# Characterizing Oxidation and Reduction

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxidation</td>
<td>Loss of electrons</td>
</tr>
<tr>
<td>Reduction</td>
<td>Gain of electrons</td>
</tr>
<tr>
<td>Redox Reaction</td>
<td>Reaction in which e(^{-}) are gained by 1 atom or ion and lost by another atom or ion</td>
</tr>
<tr>
<td>Oxidizing agent</td>
<td>Reactant that gains e(^{-}) and thus oxidizes other reactant</td>
</tr>
<tr>
<td>Reducing agent</td>
<td>Reactant that donates e(^{-}) and thus reduces another reactant</td>
</tr>
<tr>
<td>Spectator ions</td>
<td>Ions present on both sides of equation and do NOT change</td>
</tr>
<tr>
<td>Net ionic equation</td>
<td>Ionic equation in which spectator ions are REMOVED</td>
</tr>
</tbody>
</table>

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**Identify the reactant that is oxidized, reduced, is the oxidizing agent, is the reducing agent.**

### Table 3: Relative Strength of Oxidizing and Reducing Agents

<table>
<thead>
<tr>
<th>Reducing agents</th>
<th>Oxidizing agents</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li^+</td>
<td>Li</td>
</tr>
<tr>
<td>K^+</td>
<td>K</td>
</tr>
<tr>
<td>Ca^2+</td>
<td>Ca</td>
</tr>
<tr>
<td>Na^+</td>
<td>Na</td>
</tr>
<tr>
<td>Mg^2+</td>
<td>Mg</td>
</tr>
<tr>
<td>Al^3+</td>
<td>Al</td>
</tr>
<tr>
<td>Zn^2+</td>
<td>Zn</td>
</tr>
<tr>
<td>Cr^3+</td>
<td>Cr</td>
</tr>
<tr>
<td>Fe^3+</td>
<td>Fe</td>
</tr>
<tr>
<td>Ni^2+</td>
<td>Ni</td>
</tr>
<tr>
<td>Sn^2+</td>
<td>Sn</td>
</tr>
<tr>
<td>Pb^2+</td>
<td>Pb</td>
</tr>
<tr>
<td>H_2</td>
<td>H_2O</td>
</tr>
<tr>
<td>H_2S</td>
<td>S</td>
</tr>
<tr>
<td>Cu^2+</td>
<td>Cu</td>
</tr>
<tr>
<td>Cu</td>
<td>Cu^2+</td>
</tr>
<tr>
<td>MnO_2^-</td>
<td>MnO_3^-</td>
</tr>
<tr>
<td>Fe^2+</td>
<td>Fe^3+</td>
</tr>
<tr>
<td>Hg</td>
<td>Hg^2+</td>
</tr>
<tr>
<td>Ag^+</td>
<td>Ag^+</td>
</tr>
<tr>
<td>NO_3^-</td>
<td>NO_2^-</td>
</tr>
<tr>
<td>Br^-</td>
<td>Br^2-</td>
</tr>
<tr>
<td>Mn^2+</td>
<td>MnO_2^-</td>
</tr>
<tr>
<td>SO_3^-</td>
<td>H_2SO_4(conc.)</td>
</tr>
<tr>
<td>Cr^3+</td>
<td>CrO_4^-</td>
</tr>
<tr>
<td>Cl^-</td>
<td>Cl^-</td>
</tr>
<tr>
<td>Mn^3+</td>
<td>MnO_3^-</td>
</tr>
<tr>
<td>F^-</td>
<td>F^-</td>
</tr>
</tbody>
</table>

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a) \(\text{Zn(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Cu(s)} + \text{Zn}^{2+}(aq)\)

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

\[\text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu(s)} \text{reduction}\]

**oxidizing agent** \(\text{Cu}^{2+}(aq)\)

**reducing agent** \(\text{Zn(s)}\)

b) \(\text{FeSO}_4(aq) + \text{Cr(s)} \rightarrow \text{CrSO}_4(aq) + \text{Fe(s)}\)

\[\text{Fe}^{2+}(aq) + 2e^- \rightarrow \text{Fe(s)} \text{reduction}\]

\[\text{Cr(s)} \rightarrow \text{Cr}^{3+}(aq) + 2e^- \text{oxidation}\]

**Oxidizing agent** \(\text{Cr(s)}\)

**Reducing agent** \(\text{Fe}^{2+}(aq)\)
Redox Reactions Involving Ionic Compounds

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Half-Reaction</td>
<td>Equation that shows changes in only the compound that is oxidized or reduced.</td>
</tr>
</tbody>
</table>

