4 Chemical and Physical Changes

Learning Objectives

As you work through this chapter you will learn how to:

- distinguish between kinetic and potential energy.
- perform calculations involving energy units.
- distinguish between exothermic and endothermic processes.
- write balanced chemical equations.
- classify chemical reactions.
- determine whether an ionic compound is likely to dissolve in water.

4.1 Energy

In newspaper articles, television reports, consumer products and even popular music you often encounter references to energy ("energy conservation", "renewable energy sources", "energy drinks", "Taking all of my energy."). This is not surprising since essentially all changes involve the transfer of energy. **Energy** (E) comes in many forms but there are two general types - **kinetic energy** (K.E.) and **potential energy** (P.E.). Kinetic energy is associated with the motion of an object. The faster the object moves the more kinetic energy it possesses. Potential energy is stored energy, or energy that results from an object's position. For example, gravitational energy is the potential energy due to the force of gravity on an object; the stronger the gravitational attraction, the lower the potential energy. As a boulder is pushed up a hill it moves further away from the center of the earth and its potential energy increases as the force of gravity on the boulder decreases. In a similar way, energy can be stored in a spring as it is compressed. Where does this potential energy come from?

Consider a child on a swing (Fig. 4.1). In order to start the child swinging perhaps his parent pulls him upward (position (a) in Fig. 4.1) and then lets go. Energy is transferred to the child as **mechanical work**[1] is done in pulling the child and swing upward. This mechanical energy is stored as potential energy at the top of the arc. Mechanical work is also done to push a boulder up a hill or to compress a spring and this work is the source of the energy stored in these systems.

When the child is let go, the force of gravity pulls him and the swing downward and the stored potential energy is transformed into kinetic energy (motion). The kinetic energy reaches a maximum at the bottom of the arc and then decreases as it is transformed back into potential energy during the upward motion of the swing (position (b) in Fig. 4.1). At the top of the arc the kinetic energy is zero and the potential energy is at a maximum. This exchange between kinetic energy and potential energy is common; energy is readily converted from one type or form to another.

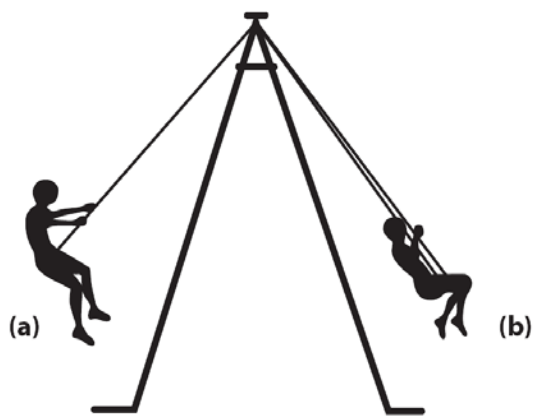

Figure 4.1 Transformation of potential and kinetic energy

[1]Perhaps you learned in a previous science class that *energy is the capacity to do work.* Mechanical work is done by exerting a force through a distance.

It is important to note that energy takes many forms in addition to the energy in the swing example. Some common energy forms are radiant (light) energy, chemical energy, thermal (heat) energy and electrical energy. Chemical and physical changes frequently absorb or release energy as heat.

Many experiments in the 1800s convinced scientists and engineers that indeed *energy is neither created nor destroyed, only converted from one form to another*. This statement is now known as the **first law of thermodynamics**. We can predict that the child on the swing will eventually come to rest at the bottom of the arc unless additional energy is transferred to the child and swing. At this point the child on the swing has minimum potential energy and zero kinetic energy. You might wonder what has happened to the energy the child and swing had previously. It has not been destroyed, only transformed into heat (thermal energy) due to friction at the pivot point and wind resistance.

The SI derived unit for energy is the **joule** (J). It takes about 20 J of heat to raise the temperature of a teaspoon of water by 1 °C. Chemical and physical processes typically release kilojoules (kJ) of energy. In the past, the **calorie** (cal) was used as a measure of heat energy and was defined by the heating of a sample of water.[2] The calorie is now defined in terms of the joule; 1 cal = 4.184 J (exactly). The calorie is a familiar unit, however, it is important to note that the energy in food is expressed in the **Calorie** (with an uppercase C) unit. One Calorie actually equals 1000 calories or 1 kcal. So the 80 Calories in that container of reduced-sugar yogurt actually corresponds to 80,000 cal or about 330 kJ.

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[2]One calorie was defined as the amount of heat needed to raise the temperature of 1 gram of water by one degree Celsius.

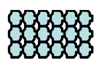
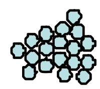
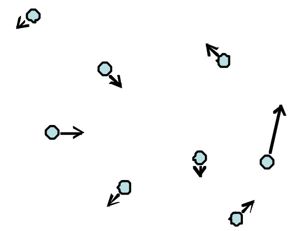
4.2 Energy Changes

When a physical or chemical process involves the absorption of heat, you might expect the temperature of the material to increase. What changes on the atomic level occur as this energy is absorbed? **Temperature** is a measure of the average kinetic energy of the atoms, molecules or ions that make up a sample of matter. The higher the temperature, the more energy of motion the particles have. Consider what happens when a given substance, like water, changes its physical state. Imagine starting with water as a solid (ice). A specific amount of water in the solid phase is characterized as having a fixed shape and fixed volume. The shape and volume are fixed because the water molecules do not have enough energy to move freely around each other. In fact, the molecules are held rather rigidly in place in a regular arrangement by electrostatic attractions. Particle motion is limited to vibrations within the structure of the solid.

As the ice is warmed, energy is absorbed and this causes the molecules to move more vigorously. As the motions become strong enough to overcome the attractions between molecules, they begin to move around each other. At this point the ice begins to melt and form liquid water. In the liquid phase the attractions between water molecules are not as strong as in the solid phase, so the liquid does not have a definite shape, however, it continues to have a definite volume.

As more energy is added to the liquid water, the temperature increases and molecular motion becomes more and more vigorous. Eventually the energy of motion is strong enough for the molecules to break away completely from each other to form the gas phase and the liquid starts to boil. Molecules in the gas phase are moving very rapidly and independently of each other. Consequently, a water vapor sample does not have a definite shape or a definite volume (it takes the shape and volume of its container). The molecules in the gas phase are also much farther apart than in the solid or liquid phase so the density of water vapor is much smaller. A representation of what happens to the particles of a substance as it is converted from a solid to a liquid to a gas is given in Table 4.1.

