Chapter 9
Making Chemical Changes

One of the most amazing aspects of chemistry is the ability to transform one substance into another. In this chapter, the fundamental ideas of chemical change will be introduced and how they can be represented by chemical equations.

9.1 Background

From the earliest times, humans have relied on chemical change. One of the first examples was using chemical change as a source of heat: that is, the heat released by the combustion of wood to produce carbon dioxide and water. This heat energy was then used to cause the biochemical changes when raw meat was cooked. Among the many subsequent chemical reactions was that to convert metal oxides to metals using heat and carbon, resulting in the Bronze Age and then the Iron Age. Thus civilization itself largely developed through the discovery of chemical change.

Though chemical changes enabled civilization to progress, the study of chemistry itself was subverted into many fruitless attempts to transform common metals into gold or to find a potion which would confer immortality on those who drank it. These early researchers were called alchemists. They performed their experiments in secret and made notes in mystical writings. Though little is known about her, one of the early alchemists who designed useful chemical equipment was Maria the Jewess, who lived in Alexandria, Egypt, possibly about 100 C.E. Among other items, she is believed to have devised the double-boiler, called *bain-marie* in French.

Figure 9.1 A still, one of the pieces of chemical apparatus possibly devised by alchemist, Maria the Jewess, about two thousand years ago

It was in the late 1700s that a group of French scientists rejected the mysticism of alchemy and founded the rational principles of modern chemistry which we use today.

9.2 Chemical Change

Our modern lives would not be possible without the materials produced by chemical reactions. Paints, cosmetics, and the fibres for most clothing are all synthesized by chemical industry. The large majority of our electrical energy is produced by chemical reaction: the combustion of coal, oil, or natural gas.

In Chapter 2, Section 2.5, physical change and chemical change were differentiated. A chemical change is the result of a chemical reaction. There are four ways in which a chemical reaction can be observed:

1. Change in Colour
2. Formation of a Gas
3. Formation of a Solid
4. Release or Absorption of Heat

Though the first three ways are reliable evidence that a chemical reaction has occurred, a change in heat may simply be a result of a physical change, such as a substance dissolving in water. If a chemical reaction causes heat to be released it is said to be ***exothermic***; if heat is absorbed, it is said to be ***endothermic***.

9.3 Writing a Chemical Equation

In Chapter 2, Section 2.8, it was shown how chemical elements can be represented by a name and a symbol. Then in Chapter 7, it was shown how compounds can be similarly represented by combinations of element names or by combinations of element symbols. This Section will build upon these earlier concepts to construct chemical equations.

DERIVING A CHEMICAL EQUATION

A chemical equation is a compact method of displaying concise information about a chemical reaction. The chemical reaction between sodium and chlorine can be used as an example. In this reaction, silvery sodium metal reacts with poisonous yellow-green chlorine gas to give solid transparent colorless crystals of sodium chloride. The substances that combine are called the ***reactants***, while the substance(s) produced are called the ***products***.

Figure 9.2 (left) silvery sodium metal reacts with (centre) yellow-green chlorine gas to give (right) colorless crystals of solid sodium chloride.

As a first step, the reaction can be written as a word equation. This is accomplished by writing the names of the reactants on the left, separated by "+" signs and the name(s) of the products on the right. The reactants and products are separated by an arrow "→". So this particular reaction can be written as the word equation:

Sodium + Chlorine → Sodium Chloride

The next step is to replace the chemical names by the chemical symbols or formulas. Sodium has the formula 'Na'. Chlorine molecules occur as pairs of atoms as was discussed in Chapter 6, Section 6.4, Table 6.1. so is written as 'Cl_2'. Sodium chloride consists of the combination of one-to-one ratio of sodium ions (Na^+) and chloride ions (Cl^-) to give sodium chloride 'NaCl'. This gives the following result:

$Na + Cl_2 \rightarrow NaCl$

However, atoms cannot be created or destroyed in a chemical reaction. On the left side, there are two atoms of chlorine and on the right, only one chloride ion. The equation is said to be unbalanced.

The first step in the balancing process is to provide two chloride ions on the right side of the equation. It is important to realize that a chemical formula cannot be altered for this purpose. For example, sodium chloride always has the formula of NaCl, balancing the +1 charge on the sodium ion with the −1 charge on the chloride ion. Instead, it is necessary to place a '2' in front of the 'NaCl' to indicate that two formula units of sodium chloride are produced.

$Na + Cl_2 \rightarrow 2\ NaCl$

Though the numbers of chlorines are now said to be 'balanced', in the process, the number of sodium atoms on the left and sodium ions on the right has become unbalanced. Placing a '2' in front of the 'Na' then results in a balancing of both elements.

$2\ Na + Cl_2 \rightarrow 2\ NaCl$

Finally, the phase of the substance is added after the formula of each element and compound. The phases are displayed in italic as follows:

Solid – (*s*)
Liquid - (*l*)
Gas – (*g*)
Aqueous (dissolved in water) – (*aq*)

The final complete balanced equation becomes:

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

BALANCING MORE COMPLICATED EQUATIONS

Balancing a chemical equation is usually straightforward, but there are some cases which are more complex. A good example is the combustion of the organic compound, propane, $C_3H_8(g)$, in atmospheric oxygen to give carbon dioxide gas and water (also a gas at that temperature).

Figure 9.3 The combustion of propane gas is the common chemical energy source for campers

First, the word equation is constructed:

Propane + Oxygen $\rightarrow$ Carbon Dioxide + Water

Second, the names of the substances are replaced by the respective formulas:

$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

It is important to leave until last the balancing of any element that appears more than once on either side of the equation. In this case, it would be oxygen. So either carbon or hydrogen should be balanced first. Here, carbon will be chosen. As there are three carbon atoms in a molecule of propane, then there will have to be three molecules of carbon dioxide produced:

$C_3H_8 + O_2 \rightarrow 3\ CO_2 + H_2O$

Thus the next element to be balanced will be hydrogen. To balance the eight atoms of hydrogen in propane, four water molecules will be produced:

$C_3H_8 + O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

Then oxygen can be balanced. On the left, there are two atoms while on the right there is a total of ten oxygen atoms. To balance, five oxygen molecules will be needed:

$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

Everything now balances: three atoms of carbon; eight atoms of hydrogen; and ten atoms of oxygen. The final task is to add the phases:

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

EXAMPLE 9.1

The airbag used in automobiles is inflated by means of a chemical reaction involving the decomposition of a chemical compound by the systematic name of sodium trinitride, $NaN_3(s)$. Write a balanced chemical equation for the reaction.

