Balancing Redox reactions in an acid or a base

Redox reactions in acidic solutions

- I will tell you if it is in an acidic solution.
- These have special rules.
- Separate the reaction into half reactions.
- Balance all elements except hydrogen and oxygen.
- Balance oxygen by adding H_2O (which is always prevalent in an acidic solution)
- Balance hydrogen by adding H^+.
- Then balance the charge by adding electrons to whichever side is more positive.
- Recombine your two half equations.

Example

- In an acidic solution
- $Cr_2O_7^{2-} + Cl^- \rightarrow Cr^{3+} + Cl_2$
- Half reactions
- $Cr_2O_7^{2-} \rightarrow Cr^{3+}$
- $Cl^- \rightarrow Cl_2$

Here we go

- $Cr_2O_7^{2-} \rightarrow Cr^{3+}$
- $Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$
- $Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$
- $Cr_2O_7^{2-} + 14\ H^+ \rightarrow 2\ Cr^{3+} + 7\ H_2O$
- $Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Other side

- $Cl^- \rightarrow Cl_2$
- $2\ Cl^- \rightarrow Cl_2$
- $2\ Cl^- \rightarrow Cl_2 + 2\ e^-$
- I have to equal 6 e^- so multiply by 3
- $6\ Cl^- \rightarrow 3\ Cl_2 + 6\ e^-$

Combine my half reactions

- $Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$
- $6\ Cl^- \rightarrow 3\ Cl_2 + 6\ e^-$
- And you get
- $Cr_2O_7^{2-} + 14\ H^+ + 6\ Cl^- \rightarrow 2\ Cr^{3+} + 3\ Cl_2 + 7\ H_2O$
- The electrons cancel out .

Balance in an acidic solution

- $NO_2 + ClO_3^- \rightarrow NO_3^- + Cl_2$

Balance in an acidic solution

- $NO_2 + ClO_3^- \rightarrow NO_3^- + Cl_2$

- Half reactions
- $NO_2 \rightarrow NO_3^-$
- $ClO_3^- \rightarrow Cl_2$

Nitrate side

- $NO_2 \rightarrow NO_3^-$
- $NO_2 + H_2O \rightarrow NO_3^-$
- $NO_2 + H_2O \rightarrow 2\ H^+ + NO_3^-$
- $NO_2 + H_2O \rightarrow 2\ H^+ + NO_3^- + 1\ e^-$

Chlorate

- $ClO_3^- \rightarrow Cl_2$
- $2\ ClO_3^- \rightarrow Cl_2$
- $2\ ClO_3^- \rightarrow Cl_2 + 6\ H_2O$
- $2\ ClO_3^- + 12\ H^+ \rightarrow Cl_2 + 6\ H_2O$
- $2\ ClO_3^- + 12\ H^+ + 10\ e^- \rightarrow Cl_2 + 6\ H_2O$

- You will have 10 x the first reaction

- $10\ NO_2 + 10\ H_2O \rightarrow 20\ H^+ + 10\ NO_3^- + 10\ e^-$

Put them together

- $2\ ClO_3^- + 12\ H^+ + 10\ NO_2 + 10\ H_2O$
- $\rightarrow Cl_2 + 6\ H_2O + 20\ H^+ + 10\ NO_3^-$

- Notice the H^+ and the water can also cancel out
- $2\ ClO_3^- + 10\ NO_2 + 4\ H_2O \rightarrow Cl_2 + 8\ H^+ + 10\ NO_3^-$

Example

- In an acidic solution
- $MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2$

Example

- In an acidic solution
- $MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2$

- Half reactions
- $MnO_4^- \rightarrow Mn^{2+}$
- $H_2O_2 \rightarrow O_2$

Top Equation

- $MnO_4^- \rightarrow Mn^{2+}$
- $MnO_4^- \rightarrow Mn^{2+} + 4\ H_2O$
- $MnO_4^- + 8\ H^+ \rightarrow Mn^{2+} + 4\ H_2O$
- $MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

Bottom Equation

- $H_2O_2 \rightarrow O_2$
- $H_2O_2 \rightarrow O_2 + 2\ H^+$
- $H_2O_2 \rightarrow O_2 + 2\ H^+ + 2\ e^-$
- I need to equal 5 e^- so…
- That won't work…
- $2MnO_4^- + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$
- $5\ H_2O_2 \rightarrow 5\ O_2 + 10\ H^+ + 10\ e^-$

Add them together

- $2MnO_4^- + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$
- $5\ H_2O_2 \rightarrow 5\ O_2 + 10\ H^+ + 10\ e^-$
- And you get
- $2\ MnO_4^- + 6\ H^+ + 5\ H_2O_2 \rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O$
- Notice the H^+ canceled out as well.

Balancing Redox Equations in a basic solution

- Follow all rules for an acidic solution.
- After you have completed the acidic reaction add OH^- to each side to neutralize any H^+.
- Combine OH^- and H^+ to make H_2O.
- Cancel out any extra waters from both sides of the equation.

Example

- We will use the same equation as before
- In a basic solution
- $MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2$

- Balanced in an acidic solution
- $2\ MnO_4^- + 6\ H^+ + 5\ H_2O_2 \rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O$

Basic solution

- Since this is a basic solution we can't have excess H^+.
- We will add OH^- to each side to neutralize all H^+
- $2\ MnO_4^- + 6\ H^+ + 5\ H_2O_2 + 6OH^-$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O + 6OH^-$
- We added 6 OH^- because there were $6H^+$

Cont.

- $H^+ + OH^- \rightarrow H_2O$
- Combine the hydroxide and hydrogen on the reactant side to make water
- $2\ MnO_4^- + 6\ H_2O + 5\ H_2O_2$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O + 6OH^-$
- Cancel out waters on both sides
- $2\ MnO_4^- + 5\ H_2O_2$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 2\ H_2O + 6OH^-$

Another example

- In a basic solution
- $MnO_4^- + SO_3^{2-} \rightarrow MnO_4^{2-} + SO_4^{2-}$

Another example

- In a basic solution
- $MnO_4^- + SO_3^{2-} \rightarrow MnO_4^{2-} + SO_4^{2-}$
- Half reactions
- $MnO_4^- \rightarrow MnO_4^{2-}$
- $SO_3^{2-} \rightarrow SO_4^{2-}$

Half reactions

- $MnO_4^- \rightarrow MnO_4^{2-}$
- $MnO_4^- + e^- \rightarrow MnO_4^{2-}$
- $SO_3^{2-} \rightarrow SO_4^{2-}$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-}$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2\ H^+$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2\ H^+ + 2e^-$
- Double the top reaction

- $2\ MnO_4^- + 2\ e^- \rightarrow 2\ MnO_4^{2-}$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2\ H^+ + 2e^-$
- Combine them
- $2\ MnO_4^- + H_2O + SO_3^{2-} \rightarrow 2\ MnO_4^{2-} + SO_4^{2-} + 2\ H^+$
- Add OH^-
- $2\ MnO_4^- + H_2O + SO_3^{2-} + 2\ OH^- \rightarrow 2\ MnO_4^{2-} + SO_4^{2-} + 2\ H^+ + 2\ OH^-$

- $2\ MnO_4^- + H_2O + SO_3^{2-} + 2\ OH^-$
  $\rightarrow\ 2\ MnO_4^{2-} + SO_4^{2-} + 2\ H_2O$
- finishing
- $2\ MnO_4^- + SO_3^{2-} + 2\ OH^-$
  $\rightarrow\ 2\ MnO_4^{2-} + SO_4^{2-} + H_2O$