LibreTexts

4.2: Balancing Redox Reactions

In studying redox chemistry, it is important to begin by learning to balance electrochemical reactions. Simple redox reactions (for example, $H_2 + I_2 \rightarrow 2\ HI$) can be balanced by inspection, but for more complex reactions it is helpful to have a foolproof, systematic method. The **ion-electron method** allows one to balance redox reactions regardless of their complexity. We illustrate this method with two examples.

Example 1:
$\mathbf{I^-}$ is oxidized to $\mathbf{IO_3^-}$ by $\mathbf{MnO_4^-}$, which is reduced to $\mathbf{Mn^{2+}}$.

How can this reaction be balanced? In the ion-electron method we follow a series of four steps:

Step 1A: Write out the (unbalanced) reaction and identify the elements that are undergoing redox.

$MnO_4^- + I^- \rightarrow IO_3^- + Mn^{2+}$ (The elements undergoing redox are **Mn** and **I**)

Step 1B: Separate the reaction into two **half reactions**, balancing the element undergoing redox in each.

$MnO_4^- \rightarrow Mn^{2+}$
$I^- \rightarrow IO_3^-$

Step 2A: Balance the **oxygen** atoms by adding water to one side of each half reaction.

$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
$3H_2O + I^- \rightarrow IO_3^-$

Step 2B: Balance the **hydrogen** atoms by adding H^+ ions.

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

The left side has a net charge of **+7** and the right side has a net charge of **+2**

$3H_2O + I^- \rightarrow IO_3^- + 6H^+$

The left side has a net charge of **-1** and the right side has a net charge of **+5**

Step 2C: Balance the overall charge by adding **electrons**

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

The left side has a charge of **+2** while the right side has a charge of **+2**. They are balanced.

$3H_2O + I^- \rightarrow IO_3^- + 6H^+ + \mathbf{6e^-}$

The left side has a charge of **-1** while the right side has a charge of **-1**. They are balanced.
Note: We did not need to explicitly determine the oxidation states of Mn or I to arrive at the correct number of electrons in each half reaction.

Step 3: Combine the half reactions so that there are equal numbers of electrons on the left and right sides

$\underline{6}\ (8H^+ + \mathbf{5e^-} + MnO_4^- \rightarrow Mn^{2+} + 4H_2O)$
$\underline{5}\ (3H_2O + I^- \rightarrow IO_3^- + 6H^+ + \mathbf{6e^-})$

$48H^+ + \mathbf{30e^-} + 15H_2O + 6MnO_4^- + 5I^- \rightarrow 5IO_3^- + 6Mn^{2+} + 24H_2O + \mathbf{30e^-} + 30H^+$
Cancel the $\mathbf{H^+}$, **electrons**, and **water**:
$\cancel{48H^+} + \cancel{30e^-} + \cancel{15H_2O} + 6MnO_4^- + 5I^- \rightarrow 5IO_3^- + 6Mn^{2+} + \cancel{24H_2O} + \cancel{30e^-} + \cancel{30H^+}$

The overall balanced reaction is therefore:
$18H^+ + 6MnO_4^- + 5I^- \rightarrow 5IO_3^- + 6Mn^{2+} + 9H_2O$

Check your work by making sure that all **elements** and **charges** are balanced.

Step 4: If the reaction occurs under basic conditions, we add **OH^-** to each side to cancel **H^+**
$18H^+ + 18OH^- + 6MnO_4^- + 5I^- \rightarrow 5IO_3^- + 6Mn^{2+} + 9H_2O + 18OH^-$

The **$18H^+$** + **$18OH^-$** will become **$18H_2O$** so the overall balanced reaction is:

$9H_2O + 6MnO_4^- + 5I^- \rightarrow 5IO_3^- + 6Mn^{2+} + 18OH^-$

Again, it is a good idea to check and make sure that all of the elements are balanced, and that the charge is the same on both sides. If this is not the case, you need to find the error in one of the earlier steps.

Example 2:
Redox reaction of **$S_2O_3^{2-}$** and **H_2O_2**
$S_2O_3^{2-} + H_2O_2 \rightarrow S_4O_6^{2-} + H_2O$

Which elements are undergoing redox? **S** and **O**

Step 1: Write out half reactions, balancing the element undergoing redox

$2S_2O_3^{2-} \rightarrow S_4O_6^{2-}$
$H_2O_2 \rightarrow 2H_2O$

Step 2A: Balance oxygen (already balanced)

Step 2B: Balance hydrogen:

$2S_2O_3^{2-} \rightarrow S_4O_6^{2-}$
$H_2O_2 + 2H^+ \rightarrow 2H_2O$

Step 2C: Balance charge by adding electrons:

$2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$

Step 3: Combine the half reactions so that there are equal numbers of electrons on the left and right sides (already equal)
Overall balanced reaction:

$2S_2O_3^{2-} + H_2O_2 + 2H^+ \rightarrow S_4O_6^{2-} + 2H_2O$

Note that again, we did not need to know the formal oxidation states of S or O in the reactants and products in order to balance the reaction. In this case, assigning the oxidation states would be rather complex, because $S_2O_3^{2-}$ and $S_4O_6^{2-}$ both contain sulfur in more than one oxidation state.