Oxidation-Reduction Reactions

\[ \text{Fe}_2\text{O}_3(s) + 2 \text{Al}(s) \rightarrow 2 \text{Fe}(s) + \text{Al}_2\text{O}_3(s) \]

This is called the thermite reaction

Reactions like this change the ionization state of the reactants. They are said to be oxidized and reduced.

Oxidation-Reduction Reactions

(also called Redox Reactions)

Redox reactions are characterized by ELECTRON TRANSFER between an electron donor and electron acceptor.

Transfer leads to —

1. decrease in number of e\(^{-}\) of some atom = OXIDATION
2. increase in number of e\(^{-}\) of another atom = REDUCTION
Oxidation-Reduction Reactions

\[
Fe_2O_3(s) + 2 \text{Al}(s) \rightarrow 2 \text{Fe}(s) + \text{Al}_2O_3(s)
\]

In the thermite reaction, Fe\(^{3+}\) gains 3 electrons while Al loses three:

\[
Fe^{3+} + 3 \text{e}^- \rightarrow \text{Fe} \\
\text{Al} \rightarrow Al^{3+} + 3 \text{e}^-
\]

Net: \( Fe^{3+} + Al \rightarrow Fe + Al^{3+} \)

REDOX REACTIONS

In the thermite reaction, electrons were added to Fe\(^{3+}\). It is reduced.

Electrons were removed from Al. It is oxidized.

Fe\(^{3+}\) is the oxidizing agent.

Al is the reducing agent.

This is always the case. In a redox reaction, the oxidizing agent is reduced and the reducing agent is oxidized.

To tell which is which, follow the movement of the electrons.
REDOX REACTIONS

Notice that in all redox reactions if something is oxidized then something else is reduced.

\[ \text{Always.} \]

\[ \text{Cu(s)} + 2 \text{Ag}^+(aq) \rightarrow \text{Cu}^{2+}(aq) + 2 \text{Ag(s)} \]

OXIDATION NUMBERS

The oxidation number of an atom in an element or compound is the electric charge that atom has or APPEARS to have.

There are a set of rules used to calculate oxidation number:

1. Each atom in a pure element has ox. no. = 0.
   - Zn, O\(_2\), I\(_2\), S\(_8\)

2. In simple, single element ions, ox. no. = the charge on the ion.
   - -1 for Cl\(^-\) \hspace{1cm} +2 for Mg\(^{2+}\)

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3. F always has ox. no. = -1 in compounds.
4. O usually has ox. no. = -2
   (except in peroxides: in $\text{H}_2\text{O}_2$, O = -1)
5. Halogens have ox. no. of -1 except when
    bound to F or O.
6. Ox. no. of H = +1
   (except when H is associated with a metal as
    in NaH where it is -1)
7. The algebraic sum of oxidation numbers
   = overall charge on the compound

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<table>
<thead>
<tr>
<th>OXIDATION NUMBERS</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH$_3$</td>
</tr>
<tr>
<td>ClO$^-$</td>
</tr>
<tr>
<td>H$_3$PO$_4$</td>
</tr>
<tr>
<td>MnO$_4^-$</td>
</tr>
<tr>
<td>Cr$_2$O$_7^{2-}$</td>
</tr>
<tr>
<td>CaH$_2$</td>
</tr>
</tbody>
</table>

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Recognizing a Redox Reaction

Corrosion of aluminum

\[ 2 \text{Al}(s) + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu}(s) \]

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

- Ox. no. of Al increases as e\(^{-}\) are donated by the metal.
- Therefore, Al is OXIDIZED and is the REDUCING AGENT in this balanced half-reaction.

Notice that the 2 half-reactions add up to give the overall reaction if we use 2 mol of Al and 3 mol of Cu\(^{2+}\).

\[ 2 \text{Al}(s) \rightarrow 2 \text{Al}^{3+}(aq) + 6 \text{e}^- \]

\[ 3 \text{Cu}^{2+}(aq) + 6 \text{e}^- \rightarrow 3 \text{Cu}(s) \]

\[ 2 \text{Al}(s) + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu}(s) \]

Final, net equation is balanced for mass and charge. Balance by “cancelling” electrons.
Recognizing a Redox Reaction

<table>
<thead>
<tr>
<th>Reaction Change</th>
<th>Oxidation</th>
<th>Reduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change in # of oxygens</td>
<td>gain</td>
<td>loss</td>
</tr>
<tr>
<td>Change in # of halogens</td>
<td>gain</td>
<td>loss</td>
</tr>
<tr>
<td>Change in # of electrons</td>
<td>loss</td>
<td>gain</td>
</tr>
</tbody>
</table>

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