5.4: Classifying Chemical Reactions

The reactions we have examined in the previous sections can be classified into a few simple types. Organizing reactions in this way is useful because it will assist us in predicting the products of unknown reactions. There are many different classifications of chemical reactions, but here we will focus on the following types: **synthesis**, **decomposition**, **single replacement** and **double replacement**. In addition, we will see that some of these reactions involve changes in the oxidation numbers of the reactants and products; these will be referred to as **oxidation-reduction**, or "**redox**" reactions.

The first type of reaction we will consider is a **synthesis** reaction (also called a *combination* reaction). In a synthesis reaction, elements or compounds undergo reaction and combine to form a single new substance. The reaction of sodium metal with chlorine gas to give sodium chloride is an example of a synthesis reaction where both reactants are elements.

$$2\,Na\,(s) + Cl_2\,(g) \rightarrow 2\,NaCl\,(s)$$

In this reaction, sodium metal and chlorine gas have *combined* to yield (synthesize) the more complex molecule, sodium chloride. Another example of a synthesis reaction, where compounds are involved as reactants, is the reaction of the *organic* molecule ethene (a carbon-based molecule – covered in more depth in Chapter 14) with HBr to yield the organic molecule bromoethane.

In this example, two molecular compounds (the organic compound, **ethene**, and hydrogen bromide) have combined to yield (synthesize) the new molecule, **bromoethane**. In a similar manner, synthesis reaction can also involve elements reacting with compounds.

In **decomposition** reactions, a single compound will break down to form two or more new substances. The substances formed can be elements, compounds, or a mixture of both elements and compounds. Two simple examples of decomposition reactions are shown below.

$$Cu_2S\,(s) \rightarrow 2\,Cu\,(s) + S\,(s)$$

$$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$$

In a **single-replacement reaction** (also called a single-displacement reaction) an element and a compound will react so that their elements are switched. In other words, an element will typically displace another element from within a compound. As a general rule, **metals** will replace **metals** in compounds and **non-metals** will typically replace **non-metals**. An example of a single replacement reaction is shown below.

$$Zn\,(s) + CuCl_2\,(s) \rightarrow ZnCl_2\,(s) + Cu\,(s)$$

In this example, elemental zinc has *displaced* the metal, copper, from copper(II) chloride to form zinc chloride and elemental copper. In the reactants, zinc was elemental and in the products, it is present within the compound, zinc chloride. Likewise, copper was present in a compound in the reactants and is elemental in the products.

In another example, iron metal will react with an aqueous solution of copper sulfate to give copper metal and iron(II) sulfate.

$$Fe\,(s) + CuSO_4\,(aq) \rightarrow Cu\,(s) + FeSO_4\,(aq)$$

In this reaction, elemental iron replaces copper in a compound with sulfate anion and elemental copper metal is formed; metal replaces metal. The tendency of metals to replace other metals in single-replacement reactions is often referred to as an activity series. In the activity series shown below, any metal will replace any other metal which is to the right of itself in the series. Thus, iron will replace copper, as shown above, but copper metal would not replace iron from iron sulfate (iron is more active than copper). The activity series is useful in predicting whether a given single-replacement reaction will occur or not. Note that hydrogen is also included in the table, although it is not generally considered a "metal". Within this table, metals which occur before hydrogen will react with acids to form hydrogen gas and metal salts. Copper, silver, mercury and gold are less active than hydrogen and will not liberate hydrogen from simple acids. The metals in Groups I & II of the periodic table (Li, Na, K, Rb, Cs, Ca, Sr and Ba) are so reactive that they will directly react with water to liberate hydrogen gas and form metal hydroxides. These are generally referred to as "**active metals**".

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A **double-replacement** reaction (or double-displacement) two ionic compounds in aqueous solution switch anions and form two new compounds. In order for a chemical reaction to occur in a double-replacement reaction, one of the new compounds that is formed must be *insoluble* in water, forming a solid **precipitate** (or a gas, to be covered in Chapter 6). If both of the new compounds which are formed are water-soluble, then no reaction has occurred. In the reactants, there were two cations and two anions in solution, and in the products there are the *same* two cations and two anions, in the *same* solution; nothing has happened. An example of a double-replacement reaction is shown below.

$$BaCl_2\ (aq) + Na_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + 2\ NaCl\ (aq)$$

In this reaction, solid barium sulfate is formed as a precipitate. This is a chemical change and this is a valid chemical reaction. In order to predict whether a double-replacement reaction will occur or not you need to understand rules for predicting solubility of ionic compounds. These rules will be covered in a later section, but are not related to the activity series discussed above.

Table 5.2 Activity Series for Common Metals

Li >	K >	Ba >	Sr >	Ca >
Na >	Mg >	Al >	Mn >	Zn >
Fe >	Cd >	Co >	Ni >	Sn >
Pb >	H >	Cu >	Ag >	Hg >
Au				