Table 4.1 Macroscopic and Microscopic Characteristics of Solids, Liquids and Gases

Solid	Liquid	Gas
		
- fixed shape and volume - not easily compressed - particles close together and locked in place by strong attractions	- fixed volume, takes shape of container - not easily compressed - particles close together but able to move around each other	- takes shape and volume of container - compressible - particles widely spaced and moving rapidly
increasing energy →		

It is convenient to use the following shorthand notation to describe both physical and chemical changes.

	A	→	B
physical change	(initial state)		(final state)
chemical change	(reactants)		(products)

Physical and chemical changes can be thought of as a conversion between some initial state and some final state. In a chemical change, the initial materials (**reactants**) undergo a reaction to form new substances (**products**). In this shorthand notation, *A* and *B* are usually the chemical formulas of the starting and final materials. The physical process of ice melting can be represented as:

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

As noted above, the conversion of ice into liquid water involves the absorption of energy (heat) as the molecules acquire more energy of motion. Processes that involve the absorption of heat are called **endothermic**. This characteristic of a physical or chemical change can be noted in the shorthand notation by indicating that heat is being added to the initial state. That is,

$$\text{heat} + H_2O(s) \rightarrow H_2O(l) \qquad \text{(endothermic)}$$

Notice that if we reverse this process, that is, if we consider the conversion of liquid water into ice, heat appears on the product's side. Processes that release heat are called **exothermic**.

$$H_2O(l) \rightarrow H_2O(s) + \text{heat} \qquad \text{(exothermic)}$$

For many physical processes it is often clear in which direction (forward or reverse) a process absorbs heat. For example, it may not be obvious that freezing is an exothermic process, however, it is much easier to determine that the reverse process (melting) is endothermic (ice melts when it is warmed). Consider both the forward and reverse processes in order to identify which is exothermic and which is endothermic. In summary,

$$\boxed{\text{exothermic process:} \quad A \rightarrow B + \text{heat}}$$

$$\boxed{\text{endothermic process:} \quad \text{heat} + A \rightarrow B}$$

Check for Understanding 4.1 — Solutions

1. Write the shorthand notation for the conversion of liquid water into water vapor. Is this process exothermic or endothermic? Explain.

2. When water and sulfuric acid (H_2SO_4) are mixed together the mixture becomes very warm to the touch. Is this mixing process exothermic or endothermic? Explain.

Another way to indicate the relative energy of state A and state B is by using what is known as a **relative energy diagram**. To see how this is done, imagine that we want to indicate the relative height of two individuals (Karen and Ali) in the class. First, we create a vertical axis with arbitrary height units and then suggest that Ali is taller than Karen by marking his height at a higher point along the vertical axis than that for Karen. The result might look like Figure 4.2.

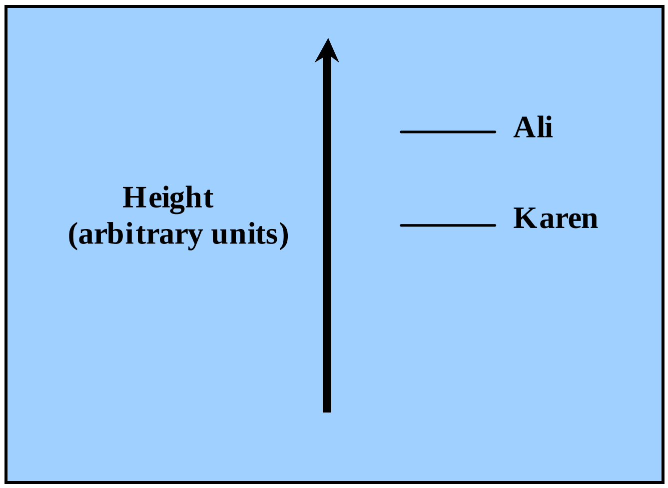


Figure 4.2 Relative height diagram

We commonly use a similar diagram to indicate the relative energy states involved in physical and chemical changes. Notice that the relative energy diagram for ice and liquid water in Figure 4.3 shows that the energy increases in going from the solid to the liquid and decreases in the reverse direction, and that the exact same amount of energy change is involved. This is a general pattern; the energy absorbed in an endothermic process is exactly equal in magnitude to the energy released in the reverse process.

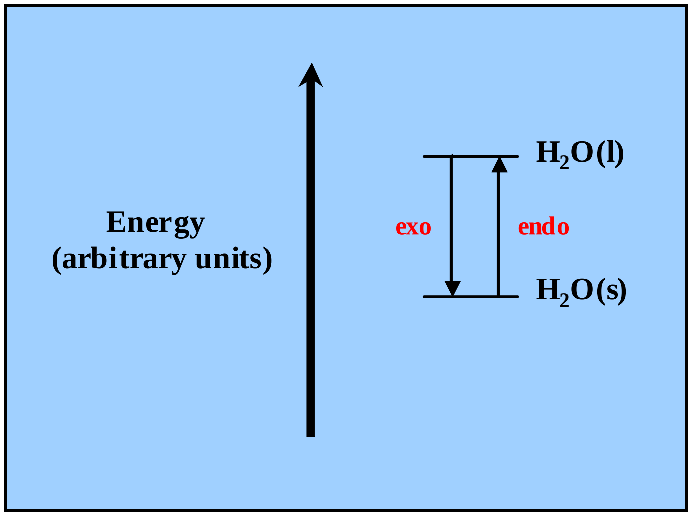


Figure 4.3 Relative energy diagram for ice and liquid water

The amount of energy change as you go from one state to another can be evaluated by taking the difference between the energy levels. This is expressed using the **delta function** (Δ). For a quantity X, ΔX is defined as $X_{\text{final}} - X_{\text{initial}}$. If the quantity is energy, $\Delta E = E_{\text{final}} - E_{\text{initial}}$. Notice that if the final state has a higher energy than the initial state ($E_{\text{final}} > E_{\text{initial}}$), it is an endothermic process and ΔE is positive. When the reverse is true ($E_{\text{final}} < E_{\text{initial}}$), it is an exothermic process and ΔE is negative.

Check for Understanding 4.2

Solution

1. On an energy level diagram indicate the relative energies of $H_2O(s)$ and $H_2O(g)$. Is the conversion of the solid to the gas phase, a process called **sublimation**, exothermic or endothermic? Explain.

We have already seen that as we heat up liquid water its temperature increases and molecular motion becomes more vigorous. Thus, the change in temperature (Δt) is directly proportional to the **heat** (q) added. Mathematically we can write

$$\Delta t \propto q$$

where $\propto$ means *proportional to*. The exact mathematical connection between the heat added and the change in temperature is given by

$$\begin{array}{ccccccc} \text{heat absorbed/released} & = & \text{mass} & \text{x specific heat} & \text{x} & \text{change in temperature} & \\ q & = & m & \text{x} \quad c & \text{x} & \Delta t & (4.1) \end{array}$$

where q is the heat added or released, in J;

m is the mass, in grams (g), of the sample;

c is the **specific heat** of the sample (J/g $\cdot$°C);

Δt is the temperature change (°C).

