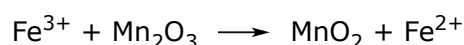


Balancing an electrochemical reaction in acidic aqueous solution is performed in the presence of $\text{H}_2\text{O}(\ell)$ and $\text{H}^+(\text{aq})$. In this case, any of these two species can be added if necessary to balance the reaction.

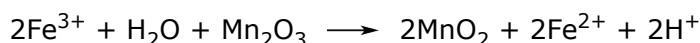
Steps for balancing a redox reaction (in an acidic solution)

1. Separate the reactions in two half-reactions that will both be balanced individually.
2. Balance the elements other than "O" and "H"
3. Balance "O" by adding $\text{H}_2\text{O}(\ell)$
4. Balance "H" by adding $\text{H}^+(\text{aq})$
5. Balance the charges by adding the appropriate number of electrons (e^-)
6. Make sure that the number of electrons is the same in both half reactions
7. Add the two half reactions together and cancel-out the electrons

Example: Balance the following reaction using the half reaction method

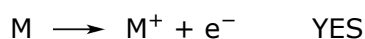
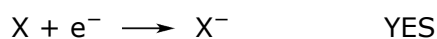


Step 1	$\text{Mn}_2\text{O}_3 \longrightarrow \text{MnO}_2$	$\text{Fe}^{3+} \longrightarrow \text{Fe}^{2+}$
Step 2	$\text{Mn}_2\text{O}_3 \longrightarrow 2\text{MnO}_2$	
Step 3	$\text{H}_2\text{O} + \text{Mn}_2\text{O}_3 \longrightarrow 2\text{MnO}_2$	
Step 4	$\text{H}_2\text{O} + \text{Mn}_2\text{O}_3 \longrightarrow 2\text{MnO}_2 + 2\text{H}^+$	
Step 5	$\text{H}_2\text{O} + \text{Mn}_2\text{O}_3 \longrightarrow 2\text{MnO}_2 + 2\text{H}^+ + 2\text{e}^-$	$\text{e}^- + \text{Fe}^{3+} \longrightarrow \text{Fe}^{2+}$
Step 6	$\text{H}_2\text{O} + \text{Mn}_2\text{O}_3 \longrightarrow 2\text{MnO}_2 + 2\text{H}^+ + 2\text{e}^-$	$2\text{e}^- + 2\text{Fe}^{3+} \longrightarrow 2\text{Fe}^{2+}$
Step 7	$2\text{e}^- + 2\text{Fe}^{3+} + \text{H}_2\text{O} + \text{Mn}_2\text{O}_3 \longrightarrow 2\text{MnO}_2 + 2\text{Fe}^{2+} + 2\text{H}^+ + 2\text{e}^-$	



IMPORTANT

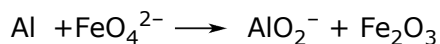
Electrons are negative particles that carry electrical charge. Since it is an actual physical entity, it cannot be "mathematically" subtracted in a chemical reaction.



Steps to balance a redox reaction in a basic aqueous solution

To balance the reaction in a basic solution, the initial steps are the same as those for an acidic solution. However, an additional step is added at the end to neutralize the acid.

Example: Balance the following reaction in a basic solution using the half reaction method.



Step 1	$\text{Al} \longrightarrow \text{AlO}_2^-$	$\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3$
Step 2	$\text{Al} \longrightarrow \text{AlO}_2^-$	$2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3$
Step 3	$2\text{H}_2\text{O} + \text{Al} \longrightarrow \text{AlO}_2^-$	$2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 5\text{H}_2\text{O}$
Step 4	$2\text{H}_2\text{O} + \text{Al} \longrightarrow \text{AlO}_2^- + 4\text{H}^+$	$10\text{H}^+ + 2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 5\text{H}_2\text{O}$
Step 5	$2\text{H}_2\text{O} + \text{Al} \longrightarrow \text{AlO}_2^- + 4\text{H}^+ + 3\text{e}^-$	$6\text{e}^- + 10\text{H}^+ + 2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 5\text{H}_2\text{O}$
Step 6	$4\text{H}_2\text{O} + 2\text{Al} \longrightarrow 2\text{AlO}_2^- + 8\text{H}^+ + 6\text{e}^-$	$6\text{e}^- + 10\text{H}^+ + 2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 5\text{H}_2\text{O}$
Step 7	$4\text{H}_2\text{O} + 2\text{Al} + 2\text{FeO}_4^{2-} + 10\text{H}^+ + 6\text{e}^- \longrightarrow \text{Fe}_2\text{O}_3 + 2\text{AlO}_2^- + 5\text{H}_2\text{O} + 8\text{H}^+ + 6\text{e}^-$	
Final:	$2\text{Al} + 2\text{FeO}_4^{2-} + 2\text{H}^+ \longrightarrow \text{Fe}_2\text{O}_3 + 2\text{AlO}_2^- + \text{H}_2\text{O}$	

In an **alkaline solution**, there must be no $\text{H}^+(\text{aq})$ present anywhere. If there is, it is neutralized by an acid-base neutralization with the addition of OH^- .

