Redox reactions – half equations to overall equations and overall to half
Lesson 3
Revise writing balanced half equations

Every redox reaction consists of an oxidant and reductant pair that form their respective conjugates. Every time an oxidising agent gains electrons, it forms an reducing agent that could give electrons if the reaction were reversed, the same applies to reducing agents.

1. Write the balanced half equations for the following overall reactions taking place in an acidic solution.

   a. $Cr(OH)_3 + Br_2 \rightarrow CrO_4^{2-} + Br^-$

   Although states are important lets simplify things for the moment and ignore them.

   Step 1 identify oxidant and reductant reactants and their conjugates using oxidation numbers. An oxidant has its oxidation number reduced while a reductant increases in oxidation number.

   Oxidant $Br_2 \rightarrow Br^-$ ***---- > Br goes from 0 to -1***

   Reductant $Cr(OH)_3 \rightarrow CrO_4^{2-}$ ***----- > Cr goes from +3 to +6***

   Step 2 Balance each half equation

   Oxidant $2e + Br_2 \rightarrow 2Br^-$

   Reductant $H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+ + 3e$

   Write balanced reduction and oxidation half equations for the reactions below.

   b. $O_2 + Sb \rightarrow H_2O_2 + SbO_2^-$

   c. $HCOOH + MnO_4^- \rightarrow CO_2 + Mn^{2+}$

   d. $ClO_2^- \rightarrow ClO_2 + Cl^-$

Write the balanced equation for the following reactions by first writing the oxidation and reduction half reactions and using these to write the overall reaction equation that occurs in an acidic environment except where stated.

Example

$Cr(OH)_3 + Br_2 \rightarrow CrO_4^{2-} + Br^-$

Step 1 identify oxidant and reductant reactants and their conjugates using oxidation numbers. An oxidant has its oxidation number reduced while a reductant increases in oxidation number.

Oxidant $Br_2 \rightarrow Br^-$ ***---- > Br goes from 0 to -1***

Reductant $Cr(OH)_3 \rightarrow CrO_4^{2-}$ ***----- > Cr goes from +3 to +6***

Step 2 Balance each half equation

Oxidant $2e + Br_2 \rightarrow 2Br^-$

Reductant $H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+ + 3e$

Step 3 eliminate the electrons by multiplying each equation so that there are 6 electrons in each.

Oxidant $(2e + Br_2 \rightarrow 2Br^-) \times 3$

$\Rightarrow 6e + 3Br_2 \rightarrow 6Br^-$

Reductant $(H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+ + 3e) \times 2$

$\Rightarrow 2H_2O + 2Cr(OH)_3 \rightarrow 2CrO_4^{2-} + 10H^+ + 6e$

Step 4 add the two equations and eliminate the electrons and cancel any other species that appears on both sides, such as H⁺ or H₂O.=>

$2H_2O + 2Cr(OH)_3 + 3Br_2 \rightarrow 2CrO_4^{2-} + 10H^+ + 6Br^-$

Balance the equations below.

a) $NiO_2 + H_2O + Fe \rightarrow Ni(OH)_2 + Fe(OH)_2$

b) $CO_2 + NH_2OH \rightarrow CO + N_2 + 3\ H_2O$

c) $H^+ + H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + H_2O$

d) $H^+ + H_2O + MnO_4^- + SO_2 \rightarrow Mn^{2+} + HSO_4^-$

e) $ClO_2 + OH^- \rightarrow ClO_2^- + ClO_3^- + H_2O$ (in an alkaline solution) Click for more information.

f) $Cr_2O_7^{-2} + H^+ + SO_2 \rightarrow Cr^{+3} + H_2O + SO_4^{-2}$