

Balancing Redox Equations by the Half-Reaction Method

Step 0 **Choose the skeletal equation.** Eliminate spectator ions such as Na^+ , K^+ . Eliminate ions such as sulfate, nitrate, and chloride **if** they are not involved in the redox process. Leave out H^+ and OH^- . Be careful to keep appropriate charges on the ions remaining in the skeletal equation.

Step 1 **Divide the skeletal equation into two half-reactions, one below the other.** One will be the oxidation process, the other the reduction process. Either half-reaction can include any reactant (on the reactant side) or product (on the product side) as needed to balance atoms in that half-reaction.

Step 2 In each half-reaction **balance all atoms other than H and O** by appropriate coefficients.

Step 3 In each half-reaction **balance O by adding H_2O molecules** to the side deficient in O atoms.

Step 4 In each half-reaction **balance H by adding H^+ ions** to the side deficient in H atoms.

Step 5 In each half-reaction **balance charges by adding electrons, e^- , to the more positive side.**
Note: Charges are counted by multiplying by coefficients: 3SO_4^{2-} contribute charge = -6 . Charges are balanced by making the two sides equal in charge but not necessarily equal to zero. [Note: $+2 = +2$ but $+1 \neq -1$] Subscripts do not multiply charges.

Step 6 **Choose multipliers for the two half-reactions to balance electrons gained and lost.** It will work to multiply the first equation by the number of e^- in the second equation and to multiply the second equation by the number of e^- in the first equation, but least common multipliers are better. If both half-reactions have electrons on the same side, something is wrong in at least one of the previous steps. Make certain ions have the appropriate charges; double check all steps--especially step 5.

Step 7 **Multiply the half-reactions by the chosen multipliers and add them together into one equation.** Cancel e^- and excess H^+ and/or H_2O that end up on both sides of the equation. Divide through where possible to reduce coefficients to smallest whole numbers.

Step 8 **Check the final equation to make certain all atoms and charges are balanced.**

In **acidic** solutions the above steps complete the balancing process. H^+ may be on either side of the equation.

In **basic** solutions there must be **no H^+ ions in the final equation.** Use step 9 to adjust as necessary.

In **neutral** solutions there must be **no H^+ or OH^- on the reactant side** of the equation. Use step 9 to adjust.

Step 9 If one side of the equation has H^+ ions, **add exactly that number of OH^- to both sides** of the equation. If both sides have H^+ , cancel excess. On the side with both H^+ and OH^- , **replace each pair ($\text{H}^+ + \text{OH}^-$) with one H_2O .** Cancel excess water. Rewrite the equation and repeat step 8.

Note: In basic solutions it may be straightforward to balance one or both half-reactions for H and O by adding OH^- directly to one side instead of following steps 3 and 4. It may still be necessary to apply step 9. Cancel excess OH^- if they end up on both sides of the final equation.

Example of balancing a redox reaction using these steps.

When $\text{K}_2\text{Cr}_2\text{O}_7(\text{aq})$ [orange] is acidified with $\text{H}_2\text{SO}_4(\text{aq})$ and reacted with $\text{H}_2\text{O}_2(\text{aq})$ [colorless] the solution turns green, indicating formation of Cr^{3+} , and bubbles are formed indicating the formation of $\text{O}_2(\text{g})$.

The redox reaction **before step 0** is $\text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \longrightarrow \text{Cr}^{3+} + \text{O}_2(\text{g})$

Written as an ionic equation this is: $2 \text{K}^+ + \text{Cr}_2\text{O}_7^{2-} + 2 \text{H}^+ + \text{SO}_4^{2-} + \text{H}_2\text{O}_2 \longrightarrow \text{Cr}^{3+} + \text{O}_2$

Step 0 Choose the skeletal equation. Eliminate spectator ions such as Na^+ , K^+ . Eliminate ions such as sulfate, nitrate, and chloride if they are not involved in the redox process. Leave out H^+ and OH^- . Be careful to keep appropriate charges on the ions remaining in the skeletal equation.