Answer

First, the word equation can be written. As there are just the two elements in the compound, the product must be the elements themselves.

Sodium trinitride $\rightarrow$ sodium + nitrogen

Sodium is a solid metal, $Na(s)$, while nitrogen is a diatomic gas, $N_2(g)$. So an unbalanced formula equation can be written:

$NaN_3 \rightarrow Na + N_2$

There are three atoms of nitrogen on the left and two on the right. The solution to any odd-even balancing is to first multiply the 'odd' number by two:

$2\ NaN_3 \rightarrow Na + N_2$

To balance, two atoms of sodium will be required as a product.

$2\ NaN_3 \rightarrow 2\ Na + N_2$

And to balance the six atoms of nitrogen on the left, three N_2 molecules of nitrogen will be required as a product.

$2\ NaN_3 \rightarrow 2\ Na + 3\ N_2$

Finally, the phases are added:

$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$

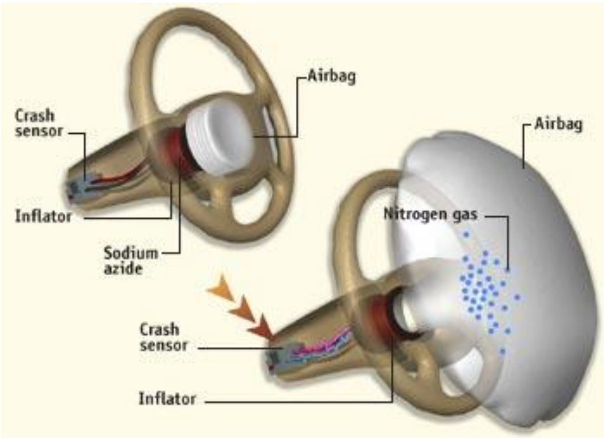

Figure 9.4 In an automobile airbag system, a crash sensor detects a sudden change in speed and sends an electrical signal to the inflator. The inflator generates a spark which ignites the sodium trinitride (azide). The gas produced rapidly inflates the bag.

ARROWS AND CHEMICAL EQUATIONS

In some equations, the arrow separating reactants and products is used to provide additional information about the chemical reaction. For example, some chemical reactions proceed very slowly.

One way to make some chemical reactions occur more quickly is to heat them. To indicate the need for heating, a Greek upper-case delta symbol, Δ, is placed over the reaction arrow. One such reaction is that between iron and oxygen gas to give iron(III) oxide:

$$2\ Fe(s) + 3\ O_2(g) \xrightarrow{\Delta} Fe_2O_3(s)$$

A few chemical reactions can be made to occur more quickly is to use a catalyst. A ***catalyst*** is a substance that alters the rate of a chemical reaction but is, itself, unchanged at the end of the reaction. The chemical formula of the catalyst is placed over the arrow. An example of a catalyzed reaction is that of the breaking-apart of a solution of hydrogen peroxide to give water and oxygen gas. A solution of hydrogen peroxide will stay stable for many years, but if a very small mass of manganese(IV) oxide is added to the solution, bubbles of oxygen gas are produced immediately:

$$2\ H_2O_2(aq) \xrightarrow{MnO_2} 2\ H_2O(l) + O_2(g)$$

Sometimes chemical reactions do not go to completion. Some of the product is formed, but there is still some of the reactants present. To indicate this, a double-headed arrow is used. An example of an equilibrium reaction is that between nitrogen gas and hydrogen gas to form ammonia gas:

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$$

Figure 9.5 In this factory, nitrogen gas and hydrogen gas is combined to form ammonia, a key fertilizer for farmers.

Classes of Chemical Reactions

There are many different types of chemical reactions, but here the focus will be on four types for which the products are predictable. The four types are shown in Figure 9.6 and they will be discussed in more detail in the next four Sections.

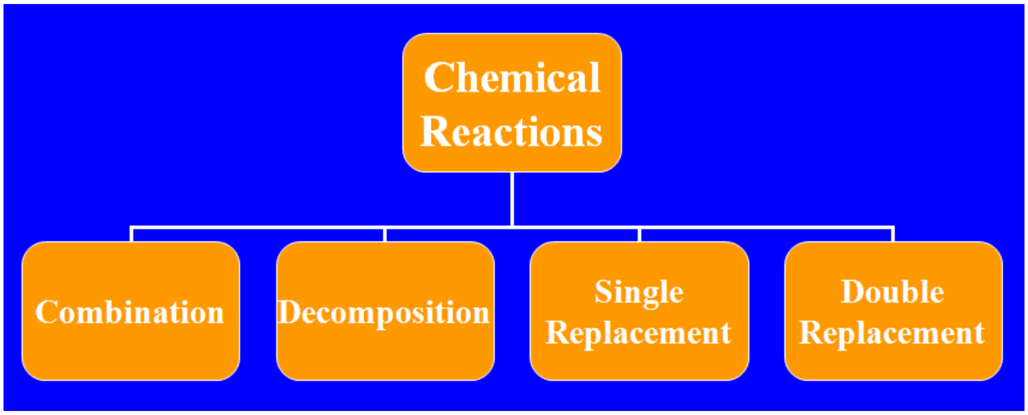

Figure 9.6 The Different Categories of Simple Chemical Reactions

9.4 Combination Reactions

The first type of reaction to be discussed will be combination reactions. ***Combination reactions*** are those in which two substances combine to form a single substance.

Combination Reactions of Elements

The simplest combination reaction is that of two elements. As an example, aluminum metal reacts with solid iodine to give solid aluminum iodide.

$$2\ Al(s) + 3\ I_2(s) \rightarrow 2\ AlI_3(s)$$

Combination reactions can also occur between two nonmetals. For example, solid sulfur burns in oxygen gas to give sulfur dioxide gas:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Figure 9.7 When sulfur burns in oxygen gas, a blue flame is observed as sulfur dioxide gas is formed

EXAMPLE 9.2

Predict the product from the reaction of lithium metal and oxygen gas.

