Electron transfer is one of the most basic processes that can happen in chemistry. It simply involves the movement of an electron from one atom to another. Many important biological processes rely on electron transfer, as do key industrial transformations used to make valuable products. In biology, for example, electron transfer plays a central role in respiration and the harvesting of energy from glucose, as well as the storage of energy during photosynthesis. In society, electron transfer has been used to obtain metals from ores since the dawn of civilization.

Oxidation state is a useful tool for keeping track of electron transfers. It is most commonly used in dealing with metals and especially with transition metals. Unlike metals from the first two columns of the periodic table, such as sodium or magnesium, transition metals can often transfer different numbers of electrons, leading to different metal ions (e.g., sodium is generally found as $\text{Na}^+$ and magnesium is almost always $\text{Mg}^{2+}$, but manganese could be $\text{Mn}^{2+}$, $\text{Mn}^{3+}$, and so on, as far as $\text{Mn}^{7+}$). Oxidation state is a number assigned to an element in a compound according to some rules. This number enable us to describe oxidation-reduction reactions, and balancing redox chemical reactions. When a covalent bond forms between two atoms with different electronegativities the shared electrons in the bond lie closer to the more electronegative atom:

The **oxidation number** of an atom is the charge that results when the electrons in a covalent bond are assigned to the more electronegative atom and is the charge an atom would possess if the bonding were ionic. In HCl (above) the oxidation number for the hydrogen would be +1 and that of the Cl would be -1.

**Example**

Determine which element is oxidized and which element is reduced in the following reactions (be sure to include the Oxidation State of each):

1. $\text{Zn + 2H}^{+} \rightarrow \text{Zn}^{2+} + \text{H}_2$
2. $\text{2Al + 3Cu}^{2+} \rightarrow \text{2Al}^{3+} + \text{3Cu}$
3. $\text{CO}_3^{2-} + 2\text{H}^{+} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

**SOLUTIONS**

1. Zn is oxidized (Oxidation number: 0 → +2); H$^+$ is reduced (Oxidation number: +1 → 0)
2. Al is oxidized (Oxidation number: 0 → +3); Cu$^{2+}$ is reduced (+2 → 0)
3. This is not a redox reaction because each element has the same oxidation number in both reactants and products: O= -2, H= +1, C= +4.

An atom is oxidized if its oxidation number increases, and an atom is reduced if its oxidation number decreases. The atom that is oxidized is the reducing agent, and the atom that is reduced is the oxidizing agent. (Note: the oxidizing and reducing agents can be the same element or compound).
Oxidation Numbers and Nomenclature

Compounds of the alkali (oxidation number +1) and alkaline earth metals (oxidation number +2) are typically ionic in nature. Compounds of metals with higher oxidation numbers (e.g., tin +4) tend to form molecular compounds.

- In ionic and covalent molecular compounds usually the less electronegative element is given first.
- In ionic compounds the names are given which refer to the oxidation (ionic) state.
- In molecular compounds the names are given which refer to the number of molecules present in the compound.

<table>
<thead>
<tr>
<th>Ionic</th>
<th>Molecular</th>
</tr>
</thead>
<tbody>
<tr>
<td>MgH₂</td>
<td>magnesium hydride</td>
</tr>
<tr>
<td>FeF₂</td>
<td>iron(II) fluoride</td>
</tr>
<tr>
<td>Mn₂O₃</td>
<td>manganese(III) oxide</td>
</tr>
<tr>
<td>H₂S</td>
<td>dihydrogen sulfide</td>
</tr>
<tr>
<td>OF₂</td>
<td>oxygen difluoride</td>
</tr>
<tr>
<td>Cl₂O₃</td>
<td>dichlorine trioxide</td>
</tr>
</tbody>
</table>

An oxidation-reduction (redox) reaction is a type of chemical reaction that involves a transfer of electrons between two species. An oxidation-reduction reaction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron. Redox reactions are common and vital to some of the basic functions of life, including photosynthesis, respiration, combustion, and corrosion or rusting.

Oxidation-Reduction Reaction Examples

Redox reactions are comprised of two parts, a reduced half and an oxidized half, that always occur together. The reduced half gains electrons and the oxidation number decreases, while the oxidized half loses electrons and the oxidation number increases. Simple ways to remember this include the mnemonic devices OIL RIG, meaning "oxidation is loss" and "reduction is gain," and LEO says GER, meaning "loss of e⁻ = oxidation" and "gain of e⁻ = reduced." There is no net change in the number of electrons in a redox reaction. Those given off in the oxidation half reaction are taken up by another species in the reduction half reaction.

The two species that exchange electrons in a redox reaction are given special names. The ion or molecule that accepts electrons is called the oxidizing agent; by accepting electrons it causes the oxidation of another species. Conversely, the species that donates electrons is called the reducing agent; when the reaction occurs, it reduces the other species. In other words, what is oxidized is the reducing agent and what is reduced is the oxidizing agent. (Note: the oxidizing and reducing agents can be the same element or compound, as in disproportionation reactions).
A good example of a redox reaction is the thermite reaction, in which iron atoms in ferric oxide lose (or give up) O atoms to Al atoms, producing $\text{Al}_2\text{O}_3$ (Figure 20.1.1).

$$\text{Fe}_2\text{O}_3(s) + 2\text{Al}(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2\text{Fe}(l)$$

Another example of the redox reaction is the reaction between zinc and copper sulfate.

Example (Figure 20.1.1): Identifying Oxidized Elements

Using the equations from the previous examples, determine what is oxidized in the following reaction.

$$\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$$

**SOLUTION**

The Oxidation State of H changes from +1 to 0, and the Oxidation State of Zn changes from 0 to +2. Hence, Zn is oxidized and acts as the reducing agent.

