through an **external circuit**. If we simply placed Zn in Cu^{2+} ions a reaction would occur, but electricity would not be created.



Here is how an electrochemical cell can be created with this reaction:

1. We need two beakers containing *electrolytic solutions*. One beaker will contain $Zn(NO_3)_2$ and the other will contain $Cu(NO_3)_2$. Zn metal will be placed in the first beaker while Cu metal will be in the second (these are called the **electrodes; electrically conducting solids which are placed in contact with electrolyte solutions**).

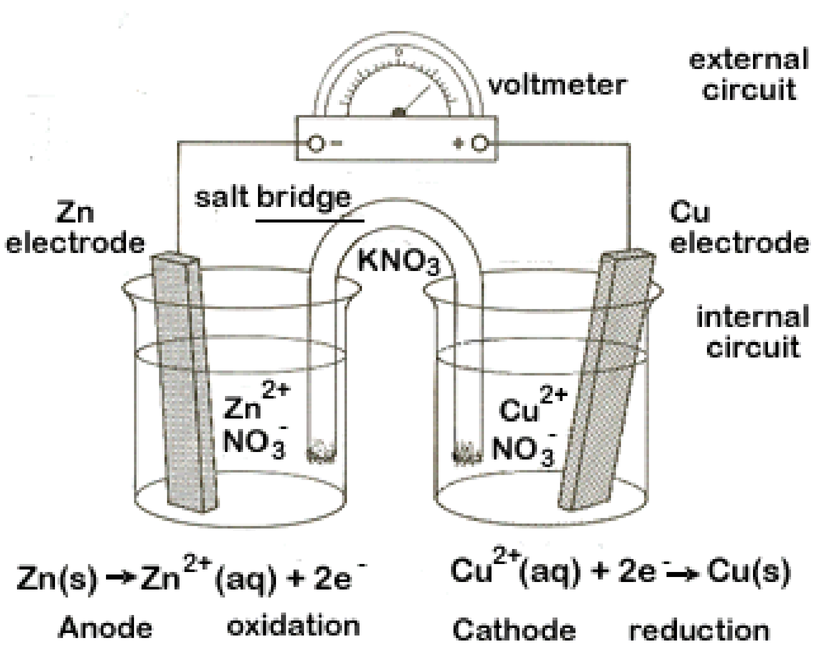


Since the electrons cannot move between the substances yet our redox reaction cannot occur yet. Therefore each beaker contains half a cell.

2. The two half cells must be connected in two ways.

First we must connect our electrodes with wire. Connected to the wire should be a voltmeter (a device that detects electric current). This is known as the **external circuit**.

Second we add a **salt bridge**, which is a U-shaped tube that also contains an electrolytic solution (KNO_3 in this case). This solution will allow electrons to flow freely between the two beakers. This is known as the **internal circuit**.

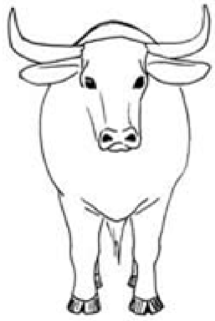


3. In the Zn half cell, the Zn electrode will disintegrate which forms Zn^{2+} ions and releases electrons. Therefore oxidation occurs here.

The half cell that undergoes oxidation is called the **anode**. This is the producer of electrons making it the negative post.

A mnemonic:

'An Ox' Anode = Oxidation

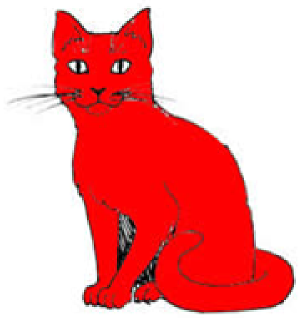


Conversely, the Cu half of the cell the Cu electrode gets Cu deposited on it. This requires electrons. Therefore, reduction is happening here.

The half cell that undergoes reduction is called the **cathode**. This is the positive post of the cell as it consumes electrons.