Let's think about equation 4.1 and why it is reasonable. Rearranging equation 4.1 gives:

$$\Delta t = \frac{q}{mc}$$

This indicates that Δt is directly proportional to the heat involved and inversely proportional to the mass of the sample and its specific heat. We have already noted the direct connection between the change in temperature (Δt) and the amount of heat absorbed or released (q). What about the connection between Δt and the two other quantities? Certainly the amount of the substance present is important. If you add a fixed amount of heat to a glass of water, it will increase its temperature much more than if this same amount of heat is added to a swimming pool of water. If the amount of substance is characterized by its mass, the temperature change is inversely proportional to the mass of the sample. Mathematically we can write:

$$\Delta t \propto \frac{1}{m}$$

This means that when the mass of a substance is large (like the swimming pool of water) the change in temperature will be small for a given amount of heat absorbed or released.

Another factor that is important to consider is how strong the attractions are between particles. If the attractions are strong, then it will take more energy to cause the particles to move vigorously. Specific heat (c) is a reflection of these interparticle attractions.[3] The specific heat of a substance is the amount of energy (heat) that must be added to 1 gram of the substance in order to increase its temperature by 1 °C. Substances with a high specific heat require more energy to increase their temperature than substances with a low specific heat. The specific heat values of some common pure substances are given in Table 4.2.

Table 4.2 Specific heat values of some common pure substances

Substance	Specific heat (J/g ·°C)
water (liquid)	4.18
aluminum	0.902
copper	0.385
gold	0.128
iron	0.450

The rather large value of specific heat for water means that it is a good heat sink. You need to add more energy to it before its temperature increases by one degree. Think about heating up water in an aluminum pan. The pan with its lower specific heat becomes hot much more rapidly than the water inside it. So temperature change and specific heat are inversely proportional.

$$\Delta t \propto \frac{1}{c}$$

specific heat example

[3]The specific heat of a substance depends on all of the ways in which energy, not just kinetic energy, can be stored in the substance, including the vibrations and rotations of molecules. This is one reason why the specific heat of water is much higher than that of the metals listed in Table 4.2.

The large specific heat of water is an important climate factor. It accounts for the mild temperature fluctuations in areas near large bodies of water (for example, coastal communities) and the harsh temperature changes in areas that lack water such as deserts.

Check for Understanding 4.3 Solutions

1. How many joules of energy are needed to raise the temperature of a hot tub containing 1500 kg of water by 5 °C?

2. Calculate the specific heat of a substance in J/g·°C if 4478 J raises the temperature of 1168 g of the material from 22.9 °C to 28.3 °C.

4.3 More Conservation Laws

The distinguishing characteristic of a chemical change is the formation of at least one new substance. This occurs when the atoms of the starting material (reactants) are rearranged to form new substances (products). Careful measurements in experiments during the 1780s by the French scientist Antoine Lavoisier led him to propose that in a chemical change mass is neither created nor destroyed (**law of conservation of mass**). For this important work and his many other scientific contributions, Lavoisier is known as the *father of modern chemistry*.

It wasn't until the twentieth century that instrumentation was precise enough to detect the very slight change in mass that is actually associated with chemical processes. When energy, such as heat or light, is released in a chemical reaction there is a very small decrease in the mass of the products compared to the mass of the starting material. This was first explained by Albert Einstein in his special theory of relativity in 1905 in which he proposed that mass and energy are two forms of the same thing. His famous $E = mc^2$ equation is a quantitative statement about mass-energy equivalence. This thinking required a restatement of the two separate conservation laws as the **law of conservation of mass and energy**: *in any change, the total amount of mass plus energy is a constant*.

Chemical changes are described in a similar manner as physical changes when heat is absorbed or released.

exothermic reaction:	reactants $\rightarrow$ products + heat

endothermic reaction:	heat + reactants $\rightarrow$ products

At this point it is rather difficult to determine whether a reaction is exothermic or endothermic just by looking at the reactants and products. In subsequent chemistry courses you will learn ways to calculate whether energy is released or absorbed in a chemical change; for now this information will be given.

4.4 Chemical Reactions and Balanced Chemical Equations

Let's look at chemical changes more carefully and learn how to use a type of shorthand to represent them. The initial step is to write a **word equation** for the chemical reaction. For example, when hydrogen gas and oxygen gas are heated the hydrogen gas burns and water is formed. This chemical change can be represented by the following word equation.

$$\text{hydrogen gas} + \text{oxygen gas} \rightarrow \text{liquid water} \qquad \text{(word equation)}$$

Chemists use a **chemical equation** as a shorthand form of a word equation. The first step in converting a word equation into a chemical equation is to substitute the correct chemical formula for each reactant and each product name; the result is an unbalanced chemical equation. For the reaction above,

$$H_2(g) + O_2(g) \rightarrow H_2O(l) \qquad \text{(unbalanced chemical equation)}$$

Notice that elemental hydrogen and oxygen are represented as diatomic (2-atom) molecules, the common stable form in which they occur. The physical state of each reactant and product is indicated by a letter designation in parentheses following the chemical formula; that is, by (s), (l), (g) or (aq).

The very slight change in mass associated with a chemical change does not invalidate the idea that each particular type of atom involved in a chemical reaction is neither created nor destroyed. That is, in a chemical change atoms are conserved, and a chemical equation must reflect this fact. A **balanced chemical equation** has the same number of atoms of each type on both sides of the equation. An equation is balanced by placing numerical coefficients in front of the reactant and product formulas to achieve equal numbers of each type of atom.[4] For example, in the unbalanced equation above, there are 2 hydrogen atoms on each side of the equation, however, there are 2 oxygen atoms on the reactant side but only 1 oxygen atom on the product side. This equation can be balanced by placing a coefficient of ½ in front of the O_2 as shown below. Now there is 1 O atom along with 2 H atoms on each side of the equation.

$$H_2(g) \;+\; \tfrac{1}{2}O_2(g) \;\rightarrow\; H_2O(l)$$

Generally, however, one wants to use the smallest set of whole-number coefficients to achieve atom balance. The fractional coefficient above can be converted into a whole number by multiplying it by 2. In order to maintain the atom balance, the other coefficients must also be multiplied by this same factor (2). Thus, the balanced chemical equation is:

$$2H_2(g) \;+\; O_2(g) \;\rightarrow\; 2H_2O(l) \qquad \text{(balanced chemical equation)}$$

Now we can see there are 4 H atoms and 2 O atoms on each side of the equation. Always

[4]A coefficient of one is implied and does not need to be written.

check to see that you have the same number of atoms of each type on both sides of the equation. Remember,

> The formulas of the reactants and products may not be altered in order to balance the equation, but you may add (whole-number) coefficients in front of the chemical formulas to balance the equation.