Answer

First, write the reactants in their normal states:

$$Li(s) + O_2(g) \rightarrow$$

Now you must deduce the formula of the product. It is a metal and non-metal combining together, so the ionic rules of constructing a formula apply. Lithium forms the cation, Li^+, and oxygen forms the anion, O^{2-}. To balance the charges, two lithium ions will be needed to balance one oxide ion, so the formula of the product will be Li_2O. This formula can be inserted in the equation:

$$Li(s) + O_2(g) \rightarrow Li_2O$$

Ionic compounds are solid at room temperature, so the phase can be added:

$$Li(s) + O_2(g) \rightarrow Li_2O(s)$$

The equation now needs balancing. We could start by balancing either element. To balance oxygen, the oxygen molecule on the left contains two atoms while there is only one oxide ion on the right. Two formula units of lithium oxide are needed on the right:

$$Li(s) + O_2(g) \rightarrow 2\ Li_2O(s)$$

Finally, four atoms of lithium metal are required on the left to balance the four lithium ions on the right.

$$4\ Li(s) + O_2(g) \rightarrow 2\ Li_2O(s)$$

If the combination reaction involves two non-metals, then you would be given the formula of the product as you would not be able to predict the formula.

COMBINATION REACTIONS THAT PRODUCE ACIDS

In Chapter 2, Section 2.6, one of the categories of compounds was that of an acid. Then in Chapter 7, Section 7.10, the formulas and corresponding names of acids were discussed.

One of those acids was carbonic acid, $H_2CO_3(aq)$. Carbonic acid is formed when carbon dioxide gas dissolves in water. It is the dissolved carbonic acid which makes even ‘normal’ rainwater slightly acidic. It is this acidity of rainwater which has contributed significantly to the erosion of rocks on Earth since the planet was formed.

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

As carbon dioxide reacts to give an acid, carbon dioxide is called an ***acid oxide***. Many nonmetal oxides react similarly with water to give acids.

Example 9.3

Complete the balanced equation for the reaction of sulfur dioxide gas with water.

Answer

The reactants written as formulas are:

$SO_2(g) + H_2O(l) \rightarrow$

In Chapter 7, Section 7.9, you learned of two sulfur-containing acids: sulfurous acid, $H_2SO_3(aq)$; and sulfuric acid, $H_2SO_4(aq)$. Looking at the numbers of atoms of each element in the equation, there are two hydrogen atoms, one sulfur atom, and a total of three oxygen atoms. The product, therefore, must be sulfurous acid, $H_2SO_3(aq)$. So the equation can be completed as:

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

Combination Reactions that Produce Bases

In Chapter 2, Section 2.6, another category of compounds was bases. The simple bases are ionic compounds containing the polyatomic ion, hydroxide, OH^-. Hydroxides can be formed by the reaction of certain metal oxides with water. These metal oxides are called ***basic oxides***. An example is the reaction of solid calcium oxide with water to give a solution of calcium hydroxide. A solution of calcium hydroxide is commonly called limewater.

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

Example 9.4

Deduce the product from the reaction of solid sodium oxide with water.

Answer

First, you need the formula of sodium oxide. Sodium forms a cation, Na^+, while oxygen forms the oxide anion, O^{2-}. The formula of sodium oxide will therefore be Na_2O. The reactant side of the equation will be:

$Na_2O(s) + H_2O(l) \rightarrow$

The product will be sodium hydroxide. The hydroxide ion is OH^- so the formula of sodium hydroxide is NaOH. Assuming excess water is used, a solution of the compound is formed.

The unbalanced equation will be:

$Na_2O(s) + H_2O(l) \rightarrow NaOH(aq)$

Finally, to balance, as the reactant side has two sodiums, two oxygens, and two hydrogens while the right has one of each, then two formula units of sodium hydroxide must be formed.

$Na_2O(s) + H_2O(l) \rightarrow 2\ NaOH(aq)$

9.5 Decomposition Reactions

The opposite of a combination reaction is a decomposition reaction. A ***decomposition reaction*** is a reaction in which a single compound decomposes to give two or more substances. If heat is necessary to cause the decomposition, the reaction is said to be a ***thermal decomposition reaction***.

There are three sub-categories of simple decomposition reactions corresponding to those of the combination reactions. Unfortunately, not all thermal decomposition reactions fit into one of the categories but all those discussed here will do so.

DECOMPOSITION REACTIONS THAT PRODUCE ELEMENTS

A classic example of a decomposition reaction to give elements is that of heating mercury(II) oxide to give mercury metal and oxygen gas. In this reaction, heating the red powdery mercury(II) oxide results in globules of shiny liquid metallic mercury. Unfortunately, the vapour of mercury metal is very poisonous so this reaction is no longer demonstrated.

$$2\ HgO(s) \xrightarrow{\Delta} 2\ Hg(l) + O_2(g)$$

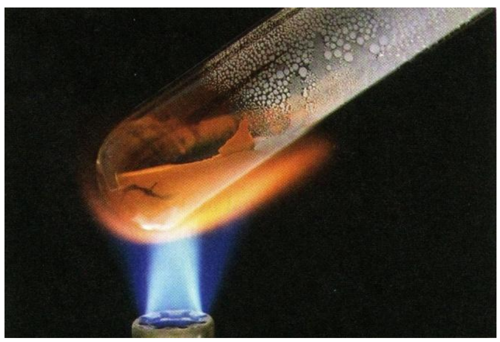

Figure 9.8 Heating solid red-orange mercury(II) oxide causes the compound to decompose to its constituent elements. Globules of liquid mercury can be seen depositing on the wall of the test tube.

EXAMPLE 9.5

Dichlorine monoxide gas decomposes on warming to chlorine gas and oxygen gas. Write the corresponding molecular equation.

