Chapter Map

Classification of substances (Section 5.3)

Formation of binary ionic compounds

Oxidation

Reduction

Formation of molecular compounds

Oxidation-Reduction reactions

Oxidation numbers

Types of chemical reactions

Combination

Combustion

Decomposition

Single displacement

Redox?

What’s oxidized?

What’s reduced?

What’s oxidizing agent?

What’s reducing agent?
Historically, oxidation meant reacting with oxygen.

\[2\text{Zn}(s) + \text{O}_2(g) \rightarrow 2\text{ZnO}(s)\]

\[\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-\]

or

\[2\text{Zn} \rightarrow 2\text{Zn}^{2+} + 4e^-\]

\[\text{O} + 2e^- \rightarrow \text{O}^{2-}\]

or

\[\text{O}_2 + 4e^- \rightarrow 2\text{O}^{2-}\]
Many reactions that are similar to the reaction between zinc and oxygen were not considered oxidation.

For example, both the zinc-oxygen reaction and the reaction between sodium metal and chlorine gas (described on the next slide) involve the transfer of electrons.
Oxidation and Formation of Binary Ionic Compounds

**Formation of NaCl**

Sodium atoms + Chlorine molecule $\rightarrow$ Sodium ions + Chloride ions

**Oxidation of zinc**

Zinc atoms + Oxygen molecule $\rightarrow$ Zinc ions + Oxide ions
Similar to Oxidation of Zinc

\[ 2Na(s) + Cl_2(g) \rightarrow 2NaCl(s) \]

\[ Na \rightarrow Na^+ + e^- \]

or \[ 2Na \rightarrow 2Na^+ + 2e^- \]

\[ Cl + e^- \rightarrow Cl^- \]

or \[ Cl_2 + 2e^- \rightarrow 2Cl^- \]

Oxidation = Loss of Electrons
To include the similar reactions in the same category, *oxidation* was redefined as any chemical change in which at least one element loses electrons.
Zinc Oxide Reduction

- The following equation describes one of the steps in the production of metallic zinc.

\[
\text{ZnO(s)} + \text{C(g)} \rightarrow \text{Zn(s)} + \text{CO(g)}
\]

- Because zinc is reducing the number of bonds to oxygen atoms, historically, zinc was said to be reduced.

- When we analyze the changes taking place, we see that zinc ions are gaining two electrons to form zinc atoms.

\[
\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}
\]

- The definition of reduction was broadened to coincide with the definition of oxidation. According to the modern definition, when something gains electrons, it is reduced.
Reduction

- The loss of electrons (oxidation) by one substance is accompanied by the gain of electrons by another (reduction). **Reduction** is any chemical change in which at least one element gains electrons.
Oxidizing and Reducing Agents

- A **reducing agent** is a substance that loses electrons, making it possible for another substance to gain electrons and be reduced. The oxidized substance is always the reducing agent.

- An **oxidizing agent** is a substance that gains electrons, making it possible for another substance to lose electrons and be oxidized. The reduced substance is always the oxidizing agent.
Identifying Oxidizing and Reducing Agents

2Zn(s) + O_2(g) → 2ZnO(s)

Zn → Zn^{2+} + 2e^{-}

O + 2e^{-} → O^{2-}

- Zinc atoms lose electrons, making it possible for oxygen atoms to gain electrons and be reduced, so zinc is the reducing agent.
- Oxygen atoms gain electrons, making it possible for zinc atoms to lose electrons and be oxidized, so O_2 is the oxidizing agent.
The N-O bond is a polar covalent bond in which the oxygen atom attracts electrons more than the nitrogen atom.

Thus the oxygen atoms gain electrons partially and are reduced.

The nitrogen atoms lose electrons partially and are oxidized.

N$_2$(g) + O$_2$(g) → 2NO(g)

• N$_2$ is the reducing agent.
• O$_2$ is the oxidizing agent.
Redox Terms (1)

- The reducing agent loses electrons and thus is oxidized in the reaction.
- The oxidizing agent gains electrons and thus is reduced in the reaction.