The chemical equations you will be asked to balance in this class are relatively simple and you can use a trial-and-error approach to determine the coefficients. In subsequent chemistry courses you will discover that such a method is inadequate and you will learn a more systematic approach to balancing chemical equations. You will also encounter chemical equations that have ions as reactants and/or products. In such cases a balanced equation has the same net electric charge on both sides. An example of how to write a balanced chemical equation from information about the chemical reaction is given below.

Example 4.1

Problem

Convert the following information about a chemical reaction into a balanced chemical equation with whole-number coefficients. Be sure to indicate the state of each reactant and product.

magnesium oxide is formed when magnesium metal reacts with oxygen gas

Solution

First, write a word equation for the reaction.

magnesium metal + oxygen gas → magnesium oxide

Next, substitute the correct chemical formula, including information about the state, for each reactant and product. Since the product is an ionic compound (How do you know this?), it is expected to be a solid composed of Mg^{2+} cations and O^{2-} anions.

$$Mg(s) + O_2(g) \rightarrow MgO(s)$$

Now balance atoms. Notice that the Mg atoms are balanced. Since there are 2 O atoms on the reactant side and only 1 O atom on the product side, a coefficient of ½ in front of O_2 will balance oxygen atoms.

$$Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s) \quad \text{(all atoms balanced)}$$

The fractional coefficient in front of O_2 can be converted to the smallest whole number by multiplying by 2. This requires that all other coefficients be multiplied by 2 in order to retain the atom balance. The resulting balanced equation is:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s) \quad \text{(all atoms balanced using whole-number coefficients)}$$

Check for Understanding 4.4 **Solutions**

1. Convert each of the following descriptions of chemical reactions into a chemical equation and balance it using whole-number coefficients. Be sure to indicate the physical state of each reactant and product.

   a) hydrochloric acid reacts with an aqueous solution of sodium hydroxide to form water and an aqueous solution of sodium chloride

   b) elemental carbon reacts with iron(III) oxide to form elemental iron and carbon monoxide gas

   c) lithium metal reacts with oxygen gas to form lithium oxide

4.5 Classification of Chemical Reactions

Chemical reactions are conveniently organized into general categories based on the type of chemical change that is occurring. Reactions may be classified as combination, decomposition, combustion or displacement reactions. In many instances, by determining the type of reaction that is occurring, you will be able to anticipate the reaction products. Now let's study the different reaction types and learn how to identify them.

Combination Reactions

In a **combination reaction** (or **synthesis reaction**) two or more substances react to form a single product. The general chemical equation for a combination reaction is

$$A + B \rightarrow C \qquad \text{(combination reaction)}$$

where A and B are either elements or compounds and C is a compound. The formation of water from hydrogen gas and oxygen gas is a combination reaction. Other examples of combination reactions are:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$
(sodium metal) (chlorine gas) (sodium chloride)

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$
(nitrogen gas) (hydrogen gas) (ammonia)

$$H_2O(l) + CO_2(g) \rightarrow H_2CO_3(aq)$$
(water) (carbon dioxide) (carbonic acid)

When the product of a combination reaction is a molecular compound, the formula is often difficult to anticipate and you will generally be given this information. However, when the product is an ionic compound you should be able to anticipate the correct formula of the product based on what you have learned about ions and ionic compounds in Chapter 3.

Decomposition Reactions

A **decomposition reaction** involves the breakdown of a single compound to form two or more products. The general chemical equation for a decomposition reaction is

$$A \rightarrow B + C \qquad \text{(decomposition reaction)}$$

where A is a compound and B and C are either elements or compounds or both. Some specific examples of decomposition reactions are:

$$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$$
(mercury(II) oxide) (mercury metal) (oxygen gas)

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

(calcium carbonate) (calcium oxide) (carbon dioxide)

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

(potassium chlorate) (potassium chloride) (oxygen gas)

The formulas of the products in a decomposition reaction are often difficult to anticipate. You can look for patterns to guide you, such as the above examples involving the decomposition of metal oxides, carbonates and chlorates, but generally you will be given this information.

Combustion Reactions

In this class we will define a **combustion reaction** as the reaction of a **hydrocarbon** (C_xH_y), a compound containing only carbon and hydrogen, with oxygen, or the reaction of a compound containing only carbon, hydrogen and oxygen ($C_xH_yO_z$) with oxygen.[5] The general chemical equations for a combustion reaction are:

$$C_xH_y + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

(combustion reactions)

$$C_xH_yO_z + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

The reactants C_xH_y or $C_xH_yO_z$ are molecular compounds and can exist as a solid, liquid or gas. This information will be specified. Combustion reactions are the processes involved in the burning of fuel such as natural gas (mostly CH_4) and gasoline (which contains compounds like octane, C_8H_{18}). The sugar in food (sucrose, $C_{12}H_{22}O_{11}$) also serves as a fuel when it is metabolized. Since fuels are a source of energy, we expect combustion reactions to be exothermic. The carbon dioxide and water produced in combustion reactions are not the only substances formed when fuels react with oxygen.

[5]In a broader sense, a combustion reaction can be thought of as any reaction in which an element or compound reacts with oxygen.

Compounds like CO (carbon monoxide) also form and are one source of pollutants associated with burning fuel. We will write rather simple balanced equations for combustion reactions with only CO_2 and H_2O as the products. Some specific examples of combustion reactions are:

$CH_4(g)$	+	$2O_2(g)$	→	$CO_2(g)$	+	$2H_2O(l)$
(methane)		(oxygen gas)		(carbon dioxide)		(water)

$C_{12}H_{22}O_{11}(s)$	+	$12O_2(g)$	→	$12CO_2(g)$	+	$11H_2O(l)$
(sucrose)		(oxygen gas)		(carbon dioxide)		(water)

When balancing combustion reaction equations, start by balancing the carbon atoms and hydrogen atoms before attempting to balance the oxygen atoms. Then adjust the coefficient in front of O_2 to balance the oxygen atoms. If you need a fractional coefficient in front of O_2 to balance the O atoms, convert this to the smallest whole number possible by multiplying this (and all other coefficients) by the appropriate whole number. This is illustrated in Example 4.2. When balancing equations that involve the combustion of a $C_xH_yO_z$ compound, don't forget to count the oxygen atoms in this compound for O atom balance.

