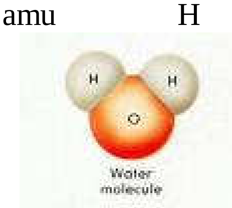
CHAPTER 6. CHEMICAL EQUATIONS AND THE MOLE CONCEPT

What Is a Mole?

Now that we are familiar with some important types of chemical compounds and their chemical formulas, we will find that we already know quite a bit about the masses of these substances. Consider, for example, water, that ubiquitous and important chemical compound formed from the elements hydrogen and oxygen. We already know from our study of atoms and the periodic table that oxygen is sixteen times heavier than hydrogen. Hydrogen, the lightest atom in the periodic table with only one proton, has an atomic mass of one. Oxygen, with eight protons and eight neutrons in the most common isotope, has an atomic mass of about 16. The covalent compound formed by the combination of these two elements has the formula H_2O, with two hydrogen atoms and one oxygen atom. The overall weight of the H_2O molecule is the total of the weights of the two hydrogen atoms and the oxygen atom, or 18 atomic mass units.

Fig. 6-1: *The water molecule has a mass of 18 amu.*

H 1.0 amu H 1.0 amu



O 16.0 amu

Atomic mass units are convenient units to use when considering the masses of atoms. But, although chemical reactions involve atoms and molecules, it would hardly be practical to measure atoms and molecules when making new compounds in the laboratory. As we have seen, it would take 6×10^{23} hydrogen atoms to make up a single gram! Chemists have chosen a very simple way to deal with the masses of the elements and their compounds by expressing the atomic masses in grams. The gram atomic mass of hydrogen, for instance, is about one gram, and the gram atomic mass of oxygen is about 16 grams. Any other unit of mass would work as well, of course, as long as oxygen is sixteen times as heavy as hydrogen... we could have the pound atomic mass, for instance. As scientists have agreed on the metric system, and our laboratory balances are calibrated in gram units, the **gram atomic mass** is a convenient unit.

No matter how complex a compound may be, it is easy to find the **gram molecular mass** through addition of the gram atomic weights of the elements it contains. But it is best to do so in an organized way, or even the simplest arithmetic can have a way of giving incorrect results. (Have you ever had trouble getting a checkbook to balance?) Here is a suggested way of adding up the gram molecular mass for H_2O:

Problem example 6-1. Finding the gram molecular mass of H_2O

2 H =	2	x	1.0 g	=	2.0 g
1 O =	1	x	16.0 g	=	<u>16.0 g</u>
					18.0 g

This organized approach will be especially helpful for a more complex substance like $(NH_4)_2SO_4$. The gram atomic masses for each element can always be found in that extraordinarily useful source of chemical information, the periodic table.

Problem Example 6-2. $(NH_4)_2SO_4$

2 N	=	2	x	14.0 g	=	28.0 g
8 H	=	8	x	1.0 g	=	8.0 g
1 S	=	1	x	32.0 g	=	32.0 g
4 O	=	4	x	16.0 g	=	<u>64.0 g</u>
						132.0g

This simple and useful concept of gram molecular mass is often abbreviated to the term "**mole**," which conjures up images of furry animals in the minds of the uninitiated. Most often, for instance, chemists will say that water has 18 grams per mole. There is a technical problem with the use of this term, however. Water is a covalently bonded molecule; but ammonium sulfate, the compound in Problem Example 6-2, is an ionic compound, so the use of a term based on the word "molecular" is somewhat misleading. For that reason, the term "gram formula mass" is often used instead; it has, however, no convenient abbreviated form, and chemists will often be found referring to "grams per mole ammonium sulfate" even though purists may frown.

