Describing Chemical Reactions

CK12 Editor

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CHAPTER 1 Describing Chemical Reactions

CHAPTER OUTLINE

1.1 Physical and Chemical Changes

Lesson Objectives

- Describe the difference between physical and chemical changes
- Identify whether a given transformation is physical or chemical

Introduction

Chemists learn a lot about the nature of matter by studying the changes that it can undergo. To help classify some of these changes, a distinction is drawn between physical change and chemical change. **Physical changes** are those in which no chemical bonds are broken or formed. Even if the substance has a new form after the change, it is still considered the same chemical substance, retaining certain fundamental properties. **Chemical changes** involve breaking and forming chemical bonds. The product of the change is a completely different substance than the starting material, with its own set of physical and chemical properties.

FIGURE 1.1
Firework displays are an example of a chemical change.

Physical Changes

One type of physical change is a physical deformation. If you rip up a piece of paper, you have not made any fundamental changes to its chemical makeup. The smaller pieces of paper still have the same basic properties as the larger one.

Creating a mixture is another type of physical change. If you have a pile of pennies and a pile of nickels, they do not lose their identity by mixing them together. One special type of mixture that you learned about in the previous chapter is a solution. Dissolving a sugar packet in a glass of water does not change the chemical identity of the sugar or the water, they are simply mixed together too well for us to see that the sugar is still there. If you were to separate the mixture (e.g. by letting the water evaporate), the sugar would still be there, fully intact.

One important type of physical change that is often confusing to introductory chemistry students is the transition between different states of matter. When ice melts, there is a clear change in its appearance and some of its physical properties. However, there is no change being made to the water molecules themselves. The difference is that the water molecules have more energy, so they are moving around freely.

FIGURE 1.2
Melting snow is an example of a physical change.

Similarly, when water boils, the vapor that comes off is still composed of water molecules - one oxygen atom covalently bonded to two hydrogen atoms. The chemical identity is the same, but now, the water molecules are energetic enough to overcome the intermolecular attractions that keep them together in the liquid form.

Chemical Changes

The defining feature of a chemical change is that bonds are broken and formed to create a new substance with a different chemical identity than the starting material. One example of a chemical change is the combustion of fuels,

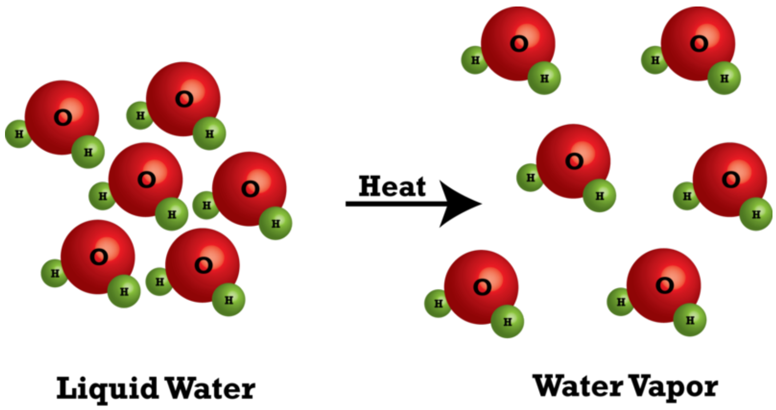


FIGURE 1.3
Water changing from a liquid to gas (boiling) is a physical change, because H_2O molecules are present at the beginning and end of the change.

such as wood, gasoline, or methane (natural gas). The combustion of methane is shown in Figure 1.4. When methane is burned in oxygen, a chemical reaction takes place that produces carbon dioxide and water, two completely different entities with fundamentally different chemical properties.

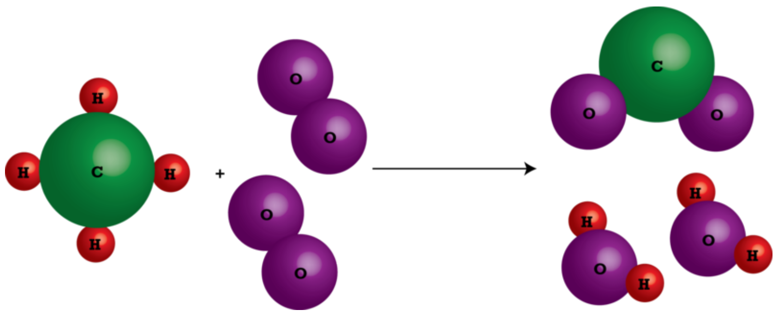


FIGURE 1.4
Natural gas burning is an example of a chemical change, because bonds are broken and formed to make different molecules.

In general, chemical changes tend to be less easy to reverse than physical ones. For example, water can easily be changed back and forth between the solid, liquid, and gas phases, but once you burn a piece of firewood, there is no going back. However, there are still many examples of reversible chemical changes and irreversible physical changes.

Because we cannot physically see whether bonds are being broken or formed, it is often difficult to tell from simple observation whether a change is physical or chemical. However, some clues that might suggest a chemical change include the following:

- Light is emitted
- Bubbles are formed, but the temperature is below the boiling point of your original substance. This may indicate that you have created a new substance that is a gas at this temperature.
- Changes in certain properties (such as color, smell, or taste) that cannot be explained by simply mixing or adding together the properties of the original substances.

Lesson Summary

- Chemists make a distinction between two different types of changes that they study – physical changes and chemical changes.

- Physical changes are changes that do not alter the identity of a substance
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes.

Vocabulary

Physical changes Changes that do not alter the identity of a substance, the same types of molecules are present at the beginning and end of the change.

Chemical changes Changes that occur when one substance is turned into another substance by making and breaking chemical bonds; different types of molecules are present at the beginning and end of the change.

Review Questions

Label each of the following as a physical or chemical change.

1. Water boils at 100°C.
2. Diamonds cut glass.
3. Water is separated by electrolysis (running electricity through it) into hydrogen gas and oxygen gas.
4. Sugar dissolves in water.
5. Vinegar and baking soda react to produce a gas.
6. Yeast acts on sugar to form carbon dioxide and ethanol.
7. Wood burns producing several new substances.
8. A cake is baked.

1.2 Collision Theory

Lesson Objectives

- Describe the conditions necessary for a reaction to occur.
- Describe reaction rate in terms of the frequency of successful collisions.

How Reactions Occur

A chemical system is made up of particles (atoms, molecules, ions, etc.) that are always moving around in random motion. Collisions between particles are constantly occurring, but most of these collisions do not lead to a chemical reaction. **Collision theory** proposes several conditions that are required for a successful reaction to take place:

a. The particles must collide with each other.
b. The particles must have the proper orientation.
c. The particles must collide with sufficient energy to break the old bonds.

A chemical reaction involves breaking bonds in the reactants, re-arranging the atoms, and forming new bonds to make the products. Therefore, not only must a collision occur between reactant particles, but the collision has to have sufficient energy to break the reactant bonds that need to be broken in order to form the products. Some reactions need less collision energy than others. The amount of energy the reactant particles must have in order for a reaction to occur is called the **activation energy**. One way to think of this is to look at an energy diagram, as shown in Figure 1.5. In order for a reaction to occur, particles must have enough energy to get over the "bump" that corresponds to the activation energy. If the particles do not have enough energy to cross this barrier, then they will bounce off of each other and no reaction will take place.

The rate at which a reaction occurs depends on the frequency of successful collisions. Remember, a successful collision occurs when two reactants collide with enough energy and with the right orientation. This suggests that there are three ways for us to increase the rate of a reaction:

a. Increase the *frequency* of collisions.
b. Increase the *energy* of the particles (so that more will have the ability to get over the activation barrier).
c. Increase the likelihood of particles colliding with the correct *orientation*.

Lesson Summary

- The collision theory explains what conditions are necessary for a successful reaction to take place.
- In order for a reaction to be effective, particles must collide in the correction orientation with enough energy to get over the activation barrier.

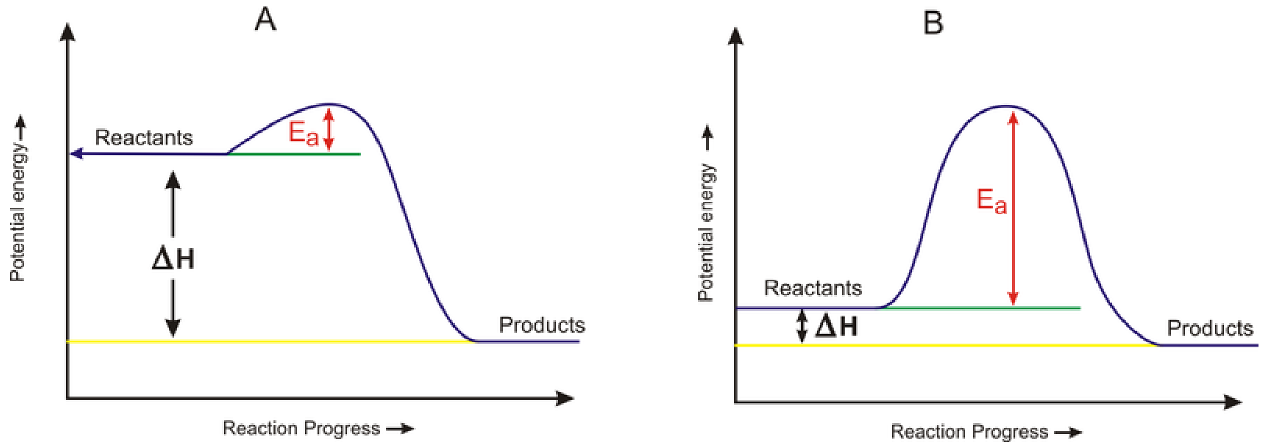


FIGURE 1.5

Each reaction has its own activation energy, E_a. The smaller the "bump", the less energy particles must have to react.

Further Readings / Supplemental Links

- Activation Energy: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/activa2.swf

Review Questions

1. According to the collision theory, it is not enough for particles to collide in order to have a successful reaction to produce products. Explain.
2. According to the collision theory, what must happen in order for a reaction to be successful? (Select all that apply.)
   a. particles must collide
   b. particles must have proper geometric orientation
   c. particles must have collisions with enough energy
3. What would happen in a collision between two particles if there was insufficient kinetic energy and improper geometric orientation?
   a. the particles would rebound and there would be no reaction
   b. the particles would keep bouncing off each other until they eventually react, therefore the rate would be slow
   c. the particles would still collide but the byproducts would form
   d. the temperature of the reaction vessel would increase

1.3 Factors That Affect Reaction Rates

Lesson Objectives

- Describe how changes to each of the following variables affects the rate of a chemical reaction:
  - Temperature
  - Concentration
  - Surface area
  - Addition of a catalyst
- Define a catalyst and describe how it affects the potential energy diagram of a reaction

Introduction

Chemists use reactions to generate a desired product. For the most part, a reaction is only useful if it occurs at a reasonable rate. For example, a reaction that took 8,000 years to complete would not be a desirable way to produce brake fluid. However, a reaction that proceeded so quickly that it caused an explosion would also not be useful (unless the explosion, and not the resulting chemical, was the desired result). For these reasons, chemists wish to be able to control reaction rates. In order to gain this control, we must first know what factors affect the rate of a reaction. We will discuss some of these factors in this section.

