EXPERIMENT 10

Chemical Equations: A Study Assignment

Performance Goal

10–1 Given information from which you can write formulas for all reactants and all products for each of the following types of reactions, write the balanced chemical equation for the reactions:

Double replacement

Double replacement—precipitation

Double replacement—acid/base (neutralization)

Combination (synthesis)

Decomposition

Complete oxidation or burning of organic compounds (combustion)

Single replacement (redox)

Other reactions in which reactants and products are identified

CHEMICAL OVERVIEW

A chemist uses a chemical equation to describe a chemical change. The general form of a chemical equation is

$$\text{Reactant 1} + \text{Reactant 2} + \ldots \rightarrow \text{Product 1} + \text{Product 2} + \ldots$$

The substances that enter into the reaction are called **reactants**. They are identified by their chemical formulas, written on the left side of the equation, and separated from each other by plus signs. The formulas of the new substances produced in the reaction, called **products**, are written on the right side, again separated by plus signs. The two sides of the equation are separated by an arrow pointing from the reactants to the products, indicating that the reactants are changed into the products. In reading a chemical equation, or expressing it in words, the arrow is frequently read as ''yields,'' ''produces,'' or ''forms''; any other term that suggests the creation of a substance not originally present is equally satisfactory.

Symbols are frequently added to chemical equations to indicate the conditions under which the reaction occurs. The symbols (s), (*l*), or (g) immediately after the formula of a substance indicate that the substance is in the solid, liquid, or gaseous state, respectively. A substance that is in aqueous (water) solution may have (aq) after its formula. Sometimes the arrow between reactants and products is lengthened, and words, formulas, temperatures, or other symbols are written above (or above and below) the arrow to indicate reaction conditions or other substances in the reaction vessel. None of these supplementary items will be used in this exercise; however, if your instructor requests that you use them, you should, of course, follow his or her directions.

A chemical equation does two things. First, it tells you what substances are involved in a chemical change. To do this accurately, it is essential that the substances be represented by their correct chemical formulas. It is assumed in this exercise that, given the name of a chemical, you are able to write its formula. Second, an equation has quantitative significance. It obeys the law of conservation of mass, which indicates that the total mass of all the reactants is equal to the total mass of all the products in an ordinary chemical change. In order for this to be true, the equation must have equal numbers of atoms of each individual element on the two sides of the equation. The equation is then said to be **balanced**.

These two characteristics of an equation lead to a simple two-step procedure by which an equation may be written:

1. Write the correct chemical formula for each reactant on the left and each product on the right.
2. *Using coefficients only,* balance the number of atoms of each element on each side of the equation. If no coefficient is written, its value is assumed to be one (1).

It is impossible to overemphasize the importance of following these two steps literally, and keeping them independent. In Step 1, write the correct formulas without concern about where the atoms come from, or how many atoms of an element may be present in some species on the other side of the equation. In Step 2, be sure that you balance the atoms of each element by placing whole-number coefficients in front of chemical formulas, and by no other means. Specifically,

DO NOT change a correct chemical formula in order to balance an element;

DO NOT add some real or imaginary chemical species to either side to make an element balance.

Quite often the word description of a chemical reaction will not identify all of the species that must be included in the equation. If you are familiar with the kinds of reactions described in the performance goal, you will be able to identify the substances not mentioned. The reaction types will be discussed as they are encountered.

EXAMPLES

The following examples are in the form of a program in which you learn by answering a series of questions. Obtain an opaque shield (a piece of cardboard, or a folded piece of paper you cannot see through) that is wide enough to cover this page. In each example place the shield on the book page so it covers everything beneath the first dotted line that runs across the page. Read to that point, and write in the space provided whatever is asked. Then lower the shield to the next dotted line. The material exposed will begin with the correct response to the question you have just answered. Compare this answer to yours, looking back to correct any misunderstanding if the two are different. When you fully understand the first step, read to the next dotted line and proceed as before.

A. Combination Reactions

A combination reaction occurs when two or more substances combine to form a single product. The reactants may be elements or compounds, perhaps one or more of each. Quite often the description of the reaction will give the chemical name of the product only. For example, the equation for the reaction in which sodium chloride is formed from its elements is $2\ Na + Cl_2 \rightarrow 2\ NaCl$. An example of a combination reaction between compounds is $CaO + H_2O \rightarrow Ca(OH)_2$.