Example 4.2

Problem

Complete and balance the following equation.

$$C_2H_6(g) \ + \ O_2(g) \ \rightarrow$$

Solution

This is a reaction between a hydrocarbon and oxygen so we know it is a combustion reaction. Therefore the products are $CO_2(g)$ and $H_2O(l)$.

$$C_2H_6(g) \ + \ O_2(g) \ \rightarrow \ CO_2(g) \ + \ H_2O(l)$$

First, balance the carbon atoms. Since there are 2 C atoms on the reactant side and only 1 C atom on the product side, a coefficient of 2 in front of CO_2 will balance carbon atoms.

$$C_2H_6(g) \ + \ O_2(g) \ \rightarrow \ 2CO_2(g) \ + \ H_2O(l) \qquad \text{(C atoms balanced)}$$

Now balance the hydrogen atoms. Since there are 6 H atoms on the reactant side and only 2 H atoms on the product side, a coefficient of 3 in front of H_2O will balance hydrogen atoms.

$$C_2H_6(g) \ + \ O_2(g) \ \rightarrow \ 2CO_2(g) \ + \ 3H_2O(l) \qquad \text{(C and H atoms balanced)}$$

Finally, balance the oxygen atoms. Since there are 2 O atoms on the reactant side and 7 O atoms on the product side, multiplying O_2 by 3½, or 7/2, will balance oxygen atoms.

$$C_2H_6(g) \ + \ 7/2O_2(g) \ \rightarrow \ 2CO_2(g) \ + \ 3H_2O(l) \qquad \text{(all atoms balanced)}$$

The fractional coefficient in front of O_2 can be converted to the smallest whole number by multiplying by 2. This requires that all other coefficients be multiplied by 2 in order to retain the atom balance. The resulting balanced equation is:

$$2C_2H_6(g) \ + \ 7O_2(g) \ \rightarrow \ 4CO_2(g) \ + \ 6H_2O(l) \qquad \text{(all atoms balanced using whole-number coefficients)}$$

Displacement Reactions

Displacement reactions may be one of two types - either single displacement or double displacement. The general equation for a **single displacement reaction** is

$$A + BC \rightarrow AC + B \qquad \text{(single displacement reaction)}$$

A and B are elemental forms and BC and AC are compounds. In the general reaction, element A displaces element B from BC to form AC. Typically both A and B are metals (including hydrogen) or both are halogens. Some specific examples of single displacement reactions are:

$$\underset{\text{(zinc metal)}}{Zn(s)} + \underset{\text{(hydrochloric acid)}}{2HCl(aq)} \rightarrow \underset{\text{(zinc chloride)}}{ZnCl_2(aq)} + \underset{\text{(hydrogen gas)}}{H_2(g)}$$

$$\underset{\text{(lithium metal)}}{2Li(s)} + \underset{\text{(water)}}{2H_2O(l)} \rightarrow \underset{\text{(lithium hydroxide)}}{2LiOH(aq)} + \underset{\text{(hydrogen gas)}}{H_2(g)}$$

$$\underset{\text{(chlorine gas)}}{Cl_2(g)} + \underset{\text{(potassium iodide)}}{2KI(aq)} \rightarrow \underset{\text{(iodine)}}{I_2(s)} + \underset{\text{(potassium chloride)}}{2KCl(aq)}$$

The general equation for a **double displacement reaction** is

$$AB + CD \rightarrow AD + CB \qquad \text{(double displacement reaction)}$$

where A and C are usually cations and B and D are usually anions. Some specific examples of double displacement reactions are:

$$\underset{\text{(lithium hydroxide)}}{LiOH(aq)} + \underset{\text{(hydrobromic acid)}}{HBr(aq)} \rightarrow \underset{\text{(lithium bromide)}}{LiBr(aq)} + \underset{\text{(water)}}{H_2O(l)}$$

$$3CaCl_2(aq) \quad + \quad 2Na_3PO_4(aq) \quad \rightarrow \quad Ca_3(PO_4)_2(s) \quad + \quad 6NaCl(aq)$$

(calcium chloride) (sodium phosphate) (calcium phosphate) (sodium chloride)

$$(NH_4)_2S(aq) \quad + \quad 2HCl(aq) \quad \rightarrow \quad H_2S(g) \quad + \quad 2NH_4Cl(aq)$$

(ammonium sulfide) (hydrochloric acid) (hydrogen sulfide) (ammonium chloride)

4.6 Solubility of Ionic Compounds

Notice that double displacement reactions generally involve reactants in aqueous solution and frequently produce ionic compounds. We have noted that pure ionic compounds are solids at room temperature, however, when placed in contact with water an ionic compound may dissolve (it is **soluble** in water) or remain largely undissolved as a solid (it is **insoluble**). In the second double displacement example above, both products are ionic compounds, however, one of them ($Ca_3(PO_4)_2$) is insoluble and the other (NaCl) is soluble. An insoluble product formed in aqueous solution is called a **precipitate**.

You might wonder if there is any way to anticipate whether an ionic compound is likely to be soluble or insoluble in water. The answer is yes. In fact there are two ways to determine if a compound will be soluble in water. One approach is to memorize the behavior of different categories of ionic compounds (for example, compounds containing nitrate anion). A list these trends is called the **solubility rules**. One such rule is that *all nitrates are soluble in water*. A typical list of rules covers a dozen categories of compounds and lists important exceptions. However, there is another approach you can take that will allow you to predict whether an ionic compound is soluble or insoluble in water without having to memorize a lengthy list. In fact, it makes use of the information you learned in Chapter 3.

First, let's consider what happens when an ionic compound dissolves in water. Recall that an ionic compound consists of a 3-dimensional array of oppositely-charged ions held in place by strong electrostatic attractions (see Fig. 3.7). When an ionic compound dissolves the ions are separated from each other so that they are free to move around in solution. The difficulty of separating these ions depends on the strength of the electrostatic attractions holding them together. Although other factors are involved, the solubility of ionic compounds generally correlates with the force of attraction between cations and anions. For many ionic compounds, the stronger the force of attraction the less soluble the ionic compound is in water. This simplified thinking allows us to predict whether an ionic compound is soluble in water or not by considering the charges of the ions involved.

If one looks carefully at the various ‘solubility rules’, it becomes clear that ionic compounds containing either (+1) or (-1) ions, or both, tend to be soluble; these ions are held together by weaker attractions. Those compounds that contain combinations of (+2) or (+3) cations with either (-2) or (-3) anions tend to be insoluble; these ions are held together by stronger attractions. Although there are many exceptions to these guidelines, the most important ones are: (i) ionic compounds containing hydroxide ion (OH^-) in combination with (+2) or (+3) ions, for example $Cu(OH)_2$, tend to be insoluble, and (ii) sulfates (SO_4^{2-}) tend to be soluble with rare exceptions. You can use these guidelines to anticipate whether an ionic compound formed in solution will remain dissolved or form a precipitate.

