Redox reactions – half equations to overall equations and overall to half

Lesson 3

Refresh your knowledge of half-equations by visiting

www.dynamicscience.com.au/tester/solutions1/chemistry/redox/buildinghalfreactions.htm

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.

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

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

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

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

reduction $2e + O_2 + 2H^+ => H_2O_2$

oxidation $Sb + 2H_2O => SbO_2^- + 4H^+ + 3e$

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

oxidation $HCOOH => CO_2 + 2H^+ + 2e$

reduction $5e + 8H^+ + MnO_4^- => Mn^{2+} + 4H_2O$

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

oxidation $ClO_2^- => ClO_2 + e$

reduction $4e + 4H^+ + ClO_2^- => Cl^- + 2H_2O$

Write the balanced equation for the following reactions by first writing the oxidation and reduction half reactions and using these write the overall reaction equation.

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

Step 1 identify oxidant and reductant reactants and their conjugates using oxidation numbers. Ignore the H_2O, H^+ and OH^- present. An oxidant has its oxidation number reduced while a reductant increases in oxidation number.

Oxidant $NiO_2 \rightarrow Ni(OH)_2$ ***---- > Ni goes from +4 to +2***

Reductant $Fe \rightarrow Fe(OH)_2$ ***----- > Fe goes from 0 to +2***

Step 2 Balance each half equation

Oxidant $2e + 2H^+ + NiO_2 \rightarrow Ni(OH)_2$

Reductant $Fe + 2H_2O \rightarrow Fe(OH)_2 + 2H^+ + 2e$

Step 3 Add the two equations by first eliminating the electrons and cancel any H^+ or H_2O that appears on both sides of the equation.

$Fe + 2H_2O + NiO_2 \rightarrow Fe(OH)_2 + Ni(OH)_2$

a. $CO_2 + NH_2OH \rightarrow CO + N_2 + 3\ H_2O$

reduction	$2e + 2H^+ + CO_2 \Rightarrow CO + H_2O$
oxidation	$2NH_2OH \Rightarrow N_2 + 2H_2O + 2H^+ + 2e$
Overall	$CO_2 + 2NH_2OH \rightarrow CO + N_2 + 3\ H_2O$

b. $H^+ + H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + 2\ H_2O$

reduction	$2e + 2H^+ + H_2O_2 \Rightarrow 2H_2O$
oxidation	$Fe^{2+} \Rightarrow Fe^{3+} + e$
Overall	$2Fe^{2+} + 2H^+ + H_2O_2 \Rightarrow 2H_2O + 2Fe^{3+}$

c. $H^+ + H_2O + MnO_4^- + SO_2 \rightarrow Mn^{2+} + HSO_4^-$

reduction	$5e + 8H^+ + MnO_4^- \Rightarrow Mn^{2+} + 4H_2O$
oxidation	$2H_2O + SO_2 \Rightarrow HSO_4^- + 3H^+ + 2e$
Overall	$2H_2O + H^+ + 2MnO_4^- + 5SO_2 \Rightarrow 2Mn^{2+} + 5HSO_4^-$

d. $ClO_2 + OH^- \rightarrow ClO_2^- + ClO_3^- + H_2O$ (in an alkaline solution)

reduction	$e + ClO_2 \Rightarrow ClO_2^-$
oxidation	$H_2O + ClO_2 \Rightarrow ClO_3^- + 2H^+ + e$
Overall	$2ClO_2 + 2OH^- \rightarrow ClO_2^- + ClO_3^- + H_2O$

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

reduction	$6e + 14H^+ + Cr_2O_7^{-2} \Rightarrow 2Cr^{3+} + 7H_2O$
oxidation	$2H_2O + SO_2 \Rightarrow SO_4^{-2} + 4H^+ + 2e$
Overall	$Cr_2O_7^{-2} + 2H^+ + 3SO_2 \Rightarrow 2Cr^{+3} + H_2O + 3SO_4^{-2}$