Reduction & Oxidation Reactions

In the reduction/oxidation (red/ox), electrons are transferred from one atom to another atom.

-The atom that accepted the electrons is reduced, it is also called the oxidizing agent.
$Fe^{+3} + e^{-} \rightarrow Fe^{+2}$ (Fe^{+3} was reduced, it is also called the oxidizing agent)

-The atom that denoted the electrons is oxidized, it is also the reducing agent.
$Cr^{+3} - 3e^{-} \rightarrow Cr^{+6}$ (Cr^{+6} was oxidized, it is called the reducing agent)

The vocabulary can get confusing, as in the example Fe^{+3} was reduced to Fe^{+2} and it is the oxidizing agent. A good place to start in red/ox reaction is to look at the change in the charges on the atoms. The atom that has a charge that goes down in numerical value is the atom that is reduced. The other atom that changed is the one that is oxidized.

Steps to solve Red/Ox Problems

Assign oxidation numbers to each element
-Hydrogen (H) = +1
-Oxygen (O) = -2

-Elements that have a charge, have the same oxidation number (Ex. Cr^{+3} = +3)

-Oxidation number by group

	Group IA	Group IIA	Group IIIA	Group VA	Group VIA	Group VIIA
	+1	+2	+3	-3	-2	-1
ex:	Na	Mg	Al	N	S	Cl

-Any other element that is not bonded to another has the value 0. (ex. e, there is no charge therefore the oxidation number is 0.)

-For all the other elements not numbered yet, use the oxidation numbers on the elements bonded to it and the total charge on the compound to solve for the charge on the element.
A(x) + B(y) = total charge of the compound

EXAMPLES

What is the oxidation number on iron in FeO?

Oxidation number on Oxygen (O) is -2
1(Fe) + 1(-2) = 0
Fe + -2 = 0
Fe = +2
Oxidation number on Fe is +2

ANOTHER EXAMPLE

What is the oxidation number of Cr in the compound $(Cr_2O_7)^{-2}$.

The Oxygen (O) has an oxidation number of -2 and there are seven Oxygens, which add up to a -14, with two chromiums (Cr) and -14 totaling the charge of -2, the oxidation number on chromium must be a +6.

$$2(Cr) + 7(-2) = -2$$
$$2Cr + -14 = -2$$
$$2Cr = +12$$
$$Cr = +6$$

Working Red/Ox Reactions

1) Assign oxidation numbers to each element.

2) Find the element that is reduced (the oxidation number went down), find the element that is oxidized.

3) Write down the half-reactions for the reduced and the oxidized elements.

4) Balance the half reactions
   a) Balance the reduction or oxidation **atoms**.
   b) Balance the oxygen (O) by **adding water H_2O.**
   c) Balance the hydrogen (H) by adding **H^+**.
   d) Balance the charge by **adding electrons.**

5) Multiply the half reactions by any number(2) that will result in the same number of electrons in each half reaction.

6) Combine the half reactions and cancel items that are on both sides of the equation.

The rest of this is not required for balancing all of the red/ox reactions

This is now balanced for an acidic solution, if the question requires the solution to be basic, continue through the next steps.

7) Determine the amount of ^-OH needed to neutralize the H^+ that is on one side of the equation. Add this amount of **^-OH to both sides** of the equation.

8) The one side that has H^+ and the new ^-OH will combine these two to form (H-OH) H_2O.

9) Cancel this amount of water found on both sides of the equation.

The equation is now balanced for a basic solution.