Example 1

Write the equation showing how magnesium oxide is formed from its elements.

''magnesium oxide is formed'' indicates that magnesium oxide is the product of the reaction, so its formula will appear on the right side of the equation. ''from its elements'' identifies magnesium and oxygen as the reactants whose formulas will be written to the left of the arrow. Complete Step 1 of the procedure by writing the unbalanced equation.

....

1a. $Mg + O_2 \rightarrow MgO$

Remember, oxygen is a diatomic element; its correct formula is therefore O_2 and not simply O.

Step 2 calls for you to balance the atoms of each element on the two sides of the equation. As it stands, there is one magnesium atom on each side; magnesium is in balance. The left side of the equation has two atoms of oxygen, and the right side only one. What must you do to balance oxygen? Remember, there is only one way to do it, and watch out for the DON'Ts listed earlier. Balance the oxygen.

$$Mg + O_2 \rightarrow MgO$$

....

1b. $Mg + O_2 \rightarrow 2\ MgO$

Remember, your only way to balance atoms of an element in Step 2 is to use coefficients in front of substances in the unbalanced equation. Some common WRONG responses to the above—and what is wrong with them—are:

$Mg + O_2 \rightarrow MgO + O$	There are no oxygen atoms in the reaction. This violates the second DON'T.

$Mg + O_2 \rightarrow MgO_2$	MgO_2 happens to be a real substance, but it is *not* the product of this reaction. This violates the first DON'T.
$Mg + O_2 \rightarrow Mg\ 2\ O$	Mg 2 O is not a chemical formula. Coefficients are placed in front of a formula, not in the middle, and they affect the entire formula.

The last of the three wrong balancing methods points out that in balancing oxygen we have *un*balanced magnesium. There is now one magnesium atom on the left and two on the right. Correct this now.

....

1c. $2\ Mg + O_2 \rightarrow 2\ MgO$

There is another way you might have balanced $Mg + O_2 \rightarrow MgO$. You could have introduced the fractional coefficient 1/2 in front of oxygen: $Mg + 1/2\ O_2 \rightarrow MgO$. Fractional coefficients should not be used in this exercise. They may be used as a means to the final equation, however. If you do choose to balance the equation with a fractional coefficient, as above, you can then multiply the entire equation by 2 (doubling *each* coefficient), giving $2\ Mg + O_2 \rightarrow 2\ MgO$. Incidentally, equations should be written with the *smallest* whole-number coefficients. If, in your balancing procedure, you happened to arrive at $4\ Mg + 2\ O_2 \rightarrow 4\ MgO$, you could divide the entire equation—each coefficient—by 2 to get the desired result.

Example 2

Write the equation for the formation of iron(III) oxide from its elements.

Complete the first step by writing the formulas of the reactants on the left and the formula of the product on the right.

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2a. $Fe + O_2 \rightarrow Fe_2O_3$

Start with the iron; balance it first and leave oxygen unbalanced.

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2b. $2\ Fe + O_2 \rightarrow Fe_2O_3$

There are two thought processes by which balancing may be completed, both leading to the same result. Both will be discussed—after you have balanced the rest of the equation yourself.

....

2c. $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$

Oxygen atoms come two to the package in O_2 molecules, and three to the package in Fe_2O_3 units. Six atoms—2 times 3—is the smallest number of atoms by which a 3-and-2 combination can be equalized. If you take 3 packages of 2 each, you will have the same number as 2 packages of 3 each. This fixes the coefficients of O_2 and Fe_2O_3. The coefficient of iron is adjusted to correspond with the iron atoms in 2 Fe_2O_3.

A second way of reaching the final equation is to select the fractional coefficient of O_2 that will give the proper number of oxygen atoms to balance the three on the right side of $2\ Fe + O_2 \rightarrow Fe_2O_3$. With three oxygens on the right, we need three on the left, where they come two to a package in O_2. We therefore need $1\frac{1}{2}$ packages, or 3/2, yielding $2\ Fe + 3/2\ O_2 \rightarrow Fe_2O_3$. This balanced equation can be cleared of fractions by multiplying all coefficients by 2, giving $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$.