**STEPS**
1. Identify element oxidized and reduced
2. Make sure amount of $e^-$ gained and lost equal. If not multiply by WNR to balance $e^-$

For example:
Write the redox half-reactions for the following equation. Make sure to identify the oxidizing and reducing agents.

a) Solid potassium reacts with chlorine gas to produce solid potassium chloride.

$$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$$

$$2K_0 \rightarrow 2K^{+1} + 2e^- \quad \text{oxid}$$

$$Cl_2(0) + 2e^- \rightarrow 2Cl^- \quad \text{red}$$

b) $3Zn(s) + Fe_2(SO_4)_{3(aq)} \rightarrow 3ZnSO_4_{(aq)} + 2Fe(s)$

$$3Zn_0 \rightarrow 3Zn^{+2}_{(aq)} + 6e^- \quad \text{oxid}$$

$$2Fe^{+3}_{(aq)} + 6e^- \rightarrow 2Fe \quad \text{red}$$
Balancing Equations in Acidic Solutions

Example:
Sulphur is oxidized by nitric acid in an aqueous solution, producing sulphur dioxide, nitrogen monoxide, and water, as shown by the unbalanced equation. Use the half-reaction method to balance the following equation:

\[ \text{S}_8(s) + \text{HNO}_3(aq) \rightarrow \text{SO}_2(g) + \text{NO}(g) + \text{H}_2\text{O}(l) \]

**Step 1:** Write oxidation and reduction half-reactions

- \[ \text{S}_8(s) \rightarrow \text{SO}_2(g) \]
- \[ \text{NO}_3^-(aq) \rightarrow \text{NO}(g) \]

**Step 2:** Balance any atoms other than oxygen or hydrogen

- \[ 8 \text{S}_8(s) \rightarrow 8 \text{SO}_2(g) \]
- \[ \text{NO}_3^-(aq) \rightarrow \text{NO}(g) + 2 \text{H}_2\text{O}(l) \]

**Step 3:** Balance any oxygen atoms by adding \( \text{H}_2\text{O} \) molecules

\[ 16 \text{H}_2\text{O}(l) + 8 \text{S}_8(s) \rightarrow 8 \text{SO}_2(g) + 32 \text{H}^+(aq) \]

**Step 4:** Balance any hydrogen atoms by adding hydrogen ions, \( \text{H}^+(aq) \)

\[ 16 \text{H}_2\text{O}(l) + 8 \text{S}_8(s) \rightarrow 8 \text{SO}_2(g) + 32 \text{H}^+(aq) \]

**Step 5:** Balance the charges by adding electrons, \( e^- \)

\[ 32e^- + 16 \text{H}_2\text{O}(l) + 8 \text{S}_8(s) \rightarrow 8 \text{SO}_2(g) + 32 \text{H}^+(aq) \]

**Step 6:** Determine lowest common multiple, LCM, so \( e^- \) will cancel in half-rx'n.

\[ 48 \text{H}_2\text{O}(l) + 32 \text{S}_8(s) \rightarrow 74 \text{SO}_2(g) + 16 \text{H}^+(aq) + 96e^- \]

**Step 7:** Multiply by LCM so \( e^- \) will cancel and add half-rx'n together.

\[ 32 \text{H}^+(aq) + 32 \text{S}_8(s) + 32 \text{NO}_3^-(aq) \rightarrow 24 \text{SO}_2(g) + 32 \text{NO}(g) + 16 \text{H}_2\text{O}(l) \]
Balancing Equations in Basic Solutions

Example:
Cyanide, CN\(^{\text{aq}}\), is oxidized by permanganate, MnO\(_4^{\text{aq}}\), in a basic solution, as shown in the following unbalanced equation. Use the half-reaction method to balance the equation:

\[
\text{CN}^{\text{aq}} + \text{MnO}_4^{\text{aq}} \rightarrow \text{CNO}^{\text{aq}} + \text{MnO}_2(s)
\]

Step 1: Write unbalanced oxidation & reduction half-rx'ns.

\[
\begin{align*}
\text{MnO}_4^{\text{aq}} & \rightarrow \text{MnO}_2(s) \\
\text{CN}^{\text{aq}} & \rightarrow \text{CNO}^{\text{aq}}
\end{align*}
\]

Step 2: Balance any atoms other than oxygen and hydrogen.