Temperature

When temperature is increased, the rate of a reaction generally increases as well. Raising the temperature increases the average velocity of the particles. The result is an increase in collision frequency, because faster particles will encounter more other particles each second. However, this is not the only reason for the higher reaction rate. Kinetic energy depends only on the mass of a particle and its velocity, so if the velocity is increased (and the mass is not reduced), the kinetic energy is increased as well. Having a higher average kinetic energy means that more particles will have the energy necessary for reaction when collisions do occur.

At any one moment in the atmosphere, there are many collisions occurring between molecules of hydrogen and oxygen. However, water is not formed because the activation energy is just too high, so all the collisions are resulting in a rebound.

$$O_2(g) + H_2(g) \rightarrow \text{No reaction}$$

When energy is added by increasing the temperature (or some other method), the molecules are able to react, producing water:

$$O_2(g) + H_2(g) \rightarrow H_2O(l)$$

In addition to accelerating desired reactions, temperature changes can also be used to hinder undesired reactions. As you might expect from the previous discussion, lowering the temperature will decrease the kinetic energy of the particles and the frequency of collisions, thus slowing the reaction rate. When you put something in the fridge or freezer to keep it from spoiling, you are taking advantage of this fact.

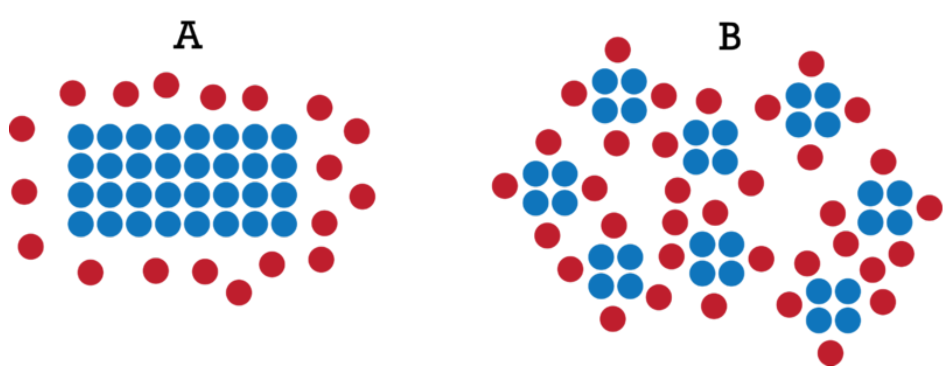
Concentration

Increasing the concentration of one or more of the reactants in a chemical reaction will raise the collision frequency, thus speeding up the reaction. In order to intuitively grasp the effect of concentration on collision frequency, it may help to consider the following thought experiment.

Imagine one red ball and one green ball bouncing randomly around a large empty room (ignore gravity for now). Every now and then, the two balls will collide, and over time, there will be some average frequency for these collisions. Now imagine adding another red ball to the room. The green ball will now collide with a red ball at twice the original frequency, because at any moment in time, there are two spots in the room that could result in such a collision rather than just one. Concentration is an amount divided by a volume. We have discussed this concept mainly in terms of molarity (moles per liter), but concentration can just as easily be discussed in terms of other units, such as balls per room. By adding another red ball, we have doubled the concentration of red balls in the room, thus doubling the rate of collision between one red and one green ball. The same is true on a molecular level; doubling the amount of reactant A in a given volume will double the frequency with which one molecule of A collides with one molecule of B. Assuming the percent of collisions that are successful is not affected by concentration (generally a good assumption), the reaction rate should double as well.

Surface Area

The previous section pointed out how increasing the concentration of the reactants increases reaction rate by increasing the frequency of collisions between reactant particles. However, collision frequency is also dependent on the surface area of reactants. Consider a reaction between reactant RED and reactant BLUE in which reactant blue is in the form of a single lump. Then compare this to the same reaction where reactant blue has been broken up into many smaller pieces.



Only the blue particles on the outside surface of the lump are available for collision with reactant red. The blue particles on the interior of the lump are protected by the blue particles on the surface. In **Figure A**, if you count the number of blue particles available for collision, you will find that only 20 blue particles could be struck by a particle of reactant red. In **Figure B**, however, the lump has been broken up into smaller pieces and all the interior blue particles are now on a surface and available for collision. In diagram B, more collisions between blue and red will occur, and therefore, the reaction in **Figure B** will occur at faster rate than the same reaction in **Figure A**. Increasing the surface area of a reactant increases the frequency of collisions and increases the reaction rate.

The more surface area that is available for particles to react, the faster the reaction will occur. You can see an example of this in everyday life if you have ever tried to start a fire in the fireplace. If you hold a match up against a large log in an attempt to start the log burning, you will find it to be an unsuccessful effort. Flammable material like wood requires a significant input of activation energy for the reaction to occur. The reaction between wood and oxygen is an exothermic reaction, and once the fire has been started, the heat released by the first reactions provides the activation energy for the subsequent reactions. In order to start a wood fire, it is common to break a log up into many small, thin sticks called kindling. These thinner sticks of wood provide many times the surface area of a single log. By accelerating the reaction in this way, we are ensuring that sufficient heat is generated to provide activation energy for further reactions.

There have been, unfortunately, cases where serious accidents were caused by the failure to understand the relationship between surface area and reaction rate. One such example occurred in flour mills. A grain of wheat is not very flammable. It takes a significant effort to get a grain of wheat to burn. If the grain of wheat, however, is pulverized and scattered through the air, only a spark is necessary to cause an explosion. When the wheat is ground to make flour, it is pulverized into a fine powder and some of the powder gets scattered around in the air. At that point, a small spark is sufficient to start a very rapid reaction capable of destroying the entire flour mill. In a 10-year period from 1988 to 1998, there were 129 grain dust explosions in mills in the United States. Efforts are now made in flour mills to have huge fans circulate the air in the mill through filters to remove the majority of the flour dust particles.

Another example is found in the operation of coal mines. Coal will burn but it takes an effort to get the coal started, and once it is burning, it burns slowly because only the surface particles are available to collide with oxygen particles. The interior particles of coal have to wait until the outer surface of the coal lump burns off before they can collide with oxygen. In coal mines, huge blocks of coal must be broken up before the coal can be brought out of the mine. To accomplish this, drills are used to drill into the walls of coal. This drilling produces fine coal dust that mixes into the air. Then, a spark from a tool can cause a massive explosion. There are explosions in coal mines for other reasons, but coal dust explosions have contributed to the death of many miners. In modern coal mines, lawn sprinklers are used to spray water through the air in the mine, reducing the coal dust in the air and reducing the risk of coal dust explosions.

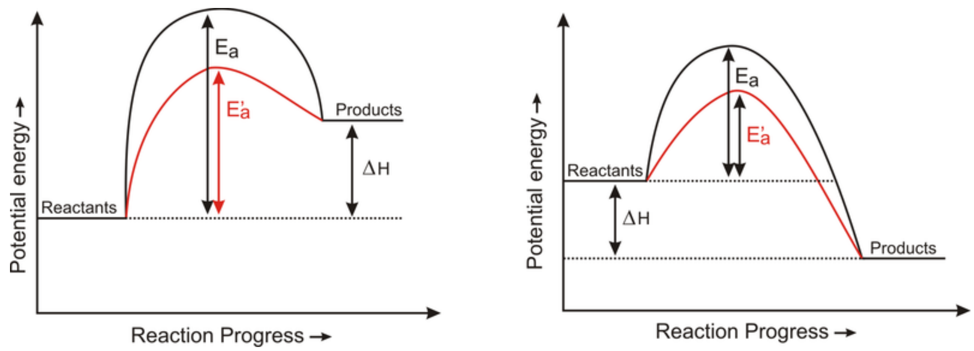
Catalyst

One final factor that affects reaction rates is the addition of a catalyst. A **catalyst** is a substance that speeds up the rate of the reaction without being consumed by the reaction.

There are a number of different types of catalysts. Some catalysts merely provide a surface for intermediate products to adhere to. Others are consumed at the beginning of a reaction but are re-formed at the end. Enzymes are proteins that catalyze numerous chemical reactions in the body. Many commercial preparations of chemicals rely on catalysts to prepare their products more efficiently and cost effectively.

The decomposition of potassium chlorate into potassium chloride and oxygen gas is greatly accelerated by the addition of a manganese dioxide catalyst. When the reaction is over, most of the MnO_2 can be recovered, because as a catalyst, it is not consumed by the reaction. To indicate the use of a catalyst in a reaction, it is generally written over the reaction arrow rather than with the reactants or products.

$$2\ KClO_3(s) \xrightarrow{MnO_2(s)} 2\ KCl(s) + 3\ O_2(g)$$



Some reactions occur very slowly without the presence of a catalyst. In other words the activation energy for these reactions is very high. When the catalyst is added, the activation energy is lowered because the catalyst provides a new reaction pathway with lower activation energy.

The left diagram in the figure above shows the energy diagram for an endothermic reaction both with (red) and without (black) the addition of a catalyst. The new reaction pathway has a lower activation energy (E_a) but has no effect on the energy of the reactants or the products.

The same is true for the exothermic reaction in the right diagram above. The activation energy of the catalyzed reaction (again designated by E'_a) is lower than that of the uncatalyzed reaction. The new reaction pathway provided by the catalyst for the exothermic reaction in this diagram affects the energy required for reactant bonds to break and product bonds to form.

While many reactions in the laboratory can be sped up by increasing the temperature, that is not possible for the reactions that occur in our bodies. To function properly, the body needs to be maintained at a very specific temperature: $98.6°F$ ($37°C$). There are times when the body temperature may be increased, such as when the body is fighting an infection, but in a healthy person, the temperature is usually quite consistent. However, many of the reactions that a healthy body depends on could never occur at body temperature. The answer to this dilemma is catalysts referred to as enzymes. These enzymes are proteins made by your cells, directed by the templates carried in your DNA.

Lesson Summary

- With an increase in temperature, there is an increase in energy that can be used as activation energy in a collision, causing an increase in the reaction rate. A decrease in temperature would have the opposite effect.
- With an increase in temperature there is an increase in the frequency of collisions, which increases the rate of reaction. However, the major reason reaction rate increases with increasing temperature is because more particles have sufficient activation energy to overcome the activation barrier.
- Increasing the concentration of a reactant increases the frequency of collisions between reactants, raising the reaction rate.
- Increasing the surface area of a reactant increases the number of particles available, thus increasing the collision frequency.
- A catalyst is a substance that speeds up the rate of a reaction without itself being consumed by the reaction. The rate is increased because the catalyst provides a new reaction pathway with a lower activation energy.
- The new reaction pathway has lower activation energy but has no effect on the energy of the reactants or the products.

Further Reading / Supplemental Links

- http://learner.org/resources/series61.html

The ***learner.org*** website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called ***Molecules in Action***.

- http://www.vitamins-guide.net
- http://en.wikipedia.org/wiki
- Observing molecules during chemical reactions helps explain the role of catalysts. Dynamic equilibrium is also demonstrated. Molecules in Action http://www.learner.org/vod/vod_window.html?pid=806
- Surface science examines how surfaces react with each other at the molecular level. On the Surface http://www.learner.org/vod/vod_window.html?pid=812)
- http://en.wikipedia.org/wiki

Vocabulary

Catalyst A substance that speeds up the rate of a reaction without itself being consumed by the reaction

Surface area to volume ratio The comparison of the volume inside a solid to the area exposed on the surface.

