BALANCING REDOX REACTIONS

A Modified Half-Reaction Method

1. Assign oxidation numbers to all atoms in the equation. Determine which element is oxidized and which is reduced.

2. Write two *un*balanced half-equations--one for oxidation, the other for reduction.

3. Separately, for each of the half-reactions, complete the following sequence of steps:

   a. Balance all the elements in the half reaction *except hydrogen and oxygen*; [**EXCEPTION:** if either hydrogen or oxygen is being oxidized or reduced, the element involved must be balanced at this time.]

   b. Add the correct number of electrons (e^-) to reflect the number of electrons involved in the oxidation or reduction. To find this number, multiply (the change in oxidation number) **X** (the number of atoms undergoing that change). For *oxidation*, the electrons are added to the right side of the equation (electrons are "lost", so they are a *product* of the reaction); for *reduction*, they are added to the left (electrons are "gained", so they are a *reactant*).

   c. Balance the net charge by adding H^+ if the solution is acidic, or OH^- if basic.

   d. Balance the mass [H and O atoms] by adding the appropriate number of H_2O.

4. CHECK: Both half-reactions should be separately balanced (atoms and charges).

5. Multiply each of the half-reaction equations by an appropriate number, so that the number of electrons in each of the equations is identical.

6. Add the two half-reaction equations. The electrons *must* cancel! If H^+, OH^- and/or H_2O appear on both sides of the equation, cancel them algebraically. Reduce coefficients, if appropriate, to give simplest whole number ratios.

7. CHECK YOUR WORK. Both mass (atoms) and net charge must balance. Electrons will never appear in the net reaction equation.

Revised: Spring 2004
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