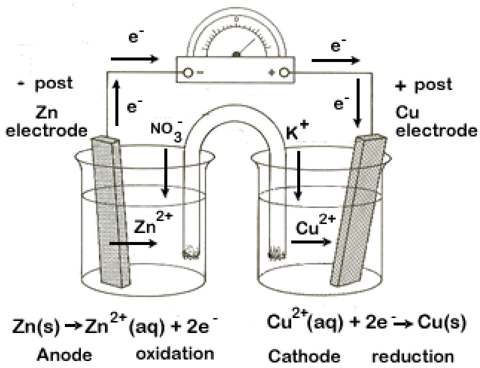
A mnemonic:

'red cat' reduction = cathode



The external circuit is where all the electrical work is done as the electrons flow from the anode to the cathode.

The internal circuit is used to keep each half cell electrically neutral.



If the anode's half cell is producing Zn^{2+} then it will attract the NO_3^- from the salt bridge.

Likewise, if Cu^{2+} is leaving the anode's solution then it will attract the K^+ from the bridge.

Therefore, **electrons flow from the anode to the cathode.**

- Anions flow to the anode side of the cell.
- Cations flow to the cathode side of the cell.

At this point, the cell is complete and electrons can flow allowing the redox reaction to proceed.

When the mass of each electrode is measured before and after the reaction, it is found to be different. The copper electrode will have a greater mass after the experiment and the zinc electrode will have less mass.

Standard Electrode Potentials

Why was Zn oxidized and Cu reduced in the hypothetical electrochemical cell we made last section? The answer is due to **valence electrons.**

Metals, in the first place, only have a few valence electrons which means they like to get rid of them (become oxidized).

However, metals differ in how easily they can lose their electrons. A list of how easily a metal can lose its electrons is known as the **activity series**. *The higher the metal is on the chart, the more likely it is to lose electrons (become oxidized).*

Table 20

Metal Activity Series

Metal	Metal Ion	Reactivity
Lithium	Li+	Most Reactive
Potassium	K+	-
Calcium	Ca2+	-
Sodium	Na+	-
Magnesium	Mg2+	-
Aluminum	Al3+	-
Manganese	Mn2+	-
Zinc	Zn2+	-
Chromium	Cr2+, Cr3+	-
Iron	Fe2+, Fe3+	-
Lead	Pb2+	-
Copper	Cu2+	-
Mercury	Hg2+	-
Silver	Ag+	-
Platinum	Pt2+	-
Gold	Au+, Au3+	Least Reactive

This difference in ability to lose electrons is what drives electrochemical cells and it is the force that allows electrons to flow from the anode to the cathode.

- This force is known as **potential difference** or **electromotive force (emf or E)**.

Potential difference is measured in **volts (V)** and will be referred to as **voltage**. This is a measure of the tendency of electrons to flow through the external circuit.

The higher the voltage the greater the tendency for electrons to flow from the anode to the cathode. That is, more current is produced with a higher voltage.

Calculating Voltages

The voltage for electrochemical cells can be found using a **standard electrode potential for half reaction** table (*Table 18 and 19*). Note that the table may have more than one half reaction for a given element. Therefore, pay close attention to the specific **ion** listed.

The standard reduction potential of a half-cell is a measure of the tendency to GAIN electrons; higher number = greater oxidizing agent.

Chemists have decided to use the hydrogen half-cell as the standard to which all other half-cell potentials will be compared. The hydrogen half-cell potential is written as $E^o = 0.00V$ (E^o = 1 M solutions are standard temp and pressure)

The table shows the voltage for a number of half reactions. Notice that the half reactions listed are all **reduction** reactions.

If we know the substances of two half cells we can use the table to determine which will be the anode (oxidation) and which will be the cathode (reduction). <u>You will need to change the sign of the given E ° value for the oxidation reaction.</u>

When we know this, we can calculate the voltage of the electrochemical cell adn determine if the reaction is spontaneous.

****If the total E^o is positive, your redox reaction is spontaneous in the forward direction.**

****If the total E^o is negative, your redox reaction is spontaneous in the reverse direction.**

Ex: An electrochemical cell is produced from the following two half-cells: (Pb^{2+} | $Pb_{(s)}$) and ($Al_{(s)}$ | Al^{3+})

a. Determine E^o

b. Identify the anode and the cathode

c. State whether the forward reaction will be spontaneous.