Review Questions

1. Why does higher temperature increase the reaction rate?
   a. more of the reacting molecules will have higher kinetic energy
   b. increasing the temperature causes the reactant molecules to heat up
   c. the activation energy will decrease
   d. increasing the temperature causes the potential energy to decrease
2. When the temperature is increased, what does not change?
   a. number of collisions
   b. activation energy requirement
   c. number of successful collisions
   d. all of the above change
3. Explain how concentration affects reaction rate using the collision theory. You may want to include a diagram to help illustrate your explanation.
4. Why is the increase in concentration directly proportional to the rate of the reaction?
   a. The kinetic energy increases.
   b. The activation energy increases.
   c. The number of successful collisions increases.
   d. All of the above.
5. Why, using the collision theory, do reactions with higher surface area have faster reaction rates?
6. When does an increase in surface area not increase the rate of reaction?
   a. The rate will not be increased if there is insufficient activation energy present.

   b. The rate will not increase if there is not an increase in collisions.
   c. The rate will not increase if the concentration doesn't change.

7. Choose the substance with the greatest surface in the following groupings:
   a. a block of ice or crushed ice
   b. sugar crystals or sugar cubes
   c. a piece of wood or wood shavings
8. The main function of a catalyst is to
   a. provide an alternate reaction pathway
   b. change the kinetic energy of the reacting particles
   c. add another reactant
9. What happens when a catalyst is added?
10. Limestone (calcium carbonate) reacts with hydrochloric acid in an irreversible reaction, to form carbon dioxide and water as described by the following **Equation**

$$CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

    What is the effect on the rate if

    a. The temperature is lowered?
    b. Limestone chips are used instead of a block of limestone?
    c. A more dilute solution of HCl is used?

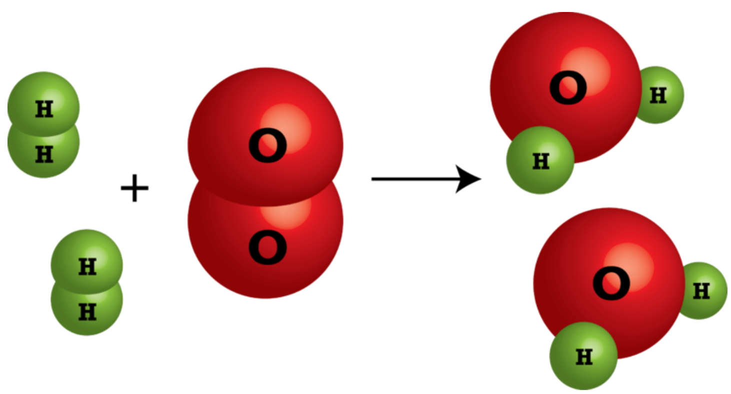
1.4 Chemical Reactions and Equations

Lesson Objectives

- Identify the reactants and products in any chemical reaction.
- Convert verbal descriptions into chemical equations and vice versa.
- Use the common symbols $(s), (l), (g), (aq)$, and $\rightarrow$ appropriately when writing a chemical reaction
- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Demonstrate the Law of Conservation of Mass for a chemical reaction.

Introduction

Chemists have multiple ways of describing a chemical reaction. They can draw a picture of the reaction:



They can describe the reaction in words: "Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

Or, they can write a chemical equation:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

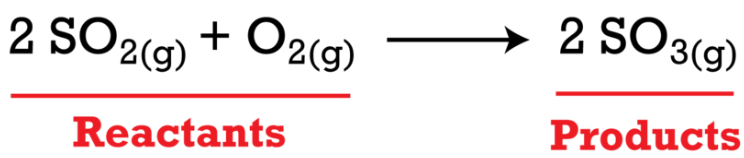
In the equation, chemical symbols are used instead of names, and other symbols are used to indicate the phase of each substance. It should be apparent that this is the quickest and clearest method for describing a chemical reaction.

Reactants and Products

In order to describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

The general equation for a reaction is "Reactants [U+F0E0] Products". Let's look at an example.

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $SO_2 + O_2$, are reactants and sulfur trioxide, SO_3, is the product.



Describing Reactions in Multiple Ways

It is important to be able to translate between written descriptions for reactions and chemical equations. For example, an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. The corresponding chemical equation would be:

$$Ca(NO_3)_2(aq) + 2\ NaOH(aq) \rightarrow Ca(OH)_2(s) + 2\ NaNO_3(aq)$$

Now let's try that in reverse. Consider the following equation:

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

A verbal description of this reaction might read, "Solid copper reacts with an aqueous solution of silver nitrate to produce solid silver and an aqueous solution of copper (II) nitrate."

In addition to the chemical formulas for reactants and products, there are a few symbols that we need to know in order to communicate using chemical equations. Table 1.1 summarizes some of the more commonly used symbols.

TABLE 1.1: Common Symbols in Chemical Reactions

Symbol	Meaning	Example
$\rightarrow$	used to separate reactants from products	$2\ H_2 + O_2 \rightarrow 2\ H_2O$
+	used to separate reactants from each other or products from each other	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
(s)	in the solid state	sodium in the solid state $= Na(s)$
(l)	in the liquid state	water in the liquid state $= H_2O(l)$
(g)	in the gaseous state	carbon dioxide in the gaseous state $= CO_2(g)$
(aq)	in the aqueous state	sodium chloride solution $= NaCl(aq)$

Sample Question

Translate the following symbolic equations into verbal descriptions or vice versa.

(a) $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

(b) Gaseous propane, C_3H_8, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.

(c) Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

(a) An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

(b) $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

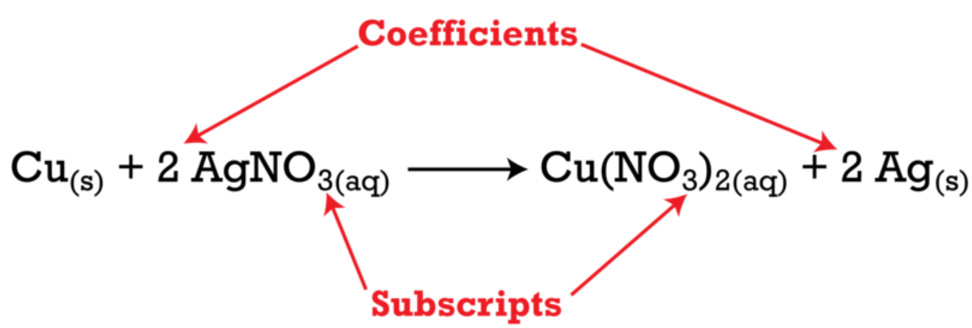
(c) $HF(g) + K_2CO_3(aq) \rightarrow KF(aq) + H_2O(l) + CO_2(g)$

Balancing Equations

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, individual atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products – they are merely re-organized into different arrangements. In a complete chemical equation, the two sides of the equation must be balanced. That is, in a complete chemical equation, the same number of each atom must be present on each side of the equation.

It is very important to note that these steps must be carried out *in the correct order*. You must have the correct formulas for your reactants and products before you can use coefficients to balance the equation.

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products, and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance are used or produced.



The subscripts are part of the formulas, and once the formulas for the reactants and products are determined, the subscripts may not be changed. The coefficients indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two moles of solid silver.

When you learned how to write chemical formulas, it was made clear that when only one atom of an element is present, the subscript "1" is not written - when no subscript appears for an atom in a formula, you read that as one atom. The same is true in writing balanced chemical equations. If only one atom or molecule is involved in the reaction, the coefficient "1" is omitted. Coefficients are inserted into the chemical equation in order to balance it, making the total number of each atom on the two sides of the equation equal. Equation balancing is accomplished by changing coefficients, never by changing subscripts.

The process of writing a balanced chemical equation involves three steps.

- **Step 1:** Identify the reactants and products.

- **Step 2:** Write the formulas for all the reactants and products. Be careful about diatomic elements and remember that the subscripts of an ionic compound are used to balance out the charge.
- **Step 3:** Adjust the coefficients to balance the equation.

Sample Problem 1: Write a balanced equation for the reaction that occurs between chlorine gas and aqueous sodium bromide to produce liquid bromine and aqueous sodium chloride.

Solution

Step 1: Write the equation out in words (keeping in mind that chlorine and bromine refer to the diatomic molecules).

$$\text{Chlorine} + \text{sodium bromide} \rightarrow \text{bromine} + \text{sodium chloride}$$

Step 2: Substitute the correct formulas into the equation.

$$Cl_2 + NaBr \rightarrow Br_2 + NaCl$$

Step 3: Insert coefficients where necessary to balance the equation.

By placing a coefficient of 2 in front of the $NaBr$, we can balance the bromine atoms and by placing a coefficient of 2 in front of the $NaCl$, we can balance the chloride atoms.

$$Cl_2 + 2\ NaBr \rightarrow Br_2 + 2\ NaCl$$

A final check (always do this) shows that we have the same number of each atom on the two sides of the equation, so this equation is properly balanced.

Sample Problem 2: Write a balanced equation for the reaction between aluminum sulfate and calcium bromide to produce aluminum bromide and calcium sulfate. (You may need to refer to a chart of polyatomic ions.)

Solution

Step 1: Write the equation in words.

$$\text{Aluminum sulfate} + \text{calcium bromide} \rightarrow \text{aluminum bromide} + \text{calcium sulfate}$$

Step 2: Replace the names of the substances with formulas.

$$Al_2(SO_4)_3 + CaBr_2 \rightarrow AlBr_3 + CaSO_4$$

Step 3: Insert coefficients to balance the equation.

In order to balance the aluminum atoms, we must insert a coefficient of 2 in front of the aluminum compound in the products.

$$Al_2(SO_4)_3 + CaBr_2 \rightarrow 2\ AlBr_3 + CaSO_4$$

In order to balance the sulfate ions, we must insert a coefficient of 3 in front of $CaSO_4$.

$$Al_2(SO_4)_3 + CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4$$

In order to balance the bromine atoms, we must insert a coefficient of 3 in front of $CaBr_2$.

$$Al_2(SO_4)_3 + 3\ CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4$$

The insertion of the 3 in front of the $CaBr_2$ in the reactants also balances the calcium atoms. A final check shows that each side contains 2 aluminum atoms, 3 sulfur atoms, 12 oxygen atoms, 3 calcium atoms, and 6 bromine atoms. This equation is balanced.

Sample Problem 3: Balance the following skeletal equation. (The term "skeletal equation" refers to an equation that has the correct formulas but has not yet had coefficients added.)

$$Fe(NO_3)_3 + NaOH \rightarrow Fe(OH)_3 + NaNO_3$$

Solution

We can balance the hydroxide ion by inserting a coefficient of 3 in front of the NaOH on the reactant side.

$$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + NaNO_3$$

Then we can balance the nitrate ions by inserting a coefficient of 3 in front of the sodium nitrate on the product side.

$$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaNO_3$$

Counting the number of each type of atom on the two sides of the equation will now show that this equation is balanced.

Chemical equations should be balanced with the smallest whole number coefficients possible. Here is the properly balanced equation from the previous section:

$$Al_2(SO_4)_3 + 3\ CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4$$

You should note that the following equation would also have the same number of atoms of each type on each side of the equation:

$$2\ Al_2(SO_4)_3 + 6\ CaBr_2 \rightarrow 4\ AlBr_3 + 6\ CaSO_4$$

While this set of coefficients does "balance" the equation, they are not the lowest set of coefficients possible. We could divide each of the coefficients in this equation by 2 and get another set of coefficients that are whole numbers and also balance the equation. Since it is required that an equation be balanced with the lowest whole number coefficients, the last equation is NOT considered to be properly balanced. When you have finished balancing an equation, you should not only check to make sure it is balanced, but also check to make sure that it is balanced with the simplest set of whole number coefficients possible.