Step 3: Balance oxygen atoms by adding H\(_2\)O molecules.

Step 4: Balance hydrogen atoms by adding hydrogen ions, H\(^{+\text{(aq)}}\).

Step 5: Adjust for basic conditions by adding same # if hydroxide ions, OH\(^{\text{-aq}}\) as H\(^{+\text{(aq)}}\) to both sides.

Step 6: Combine H\(^{+}\) and OH\(^{-}\) to make water and cancel water molecules on both sides.

Step 7: Balance charges by adding e\(^{-}\).

Step 8: Determine LCM and multiply necessary half-rx'n.

Step 9: Add half-rx'ns and cancel e\(^{-}\) and any other identical molecules.
Redox for Molecular Compounds

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxidation Number</td>
<td>The charge of an atom once it has gained or lost electrons to satisfy a stable octet</td>
</tr>
</tbody>
</table>

Table 9.3 Rules for Assigning Oxidation Numbers.

<table>
<thead>
<tr>
<th>Rules</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. A pure element has an oxidation number of 0.</td>
<td>Na in Na₂Br, Br in Br₂, and P in P₄ all have an oxidation number of 0.</td>
</tr>
<tr>
<td>2. The oxidation number of an element in a monatomic ion equals the charge of the ion.</td>
<td>The oxidation number of Al in Al²⁺ is +3. The oxidation number of Se in Se²⁻ is -2.</td>
</tr>
<tr>
<td>3. The oxidation number of hydrogen in its compounds is +1, except in metal hydrides, where the oxidation number of hydrogen is -1.</td>
<td>The oxidation number of H in H₂S or CH₄ is +1. The oxidation number of H in NaH or in CaH₂ is -1.</td>
</tr>
<tr>
<td>4. The oxidation number of oxygen in its compounds is usually -2, but there are exceptions. These include peroxides, such as H₂O₂, and the compound OF₂.</td>
<td>The oxidation number of O in Li₂O or in KNO₃ is -2.</td>
</tr>
<tr>
<td>5. In covalent compounds that do not contain hydrogen or oxygen, the more electronegative element is assigned an oxidation number that equals the negative charge it usually has in its ionic compounds.</td>
<td>The oxidation number of Cl in PCl₃ is -1. The oxidation number of S in CS₂ is -2.</td>
</tr>
<tr>
<td>6. The sum of the oxidation numbers of all the elements in a compound is 0.</td>
<td>In CF₄, the oxidation number of F is -1, and the oxidation number of C is +4. (4) + (4) = 0</td>
</tr>
<tr>
<td>7. The sum of the oxidation numbers of all the elements in a polyatomic ion equals the charge on the ion.</td>
<td>In NO₂⁻, the oxidation number of O is -2, and the oxidation number of N is +3. (3) + (2) = -1</td>
</tr>
</tbody>
</table>

Assign oxidation numbers to:

a) SiBr₄(1)
\[ Si + 4(-1) = 0 \]
\[ Si = +4 \]

b) HClO₄(aq)
\[ +1 + Cl + 4(-2) = 0 \]
\[ Cl = +7 \]

c) Cr₂O₇²⁻(aq)
\[ 2Cr + 7(-2) = -7 \]
\[ 2Cr = +12 \]

\[
\begin{align*}
\text{d) Fe₃O₄(s)} & \\
3Fe + 4(-2) &= 0 \\
3Fe &= +8 \\
Fe &= +\frac{8}{3}
\end{align*}
\]

e) H₆C₃O(1)
\[
\begin{align*}
\text{e) H₆C₃O(1)} & \\
0 &= 0 \\
\end{align*}
\]
Is it a Redox Reaction?

In order for a reaction to be a redox reaction, it must involve both the gain and loss of electrons. That is one reactant must be oxidized and the other reactant reduced.

Identify if the following equations are redox reactions or not. In order to do this you must be able to show that one reactant gains electrons and the other reactant loses electrons.

a) \[ CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g) \]

\[ \text{C is oxidized and Cl is reduced.} \]

b) \[ CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g) \]

c) \[ 3HNO_2(aq) \rightarrow HNO_3(aq) + 2NO(g) + H_2O(l) \]

\[ N \text{ is being oxidized and reduced it is a redox reaction.} \]
Balancing Equations Using the Oxidation # Method

Example:
The dichromate ion reacts with ethanol in an acidic solution to produce the chromium(III) ion and carbon dioxide. Write a balanced equation for the reaction using the oxidation number method.