Chemists make frequent use of the concept of gram molecular mass. But one need not be a practicing chemist to make good use of the information in molecular masses. In Chapter 1 sodium compounds were examined as components of packaged foods like macaroni and cheese. Because sodium is found in so many substances and its presence may be an important health issue, the total amount of sodium from all sources is listed on the box label. It can be instructive, however, to visualize the quantity of table salt, or sodium chloride, that would correspond to this sodium content. Nutritional information on a box of macaroni and cheese, for example, states that a 3/4 cup serving prepared from the box contents contains 530 mg of sodium. How much salt would you have to add from a salt shaker to equal this amount? To make this comparison, we need to find the number of grams in a mole of sodium chloride, or table salt, and compare that with the number of grams in a mole of sodium.

Problem Example 6-3: Find the number of grams in a mole of sodium chloride, NaCl.

1 Na	=	1 x 23.0 g	=	23.0 g
1 Cl	=	1 x 35.5 g	=	<u>35.5 g</u>
				58.5 g per mole NaCl

We see, then, that each 23.0 gram quantity of sodium corresponds to a weight of 58.5 grams of sodium chloride. The weight ratio of sodium chloride to sodium is 58.5 to 23.0, simplifying to 2.54. The weight of sodium listed on a food label, then, corresponds to over 2 and 1/2 that quantity by weight of table salt. The serving of macaroni and cheese that contains 530 mg of Na corresponds to adding table salt weighing 530 mg x 2.54, or 1350 mg. This quantity of table salt, 1.35 grams, is much more than most people would feel like adding from a salt shaker. If you are not yet familiar enough with metric units to be able to visualize this amount, it might be worthwhile to weigh out 1.35 grams of salt on a laboratory balance.

Sodium compounds are not the only everyday substances for which consumers can benefit from knowing about the concept of gram molecular mass. The label on a detergent box, for example, may carry the information "contains 15% phosphate as phosphorus." If phosphate-containing sewage finds its way into a lake or pond, the phosphate, acting as a plant nutrient, can cause an overgrowth of algae and result in the overall deterioration of water quality known as eutrophication. If consumers want to minimize the use of phosphate, how are they to interpret this detergent box label? Phosphate is an anion with the formula PO_4^{3-}; it is one of the common anions listed in Chapter 5. The phosphate anion, of course, is only part of the formula of an ionic compound; there must be a cation as well. For instance, is the phosphate in the detergent box may be in the form of sodium phosphate, Na_3PO_4. In order to find out how this affects the overall weight of the phosphate compound in the box, we need to calculate the weight of a mole, or a gram formula mass, of sodium phosphate.

Problem Example 6-4: Find the gram formula mass of sodium phosphate, Na_3PO_4.

3 Na	=	3	x	23.0 g	=	69.0 g
1 P	=	1	x	31.0 g	=	31.0 g
4 O	=	4	x	16.0 g	=	64.0 g
						164.0 g

The active phosphate-containing ingredient in the detergent, sodium phosphate, weighs 164 g, or more than five times as much as phosphorus, with an atomic weight of only 31.0 g. In other words, the detergent formulation which contains "15% phosphate as phosphorus" is really over 75% sodium phosphate by weight! Clearly, a simple molecular mass calculation can be a powerful source of chemical information.

Problem Solving with the Mole Concept

Familiarity with the mole concept is a useful chemical skill that can be applied in a variety of problem-solving situations. First, consider one of the simpler situations that might arise. If someone asked you how many eggs in half a dozen, you would probably have a ready answer. But what if you were asked to find the number of grams in half a mole of carbon? First, looking at the periodic table, find the number of grams in a mole of carbon, 12.0 g. Now can you tell how many grams in half a mole of carbon? The reasoning process is really the same.

Problem Example 6-5. Find the number of grams in 0.5 moles of carbon.

Let us look at this simple problem in terms of our general problem-solving strategy. As always, first define the problem objective, or **what we want** to find; it is the number of **grams**, and will appear at the right of the problem as our final solution. Now determine **what is given** as information; it is the number of **moles of carbon**, one-half or 0.5, and will appear at the left of our equation.