Sample Problems

Balance the following skeletal equations.

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

(b) $H_2SO_4(aq) + Al(OH)_3(aq) \rightarrow Al_2(SO_4)_3(aq) + H_2O(l)$

(c) $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(aq) + NaNO_3(aq)$

(d) $C_2H_4(g) + O_2 \rightarrow CO_2(g) + H_2O(l)$

Solutions

(a) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ (sometimes the skeletal equation is already balanced with all coefficients equaling 1)

(b) $3H_2SO_4(aq) + 2Al(OH)_3(aq) \rightarrow Al_2(SO_4)_3(aq) + 6H_2O(l)$

(c) $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(aq) + 2NaNO_3(aq)$

(d) $C_2H_4(g) + 3O_2 \rightarrow 2CO_2(g) + 2H_2O(l)$

Conservation of Mass in Chemical Reactions

The Law of Conservation of Mass states that mass is conserved in chemical reactions. An examination of a properly balanced equation will demonstrate this property. Consider the following reaction.

$$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaNO_3$$

You should check that this equation is balanced by counting the number of each type of atom on each side of the equation.

We can demonstrate that mass is conserved in this reaction by determining the total mass on the two sides of the equation.

Reactant Side Mass

$$\begin{gathered}
1 \text{ mole of } Fe(NO_3)_3 \times \text{molar mass} = (1 \text{ mol})(241.9 \text{ g/mol}) = 241.9 \text{ g} \\
3 \text{ moles of NaOH} \times \text{molar mass} = (3 \text{ mol})(40.0 \text{ g/mol}) = 120. \text{ g} \\
\text{Total mass for reactants} = 241.9 \text{ g} + 120. \text{ g} = 361.9 \text{ g}
\end{gathered}$$

Product Side Mass

$$\begin{gathered}
1 \text{ mole of } Fe(OH)_3 \times \text{molar mass} = (1 \text{ mol})(106.9 \text{ g/mol}) = 106.9 \text{ g} \\
3 \text{ moles of } NaNO_3 \times \text{molar mass} = (3 \text{ mol})(85.0 \text{ g/mol}) = 255 \text{ g} \\
\text{Total mass for products} = 106.9 \text{ g} + 255 \text{ g} = 361.9 \text{ g}
\end{gathered}$$

Lesson Summary

- Chemical reactions are represented by chemical equations.

- Chemical equations have reactants on the left and the products on the right separated by an arrow.
- To be useful, chemical equations must always be balanced.
- Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio.
- Mass is always conserved in chemical reactions.

Vocabulary

Reactants The starting materials in a reaction.
Products Materials present at the end of a reaction.
Balanced chemical equation A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
Subscripts Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
Coefficient A whole number that appears in front of a formula in a balanced chemical equation.

Further Reading / Supplemental Links

- For a Bill Nye video on reactions, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "Bill Nye reactions". (username to pioneer: pioneer; password: time)
- For videos and clips on reactions, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "reactions" or "chemical equations". (username to pioneer: pioneer; password: time)
- Vision Learning: Chemical Equations http://visionlearning.com/library/module_viewer.php?mid=56&l=&c3=
- Balancing Equations Tutorial: http://www.mpcfaculty.net/mark_bishop/balancing_equations_tutorial.htm
- Balancing Equations Tutorial: http://www.wfu.edu/ ylwong/balanceeq/balanceq.html
- Law of Conservation of Mass (YouTube): http://www.youtube.com/watch%3Fv%3DdExpJAECSL8

Review Questions

Balance the following chemical equations and translate them into words.

1. $H_2SO_4(aq) + NaCN(aq) \rightarrow HCN(aq) + Na_2SO_4(aq)$
2. $Cu(s) + AgNO_3(aq) \rightarrow Ag(s) + Cu(NO_3)_2(aq)$
3. $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

Write balanced chemical equations for hte following reactions. Be sure to write the correct formula for ionic compounds by balancing the charges for the cations and anions.

4. Solid calcium metal is placed in liquid water to produce aqueous calcium hydroxide and hydrogen gas.
5. Aqueous sodium hydroxide is mixed with gaseous chlorine to produce aqueous solutions of sodium chloride and sodium hypochlorite plus liquid water.
6. Solid xenon hexafluoride (XeF_6) is mixed with liquid water to produce solid xenon trioxide (XeO_3) and gaseous hydrogen fluoride.
7. Iron reacts with sulfur when heated to form iron(II) sulfide.

8. When aluminum is added to sulfuric acid H_2SO_4)</math>, the solution reacts to form hydrogen gas and aluminum sulfate.
9. When aluminum is mixed with iron(III) oxide, they react to produce aluminum oxide and iron.
10. Fluorine gas (F_2) is mixed with sodium hydroxide to form sodium fluoride, oxygen gas, and water.
11. A solid chunk of iron is dropped into a solution of copper(I) nitrate, forming iron(II) nitrate and solid copper.
12. Explain in your own words why it is essential that subscripts remain constant but coefficients can change.
13. When properly balanced, what is the sum of all the coefficients in the following chemical equation? $SF_4 + H_2O \rightarrow H_2SO_3 + HF$

    a. 4
    b. 7
    c. 9
    d. None of the above

14. When the following equation is balanced, what is the coefficient found in front of the O_2? $P_4 + O_2 + H_2O \rightarrow H_3PO_4$

    a. 1
    b. 3
    c. 5
    d. 7

Balance the following equations.

15. $XeF_6 + H_2O \rightarrow XeO_3 + HF$
16. $Cu + AgNO_3 \rightarrow Ag + Cu(NO_3)_2$
17. $Fe + O_2 \rightarrow Fe_2O_3$
18. $Al(OH)_3 + Mg_3(PO_4)_2 \rightarrow AlPO_4 + Mg(OH)_2$
19. $Cu + O_2 \rightarrow CuO$
20. $H_2O \rightarrow H_2 + O_2$
21. $Fe + H_2O \rightarrow H_2 + Fe_2O_3$
22. $AsCl_3 + H_2S \rightarrow As_2S_3 + HCl$
23. $Fe_2O_3 + H_2 \rightarrow Fe + H_2O$
24. $CaCO_3 \rightarrow CaO + CO_2$
25. $H_2S + KOH \rightarrow H_2O + K_2S$
26. $NaCl \rightarrow Na + Cl_2$
27. $Al + H_2SO_4 \rightarrow H_2 + Al_2(SO_4)_3$
28. $H_3PO_4 + NH_4OH \rightarrow H_2O + (NH_4)_3PO_4$
29. $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
30. $Al + O_2 \rightarrow Al_2O_3$
31. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

1.5 Types of Reactions

Lesson Objectives

- Classify a chemical reaction as a synthesis, decomposition, single replacement, double replacement, or a combustion reaction.
- Predict the products of simple reactions.

Introduction

Chemical reactions are classified into types to help us analyze them and also to help us predict what the products of the reaction will be. The five major types of chemical reactions are synthesis, decomposition, single replacement, double replacement, and combustion.

Synthesis Reactions

A synthesis reaction is one in which two or more reactants combine to make one type of product.

$$General\ equation:\ A + B \rightarrow AB$$

$$Example:\ 2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

We can easily identify a synthesis reaction because there is only one product of the reaction.

You should be able to write the chemical equation for the synthesis of a compound from its individual elements. Also, if you are given reactants and told that the reaction is a synthesis reaction, you should be able to predict the products.

Sample Question

(a) Write the chemical equation for the synthesis of silver bromide, AgBr, from elemental silver and bromine.

(b) Predict the products for the following reaction: $CO_2(g) + H_2O(l)$

Solution

(a) $2\ Ag + Br_2 \rightarrow 2\ AgBr$

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

Decomposition Reactions

When one type of reactant breaks down to form two or more products, we have a decomposition reaction. The best way to remember a decomposition reaction is that for all reactions of this type, there is only one reactant.

$$General\ Equation:\ AB \rightarrow A + B$$

$$Example:\ NH_4NO_3 \rightarrow N_2O + 2\ H_2O$$

Notice the only reactant, NH_4NO_3, is on the left of the arrow, but there is more than one product on the right side of the arrow. This is the exact opposite of the synthesis reaction type.

We can predict the reactants or products for certain decomposition reactions in the same way that we did for synthesis reactions. For example, what would be the equation for the decomposition of magnesium nitride into its constituent elements? The answer is shown below.

$$Mg_3N_2 \rightarrow 3\ Mg + N_2$$

Sample Question Write the chemical equation for the decomposition of:

(a) Al_2O_3

(b) Ag_2S

(c) MgO

Solution

(a) $2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2$

(b) $Ag_2S \rightarrow 2\ Ag + S$

(c) $2\ MgO \rightarrow 2\ Mg + O_2$

Single Replacement Reactions

A third type of reaction is the single replacement reaction. In single replacement reactions, one reactant trades place with some portion of the other reactant. Metal elements will always replace other metals in ionic compounds or hydrogen in acid. Non-metal elements, except for hydrogen, will replace the anionic component of an ionic compound.

$$General\ equation:\ A + BC \rightarrow B + AC$$

Consider the following examples.

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

Notice that the metal element, Zn, replaced the metal in the compound $AgNO_3$. A metal element will always replace another metal in an ionic compound. Also, note that the charges of the ionic compounds must equal zero. To correctly predict the formula of the ionic product, you must know the charges of the ions you are combining.

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

When a metal element is mixed with acid, the metal will replace the hydrogen in the acid and release hydrogen gas a product. Once again, note that the charges of the ionic compounds must equal zero. To correctly predict the formula of the ionic product, you must know the charges of the ions you are combining, in this case, Zn^{2+} and Cl^-.

$$Cl_2(g) + 2\ KI(aq) \rightarrow 2\ KCl(aq) + I_2(s)$$

When a nonmetal element is added to an ionic compound, the element will replace the nonmetal in the compound. Also, to correctly write the formulas of the products, you must first identify the charges of the ions that will be in the ionic compound.

Sample Question

What would be the products of the reaction between solid aluminum and iron(III) oxide?

Solution

In order to predict the products, we need to know that aluminum will replace iron and form aluminum oxide (the metal will replace the metal ion in the compound). Aluminum has a charge of +3 and oxygen has a charge of -2. The compound formed between aluminum and oxygen, therefore, will be Al_2O_3. Since iron is replaced in the compound by aluminum, the iron will now be the single element in the products. The unbalanced equation will be:

$$Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$$

and the balanced equation will be:

$$2\ Al + Fe_2O_3 \rightarrow Al_2O_3 + 2\ Fe$$

Sample Question

(a) Write the chemical equation for the single replacement reaction between solid zinc and lead(II) nitrate solution to produce a zinc nitrate solution and solid lead. (*Note: zinc forms ions with a +2 charge)

(b) Predict the products for the following reaction: $Fe + CuSO_4$ (in this reaction, assume iron forms ions with a +2 charge)

(c) Predict the products for the following reaction: $Al + CuCl_2$

(d) Complete the following reaction, then balance the equation: $Al + HNO_3 \rightarrow$

Solution

(a) $Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2$

(b) $Fe + CuSO_4 \rightarrow Cu + FeSO_4$

(c) $2\ Al + 3\ CuCl_2 \rightarrow 3\ Cu + 2\ AlCl_3$

(d) $2\ Al + 6\ HNO_3 \rightarrow 2\ Al(NO_3)_3 + 3\ H_2$

Double Replacement

For double replacement reactions, two ionic compounds react by having their cations exchange places, forming two new ionic compounds. Precipitation and neutralization reactions are two of the most common double replacement reactions. **Precipitation reactions** are ones where both reactants are dissolved in solution, ut one or more of the products is an insoluble solid. In a **neutralization reaction**, an acid and a base react to produce salt and water.

$$General\ equation: \ AB+CD \rightarrow AD+CB$$

For example, when solutions of silver nitrate and sodium chloride are mixed, the following reaction occurs:

$$AgNO_3(aq)+NaCl(aq) \rightarrow AgCl(s)+NaNO_3(aq)$$

This is an example of a precipitation reaction. Two aqueous reactants form one solid, the precipitate, and another aqueous product.

