Balancing an electrochemical reaction in acidic aqueous solution is performed in the presence of $H_2O(\ell)$ and $H^+(aq)$. 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 $H_2O(\ell)$
4. Balance "H" by adding $H^+(aq)$
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

$$Fe^{3+} + Mn_2O_3 \longrightarrow MnO_2 + Fe^{2+}$$

Step 1	$Mn_2O_3 \longrightarrow MnO_2$	$Fe^{3+} \longrightarrow Fe^{2+}$
Step 2	$Mn_2O_3 \longrightarrow \mathbf{2}MnO_2$	
Step 3	$\mathbf{H_2O} + Mn_2O_3 \longrightarrow 2MnO_2$	
Step 4	$H_2O + Mn_2O_3 \longrightarrow 2MnO_2 + \mathbf{2H^+}$	
Step 5	$H_2O + Mn_2O_3 \longrightarrow 2MnO_2 + 2H^+ + \mathbf{2e^-}$	$\mathbf{e^-} + Fe^{3+} \longrightarrow Fe^{2+}$
Step 6	$H_2O + Mn_2O_3 \longrightarrow 2MnO_2 + 2H^+ + \mathbf{2e^-}$	$\mathbf{2e^-} + \mathbf{2}Fe^{3+} \longrightarrow \mathbf{2}Fe^{2+}$
Step 7	$\mathbf{2e^-} + 2Fe^{3+} + H_2O + Mn_2O_3 \longrightarrow 2MnO_2 + 2Fe^{2+} + 2H^+ + \mathbf{2e^-}$	

$$2Fe^{3+} + H_2O + Mn_2O_3 \longrightarrow 2MnO_2 + 2Fe^{2+} + 2H^+$$

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.

$X + e^- \longrightarrow X^-$	YES
$M \longrightarrow M^+ + e^-$	YES
$M - e^- \longrightarrow M^+$	NO, NEVER!

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.

$$Al + FeO_4^{2-} \longrightarrow AlO_2^- + Fe_2O_3$$

Step 1	$Al \longrightarrow AlO_2^-$	$FeO_4^{2-} \longrightarrow Fe_2O_3$
Step 2	$Al \longrightarrow AlO_2^-$	$\mathbf{2}FeO_4^{2-} \longrightarrow Fe_2O_3$
Step 3	$\mathbf{2H_2O} + Al \longrightarrow AlO_2^-$	$2FeO_4^{2-} \longrightarrow Fe_2O_3 + \mathbf{5H_2O}$
Step 4	$2H_2O + Al \longrightarrow AlO_2^- + \mathbf{4H^+}$	$\mathbf{10H^+} + 2FeO_4^{2-} \longrightarrow Fe_2O_3 + 5H_2O$
Step 5	$2H_2O + Al \longrightarrow AlO_2^- + 4H^+ + \mathbf{3e^-}$	$\mathbf{6e^-} + 10H^+ + 2FeO_4^{2-} \longrightarrow Fe_2O_3 + 5H_2O$
Step 6	$\mathbf{4}H_2O + \mathbf{2}Al \longrightarrow \mathbf{2}AlO_2^- + \mathbf{8}H^+ + \mathbf{6}e^-$	$6e^- + 10H^+ + 2FeO_4^{2-} \longrightarrow Fe_2O_3 + 5H_2O$

Step 7 $\mathbf{4H_2O} + 2Al + 2FeO_4^{2-} + \mathbf{10H^+} + \mathbf{6e^-} \longrightarrow Fe_2O_3 + 2AlO_2^- + \mathbf{5H_2O} + \mathbf{8H^+} + \mathbf{6e^-}$

Final: $2Al + 2FeO_4^{2-} + \mathbf{2H^+} \longrightarrow Fe_2O_3 + 2AlO_2^- + \mathbf{H_2O}$

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

a. Use the equation from "Step 7" and add the equivalent number of moles of $OH^-(aq)$ to that of H^+ ON BOTH SIDES of the reaction to maintain equilibrium.

b. H^+ is consumed and $H_2O(\ell)$ is produced (acid-base reaction).

c. any excess H_2O, present on both side of the reaction is removed.

