Balancing Redox Reactions by the Ion-Electron Method

1. Write the skeletal equation containing the oxidizing and reducing agents and the products in ionic form. Assign oxidation numbers to each atom.

2. Separate the equation into two half-reactions.

3. Balance the atoms (except the O and H) in each half reaction separately. This is called *mass balance*.

4a. For reactions in *acidic* medium, add H_2O to balance the oxygen atoms and H^+ to balance the hydrogen atoms.

4b. For reactions in *basic* medium, first balance the atoms as you would for an acidic solution. Then, for each H^+ ion, add OH^- ion to both sides of the half-reaction. Whenever H^+ and OH^- appear on the same side, combine them to make H_2O.

5. Add electrons to one side of each half-reaction to equalize the charges. Electrons are added to the reactant side of a reduction half-reaction and to the product side of an oxidation half-reaction. The number of electrons added to the one side of the half-reaction should make the total charge of that side equal to the charge on the other side. This procedure is called a *charge balance*.

6. Multiply the half-reactions through by the appropriate coefficients to balance the number of electrons as needed.

7. Add the two half-reactions, and check the final equation by inspection. Reduce the balanced equation to lowest terms (coefficients).

Balancing Redox reactions

Balance the following redox reactions:

1. $Sn + NO_3^- \rightarrow SnO_2 + NO_2$ (acidic solution)

2. $BrO_3^- + I^- \rightarrow I_2 + Br^-$ (acidic solution)

3. $H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + H_2O$ (acidic solution)

4. $MnO_2 + BiO_3^- \rightarrow MnO_4^{2-} + Bi^{3+}$ (basic solution)

5. $Cl_2 \rightarrow ClO_4^- + Cl^-$ (basic solution)

6. $P_4 \rightarrow PH_3 + HPO_3^{2-}$ (basic solution)

7. $Sb + NO_3^- \rightarrow Sb_2O_5 + NO$ (acidic solution)