Balancing Redox Equations

Strategy for redox equations in **acidic** *solutions*

Problem: When dilute nitric acid is poured on a piece of copper metal, copper(II) ions and the gas nitric oxide, NO, are formed. Write the balanced equation for the reaction.

Step 0: *Write the "skeleton equation"*

$$\mathrm{Cu(s) + H^+(aq) + NO_3{}^-(aq) \longrightarrow Cu^{2+}(aq) + NO(g)}$$

and determine the oxidation state (the oxidation state is *per* atom!)

$$\overset{0}{\mathrm{Cu}}\mathrm{(s)} + \overset{+1}{\mathrm{H}}{}^+\mathrm{(aq)} + \overset{+5}{\mathrm{N}}\overset{-2}{\mathrm{O}}{}_3{}^-\mathrm{(aq)} \longrightarrow \overset{+2}{\mathrm{Cu}}{}^{2+}\mathrm{(aq)} + \overset{+2}{\mathrm{N}}\overset{-2}{\mathrm{O}}\mathrm{(g)}$$

So we have

$$\begin{aligned} \mathrm{Cu} &: 0 \rightarrow +2 \qquad \text{ox} \\ \mathrm{N} &: +5 \rightarrow +2 \qquad \text{red} \end{aligned}$$

Step 1: *Write the "skeleton" half-reactions*

$$\begin{aligned} \text{ox:} &\quad \mathrm{Cu \longrightarrow Cu^{2+}} \\ \text{red:} &\quad \mathrm{NO_3{}^- \longrightarrow NO} \end{aligned}$$

Step 2: *Balance each half-reaction "atomically"*

- all atoms other than H and O (You can use *any* of the species that appear in the skeleton equation—Step 0—for this purpose.)
- balance O atoms by adding H_2O
- balance H atoms by adding H^+

$$\text{ox:}\quad \mathrm{Cu} \longrightarrow \mathrm{Cu^{2+}}$$

$$\text{red:}\quad \mathrm{NO_3{}^-} \longrightarrow \mathrm{NO} + 2\,\mathrm{H_2O}$$

$$\text{red:}\quad \mathrm{NO_3{}^-} + 4\,\mathrm{H^+} \longrightarrow \mathrm{NO} + 2\,\mathrm{H_2O}$$

Step 3: *Balance the electric charges by adding electrons*

$$\text{ox:}\quad \mathrm{Cu} \longrightarrow \mathrm{Cu^{2+}} + 2\,\mathrm{e^-}$$

$$\text{red:}\quad \mathrm{NO_3{}^-} + 4\,\mathrm{H^+} + 3\,\mathrm{e^-} \longrightarrow \mathrm{NO} + 2\,\mathrm{H_2O}$$

The electrons have to appear on the *right* hand side for the *oxidation* half-reaction, and on the *left* hand side for the *reduction* half-reaction.

Step 4: *Prepare the two half-equations for summation by making the number of electrons the same in both, i.e, find the least common multiple*

$$3 \times \text{ox:}\quad 3\,\mathrm{Cu} \longrightarrow 3\,\mathrm{Cu^{2+}} + 6\,\mathrm{e^-}$$

$$2 \times \text{red:}\quad 2\,\mathrm{NO_3{}^-} + 8\,\mathrm{H^+} + 6\,\mathrm{e^-} \longrightarrow 2\,\mathrm{NO} + 4\,\mathrm{H_2O}$$

Step 5: *Combine the two half-reactions*

$$3\,\mathrm{Cu} + 2\,\mathrm{NO_3{}^-} + 8\,\mathrm{H^+} + 6\,\mathrm{e^-} \longrightarrow 3\,\mathrm{Cu^{2+}} + 6\,\mathrm{e^-} + 2\,\mathrm{NO} + 4\,\mathrm{H_2O}$$

Step 6: *Simplify*

$$3\,\mathrm{Cu} + 2\,\mathrm{NO_3{}^-} + 8\,\mathrm{H^+} \longrightarrow 3\,\mathrm{Cu^{2+}} + 2\,\mathrm{NO} + 4\,\mathrm{H_2O}$$

Step 7: *Indicate the state of each species*

$$3\,\mathrm{Cu(s)} + 2\,\mathrm{NO_3{}^-(aq)} + 8\,\mathrm{H^+(aq)} \longrightarrow 3\,\mathrm{Cu^{2+}(aq)} + 2\,\mathrm{NO(g)} + 4\,\mathrm{H_2O(l)}$$

This is the fully balanced net ionic equation.

