Chemistry 110  
Oxidation Reduction Reactions

Oxidation Number  
The oxidation number, or oxidation state, of an atom may be calculated using the following rules.

1. The oxidation number is 0 in any neutral substance that contains atoms of only one element.  
   Examples:  
   The oxidation number of the iron in Fe (a piece of iron metal) is zero.  
   The oxidation number of the oxygen in O_2 is zero.  
   The oxidation number of the hydrogen in H_2 is zero.

2. The sum of the oxidation numbers of all the atoms in a compound is zero. The sum of the  
   oxidation numbers of all the atoms in an ion is the charge on the ion.

3. The oxidation number of oxygen is -2 when it is in most compounds.  
   Examples:  
   The oxidation number of oxygen in H_2O is -2.  
   The oxidation number of oxygen in CH_4O is -2.  
   The oxidation number of oxygen in Al_2O_3 is -2.

4. The oxidation numbers of the metals in Groups 1 and 2 are +1 and +2 respectively in  
   compounds.  
   Examples:  
   The oxidation number of the potassium in KCl is +1.  
   The oxidation number of the calcium in CaCO_3 is +2.

5. The oxidation number of hydrogen is +1 in most compounds.  
   Examples:  
   The oxidation number of the hydrogen in HCl is +1.  
   The oxidation number of the hydrogen in CH_4 is +1.

The oxidation number of other atoms in a compound or ion can usually be calculated using these  
five rules.  
Example 1:  
What is the oxidation number of carbon in CH_4?  
Using rules #2 and #5:  
(oxidation number of carbon) + 4 × (oxidation number of H) = 0  
(oxidation number of carbon) + 4 × (+1) = 0  
(oxidation number of carbon) + 4 = 0  
(oxidation number of carbon) = -4

Example 2:  
What is the oxidation number of carbon in Na_2CO_3?  
Using rules #2, #3, and #4:  
2 × (oxidation number of Na) + (oxidation number of C) + 3 × (oxidation number of O) = 0  
2 × (+1) + (oxidation number of C) + 3 × (-2) = 0  
2 + (oxidation number of C) − 6 = 0  
(oxidation number of C) = +4
Example 3:
What is the oxidation number of nitrogen in $\text{N}_2\text{O}_5$?
Using rules #2 and #3:
\[
2 \times (\text{oxidation number of N}) + 5 \times (\text{oxidation number of O}) = 0
\]
\[
2 \times (\text{oxidation number of N}) + 5 \times (-2) = 0
\]
\[
2 \times (\text{oxidation number of N}) - 10 = 0
\]
\[
2 \times (\text{oxidation number of N}) = 10
\]
\[
\text{(oxidation number of N)} = 5
\]

Example 4:
What is the oxidation number of sulfur in $\text{SO}_4^{2-}$?
Using rules #2 and #3:
\[
(\text{oxidation number of S}) + 4 \times (\text{oxidation number of O}) = -2
\]
\[
(\text{oxidation number of S}) + 4 \times (-2) = -2
\]
\[
(\text{oxidation number of S}) - 8 = -2
\]
\[
(\text{oxidation number of S}) = 6
\]

Practice Problem 1:
What is the oxidation number of chlorine in $\text{Cl}_2$?

Practice Problem 2:
What is the oxidation number of nitrogen in $\text{NH}_3$?

Practice Problem 3:
What is the oxidation number of nitrogen in $\text{KNO}_3$?

Practice Problem 4:
What is the oxidation number of carbon in $\text{C}_2\text{H}_6\text{O}$?

Practice Problem 5:
What is the oxidation number of carbon in $\text{C}_2\text{H}_4\text{O}_2$?

Practice Problem 6:
What is the oxidation number of phosphorus in $\text{PO}_4^{3-}$?

Oxidation and Reduction
In a chemical reaction, an element is oxidized when the oxidation number of the element in a product is higher than its oxidation number in a reactant (oxidation number increases).

In a chemical reaction, an element is reduced when the oxidation number of the element in a product is lower than its oxidation number in a reactant (oxidation number decreases).

Example 1:
In the chemical reaction $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$, what element is oxidized and what element is reduced?

Solution: To answer this question, we must first calculate the oxidation number for each element in each substance using the rules above. So, we first use the rules above to find the oxidation numbers of carbon and hydrogen in $\text{CH}_4$. They are -4 and +1 respectively. Then
we find that the oxidation number of oxygen in \( O_2 \) is zero. Then we find the oxidation numbers of carbon and oxygen in \( CO_2 \). They are +4 and -2. And finally we find the oxidation numbers of hydrogen and oxygen in \( H_2O \). They are +1 and -2.

It is traditional to place the oxidation number of each element above the symbol for the element, as shown here.

\[
\begin{array}{cccccc}
-4 & +1 & 0 & +4 & -2 & +1 & -2 \\
\end{array}
\]

\[\text{CH}_4 + O_2 \rightarrow CO_2 + H_2O\]

We can see that the oxidation number of C increases from -4 to +4 in this reaction, so C is oxidized. We can also see that the oxidation number of O decreases from zero (0) to -2, so O is reduced. Notice that the oxidation number of hydrogen does not change.

It is always the case that if any element is oxidized or reduced, then one element will be oxidized and one element will be reduced. There is never oxidation without reduction. There is never reduction without oxidation. And there is never more than one element oxidized. And there is never more than one element reduced. However, there are chemical reactions in which nothing is oxidized or reduced.

The equations in these practice problems are modified from Microbial Metabolic Diversity – The Tip of the Iceberg, by Dr. Ruth Ann Mikels, May 2008, which states, “Each of the following equations represents a metabolic activity carried out by various species of bacteria.”

In each of these reactions, what element is oxidized and what element is reduced?

Practice Problem 7: \( H_2 + O_2 \rightarrow H_2O \)

Practice Problem 8: \( CO_2 + H_2S \rightarrow C_6H_{12}O_6 + S + H_2O \)

Practice Problem 9: \( NH_4^+ + O_2 \rightarrow NO_2^- + H^+ + H_2O \)

Answers:

1. Cl is 0
2. N is -3, H is +1
3. K is +1, N is +5, O is -2
4. C is -4, H is +1, O is -2
5. C is 0, H is +1, O is -2
6. P is +5, O is -2
7. H is oxidized from 0 to +1. O is reduced from 0 to -2
8. S is oxidized from -2 to 0. C is reduced from +4 to 0.
9. N is oxidized from -3 to +3. O is reduced from 0 to -2.