WE ARE GIVEN	*WE WANT*
0.5 moles C	grams

The **conversion factor**, in order to produce a final result of grams C, must have the units of grams C on the top part of the factor. The bottom part of the factor will be in units of moles C and will then cancel out the units of the given quantity. The units of the conversion factor, then will be in grams/mole C, or grams per one mole of carbon. We know the number of grams in a mole of carbon; it is the gram atomic mass, which we find in the periodic table or other reference source as 12 g/mole. The final conversion factor, then, which appears in our equation is 12 g/mole C. This can also be written as 12 g/ 1 mole C.

WE ARE GIVEN		*CONVERSION FACTOR*		*WE WANT*
0.5 moles C	x	$\frac{12\text{ g}}{1\text{ mole C}}$	=	g

Solving the problem, we find the answer, 6.0 g C. Recalling out original question of how many eggs are in 1/2 dozen eggs, do you now see the similarity to this question? Does the answer to the question of how many grams in 1/2 mole of carbon begin to seem intuitively obvious?

Problem Example 6-6. Find the number of grams in 0.100 moles of NaCl.

This is really a two-part problem. In order to solve the problem, we need to find the number of grams in one mole of NaCl.

1 Na	=	1	x	23.0 g	=	23.0 g
1 Cl	=	1	x	35.5 g	=	<u>35.5 g</u>
						58.5 g per mole NaCl

Then set up the problem by writing what we are given, 0.100 moles NaCl, on the left, and what we want, grams, on the right.

WE ARE GIVEN	*WE WANT*
0.100 moles NaCl	grams

The conversion factor, as in the previous example, must have units of grams in the top of the factor and moles in the bottom. Do we know the g/mole NaCl, or grams per mole? It is the gram molecular mass which we have calculated above.

WE ARE GIVEN	*CONVERSION FACTOR*	*WE WANT*
0.100 mole NaCl	$\frac{58.5\ \text{g}}{1\ \text{mole NaCl}}$	g

Solving the problem, we find the answer, 5.85 grams.

Problem Example 6-7. Find the number of moles in 117 g of NaCl.

This time **what is given** is the number of grams of NaCl. The problem objective, or **what we want**, is to find the number of moles. As you set up the problem you will notice that the conversion factor must have moles NaCl on top and grams on the bottom. We can still use the relationship between moles NaCl and grams we calculated in the example above. There are 58.5 g in 1 mole of NaCl.

WE ARE GIVEN		*CONVERSION FACTOR*		*WE WANT*
117 g NaCl	x	$\frac{1\ \text{mole}}{58.5\ \text{g NaCl}}$	=	moles NaCl

Solving the problem, we find the answer, 2.00 moles NaCl..

Chemical Equations: Our Shorthand for Chemical Reactions

One of the most rewarding aspects of chemistry both intellectually and financially is the ability to make new substances through chemical reactions and to understand the reactions that occur around us constantly as a part of everyday life. As an example, let us again look at a reaction which produces the important everyday chemical substance water, or H_2O. The reaction in which water is formed from hydrogen and oxygen is easy to perform. As a matter of fact, it is a popular classroom demonstration because of the spectacular energy effects that accompany it. Hydrogen gas in a balloon is touched with a candle; the spark bursts the balloon and triggers an explosive reaction with the oxygen which is present in the air, or, more spectacularly, in the balloon.

Fig. 6-2: *Dr. Shakashiri demonstrates that when the candle is touched to the balloon full of hydrogen and oxygen, a loud explosion will result.*

Dirigibles like the Goodyear blimp we sometimes see today were once considered as a form of passenger transportation; these early dirigible models used hydrogen gas inside the large inflated body because hydrogen is the lightest gas. A famous tragedy occurred at Lakehurst, New Jersey, in 1937 when the hydrogen gas in the dirigible Hindenberg exploded, killing all aboard and ending forever the use of hydrogen as a dirigible gas. http://www.archive.org/details/hindenberg_explodes Now the light but totally unreactive gas helium is used instead.

