Chapter 4: Redox Reactions

Section 1: Oxidation and Reduction

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Balancing Redox Equations Using Half-Reactions

- **A species** is any kind of chemical unit involved in a process.
- In the equilibrium equation $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$
- there are four species: the two molecules NH_3 and H_2O and the two ions NH_4^+ and OH^-.
- Oxidation-reduction reactions occur whenever a species that can give up electrons (reducing agent) comes in contact with another species that can accept them (oxidizing agent).
- For example, iron can reduce many species that are oxidizing agents, including chlorine.

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

- In this reaction, each iron atom is oxidized by losing three electrons to become an Fe^{3+} ion.
- At the same time, each chlorine atom in Cl_2 is reduced by accepting one electron to become a Cl^- ion.
- Oxidation: $Fe \rightarrow Fe^{3+} + 3e^-$
- Reduction: $Cl_2 + 2e^- \rightarrow 2Cl^-$
- Equations such as these represent half-reactions.
- **A half-reaction** is one of the two parts of a redox reaction - the oxidation half or the reduction half.
- **Table shows a variety of reduction half-reactions that involve the oxidation of Fe to Fe^{3+}.**

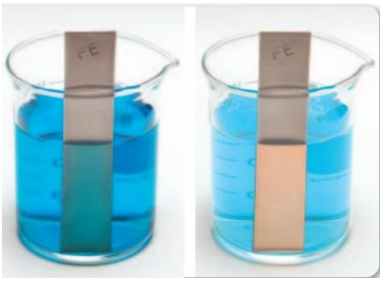
Redox Reactions that Oxidize Iron

Overall Reaction (unbalanced)	Oxidation Half-Reaction	Reduction Half-Reaction
$Fe + O_2 \rightarrow Fe_2O_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$O_2 + 4e^- \rightarrow 2O^{2-}$
$Fe + F_2 \rightarrow FeF_3$		$F_2 + 2e^- \rightarrow 2F^-$
$Fe + HBr \rightarrow H_2 + FeBr_3$		$2H^+ + 2e^- \rightarrow H_2$
$Fe + AgNO_3 \rightarrow Ag + Fe(NO_3)_3$		$Ag^+ + e^- \rightarrow Ag$
$Fe + CuSO_4 \rightarrow Cu + Fe_2(SO_4)_3$		$Cu^{2+} + 2e^- \rightarrow Cu$

How to use half-reactions to balance a redox equation?

- For example, the following unbalanced equation represents the reaction that occurs when you put an iron nail into a solution of copper(II) sulfate, as shown in Figure .

$$Fe(s) + CuSO_4\,(aq) \rightarrow Cu(s) + Fe_2(SO_4)_3\,(aq)$$



1. **Write the net ionic equation for the reaction, omitting spectator ions.** $Fe + Cu^{2+} + SO4^{2-} \rightarrow Cu + 2Fe^{3+} + 3SO_4^{2-}$

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

2. **Write the oxidation and reduction half-reactions for the net ionic equation.**

$Fe \rightarrow 2Fe^{3+} + 6e^-$ $\qquad$ $Cu^{2+} + 2e^- \rightarrow Cu$

3. **Balance the atoms and charges in each half-reaction.**

$2Fe \rightarrow 2Fe^{3+} + 6e^-$ $\qquad$ $Cu^{2+} + 2e^- \rightarrow Cu$

4. **Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction.**

$2Fe \rightarrow 2Fe^{3+} + 6e^-$ $\qquad$ $3Cu^{2+} + 6e^- \rightarrow 3Cu$

5. **Add the balanced half-reactions and return spectator ions.**

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

$2Fe(s) + 3CuSO_4\,(aq) \rightarrow 3Cu(s) + Fe_2(SO_4)_3\,(aq)$

Important Rules

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1- Balance a Redox Equation by Using Half-Reactions.

$KMnO_4(aq) + SO_2(g) \rightarrow MnSO_4(aq) + K_2SO_4(aq)$ (in acid solution)

2- Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

A. $Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)

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B. $Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4^-(aq) + Bi^{2+}(aq)$ (in acid solution)

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C. Challenge $N_2O(g) + ClO^-(aq) \rightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution)

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D. $MnO_4^{2-} + MnO_2 \rightarrow MnO_4^-$ (in basic solution)

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E. $MnO_4^- + IO_3^- \rightarrow MnO_2 + IO_4^-$ (in basic solution)

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F. $MnO_4^- + Cl^- \rightarrow Mn^{2+} + ClO^-$ (in acidic solution)

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H. $ClO_4^-(aq) + Br^-(aq) \rightarrow Cl^-(aq) + Br_2$ (in acidic solution)

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Use the half-reaction method to balance the redox equations.

1- $H_2S(g) + NO_3^-(aq) \rightarrow S(s) + NO(g)$ (in acid solution)

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2- $Sb^{3+} + MnO_4^- \rightarrow SbO_4^{3-} + Mn^{2+}$ (in acid solution)

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3- $Zn + NO_3^- \rightarrow Zn^{2+} + NO_2$ (in acid solution)

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4- Challenge $I^-(aq) + MnO_4^-(aq) \rightarrow I_2(s) + MnO_2(s)$ (in basic solution)

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Revision:

1- Write the oxidation and reduction half-reactions for the redox equation.

$Pb(s) + Pd(NO_3)_2(aq) \rightarrow Pb(NO_3)_2(aq) + Pd(s)$

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2- Determine The oxidation half-reaction of a redox reaction is $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$, and the reduction half-reaction is Au3+ + 3e- → Au. What minimum numbers of tin(II) ions and gold(III) ions would have to react in order to have zero electrons left over?

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3- Apply Balance the following equations.

a. $HClO_3\ (aq) \rightarrow ClO_2\ (g) + HClO_4\ (aq) + H_2O\ (l)$

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b. $H_2SeO_3\ (aq) + HClO_3\ (aq) \rightarrow H_2SeO_4\ (aq) + Cl_2\ (g) + H_2O\ (l)$

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c. $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$ (in acid solution)

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4- Explain what a spectator ion is.

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5- Define the term species as it is used in describing redox reactions.

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6- How can you tell that the equation below is not balanced?

$Fe(s) + Ag^{+}(aq) \rightarrow Fe^{2+}(aq) + Ag(s)$

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7- Balance the redox reaction by half-reaction method .

a. $PbO(s) + NH_3(g) \rightarrow N_2(g) + H_2O(l) + Pb(s)$ (in acid solution)

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a. $I_2(s) + Na_2S_2O_3(aq) \rightarrow Na_2S_2O_4(aq) + NaI(aq)$ (in acid solution)

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b. $Sn(s) + 2HCl(aq) \rightarrow SnCl_2(aq) + H_2(g)$ (in acid solution)

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c. $Cl^-(aq) + NO_3^-(aq) \rightarrow ClO^-(aq) + NO(g)$ (in acid solution)

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e. IO_3^- (aq) + Br^- (aq) → Br_2 (l) + IBr (s) (in acid solution)

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8- Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed.

a. NH_3 (g) + NO_2 (g) → N_2 (g) + H_2O(l) (in acidic solution)

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b. $Br_2 \rightarrow Br^- + BrO_3^-$ (in basic solution)

c. $N_2O + ClO^- \rightarrow Cl^- + NO_2^-$ (in basic solution)