After step 0, this equation becomes: $\text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}_2 \longrightarrow \text{Cr}^{3+} + \text{O}_2$

Step 1 Divide the skeletal equation into two half-reactions, one below the other. One will be the oxidation process, the other the reduction process.

After step 1, this looks like: $\begin{array}{l} \text{Cr}_2\text{O}_7^{2-} \longrightarrow \text{Cr}^{3+} \\ \text{H}_2\text{O}_2 \longrightarrow \text{O}_2 \end{array}$ Cr atoms must go in the same half reaction

Step 2 In each half-reaction **balance all atoms other than H and O** by appropriate coefficients.

After step 2, this looks like $\begin{array}{l} \text{Cr}_2\text{O}_7^{2-} \longrightarrow 2 \text{Cr}^{3+} \\ \text{H}_2\text{O}_2 \longrightarrow \text{O}_2 \end{array}$ Charges will be balanced later.
H and O will be balanced later.

Step 3 In each half-reaction **balance O by adding H_2O molecules** to the side deficient in O atoms.

After step 3, this becomes: $\begin{array}{l} \text{Cr}_2\text{O}_7^{2-} \longrightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O} \\ \text{H}_2\text{O}_2 \longrightarrow \text{O}_2 \end{array}$ No change here; O are balanced

Step 4 In each half-reaction **balance H by adding H^+ ions** to the side deficient in H atoms.

After step 4, this becomes: $\begin{array}{l} 14 \text{H}^+ + \text{Cr}_2\text{O}_7^{2-} \longrightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O} \\ \text{H}_2\text{O}_2 \longrightarrow \text{O}_2 + 2 \text{H}^+ \end{array}$ All atoms are balanced
Charges are not balanced.

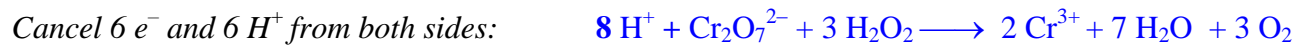
Step 5 In each half-reaction **balance charges by adding electrons, e^- , to the more positive side.**

After step 5, this is: $\begin{array}{l} 6 e^- + 14 \text{H}^+ + \text{Cr}_2\text{O}_7^{2-} \longrightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O} \\ \text{H}_2\text{O}_2 \longrightarrow \text{O}_2 + 2 \text{H}^+ + 2 e^- \end{array}$ $6 \times (-1) + 14 \times (+1) + (-2) = +6 = 2 \times (+3)$
 $0 = 2 \times (+1) + 2 \times (-1)$

Step 6 Choose multipliers for the two half-reactions to balance electrons gained and lost.

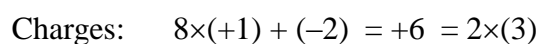
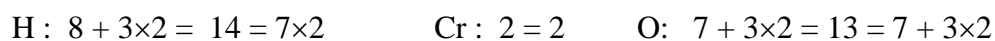
Step 6 gives: $\begin{array}{l} 6 e^- + 14 \text{H}^+ + \text{Cr}_2\text{O}_7^{2-} \longrightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O} \quad \times 1 \\ \text{H}_2\text{O}_2 \longrightarrow \text{O}_2 + 2 \text{H}^+ + 2 e^- \quad \times 3 \end{array}$ Multiply the whole equation by one.
To give 6 (2×3) e^- on the right.

Step 7 **Multiply the half-reactions by the chosen multipliers and add them together into one equation.**



After step 7 the equation is balanced.

Step 8 **Check the final equation to make certain all atoms and charges are balanced.**



The balanced redox equation in acid solution is: $8 H^{+} + Cr_2O_7^{2-} + 3 H_2O_2 \longrightarrow 2 Cr^{3+} + 7 H_2O + 3 O_2$

You should not need to rewrite equations between steps 1 and 7.