7.4 - Introduction to Electrochemistry and Calculating Voltages

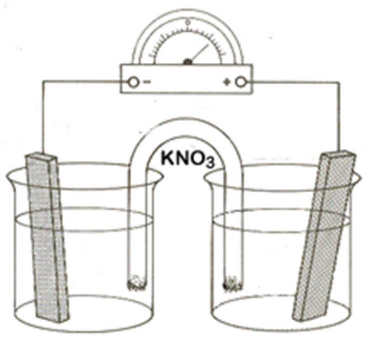
Steps to calculate voltage:

1. Write down the half reactions.
2. Determine which is the anode and which is the cathode. The reaction with the largest voltage will be the cathode since it will be reduced. Write down the voltage associated with the cathode.
3. Write down the voltage associated with the anode. We will need to switch the sign on this value as it is being oxidized instead of reduced.
4. Get the same number of electrons in each half reaction and add the two half reactions together. This allows us to cancel out our electrons and find the overall full redox reaction and the voltage.

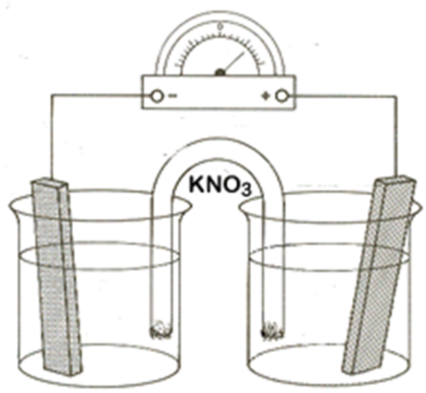
Ex) Calculate the voltage for the following electrochemical cells and label the following items on the diagrams for each:

i) A cathode and the substance it is made of
ii) An anode and the substance it is made of
iii) The correct half reaction under each beaker
iv) The ions coming off or attaching to each electrode
v) The direction of the ions in the salt bridge
vi) The direction of the flow of electrons

a) copper-copper(II) and zinc-zinc(II) half cell.



b) (Al | Al^{3+}) and (Pb|Pb^{2+})



7.4 Electrochemistry Worksheet

1. Calculate the voltage produced for each of the electrochemical cells containing the substances below. On each diagram, label:

   i. A cathode and the substance it is made of

   ii. An anode and the substance it is made of

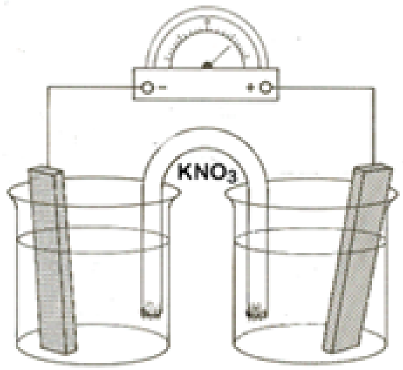
   iii. The correct half reaction under each beaker

   iv. The ions coming off or attaching to each electrode

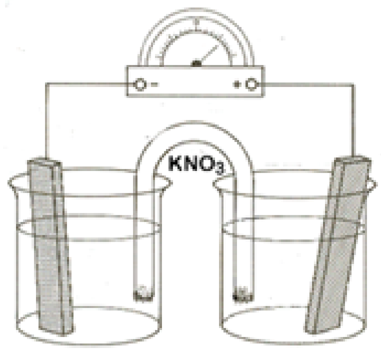
   v. The direction of the ions in the salt bridge

   vi. The direction of the flow of electrons

a) $(Fe \mid Fe^{2+})$ and $(Pb \mid Pb^{2+})$



b) $(Cr \mid Cr^{3+})$ and $(Li \mid Li^{+})$



2. Determine the E° for a (Ag | Ag^+) and (Zn|Zn^{2+}) electrochemical cell

3. Determine E° for a $Fe_{(s)}$ | $Fe^{2+}_{(aq)}$||$Cu_{(s)}$|$Cu^{2+}_{(aq)}$

4. a) An electrochemical cell is created using gold and magnesium half-cells. Write the redox reaction. Determine which half-cell will undergo oxidation and which will undergo reduction, identify which substance is the anode and cathode, and calculate the voltage for the cell. You do not need to draw a diagram of the cell.

b. If the mass of the magnesium electrode changes by 5.0 g, what will be the change in mass of the gold electrode, and will its mass increase or decrease? (hint – use mole to mole stoichiometry using the redox equation found in part a)

7.5 Electrolysis Assignment

1. What is the difference between an anode and a cathode in a electrolytic vs. an electrochemical cell?

	anode	cathode
electrochemical cell	OX −	RED +
electrolytic cell	OX +	RED −

2. Predict the anode and cathode half reactions during the electrolysis of a 1.0M solution of $KI_{(aq)}$. Assume that inert electrodes are used.