Answer

The reactant can be placed in the equation:

$$Cl_2O(g) \xrightarrow{\Delta}$$

Both chlorine gas and oxygen gas are diatomic molecules. So the unbalanced equation can be written as:

$$Cl_2O(g) \xrightarrow{\Delta} Cl_2(g) + O_2(g)$$

To balance the two oxygen atoms on the right, two molecules of dichlorine monoxide must be used:

$$2\ Cl_2O(g) \xrightarrow{\Delta} Cl_2(g) + O_2(g)$$

Finally, to balance the four chlorine atoms on the left, two molecules of chlorine must be produced:

$$2\ Cl_2O(g) \xrightarrow{\Delta} 2\ Cl_2(g) + O_2(g)$$

DECOMPOSITION REACTIONS OF ACIDS

Just as carbon dioxide will dissolve in water to form carbonic acid, so warming a solution of carbonic acid will result in bubbles of carbon dioxide gas (and water). The corresponding chemical equation will be:

$$H_2CO_3(aq) \xrightarrow{\Delta} H_2O(l) + CO_2(g)$$

When a reaction can proceed in either direction, it is called a ***reversible reaction***. The combination of carbon dioxide and water and the corresponding decomposition of carbonic acid is one such example. To indicate a reversible reaction, a double-headed arrow is used. So this particular reaction is sometimes written as:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

EXAMPLE 9.6

Predict the other gaseous product from very strongly heating sulfuric acid:

$$H_2SO_4(aq) \xrightarrow{\Delta} H_2O(g) + ?$$

Answer

Taking two “H” and one “O” out of the acid gives “SO_3”. The equation will therefore be:

$$H_2SO_4(aq) \xrightarrow{\Delta} H_2O(g) + SO_3(g)$$

DECOMPOSITION REACTIONS OF BASES

Just as some acids can decompose, so can some bases. As an example, gently heating blue-green copper(II) hydroxide gives black copper(II) oxide (and water):

$$Cu(OH)_2(s) \xrightarrow{\Delta} CuO(s) + H_2O(l)$$

EXAMPLE 9.7

Write a balanced equation for the heating of solid lithium hydroxide, LiOH(*s*).

Answer

One of the products of heating a hydroxide is water:

$$LiOH(s) \xrightarrow{\Delta} H_2O(g) + ?$$

The other product must contain "Li" and "O."

As lithium ion is Li^+ and oxide is O^{2-}, the other product has to be Li_2O (and inorganic compounds will always be in the solid phase).

$$LiOH(s) \xrightarrow{\Delta} H_2O(g) + Li_2O(s)$$

Finally, the equation needs balancing:

$$2\ LiOH(s) \xrightarrow{\Delta} H_2O(g) + Li_2O(s)$$

9.6 Single Replacement Reactions

A ***single replacement reaction*** is a reaction in which an element reacts with an ionic compound displacing an element from it. That is, a metallic element will react with an ionic compound to give a different combination of metal and ionic compound.

As an example, if a strip of shiny red-brown copper metal is placed in a beaker of colourless silver nitrate solution, silvery metallic crystals of silver form on the copper strip while the solution turns a pale blue colour, the characteristic colour of the copper(II) ion, $Cu^{2+}(aq)$. In the reaction, each atom of copper metal loses two electrons to become the copper(II) ion while each silver ion gains one electron to become a silver atom. The corresponding chemical equation is:

$$Cu(s) + 2\ AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\ Ag(s)$$

From experimentation, it is known that copper metal reacts with silver nitrate solution. However, silver metal does not react with copper(II) nitrate solution. To account for these observations, it was proposed that some elements are more reactive than others. It is a more reactive element which replaces a less-reactive element from its compound. There are two series of such elements, one for metals and one for halogens. The metal series will be discussed first.

THE ACTIVITY SERIES FOR METALS

Figure 9.9 shows the Activity Series for a few relevant metals (the full table would include all metals). The most reactive metals are at the top and the least reactive metals are at the bottom of the Table. Thus copper is above silver in the Table and, as discussed above, copper metal can 'push out' silver ion from its compound. Copper metal can similarly 'push out' gold from its compounds. However, copper metal cannot 'push out' any metal ion above it in the Table.

Sodium
Lithium
Calcium
Magnesium
Aluminum
Zinc
Iron
Tin
Lead
Copper
Silver
Gold

Figure 9.9 Activity Series for some metals.

As another example, zinc metal can replace copper(II) ion from a blue solution of copper(II) sulfate. In this case, red-brown crystals of copper metal will form while the blue colour of the solution will become paler and paler as the blue copper(II) ion is replaced by the colourless zinc ion.

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Example 9.8

Will a solution of aluminum chloride react with copper metal?

Answer

For a single replacement reaction, the element must be higher in the series than the metal ion. On consulting the activity series, you can see that the metal cation, Al^{3+}, in $AlCl_3(aq)$ is higher in the Series than the metal, $Cu(s)$. Therefore there will be no chemical reaction.

Example 9.9

Suppose aluminum metal was placed in a solution of silver nitrate, what would be formed?

Answer

First, you need to write the formulas of the reactants.

Aluminum metal will be $Al(s)$

Silver forms the Ag^+ cation while nitrate is NO_3^-. Thus the formula is $AgNO_3$.

The reactants can be written as:

$$Al(s) + AgNO_3(aq) \rightarrow$$

In a single replacement reaction, the metal and metal ion ‘switch over.’ Thus we can write the formulas of the products:

Silver metal will be $Ag(s)$

Aluminum forms the Al^{3+} cation while nitrate is NO_3^-. Thus the formula is $Al(NO_3)_3$. In Section 9.7, we will see that all nitrate compounds are soluble, so the product will be $Al(NO_3)_3(aq)$.

The equation becomes:

$Al(s) + AgNO_3(aq) \rightarrow Ag(s) + Al(NO_3)_3(aq)$

Finally, we need to balance:

$Al(s) + 3\ AgNO_3(aq) \rightarrow 3\ Ag(s) + Al(NO_3)_3(aq)$

More on the Activity Series for Metals

The Activity Series for metals is more versatile than just predicting which metal will replace which metal ion. Other information can be gained from Table 9.5. To do this, it is necessary to divide the Table into three sections as in Figure 9.10.

Sodium
Lithium
Calcium
Magnesium
Aluminum
Zinc
Iron
Tin
Lead
Copper
Silver
Gold

Figure 9.10 Other uses of the Activity Series for some metals: those coloured yellow react with water; while those in white react with acid.

First, the three metals at the bottom of the Table are particularly chemically unreactive. It is for this reason that, throughout history, these three metals have been used for coinage.

Figure 9.11 The preferred metals for coinage are those lowest in the activity series so they will last long. Here (from left to right) are a copper one-cent coin; a silver one-dollar coin; and a gold 100 dollar coin.