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

a. $\mathbf{2OH^-} + \mathbf{2H^+} + 2Al + 2FeO_4^{2-} \longrightarrow Fe_2O_3 + 2AlO_2^- + H_2O + \mathbf{2OH^-}$

b. $\mathbf{2H_2O} + 2Al + 2FeO_4^{2-} \longrightarrow Fe_2O_3 + 2AlO_2^- + \mathbf{H_2O} + 2OH^-$

c. $\mathbf{H_2O} + 2Al + 2FeO_4^{2-} \longrightarrow Fe_2O_3 + 2AlO_2^- + 2OH^-$

The reaction is now balanced in alkaline solution.

Exercises

1. Balance each of the the following half-reactions in acidic aqueous solution.
Tell whether it is a reduction or an oxidation reaction.

a. $SiO_2 \longrightarrow SiO_4$ d. $HN_3 \longrightarrow H_3N$

b. $CO_2 \longrightarrow CO$ e. $HNO_2 \longrightarrow NO_3^-$

c. $Cr^{3+} \longrightarrow Cr_2O_7^{2-}$

2. Balance the following redox reactions in acidic (H^+) aqueous solution.
Identify the reducing agent.

a. $Cl^- + Sn + NO_3^- \longrightarrow SnCl_6^{2-} + NO_2$

b. $Fe^{2+} + NO_2 \longrightarrow NH_3 + Fe^{3+}$

c. $MnO_4^- + Sn^{2+} \longrightarrow Sn^{4+} + Mn^{2+}$

d. $N_2O_4 \longrightarrow NO_3^- + NO_2^-$ *(hint : The two half-reactions use the same reactant.)*

3. Balance the following redox reactions in basic (OH^-) aqueous solution.
Identify the oxidizing agent.

a. $Ag + CN^- + O_2 \longrightarrow Ag(CN)_2^-$

b. $NO_2^- + Al \longrightarrow NH_3 + AlO_2^-$

Answer

1. a. $2H_2O + SiO_2 \longrightarrow SiO_4 + 4H^+ + 4e^-$ (oxidation)
   b. $2e^- + 2H^+ + CO_2 \longrightarrow CO + H_2O$ (reduction)
   c. $7H_2O + 2Cr^{3+} \longrightarrow Cr_2O_7^{2-} + 14H^+ + 6e^-$ (oxidation)
   d. $8e^- + 8H^+ + HN_3 \longrightarrow 3H_3N$ (reduction)
   e. $H_2O + HNO_2 \longrightarrow NO_3^- + 3H^+ + 2e^-$ (oxidation)

2. a. $8H^+ + 6Cl^- + Sn + 4NO_3^- \longrightarrow SnCl_6^{2-} + 4NO_2 + 4H_2O$ red. agent: Sn
   b. $7Fe^{2+} + NO_2 + 7H^+ \longrightarrow NH_3 + 7Fe^{3+} + 2H_2O$ red. agent: Fe^{2+}
   c. $2MnO_4^- + 5Sn^{2+} + 16H^+ \longrightarrow 5Sn^{4+} + 2Mn^{2+} + 8H_2O$ red. agent: Sn^{2+}
   d. $2N_2O_4 + 2H_2O \longrightarrow 2NO_3^- + 2NO_2^- + 4H^+$ red. agent: N_2O_4

3. a. $2H_2O + 4Ag + 8CN^- + O_2 \longrightarrow 4Ag(CN)_2^- + 4OH^-$ ox. agent: O_2
   b. $NO_2^- + 2Al + OH^- + H_2O \longrightarrow NH_3 + 2AlO_2^-$ ox. Agent: NO_2^-