Species: K^+, I^-, H_2O

(red) cathode: $K^+ + e^- \rightarrow K_{(s)}$ −2.9

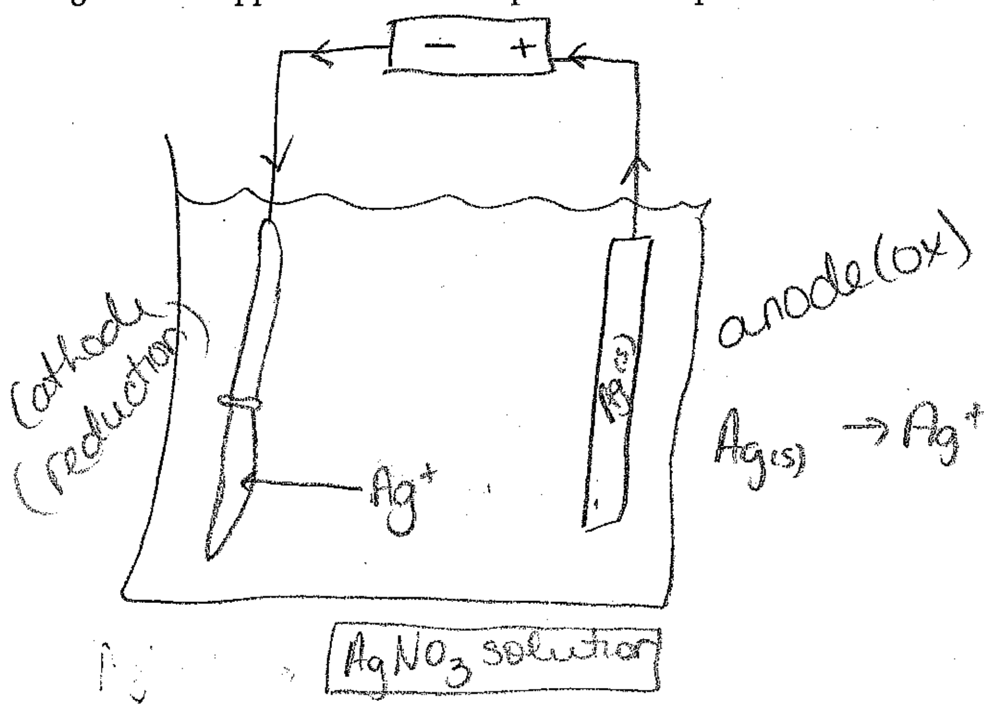
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ (10^{-7}) −0.41 (red)

(ox) anode: $2I^- \rightarrow I_2 + 2e^-$ −0.53 (ox)

$2H_2O \rightarrow \frac{1}{2}O_{2(g)} + 2H^+ (10^{-7}M) + 2e^-$ −0.82

$\boxed{\therefore \text{anode: } 2I^- \rightarrow I_2 + 2e^- \quad \text{cathode: } 2H_2O + 2e^- \rightarrow H_2 + 2OH^- (10^{-7} M)}$

3. Draw a labelled diagram of an apparatus that would plate a nickel-plated knife with silver.



$Ag_{(s)} \rightarrow Ag^+ + 1e^-$

$Ag^+ + 1e^- \rightarrow Ag_{(s)}$

4. Will silver metal in a solution of chloride ions produce silver ions and chlorine gas? Explain.

$Ag_{(s)} \rightarrow Ag^{+} + 1e^{-}$ -0.80

$2Cl^{-} \rightarrow Cl_{2(g)} + 2e^{-}$ $+1.36$

* no b/c 2 oxidation equations *

5. Can a 1M solution of iron (III) sulfate be stored in a container of nickel metal? Explain.

$Ni_{(s)} \rightarrow Ni^{2+} + 2e^{-}$ $+0.23$

$(Fe^{3+} + e^{-} \rightarrow Fe^{2+})2$ $+0.77$

$Ni_{(s)} + 2Fe^{3+} \rightarrow Ni^{2+} + 2Fe^{2+}$ $+0.90$

∴ reaction will occur so cannot be stored.