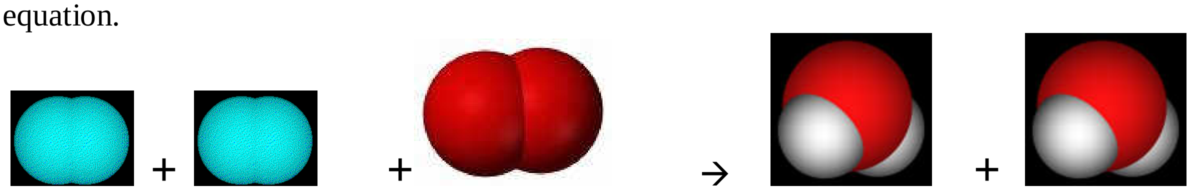
Fig. 6-3: *Helium has a very low gram atomic mass (2.0 g/mol) and is an unreactive noble gas. Because of these properties, it is used today in blimps.*

The same reaction can be performed in a closed, explosion-proof container, which makes it easier to examine the situation before and after the reaction. After the reaction occurs, the hydrogen gas has disappeared, along with the oxygen gas that reacted with it. The only substance left in the closed container is a small amount of water.

An important problem-solving tool for understanding chemical reactions like this one is the **chemical equation**, in which a great deal of useful information is contained in a simple form. In our chemical reaction of hydrogen and oxygen to form water, we say that the hydrogen and oxygen are the **reactants**, and water is the reaction **product**. The reactants appear to the left of our chemical equation, and the products to the right. An arrow indicates the change from reactants to products. So we can easily write a chemical equation that shows the reactants and products in the formation of water:

$$H_2 \quad + \quad O_2 \quad \text{-->} \quad H_2O \qquad\qquad \text{unbalanced} \quad (6\text{-}1)$$

This chemical equation indicates in a very efficient way what substances are present both before the reaction occurs and afterwards. There is even more chemical information contained in this brief line. Hydrogen, as we have seen, occurs in nature as two hydrogen atoms covalently bonded as the H_2 molecule. Oxygen also occurs naturally as a diatomic molecule O_2. The water molecule is composed of two hydrogen atoms and one oxygen atom. All this information is contained in the chemical equation.



: *When hydrogen molecules and oxygen molecules react to form water molecules, the atoms form different bonds to make new molecules. The total number of atoms remains the same because the same atoms are present before and after the reaction.*

But this equation as we have written it is an **unbalanced equation**; there are two oxygen atoms on the left, but only one on the right. This is an impossibility, and let us examine exactly why this is so. When the hydrogen and oxygen disappear, and water appears along with a big bang, it seems like a magic trick to the observer. But our knowledge of the existence of atoms has given us an understanding of what actually occurs. The hydrogen atoms have attached themselves to oxygen atoms instead of being bonded to other hydrogen atoms. The oxygen atoms have become attached to the hydrogen atoms instead of being bonded to other oxygen atoms. The same number of atoms is present before and after the reaction. They have simply rearranged themselves into new molecules. That is how new substances are formed. Molecular models are helpful in visualizing the way the same atoms have formed different molecules before and after the reaction (Figure 6-4). Our chemical equation, which describes exactly how the "magic" takes place when new substances are formed out of old ones, must clearly show this very important principle; it must be a "balanced" equation with the same number of each atom after the reaction in the products (to the right of the arrow) as there were before the reaction in the reactants (to the left of the arrow).

A balanced chemical equation, then, has its beginnings in the insights of the ancient Greek and Roman philosophers. In the balanced equation the ancient concepts of the atoms and the elements are used today to explain the seeming magic in the transformations of substances. There is a very practical reason for the balanced equation, as well. In performing chemical reactions it is important to know not only what chemicals are involved, but exactly how much. Chemical plants which produce pharmaceuticals, plastics, and the other substances of our daily lives must use exactly the right amounts of reacting materials in order to produce chemical products efficiently. Chemistry students performing reactions in the laboratory must know how much of the reactants to weigh out. The chemical equation, combined with our knowledge about the weights of the elements, provides this information.

