



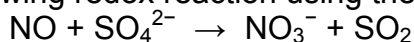
Balancing Redox Reactions 2: The Ion-Electron Method

In the first redox reaction worksheet, we saw the oxidation number method of balancing equations. This worksheet shows you another method.

The steps for balancing a redox reaction using the *ion-electron method* are:

- [1] **Break the equation into two half-reactions**, one for the oxidation step (loss of electrons) and one for the reduction step (gain of electrons). You will still need to use oxidation numbers to know which is which.
- [2] Obtain material balance (i.e. **balance the atoms**) in each half-reaction.
 - [a] Balance everything other than hydrogen and oxygen.
 - [b] Balance oxygen by adding H₂O to the other side.
 - [c] Balance hydrogen by adding H⁺ to the other side.
 - [d] IF THE REACTION IS IN BASIC SOLUTION, add equal amounts of OH⁻ to both sides to neutralize the H⁺. The OH⁻ and H⁺ combine to form water and leave excess OH⁻ on the other side. Cancel any water that appears on both sides. (ignore step d if solution is acidic)
- [3] Obtain charge balance for each half-reaction by **adding electrons as a product/reactant to the more positive side**.
- [4] **Combine the half-reactions to cancel the electrons**. You may have to multiply the equations by whole numbers to do this.

Example 1: Balance the following redox reaction using the ion-electron method:

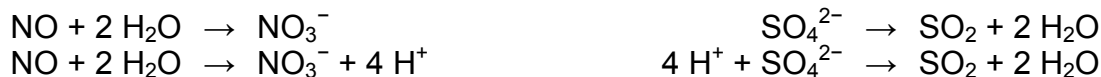


Solution: Following the steps above:

- [1] Nitrogen gets oxidized, and sulphur is reduced, so the half-reactions are:



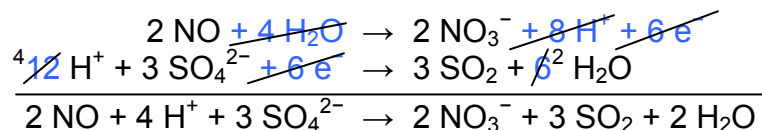
- [2] We balance the atoms:



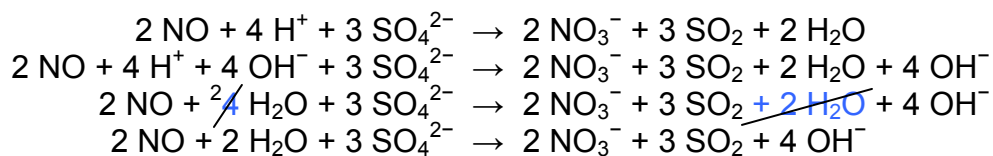
- [3] We add electrons so that the charge balances:



- [4] And finally we cancel the electrons:

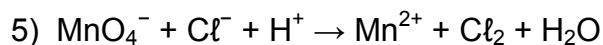
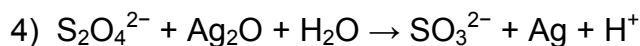
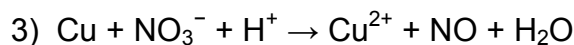
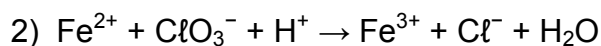
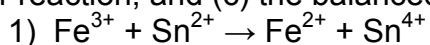


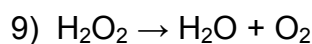
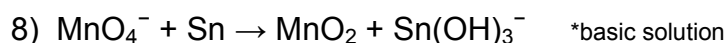
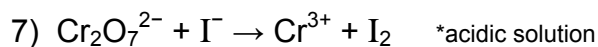
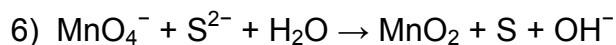
[5] With sulphates and nitrates, this reaction is not likely to take place in basic solution, but if it were, we would not be able to have H^+ in the final equation. We would add OH^- to both sides to cancel the H^+ that is there:



EXERCISES

A. For each redox equation, determine (a) the oxidation half-reaction, (b) the reduction half-reaction, and (c) the balanced redox reaction.