6. How many grams of hydrogen gas would be produced from the oxidation of 5.00g of magnesium metal?

$Mg_{(s)} \rightarrow Mg^{2+} + 2e^{-}$ $+2.37$

$2H^{+} + 2e^{-} \rightarrow H_{2(g)}$ 0.00

$Mg_{(s)} + 2H^{+} \rightarrow Mg^{2+} + H_{2(g)}$ $+2.37$

$n = \frac{wt}{mm} = \frac{5.0g}{24.305g/mol} = 0.206\ mol\ Mg_{(s)}$ = mols H_2

$wt = n \cdot mm = (0.206mol)(2.02g/mol) = 0.416g\ H_2$

7.5 - Electrolysis

Electrolysis and electrolytic cells make up the second branch of electrochemistry.

We can think of these two things as opposite of electrochemical cells.

Electrolytic cells convert electrical energy into chemical energy (the electricity is needed to cause a non spontaneous reaction to occur)

The E° value is negative instead of positive.

Electrolysis of Molten NaCl

Electrolysis can be used to break up ionic compounds into its component elements.

For example,

$2NaCl_{(aq)} \rightarrow 2Na_{(s)} + Cl_{2\,(g)}$

Note that this is a redox reaction. Therefore, the two half reactions can be used to find the total voltage:

E°

reduction $2Na^{+}_{(aq)} + 2e^{-} \rightarrow Na_{(s)}$ -2.71 V

oxidation $Cl^{-}_{(aq)} \rightarrow Cl_{2\,(g)} + 2\,e^{-}$ -1.36 V

----

net voltage required - 4.07 V

This negative voltage tells us that the overall reaction will **not** be spontaneous and it will take 4.07 V to cause the reaction to occur.

7.5 - Electrolysis

Some key differences between electrolytic and electrochemical set-ups:

1. There is no salt bridge separating the two half reactions.
2. A source of current is **needed**.
3. The anode is positively charged and the cathode is negatively charged, but the anode is still the site of **oxidation** and the cathode is still the site of **reduction**.

Here is a diagram, with explanations, of an electrolytic cell.

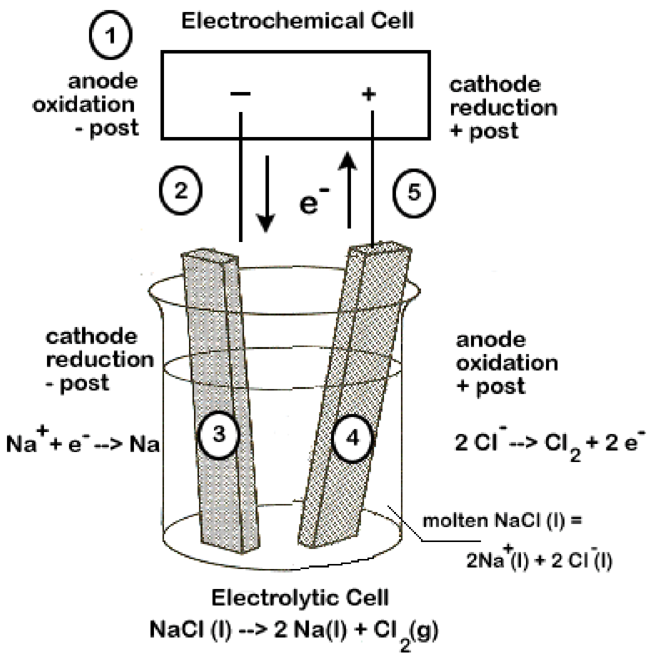
1. Electrons are "produced" in the battery at the anode, the site of oxidation.