The six metals above them, magnesium through lead, are more chemically reactive. In particular, each one of these metals reacts with acid in a single replacement reaction. For example, zinc metal will react with hydrochloric acid, the products being a solution of zinc chloride and hydrogen gas.

$$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

For the three metals at the top of the Activity Series, they are so reactive that they will react with the hydrogen in water molecules. The product in each case will be the metal hydroxide and hydrogen gas. As an example, lithium metal will be used. To make it clearer that only one of the hydrogen atoms in a water molecule is displaced, the water molecule will be written as 'HOH.'

$$2\ Li(s) + 2\ HOH(aq) \rightarrow 2\ LiOH(aq) + H_2(g)$$

THE ACTIVITY SERIES FOR HALOGENS

Just as there is a series of reactivity for metals, so there is for the halogen elements. This series is easy to recall as it is the order of halogen elements in the Periodic Table. The list is shown in Figure 9.12.

Fluorine
Chlorine
Bromine
Iodine

Figure 9.12 The Activity Series for Halogens

To illustrate this series, chlorine is above bromine. Thus the element chlorine will replace both bromide ion and iodide ion from solution. Each of the halogens in element form exists as a diatomic molecule. Though fluorine, F_2, and chlorine, Cl_2, are gases; bromine, Br_2, a liquid; and iodine, I_2, a solid; it is more convenient in the laboratory to dissolve the element in water. For example, molecular bromine dissolved in water would be represented as $Br_2(aq)$.

In the reaction below, aqueous chlorine is reacted with a solution of potassium bromide. Chlorine, being above bromine in the Activity Series for halogens, will replace the bromide ion of potassium bromide. The products will be an aqueous solution of molecular, diatomic bromine and aqueous potassium chloride.

$$Cl_2(aq) + 2\ KBr(aq) \rightarrow Br_2(aq) + 2\ KCl(aq)$$

SUMMARY OF SINGLE REPLACEMENT REACTIONS

In this Section, the different categories of single replacement reactions have been discussed. So it will be useful to review them here. The major category is that of metal replacement. A metal, as the element, can: replace any metal ion below it in the Activity Series; replace hydrogen in an acid if the metal is between magnesium and lead (inclusive) in the Activity Series; replace one hydrogen atom in water if the metal is one of the top three metals in the Activity Series.

Then there is the separate category of halogen replacement where it is the halogen higher in the Periodic Table which will displace any halide ion below it. The categories are shown as a flow chart in Figure 9.13.

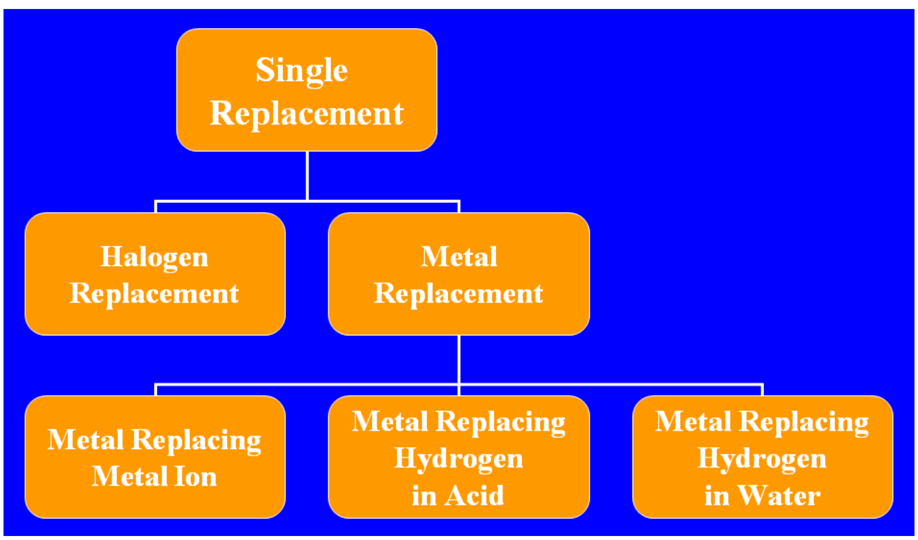

Figure 9.13 The Sub-categories of Single Replacement Reactions

9.7 Double Replacement Reactions

A ***double replacement reaction*** is a reaction in which two ionic compounds react together exchanging cations and anions. The reaction type can be written generically as:

$$AB + CD \rightarrow AD + CB$$

The reasons why a double replacement reaction occurs is quite different to that for a single replacement reaction. There are three ways in which a double replacement reaction is caused: the formation of a precipitate; the formation of a gas; or the formation of water. Each of these categories will be discussed in detail.

REACTIONS PRODUCING A PRECIPITATE

An example of this type of reaction is that between a calcium iodide solution and a silver nitrate solution. When these two solutions are mixed together, there is a yellow solid formed which settles to the bottom of the container. The solid formed as a result of a chemical reaction in solution is called a ***precipitate***. In this case, the precipitate is silver iodide and the other product consists of the other two ions, calcium nitrate. The chemical equation is:

$$CaI_2(aq) + 2\ AgNO_3(aq) \rightarrow 2\ AgI(s) + Ca(NO_3)_2(aq)$$

To be able to decide if a precipitate is formed when two solutions of ionic compounds are mixed, it is necessary to know which inorganic compounds do not dissolve in water. Such compounds are called ***insoluble*** compounds, while those which dissolve are said to be ***soluble***. Figure 9.14 shows some general rules. The use of this Table is shown in Example 9.2.

SOLUBILITY RULES

1. All alkali metal and ammonium compounds are soluble.
2. All nitrates are soluble.
3. All halides are soluble
   [exceptions: Ag^+, Hg^+, Pb^{2+}]
4. All sulfates are soluble
   [exceptions: Ba^{2+}, Hg^{2+}, Pb^{2+}]
5. All carbonates, phosphates, and sulfides are insoluble
   [exceptions: alkali metal and ammonium compounds]
6. All hydroxides are insoluble
   [exceptions: alkali metal, ammonium, and barium compounds]
7. All inorganic acids are soluble

Figure 9.14 Some general rules on the solubility of ionic compounds

Figure 9.15 Calcium phosphate being mined in central Florida. Like nearly all phosphates, calcium phosphate is insoluble, and must be changed chemically into a water-soluble compound for use as a phosphate fertilizer.

