Examples of Balancing a Redox Reaction Using Half-Reaction Method 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.**

***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 now 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 this whole equation by one.

$H_2O_2 \longrightarrow O_2 + 2\,H^+ + 2\,e^-$ $\times\mathbf{3}$ This one by 3 to give $2\times3 = 6\,e^-$

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. When you have finished balancing this equation, your work should look approximately like this:

$$6\ e^{-} + 14\ H^{+} + Cr_2O_7^{2-} \longrightarrow 2\ Cr^{3+} + 7\ H_2O \qquad \times1$$

$$H_2O_2 \longrightarrow O_2 + 2\ H^{+} + 2\ e^{-} \qquad \times3$$

multiply these in your head and add them into one equation

$$\cancel{6\ e^{-}} + \underset{8}{\cancel{14}}\ H^{+} + Cr_2O_7^{2-} + 3\ H_2O_2 \longrightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ O_2 + \cancel{6\ H^{+}} + \cancel{6\ e^{-}}$$

$$8\ H^{+} + Cr_2O_7^{2-} + 3\ H_2O_2 \longrightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ O_2$$

rewrite the equation if unclear

Balance a Redox Equation in Basic Solution by the Half-Reaction Method Steps

This redox reaction can occur in either an acid or a base:

(1) $IO_3^- + I^- \longrightarrow I_2$

One might mix KIO_3(aq) with H_2SO_4(aq) and KI(aq). The equation would be balanced by the half-reaction method for acidic solutions:

(2) $10\ e^- + 12H^+ + 2\ IO_3^- \longrightarrow I_2 + 6\ H_2O \quad \times 1$

(3) $2\ I^- \longrightarrow I_2 + 2\ e^- \quad \times 5$

(4) $12H^+ + 2\ IO_3^- + 10\ I^- \longrightarrow 6\ I_2 + 6\ H_2O$

The same reaction, equation (1), could take place in basic solution. One might mix KIO_3(aq) with NaOH(aq) and KI(aq). The equation **cannot** be balanced in exactly the same way because there is essentially no H^+ in a basic solution, such as a solution with NaOH. Since acids and bases neutralize each other, they cannot exist together. They react to give water.

(5) $H^+ + OH^- \longrightarrow H_2O$

A solution can have an excess of H^+ or OH^- but not both at once. So the equation (4) is not correct in a basic solution because there is no H^+ reactant in a basic solution. We were able to balance oxygen and hydrogen in an acidic using H_2O and H^+, respectively, but technically we must use H_2O and OH^- in a basic solution to balance oxygen and hydrogen. This is more difficult to do directly since to add either one changes both oxygen and hydrogen. However, we can balance oxygen and hydrogen as if the solution were acidic and then use equation (5) to correct for a basic solution. Since equation (4) is balanced, it will still be balanced if we add **exactly the same molecules** to both sides of the equation. Since equation (4) has $12\ H^+$ that should not be there, this can be fixed by adding $12\ OH^-$ to both sides of the equation:

(6) $12\ OH^- + 12\ H^+ + 2\ IO_3^- + 10\ I^- \longrightarrow 6\ I_2 + 6\ H_2O + 12\ OH^-$

Then since equation (5) is true, the $12\ OH^- + 12\ H^+$ can be replaced with $12\ H_2O$:

(7) $12\ H_2O + 2\ IO_3^- + 10\ I^- \longrightarrow 6\ I_2 + 6\ H_2O + 12\ OH^-$

The resulting equation (7) can be improved by canceling 6 H_2O from both sides to give:

(7) $6\ H_2O + 2\ IO_3^- + 10\ I^- \longrightarrow 6\ I_2 + 12\ OH^-$

This is balanced: $6{\times}2\ H = 12\ H$; $6\ O + 2{\times}3\ O = 12\ O$; $2\ I + 10\ I = 6{\times}2\ I$; $2(-) + 10(-) = 12\ (-)$

The overall work for the reaction, $IO_3^- + I^- \longrightarrow I_2$ in basic solution, might look like:

$$10\ e^- + 12H^+ + 2\ IO_3^- \longrightarrow I_2 + 6\ H_2O \quad \times 1$$
$$2\ I^- \longrightarrow I_2 + 2\ e^- \quad \times 5$$

$$12OH^- + 12H^+ + 2\ IO_3^- + 10\ I^- \longrightarrow 6\ I_2 + 6\ H_2O + 12\ OH^-$$
$$12\ H_2O \quad 6H_2O$$
$$6\ H_2O + 2\ IO_3^- + 10\ I^- \longrightarrow 6\ I_2 + 12\ OH^-$$