Ionic compound solubility guidelines

soluble ionic compounds contain (+1) and/or (-1) ions
(main exception: compounds containing (+2) or (+3) cations and OH^-)

insoluble ionic compounds contain combinations of (+2) or (+3) cations with either (-2) or (-3) anions
(main exception: most compounds containing SO_4^{2-})

Check for Understanding 4.5

Solutions

1. Classify each of the following reactions as combination, decomposition, combustion, single displacement or double displacement, then complete and balance each equation. Indicate the physical state of each product.

   a) $CH_3OH(l) + O_2(g) \rightarrow$

   b) $Pb(s) + AgNO_3(aq) \rightarrow$

   c) $Bi(NO_3)_3(aq) + H_2S(aq) \rightarrow$

2. Indicate whether each of the following compounds is expected to be soluble or insoluble in water.

   a) $Cu(NO_3)_2$ b) Fe_2S_3 c) $Zn(OH)_2$ d) K_3PO_4

Chapter 4 Keywords

Glossary

energy	exothermic	synthesis reaction
kinetic energy	relative energy diagram	decomposition reaction
potential energy	delta function	combustion reaction
mechanical work	sublimation	hydrocarbon
first law of thermodynamics	heat	single displacement reaction
joule	specific heat	double displacement reaction
calorie	law of conservation of mass	soluble
Calorie	word equation	insoluble
temperature	chemical equation	solubility rules
reactants	balanced chemical equation	precipitate
products	combination reaction	
endothermic	law of conservation of mass and energy	

Supplementary Chapter 4 *Check for Understanding* questions

Chapter 4 Exercises

Answers

(You may use a periodic table as needed.)

1. What happens to the kinetic energy of a moving bus as it comes to a stop?

2. If energy is never destroyed, why are we encouraged to conserve energy?

3. An example of an endothermic process is:

   A. digestion of food
   B. burning of coal
   C. melting of ice
   D. all of the above

4. In an experiment, if the initial temperature is 72.1 °C and the final temperature is 26.9 °C, what is the temperature change (Δt) for this experiment?

5. In which of the following is energy (heat) released?

A. evaporation of water from your skin
B. conversion of carbon dioxide gas into dry ice (solid carbon dioxide)
C. digestion of food
D. all of the above
E. only B and C

6. If 25 g of iron (c = 0.45 J/g·°C) and 25 g of silver (c = 0.24 J/g·°C), both at room temperature, are dropped into separate glasses of 50 mL of water at 90 °C, which metal will reach the higher temperature? Explain.

7. How does the amount of energy needed to heat a sample of water from 16 °C to 42 °C compare to that needed to heat this same sample from 51 °C to 77 °C? Explain.

8. A 280-g metal bar requires 5.31 kJ to change its temperature from 21 °C to 100 °C. What is the specific heat of the metal?

9. Consider a sample of water in the gaseous state. Describe what happens to the molecules in the sample as the sample is slowly cooled until it liquefies and then solidifies.

10. What important idea is behind the practice of writing balanced chemical equations?

11. Convert each of the following word equations into a chemical equation and then balance it. Indicate the physical state of each reactant and product.

a) sodium metal + oxygen gas → sodium oxide

b) sodium bicarbonate → sodium carbonate + water + carbon dioxide

c) hydrochloric acid + nickel metal → hydrogen gas + aqueous nickel chloride

12. Balance each of the following chemical equations. Classify each reaction as combination, decomposition, combustion, single displacement or double displacement.

a) $Na_3PO_4(aq) + BaCl_2(aq) \rightarrow NaCl(aq) + Ba_3(PO_4)_2(s)$

b) $SO_2(g) + O_2(g) \rightarrow SO_3(g)$

c) $H_2CO_3(aq) + KOH(aq) \rightarrow K_2CO_3(aq) + H_2O(l)$

d) $Fe_2O_3(s) + Al(s) \rightarrow Fe(s) + Al_2O_3(s)$

e) $C_3H_5(NO_3)_3(l) \rightarrow CO_2(g) + H_2O(g) + N_2(g) + O_2(g)$

13. In which of the following compounds are the electrostatic attractions between oppositely-charged ions likely to be the strongest?

A. LiCl
B. Li_2O
C. FeO
D. $FeCl_2$

14. Which of the following substances is likely to be insoluble in water?

A. potassium bromide
B. sodium carbonate
C. lead(II) sulfide
D. copper(II) acetate
E. all of these

15. When FeS(s) reacts with HCl(aq) the products of this reaction are:

A. FeCl(aq) and HS(aq)
B. $FeCl_2(s)$ and HS(aq)
C. FeCl(aq) and $H_2S(aq)$
D. $FeCl_2(aq)$ and $H_2S(aq)$

16. Classify each of the following reactions as combination, decomposition, combustion, single displacement or double displacement, then complete and balance each equation. Indicate the physical state of each product.

a) $Mg(OH)_2(s) + H_3PO_4(aq) \rightarrow$

b) $C_2H_6O_2(l) + O_2(g) \rightarrow$

c) $Cl_2(aq) + NaI(aq) \rightarrow$

d) $Ag(s) + O_2(g) \rightarrow$

17. When each of the following pairs of aqueous solutions is mixed, a precipitate forms. Write the correct chemical formula for the precipitate in each case.

iron(III) perchlorate and potassium hydroxide ____________

nickel acetate and sodium sulfide ____________

Chapter 4

Check for Understanding 4.1

1. Write the shorthand notation for the conversion of liquid water into water vapor. Is this process exothermic or endothermic? Explain.

Answer: heat + $H_2O(l) \rightarrow H_2O(g)$ (endothermic)

Solution

Energy must be added to cause molecules in the liquid phase to move fast enough to break away from each other and form a gas. Since heat is added this process is endothermic.

2. When water and sulfuric acid (H_2SO_4) are mixed together the mixture becomes very warm to the touch. Is this mixing process exothermic or endothermic? Explain.

Answer: exothermic

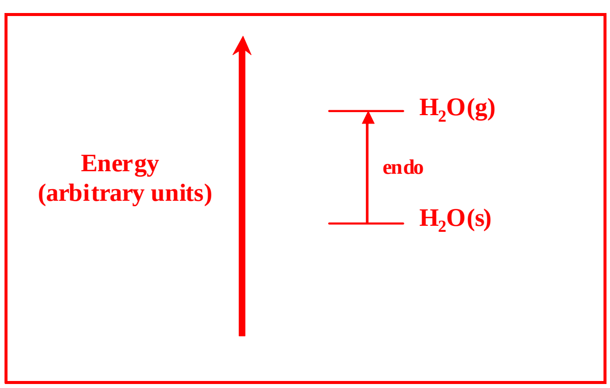
Solution

Since the mixture becomes warm, heat must be released during the mixing. Such a process is exothermic.