An example of a neutralization reaction occurs when sodium hydroxide, a base, is mixed with sulfuric acid:

$$2\ NaOH(aq)+H_2SO_4(aq) \rightarrow Na_2SO_4(aq)+2\ H_2O(l)$$

In order to write the products for a double displacement reaction, you must be able to determine the correct formulas for the new compounds. Remember, the total charge of all ionic compounds is zero. To correctly write the formulas of the products, you must know the charges of the ions in the reactants. Let's practice with an example or two.

Sample Question

A common laboratory experiment involves the reaction colorless solutions of between lead(II) nitrate and sodium iodide. The reactants are given below. Predict the products.

$$Pb(NO_3)_2(aq)+NaI(aq) \rightarrow$$

Solution

We know that the cations will exchange anions. We now have to look at the charges of each of the cations and anions to see what the products will be. In $Pb(NO_3)_2$, the nitrate, NO_3^- has a charge of -1. This means the lead must be +2, Pb^{2+}. In the sodium iodide, we are combining Na^+ and I^-.

Now we switch ions and write the correct subscripts so the total charge of each compound is zero. The Pb^{2+} will combine with the I^- to form PbI_2. The Na^+ will combine with the NO_3^- to form $NaNO_3$.

Only after we have the correct formulas can we worry about balancing the two sides of the reaction. The final balanced reaction will be:

$$Pb(NO_3)_2(aq)+2\ NaI(aq) \rightarrow PbI_2(s)+2\ NaNO_3(aq)$$

Sample Question

(a) Write a chemical equation for the double replacement reaction between calcium chloride solution and potassium hydroxide solution to produce potassium chloride solution and a precipitate of calcium hydroxide.

(b) Predict the products for the following reaction: $AgNO_3(aq) + NaCl(aq) \rightarrow$

(c) Predict the products for the following reaction: $FeCl_3(aq) + KOH(aq) \rightarrow$

Solution

(a) $CaCl_2(aq) + 2\ KOH(aq) \rightarrow Ca(OH)_2(s) + 2\ KCl(aq)$

(b) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

(c) $FeCl_3(aq) + 3KOH(aq) \rightarrow Fe(OH)_3(s) + 3KCl(aq)$

Combustion

In a combustion reaction, oxygen reacts with another substance to produce carbon dioxide and water. This is what happens when fuel burns. In a particular branch of chemistry, known as organic chemistry, we study compounds known as hydrocarbons. A **hydrocarbon** is compound consisting of only hydrogen and carbon. Combustion reactions usually have the same products, CO_2 and H_2O, and one of its reactants is always oxygen. In other words, the only part that changes from one combustion reaction to the next is the actual hydrocarbon that burns. The general equation is given below. Notice the oxygen, carbon dioxide, and water parts of the reaction are listed for you to show you how these reactants and products remain constant.

$$General\ equation:\ C_xH_y\ (hydrocarbon) + O_2 \rightarrow CO_2 + H_2O$$

$$Example:\ 2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

This reaction is referred to as complete combustion. Complete combustion reactions occur when there is enough oxygen to burn the entire hydrocarbon. If combustion is not complete, there may be products other than carbon dioxide and water.

Sample Problem

Write the balanced reaction for the complete combustion of propane, C_3H_8.

Solution

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

Lesson Summary

TABLE 1.2: The Five Types of Chemical Reactions

Reaction Name	Reaction Description
Synthesis:	two or more reactants form one product.
Decomposition:	one reactant forms two or more products.

TABLE 1.2: (continued)

Reaction Name	Reaction Description
Single replacement:	one element reacts with one compound to form products.
Double replacement:	two ionic compounds exchange cations.
Combustion:	a fuel reacts with oxygen gas to produce carbon dioxide and water.

Vocabulary

Synthesis reaction A reaction in which two or more reactants combine to make one product.

Decomposition reaction A reaction in which one reactant breaks down to form two or more products.

Single replacement reaction A reaction in which an element reacts with a compound to form products.

Double replacement reaction A reaction in which two ionic reactants form products by having the cations exchange places.

Combustion reaction A reaction in which oxygen reacts with another substance to produce carbon dioxide and water.

Hydrocarbon An organic substance consisting of only hydrogen and carbon.

Review Questions

Balance each equation and classify each type of reaction as synthesis, decomposition, single replacement, double replacement, or combustion.

1. $Cu + O_2 \rightarrow CuO$
2. $H_2O \rightarrow H_2 + O_2$
3. $Fe + H_2O \rightarrow H_2 + Fe_2O_3$
4. $AsCl_3 + H_2S \rightarrow As_2S_3 + HCl$
5. $Fe_2O_3 + H_2 \rightarrow Fe + H_2O$
6. $CaCO_3 \rightarrow CaO + CO_2$
7. $H_2S + KOH \rightarrow H_2O + K_2S$
8. $NaCl \rightarrow Na + Cl_2$
9. $Al + H_2SO_4 \rightarrow H_2 + Al_2(SO_4)_3$
10. $CH_4 + O_2 \rightarrow CO_2 + H_2O$
11. Distinguish between synthesis and decomposition reactions.
12. When dodecane, $C_{10}H_{22}$, burns in excess oxygen, what will be the products?
13. In the decomposition of antimony trichloride, $SbCl_2$, what will be the products? Write and balance the reaction.
14. When iron rods are placed in liquid water, a reaction occurs. Hydrogen gas evolves from the container and iron(III) oxide forms onto the iron rod.
    a. Write a balanced chemical equation for the reaction.
    b. What type of reaction is this?

Classify each of the following reactions and predict products for each reaction.

15. $H_3PO_4 + NH_4OH \rightarrow$

16. $C_3H_8 + O_2 \rightarrow$
17. $Al + O_2 \rightarrow$
18. $BaCl_2 + Na_2SO_4 \rightarrow$
19. $Ca + HCl \rightarrow$
20. $FeS + HCl \rightarrow$
21. $NaI + Br_2 \rightarrow$

1.6 Stoichiometry

Lesson Objectives

- Explain the meaning of the term “stoichiometry.”
- Determine mole ratios in chemical equations.
- Calculate the number of moles of any reactant or product from a balanced equation, given the amount of one reactant or product.
- Calculate the mass of any reactant or product from a balanced equation, given the amount of one reactant or product.

Introduction

Chemical equations provide us with information about the types of compounds that react to form products. They also provide us with the ratios in which these compounds react to form products. In this chapter you will explore the quantitative relationships that exist between the reactants and products in a balanced equation. This is known as stoichiometry.

Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in a balanced chemical equation. The word stoichiometry comes from the Greek words stoikheion, which means element, and metron, which means measure.

Interpreting Chemical Equations

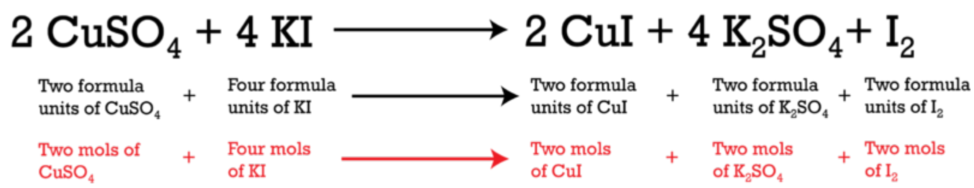
Consider the following reaction:

$$N_2O_3 + H_2O \rightarrow 2\,HNO_2$$

We have learned that the coefficients in a chemical equation tell us the relative amounts of each substance. Therefore, one way to describe the ratios involved in this reaction would be, “One molecule of dinitrogen trioxide plus one molecule of water yields two molecules of nitrous acid.” However, because these are only ratios, this statement would be equally valid using units other than molecules. Thus, we could also say, “One *mole* of dinitrogen trioxide plus one mole of water yields two moles of nitrous acid.” We can use moles instead, because a mole is simply an amount equal to Avogadro’s number, just like a dozen is an amount equal to 12.

It is important that you do not confuse this with units that describe properties other than amount. For example, it would NOT be correct to say that one gram of dinitrogen trioxide plus one gram of water yields two grams of nitrous acid, because each of these molecules has a different mass.

Now consider this reaction:



Here, we can say, "2 moles of copper (II) sulfate react with 4 moles of potassium iodide, yielding two moles of copper (I) iodide, 4 moles of potassium sulfate, and 1 mole of molecular iodine." However, because ionic substances exist as crystal lattices instead of discrete molecules, it is generally not correct to refer to "molecules" of something like KI. Instead, the term "formula unit" is used to describe one potassium ion plus one iodide ion.

Sample Question: Indicate the ratios of compounds involved in each of these balanced chemical equations in both moles and molecules/formula units.

(a) $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

(b) $KBrO_3 + 6\ KI + 5\ HBr \rightarrow 7\ KBr + 3\ I_2 + 3\ H_2O$

Solution:

(a) Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.

Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.

(b) Two formula units of $KBrO_3$ plus six formula units of KI plus six formula units of HBr yields seven formula units of KBr plus three molecules of I_2 and three molecules of H_2O.

Two moles of $KBrO_3$ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I_2 and three moles of H_2O.

Mole Ratios

A mole ratio is the relationship between the number of moles of the substances in a reaction.

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

For example, the following mole ratios can be obtained from the reaction above:

$$\frac{2\ mol\ H_2}{1\ mol\ O_2}\ or\ \frac{1\ mol\ O_2}{2\ mol\ H_2}\ or\ \frac{2\ mol\ H_2O}{2\ mol\ H_2}\ or\ \frac{2\ mol\ H_2O}{1\ mol\ O_2}\ or\ \frac{2\ mol\ H_2}{2\ mol\ H_2O}\ or\ \frac{1\ mol\ O_2}{2\ mol\ H_2O}$$

Using the coefficients of a balanced reaction, you can compare the molar ratios of any two substances in the reaction you are interested in, whether they are reactants or products. The chemical equation MUST always be balanced before the mole ratios are used for calculations.

Mole-Mole Calculations

We have already learned the process through which chemists solve many math problems, dimensional analysis. The mole-mole ratio we obtain from a balanced reaction can be used as a ratio in part of that process.

Example: If only 0.050 mol of magnesium hydroxide, $Mg(OH)_2$, is present, how many moles of phosphoric acid, H_3PO_4, would be required for the reaction?

$$2\ H_3PO_4 + 3\ Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6\ H_2O$$

Solution: We need to set up this problem using dimensional analysis.