Step 1: Write an unbalanced equation from the given information.

\[ \text{Cr}_2\text{O}_7^{\text{-2}}(\text{aq}) + \text{C}_2\text{H}_5\text{OH}(\text{l}) \rightarrow \text{Cr}^{\text{3+}}(\text{aq}) + \text{CO}_2(g) \]

Step 2: Assign an oxidation number to each atom in the eq’n & determine whether it’s redox.

\[ \begin{align*}
\text{Cr} & : 6 \rightarrow +6 \\
\text{O} & : 2 \rightarrow 0 \\
\text{C} & : 4 \rightarrow -2 \\
\text{H} & : 1 \rightarrow +1 \\
\text{O} & : 2 \rightarrow 0 \\
\end{align*} \]

Step 3: Identify the atom(s) that ↑ in oxidation # or atom(s) that ↓ in oxidation #.

\[ \text{Cr} \downarrow \quad \text{C} \uparrow \]

Step 4: Determine numerical values of ↑ & ↓ in oxidation #.

\[ \text{Cr} \downarrow 3 \quad \text{C} \uparrow 6 \]

Step 5: Determine LCM of ↑ & ↓ in oxidation #

\[ 6 \]

Step 6: Multiply by LCM to atoms oxidized and reduced in order to balance loss/gain of e–

\[ 6\left[ \text{Cr}_2\text{O}_7^{\text{-2}}(\text{aq}) + 2\text{C}_2\text{H}_5\text{OH}(\text{l}) \rightarrow 2\text{Cr}^{\text{3+}}(\text{aq}) + 2\text{CO}_2(g) + 11\text{H}_2\text{O}(\text{aq}) \right] \]

Step 7: Balance # of atoms of all elements by inspection unless solution is acidic or basic.

Step 8: If acidic/basic, then include water molecules, H\text{+}(\text{aq}), \text{OH}^{-}(\text{aq})
Galvanic Cells

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Galvanic Cell</td>
<td>A device that uses redox rx'n to transform chemical energy into electrical</td>
</tr>
<tr>
<td></td>
<td>energy.</td>
</tr>
<tr>
<td>Salt bridge</td>
<td>An electrical connection between half-cells that contains an electrolyte</td>
</tr>
<tr>
<td></td>
<td>solution, allowing a current to flow but preventing contact between the</td>
</tr>
<tr>
<td></td>
<td>oxidizing agent and the reducing agent.</td>
</tr>
<tr>
<td>Electrode</td>
<td>A conductor that carries electric current into and out of electrochem cells</td>
</tr>
<tr>
<td>Electrolyte</td>
<td>A substance that, when dissolved in water conducts electricity.</td>
</tr>
<tr>
<td>Anode</td>
<td>The electrode at which oxidation occurs</td>
</tr>
<tr>
<td>Cathode</td>
<td>The electrode at which reduction occurs</td>
</tr>
</tbody>
</table>

Example:
Sketch a galvanic cell based on the half-reactions below.

\[
\begin{align*}
\text{Anode:} & & \text{Cathode:} \\
\text{Al}^{3+} (aq) + 3e^- & \rightarrow & \text{Al} (s) \\
\text{Ni}^{2+} (aq) + 2e^- & \rightarrow & \text{Ni} (s)
\end{align*}
\]

a. Label the anode and the cathode.
b. Indicate where oxidation and reduction are occurring.
c. Show the direction of flow of electrons.
d. Show the direction of the movement of ions.
e. Write a balanced ionic equation for the reaction.
In the galvanic cell, chemists use a shorthand method to describe the cell, this method is called cell notation.

\[
\begin{array}{c}
\text{Zn(s)} \mid \text{ZnSO}_4(\text{aq}) \mid \text{CuSO}_4(\text{aq}) \mid \text{Cu(s)}
\end{array}
\]

A single horizontal line, \( | \), represents a phase change i.e. from solid Zn to aq ZnSO\(_4\).

A double horizontal line, \( || \), represents the salt bridge between half cells, \( \text{KNO}_3(\text{aq}) \).

