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, $e^–$, <u>to the more positive side</u>**. Note: Charges are counted by multiplying by coefficients: 3 $SO_4^{2–}$ 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 ≠ –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 <u>all atoms</u> and <u>charges</u> 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 $OH^–$ <u>to both sides</u>** of the equation. If both sides have H^+, cancel excess. On the side with both H^+ and $OH^–$, **replace each pair** ($H^+ + OH^–$) **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 $K_2Cr_2O_7(aq)$ [orange] is acidified with $H_2SO_4(aq)$ and reacted with $H_2O_2(aq)$ [colorless] the solution turns green, indicating formation of Cr^{3+}, and bubbles are formed indicating the formation of $O_2(g)$.

The redox reaction **before step 0** is $K_2Cr_2O_7 + H_2SO_4(aq) + H_2O_2(aq) \longrightarrow Cr^{3+} + O_2(g)$

Written as an ionic equation this is: $2\,K^+ + Cr_2O_7^{2-} + 2\,H^+ + SO_4^{2-} + H_2O_2 \longrightarrow Cr^{3+} + O_2$

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*: $Cr_2O_7^{2-} + H_2O_2 \longrightarrow Cr^{3+} + O_2$

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:*

$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$ Cr atoms must go in the same half reaction

$H_2O_2 \longrightarrow O_2$

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

***After step 2**, this looks like*

$Cr_2O_7^{2-} \longrightarrow \mathbf{2}\,Cr^{3+}$ Charges will be balanced later.

$H_2O_2 \longrightarrow O_2$ 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*:

$Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+} + \mathbf{7\,H_2O}$

$H_2O_2 \longrightarrow O_2$ 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*:

$\mathbf{14\,H^+} + Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+} + 7\,H_2O$ All atoms are balanced

$H_2O_2 \longrightarrow O_2 + \mathbf{2\,H^+}$ Charges are not balanced.

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

***After step 5**, this is*:

$\mathbf{6\,e^-} + 14\,H^+ + Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+} + 7\,H_2O$ $6\times(-1) + 14\times(+1) + (-2) = +6 = 2\times(+3)$

$H_2O_2 \longrightarrow O_2 + 2\,H^+ + \mathbf{2\,e^-}$ $0 = 2\times(+1) + 2\times(-1)$

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

***Step 6** gives*:

$6\,e^- + 14\,H^+ + Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+} + 7\,H_2O$ $\times\mathbf{1}$ Multiply the whole equation by one.

$H_2O_2 \longrightarrow O_2 + 2\,H^+ + 2\,e^-$ $\times\mathbf{3}$ 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**.

Multiply and add: $6\ e^{-} + 14\ H^{+} + Cr_2O_7^{2-} + \mathbf{3}\ H_2O_2 \longrightarrow 2\ Cr^{3+} + 7\ H_2O + \mathbf{3}\ O_2 + \mathbf{6}\ H^{+} + \mathbf{6}\ e^{-}$

Cancel 6 e^{-} and 6 H^{+} from both sides: $\mathbf{8}\ H^{+} + Cr_2O_7^{2-} + 3\ H_2O_2 \longrightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ O_2$

After step 7 *the equation is balanced.*

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

H : $8 + 3\times2 = 14 = 7\times2$ Cr : $2 = 2$ O: $7 + 3\times2 = 13 = 7 + 3\times2$

Charges: $8\times(+1) + (-2) = +6 = 2\times(3)$

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.