BALANCING REDOX REACTIONS

Half-Reaction Method

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An overall redox reaction can be balanced using half-reactions once the correct number of electrons have been accounted for.

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$$\overset{+2}{Cu}^{2+}_{(aq)} + \overset{0}{Al}_{(s)} \rightarrow \overset{0}{Cu}_{(s)} + \overset{+3}{Al}^{3+}_{(aq)}$$

$$3Cu^{2+}_{(aq)} + \cancel{6e^-} \rightarrow 3Cu_{(s)} \qquad \text{x 3}$$

$$2Al_{(s)} \rightarrow 2Al^{3+}_{(aq)} + \cancel{6e^-} \qquad \text{x 2}$$

$$3Cu^{2+}_{(aq)} + 2Al_{(s)} \rightarrow 3Cu_{(s)} + 2Al^{3+}_{(aq)}$$

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Example #1

Write a balanced equation to describe the reaction between zinc metal with aqueous lead (II) nitrate.

$$\overset{0}{Zn}_{(s)} + \overset{+2}{Pb}(\overset{+5\ -2}{NO_3})_{2(aq)} \rightarrow \overset{0}{Pb}_{(s)} + \overset{+2}{Zn}(\overset{+5\ -2}{NO_3})_{2(aq)}$$

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

$$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$$

Redox reaction: $Pb^{2+}_{(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)}$

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Example #2

Write the balanced redox equation for the reaction of MnO_4^- in an acidic solution with H_2SO_3 resulting with the products of Mn^{2+} and SO_4^{2-}.

You can use half-reactions from the Half Reactions Table

Reduction: $2MnO_4^-{}_{(aq)} + 16H^+{}_{(aq)} + 10e^- \rightarrow 2Mn^{2+}{}_{(aq)} + 8H_2O_{(l)}$ x 2

Oxidation: $5H_2SO_{3(aq)} + 5H_2O_{(l)} \rightarrow 5SO_4^{2-}{}_{(aq)} + 20H^+{}_{(aq)} + 10e^-$

$2MnO_4^-{}_{(aq)} + 5H_2SO_{3(aq)} \rightarrow 2Mn^{2+}{}_{(aq)} + 3H_2O_{(l)} + 5SO_4^{2-}{}_{(aq)} + 4H^+{}_{(aq)}$

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Sometimes, the half-reaction you need is not part of the given tables. Therefore, you need to come up with the reaction yourself.

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Example #3

Write the balanced equation for the half-reaction where NO is reduced to N_2O in an acidic solution.

1. Write down the reactant and product and balance atoms other than oxygen and hydrogen.
2. Add **H_2O** molecules to balance the oxygen atoms.
3. Add **H+** ions to balance the hydrogen atoms.
4. Add **electrons** to balance the charge on both sides of the equation.

$$NO_{(g)} \rightarrow N_2O_{(g)}$$

Step 1: $$2NO_{(g)} \rightarrow N_2O_{(g)}$$

Step 2: $$2NO_{(g)} \rightarrow N_2O_{(g)} + H_2O_{(l)}$$

Step 3: $$2NO_{(g)} + 2H^+_{(aq)} \rightarrow N_2O_{(g)} + H_2O_{(l)}$$

Step 4: $$2NO_{(g)} + 2H^+_{(aq)} + 2e^- \rightarrow N_2O_{(g)} + H_2O_{(l)}$$

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Example #4

Write the balanced equation for the oxidation half-reaction where Cl_2 is oxidized to ClO_3^- in a basic solution.