Chemists can then use this notation in combination with a table of standard reduction potentials to calculate the standard cell potential, \( E^\circ_{\text{cell}} \), which we often refer to as voltage of the cell or the electrical potential difference.

\[
E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}
\]

Example:

a) Calculate the \( E^\circ_{\text{cell}} \) potential for the galvanic cell:
\[
2I^-_{(aq)} + Br_2(l) \rightarrow I_2(s) + 2Br^-_{(aq)}
\]

b) Predict whether the reaction will be spontaneous

\[
E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}
\]

\[
= (+1.07) - (+0.54)
\]

\[
= +0.53 \text{V}
\]

\( \Rightarrow E^\circ_{\text{cell}} > 0 \)

\( \therefore \) the cell is spontaneous
**Electrolytic Cells**

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electrolytic Cell</td>
<td>Electrochem cell that uses an external source of energy to drive a</td>
</tr>
<tr>
<td></td>
<td>non-spontaneous redox reaction</td>
</tr>
<tr>
<td>Electrolysis</td>
<td>Process of using electrical energy to drive a non-spontaneous rx'n</td>
</tr>
</tbody>
</table>

**Table 10.2 Galvanic vs Electrolytic Cells**

<table>
<thead>
<tr>
<th>Galvanic</th>
<th>Electrolytic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Spontaneous rx'n, $E_{\text{cell}}^o$ = positive</td>
<td>Non-spontaneous rx'n, $E_{\text{cell}}^o$ = negative</td>
</tr>
<tr>
<td>Anode (negatively charged): zinc electrode</td>
<td>Anode (positively charged): copper electrode</td>
</tr>
<tr>
<td>Cathode (positively charged): copper electrode</td>
<td>Cathode (negatively charged): zinc electrode</td>
</tr>
<tr>
<td>Oxidation (at anode) $\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^-$</td>
<td>Oxidation (at anode) $\text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2e^-$</td>
</tr>
<tr>
<td>Reduction (at cathode) $\text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s)$</td>
<td>Reduction (at cathode) $\text{Zn}^{2+}(aq) + 2e^- \rightarrow \text{Zn}(s)$</td>
</tr>
<tr>
<td>Cell reaction: $\text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s)$</td>
<td>Cell reaction: $\text{Cu}(s) + \text{Zn}^{2+}(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{Zn}(s)$</td>
</tr>
</tbody>
</table>

**Draw the electrolytic cell below**

![Diagram of electrolytic cell with labels and reactions]

**Cell reaction**: $\text{Cu}(s) + \text{Zn}^{2+}(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{Zn}(s)$

$E_{\text{cell}}^o$ = negative
When your electrolytic cell doesn't work!

working car

not working
**Faraday's Law**

- Much of the early work in electrochemistry was performed by British scientist Michael Faraday. He determined that the amount of chemical change that occurs during electrolysis is directly proportional to the amount of electricity that is passed through the cell. Consider the following reduction of iron:
  \[ \text{Fe}^{3+} (aq) + 3 \, \text{e}^- \rightarrow \text{Fe} (s) \]

- The reduction of 1 mol of iron (III) requires 3 moles of electrons. If we required the reduction of 3 moles of iron (III) we would need 9 moles of electrons.

- Faraday was honoured with the unit for 1 mol of electrons, the Faraday \( F \).

- He discovered that the mass of an element produced or consumed at an electrode was directly proportional to the time the cell operated, as long as the current was constant.

\[
\frac{n_e}{F} = \frac{It}{F} = Q = It
\]

\( n_e = \text{number of electrons} \)
\( I = \text{current strength in Ampres} \)
\( F = \text{Faraday's Constant} = 6.02 \times 10^{23} \times 1.6 \times 10^{-19} \text{C} \)

\( F = 96500 \text{ Coulombs/mole} \)

**Question:** How many grams of copper are deposited on the cathode of an electrolytic cell if an electric current of 2.00A is run through a solution of CuSO₄ for 20 minutes?

\[
\sqrt{2} \, n_e = \frac{It}{F} = \frac{(2)(1200)}{96500}
\]

\[
= \frac{(2)(1200)}{96500} \times 0.0249 \text{ mol} \times 5 \times 0.01245 \text{ mol}
\]

\[
= 2.49 \times 10^{-2} \text{ mol} \times m = n \text{ M}
\]

\[\sqrt{2} \, n_e = \frac{It}{F} = \frac{(2)(1200)}{96500} \times 0.0249 \text{ mol} \times 5 \times 0.01245 \text{ mol} \]

\[= 2.49 \times 10^{-2} \text{ mol} \times m = n \text{ M} \]
\[ t = (0.01245)(63.1576) \]
\[ \sqrt{t} = 0.79g \]

\[ \sqrt{8} \cdot 0.79g \text{ of Cu is plated over 20 min using 2.0A} \]