EXAMPLE 9.10

If solutions of barium nitrate and sodium sulfate are mixed, will there be a chemical reaction? If so, what will be the products?

Answer

A good place to start is the theoretical word equation. A double replacement reaction involves the switching of ions, so the equation would be:

barium chloride + sodium sulfate → sodium nitrate + barium sulfate

Thus, if a reaction occurred, the two products would be sodium nitrate and barium sulfate.

From solubility rule 1, all alkali metal compounds – which includes sodium – are soluble. So sodium nitrate would not form a precipitate.

From solubility rule 5, all sulfates are soluble, one of the exceptions being barium. So barium sulfate will form a precipitate. Reaction will occur to give a precipitate of barium sulfate.

The next step is to write the unbalanced formula equation. You can start with the reactants.

The barium cation is Ba^{2+} and the nitrate anion is NO_3^- so the compound has the formula $Pb(NO_3)_2$.

The sodium cation is Na^+ and the sulfate anion is SO_4^{2-} so the compound has the formula Na_2SO_4.

Thus the reactants can be written as:

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow$

You have already predicted the products as barium sulfate and sodium nitrate.

The barium cation is Ba^{2+} and the sulfate anion is SO_4^{2-} so the compound has the formula $BaSO_4$.

The sodium cation is Na^+ and the nitrate anion is NO_3^- so the compound has the formula $NaNO_3$.

Thus the equation can be written as:

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + NaNO_3(aq)$

Finally, the equation needs balancing with two ‘$NaNO_3$.’:

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaNO_3(aq)$

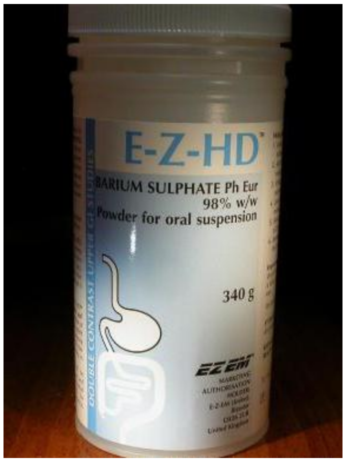

Figure 9.16 Barium ion is an excellent absorber of X-rays, so it is used for highlighting internal organs. Barium ion is very toxic, so to make it safe, the highly insoluble barium sulfate is used.

REACTIONS PRODUCING A GAS

There are only four examples of this type of double replacement reaction. The first of these is the production of hydrogen sulfide gas from reaction of a metal sulfide with an acid in a standard double replacement reaction. For example, sodium sulfide solution reacts with dilute hydrochloric acid to produce hydrogen sulfide gas and sodium chloride solution.

$Na_2S(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2S(g)$

The other three examples involve a double replacement reaction followed by a decomposition reaction. The first of these is the reaction of any metal carbonate with a dilute acid. As an example, a solution of sodium carbonate first reacts with dilute sulfuric acid to give a solution of sodium sulfate and carbonic acid:

$Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2CO_3(aq)$

However, as was described in Section 9.4, carbonic acid decomposes to give carbon dioxide gas and water. Thus a second reaction then occurs:

$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$

If required, the two equations can be combined together:

$Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

The second is the parallel reaction of any metal sulfite with a dilute acid. As an example, a solution of potassium sulfite first reacts with dilute nitric acid to give a solution of sodium nitrate and sulfurous acid:

$K_2SO_3(aq) + 2\ HNO_3(aq) \rightarrow 2\ KNO_3(aq) + H_2SO_3(aq)$

However, sulfurous acid, like carbonic acid, decomposes to give sulfur dioxide and water. Thus a second reaction then occurs:

$H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g)$

If required, the two equations can be combined together:

$K_2SO_3(aq) + 2\ HNO_3(aq) \rightarrow 2\ KNO_3(aq) + H_2O(l) + SO_2(g)$

The last of these is the reaction of any ammonium compound with any soluble hydroxide. As an example, a solution of ammonium chloride first reacts with sodium hydroxide solution to give a solution of sodium chloride and ammonium hydroxide:

$NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + NH_4OH(aq)$

Ammonium hydroxide, too, undergoes a decomposition reaction, forming water and ammonia gas, $NH_3(g)$:

$NH_4OH(aq) \rightarrow H_2O(l) + NH_3(g)$

If required, the two equations can be combined together:

$NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$

REACTIONS PRODUCING WATER (NEUTRALIZATION REACTIONS)

In this type of double replacement reaction, the cause of the reaction is the formation of water molecules. By far the most common form of this type is the reaction between an acid and a base. As an example, dilute nitric acid reacts with a solution of sodium hydroxide to give sodium nitrate solution and water:

$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

Thus it is the combining of the 'H^+' of the nitric acid with the 'OH^-' of the base to give 'HOH' which causes the reaction. Neutralization reactions can occur even when the base is in the solid phase. For example, the antacid, Milk of Magnesia™, consists of small particles of solid magnesium hydroxide, $Mg(OH)_2(s)$ in suspension. This base reacts with stomach acid – mostly hydrochloric acid – to neutralize the excess acid:

$2\ HCl(aq) + Mg(OH)_2(s) \rightarrow MgCl_2(aq) + 2\ H_2O(l)$

Neutralization reactions can also happen when a basic (metal) oxide reacts with an acid. As an example of this type, nitric acid reacts with zinc oxide to give zinc nitrate solution and water:

$2\ HNO_3(aq) + ZnO(s) \rightarrow Zn(NO_3)_2(aq) + H_2O(l)$

Likewise, an acidic (nonmetal) oxide can react with a base. An example of this sub-category is the reaction between carbon dioxide and calcium hydroxide solution. This reaction is the limewater test used in chemistry and biology: exhaled breath is bubbled through the solution of calcium hydroxide, and in addition to water, a white precipitate of calcium carbonate is formed:

$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$

SUMMARY OF DOUBLE REPLACEMENT REACTIONS

At the end of Section 9.5, the categories and sub-categories of single replacement reactions were reviewed. Here, the categories of double replacement reactions will be similarly summarised. The criteria for double replacement reactions are very different.

There are three ways in which a double replacement reaction may be caused. The most common reason is the formation of a precipitate when two solutions of ionic compounds are mixed. In order to predict whether or not one of these types of reactions will occur, it is necessary to know the solubility rules.

