Skills to Develop

• Define three common types of chemical reactions (precipitation, acid-base, and oxidation-reduction)
• Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
• Compute the oxidation states for elements in compounds

Oxidation-Reduction Reactions

Earth’s atmosphere contains about 20% molecular oxygen, $O_2$, a chemically reactive gas that plays an essential role in the metabolism of aerobic organisms and in many environmental processes that shape the world. The term oxidation was originally used to describe chemical reactions involving $O_2$, but its meaning has evolved to refer to a broad and important reaction class known as oxidation-reduction (redox) reactions. A few examples of such reactions will be used to develop a clear picture of this classification.

Some redox reactions involve the transfer of electrons between reactant species to yield ionic products, such as the reaction between sodium and chlorine to yield sodium chloride:

\[
\text{\ce{2Na}(s)+\ce{Cl}_2(g)\rightarrow \ce{2NaCl}(s)}
\]

It is helpful to view the process with regard to each individual reactant, that is, to represent the fate of each reactant in the form of an equation called a half-reaction:

\[
\text{\ce{2Na}(s)\rightarrow \ce{2Na}^+(s)+\ce{2e^-}}
\]
\[
\text{\ce{Cl}_2(g)+\ce{2e^-}\rightarrow \ce{2Cl}^-(s)}
\]

These equations show that Na atoms lose electrons while Cl atoms (in the $\text{Cl}_2$ molecule) gain electrons, the "s" subscripts for the resulting ions signifying they are present in the form of a solid ionic compound. For redox reactions of this sort, the loss and gain of electrons define the complementary processes that occur:

\[
\begin{align}
\textbf{oxidation} &= \text{loss of electrons} \\
\textbf{reduction} &= \text{gain of electrons}
\end{align}
\]

Remembering Oxidation and Reduction

It is common to remember the difference between oxidation and reduction using one of two pneumonic devices:

1. "LeO goes GeR"
   In Greek Mythology, Leo was a lion (like the zodiac sign). And what do lion's do? Well, they roar but that doesn't really fit, so we go with "Ger".
   - Loses electrons = Oxidation
   - Gains electrons = Reduction
2. "Oil Rig"
   Oxidation is loss
   Reduction is gain

In this reaction, then, sodium is oxidized and chlorine undergoes reduction. Viewed from a more active perspective, sodium functions as a reducing agent (reductant), since it provides electrons to (or reduces) chlorine. Likewise, chlorine functions as an oxidizing agent (oxidant), as it effectively removes electrons from (oxidizes) sodium.

Some redox processes, however, do not involve the transfer of electrons. Consider, for example, a reaction similar to the one yielding NaCl:

\[
\ce{H2(g) + Cl2(g) -> 2HCl(g)}
\]

The product of this reaction is a covalent compound, so transfer of electrons in the explicit sense is not involved. To clarify
the similarity of this reaction to the previous one and permit an unambiguous definition of redox reactions, a property called *oxidation number* has been defined. The oxidation number (or oxidation state) of an element in a compound is the charge its atoms would possess if the compound was ionic. The following guidelines are used to assign oxidation numbers to each element in a molecule or ion. These rules are hierarchical; if two rules conflict, the rule that is higher up on the list takes precedence:

1. The oxidation number of an atom in an elemental substance (a free element) is zero.
2. The oxidation number of a monatomic ion is equal to the ion’s charge.
3. The sum of oxidation numbers for all atoms in a molecule or polyatomic ion equals the charge on the molecule or ion.
4. Metals always have positive oxidation states when in a compound.
   - Group 1 (hydrogen) is ALWAYS +1
   - Group 2 (beryllium) is ALWAYS +2
5. Oxidation numbers for common nonmetals are usually assigned as follows:
   - Fluorine: always -1
   - Hydrogen: +1 when combined with nonmetals, −1 when combined with metals
   - Oxygen: −2 in most compounds, sometimes −1 (so-called peroxides, $\ce{O2^2-}$), very rarely $-\dfrac{1}{2}$ (so-called superoxides, $\ce{O2-}$), positive values when combined with F (values vary)
   - Halogens: −1 for other halogens except when combined with oxygen or other halogens (positive oxidation numbers in these cases, varying values)
   - Group 6 (oxygen) is -2
   - Group 5 (nitrogen) is -3

Note: The proper convention for reporting charge is to write the number first, followed by the sign (e.g., 2+), while oxidation number is written with the reversed sequence, sign followed by number (e.g., +2). This convention aims to emphasize the distinction between these two related properties.