It is worthwhile to become familiar with both methods. The 3-and-2 combination appears frequently enough to justify the routine 2-of-3 and 3-of-2 thought processes. It is convenient to realize that if you need X atoms of oxygen from O_2 molecules, the number of molecules required is $X/2$. Doubling the equation yields whole-number coefficients.

B. Decomposition Reactions

The chemical change in which a single reactant decomposes into two or more products is a decomposition reaction. This is just the opposite of a combination reaction; indeed, many combination reactions can be reversed, as $2\ NaCl \rightarrow 2\ Na + Cl_2$. The reaction $2\ Al(OH)_3 \rightarrow Al_2O_3 + 3\ H_2O$ illustrates a decomposition of a compound into two simpler compounds. Another type of decomposition reaction occurs when hydrates (compounds containing water of hydration) are heated. $Na_2CO_3 \cdot 10\ H_2O \xrightarrow{\Delta} Na_2CO_3 + 10\ H_2O$ illustrates such a reaction, where Δ written over the arrow generally means ''applying heat.''

Example 3

Calcium carbonate is decomposed into calcium oxide and carbon dioxide by heat. Write the equation.

The first step is to write the formulas of reactants and products in their proper places. Proceed that far.

....

3a. $CaCO_3 \rightarrow CaO + CO_2$

Now Step 2: balance the atoms of each element on the two sides of the equation.

....

3b. $CaCO_3 \rightarrow CaO + CO_2$

Sometimes balancing an equation is easy—particularly when all coefficients are 1!

C. Complete Oxidation or Burning of Organic Compounds

Other than the oxides of carbon, carbonates, and a few other substances, the compounds of carbon are classified as **organic compounds**. Hydrogen is almost always present in an organic compound, and oxygen is a third very common element. When compounds containing carbon and hydrogen, or carbon, hydrogen, and oxygen, react *completely* with an excess of oxygen, the products are always carbon dioxide and water. Such a reaction may occur with the oxygen in the air, giving heat and light, in which case the process is called **burning**; and it may occur in living organisms, again giving off heat and other forms of energy, in which case it is referred to as **oxidation**. The description of such a reaction may be very brief: Compound X is completely oxidized, or Compound Y is burned in air. In both reactions you must recognize oxygen as an unnamed reactant to be included in the equation, and write the formulas of carbon dioxide and water as the products. $2\ C_6H_{14} + 19\ O_2 \rightarrow 12\ CO_2 + 14\ H_2O$ is an example of a burning reaction. Because organic compounds are frequently quite large, equations may have large coefficients; but don't let that bother you, because they are reached by the same method outlined above.

Example 4

Write the equation for the complete oxidation of methyl ethyl ketone, $CH_3COC_2H_5$.

Methyl ethyl ketone has been chosen for this example because its equation includes all the little things you must look out for in writing oxidation equations. First, notice that organic chemists sometimes write formulas in ways that seem strange to the beginning student. This is because the sequences of elements and certain combinations in the formula suggest how atoms are arranged in the molecule and identify the kind of compound it is. Second, you must be sure to count *all* the atoms of a given element in a molecular formula when balancing, such as 4 carbon atoms, 8 hydrogen atoms, and 1 oxygen atom in a molecule of $CH_3COC_2H_5$. A third point will show up later. Right now, complete Step 1 by writing the formulas of reactants and products in their proper places in an unbalanced equation.

....

4a. $CH_3COC_2H_5 + O_2 \rightarrow CO_2 + H_2O$

Always remember that, although it is unnamed in the statement of the reaction, oxygen is a second reactant and the products are carbon dioxide and water.

To begin Step 2, you balance both carbon and hydrogen. With the warning already given, add those coefficients to the equation.

....

4b. $CH_3COC_2H_5 + O_2 \rightarrow 4CO_2 + 4H_2O$

With carbon and hydrogen balanced, and oxygen in its elemental form on the left, oxygen can be balanced simply by placing in front of oxygen the coefficient that does the job. Sounds simple, but be careful . . .

....