A more limited type of double replacement reaction is that caused by the production of a gas. There are only four gases that can be produced: hydrogen sulfide, carbon dioxide, sulfur dioxide, and ammonia. For the last three, one of the products of the double replacement reaction itself undergoes a decomposition reaction to produce the gas and water.

The third category of double replacement reactions is that producing water. More commonly known as neutralization reactions, they usually involve reaction of an acid with a base. In this category, sometimes the reaction is of an acid oxide with a base, or an acid with a basic oxide. The categories are shown as a flow chart in Figure 9.17.

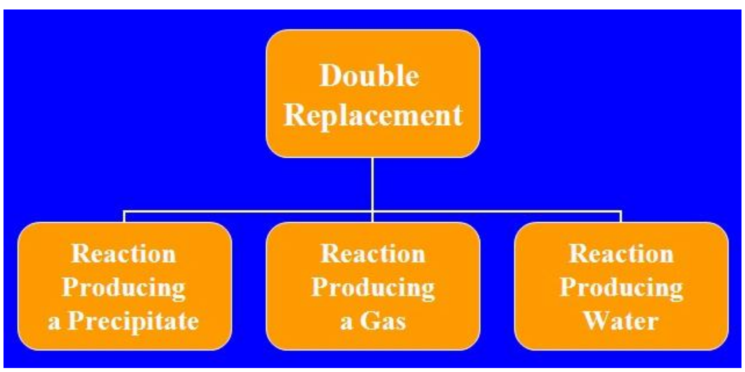

Figure 9.17 The Categories of Double Replacement Reactions

9.8 Net Ionic Equations

Particularly for single replacement and double replacement reactions, it is useful to understand exactly why the chemical reaction is occurring. To do so, the actual species in solution need to be considered.

IONS IN SOLUTION

For an ionic compound, in the solid phase, as discussed in Chapter 6, Section 6.4, the ions are arranged in a crystal lattice. When an ionic compound, such as sodium chloride, dissolves in water, water molecules surround each of the ions on the surface of the crystal and pull the ions off into the surrounding water. The process usually continues, layer by layer, until the crystal is all dissolved. Though it is not a 'true' chemical reaction, the process can be written as:

$$Na^+Cl^-(s) \xrightarrow{\text{dissolve in water}} Na^+(aq) + Cl^-(aq)$$

Thus, when a solution of sodium chloride is written as: '$NaCl(aq)$' it is actually a compact way of writing '$Na^+(aq) + Cl^-(aq)$'. The sodium cations and chloride anions, each surrounded by a layer of water molecules are floating freely and independently in the solution.

Compounds that form acids in solution are different. In the pure compound, the atoms are all connected by covalent bonds. The solution process again results in water molecules surrounding the compound, but in these cases, the covalent bond between the hydrogen and the other part of the molecule breaks, to give a hydrogen ion, $H^+(aq)$, and the respective anion. As an example, hydrogen chloride, $HCl(g)$, dissolves in water to form a solution of hydrochloric acid, $HCl(aq)$. This process can be represented as:

$$H{-}Cl(g) \xrightarrow{\text{dissolve in water}} H^+(aq) + Cl^-(aq)$$

Thus, when hydrochloric acid is written as: '$HCl(aq)$' it is actually a compact way of writing '$H^+(aq) + Cl^-(aq)$'. The hydrogen cations and chloride anions, each surrounded by a layer of water molecules are floating freely and independently in the solution.

WRITING NET IONIC EQUATIONS

Writing a net ionic equation is a two-stage process. First, the ionic compounds in solution are re-written as independent ions. The following categories are left unchanged:

Any compound in the solid phase (*s*). This rule even applies to solid ionic compounds, as in the solid phase, the ions are not free to move.

Any compound in the liquid phase (*l*). The most common example of this is the water molecule, H_2O.

Any compound in the gaseous phase (*g*). As examples would be the four gases discussed in Section 9.6: hydrogen sulfide, carbon dioxide, sulfur dioxide, and ammonia.

Any molecular element or compound in solution (*aq*). The most common examples are those of the halogens dissolved in water as discussed in Section 9.5: $F_2(aq)$, $Cl_2(aq)$, $Br_2(aq)$, and $I_2(aq)$. Also, most organic compounds (see Chapters 15 and 16), such as ethanol, $C_2H_5OH(aq)$.

This step gives the ***total ionic equation***.

Second, the species which are identical on each side of the chemical equation cannot have taken part in the reaction. For the ***net ionic equation***, these species, called ***spectator ions***, are deleted. The procedure is summarized in Figure 9.18.

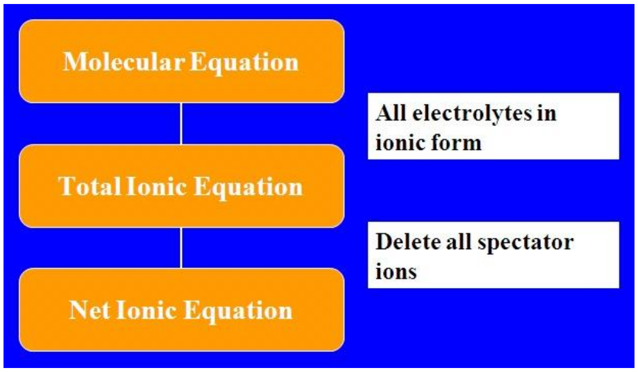

Figure 9.18 A summary of the procedure to convert a molecular equation to give a net ionic equation.

NET IONIC EQUATIONS FOR SINGLE REPLACEMENT REACTIONS

Though net ionic equations can be written for any chemical reaction that occurs in water, in practice, most of the reactions are single or double replacement reactions. Here some single replacement reactions will be converted from molecular to net ionic form.

EXAMPLE 9.11

An example of a metal replacement reaction is given below. Re-write this equation in net ionic form.

$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$

Answer

To rewrite the formula equation in ionic form, it is necessary to break the formulas apart into the constituent ions.

$Cu(s)$ is a solid and must stay as the molecular form.

$AgNO_3(aq)$ is comprised of $Ag^+(aq)$ and $NO_3^-(aq)$ and as there are two formula units in the equation, they need writing as $2\ Ag^+(aq)$ and $2\ NO_3^-(aq)$.

$Ag(s)$ is a solid and must stay as the molecular form and as there are two atoms in the equation, it needs writing as $2\ Ag(s)$.