1. Write down the reactant and product and balance atoms other than oxygen and hydrogen.
2. Add **H_2O** molecules to balance the oxygen atoms.
3. Add **H+** ions to balance the hydrogen atoms.
4. Add **OH^-** ions equal in number to the H^+ ions to **both sides** of the reaction.
5. Combine H^+ and OH^- ions on the same side of the equation to **form H_2O** molecules.
6. Add **electrons** to balance the charge.

$$Cl_{2(g)} \rightarrow ClO_3^-{}_{(aq)}$$

Step 1: $$Cl_{2(g)} \rightarrow 2ClO_3^-{}_{(aq)}$$

Step 2: $$Cl_{2(g)} + 6H_2O_{(l)} \rightarrow 2ClO_3^-{}_{(aq)}$$

Step 3: $$Cl_{2(g)} + 6H_2O_{(l)} \rightarrow 2ClO_3^-{}_{(aq)} + 12H^+_{(aq)}$$

Step 4: $$12OH^-_{(aq)} + Cl_{2(g)} + 6H_2O_{(l)} \rightarrow 2ClO_3^-{}_{(aq)} + 12H^+_{(aq)} + 12OH^-_{(aq}$$

Step 5: $$12OH^-_{(aq)} + Cl_{2(g)} + 6H_2O_{(l)} \rightarrow 2ClO_3^-{}_{(aq)} + 12H_2O_{(l)}$$

(-12) (-2)

Step 6: $$12OH^-_{(aq)} + Cl_{2(g)} + 6H_2O_{(l)} \rightarrow 2ClO_3^-{}_{(aq)} + 12H_2O_{(l)} + 10e^-$$

$$12OH^-_{(aq)} + Cl_{2(g)} \rightarrow 2ClO_3^-{}_{(aq)} + 6H_2O_{(l)} + 10e^-$$

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For the most part, the complete redox reactions you will need to balance will require you to develop your own half-reactions.

The method used is the ion-electron method or the half-reaction method.

BALANCING REDOX REACTIONS

It is important to know whether the reaction occurs in an acidic or basic environment, as different final products will be produced.

This information will always be provided.

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Example #5

Balance the following redox reaction which occurs in an acidic solution:

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$$

Which half-reaction is reduction? Which is oxidation?

Separate into two skeleton half equations

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

$$Fe^{2+} \rightarrow Fe^{3+}$$

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

$$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

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

Reduction

$$Fe^{2+} \rightarrow Fe^{3+}$$

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

Oxidation

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Example #5

Balance the following redox reaction which occurs in an acidic solution:

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$$

Which half-reaction is reduction? Which is oxidation?

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

$$Fe^{2+} \rightarrow Fe^{3+}$$

Reduction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Oxidation: $6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$ x 6

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$$

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Example #6

Balance the following redox reaction which occurs in a basic solution:

$$SO_3^{2-} + MnO_4^- \rightarrow SO_4^{2-} + MnO_2$$

Which half-reaction is reduction? Which is oxidation?

Separate into two skeleton half equations

$SO_3^{2-} \rightarrow SO_4^{2-}$

$SO_3^{2-} + H_2O \rightarrow SO_4^{2-}$

$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+$

$2OH^- + SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2OH^-$

$2OH^- + SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H_2O$

$2OH^- + SO_3^{2-} \rightarrow SO_4^{2-} + H_2O + 2e^-$

$MnO_4^- \rightarrow MnO_2$

$MnO_4^- \rightarrow MnO_2 + 2H_2O$

$4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$

$MnO_4^- \rightarrow MnO_2$ $4H^+ + MnO_4^- + 4OH^- \rightarrow MnO_2 + 2H_2O + 4OH^-$

$4H_2O + MnO_4^- \rightarrow MnO_2 + 2H_2O + 4OH^-$

$2H_2O + MnO_4^- + 3e^- \rightarrow MnO_2 + 4OH^-$

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Example #6

Balance the following redox reaction which occurs in a basic solution:

$$SO_3^{2-} + MnO_4^- \rightarrow SO_4^{2-} + MnO_2$$

Which half-reaction is reduction? Which is oxidation?

Oxidation: $6OH^- + 3SO_3^{2-} + 3H_2O \rightarrow 3SO_4^{2-} + 6H_2O + \cancel{6e^-}$ **x 3**

Reduction: $8H_2O + 2MnO_4^- + \cancel{6e^-} \rightarrow 2MnO_2 + 4H_2O + 8OH^-$ **x 2**

$$3SO_3^{2-} + H_2O + 2MnO_4^- \rightarrow 3SO_4^{2-} + 2MnO_2 + 2OH^-$$