Electrochemistry:

Definition:

The last unit was about the transfer of protons. This unit is about the transfer of electrons.

When the number of electrons in an atom changes:

Therefore, in order to recognise reduction-oxidation reactions (redox reactions):

Oxidation Number: "The charge an atom in a molecule would have if all the electrons in its bond belonged entirely to the more electronegative atoms" OR "A number assigned to an element to indicate its position on a scale of oxidation levels defined by an arbitrary set of rules"

Determining the oxidation number:

The Rules:

1)

2)

3)

4)

	5) Ex: Ex: Try one: **Read pg 193-194 and do #3** **<u>Oxidation Reaction</u>:** * Used to be: * Now is: * Involves:

<u>Reduction Reactions</u>:

* Used to be:

* Now is:

* Involves:

* These two processes happen at the same time. In order for there to be a reduction, there must also be an oxidation and vice versa. One reactant loses an electron and the other gains it.

Ex:

Ex:

Ex:

Try one: (identify all the O.N.'s and say which is oxidized and which is reduced)

Remember that one is oxidized and when it oxidizes, it reduces the other? It is called the:
Remember that one is reduced and when it reduces it oxidizes the other? It is called the:

*For it to be a redox reaction, there must be a gain AND a loss

Assignment: Read pages 193-194 & 189-191 and do questions # 4, 5, & 1

Spontaneity of Red-ox Reactions:

When 2 reactants are mixed, the questions are:
1) will they react? and if so,
2) will the reaction not be spontaneous? or
3) will the reaction be spontaneous

Let's look at our last table

Standard Potential of Half Cells

Some points to consider:
- All reactions are written as reductions

- the ones that reduce strongly (a.k.a. strong oxidizing agents) are at the top left. THEY

- the ones that oxidize strongly (a.k.a. strong reducing agents) are at the bottom right. THEY

Some tips:

- metals are usually at the bottom right
- halogens and oxygen containing things are at the top left
- CAREFUL: some things appear more than once and metals may have more than one charge.

- Reactions can go both ways. Reactions that are near the top of the table will tend to go forward, near the bottom will tend to go backward.

***** be really careful ……… many are on both sides

The Possibilites:

1/ Both reactants are on the same side:

Ex: $F_2 + H_3PO_4 \rightarrow$

Ex: $H_2S + Ag \rightarrow$

2/ Reactants are on opposite sides:

a) the reactant to be reduced (on the left) is lower in the table than the one to be oxidized (on the right)

Ex: $SO_4^{2-} + Br_2 \rightarrow$

b) The one to be reduced is higher than the one to be oxidized

Ex: $MnO_4^- + Ag \rightarrow$

	Summary:
	One more thing: - some things require H^+/H_2O (acidic conditions) - some things require OH^-/H_2O (basic conditions) **Read pages 195-199 and do questions #7, 8, 11, & 12** **<u>Balancing Half-Reactions</u>:** Steps: 1/ Redox? 2/ Split into half reactions 3/ Balance the number of atoms 4/ balance the charge 5/ balance the number of electrons 6/ net reaction 7/ check Ex: $IO_3^- + H^+ + H_2SO_3 + H_2O \rightarrow I_2 + H_2O + SO_4^{-2} + H^+$

	Ex:$ClO_4^- + H^+ + MnO_2 + OH^- \rightarrow MnO_4^- + H_2O + Cl^- + H_2O$ Try this one: Another tip: We need to make sure that when we are doing the half reactions, we include all atoms involved with the atom that has changed.

	Balancing in Acidic Conditions: Steps: 1/ Redox? 2/ Split into half reactions 3/ Balance the number of atoms (add H_2O to side with less Ox) 4/ balance the charge 5/ balance the number of electrons 6/ net reaction 7/ check Ex: $Cr_2O_7^{-2} + H_3PO_3 \rightarrow H_3PO_4 + Cr^{+3}$ Try one: $IO_3^- + N_2O_4 \rightarrow I_2 + NO_3^-$

Balancing in Basic Conditions:

Steps:
1/ Redox?
2/ Split into half reactions
3/ Balance the number of atoms (add OH^- to side with less Ox)
4/ balance the charge
5/ balance the number of electrons
6/ net reaction
7/ check

Ex: $ClO_4^- + N_2O_4 \rightarrow NO_3^- + Cl^-$

First Way:

Second Way:

Try one:

$Bi(OH)_2 + SnO_2^{-2} \rightarrow Bi + SnO_3^{-2}$

- make sure you show each step for these ones.

A Nasty: (on Feb 08's provincial)

Balance in Basic Conditions:

$IPO_4 \rightarrow I_2 + H_2PO_4^- + IO_3^-$
(by the way, this is called a **disproportionation reaction** – when the reactant both reduces and oxidizes)

Assignment: Read pages 201-207 and do questions 24 a, f, k, n, u

Balancing Using The Whole O.N. Method:

Ex:

$$S + HNO_3 \rightarrow SO_2 + NO + H_2O$$

Read pages 208-209 and do questions #25 a, e, & m

<u>Redox Titrations</u>:

* the slow reaction of a reducing agent and an oxidizing agent.

* Equivalence is viewed (endpoint) via an indicator OR as a huge change in voltage. A redox indicator changes colour when it goes from its oxidized to its reduced form.

Ex: ferroin

Titration Curve would look much the same as for an acid base reaction.

Picking an oxidizing agent:

* You want to pick one that can oxidize many things:

Ex:

* another reason to pick MnO_4^-:

Picking a Reducing agent:

* I^- is used a lot because it can be oxidized by many things and we can use it (indirectly) as an indicator.

Two Steps to this titration:

1/ I^- is oxidized to I_2 by another reactant
2/ I_2 is reduced back to I^- with a second reactant to get a colour change.

It's called a back titration and you really don't have to know anything about it. Just read that section in the text.

	Read Pages 210-212