4c. $2\ CH_3COC_2H_5 + 11\ O_2 \rightarrow 8\ CO_2 + 8\ H_2O$

Starting from $CH_3COC_2H_5 + O_2 \rightarrow 4\ CO_2 + 4\ H_2O$, you count 12 oxygen atoms on the right side of the equation. On the left, *one of the required 12 oxygen atoms comes from the reactant,* and the remaining 11 come from O_2. This is the third thing you must look out for in balancing oxidation equations, being sure not to overlook oxygen present in the compound being oxidized. With 11 oxygen atoms to come from O_2, you can balance the equation with a fractional coefficient: $CH_3COC_2H_5 + 11/2\ O_2 \rightarrow 4\ CO_2 + 4\ H_2O$. Doubling the entire equation gives whole-number coefficients, as required.

D. Oxidation–Reduction Reactions

The words **oxidize** and **oxidation** have meaning in chemistry other than ''reaction with oxygen,'' as suggested in the foregoing section. In its broader meaning, oxidation means loss of electrons. If one reactant loses electrons, another reactant must gain those electrons. The process of gaining electrons is called **reduction**. A reaction in which oxidation and reduction occur—and they must always occur simultaneously—is called an **oxidation–reduction reaction**, frequently shortened to ''redox'' reaction.

In this exercise we will be concerned with only one kind of redox reaction. The equation has the appearance of an element reacting with a compound in such a manner that the element replaces one of the elements in the compound. $Zn + Cu(NO_3)_2 \rightarrow Cu + Zn(NO_3)_2$ is such a reaction. It appears as if elemental zinc has replaced copper from $Cu(NO_3)_2$. This kind of equation is frequently called a **single replacement equation**. Given an element and an ionic compound as reactants, you should recognize the possibility of a redox reaction and be able to write the single replacement equation for that reaction. Whether or not the reaction actually occurs requires laboratory confirmation, of course.

Example 5

Gaseous hydrogen is released when zinc reacts with hydrochloric acid. Write the equation for the reaction.

The reactants and one of the products are identified. As you write the unbalanced equation (Step 1) for these three species, see if you can recognize the single replacement character of that equation and then figure out the formula of the second product.

....

5a. $Zn + HCl \rightarrow H_2 + ZnCl_2$

In the reaction zinc is releasing, or replacing, hydrogen in HCl. The second product is therefore zinc chloride. Balancing the equation is straightforward . . .

....

5b. $Zn + 2\ HCl \rightarrow H_2 + ZnCl_2$

Example 6

Write the equation for the reaction between aluminum and nickel nitrate.

This time you are given only the names of two reactants. Write their formulas on the left side of the arrow, leaving the product side blank.

....

6a. $Al + Ni(NO_3)_2 \rightarrow$

Here's where your skill in recognizing the possibility of a redox reaction comes into play. What possible products could come from these reactants? There is no indication that the nitrate ion decomposes. Ions of aluminum and nickel are both positively charged, so there is no way they could form a compound. The nitrate ion is negatively charged, so it could form a compound with an aluminum ion. It all points to a single replacement equation in which aluminum bumps nickel out of the compound. Complete Step 1 by writing the formulas of the products on the right side of the equation.

....

6b. $Al + Ni(NO_3)_2 \rightarrow Ni + Al(NO_3)_3$

This example gives us an opportunity to introduce an important technique in balancing equations. A quick glance shows that aluminum and nickel are balanced, but nitrogen and oxygen are not. You could balance them individually, but there is an easier way. It was noted above that the *nitrate ion* does not decompose in the reaction; in other words, the nitrate ion is the same on the product side of the equation as it is on the reactant side. Any time a polyatomic (many atom) ion is unchanged in a chemical reaction, that ion may be balanced *as a unit* in the equation. In other words, your thought process should be, ''There are two nitrate ions on the left, and three nitrate ions on the right. How do I balance them?'' How *do* you balance a 3-and-2 combination? You already know that, so go ahead. While you're at it, be sure to do whatever is necessary to keep the aluminum and nickel in balance.

....

6c. $2\ Al + 3\ Ni(NO_3)_2 \rightarrow 3\ Ni + 2\ Al(NO_3)_3$

You need 3 nickel nitrate units, where nitrate ions appear in packages of two, to balance 2 aluminum nitrate units, where the nitrate ions appear in packages of three, giving 6 nitrate ions on each side of the equation. The coefficients for the metals complete the equation.