- Complete transfer of electrons
  - Ionic bond
- Partial transfer of electrons
  - Polar covalent bond
• **Oxidation-Reduction Reaction**
  – an electron transfer reaction
• **Oxidation**
  – complete or partial loss of electrons
• **Reduction**
  – complete or partial gain of electrons
• **Oxidizing Agent**
  – the substance reduced; gains electrons, making it possible for something to lose them.
• **Reducing Agent**
  – the substance oxidized; loses electrons, making it possible for something to gain them.
**Questions Answered by Oxidation Numbers**

<table>
<thead>
<tr>
<th>Question</th>
<th>Answer</th>
</tr>
</thead>
<tbody>
<tr>
<td>Is the reaction redox?</td>
<td>If any atoms change their oxidation number, yes.</td>
</tr>
<tr>
<td>What’s oxidized?</td>
<td>The element that increases its oxidation number</td>
</tr>
<tr>
<td>What’s reduced?</td>
<td>The element that decreases its oxidation number</td>
</tr>
<tr>
<td>What’s the reducing agent?</td>
<td>The substance with the element oxidized</td>
</tr>
<tr>
<td>What’s the oxidizing agent?</td>
<td>The substance with the element reduced</td>
</tr>
</tbody>
</table>
Steps for Determination of Oxidation Numbers

- **Step 1:** Assign oxidation numbers to as many atoms as you can using the guidelines described on the next slide.
- **Step 2:** To determine oxidation numbers for atoms not described on the previous slide, use the following guideline.
  - The sum of the oxidation numbers for each atom in the formula is equal to the overall charge on the formula. (This includes uncharged formulas where the sum of the oxidation numbers is zero.)
<table>
<thead>
<tr>
<th>uncharged element</th>
<th>0</th>
<th>no exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>monatomic ions</td>
<td>charge on ion</td>
<td>no exceptions</td>
</tr>
<tr>
<td>combined fluorine</td>
<td>-1</td>
<td>no exceptions</td>
</tr>
<tr>
<td>combined oxygen</td>
<td>-2</td>
<td>-1 in peroxides</td>
</tr>
<tr>
<td>covalently bonded hydrogen</td>
<td>+1</td>
<td>no exceptions</td>
</tr>
</tbody>
</table>
More Types of Chemical Reactions

- Combination
- Decomposition
- Combustion
- Single Displacement
In combination reactions, two or more elements or compounds combine to form one compound.

\[2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s)\]
\[\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)\]
\[\text{MgO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Mg(OH)}_2(s)\]
Decomposition Reactions

• In *decomposition reactions*, one compound is converted into two or more simpler substances.

  Electric current

  \[2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g)\]

  Electric current

  \[2\text{NaCl}(l) \rightarrow 2\text{Na}(l) + \text{Cl}_2(g)\]
• A *combustion reaction* is a redox reaction in which the reaction is very rapid and is accompanied by heat and usually light. The combustion reactions that you will be expected to recognize have oxygen, $O_2$, as one of the reactants.

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$
• When any substance that contains carbon is combusted (or burned) completely, the carbon forms carbon dioxide.

• When a substance that contains hydrogen is burned completely, the hydrogen forms water.

\[
C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)
\]
• The complete combustion of a substance, like ethanol, C$_2$H$_5$OH, that contains carbon, hydrogen, and oxygen also yields carbon dioxide and water.

\[
\text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l)
\]
Combustion Products (3)

- When any substance that contains sulfur burns completely, the sulfur forms sulfur dioxide gas.

\[
\text{CH}_3\text{SH}(g) + 3\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l) + \text{SO}_2(g)
\]
Example

• Write the complete balanced equation for the complete combustion of liquid dimethyl sulfoxide, DMSO, which is a solvent that penetrates the skin without doing damage. It is used as a topical anesthetic and to aid the transfer of pharmaceuticals, such as antifungal compounds, through the skin. The condensed formula CH$_3$SOCH$_3$ can be used to describe DMSO.
Example

- $\text{CH}_3\text{SOCH}_3(\text{l}) + \text{O}_2(\text{g}) \rightarrow$

- $\text{CH}_3\text{SOCH}_3(\text{l}) + \text{O}_2(\text{g})$
  $\rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g})$

- $\text{CH}_3\text{SOCH}_3(\text{l}) + 4\text{O}_2(\text{g})$
  $\rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g})$
Single Displacement

Pure element displaces element in compound

\[ \text{Zn}(s) + \text{CuSO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{Cu}(s) \]

\[ \text{Cd}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CdSO}_4(aq) + \text{H}_2(g) \]

\[ \text{Cl}_2(g) + 2\text{NaI}(aq) \rightarrow 2\text{NaCl}(aq) + \text{I}_2(s) \]
Single Displacement Reaction

\[ \text{Zn(s)} + \text{CuSO}_4(aq) \rightarrow \]
Single Displacement Reaction Example

Cu^{2+} ions collide with Zn atoms, producing Cu atoms and Zn^{2+} ions.

Sulfate ion, SO_4^{2-}

Copper ions, Cu^{2+}

2e^- + Zn^{2+} ions move into solution.