Given: 0.050 mol $Mg(OH)_2$

Find: mol H_3PO_4

The ratio we need is one that compares mol $Mg(OH)_2$ to mol H_3PO_4. This is the ratio obtained in the balanced reaction. Note that there are other reactants and products in this reaction, but we don't need to use them to solve this problem.

$$0.050\ \cancel{mol\ Mg(OH)_2} \cdot \frac{2\ mol\ H_3PO_4}{3\ \cancel{mol\ Mg(OH)_2}} = 0.033\ mol\ H_3PO_4$$

Notice that if the equation was not balanced, the amount of H_3PO_4 would have been different. The reaction MUST be balanced to use the reaction in any calculations.

Mass-Mass Calculations

A mass-mass calculation would allow you to solve one of the following types of problems:

- Determine the mass (in grams, kg, etc.) of reactant necessary to product a given amount of product.
- Determine the mass of product that would be produced from a given amount of reactant.
- Determine the mass of reactant necessary to react completely with a second reactant.

These types of problems are generally solved by using dimensional analysis. Let's look at an example.

Example: 15.0 g of chlorine gas is bubbled over liquid sulfur to produce disulfur dichloride. How much product is produced, in grams, according to the balanced equation:

$$Cl_2(g) + 2\ S(l) \rightarrow S_2Cl_2(l)$$

Solution:

a. Identify the **given:** 15.0 $g\ Cl_2$
b. Identify the **find:** $g\ S_2Cl_2$
c. Use the **ratios** that allow you to cancel the units you don't want and get to the unit you are calculating for.

This is the molar ratio from the balanced equation, which converts moles of known to moles of unknown.

This is the known quantity.

This is the final answer.

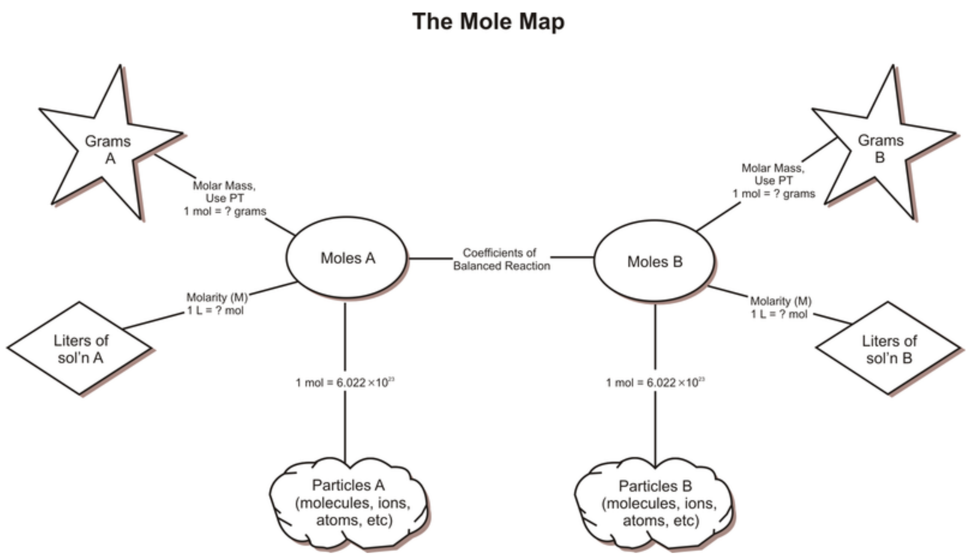
$$15.0\ \mathrm{g\ Cl_2} \times \frac{1\ \mathrm{mol\ Cl_2}}{71.0\ \mathrm{g\ Cl_2}} \times \frac{1\ \mathrm{mol\ S_2Cl_2}}{1\ \mathrm{mol\ Cl_2}} \times \frac{135\ \mathrm{g\ S_2Cl_2}}{1\ \mathrm{mol\ S_2Cl_2}} = 28.5\ \mathrm{g\ S_2Cl_2}$$

This is the molar mass of the known, which converts the given grams to moles.

This is the molar mass of the unknown, which converts moles to grams.

Calculations Using a Mole Map

If we combine the mole-mole ratio with other ratios we have already learned, we have several tools we can use to solve a wide variety of problems. The mole map is a diagram we can use to figure out which ratios to use when solving problems.



You use this map much like you would use a road map. You must first find out where you are on the map (your given units) and where you would like to go (your "find" units). The map will then let you know which roads (ratios) to take to get there. Let's see how this works with a couple of example problems.

Example

The thermite reaction is a very exothermic reaction which produces liquid iron, given by the following balanced equation:

$$Fe_2O_3(s)+2\,Al(s) \rightarrow 2\,Fe(l)+Al_2O_3(s)$$

If 5.00 g of iron is produced, how much iron(III) oxide was placed in the original container?

Solution

a. Identify the "given": 5.00 g iron. (Even though this is a product, it is still the measurement given to us in the problem. The given may be a reactant or a product.)
b. Identify the units of the "find": $g\ Fe_2O_3$ (remember, mass is measured in grams)
c. Ratios: This is where the map comes in handy. To start with, we are at 5.00 $g\ Fe$. For this problem, then, "A" on the map stands for Fe. We start at grams A.

We want to know $g\ Fe_2O_3$. For this problem, "B" stands for Fe_2O_3. We are heading to grams B.

Our map tells us this problem will take 3 ratios (3 roads from $g\ A$ to $g\ B$): molar mass of A, mol:mol ratio from a balanced reaction, and molar mass of B. To solve our problem, the work will look like:

$$5.00\ \cancel{g\ Fe} \cdot \frac{1\ \cancel{mol\ Fe}}{55.85\ \cancel{g\ Fe}} \cdot \frac{1\ \cancel{mol\ Fe_2O_3}}{2\ \cancel{mol\ Fe}} \cdot \frac{159.7\ g\ Fe_2O_3}{1\ \cancel{mol\ Fe_2O_3}} = 14.3\ g\ Fe_2O_3$$

Sample Question

Ibuprofen is a common painkiller used by many people around the globe. It has the formula $C_{13}H_{18}O_2$. If 200.g of Ibuprofen is combusted how much carbon dioxide is produced? The balanced reaction is:

$$2\ C_{13}H_{18}O_2 + 33\ O_2 \rightarrow 26\ CO_2 + 18\ H_2O$$

Solution

Given: 200.$g\ C_{13}H_{18}O_2$ (g A on the map)

Find: $g\ CO_2$ (g B on the map)

Ratios: The map says we need to use the molar mass of $C_{13}H_{18}O_2$, then the coefficients of the balanced reaction, then the molar mass of CO_2.

$$200.\cancel{g\ C_{13}H_{18}O_2} \cdot \frac{1\ \cancel{mol\ C_{13}H_{18}O_2}}{206.3\ \cancel{g\ C_{13}H_{18}O_2}} \cdot \frac{26\ \cancel{mol\ CO_2}}{2\ \cancel{mol\ C_{13}H_{18}O_2}} \cdot \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} = 555\ g\ CO_2$$

Sample Question

If sulfuric acid is mixed with sodium cyanide, the deadly gas hydrogen cyanide is produced. How many moles of sulfuric acid are required to produce 12.5 g of hydrogen cyanide? The balanced reaction is:

$$2\ NaCN + H_2SO_4 \rightarrow Na_2SO_4 + 2\ HCN$$

Solution

Given: 12.5 $g\ HCN$ (g A on map)

Find: mol H_2SO_4 (mol A on map)

Ratios: The mole map says we need the molar mass of HCN and the coefficients of the balanced reaction.

$$12.5\ g\ HCN \cdot \frac{1\ mol\ HCN}{27.0\ g\ HCN} \cdot \frac{1\ mol\ H_2SO_4}{2\ mol\ HCN} = 0.231\ mol\ H_2SO_4$$

Lesson Summary

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- When the moles of one substance in a reaction is known, the coefficients of the balanced equation can be used to determine the moles of all other substances in the reaction.
- Mass-mass calculations can be done using dimensional analysis, with help from a mole map if needed.

Vocabulary

Stoichiometry The calculation of quantitative relationships of the reactants and products in a balanced chemical equation.

Mole ratio The ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation.

Further Reading / Supplemental Links

- Stoichiometry: http://www.lsua.us/chem1001/stoichiometry/stoichiometry.html

Review Questions

1. Distinguish between formula unit, molecule, and mole. Use an example in your answer.
2. Given the reaction between ammonia and oxygen to produce nitrogen monoxide, how many moles of water vapor can be produced from 2 mol of ammonia? The balanced reaction is: $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$
3. When properly balanced, how many moles of bismuth(III) oxide can be produced from 0.625 mol of bismuth? The unbalanced reaction is: $Bi(s) + O_2(g) \rightarrow Bi_2O_3(s)$
4. For the following reaction, balance the equation and then determine the ratio of moles of $B(OH)_3$ to moles of water? $B_2O_3(s) + H_2O(l) \rightarrow B(OH)_3(s)$

Write the balanced chemical equation for the reactions below. When written, find the mol ratios indicated.

5. Gaseous propane, C_3H_8, combusts to form gaseous carbon dioxide and water; find ratio of mol O_2 to mol CO_2.
6. Solid lithium reacts with an aqueous solution of aluminum chloride to produce aqueous lithium chloride and solid aluminum; find ratio of mol $AlCl_3(aq)$ to mol $LiCl(aq)$.
7. An aqueous solution of ammonium hydroxide reacts with an aqueous solution of phosphoric acid to produce aqueous ammonium phosphate and water; find ratio of mol $H_3PO_4(aq)$ to mol $H_2O(l)$.
8. Solid rubidium reacts with solid phosphorous to produce solid rubidium phosphide; find ratio of mol $Rb(s)$ to mol $P(s)$.
9. For the given reaction (balanced): $Ca_3(PO_4)_2 + 3\ SiO_2 + 5\ C \rightarrow 3\ CaSiO_3 + 5\ CO + 2\ P$
   a. how many moles of silicon dioxide are required to react with 0.35 mol of carbon?
   b. how many moles of calcium phosphate are required to produce 0.45 mol of calcium silicate?
10. For the given reaction (balanced), $4\ FeS + 7\ O_2 \rightarrow 2\ Fe_2O_3 + 4\ SO_2$
    a. how many moles of iron(III) oxide are produced from 1.27 moles of oxygen?
    b. how many moles of iron(II) sulfide are required to produce 3.28 moles of sulfur dioxide?
11. Given the reaction between copper (II) sulfide and nitric acid, how many grams of nitric acid will react with 2.00 g of copper(II) sulfide?

$$3\ CuS(s) + 8\ HNO_3(aq) \rightarrow 3\ Cu(NO_3)_2(aq) + 2\ NO(g) + 4\ H_2O(l) + 3\ S(s)$$

12. When properly balanced, what mass of iodine is needed to produce 2.5 g of sodium iodide in the equation below? $I_2(aq) + Na_2S_2O_3(aq) \rightarrow Na_2S_4O_6(aq) + NaI(aq)$
13. Determine the mass of lithium hydroxide produced when 0.38 grams of lithium nitride reacts with water according to the following equation: $Li_3N + 3\ H_2O \rightarrow NH_3 + 3\ LiOH$
14. If 3.01×10^{23} formula units of cesium hydroxide are produced according to this reaction: $2\ Cs + 2\ H_2O \rightarrow 2\ CsOH + H_2$, how many grams of cesium reacted?
15. How many liters of oxygen are necessary for the combustion of 425 g of sulfur, assuming that the reaction occurs at STP? The balanced reaction is: $S + O_2 \rightarrow SO_2$ (hint: one mole of oxygen is 22.4 Liters at STP)

16. If I have 2.0 grams of carbon monoxide, how many molecules of carbon monoxide are there?
17. What mass of oxygen is needed to burn 3.5 g of propane (C_3H_8) is burned according to the following equation: $C_3H_8 + 5\ O_2 \rightarrow 4\ H_2O + 3\ CO_2$
18. How many grams of water are produced if 5 moles of oxygen react according to the following reaction? $2\ H_2 + O_2 \rightarrow 2\ H_2O$

1.7 Introduction to Equilibrium

Lesson Objectives

- Describe what is happening in a system at equilibrium.