E. Double Replacement Reactions

As the name suggests, double replacement reactions involve the combination of ions from different sources to form a new product. If two ions are to combine, one must have a positive charge and the other must have a negative charge. The combination of a lead ion from lead(II) nitrate and a chloride ion from sodium chloride to form lead(II) chloride is a good example: $Pb(NO_3)_2 + 2\ NaCl \rightarrow PbCl_2 + 2\ NaNO_3$. If you look at the equation, it appears as if the positive and negative ions in the two reactants have simply ''changed partners'' in the products; the positive ion of the first reactant has joined up with the negative ion of the second reactant, and the negative ion of the first has combined with the positive ion of the second. Whenever you see an equation with two ionized reactants, you can make an intelligent prediction that the products will be derived from an exchange of ions, and write their formulas accordingly.

Most double replacement reactions occur in water solution. One of the driving forces for these reactions is the formation of an insoluble ionic solid, called a **precipitate**. In the example above, lead(II) chloride is insoluble in water, so it precipitates as the ions combine with each other.

The other driving force that brings ions together is the formation of a molecular product, in which a covalent (shared electron pair) bond forms between the reacting ions. The most common molecular product is water, as in $HCl + KOH \rightarrow KCl + H_2O$. This kind of reaction, in which an acid reacts with a base, is called a **neutralization reaction**. The ionic product formed (KCl, in this example) is classified as a **salt**.

Name Date Section

Experiment 10

Work Page

Equation-Writing Exercise

Write the chemical equation for each reaction described below.

A. Combination Reactions

1. Diphosphorus trioxide is formed by direct combination of its elements.

2. Ammonia and sulfuric acid combine to form ammonium sulfate.

B. Decomposition Reactions

3. Ammonium nitrite decomposes into nitrogen and water.

4. When heated, potassium chlorate decomposes into oxygen and potassium chloride.

C. Complete Oxidation or Burning of Organic Compounds

5. Propane, C_3H_8, burns in air.

6. Acetaldehyde, CH_3CHO, is completely oxidized.

D. Oxidation–Reduction Reactions

7. Hydrogen is released when aluminum reacts with hydrochloric acid.

8. Magnesium reacts with silver nitrate solution.

E. Double Replacement Reactions

9. Barium carbonate precipitates from the reaction of barium chloride and sodium carbonate solutions.

10. Sulfuric acid reacts with calcium hydroxide.

11. Sodium iodate and silver nitrate solutions are combined.

12. Potassium fluoride reacts with hydrobromic acid.

13. Zinc hydroxide reacts with hydrochloric acid.

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Work Page

F. Other Reactions

14. Copper(II) chloride and water result from the reaction of copper(II) oxide and hydrochloric acid.

15. Carbon dioxide and water are two of the three products from the reaction of sulfuric acid with sodium hydrogen carbonate.

G. Mixed Reactions

16. Hydrobromic acid reacts with potassium hydroxide.

17. Aluminum reacts with phosphoric acid.

18. Silver nitrate reacts with hydrosulfuric acid.

19. Phosphorus triiodide is formed from its elements.

20. Iron(II) chloride reacts with sodium phosphate.

21. Sugar, $C_{12}H_{22}O_{11}$, is burned in air.

22. Sugar, $C_{12}H_{22}O_{11}$, breaks down to carbon and water when heated.

23. Lithium hydroxide solution is the product of the reaction of lithium oxide and water.

24. Magnesium sulfate reacts with sodium hydroxide.

25. Chlorine reacts with a solution of sodium iodide.

26. Nickel hydroxide reacts with sulfuric acid.

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Experiment 10

Work Page

27. Barium peroxide, BaO_2, decomposes into oxygen and barium oxide.

28. Ammonia is formed from its elements.

29. Butyl alcohol, C_4H_9OH, is oxidized completely.

30. Water is driven from copper sulfate pentahydrate, $CuSO_4 \cdot 5\ H_2O$, with heat.

31. Magnesium nitride is formed from its elements.

32. Sulfuric acid reacts with potassium nitrite.

Name | Date | Section

Experiment 10

Report Sheet

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Experiment 10

Report Sheet

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