Example (Figure 20.1.1): Identifying Reduced Elements

What is reduced species in this reaction?

$$\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$$

**SOLUTION**

The Oxidation State of H changes from +1 to 0, and the Oxidation State of Zn changes from 0 to +2. Hence, H+ ion is reduced and acts as the oxidizing agent.
Combination Reactions

Combination reactions are among the simplest redox reactions and, as the name suggests, involves "combining" elements to form a chemical compound. As usual, oxidation and reduction occur together. The general equation for a combination reaction is given below:

\[ A + B \rightarrow AB \]

Example \(\PageIndex{4}\): Combination Reaction

Equation: \(H_2 + O_2 \rightarrow H_2O\)

Calculation: \(0 + 0 \rightarrow (2)(+1) + (-2) = 0\)

Explanation:

In this equation both \(H_2\) and \(O_2\) are free elements; following Rule #1, their Oxidation States are 0. The product is \(H_2O\), which has a total Oxidation State of 0. According to Rule #6, the Oxidation State of oxygen is usually -2. Therefore, the Oxidation State of hydrogen in \(H_2O\) must be +1.

Decomposition Reactions

A decomposition reaction is the reverse of a combination reaction, the breakdown of a chemical compound into individual elements:

\[ AB \rightarrow A + B \]

Example \(\PageIndex{5}\): Decomposition Reaction

Consider the decomposition of water:

\[ H_2O \rightarrow H_2 + O_2 \]

Calculation:

\([(2)(+1) + (-2) = 0 \rightarrow 0 + 0]\)

In this reaction, water is "decomposed" into hydrogen and oxygen. As in the previous example the \(H_2O\) has a total Oxidation State of 0; thus, according to Rule #6 the Oxidation State of oxygen is usually -2, so the Oxidation State of hydrogen in \(H_2O\) must be +1.

Single Replacement Reactions

A single replacement reaction involves the "replacing" of an element in the reactants with another element in the products:

\[ A + BC \rightarrow AB + C \]
Example \(\PageIndex{6}\): Single Replacement Reaction

Equation:

\[
\text{Cl}_2 + \text{Na}\underline{\text{Br}} \rightarrow \text{Na}\underline{\text{Cl}} + \text{Br}_2
\]

Calculation: \(0\) + \((+1)\) + \((-1)\) = \(0\) \(-\) \((+1)\) + \((-1)\) = \(0\) + 0

Explanation: In this equation, Br is replaced with Cl, and the Cl atoms in Cl\(_2\) are reduced, while the Br ion in NaBr is oxidized.

Double Replacement Reactions

A double replacement reaction is similar to a double replacement reaction, but involves "replacing" two elements in the reactants, with two in the products:

\[
\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}
\]

Example \(\PageIndex{7}\): Double Replacement Reaction

The reaction of gaseous hydrogen chloride and iron oxide is a double replacement reaction. Write the expected reaction for this chemistry equation.

Solution

\[
\text{Fe}_2\text{O}_3 + 6\text{HCl} \rightarrow 2\text{FeCl}_3 + 3\text{H}_2\text{O}
\]

In this equation, Fe and H trade places, and oxygen and chlorine trade places.

Combustion Reactions

Combustion reactions almost always involve oxygen in the form of O\(_2\), and are almost always exothermic, meaning they produce heat. Chemical reactions that give off light and heat and light are colloquially referred to as "burning."

\[
\text{C}_x\text{H}_y + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

Although combustion reactions typically involve redox reactions with a chemical being oxidized by oxygen, many chemicals "burn" in other environments. For example, both titanium and magnesium burn in nitrogen as well:

\[
\text{2Ti}_{(s)} + \text{N}_{2(g)} \rightarrow 2\text{TiN}_{(s)}
\]

\[
\text{3 Mg}_{(s)} + \text{N}_{2(g)} \rightarrow \text{Mg}_3\text{N}_2_{(s)}
\]

Moreover, chemicals can be oxidized by other chemicals than oxygen, such as Cl\(_2\) or F\(_2\); these processes are also considered combustion reactions.
Disproportionation Reactions

Disproportionation Reactions: In some redox reactions a single substance can be both oxidized and reduced. These are known as disproportionation reactions, with the following general equation:

\[2A \rightarrow A^{+n} + A^{-n}\]

Where n is the number of electrons transferred. Disproportionation reactions do not need begin with neutral molecules, and can involve more than two species with differing oxidation states (but rarely).

Example \(\PageIndex{8}\): Disproportionation Reaction

Disproportionation reactions have some practical significance in everyday life, including the reaction of hydrogen peroxide, \(H_2O_2\) poured over a cut. This a decomposition reaction of hydrogen peroxide, which produces oxygen and water. Oxygen is present in all parts of the chemical equation and as a result it is both oxidized and reduced. The reaction is as follows:

\[2H_2O_{2}(aq) \rightarrow 2H_2O(l) + O_{2}(g)\]

**Explanation:** On the reactant side, H has an Oxidation State of +1 and O has an Oxidation State of -1, which changes to -2 for the product \(H_2O\) (oxygen is reduced), and 0 in the product \(O_2\) (oxygen is oxidized).

Summary

Oxidation signifies a loss of electrons and reduction signifies a gain of electrons. Balancing redox reactions is an important step that changes in neutral, basic, and acidic solutions. The types of redox reactions: Combination and decomposition, Displacement reactions (single and double), Combustion, Disproportionation. The oxidizing agent undergoes reduction and the reducing agent undergoes oxidation.

References


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