Check for Understanding 4.2

1. On an energy level diagram indicate the relative energies of $H_2O(s)$ and $H_2O(g)$. Is the conversion of the solid to the gas phase, a process called **sublimation**, exothermic or endothermic? Explain.

Answer:



Solution

Energy must be added to cause the molecules in the solid phase to move fast enough to break away from each other and form a gas so the energy of the gas phase is higher than that of the solid. Since heat is added the process is endothermic.

Check for Understanding 4.3

1. How many joules of energy are needed to raise the temperature of a hot tub containing 1500 kg of water by 5 °C?

Answer: 3.1×10^7 J

Solution

The amount of heat energy (q) needed is calculated from $q = m \cdot c \cdot \Delta t$. For water, the specific heat (c) equals 4.18 J/g·°C. The mass of water (m) in grams is 1500 x 10^3 g. The temperature change (Δt) is 5 °C. Substituting these values and calculating q gives:

$$q = (1500 \times 10^3 \cancel{\text{g}})(4.18\,\text{J}/\cancel{\text{g}} \cdot \cancel{^\circ\text{C}})(5\,\cancel{^\circ\text{C}}) = 3.1 \times 10^7\,\text{J}$$

The answer is given to 2 significant figures because of the 2 significant figures in the mass. It is assumed that the temperature change is an exact number.

2. Calculate the specific heat of a substance in J/g·°C if 4478 J raises the temperature of 1168 g of the material from 22.9 °C to 28.3 °C.

Answer: 0.71 J/g·°C

Solution

The specific heat (c) can be calculated by rearranging equation 4.1 to give:

$$c = \frac{q}{m\Delta t}$$

The amount of heat added (q) is 4478 J. The mass of the substance (m) in grams is 1168 g and the temperature change (Δt) is 28.3 °C - 22.9 °C = 5.4 °C. Substituting these values and calculating c gives:

$$c = \frac{4478\,\text{J}}{(1168\,\text{g})(5.4\,^\circ\text{C})} = 0.71\frac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}$$

The answer is given to 2 significant figures because of the 2 significant figures in the temperature change.

Check for Understanding 4.4

1. Convert each of the following descriptions of chemical reactions into a chemical equation and balance it using whole-number coefficients. Be sure to indicate the physical state of each reactant and product.

   a) hydrochloric acid reacts with an aqueous solution of sodium hydroxide to form water and an aqueous solution of sodium chloride

   b) elemental carbon reacts with iron(III) oxide to form elemental iron and carbon monoxide gas

   c) lithium metal reacts with oxygen gas to form lithium oxide

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

b) $3C(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + 3CO(g)$

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

<u>Solutions</u>

First, write the chemical formula for each reactant and product, along with its physical state. Then use whole-number coefficients in front of the reactant/product formulas to balance atoms on both sides of the equation.

a) Reactant formulas:

hydrochloric acid - The *hydro-* prefix indicates that this is a binary acid and the stem *chlor-* tells you that the nonmetal chlorine is bonded to hydrogen. Thus, the binary acid formula will look like $H_nCl(aq)$. In order to determine the value for n, remember that in water this molecule will break up to form H^+ ions and chloride anions. The charge on the cation must be balanced by the charge on the anion that is formed. Since chlorine is in Group 7A we expect it to form an anion with a minus one charge (Cl^-). Thus, only one H^+ is needed for each Cl^- so $n = 1$ and the correct formula is $HCl(aq)$.

aqueous solution of sodium hydroxide - The metal cation is Na^+. The polyatomic anion formula is OH^-. Since the ion charges are the same, it takes one cation to balance the charge on the anion and the compound formula is $NaOH(aq)$.

Product formulas:

water - $H_2O(l)$

aqueous solution of sodium chloride - The metal cation is Na^+. The anion formula is Cl^-. Since the ion charges are the same, it takes one cation to balance the charge on the anion and the compound formula is NaCl(aq).

This results in the equation:

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

In this case, all of the atoms are balanced so no adjustment of the coefficients is necessary.

b) Reactant formulas:

elemental carbon - This nonmetal is a solid at room temperature with the formula C(s).

iron(III) oxide - The Roman numeral indicates that the cation is Fe^{3+}. The oxide ion formula is O^{2-}. Using the value of the anion charge as the subscript for the cation and the value of the cation charge as the subscript for the anion results in the formula Fe_2O_3. This ionic compound is a solid at room temperature.

Product formulas:

elemental iron - This metal is a solid at room temperature with the formula Fe(s).

carbon monoxide gas - The name suggests a binary molecular compound with the formula CO(g).

This results in the unbalanced equation:

$$C(s) \quad + \quad Fe_2O_3(s) \quad \rightarrow \quad Fe(s) \quad + \quad CO(g)$$

To balance this equation, place a coefficient of 3 in front of CO to balanced oxygen atoms. This requires a coefficient of 3 in front of C to balanced carbon atoms. Finally, a coefficient of 2 in front of Fe will balance iron atoms. The balanced equation is:

$$3C(s) \quad + \quad Fe_2O_3(s) \quad \rightarrow \quad 2Fe(s) \quad + \quad 3CO(g)$$

c) Reactant formulas:

lithium metal - This metal is a solid at room temperature with the formula Li(s).

oxygen gas - This element exists as a diatomic molecule with the formula $O_2(g)$.

Product formulas:

lithium oxide - The metal cation is Li^+. The oxide ion formula is O^{2-}. Since it takes 2 cations to balance the charge on each anion, the formula is Li_2O. This ionic compound is a solid at room temperature.

This results in the unbalanced equation:

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

To balance this equation, place a coefficient of 2 in front of Li_2O to balanced oxygen atoms. This requires a coefficient of 4 in front of Li to balanced lithium atoms. The balanced equation is:

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

Check for Understanding 4.5

1. Classify each of the following reactions as combination, decomposition, combustion, single displacement or double displacement, then complete and balance each equation. Indicate the physical state of each product.

   a) $CH_3OH(l) \ + \ O_2(g) \ \rightarrow$

   b) $Pb(s) \ + \ AgNO_3(aq) \ \rightarrow$

   c) $Bi(NO_3)_3(aq) \ + \ H_2S(aq) \ \rightarrow$

Answers:

a) combustion $2CH_3OH(l) \ + \ 3O_2(g) \ \rightarrow \ 2CO_2(g) \ + \ 4H_2O(l)$

b) single displacement $Pb(s) \ + \ 2AgNO_3(aq) \ \rightarrow \ 2Ag(s) \ + \ Pb(NO_3)_2(aq)$

c) double displacement $2Bi(NO_3)_3(aq) \ + \ 3H_2S(aq) \ \rightarrow \ Bi_2S_3(s) \ + \ 6HNO_3(aq)$

Solutions

a) This is a reaction between a $C_xH_yO_z$ compound and oxygen so it is a combustion reaction that forms $CO_2(g)$ and $H_2O(l)$ as products.