2. The electrons leave the electrochemical cell through the external circuit.

3. These negative electrons create a negative electrode in the electrolytic cell which attracts the positive Na^+ ions in the electrolyte. Na^+ ions combine with the free electrons and become reduced ($2Na^+ + 2e^- \rightarrow Na$)

4. Meanwhile the negative Cl^- become attracted to the positive electrode of the electrolytic cell. At this electrode chlorine is oxidized, releasing electrons ($Cl^- \rightarrow Cl_2 + 2\ e^-$

5. These electrons travel through the external circuit, returning to the electrochemical cell.



7.5 - Electrolysis

Electroplating

https://www.youtube.com/watch?v=FnJ0V7B7nKo

In electroplating, we will see that the cathodes not only **just** carry a charge, but they actively participate in the reaction.

Electroplating is when a thin layer of a desired metal is used to coat (or **plate**) another object. The purpose is to protect against corrosion or improve appearance.

For example, forks made from inexpensive metal are often coated with silver.

The requirements for electroplating:

1. An electrolytic solution which contains ions of the plating metal. $AgNO_3$ will produce sufficient Ag^+ ions.
2. A source of current (a battery).
3. Two electrodes. The first is the object we are plating (the fork), while the second must be the plating metal (silver).

Here are the half reactions:

$Ag^+ + e^- \rightarrow Ag$	cathode	reduction
$Ag \rightarrow Ag^+ + e^-$	anode	oxidation

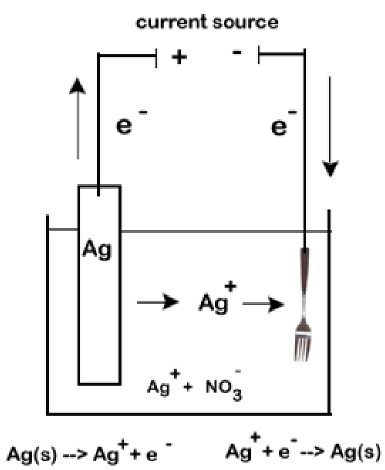
Note - there is only one metal involved.

The Ag^+ that will be deposited on the fork as pure silver once it undergoes reduction will come from the electrolytic solution.

This causes the solution to become negatively charged because there is now more NO_3^- present than Ag^+. The silver bar then undergoes oxidation and replaces the Ag^+ that has been removed.

The flow of electrons goes from the anode of the battery, through the external circuit, and into the cathode of the electrolytic cell.

Electrons are produced by the oxidation of silver in the anode. These flow up through the external circuit into the cathode of the battery.



7.5 - Electrolysis

Ex 1) First, determine the cell reaction for the electrolysis of NaI(aq) using inert electrodes. Then, calculate the minimum voltage that must be applied in order for this reaction to occur.

Ex 2) An iron nail is suspended in a solution of $CuSO_{4\,(aq)}$ by an inert platinum wire. An external power source if connected so the nail becomes the cathode and a copper electrode becomes the anode. If sufficient voltage is applied, what anode and cathode half-reactions occur in this cell?

7.5 Electrolysis Assignment

1. What is the difference between an anode and a cathode in a electrolytic vs. an electrochemical cell?

2. Predict the anode and cathode half reactions during the electrolysis of a 1.0M solution of $KI_{(aq)}$. Assume that inert electrodes are used.

3. Draw a labelled diagram of an apparatus that would plate a nickel-plated knife with silver.

7.5 - Electrolysis

4. Will silver metal in a solution of chloride ions produce silver ions and chlorine gas? Explain.

5. Can a 1M solution of iron (III) sulfate be stored in a container of nickel metal? Explain.

6. How many grams of hydrogen gas would be produced from the oxidation of 5.00g of magnesium metal?

Name: ________________________

Unit 4 Hand-In Assignment #2

This assignment covers sections 4.4-4.5. Make sure you show *all states for products and remember to balance equations*. When using a formula, write down the formula then substitute values with units. **Your answer must have the correct units and significant figures in all of your final answers.**

1. What is the solubility of a saturated solution of:
   a. Ammonium chloride at $80^{o}C$ in g/100mL of water. (1)

   b. NH_3 at $15^{o}C$ in mols/L water (show your work). (2)