Reverse Reactions and Equilibrium

Consider this hypothetical reaction: $A+B \rightarrow C+D$. Based on what we have learned so far, you might assume that the reaction will keep going forward, forming C and D until A and B run out. When this is the case, we would say that the reaction “goes to completion.” Reactions that go to completion are referred to as irreversible reactions.

However, some reactions are reversible, meaning that products can also react to re-form the reactants. In our example, this would correspond to the reaction $A+B \leftarrow C+D$. During a **reversible reaction**, both the forward and backward reactions are happening at the same time.

Equilibrium

As we learned earlier, the rate of a reaction depends on the concentration of the reactants. At the very beginning of the reaction shown above, we would not expect the reverse reaction to proceed very quickly. If only a few particles of C and D have been created in a large flask of A and B, then it is very unlikely that they will find each other, because the concentration of C and D (amount of C and D divided by the volume of the flask) is just too low. If C and D cannot find each other and collide with the correct energy and orientation, no reaction will occur.

However, as more and more C and D are created, it becomes more and more likely that they will find each other and react to re-form A and B. Conversely, as the A and B are being used up, the forward reaction slows down for the same exact reason. The concentration of A and B decreases over the course of a reaction because there are less A and B particles in the same size flask.

At some point, the rates for the forward and reverse reactions will be equal, at which point the concentrations will no longer change. If A and B are being destroyed at the same rate that they are being created, the overall amount should not change over time. The term **dynamic equilibrium** refers to a state where no net change is taking place, despite the fact that both reactions are still occurring. Individual molecules are still being formed and broken down, but the system as a whole is not changing over time.

Chemists use a double-headed arrow, $\rightleftharpoons$, to show that a reaction is at equilibrium. For our generic reaction, the chemical equation would be: $A+B \rightleftharpoons C+D$. This indicates that both directions of the reaction are occurring.

Lesson Summary

- Irreversible reactions will continue to form products until the reactants are fully consumed.

- Reversible reactions will react until a state of equilibrium is reached.
- A reaction is at equilibrium when there is no net change to the system over time.
- Dynamic equilibrium refers to an equilibrium where forward and reverse reactions are still occurring, but they are proceeding at the same rate, so there is no net change.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki/Chemical_equilibrium

Vocabulary

Irreversible reaction A reaction that continues to form products until reactants are fully consumed.
Reversible reaction A reaction that can also proceed in the reverse direction.
Dynamic equilibrium A state that occurs when the rate of the forward reaction is equal to the rate of the reverse reaction.

Review Questions

1. For the reaction $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$, describe what is happening to make this an equilibrium reaction.
2. What does the term dynamic equilibrium mean?
3. List all of the conditions of a dynamic equilibrium?
4. If the following table of concentration vs. time was provided to you for the ionization of acetic acid. How would you know when equilibrium was reached?
5. Chemical equilibrium is defined as a state where the reversible process shows that the rate of the forward reaction is equal to the rate of the reverse reaction. What does the term equal mean in this definition?

TABLE 1.3:

Time (min)	$[HC_2H_3O_2]$ mol/L
0	0.100
0.5	0.099
1.0	0.098
1.5	0.097
2.0	0.096
2.5	0.095
3.0	0.095
3.5	0.095
4.0	0.095

6. If you were to take snapshots of a system at equilibrium, how would one picture compare to the next? What would be different? What would be the same?
7. The word “equilibrium” makes people think of the word “equal”. How can a system be both dynamic and at equilibrium? What is “dynamic” in dynamic equilibrium? What is “equal”?

Indicate whether each of the following statements is true or false for a system in equilibrium.

8. The amount of products is equal to the amount of reactants.
9. The amount of product is not changing.
10. The amount of reactant is not changing.
11. Particles (atoms/molecules) are not reacting.

1.8 Equilibrium Constant

Lesson Objectives

- Write equilibrium constant expressions.
- Use equilibrium constant expressions to solve for unknown concentrations.
- Use known concentrations to solve for the equilibrium constants.
- Explain what the value of K means in terms of relative concentrations of reactants and products.

Introduction

In the previous section, you learned about reactions that can reach a state of equilibrium, in which the concentration of reactants and products aren't changing. If these amounts are changing, we should be able to make a relationship between the amount of product and reactant when a reaction reaches equilibrium.

The Equilibrium Constant

Equilibrium reactions are those that do not go to completion but are in a state where the reactants are reacting to yield products and the products are reacting to produce reactants. In a reaction at equilibrium, the equilibrium concentrations of all reactants and products can be measured. The **equilibrium constant** (K) is a mathematical relationship that shows how the concentrations of the products and reactants are related when the system is in equilibrium. Sometimes, subscripts are added to the equilibrium constant symbol K, such as K_{eq}, K_c, K_p, K_a, K_b, and K_{sp}. These are all equilibrium constants and are subscripted to indicate special types of equilibrium reactions.

There are some rules about writing equilibrium constant expressions that you must learn:

a. Construct a fraction where the concentration of each product is in the numerator and the concentration of each reactant is in the denominator.
b. If there are coefficients on any of the products or reactants in the balanced chemical equation, place them as exponents on the appropriate concentration.
c. Do not include solids or liquids in the equilibrium expression, because their concentrations do not change noticeably over the course of a reaction.

Example

Write the equilibrium expression for:

$$CO(g) + 3\,H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$$

Solution

$$K = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

Note that the coefficients become exponents. Also, note that the concentrations of products in the numerator are *multiplied*. The same is true of the reactants in the denominator.

Example

Write the equilibrium expression for:

$$2\ TiCl_3(s) + 2\ HCl(g) \rightleftharpoons 2\ TiCl_4(s) + H_2(g)$$

Solution

$$K = \frac{[H_2]}{[HCl]^2}$$

Note that the solids are left out of the expression completely.

Example

Write the equilibrium expression for:

$$P_4(s) + 6\ Cl_2(g) \rightleftharpoons 4\ PCl_3(s)$$

Solution

$$K = \frac{1}{[Cl_2]^6}$$

Note that the only product is a solid, which is left out. That leaves just 1 for the numerator.

Example

Write the equilibrium expression for:

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

Solution

$$K = [H^+][OH^-]$$

Mathematics with Equilibrium Expressions

The equilibrium expression is an equation that we can use to solve for K or for the concentration of a reactant or product.

Sample Question

$$SO_2(g)+NO_2(g) \rightleftharpoons SO_3(g)+NO(g)$$

Determine the value of K when the equilibrium concentrations are: $[SO_2] = 1.20\,M,\ [NO_2] = 0.60\,M,\ [NO] = 1.6\,M$, and $[SO_3] = 2.2\,M$.

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{[SO_3][NO]}{[SO_2][NO_2]}$$

Step 2: Substitute in given values and solve:

$$K = \frac{(2.2)(1.6)}{(1.20)(0.60)} = 4.9$$

Sample Question

Consider the following reaction: $CO(g)+H_2O(g) \rightleftharpoons H_2(g)+CO_2(g);\ K = 1.34$

If $[H_2O] = 0.100\,M,\ [H_2] = 0.100\,M$, and $[CO_2] = 0.100\,M$ at equilibrium, what is the equilibrium concentration of CO?

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{[H_2][CO_2]}{[CO][H_2O]}$$

Step 2: Substitute in given values and solve:

$$1.34 = \frac{(0.100)(0.100)}{[CO](0.100)}$$

Solving for $[CO]$, we get: $[CO] = 0.0746\,M$

The equilibrium constant value is the ratio of the concentrations of the products over the reactants. Therefore, a value of 4.9 for the equilibrium constant indicates that there are more products (numerator) than there are reactants (denominator) at equilibrium.

If the equilibrium constant is "1" or nearly "1," it indicates that the molarities of the reactants and products are about the same at equilbirum. If the equilibrium constant value was a large number, like 1×10^5, it indicates that at equilibrium, the great majority of the material is in the form of products, and we say the "products are strongly favored." If the equilibrium constant is very small, like 1×10^{-12}, it indicates that the reactants are strongly favored. For large K values, most of the material at equilibrium is in the form of products, and for small K values, most of the material at equilibrium is in the form of reactants.

Lesson Summary

- The equilibrium expression is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants.
- If the value of K is greater than 1, the products in the reaction are favored; if the value of K is less than 1, the reactants in the reaction are favored; if K is equal to 1, neither reactants nor products are favored.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki/Chemical_equilibrium
- Crabapples & Equlibrium: http://www.chem.ox.ac.uk/vrchemistry/ChemicalEquilibrium/HTML/page05.htm
- Equilibrium Animation / Applet: Dots: http://chemconnections.org/Java/equilibrium/

Vocabulary

Equilibrium constant (K) A mathematical ratio that shows the concentrations of the products divided by concentration of the reactants.

Review Questions

1. Which phases of substances do NOT affect the equilibrium expression?
2. What does a large value for K imply?
3. What does a small value of K imply?

Write an equilibrium expression for each reaction:

4. $2\ H_2(g) + O_2(g) \rightleftharpoons 2\ H_2O(g)$
5. $2\ NO(g) + Br_2(g) \rightleftharpoons 2\ NOBr(g)$
6. $NO(g) + O_3(g) \rightleftharpoons O_2(g) + NO_2(g)$
7. $CH_4(g) + Cl_2(g) \rightleftharpoons CH_3Cl(g) + HCl(g)$
8. $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3\ H_2(g)$
9. $CO(g) + 2\ H_2(g) \rightleftharpoons CH_3OH(g)$
10. $2\ C_2H_6(g) + 7\ O_2(g) \rightleftharpoons 4\ CO_2(g) + 6\ H_2O(g)$
11. $C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$
12. $Hg(g) + I_2(g) \rightleftharpoons HgI_2(g)$
13. $SnO_2(s) + 2\ CO(g) \rightleftharpoons Sn(s) + 2\ CO_2(g)$
14. $C(s) + CO_2(g) \rightleftharpoons 2\ CO(g)$
15. $FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$
16. $P_4(s) + 6\ NO(g) \rightleftharpoons P_4O_6(s) + 3\ N_2(g)$
17. $Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2\ OH^-(aq)$
18. Consider the following equilibrium system: $2\ NO(g) + Cl_2(g) \rightleftharpoons 2\ NOCl(g)$. At a certain temperature, the equilibrium concentrations are as follows: $[NO] = 0.184\ M$, $[Cl_2] = 0.165\ M$, $[NOCl] = 0.060\ M$. What is the equilibrium constant for this reaction?