$CH_3OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

Notice that the carbon atoms are already balanced, so balance the hydrogen atoms. Since there are 4 H atoms on the reactant side and only 2 H atoms on the product side, a coefficient of 2 in front of H_2O will balance hydrogen atoms.

$CH_3OH(g) + O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ (C and H atoms balanced)

Finally, balance the oxygen atoms. Since there are 3 O atoms on the reactant side and 4 O atoms on the product side, multiplying O_2 by 3/2 will balance oxygen atoms.

$CH_3OH(g) + 3/2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ (all atoms balanced)

The fractional coefficient in front of O_2 can be converted to the smallest whole number by multiplying by 2. This requires that all other coefficients be multiplied by 2 in order to retain the atom balance. The resulting balanced equation is:

$2CH_3OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$ (all atoms balanced using whole-number coefficients)

b) This reaction between an elemental form and an ionic compound should suggest a single displacement reaction of the general form A + BC → AC + B where A and B are elemental forms and BC and AC are compounds. The metal lead (Pb) displaces the silver ion in the compound to form elemental silver, Ag(s), and an ionic compound consisting of lead(II) ions and nitrate (NO_3^-) ions. Recall that the common ion of lead is Pb^{2+}. This means that the formula of the compound formed between Pb^{2+} and NO_3^- is $Pb(NO_3)_2$. Since this ionic compound contains a -1 anion that is not OH^-, this compound should be soluble in water. The unbalanced equation is:

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

At this stage Pb and Ag are balanced. Notice that the polyatomic nitrate ion does not undergo any change in composition in this reaction. This means that you can balance NO_3^- units just like an individual atom. Since there are 2 nitrate ions on the product side, a coefficient of 2 is needed in front of $AgNO_3$ on the reactant side. This upsets the Ag balance. This can be restored by putting a coefficient of 2 in front of Ag on the product side. The balanced equation is:

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

c) This reaction between an ionic compound and a binary acid should suggest a double displacement reaction of the general form AB + CD → AD + CB. In this case the metal bismuth (Bi) and hydrogen in the reactants switch places to form an ionic compound between bismuth and sulfur and an oxoacid. Notice that the ionic compound reactant ($Bi(NO_3)_3$) involves Bi^{3+} and NO_3^- ions. Since sulfur is expected to form a S^{2-} ion, the correct formula for the ionic compound formed is Bi_2S_3. This combination of a +3 cation and a -2 anion is expected to insoluble. The oxoacid formed from H^+ and NO_3^- has the formula HNO_3 and forms an aqueous solution. This results in the unbalanced equation:

$$Bi(NO_3)_3(aq) + H_2S(aq) \rightarrow Bi_2S_3(s) + HNO_3(aq)$$

To balance this equation, place a coefficient of 3 in front of H_2S to balanced sulfur atoms. This requires a coefficient of 6 in front of HNO_3 to balance hydrogen atoms. Finally, a coefficient of 2 in front of $Bi(NO_3)_3$ will balance both the bismuth atoms and the nitrate units. The balanced equation is:

$$2Bi(NO_3)_3(aq) + 3H_2S(aq) \rightarrow Bi_2S_3(s) + 6HNO_3(aq)$$

2. Indicate whether each of the following compounds is expected to be soluble or insoluble in water.

a) $Cu(NO_3)_2$ b) Fe_2S_3 c) $Zn(OH)_2$ d) K_3PO_4

Answers: a) soluble

b) insoluble

c) insoluble

d) soluble

Solutions

Recall the solubility guidelines: (I) soluble ionic compounds contain (+1) and/or (-1) ions (main exception: compounds containing (+2) or (+3) cations and OH^-), and (ii) insoluble ionic compounds contain combinations of (+2) or (+3) cations with either (-2) or (-3) anions (main exception: most compounds containing SO_4^{2-}).

a) The -1 anion, since it is not OH^-, suggests this compound is soluble in water.

b) The combination of a +3 cation and -2 anion suggests this compound is insoluble in water.

c) The combination of a +2 cation with OH^- suggests this compound is insoluble in water.

d) The +1 cation suggests this compound is soluble in water.

Chapter 4

1. It is converted into heat.

3. C

4. -45.2 °C

5. E

6. silver

7. The amount of thermal energy (heat) is the same since the same sample (same mass and specific heat) is undergoing the same temperature change ($\Delta t = 26$ °C). Recall $q = m \times c \times \Delta t$.

8. 0.24 J/g·°C

9. As water molecules in the gas phase are cooled they begin to slow down and eventually interact strongly enough so that they can only move around each other as the liquid state forms. Further cooling causes the molecules to move even more slowly, eventually forming a 3-dimensional arrangement (lattice) in which each molecule is locked into place by its strong attractions to other molecules as the liquid freezes.

10. Atoms are neither created nor destroyed in chemical reactions.

11. a) $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

    b) $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$

    c) $2HCl(aq) + Ni(s) \rightarrow H_2(g) + NiCl_2(aq)$

12. a) $2Na_3PO_4(aq) + 3BaCl_2(aq) \rightarrow 6NaCl(aq) + Ba_3(PO_4)_2(s)$ (double displacement)

b) $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ (combination)

c) $H_2CO_3(aq) + 2KOH(aq) \rightarrow K_2CO_3(aq) + 2H_2O(l)$ (double displacement)

d) $Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(s) + Al_2O_3(s)$ (single displacement)

e) $4C_3H_5(NO_3)_3(l) \rightarrow 12CO_2 + 10H_2O + 6N_2 + O_2$ (decomposition)

13. C

14. C

15. D

16. a) double displacement $3Mg(OH)_2(s) + 2H_3PO_4(aq) \rightarrow Mg_3(PO_4)_2(s) + 6H_2O(l)$

b) combustion $2C_2H_6O_2(l) + 5O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

c) single displacement $Cl_2(aq) + 2NaI(aq) \rightarrow 2NaCl(aq) + I_2(s)$

d) combination $4Ag(s) + O_2(g) \rightarrow 2Ag_2O(s)$

17. iron(III) perchlorate and potassium hydroxide ____$Fe(OH)_3$____

nickel acetate and sodium sulfide ___NiS___