Oxidation-Reduction Reactions

A simple transfer of electrons

Oxidation

- Original definition
  - Adding oxygen to something
- Later definition
  - Removing Hydrogen
- Current definition
  - Loss of Electrons

Reduction

- Original definition
  - Removing oxygen from something
- Later definition
  - Adding Hydrogen
- Current definition
  - Gain of Electrons

Oxidation-Reduction reactions (redox)

- Balancing some redox reactions is relatively simple
  - Combustion reactions are an example
- Others are more difficult and we must have a way of keeping track of where the electrons are going to/coming from.

Oxidation Numbers

- Oxidation numbers are hypothetical charges on atoms.
- Oxidation numbers are assigned to atoms following a set of rules

Rules for assigning oxidation numbers

- Ox. No. of free elements is zero.
- Ox. No. of monatomic ions is equal to the charge.
- Ox. No. of F in compounds is always -1.
- Ox. No. of O in compounds is usually -2 except in peroxides (-1).
- Ox. No of H is usually +1 except in hydrides (-1).
- Sum of all Ox. Nos. in compounds must add to the overall charge.

Balancing Redox reactions

- Two conditions
  - Acid
  - Base
- Balancing redox reactions requires knowledge of the conditions because the rules for each are slightly different.

Balancing redox reactions in acid

- Write the skeleton equation and assign ox. nos. to all elements.
- Split the reaction into two half-reactions:
  - One that contains all species being reduced
  - One that contains all species being oxidized
- Balance all atoms in each half-reaction except for H and O
- Balance O by adding water to side needing O.
- Balance H by adding H^+ to side needing H.

Balancing redox reactions in acid

- Balance charge by adding e^- to side with more positive charge.
- Balance e^- by multiplying each half-reaction by the number of electrons in the other half-reaction.
- Add the half-reactions together and cancel similar terms.
- Check to make sure atoms and charges are balanced.

Example

- Balance
  $Nb + OsO_4 \rightarrow Nb_2O_5 + Os$ (in acid)
- Assign Oxidation Numbers
  - $Nb + OsO_4 \rightarrow Nb_2O_5 + Os$
    0 +8 -2 +5 -2 0

Example Continued

- Split into half reactions
  - $Nb \rightarrow Nb_2O_5$
  - $OsO_4 \rightarrow Os$
- Balance atoms other than H and O
  - $2\ Nb \rightarrow Nb_2O_5$
  - $OsO_4 \rightarrow Os$

Example Continued

- Balance O by adding water
  - $2\ Nb + 5\ H_2O \rightarrow Nb_2O_5$
  - $OsO_4 \rightarrow Os + 4\ H_2O$
- Balance H by adding H^+
  - $2\ Nb + 5\ H_2O \rightarrow Nb_2O_5 + 10\ H^+$
  - $8\ H^+ + OsO_4 \rightarrow Os + 4\ H_2O$

Example Continued

- Balance charge by adding e^-
  - $2\ Nb + 5\ H_2O \rightarrow Nb_2O_5 + 10\ H^+ + 10\ e^-$
  - $8\ e^- + 8\ H^+ + OsO_4 \rightarrow Os + 4\ H_2O$
- Balance charge between half-reactions
  - $4\ (2\ Nb + 5\ H_2O \rightarrow Nb_2O_5 + 10\ H^+ + 10\ e^-)$
  - $5\ (8\ e^- + 8\ H^+ + OsO_4 \rightarrow Os + 4\ H_2O)$

Example Continued

- Combine half-reactions
  - $4\ (2\ Nb + 5\ H_2O \rightarrow Nb_2O_5 + 10\ H^+ + 10\ e^-)$
  - $5\ (8\ e^- + 8\ H^+ + OsO_4 \rightarrow Os + 4\ H_2O)$
  - $40\ e^- + 40\ H^+ + 8\ Nb + 20\ H_2O + 5\ OsO_4 \rightarrow 4\ Nb_2O_5 + 5\ Os + 40\ H^+ + 40\ e^- + 20\ H_2O$

Example Continued

- Cancel out like terms
  - $\cancel{40\ e^-} + \cancel{40\ H^+} + 8\ Nb + \cancel{20\ H_2O} + 5\ OsO_4 \rightarrow 4\ Nb_2O_5 + 5\ Os + \cancel{40\ H^+} + \cancel{40\ e^-} + \cancel{20\ H_2O}$
  - $8\ Nb + 5\ OsO_4 \rightarrow 4\ Nb_2O_5 + 5\ Os$
- Check for balance

Balancing redox reactions in base

- Write the skeleton equation and assign ox. nos. to all elements.
- Split the reaction into two half-reactions:
  - One that contains all species being reduced
  - One that contains all species being oxidized
- Balance all atoms in each half-reaction except for H and O
- Balance O by adding water to side needing O.
- Balance H by adding H_2O to side needing H and OH^- to the other side at the same time.

Balancing redox reactions in acid

- Balance charge by adding e^- to side with more positive charge.
- Balance e^- by multiplying each half-reaction by the number of electrons in the other half-reaction.
- Add the half-reactions together and cancel similar terms.
- Check to make sure atoms and charges are balanced.

Deriving Molecular Equations

- The chemical equation is sometimes more useful than the net ionic equation.
- To get the chemical equation from the net ionic we need to add back in the spectator ions.
- We need other information also, such as what acid or base is used.

How to work…

- Remember the previous problem…
  - $I^- + BrO_3^-$ $Br^- + I_2$ (acid)
- We need other information
  - What are the spectator ions?
  - What is the acid?
- Suppose we had instead…
  - $KI + KBrO_3$ $Br^- + I_2$ (in HI)

$KI + KBrO_3$ $Br^{1-} + I_2$ (in HI)

- First remove spectator ions…
  - In this case we can see that only K^+ are spectator ions.
  - $I^- + BrO_3^-$ $Br^{1-} + I_2$ (in HI)
- Balance as usual
  - $6H^+ + 6\ I^- + BrO_3^-$ $Br^{1-} + 3\ I_2 + 3\ H_2O$
- Add the spectator ions back in
  - $6H^+ + 6\ I^- + 6\ K^+ + 6\ I^- + K^+ + BrO_3^-$ $Br^{1-} + 3\ I_2 + 3\ H_2O + K^+ + 6\ I^- + 6\ K^+$

Continued

- Combine the ions into compounds
  - $6H^+ + 6\ I^- + 6\ K^+ + 6\ I^- + K^+ + BrO_3^-$ $Br^{1-} + 3\ I_2 + 3\ H_2O + K^+ + 6\ I^- + 6\ K^+$
  - $6\ HI + KBrO_3$ $KBr + 3\ I_2 + 3\ H_2O$