Balancing Oxidation-Reduction Equations
by the Oxidation Number Change Method

Four Easy Steps:

1. Determine the oxidation numbers of the species being oxidized and reduced (and make sure there are the same number of atoms on each side).
2. Balance the changes in oxidation numbers by multiplying by the appropriate coefficient.
3. Balance charges with:
   a. H⁺ in acidic solution.
   b. OH⁻ in basic solution.
4. Balance H (and O!) with H₂O.

Example 1:

Given the skeletal equation $\text{H}_3\text{PO}_2 + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{H}_3\text{PO}_4 + \text{Cr}^{3+}$ (in acidic solution),

Step 1: Phosphorus is being oxidized (+1 to +5) and chromium is being reduced (+6 to +3). Put a 2 in front of the Cr³⁺ to have the same number of chromiums on both sides of the equation:

$$\begin{align*}
\text{H}_3\text{PO}_2 + \text{Cr}_2\text{O}_7^{2-} & \rightarrow \text{H}_3\text{PO}_4 + 2\text{Cr}^{3+} \\
(+1) & \quad (+6) \\
\uparrow 4 & \quad \uparrow 6
\end{align*}$$

Step 2: We see that phosphorus is going up 4 in oxidation number and chromium is going down 6 (down 3 each but there are two chromiums).
We need to multiply all the phosphorus species by 3 and the chromium species by 2 to get the same number going up as going down (up 12, down 12):

\[3\text{H}_2\text{PO}_4^- + 2\text{Cr}_2\text{O}_7^{2-} \rightarrow 3\text{H}_3\text{PO}_4 + 4\text{Cr}^{3+}\]

Step 3: Now we can balance charges with H+. Note that the net charge is -4 on the left hand side of the equation, and +12 on the right. Therefore we need to add 16 H+ to the left hand side of the equation (to get a net charge of +12):

\[3\text{H}_3\text{PO}_4 + 2\text{Cr}_2\text{O}_7^{2-} + 16\text{H}^+ \rightarrow 3\text{H}_3\text{PO}_4 + 4\text{Cr}^{3+}\]

Step 4: Lastly, we balance H or O with H2O. (*Nota bene*: they'd better both balance, or you've done something wrong!) In this case, we need to add 8 waters to the right side of the equation (25 H and 20 O on each side):

\[3\text{H}_3\text{PO}_4 + 2\text{Cr}_2\text{O}_7^{2-} + 16\text{H}^+ \rightarrow 3\text{H}_3\text{PO}_4 + 4\text{Cr}^{3+} + 8\text{H}_2\text{O}\]

And voila! We are done.

---

**Example 2:**

Given the skeletal equation \[\text{MnO}_4^- + \text{C}_2\text{O}_4^{2-} \rightarrow \text{MnO}_2 + \text{CO}_2\] (in basic solution)

Step 1: Manganese is being reduced (+7 to +4) and carbon is being oxidized (+3 to +4). Put a 2 in front of the CO2 to have the same number of carbons on both sides of the equation:

\[
\begin{align*}
\text{MnO}_4^- + 2\text{C}_2\text{O}_4^{2-} & \rightarrow \text{MnO}_2 + 2\text{CO}_2 \\
(+7) & \quad (+3) \quad (4) \quad (4) \\
\text{MnO}_4^- + 2\text{C}_2\text{O}_4^{2-} & \rightarrow \text{MnO}_2 + 2\text{CO}_2
\end{align*}
\]
Step 2: We see that manganese is going down 3 in oxidation number and carbon is going up 2 (up 1 each, but there are 2 carbons):

\[
\begin{align*}
\text{MnO}_4^- + C_2O_4^{2-} & \rightarrow \text{MnO}_2 + 2\text{CO}_2 \\
\downarrow 3 & \quad \uparrow 2
\end{align*}
\]

We need to multiply all the manganese species by 2 and the carbon species by 3 to get the same number going up as down:

\[
2\text{MnO}_4^- + 3C_2O_4^{2-} \rightarrow 2\text{MnO}_2 + 6\text{CO}_2
\]

Step 3: Now we can balance charges with OH\(^-\). Notice that the net charge on the left hand side of the equation is -8, but zero on the right. Therefore we need to add 8 OH\(^-\) to the right:

\[
2\text{MnO}_4^- + 3C_2O_4^{2-} \rightarrow 2\text{MnO}_2 + 6\text{CO}_2 + 8\text{OH}^-
\]

Lastly, we balance H or O with H\(_2\)O. We need to add 4 waters to the left (8 H's and 24 O's on each side).

\[
2\text{MnO}_4^- + 3C_2O_4^{2-} + 4\text{H}_2\text{O} \rightarrow 2\text{MnO}_2 + 6\text{CO}_2 + 8\text{OH}^-
\]

Balanced!