Fig. 6-5: *Chemists use balanced chemical equations together with gram molecular masses to find out how much of each material to use in performing a chemical reaction.*

Looking once more at our unbalanced equation of the reaction of hydrogen and oxygen to form water, how can it be balanced to show the correct number of atoms for each element in the reaction? The subscripts in the chemical formulas cannot be changed, because they give important information about the molecule. If, for instance, one were to write the product as H_2O_2, the equation would be balanced, but the product would contain two hydrogen atoms and two oxygen atoms; the name of this substance is hydrogen peroxide, and it is quite different from water. Instead we can change the **coefficients**, the numbers in front of the chemical formulas, and that way change the number of molecules. We can balance the number of oxygens atoms on each side of the equation by changing the coefficient of the water molecule to 2; that way there will be two oxygen atoms both on the left and the right sides of the equation. If we do so, however, we will find that we have now put the hydrogen atoms out of balance.

$$H_2 \quad + \quad O_2 \quad \text{-->} \quad 22\ H_2O \qquad\qquad \text{unbalanced} \qquad (6\text{-}2)$$

To balance the hydrogen atoms, a coefficient of 2 can be placed in front of the H_2 molecule; the total number of hydrogen atoms in 2 H_2 is four. The balanced equation for the formation of water is

$2\ H_2$ + O_2 --> $2\ H_2O$ balanced (6-3)

As a final check, notice that the balanced equation has four hydrogen atoms on both the reactant and product sides of the equation, and two oxygen atoms on each side. Occasionally a student has difficulty in looking at the shorthand formula 2 H_2O and understanding that there are four hydrogen atoms. Looking at the molecular models in Figure it should be clear that there are two water molecules, each with two hydrogen atoms.

Balancing a chemical equation can take several steps before a final check shows that all elements in the equation have been balanced. Often trial and error is part of the problem-solving technique. There are also some strategies that are helpful in balancing chemical equations and which are best illustrated by examples.

Problem Example 6-8. Balance the following chemical equation, which describes a way to produce small amounts of oxygen in the laboratory by heating potassium perchlorate.

$KClO_3$ --> KCl + O_2 unbalanced

In this equation the oxygen atoms are unbalanced; how can the reactant side, with three oxygen atoms, be balanced with the product side, which has two oxygen atoms? The simplest number, or lowest common multiple, which contains both numbers, is six. The oxygen atoms will be balanced with six on each side by writing coefficients as follows:

$2\ KClO_3$ --> KCl + $3\ O_2$ unbalanced

Now the K and Cl atoms have become unbalanced, but a coefficient of 2 for KCl remedies this unbalance. The final balanced equation is:

$2\ KClO_3$	-->	$2\ KCl$	+	$3\ O_2$	balanced
2 K		2 K			
2 Cl		2 Cl			
6 O				6 O	

Notice that a good way of making a final check of balancing equations is to add up each element on a separate line.

Problem Example 6-9. Balance the following equation.

NH_4Cl + CaO --> $CaCl_2$ + H_2O + NH_3 unbalanced

This equation may appear complex at first. The approach to such a problem involves an important problem-solving skill: **attempt the easier parts first**. In this case, try balancing the N, Cl, and Ca atoms first as follows:

2 NH_4Cl + CaO --> $CaCl_2$ + H_2O + 2 NH_3

Now check the H and O atom balance at this point by adding carefully the atoms in each molecule, doing O and H last.

2 NH_4Cl	+	CaO	-->	$CaCl_2$	+	H_2O	+	2 NH_3	balanced
2 N								2 N	
2 Cl				2 Cl					
1 Ca				1 Ca					
8 H						2 H		6 H	
		1 O				1 O			

The H and O are already balanced! Equation balancing will not always be so easy, but it is a good general problem-solving rule to attack the simpler parts of the problem first. In equation-solving, this usually means to **balance H and O last**.