$Cu(NO_3)_2(aq)$ is comprised of $Cu^{2+}(aq)$ and $2\ NO_3^-(aq)$.

The total ionic equation will be:

$Cu(s) + 2\ Ag^+(aq) + 2\ NO_3^-(aq) \rightarrow 2\ Ag(s) + Cu^{2+}(aq) + 2\ NO_3^-(aq)$

The nitrate ion appears on each side, so it is a spectator ion. The net ionic equation will be:

$Cu(s) + 2\ Ag^+(aq) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$

Example 9.12

An example of a single replacement reaction involving an acid is given below. Re-write this equation in net ionic form.

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Answer

$Zn(s)$ is a solid and must stay as the molecular form.

$HCl(aq)$ is comprised of $H^+(aq)$ and $Cl^-(aq)$ and as there are two formula units in the equation, they need writing as $2\ H^+(aq)$ and $2\ Cl^-(aq)$.

$ZnCl_2(aq)$ is comprised of $Zn^{2+}(aq)$ and $2\ Cl^-(aq)$.

$H_2(g)$ is a gas and must stay as the molecular form.

The total ionic equation will be:

$Zn(s) + 2\ H^+(aq) + 2\ Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2\ Cl^-(aq) + H_2(g)$

The chloride ion appears on each side, so it is a spectator ion. The net ionic equation will be:

$Zn(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Example 9.13

An example of a single replacement reaction involving halogen replacement is given below. Re-write this equation in net ionic form.

$Cl_2(aq) + 2\ KBr(aq) \rightarrow Br_2(aq) + 2\ KCl(aq)$

Answer

$Cl_2(aq)$ is a covalently-bonded diatomic molecule and must stay as the molecular form.

$KBr(aq)$ is comprised of $K^+(aq)$ and $Br^-(aq)$ and as there are two formula units in the equation, they need writing as $2\ K^+(aq)$ and $2\ Br^-(aq)$.

$Br_2(aq)$ is a covalently-bonded diatomic molecule and must stay as the molecular form.

$KCl(aq)$ is comprised of $K^+(aq)$ and $Cl^-(aq)$ and as there are two formula units in the equation, they need writing as $2\ K^+(aq)$ and $2\ Cl^-(aq)$.

The total ionic equation will be:

$Cl_2(aq) + 2\ K^+(aq) + 2\ Br^-(aq) \rightarrow Br_2(aq) + 2\ K^+(aq) + 2\ Cl^-(aq)$

The potassium ion appears on each side, so it is a spectator ion. The net ionic equation will be:

$Cl_2(aq) + 2\ Br^-(aq) \rightarrow Br_2(aq) + 2\ Cl^-(aq)$

Net Ionic Equations for Double Replacement Reactions

In the following examples, some double replacement reactions will be converted from molecular to net ionic form.

EXAMPLE 9.14

An example of a double replacement reaction involving a precipitate is given below. Re-write this equation in net ionic form.

$Ca(NO_3)_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ KNO_3(aq)$

Answer

To rewrite the formula equation in ionic form, it is necessary to break the formulas apart into the constituent ions.

$Ca(NO_3)_2(aq)$ is comprised of $Ca^{2+}(aq)$ and $2\ NO_3^-(aq)$

$K_2CO_3(aq)$ is comprised of $2\ K^+(aq)$ and $CO_3^{2-}(aq)$

$CaCO_3(s)$ is a solid and must stay as the molecular form.

$KNO_3(aq)$ is comprised of $K^+(aq)$ and $NO_3^-(aq)$ and as there are two formula units in the equation, they need writing as $2\ K^+(aq)$ and $2\ NO_3^-(aq)$.

The total ionic equation becomes:

$Ca^{2+}(aq) + 2\ NO_3^-(aq) + 2\ K^+(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s) + 2\ K^+(aq) + 2\ NO_3^-(aq)$

The potassium ions and the nitrate ions are spectator ions, so the net ionic equation becomes:

$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

Thus what happens in the chemical reaction is that calcium ions and carbonate ions have such a strong electrostatic attraction that they cluster together in a crystal lattice and form a solid precipitate.

EXAMPLE 9.15

An example of a double replacement reaction involving production of a gas is given below. Re-write this equation in net ionic form.

$Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

Answer

To rewrite the formula equation in ionic form, it is necessary to break the formulas apart into the constituent ions.

$Na_2CO_3(aq)$ is comprised of $2\ Na^+(aq)$ and $CO_3^{2-}(aq)$

$H_2SO_4(aq)$ is comprised of $2\ H^+(aq)$ and $SO_4^{2-}(aq)$

$Na_2SO_4(aq)$ is comprised of $Na^+(aq)$ and $SO_4^{2-}(aq)$

$H_2O(l)$ is a liquid and must stay as the molecular form

$CO_2(g)$ is a gas and must stay in the molecular form

The total ionic equation will be:

$2\ Na^+(aq) + CO_3^{2-}(aq) + 2\ H^+(aq) + SO_4^{2-}(aq) \rightarrow 2\ Na^+(aq) + SO_4^{2-}(aq) + H_2O(l) + CO_2(g)$

The sodium ions and sulfate ions appear on each side, so they are spectator ions. The net ionic equation will be:

$CO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

EXAMPLE 9.16

An example of a double replacement reaction involving production of water – a neutralization reaction – is given below. Re-write this equation in net ionic form.

$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

Answer

To rewrite the formula equation in ionic form, it is necessary to break the formulas apart into the constituent ions.

$HNO_3(aq)$ is comprised of $H^+(aq)$ and $NO_3^-(aq)$

$NaOH(aq)$ is comprised of $Na^+(aq)$ and $OH^-(aq)$

$NaNO_3(aq)$ is comprised of $Na^+(aq)$ and $NO_3^-(aq)$

$H_2O(l)$ is a liquid and must stay as the molecular form

The total ionic equation will be:

$H^+(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + H_2O(l)$

The sodium ions and nitrate ions appear on each side, so they are spectator ions. The net ionic equation will be:

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

9.9 Where Next?

In Chapter 8, mole calculations for compounds were discussed. Now in Chapter 9, common types of chemical reactions have been described. So in the next Chapter, the two topics will be combined to show how mole calculations can be used in chemical reactions.