Balancing Equations by the Ion Electron Method

(A.k.a. Half – Reaction Method)

Balancing equations using oxidation numbers is good but it often does not give a good understanding of what is going on in the reaction. Redox reactions can be split into two half reactions – an oxidation ½ reaction, and a reduction ½ reaction; based on electron loss and gain. In a redox reaction the number of electrons "gained" must equal the number of electrons "lost." This means that the number of electrons in two ion– electron equations must be made equal before the equations can be added to give an overall redox equation.

In one ½ reaction oxidation occurs – electrons are produced
In the other ½ reaction reduction occurs – electrons are consumed.

Example: $Cu_{(s)} + 2\ Ag(NO_3)_{2\ (aq)} \longrightarrow Cu\ (NO_3)_{2\ (aq)} + 2\ Ag_{(s)}$

What half reactions are happening here?
Always remove the spectator ions when splitting this up into half reactions.

Oxidation ½ reaction:

Reduction ½ reaction:

Balancing Half Reactions In Acidic Solution

1. Write the skeleton equation.
2. Balance for species other than oxygen and hydrogen.
3. Balance for **oxygen using one water molecule** for each oxygen you require.
4. Balance for **hydrogen using a hydrogen ion** for each hydrogen you require.
5. Balance for **charge by adding electrons** to either the product or reactant side.

Examples: Balance the following:

1. $ClO_4^- \longrightarrow Cl_2$
2. $NO_3^- \longrightarrow HNO_2$
3. $MnO_4^- \longrightarrow Mn^{+2}$
4. $Cr_2O_7^{-2} \longrightarrow Cr^{+3}$
5. $NO_3^- \longrightarrow NH_3$
6. $ClO_3^- \longrightarrow Cl^-$
7. $H_2S \longrightarrow S$
8. $NO_3^- \longrightarrow NO$

Balancing The Whole Equation

To balance the whole equation you must do the following:

1. Separate the whole equation into its two ½ reactions and balance each separately.
2. Multiply each ½ reaction by a suitable factor so that the electrons produced equals the electrons consumed, (i.e. electron gained equals electron lost).
3. Add the two ½ reactions together and cancel any common species.

Examples

1. Purple MnO_4^- ions and Fe^{+2} ion in acid solution will produce pink Mn^{+2} and Fe^{+3} ion:

2. Bismuthate, BiO_3^-, is a strong oxidizing agent and can oxidize Mn^{+2} to MnO_4^- in acidic solution. The BiO_3^- ion is reduced to Bi^{+3}:

3. Hydrogen sulphide, H_2S in aqueous solution is a reducing agent , being oxidized to sulphur, whilst the nitrate ion, NO_3^-, in acidic solution is an oxidizing agent, being reduced to NO :

4. In acidic solution the hypochlorite ion, ClO^- disproportionates,

(fyi: a disproportionation reaction is in which an element in one oxidation state is both oxidized and reduced), into the chloride ion, Cl^- and the chlorate, ClO_3^- :

Balancing Equations by the Half – Reaction Method in Basic Solutions

To balance equations by the ½ reaction method for reactions occurring in basic solutions you do the following:

1. Balance the equation pretending it was in acid solution.
2. When it is balanced add hydroxide ions to both sides of the equation equivalent in number to the number of hydrogen ions. The OH^- ions will combine with the H^+ ions to form water.
3. Cancel the appropriate number of waters.

Examples

Balance the following in **basic** solution:

1. $SO_3^{-2} + MnO_4^- \longrightarrow SO_4^{-2} + MnO_2$

2. $MnO_4^- + C_2O_4^{-2} \longrightarrow MnO_2 + CO_3^{-2}$

3. $S \longrightarrow S^{-2} + SO_3^{-2}$

4. $MnO_4^- + IO_3^- \longrightarrow MnO_2 + IO_4^-$

5. $Cu + Br_2 \longrightarrow Cu_2O + Br^-$

Using all the above equation in acid medium, but now try balancing them in basic medium:

Balancing Equation in Acid and Base Medium

Balance the following oxidation-reduction equations. All reactions take place in an acidic solution unless otherwise indicated.

1. $AsH_3 + ClO_3^- \rightarrow H_3AsO_4 + Cl^-$
2. $HNO_2 + I^- \rightarrow NO_{(g)} + I_{2(g)}$
3. $MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_{2(g)}$
4. $MnO_2 + ClO_3^- \rightarrow MnO_4^{2-} + Cl^-$ (basic solution)
5. $Br_2 \rightarrow Br^- + BrO_3^-$ (basic solution)
6. $N_2O_4 + Br^- \rightarrow NO_2^- + BrO_3^-$ (basic solution)
7. $H_2PO_2^- + SbO_2^- \rightarrow HPO_3^{2-} + Sb$ (basic solution)
8. $CrO_2^- + ClO^- \rightarrow CrO_4^{2-} + Cl^-$ (basic solution)
9. $Cu(OH)_2 + HPO_3^{2-} \rightarrow Cu_2O + PO_4^{3-}$ (basic solution)
10. $HS^- + IO_3^- \rightarrow I^- + S$ (basic solution)
11. $N_2O + ClO^- \rightarrow Cl^- + NO_2^-$ (basic solution)
12. $H_2SO_3 + MnO_2 \rightarrow SO_4^{2-} + Mn^{2+}$
13. $IO_4^- + I^- \rightarrow I_{2(g)}$
14. $CrO_4^{2-} + I^- \rightarrow Cr^{3+} + I_{2(g)}$
15. $Cr^{2+} + O_2 \rightarrow Cr^{3+} + H_2O$
16. $H_3PO_3 + NO_3^- \rightarrow PO_4^{3-} + N_2O_{4(g)}$
17. $Cr_2O_7^{2-} + HNO_2 \rightarrow Cr^{3+} + NO_3^-$
18. $Sb_2O_5 + I^- \rightarrow Sb^{3+} + I_{2(g)}$
19. $H_2SO_3 + IO_3^- \rightarrow SO_4^{2-} + I^-$
20. $NO_3^- + SO_2 \rightarrow N_2O_{3(g)} + SO_4^{2-}$
21. $SbO^+ + HClO \rightarrow Sb_2O_5 + Cl^-$
22. $NO_3^- + H_2S \rightarrow NO_{(g)} + S$
23. $TeO_2 + BrO_3^- \rightarrow H_6TeO_6 + Br_{2(g)}$
24. $I^- + HClO_2 \rightarrow IO_3^- + Cl_{2(g)}$
25. $Bi_2O_3 + NO_3^- \rightarrow Bi^{3+} + NO_{(g)}$
26. $S + HNO_2 \rightarrow H_2SO_3 + N_2O_{(g)}$
27. $NO + H_5IO_6 \rightarrow NO_3^- + IO_3^-$