Chemical Equations and the Mole

In the chemical equation a great deal of information about the chemical reaction is expressed in a very efficient way. From the equation we know not only what molecules are consumed and formed, we know how many. Combining this information with the mole concept, we can learn even more about our reaction, finding how many grams of substance will react and be formed. As an example, let us look once more at the equation which describes the reaction of the elements hydrogen and oxygen to form water.
First, let us find gram molecular weights for the molecules involved in the equation.

For H_2,2 H = 2 x 1.0 = 2.0 g/mole

For O_2,2 O = 2 x 16.0 = 32.0 g/mole

For H_2O,	2 H = 2 x 1.0 =	2.0 g
	1 O = 1 x 16.0 =	<u>16.0 g</u>
		18.0g/mole

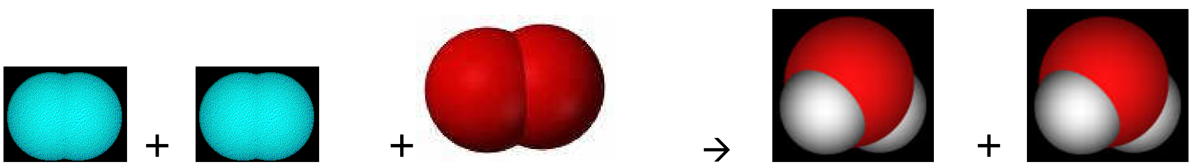
Now let us look at the chemical equation in terms of moles as well as molecules, and combine the information from the chemical equation with the weight information from the gram molecular weights.

2 H_2	+	O_2	-->	2 H_2O
2 molecules H_2	+	1 molecule O_2	-->	2 molecules H_2O
2 moles H_2	+	1 mole O_2	-->	2 moles H_2O
2 x 2.0 g H_2	+	1 x 32.0 g O_2	-->	2 x 18.0 g H_2O
4.0 g H_2	+	32.0 g O_2	-->	36.0 g H_2O

Now we have the chemical equation in the form of a chemical recipe. Four grams of hydrogen will react with 32 grams of oxygen with none left over. As with any recipe, the quantities can be increased or decreased as long as the weight relationships remain the same; two grams of hydrogen will react with 16 grams of oxygen, or one gram of hydrogen will react with 8 grams of oxygen. If one gram of hydrogen were reacted with one gram of oxygen, all the oxygen would be reacted, but most of the hydrogen would remain unreacted, even though 2 molecules of hydrogen react with each molecule of oxygen. That is because the oxygen molecule is much heavier than the hydrogen molecule, and so a greater weight of oxygen is required. All this information is provided by the chemical equation and the gram molecular weights.

Our "chemical recipe" features another relationship that you may have noticed. The weights of the substances on the reactant side of the equation, 4.0 g H_2 + 32.0 g O_2, add up to exactly the same as the weight on the product side of the equation, 36.0 g H_2. This is not a coincidence. The weights of the substances present before and after any reaction will always be the same. This principle, called the **Law of Conservation of Mass**, may seem obvious. After all, the principle embodied in the chemical equation is the idea that substances do not appear magically from nowhere, but simply occur from the recombination of the same atoms in a different way. Since the same atoms appear in the same number on both sides of the equation, the weights would logically be expected to be the same. Some of the most familiar chemical recipes are those that are used to produce chemical changes in the kitchen. The old-fashioned recipe for pound cake, for instance, calls for a pound of butter, a pound of sugar, a pound of flour, and a pound of eggs, all mixed well and baked in the oven. The product formed from a series of chemical reactions is a delicious, if rather heavy, cake. Most people would have no hesitation in predicting the final weight of the cake as four pounds, though they might be surprised to hear they were using the Law of Conservation of Mass.