*It is also important to note that it is possible to have fractional oxidation states.*

Example \(\PageIndex{2}\): Assigning Oxidation Numbers

Follow the guidelines in this section of the text to assign oxidation numbers to all the elements in the following species:

a. \(\ce{H2S}\)
b. \(\ce{SO3^2-}\)
c. \(\ce{Na2SO4}\)

**Solution**

(a) According to guideline 1, the oxidation number for H is +1.

Using this oxidation number and the compound’s formula, guideline 4 may then be used to calculate the oxidation number for sulfur:
\( \text{charge on H}_2\text{S} = 0 = (2 \times +1) + (1 \times x) \)
\( x = 0 - (2 \times +1) = -2 \)

(b) Guideline 3 suggests the oxidation number for oxygen is −2.

Using this oxidation number and the ion’s formula, guideline 4 may then be used to calculate the oxidation number for sulfur:
\( \text{charge on SO}_3^{2-} = -2 = (3 \times -2) + (1 \times x) \)
\( x = -2 - (3 \times -2) = +4 \)

(c) For ionic compounds, it’s convenient to assign oxidation numbers for the cation and anion separately.

According to guideline 2, the oxidation number for sodium is +1.

Assuming the usual oxidation number for oxygen (−2 per guideline 3), the oxidation number for sulfur is calculated as directed by guideline 4:
\( \text{charge on SO}_4^{2-} = -2 = (4 \times -2) + (1 \times x) \)
\( x = -2 - (4 \times -2) = +6 \)

Exercise \( \PageIndex{2} \)

Assign oxidation states to the elements whose atoms are underlined in each of the following compounds or ions:

a. \( \text{KNO}_3 \)

b. \( \text{AlH}_3 \)

c. \( \text{NH}_4^+ \)

d. \( \text{H}_2\text{PO}_4^- \)

Answer a

\( \text{N, +5} \)

Answer b

\( \text{Al, +3} \)

Answer c

\( \text{N, −3} \)

Answer d

\( \text{P, +5} \)

Using the oxidation number concept, an all-inclusive definition of redox reaction has been established. Oxidation-reduction (redox) reactions are those in which one or more elements involved undergo a change in oxidation number.
While the vast majority of redox reactions involve changes in oxidation number for two or more elements, a few interesting exceptions to this rule do exist as shown below). Definitions for the complementary processes of this reaction class are correspondingly revised as shown here:

\[
\begin{align}
\textbf{oxidation} &= \text{increase in oxidation number} \\
\textbf{reduction} &= \text{decrease in oxidation number} \\
\end{align}
\]

Returning to the reactions used to introduce this topic, they may now both be identified as redox processes. In the reaction between sodium and chlorine to yield sodium chloride, sodium is oxidized (its oxidation number increases from 0 in Na to +1 in NaCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl\textsubscript{2} to −1 in NaCl). In the reaction between molecular hydrogen and chlorine, hydrogen is oxidized (its oxidation number increases from 0 in H\textsubscript{2} to +1 in HCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl\textsubscript{2} to −1 in HCl).

Several subclasses of redox reactions are recognized, including combustion reactions in which the reductant (also called a fuel) and oxidant (often, but not necessarily, molecular oxygen) react vigorously and produce significant amounts of heat, and often light, in the form of a flame. Solid rocket-fuel reactions such as the one depicted below are combustion processes. A typical propellant reaction in which solid aluminum is oxidized by ammonium perchlorate is represented by this equation:

\[
\ce{10Al(s) + 6NH4ClO4(s) -> 4Al2O3(s) + 2AlCl3(s) + 12H2O(g) + 3N2(g)}
\]

Watch a brief video showing the test firing of a small-scale, prototype, hybrid rocket engine planned for use in the new Space Launch System being developed by NASA. The first engines firing at 3 s (green flame) use a liquid fuel/oxidant mixture, and the second, more powerful engines firing at 4 s (yellow flame) use a solid mixture.
Single-displacement (replacement) reactions are redox reactions in which an ion in solution is displaced (or replaced) via the oxidation of a metallic element. One common example of this type of reaction is the acid oxidation of certain metals:

\[
\ce{Zn}(s)+\ce{2HCl}(aq) \rightarrow \ce{ZnCl2}(aq)+\ce{H2}(g)
\]

Metallic elements may also be oxidized by solutions of other metal salts; for example:

\[
\ce{Cu}(s)+\ce{2AgNO3}(aq) \rightarrow \ce{Cu(NO3)2}(aq)+\ce{2Ag}(s)
\]

This reaction may be observed by placing copper wire in a solution containing a dissolved silver salt. Silver ions in solution are reduced to elemental silver at the surface of the copper wire, and the resulting \(\text{Cu}^{2+}\) ions dissolve in the solution to yield a characteristic blue color (Figure \(\PageIndex{6}\)).