Check by writing a table:

species	left	right
Cu	3	3
N	2	2
O	6	6
H	8	8
charge	+6	+6

Strategy for redox equations in **basic** *solutions*

In basic solutions, no H^+ are available to balance H!

Strategy: Pretend the solution is acidic and carry out a "neutralization reaction" at the end.

Problem: The reaction of permanganate ions with bromide ions in basic solution yields solid manganese(IV) oxide and bromate ions. Write the fully balanced net equation.

Step 0: *Write the "skeleton equation" and determine the oxidation state* (the oxidation state is *per* atom!)

$$\overset{+7\,-2}{MnO_4^-}(aq) + \overset{-1}{Br^-}(aq) + \overset{-2\,+1}{OH^-} \longrightarrow \overset{+4\,-2}{MnO_2}(s) + \overset{+5\,-2}{BrO_3^-}(aq)$$

So we have

$$\begin{aligned} Br: -1 &\rightarrow +5 \qquad \text{ox} \\ Mn: +7 &\rightarrow +4 \qquad \text{red} \end{aligned}$$

Step 1: *Write the "skeleton" half-reactions*

$$\begin{aligned} \text{ox:} \quad Br^- &\longrightarrow BrO_3^- \\ \text{red:} \quad MnO_4^- &\longrightarrow MnO_2 \end{aligned}$$

Step 2: *Balance each half-reaction "atomically"*

- all atoms other than H and O (You can use *any* of the species that appear in the skeleton equation—Step 0—for this purpose.)
- balance O atoms by adding H_2O

- balance H atoms by adding H^+

$$\text{ox:}\quad Br^- + 3\,H_2O \longrightarrow {BrO_3}^-$$

$$\text{ox:}\quad Br^- + 3\,H_2O \longrightarrow {BrO_3}^- + 6\,H^+$$

$$\text{red:}\quad {MnO_4}^- \longrightarrow MnO_2 + 2\,H_2O$$

$$\text{red:}\quad {MnO_4}^- + 4\,H^+ \longrightarrow MnO_2 + 2\,H_2O$$

Step 3: *Balance the electric charges by adding electrons*

$$\text{ox:}\quad Br^- + 3\,H_2O \longrightarrow {BrO_3}^- + 6\,H^+ + 6\,e^-$$

$$\text{red:}\quad {MnO_4}^- + 4\,H^+ + 3\,e^- \longrightarrow MnO_2 + 2\,H_2O$$

The electrons have to appear on the *right* hand side for the *oxidation* half-reaction, and on the *left* hand side for the *reduction* half-reaction.

Step 4: *Prepare the two half-equations for summation by making the number of electrons the same in both, i.e, find the least common multiple*

$$\text{ox:}\quad Br^- + 3\,H_2O \longrightarrow {BrO_3}^- + 6\,H^+ + 6\,e^-$$

$$2 \times \text{red:}\quad 2\,{MnO_4}^- + 8\,H^+ + 6\,e^- \longrightarrow 2\,MnO_2 + 4\,H_2O$$

Step 5: *Combine the two half-reactions*

$$Br^- + 3\,H_2O + 2\,{MnO_4}^- + 8\,H^+ + 6\,e^- \longrightarrow {BrO_3}^- + 6\,H^+ + 6\,e^- + 2\,MnO_2 + 4\,H_2O$$

Step 6: *Simplify*

$$Br^- + 2\,{MnO_4}^- + 2\,H^+ \longrightarrow {BrO_3}^- + +2\,MnO_2 + H_2O$$

Step 6a: *Change to* **basic** *solution by adding as many* OH^- *to* **both** *sides as there are* H^+

$$Br^- + 2\,MnO_4{}^- + 2\,H^+ + 2\,OH^- \longrightarrow BrO_3{}^- + +2\,MnO_2 + H_2O + 2\,OH^-$$

"*neutralization*": Combine the H^+ and the OH^- to form H_2O

$$Br^- + 2\,MnO_4{}^- + 2\,H_2O \longrightarrow BrO_3{}^- + +2\,MnO_2 + H_2O + 2\,OH^-$$

simplify

$$Br^- + 2\,MnO_4{}^- + H_2O \longrightarrow BrO_3{}^- + +2\,MnO_2 + 2\,OH^-$$

Step 7: *Indicate the state of each species*

$$Br^-(aq) + 2\,MnO_4{}^-(aq) + H_2O(l) \longrightarrow BrO_3{}^-(aq) + +2\,MnO_2(s) + 2\,OH^-(aq)$$

This is the fully balanced net ionic equation.

Check by writing a table:

species	left	right
Br	1	1
Mn	2	2
O	9	9
H	2	2
charge	−3	−3