1 pound + 1 pound + 1 pound + 1 pound → 4 pounds



2 grams + 2 grams + 32 grams → 18 grams + 18 grams

The law of conservation of mass states that the mass of the products in a chemical reaction is the same as the mass of the reactants. >

In chemical equations we have a powerful problem-solving tool. We know how much material to react together to make a desired new substance, and we know just how much product to expect. Moreover, we can use chemical equations to understand the consequences of the chemical reactions that take place all around us and in our own bodies every day. How much carbon dioxide will be produced in the atmosphere by the amount of fuel we burn? How much sulfuric acid will be added to the atmosphere from the sulfur in the coal we burn? These questions can be answered by using chemical equations.

Problem-Solving with Chemical Equations

As with all problem-solving, it is important to have an organized approach to solving problems using chemical equations. Notice in the sample problems below how the chemical equation comes first and organizes the problem information in a useful way.

Fig. 6-7: *When natural gas, or methane, is burned, the products are carbon dioxide and water.*

Problem Example 6-10. When natural gas (methane), with the chemical formula CH_4, is burned (reacted with oxygen) the products are carbon dioxide and water. From the chemical equation for this process determine the amount of carbon dioxide produced when 10.0 grams of methane is burned.

The balanced chemical equation for this process is

CH_4 + $2\ O_2$ --> CO_2 + $2\ H_2O$

To determine weights for the chemical substances, it will be necessary first to determine their gram molecular weights.

For CH_4,	1 x C = 1 x 12.0 g	=	12.0 g
	4 x H = 4 x 1.0 g	=	<u>4.0 g</u>
			16.0 g
For CO_2,	1 x C = 1 x 12.0 g	=	12.0 g
	2 x O = 2 x 16.0 g	=	<u>32.0 g</u>
			44.0 g

Then the molecular weight information is combined with the information contained in the chemical equation.

CH_4 +	$2\ O_2$ -->	CO_2 +		$2\ H_2O$
1 mole		1 mole		
1 x 16.0 g CH_4			1 x 44.0 g CO_2	

Finally the solution to the problem can be found. First, identify the information given in the problem; **we are given 10.0 g CH_4**. Then identify the objective of the problem; **we want** to know how many **grams of CO_2** will be formed. Though there is more than one way to solve this problem, a simple way is to observe the weight ratio of these two substances as predicted by the chemical equation. **All numbers must be labelled carefully with both units and substance.** Observe that the units in the top of the conversion factor must be the same as in the answer we want. Otherwise units will not cancel properly.

WE ARE GIVEN		*CONVERSION FACTOR*		*WE WANT*
10.0 g CH_4	x	$\frac{44\ g\ CO_2}{16.0\ g\ CH_4}$	=	g CO_2

Solving the problem, we find the answer, 27.5 g CO_2.

We have just obtained some extraordinarily useful information by the use of this chemical equation. The reaction we have just been dealing with describes the burning of methane, or natural gas. Carbon dioxide is one of the "greenhouse gases" which absorb heat and hence contribute to the warming of the earth, a source of increasing concern both to scientists and non-scientists (see Chapter 16). It is a sobering thought that for every 10 grams of methane that is burned, 27.5 grams of carbon dioxide is produced, contributing to the greenhouse effect.

Problem Example 6-11: How many moles of oxygen (O_2)are required to react with 5 moles of methane (CH_4)?

First, notice that the question refers only to moles, not grams. It will not be necessary to find weights in the calculation. The numbers of moles that react are exactly the same as the number of molecules that react in the equation.

CH_4 + $2\ O_2$ --> CO_2 + $2\ H_2O$

1 mole 2 moles

Analyzing the problem, we see that what **we are given** is the number of moles of CH_4 and what **we want** is the number of moles of O_2.

WE ARE GIVEN	*WE WANT*
5 moles CH_4	moles O_2

We find the relationship to form the conversion factor from the chemical equation.

$$5 \text{ moles } CH_4 \quad \times \quad \frac{2 \text{ moles } O_2}{1 \text{ mole } CH^4} \quad = \quad \text{moles } O_2$$

Solving the problem, we find the answer, 10 moles O_2.