2. You mix 20g of potassium nitrate with 100mL of water at $20^{o}C$. Describe the solution in terms of saturation. (1)

3. There is more dissolved oxygen in the ocean the at greater depths when compared to dissolved oxygen at the surface. Using Henry's Law, describe this observation. (2)

4. State whether the following compounds are soluble or insoluble in water. (3)
   a. Sodium hydroxide

   b. Ammonium acetate

   c. Calcium sulfate

   d. Lead (II) chloride

   e. Potassium chloride

   f. Calcium bromide

5. Why are spectator ions removed to form net ionic equations? (1)

6. Write the **balanced chemical equation** for the following reactions. Then change each reaction to ionic form (if you need); finally to **net ionic equations** (each is worth 2). If no reaction occurs, please include the spectator ions.

    a. $FeSO_{4(aq)} + (NH_4)_2S_{(aq)} \rightarrow$

    b. lead (II) nitrate and potassium bromide$\rightarrow$

    c. $KOH_{(aq)} + NaCl_{(aq)} \rightarrow$

    d. $Pb(C_2H_3O_2)_{2(aq)} + K_2SO_{4(aq)} \rightarrow$

7. Devise a procedure for selectively precipitating the following ions from each other in a common solution. Record your table, the order the solutions must be added and the precipitate that forms at each step. (4 each)

    a. Cl^- and PO_4^{3-}

    b. Ca^{2+} and Ag^+

    c. Cu^{2+} and Ca^{2+}

Name:____________________ School:____________________

Chem 30 Unit 7 Hand In Assignment (sections 7.1-7.3)

When balancing the following redox reactions with oxidation numbers, the changes in oxidation numbers must be shown and balancing coefficients placed into the equation for full marks.

When using the half reaction method, you must fully write out both half reactions such that the number of electrons in each is equal. Remember to put correct charges on molecules and ions when appropriate.

1. Balance the following reactions using the oxidation number method.

a. $NaClO + H_2S \rightarrow NaCl + H_2SO_4$

$\overline{}$ 4

b. $Sn + HNO_3 + H_2O \rightarrow H_2SnO_3 + NO$

$\overline{}$ 4

2. Balance the following reaction in an acidic solution using the half reaction method.

$CrO_4^{2-} + Cl^- \rightarrow Cr^{3+} + Cl_2$

$\overline{}$ 4

Name:___________________ School:___________________

3. Balance the following redox reaction in an acidic solution using any method you prefer (oxidation number or half reaction).

$H_2PO_2^- \quad + \quad TeO_4^{2-} \quad \rightarrow \quad PO_4^{3-} \quad + \quad Te$

$\overline{4}$

4. Balance the following redox reaction in a basic solution. Use any method you prefer (oxidation number method or half reaction method).

$Ce^{4+} \quad + \quad I^- \quad \rightarrow \quad Ce^{3+} \quad + \quad IO_3^-$

$\overline{4}$

Name:__________________ School:__________________

Chem 30 Unit 7 Hand In Assignment #2 (7.4-7.5)

1. By what mass (in grams) will a chromium cathode increase when it is coupled to a magnesium half-cell in which the magnesium anode loses 1.53 grams? (assuming chromium ions in solution are Cr^{3+}). 6 marks

2. Determine whether the following reactions will occur by determining the voltage produced. (3 marks each)

a. $Ag(s) + HCl(aq) \rightarrow$

b. $Mg(s) + FeSO_4(aq) \rightarrow$

3. The most common method of producing bromine involves oxidizing bromine ions (Br^-) to bromine liquid (Br_2) using chlorine gas (Cl_2). What is the E for this reaction? 2 marks

Name:___________________ School:___________________

4. What reaction (oxidation or reduction) occurs at an anode of . . .

a. an electrochemical cell (1mark)

b. an electrolytic cell (1 mark)

5. An iron bar is to be electroplated with zinc. Draw a diagram to do so and:
10 marks

- Identify what will act as the two electrodes for the cell
- Identify each electrode as either the anode or cathode
- Write the half-reactions occurring at each electrode
- Identify a solution that would make a suitable electrolyte for this cell
- Identify which electrode will be attached to the negative post of the battery and which will be attached to the positive post, and explain.
- Identify the flow of electrons.