SOLUTIONS

- A. (1)a) $\text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + 2 \text{e}^-$ (b) $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$
(c) a + 2b: $2 \text{Fe}^{3+} + \text{Sn}^{2+} \rightarrow 2 \text{Fe}^{2+} + \text{Sn}^{4+}$
(2)a) $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$ (b) $\text{ClO}_3^- + 6 \text{H}^+ + 6 \text{e}^- \rightarrow \text{Cl}^- + 3 \text{H}_2\text{O}$
(c) 6a + b: $6 \text{Fe}^{2+} + \text{ClO}_3^- + 6 \text{H}^+ \rightarrow 6 \text{Fe}^{3+} + \text{Cl}^- + 3 \text{H}_2\text{O}$
(3)a) $\text{Cu} \rightarrow \text{Cu}^{2+} + 2 \text{e}^-$ (b) $\text{NO}_3^- + 4 \text{H}^+ + 3 \text{e}^- \rightarrow \text{NO} + 2 \text{H}_2\text{O}$
(c) 3a + 2b: $3 \text{Cu} + 2 \text{NO}_3^- + 8 \text{H}^+ \rightarrow 3 \text{Cu}^{2+} + 2 \text{NO} + 4 \text{H}_2\text{O}$
(4)a) $\text{S}_2\text{O}_4^{2-} + 2 \text{H}_2\text{O} \rightarrow 2 \text{SO}_3^{2-} + 4 \text{H}^+ + 2 \text{e}^-$ (b) $\text{Ag}_2\text{O} + 2 \text{H}^+ + 2 \text{e}^- \rightarrow 2 \text{Ag} + \text{H}_2\text{O}$
(c) a + b: $\text{S}_2\text{O}_4^{2-} + \text{Ag}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{SO}_3^{2-} + 2 \text{Ag} + 2 \text{H}^+$
(5)a) $2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 \text{e}^-$ (b) $\text{MnO}_4^- + 5 \text{e}^- + 8 \text{H}^+ \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}$
(c) 5a + 2b: $2 \text{MnO}_4^- + 10 \text{Cl}^- + 16 \text{H}^+ \rightarrow 2 \text{Mn}^{2+} + 5 \text{Cl}_2 + 8 \text{H}_2\text{O}$
(6)a) $\text{S}^{2-} \rightarrow \text{S} + 2 \text{e}^-$ (b) $\text{MnO}_4^- + 2 \text{H}_2\text{O} + 3 \text{e}^- \rightarrow \text{MnO}_2 + 4 \text{OH}^-$
(c) 3a + 2b: $2 \text{MnO}_4^- + 3 \text{S}^{2-} + 4 \text{H}_2\text{O} \rightarrow 2 \text{MnO}_2 + 3 \text{S} + 8 \text{OH}^-$
(7)a) $2 \text{I}^- \rightarrow \text{I}_2 + 2 \text{e}^-$ (b) $\text{Cr}_2\text{O}_7^{2-} + 6 \text{e}^- + 14 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$
(c) 3a + b: $\text{Cr}_2\text{O}_7^{2-} + 6 \text{I}^- + 14 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 3 \text{I}_2 + 7 \text{H}_2\text{O}$
(8)a) $\text{Sn} + 3 \text{OH}^- \rightarrow \text{Sn}(\text{OH})_3^- + 2 \text{e}^-$ (b) $\text{MnO}_4^- + 2 \text{H}_2\text{O} + 3 \text{e}^- \rightarrow \text{MnO}_2 + 4 \text{OH}^-$
(c) 3a + 2b: $2 \text{MnO}_4^- + 3 \text{Sn} + \text{OH}^- + 4 \text{H}_2\text{O} \rightarrow 2 \text{MnO}_2 + 3 \text{Sn}(\text{OH})_3^-$
(9)a) $\text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2 \text{H}^+ + 2 \text{e}^-$ (b) $\text{H}_2\text{O}_2 + 2 \text{H}^+ + 2 \text{e}^- \rightarrow 2 \text{H}_2\text{O}$
(c) a + b: $2 \text{H}_2\text{O}_2 \rightarrow 2 \text{H}_2\text{O} + \text{O}_2$