Zn^{2+} ions

Cu atoms collect on the solid.

Zinc metal

https://preparatorychemistry.com/Zn_CuSO4_Canvas.html
Single Displacement Reaction

Zn(s) + CuSO\(_4\)(aq) \rightarrow ZnSO\(_4\)(aq) + Cu(s)
Zn(s) + Cu\(^{2+}\)(aq) \rightarrow Zn\(^{2+}\)(aq) + Cu(s)

oxidation: Zn(s) \rightarrow Zn\(^{2+}\)(aq) + 2e\(^-\)
reduction: Cu\(^{2+}\)(aq) + 2e\(^-\) \rightarrow Cu(s)

https://preparatorychemistry.com/Zn_CuSO4_Canvas.html
• The system in which two half-reactions for a redox reaction are separated allowing the electrons transferred in the reaction to be passed between them through a wire is called **voltaic cell**.
Voltaic Cell

Negative electrode (anode) Zn (−) → Zn^{2+} + 2e^−

<table>
<thead>
<tr>
<th>Electrolyte</th>
<th>Site of Oxidation</th>
</tr>
</thead>
<tbody>
<tr>
<td>SO_4^{2−}</td>
<td>Zn → Zn^{2+} + 2e^−</td>
</tr>
<tr>
<td>NO_3^{−}</td>
<td></td>
</tr>
</tbody>
</table>

Salt Bridge (+)

Positive electrode (cathode) Cu

<table>
<thead>
<tr>
<th>Electrolyte</th>
<th>Site of Reduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu^{2+}</td>
<td>Cu^{2+} + 2e^− → Cu</td>
</tr>
<tr>
<td>SO_4^{2−}</td>
<td></td>
</tr>
</tbody>
</table>

KNO_3
Electrodes

• The electrical conductors placed in the half-cells are called *electrodes*.  
• They can be *active electrodes*, which participate in the reaction, or *passive electrodes*, which transfer the electrons into or out of a half-cell but do not participate in the reaction.
The anode is the site of oxidation. Because oxidation involves loss of electrons, the anode is the source of electrons. For this reason, it is described as the negative electrode. Because electrons are lost forming more positive (or less negative) species at the anode, the surroundings tend to become more positive. Thus anions are attracted to the anode.
• The *cathode* is the site of reduction.
• By convention, the cathode is the positive electrode.
• Because electrons come to the cathode and substances gain these electrons to become more negative (or less positive), the surroundings tend to become more negative. Thus cations are attracted to the cathode.
• A device called a *salt bridge* can be used to keep the charges balanced.

• The portion of the electrochemical cell that allows ions to flow is called the *electrolyte*. 
Leclanché Cell or Dry Cell

Anode oxidation:
\[ \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^- \]

Cathode reduction:
\[ 2\text{MnO}_2(s) + 2\text{NH}_4^+(aq) + 2e^- \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{NH}_3(aq) + \text{H}_2\text{O}(l) \]

Overall reaction:
\[ \text{Zn}(s) + 2\text{MnO}_2(s) + 2\text{NH}_4^+(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Mn}_2\text{O}_3(s) + 2\text{NH}_3(aq) + \text{H}_2\text{O}(l) \]
Dry Cell Image

- Carbon Rod (cathode)
- Porous Barrier (Zinc ions, Zn\(^{2+}\), pass through)
- Paste with MnO\(_2\), NH\(_4\)Cl, and ZnCl\(_2\)
- Zinc (anode)
Anode oxidation:
\[ \text{Zn}(s) + 2\text{OH}^- (aq) \rightarrow \text{ZnO}(s) + \text{H}_2\text{O}(l) + 2\text{e}^- \]

Cathode reduction:
\[ 2\text{MnO}_2(s) + \text{H}_2\text{O}(l) + 2\text{e}^- \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{OH}^- (aq) \]

Overall reaction:
\[ \text{Zn}(s) + 2\text{MnO}_2(s) \rightarrow \text{ZnO}(s) + \text{Mn}_2\text{O}_3(s) \]
Electrolysis

- **Voltage**, a measure of the strength of an electric current, represents the force that moves electrons from the anode to the cathode in a voltaic cell.

- When a greater force (voltage) is applied in the opposite direction, electrons can be pushed from what would normally be the cathode toward the voltaic cell’s anode. This process is called *electrolysis*.