- Use the equation from "Step 7" and add the equivalent number of moles of $\text{OH}^-(\text{aq})$ to that of H^+ ON BOTH SIDES of the reaction to maintain equilibrium.
- H^+ is consumed and $\text{H}_2\text{O}(\ell)$ is produced (acid-base reaction).
- any excess H_2O , present on both side of the reaction is removed.

Step 8 (Neutralize H^+ with OH^-)

- $2\text{OH}^- + 2\text{H}^+ + 2\text{Al} + 2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 2\text{AlO}_2^- + \text{H}_2\text{O} + 2\text{OH}^-$
- $2\text{H}_2\text{O} + 2\text{Al} + 2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 2\text{AlO}_2^- + \text{H}_2\text{O} + 2\text{OH}^-$
- $\text{H}_2\text{O} + 2\text{Al} + 2\text{FeO}_4^{2-} \longrightarrow \text{Fe}_2\text{O}_3 + 2\text{AlO}_2^- + 2\text{OH}^-$

The reaction is now balanced in alkaline solution.

Exercises

- Balance each of the the following half-reactions in acidic aqueous solution.
Tell whether it is a reduction or an oxidation reaction.
 - $\text{SiO}_2 \longrightarrow \text{SiO}_4$
 - $\text{CO}_2 \longrightarrow \text{CO}$
 - $\text{Cr}^{3+} \longrightarrow \text{Cr}_2\text{O}_7^{2-}$
 - $\text{HN}_3 \longrightarrow \text{H}_3\text{N}$
 - $\text{HNO}_2 \longrightarrow \text{NO}_3^-$
- Balance the following redox reactions in acidic (H^+) aqueous solution.
Identify the reducing agent.
 - $\text{Cl}^- + \text{Sn} + \text{NO}_3^- \longrightarrow \text{SnCl}_6^{2-} + \text{NO}_2$
 - $\text{Fe}^{2+} + \text{NO}_2 \longrightarrow \text{NH}_3 + \text{Fe}^{3+}$
 - $\text{MnO}_4^- + \text{Sn}^{2+} \longrightarrow \text{Sn}^{4+} + \text{Mn}^{2+}$
 - $\text{N}_2\text{O}_4 \longrightarrow \text{NO}_3^- + \text{NO}_2^-$ (*hint : The two half-reactions use the same reactant.*)
- Balance the following redox reactions in basic (OH^-) aqueous solution.
Identify the oxidizing agent.
 - $\text{Ag} + \text{CN}^- + \text{O}_2 \longrightarrow \text{Ag}(\text{CN})_2^-$
 - $\text{NO}_2^- + \text{Al} \longrightarrow \text{NH}_3 + \text{AlO}_2^-$

Answer

- $2\text{H}_2\text{O} + \text{SiO}_2 \longrightarrow \text{SiO}_4 + 4\text{H}^+ + 4\text{e}^-$ (oxidation)
 - $2\text{e}^- + 2\text{H}^+ + \text{CO}_2 \longrightarrow \text{CO} + \text{H}_2\text{O}$ (reduction)
 - $7\text{H}_2\text{O} + 2\text{Cr}^{3+} \longrightarrow \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^-$ (oxidation)
 - $8\text{e}^- + 8\text{H}^+ + \text{HN}_3 \longrightarrow 3\text{H}_3\text{N}$ (reduction)
 - $\text{H}_2\text{O} + \text{HNO}_2 \longrightarrow \text{NO}_3^- + 3\text{H}^+ + 2\text{e}^-$ (oxidation)
- $8\text{H}^+ + 6\text{Cl}^- + \text{Sn} + 4\text{NO}_3^- \longrightarrow \text{SnCl}_6^{2-} + 4\text{NO}_2 + 4\text{H}_2\text{O}$ red. agent: Sn
 - $7\text{Fe}^{2+} + \text{NO}_2 + 7\text{H}^+ \longrightarrow \text{NH}_3 + 7\text{Fe}^{3+} + 2\text{H}_2\text{O}$ red. agent: Fe^{2+}
 - $2\text{MnO}_4^- + 5\text{Sn}^{2+} + 16\text{H}^+ \longrightarrow 5\text{Sn}^{4+} + 2\text{Mn}^{2+} + 8\text{H}_2\text{O}$ red. agent: Sn^{2+}
 - $2\text{N}_2\text{O}_4 + 2\text{H}_2\text{O} \longrightarrow 2\text{NO}_3^- + 2\text{NO}_2^- + 4\text{H}^+$ red. agent: N_2O_4
- $2\text{H}_2\text{O} + 4\text{Ag} + 8\text{CN}^- + \text{O}_2 \longrightarrow 4\text{Ag}(\text{CN})_2^- + 4\text{OH}^-$ ox. agent: O_2
 - $\text{NO}_2^- + 2\text{Al} + \text{OH}^- + \text{H}_2\text{O} \longrightarrow \text{NH}_3 + 2\text{AlO}_2^-$ ox. Agent: NO_2^-