5.5: Oxidation and Reduction Reactions

The **oxidation number** of an element represents the total number of electrons which have been *removed* (a positive oxidation state) or *added* (a negative oxidation state) to get the element into its present state. The term **oxidation** describes the *loss* of electrons by an element and an *increase* in oxidation state; the term **reduction** describes the *gain* of electrons and a *decrease* in oxidation state. Oxidation numbers for elements in compounds can be calculated using a simple set of rules, which are reproduced below in Table 5.3.

Table 5.3 Rules for Assigning Oxidation Numbers

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1. The oxidation number of an element in the free state is **zero**.
2. A monoatomic ion will have an oxidation number that is equal to its charge.
3. In compounds with metals, hydrogen will be **–1**, otherwise it will always be **+1**.
4. Oxygen, within a compound, will generally have an oxidation number of **–2**.
5. Halogens will be **–1**, except in compounds with oxygen.
6. Sulfur will generally be **–2**, except in compounds with oxygen.
7. In a molecular compound, the most electronegative element is assigned a negative oxidation number

In many chemical reactions, the oxidation number of elements change. Consider the synthesis reaction shown below. In the reactants, carbon and oxygen are both elements and their oxidation numbers are zero (Rule 1). In the product, oxygen will have an oxidation number of –2 (Rule 4), therefore, carbon in CO_2 must have an oxidation number of +4 in order to balance the four negative charges on the oxygens. During this reaction, the oxidation number of carbon has changed from zero in the reactants to +4 in the products and the oxidation number of oxygen has changed from zero to –2. This is an example of a **redox reaction**; **a chemical reaction in which the oxidation numbers of elements change on going from reactants to products.**

$$\mathbf{C}\ (s) + \mathbf{O_2}\ (g) \rightarrow \mathbf{CO_2}\ (g)$$

In a redox reaction, the element that "**loses electrons**" is said to be **oxidized** and will have *an increase in its oxidation number*. In the example above, the oxidation number of carbon *increases* from zero to +4; it has "lost electrons" and has been **oxidized**. The element that "**gains electrons**" in a redox reaction is said to be **reduced** and will have a *decrease in its oxidation number*. In the reaction above, the oxidation number of oxygen has *decreased* from zero to –2; it has "gained electrons" and has been **reduced**.

? Exercise 5.5.1

Arsonic and nitric acids react to form nitrogen monoxide, arsenic acid and water according to the equation shown below. Is this an example of a redox reaction?

$2\ HNO_3\ (aq) + 3\ H_3AsO_3\ (aq) \rightarrow 2\ NO\ (g) + 3\ H_3AsO_4\ (aq) + H_2O\ (l)$

? Exercise 5.5.1

For each of the reactions given below, calculate the oxidation number of each of the elements in the reactants and the products and determine if the reaction involves oxidation-reduction. If it is a redox reaction, identify the elements that have been oxidized and reduced.

a. $Cu_2S \rightarrow 2\ Cu + S$
   Reactants: Cu _____ S _____
   Products: Cu _____ S _____
   Element oxidized: _____ Element Reduced _____
b. $CaCO_3 \rightarrow CaO + CO_2$
   Reactants: Ca _____ C _____ O _____
   Products: Ca _____ C _____ O _____
   Element oxidized: _____ Element Reduced _____

c. $Fe_2O_3 + 3\ H_2 \rightarrow 2\ Fe + 3\ H_2O$
Reactants: Fe _____ O _____ H _____
Products: Fe _____ O _____ H _____
Element oxidized: _____ Element Reduced _____

d. $AgNO_3 + NaCl \rightarrow AgCl\ (s) + NaNO_3$
Reactants: Ag _____ N _____ O _____ Na _____ Cl _____
Products: Ag _____ N _____ O _____ Na _____ Cl _____
Element oxidized: _____ Element Reduced _____