- In a broader sense, electrolysis is the process by which a redox reaction is made to occur in the nonspontaneous direction.

\[
2\text{NaCl}(l) \rightarrow 2\text{Na}(l) + \text{Cl}_2(g)
\]
Primary and Secondary Batteries

- Batteries that are not rechargeable are called **primary batteries**.
- A rechargeable battery is often called a **secondary battery** or a **storage battery**.
Nickel-Cadmium Battery

Anode reaction:
\[ \text{Cd}(s) + 2\text{OH}^- (aq) \rightleftharpoons \text{Cd(OH)}_2(s) + 2\text{e}^- \]

Cathode reaction:
\[ \text{NiO(OH)}(s) + \text{H}_2\text{O(l)} + \text{e}^- \rightleftharpoons \text{Ni(OH)}_2(s) + \text{OH}^- (aq) \]

Net Reaction:
\[ \text{Cd}(s) + 2\text{NiO(OH)}(s) + 2\text{H}_2\text{O(l)} \rightleftharpoons \text{Cd(OH)}_2(s) + 2\text{Ni(OH)}_2(s) \]
Lead Acid Battery

\[
Pb(s) + HSO_4^-(aq) + H_2O(l) \quad \Leftrightarrow \quad PbSO_4(s) + H_3O^+(aq) + 2e^- \]

**Cathode reaction:**

\[
PbO_2(s) + HSO_4^-(aq) + 3H_3O^+(aq) + 2e^- \quad \Leftrightarrow \quad PbSO_4(aq) + 5H_2O(l) \]

**Net reaction:**

\[
Pb(s) + PbO_2(s) + 2HSO_4^-(aq) + 2H_3O^+(aq) \quad \Leftrightarrow \quad 2PbSO_4(s) + 4H_2O(l) \]
Lithium Batteries

• Because lithium metal has a very low density, and because lithium can yield high voltages in batteries, lithium batteries have very high energy to mass ratios (or energy to volume ratios), making them ideal for powering electronic devices.

• The original lithium batteries used metallic lithium. The following slide shows a typical lithium-metal battery.
Lithium-Metal Battery

- Anode:
  - Li
  - Li$^+$ + e$^-$
  - Lithium metal

- Electrolyte:
  - Li$^+$

- Cathode:
  - MnO$_2$ + nLi$^+$ + ne$^-$ → Li$_n$MnO$_2$
  - Manganese dioxide and carbon

- Current collector:

Chemical equations:

- nLi → nLi$^+$ + ne$^-$
- Li + MnO$_2$ → Li$_n$MnO$_2$
Lithium-Metal Battery (2)

- At the anode, lithium metal is oxidized to lithium ions, which migrate through the electrolyte to the cathode.
- At the cathode, lithium ions are inserted between manganese dioxide-carbon layers. Electrons must be gained at the cathode to maintain electrical neutrality.
Problems with Lithium-Metal Batteries

- Lithium metal is very reactive, making lithium metal batteries somewhat dangerous.
- For this reason, modern lithium batteries are more likely to be lithium-ion batteries.
Advantages of Lithium-Ion Batteries

- They are safe.
- Although lithium-ion batteries have a lower energy density than lithium-metal batteries, it is still high compared to other batteries...about twice the standard nickel-cadmium battery.
- Low self discharge
- Low maintenance
Limitations of Lithium-Ion Batteries

- Requires a protection circuit to maintain voltage and currents within safe limits.
- Expensive to manufacture.
- Subject to aging. A typical lifetime is two to three years.
Differences in Lithium-Ion Batteries

• The technology for lithium-ion batteries is constantly evolving, leading to changes in the cathode, anode, and electrolyte.

• One example of a lithium-ion battery is shown on the next slide.
Lithium-Ion Battery

The diagram illustrates the process of charging and discharging a lithium-ion battery. During discharge, lithium ions (Li⁺) are moved from the anode (lithiated graphite) to the cathode (cobalt dioxide) via the electrolyte. The overall reaction is:

\[ \text{Li}_n\text{C}_6 + \text{CoO}_2 \rightarrow 6\text{C} + \text{Li}_n\text{CoO}_2 \]

During charge, the process is reversed, with lithium ions moving back from the cathode to the anode.

\[ \text{nLi}^+ + 6\text{C} + \text{ne}^- \rightarrow \text{Li}_n\text{C}_6 \]

\[ \text{CoO}_2 + \text{nLi}^+ + \text{ne}^- \rightarrow \text{Li}_n\text{CoO}_2 \]
At the anode, lithium ions are trapped between graphite layers.

During discharge, lithium ions move from the anode to the cathode through the electrolyte.

When lithium ions are removed from the anode, electrons are lost to maintain electrical neutrality.

When lithium ions move to the cathode, they are inserted into layers of cobalt dioxide.

Electrons must be gained at the cathode to maintain electrical neutrality.