**Figure \(\PageIndex{6}\):** (a) A copper wire is shown next to a solution containing silver(I) ions. (b) Displacement of dissolved silver ions by copper ions results in (c) accumulation of gray-colored silver metal on the wire and development of a blue color in the solution, due to dissolved copper ions. (credit: modification of work by Mark Ott)

Example \(\PageIndex{3}\): Describing Redox Reactions

Identify which equations represent redox reactions, providing a name for the reaction if appropriate. For those reactions identified as redox, name the oxidant and reductant.
a. \( \text{ZnCO}_3(s) \rightarrow \text{ZnO}(s) + \text{CO}_2(g) \)

b. \( \text{2Ga}(l) + \text{3Br}_2(l) \rightarrow \text{2GaBr}_3(s) \)

c. \( \text{2H}_2\text{O}_2(aq) \rightarrow \text{2H}_2\text{O}(l) + \text{O}_2(g) \)

d. \( \text{BaCl}_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + \text{2KCl}(aq) \)

e. \( \text{C}_2\text{H}_4(g) + \text{3O}_2(g) \rightarrow \text{2CO}_2(g) + \text{2H}_2\text{O}(l) \)

**Solution**

Redox reactions are identified per definition if one or more elements undergo a change in oxidation number.

a. This is not a redox reaction, since oxidation numbers remain unchanged for all elements.

b. This is a redox reaction. Gallium is oxidized, its oxidation number increasing from 0 in Ga(l) to +3 in GaBr\(_3(s)\). The reducing agent is Ga(l). Bromine is reduced, its oxidation number decreasing from 0 in Br\(_2(l)\) to −1 in GaBr\(_3(s)\). The oxidizing agent is Br\(_2(l)\).

c. This is a redox reaction. It is a particularly interesting process, as it involves the same element, oxygen, undergoing both oxidation and reduction (a so-called *disproportionation reaction*). Oxygen is oxidized, its oxidation number increasing from −1 in H\(_2\text{O}_2(aq)\) to 0 in O\(_2(g)\). Oxygen is also reduced, its oxidation number decreasing from −1 in H\(_2\text{O}_2(aq)\) to −2 in H\(_2\text{O}(l)\). For disproportionation reactions, the same substance functions as an oxidant and a reductant.

d. This is not a redox reaction, since oxidation numbers remain unchanged for all elements.

e. This is a redox reaction (combustion). Carbon is oxidized, its oxidation number increasing from −2 in C\(_2\text{H}_4(g)\) to +4 in CO\(_2(g)\). The reducing agent (fuel) is C\(_2\text{H}_4(g)\). Oxygen is reduced, its oxidation number decreasing from 0 in O\(_2(g)\) to −2 in H\(_2\text{O}(l)\). The oxidizing agent is O\(_2(g)\).

**Exercise**

This equation describes the production of tin(II) chloride:

\[ \text{Sn}(s) + 2\text{HCl}(g) \rightarrow \text{SnCl}_2(s) + \text{H}_2(g) \]

Is this a redox reaction? If so, provide a more specific name for the reaction if appropriate, and identify the oxidant and reductant.

**Answer**

Yes, a single-replacement reaction. Sn\((s)\) is the reductant, HCl\((g)\) is the oxidant.
Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products. Acid-base reactions involve the transfer of hydrogen ions between reactants. Redox reactions involve a change in oxidation number for one or more reactant elements. Writing balanced equations for some redox reactions that occur in aqueous solutions is simplified by using a systematic approach called the half-reaction method.

**Glossary**

- **combustion reaction**: vigorous redox reaction producing significant amounts of energy in the form of heat and, sometimes, light
- **oxidation**: process in which an element’s oxidation number is increased by loss of electrons
- **oxidation-reduction reaction** *(also, redox reaction)*: reaction involving a change in oxidation number for one or more reactant elements
- **oxidation number** *(also, oxidation state)*: the charge each atom of an element would have in a compound if the compound were ionic
oxidizing agent
(also, oxidant) substance that brings about the oxidation of another substance, and in the process becomes reduced

reduction
process in which an element’s oxidation number is decreased by gain of electrons

reducing agent
(also, reductant) substance that brings about the reduction of another substance, and in the process becomes oxidized

single-displacement reaction
(also, replacement) redox reaction involving the oxidation of an elemental substance by an ionic species

Contributors

• Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

• Adelaide Clark, Oregon Institute of Technology

• Crash Course Chemistry: Crash Course is a division of Complexity and videos are free to stream for educational purposes.

Feedback

Have feedback to give about this text? Click here.

Found a typo and want extra credit? Click here.