Reactions in Aqueous Solution

I. Double Displacement Reactions

\[ AB + CD \rightarrow CB + AD \]

1. Acid-Base Reactions - Acid + base → salt + water

\[ HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l) \]

"salt"

The net ionic equation of any strong acid-strong base reaction is always -

\[ H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \]

2. Precipitation Reactions – use Solubility Rules to predict the precipitate

\[ AgNO_3(aq) + NaBr(aq) \rightarrow AgBr(s) + NaNO_3(aq) \]
II. Redox Reactions - you MUST assign oxidation numbers to determine if a given reaction is a redox reaction or not

NOTE: the oxidation number for each element is shown in parentheses

1. Combination Reactions - A + B → AB

\[
\begin{align*}
\text{H}_2 &+ \text{Br}_2 \rightarrow 2 \text{HBr} \\
\text{hydrogen is oxidized from 0 to +1} &
\end{align*}
\]

\[
\begin{align*}
\text{H}_2 &\text{ is the reducing agent} \\
\text{bromine is reduced from 0 to } -1 &
\end{align*}
\]

\[
\begin{align*}
\text{HgO(s)} &\rightarrow \text{Hg(l)} + \text{O}_2(\text{g}) \\
\text{mercury is reduced from +2 to 0} &
\end{align*}
\]

\[
\begin{align*}
\text{HgO is the oxidizing agent} \\
\text{oxygen is oxidized from } -2 \text{ to } 0 &
\end{align*}
\]

2. Decomposition Reactions - AB → A + B

\[
\begin{align*}
\text{mercury is reduced from +2 to 0} &
\end{align*}
\]

\[
\begin{align*}
\text{HgO is the oxidizing agent} \\
\text{oxygen is oxidized from } -2 \text{ to } 0 &
\end{align*}
\]
3. **Single Displacement Reactions** - \( A + BC \rightarrow AC + B \)

Use the Activity Series (Fig 4.15, p. 124) for the (a) hydrogen, and (b) metal displacement categories with the general mechanism:

\[
M + BC \rightarrow MC + B \quad \text{where}
\]

\[
M = \text{metal} \\
BC = \text{H}_2\text{O or acid} \\
B = H_2(g)
\]

(a) **Hydrogen Displacement**

\[
\begin{align*}
(0) & \quad (+1)(-2) \quad (+1)(-2)(+1)^* \quad (0) \quad \text{lithium is oxidized from 0 to +1} \\
Li(s) & + H_2O(l) \rightarrow LiOH(aq) + H_2(g) \quad \text{(Li is the reducing agent)} \\
*\text{OH}^- & = \text{hydroxide polyatomic ion} \\
\end{align*}
\]

hydrogen is reduced from +1 to 0

\[
\begin{align*}
(0) & \quad (+1)(-1) \quad (+2)(-1) \quad (0) \\
Zn(s) & + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g) \quad \text{zinc is oxidized from 0 to +2} \\
\text{hydrogen is reduced from +1 to 0} & \quad \text{(Zn is the reducing agent)} \\
\text{(HCl is the oxidizing agent)}
\end{align*}
\]

(b) **Metal Displacement**

\[
\begin{align*}
(0) & \quad (+1)(+5)(-2)^* \quad (0) \quad (+2)(+5)(-2)^* \quad \text{magnesium is oxidized from 0 to +2} \\
Mg(s) & + 2 \text{ AgNO}_3(aq) \rightarrow Ag(s) + Mg(NO_3)_2(aq) \quad \text{(Mg is the reducing agent)} \\
*\text{NO}_3^- & = \text{nitrate polyatomic ion} \\
\end{align*}
\]

silver is reduced from +1 to 0

\[
\begin{align*}
\text{Use the halogen Activity Series } F_2 > Cl_2 > Br_2 > I_2 \text{ for –} \quad \text{(AgNO}_3 \text{ is the oxidizing agent)}
\end{align*}
\]

(c) **Halogen Displacement**

\[
\begin{align*}
(0) & \quad (+1)(-1) \quad (0) \quad (+1)(-1) \quad \text{chlorine is reduced from 0 to –1} \\
Cl_2(g) & + 2 \text{ NaI}(aq) \rightarrow I_2(s) + 2 \text{ NaCl}(aq) \quad \text{(Cl}_2 \text{ is the oxidizing agent)} \\
\text{iodide is oxidized from –1 to 0} & \quad \text{(NaI is the reducing agent)}
\end{align*}
\]
4. **Disproportionation Reactions** – the same element is both oxidized and reduced

\[
(+1)(+3)(-2)* \quad (+1)(+5)(-2)** \quad (+1)(-2) \quad (+2)(-2)
\]

\[
3 \text{HNO}_2 \rightarrow \text{HNO}_3 + \text{H}_2\text{O} + 2 \text{NO}
\]

the nitrogen in HNO₂ is oxidized from +3 to +5 in HNO₃, and also reduced from +3 to +2 in NO
(HNO₂ is both the reducing AND oxidizing agents)

*NO₂⁻ = nitrite polyatomic ion
**NO₃⁻ = nitrate polyatomic ion