19. For the reaction: $MgCl_2(s) + \frac{1}{2}O_2(g) \rightleftharpoons MgO(s) + Cl_2(g)$. The equilibrium constant was found to be 3.86 at a certain temperature. If $[O_2] = 0.560\ M$ at equilibrium, what is the concentration of $Cl_2(g)$?
20. Consider the equilibrium: $CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$.
    a. Write an equilibrium expression for this reaction.
    b. If $[CO] = 0.200M,\ [H_2O] = 0.500M,\ [H_2] = 0.32M$ and $[CO_2] = 0.42M$, find K.
21. Hydrogen sulfide decomposes according to the equation: $2\ H_2S(g) \rightleftharpoons 2\ H_2(g) + S_2(g)$.
    a. Write an equilibrium expression for this reaction.
    b. At equilibrium, the concentrations of each gas are as follows: $[H_2S] = 7.06 \times 10^{-3}M,\ [H_2] = 2.22 \times 10^{-3}M$ and $[S_2] = 1.11 \times 10^{-3}M$. What is K_{eq}?
22. Given the following system in equilibrium: $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$
    a. Write an equilibrium expression for the reaction.
    b. If $K = 85.0$, would you expect to find more reactants or products at equilibrium? Why?
    c. If $[SO_2] = 0.0500M$ and $[O_2] = 0.0500M$, what is the concentration of SO_3 at equilibrium?

1.9 The Effects of Applying Stress to Reactions at Equilibrium

Lesson Objectives

- State Le Châtelier's Principle.
- Describe the effect of concentration changes on an equilibrium system.
- Describe the effect of temperature as a stress on an equilibrium system.

Introduction

When a reaction has reached equilibrium with a given set of conditions, if the conditions are not changed, the reaction will remain at equilibrium forever. The forward and reverse reactions continue at the same equal and opposite rates, and the macroscopic properties remain constant.

It is possible, however, to perturb that equilibrium by changing conditions. For example, you could increase the concentration of one of the products, or decrease the concentration of one of the reactants, or change the temperature. When you alter something in a reaction at equilibrium, chemists say that you put **stress** on the equilibrium. When this occurs, the reaction will no longer be in equilibrium, and the reaction will modify the concentrations of reactants and products until the reaction comes to a new position of equilibrium. Predicting how the equilibrium will be affected by various changes is the topic of this section.

Le Châtelier's Principle

In the late 1800's, a chemist by the name of Henry-Louis Le Châtelier was studying stresses that were applied to chemical equilibria. He formulated a principle from this research and, of course, the principle is called Le Chatelier's Principle. **Le Châtelier's Principle** states that when a stress is applied to a system at equilibrium, the equilibrium will shift in a direction to partially counteract the stress and once again reach equilibrium.

Le Chatelier's principle is not an explanation of what happens on the molecular level to cause the equilibrium shift, it is simply a quick way to determine which way the reaction will run in response to a stress applied to the system at equilibrium.

Effect of Concentration Changes on a System at Equilibrium

For instance, if a stress is applied by increasing the concentration of a reactant, the equilibrium position will shift toward the right, producing more products, to remove that stress. The reverse is also true. If a stress is applied by lowering a reactant concentration, the equilibrium position will shift toward the left, this time producing more reactants, and once again partially counteracting that stress. If a stress is applied by increasing the concentration of a product, the equilibrium position will shift toward the left, reducing the concentration of the product. Again the reverse is true in that if a stress is applied by reducing the concentration of a product, the equilibrium position will shift toward the right and producing more reactant to counteract that stress.

The effect of concentration on the equilibrium system according to Le Châtelier is as follows: Increasing the concentration of a reactant causes the equilibrium to shift to the right using up reactants and producing more products. Increasing the concentration of a product causes the equilibrium to shift to the left using up products and producing more reactants. The exact opposite is true when either a reactant or product is removed.

Question

For the reaction: $SiCl_4(g) + O_2(g) \rightleftharpoons SiO_2(s) + 2\ Cl_2(g)$, what would be the effect on the equilibrium system if:

(a) $[SiCl_4]$ increases

(b) $[O_2]$ increases

(c) $[Cl_2]$ increases

Solution

(a) $[SiCl_4]$ increases: The equilibrium would shift to the right

(b) $[O_2]$ increases: The equilibrium would shift to the right

(c) $[Cl_2]$ increases: The equilibrium would shift left

Question Consider this reaction at equilibrium. $A(aq) + B(aq) \rightleftharpoons C(aq) + D(aq)$

(a) Which way will the equilibrium shift if you add some A to the system without changing anything else?

(b) Which way will the equilibrium shift if you add some C to the system without changing anything else?

Solution

(a) The equilibrium will shift toward the products (forward).

(b) The equilibrium will shift toward the reactants (backward).

The Effect of Changing Temperature on a System at Equilibrium

Le Chatelier's principle also correctly predicts the equilibrium shift when systems at equilibrium are heated and cooled. An increase in temperature is the same as adding heat to the system. Consider the following equilibrium:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \quad \Delta H = -191\ kJ$$

We will learn more about this later, but ΔH has to do with the change in energy, usually heat, for this reaction. The negative sign (-) in the ΔH indicates that energy is being given off. This equation can also be written as:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)\ + 191\ kJ \text{ of heat}$$

What's important to remember about increasing the temperature of an equilibrium system, is the energy can be thought as just another product or reactant. In this example, you can clearly see that the 191 kJ are a product. Thus when the temperature of this system is raised, the effect will be the same as increasing any other product. For instance, as the temperature is increased, the equilibrium will shift away from the stress and result in increased concentration of each of the reactants and a decrease in concentration of the other product. When the temperature is increased for this reaction, the reaction will shift toward the reactants in an attempt to counteract the increased temperature. Therefore, the $[SO_3]$ will decrease and the $[SO_2]$ and $[O_2]$ will increase.

And, if the temperature for this equilibrium system is lowered, the equilibrium will shift to make up for this stress. When the temperature is decreased for this reaction, the reaction will shift toward the products in an attempt to counteract the decreased temperature. Hence, the $[SO_3]$ will increase and the $[SO_2]$ and $[O_2]$ will decrease.

TABLE 1.4: The Effect of Temperature on an Endothermic and an Exothermic Equilibrium System

	Exothermic $(-\Delta H)$	**Endothermic** $(+\Delta H)$
Increase Temperature	Shifts left, favors reactants	Shifts right, favors products
Decrease Temperature	Shifts right, favors products	Shifts left, favors reactants

Sample question: Predict the effect on the equilibrium position if the temperature is increased in each of the following.

(a) $H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g) \quad \Delta H = +40 kJ/mol$

(b) $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) + \text{energy}$

Solution:

(a) $H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g) \quad \Delta H = +40 kJ/mol$

The reaction is endothermic, because ΔH is positive, meaning heat is a reactant. With an increase in temperature for an endothermic reaction, the equilibrium will shift right, producing more products.

(b) $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) + \text{energy}$

The reaction is exothermic, meaning heat is a product. With an increase in temperature for an exothermic reaction, the equilibrium will shift left, producing more reactants.

The Haber Process

Let's look at a particularly useful reaction and how chemists applied Le Chatelier's Principle to make more of a desired product. The reaction between nitrogen gas and hydrogen gas can produce ammonia, NH_3. Under normal conditions, this reaction does not produce very much ammonia. Early in the 20^{th} century, the commercial use of this reaction was too expensive because of the small yield of ammonia. The reaction is as follows:

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)\ + \text{energy}$$

Fritz Haber, a German chemist working in the early years of the 20^{th} century, applied Le Châtelier's principle to help solve this problem. Decreasing the concentration of ammonia, for instance, by immediately removing it from the reaction container will cause the equilibrium to shift to the right and continue to produce more product.

One more factor that will affect this equilibrium system is the temperature. Since the forward reaction is exothermic (heat is released as a product), lowering the temperature will once again shift the equilibrium system to the right and increase the ammonia that is produced. Specifically the conditions that were found to produce the greatest yield of ammonia are 550°C (in commercial situations this is a "low" temperature) and 250 atm of pressure. Once the equilibrium system is producing the ammonia, the product is removed, cooled, and dissolved in water.

Lesson Summary

- Increasing the concentration of a reactant causes the equilibrium to shift to the right, producing more products.
- Increasing the concentration of a product causes the equilibrium to shift to the left, producing more reactants.
- Decreasing the concentration of a reactant causes the equilibrium to shift to the left, producing less products.
- Decreasing the concentration of a product causes the equilibrium to shift to the right, producing more products.

- For a forward exothermic reaction, an increase in temperature shifts the equilibrium toward the reactant side whereas a decrease in temperature shifts the equilibrium toward the product side. The opposite is true for an endothermic reaction.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki
- Tutorial: Le Chatlier's Principle: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/lechv17.swf

Vocabulary

Le Châtelier's Principle Applying a stress to a system at equilibrium causes the equilibrium position to shift to offset that stress and regain equilibrium.

Exothermic reaction A reaction in which heat is released, or is a product of the reaction.

Endothermic reaction A reaction in which heat is absorbed, or is a reactant of the reaction.

Review Questions

1. What is the effect on the equilibrium position if the [reactants] is increased?
2. What is the effect on the equilibrium position if the [reactants] is decreased?
3. For the reaction: $2N_2O_5(s) \rightleftharpoons 4NO_2(g) + O_2(g)$, what would be the effect on the equilibrium if:
   a. $[NO_2]$ decreases
   b. $[NO_2]$ increases
   c. $[O_2]$ increases
4. For the reaction: $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$, what would be the effect on the equilibrium system if:
   a. $[H_2O]$ increases
   b. mass of C decreases
   c. $[CO]$ increases
   d. $[H_2]$ decreases
5. Which direction will an equilibrium system shift for each of the following:
   a. adding energy to a forward exothermic equilibrium system
   b. adding energy to a reverse exothermic equilibrium system
   c. adding energy to a forward endothermic equilibrium system
   d. adding energy to a reverse endothermic equilibrium system

Predict the effect on the equilibrium position if the temperature is increased in each of the following.

6. $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g) \quad \Delta H = +51.9\ kJ$
7. $P_4O_{10}(s) + 6H_2O(l) \rightleftharpoons 4H_3PO_4(aq) + \text{heat}$
8. $Ag^+(aq) + Cl^-(aq) \rightleftharpoons AgCl(s) \quad \Delta H = -112\ kJ/mol.$
9. $NO_2(g) + NO(g) \rightleftharpoons N_2O(g) + O_2(g) \quad \Delta H = -43\ kJ$
10. $4\ NH_3(g) + 5\ O_2(g) \rightleftharpoons 4\ NO(g) + 6\ H_2O(g) + \text{heat}$

11. $2\,NOBr_{(g)} \rightleftharpoons 2\,NO(g) + Br_2(g) \quad \Delta H = +16.1kJ$. The reaction is at equilibrium in a closed container at $25°C$.
    a. What will happen to the concentration of NO if the temperature is increased?
    b. What will happen to the concentration of $NOBr$ if the temperature is increased?
    c. What will happen to the concentration of Br_2 if the temperature is increased?
12. In the following reaction, what would be the effect of each of the following changes to the system at equilibrium? $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5\ kJ/mol$
    a. increase O_2
    b. increase the temperature
13. Predict the effect on the equilibrium: $H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g) \quad \Delta H = -42\ kJ$ when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?
    a. Temperature is increased
    b. $[CO_2]$ decreases
    c. $[H_2O]$ increases
    d. $[H_2]$ decreases
14. Predict the effect on the chemical equilibrium $2\ SO_3(g) + \text{heat} \rightleftharpoons 2\ SO_2(g) + O_2(g)$, when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?
    a. Temperature is increased
    b. $[O_2]$ decreases
15. Predict the effect on the chemical equilibrium: $N_2O_4(g) + \text{heat} \rightleftharpoons 2\ NO_2(g)$, when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?
    a. Temperature is decreased
    b. $[N_2O_4]$ decreases