CONCEPTS TO UNDERSTAND FROM CHAPTER 6

Gram molecular mass is found by addition of the gram atomic masses of the elements in a compound, using the gram atomic mass once for each time the atom appears in the molecule.

The chemical equation contains much valuable information about a chemical reaction.

Chemical equations must be balanced, because the same atoms are present before the reaction and after the reaction.

The total mass present before the chemical reaction takes place is the same as the total mass after the reaction takes place. This is called the Law of Conservation of Mass.

SKILLS TO BE ACQUIRED FROM CHAPTER 6

After completing Chapter 6, you should be able to:

Find the gram atomic mass for any compound for which you are given the formula.

Find the number of grams of a substance if you are given the number of moles.

Find the number of moles of a substance if you are given the number of grams.

Balance a chemical equation.

Use the chemical equation to find how many moles of a substance are needed for a reaction or are produced in a reaction.

Use a chemical equation to find how many grams are required for a reaction or are produced in a reaction.

Name ___________________________ **Date __________________**

PROBLEMS TO SOLVE USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 6

6-1. Find gram molecular weights or gram formula weights for the following (to the nearest tenth of a gram). Show your work!

a. Baking soda, $NaHCO_3$

b. Calcium acetate (a food additive), $Ca(C_2H_3O_2)_2$

6-2. How many moles of helium in 28.0 grams of helium? Show your work always!

6-3. How many moles of calcium in 2.00 grams of calcium?

6-4. How many grams of helium in 28.0 moles of helium?

6-5. How many grams of calcium in 2.00 moles of calcium?

6-6. How many grams of calcium chloride in 2.00 moles of calcium chloride, $CaCl_2$ (used as a deicer for roads and sidewalks)?

6-7. Ammonium nitrate, NH_4NO_3, is a common component of fertilizer. How many moles of ammonium nitrate in 100 grams of ammonium nitrate?

6-8. How many grams of ammonium nitrate in 100 moles of ammonium nitrate?

6-9. Balance the following chemical equations, all of which involve the formation of pollution products in the air.

a. Sulfur impurities in coal react with oxygen from the air as the coal is burned to form sulfur dioxide gas.

__S + __O_2 --> __SO_2

b. Sulfur dioxide gas reacts with oxygen from the air to form sulfur trioxide gas.

__SO_2 + __O_2 --> __SO_3

c. Sulfur trioxide gas reacts with water vapor in the air to form sulfuric acid.

__SO_3 + __H_2O --> __H_2SO_4

d. Coal is burned in the presence of limited amounts of oxygen to form carbon monoxide.

__C + __O_2 --> __CO

e. Nitrogen gas and oxygen gas from the air react together in an automobile engine to form nitrogen oxide.

__N_2 + __O_2 --> __NO

f. Nitrogen oxide reacts with water to form nitrogen dioxide, an important component of air pollution in may cities.

__NO + __O_2 --> __NO_2

g. Nitrogen dioxide reacts with water to form nitric acid and nitrogen oxide; this reaction is a contributor to acid rain.

__NO_2 + __H_2O --> HNO_3 + __NO

6-10. Balance the following chemical equations.

a. Green plants with the help of the sun's energy make sugar with this reaction, called photosynthesis:

____ CO_2 + ____H_2O ---> ____$C_6H_{12}O_6$ + ____O_2

b. Sugar is changed to carbon dioxide gas and alcohol in the reaction, which is usedin producing wine and requires the presence of yeast:

$C_6H_{12}O_6$ --> CO_2 + C_2H_5OH

c. Chlorine gas used to disinfect a swimming pool escapes from a defective tank and reacts with iron beams in this reaction:

____Fe + ____Cl_2 ---> ____$FeCl_3$

d. Copper roofing trim reacts with acid rain in this equation:

____Cu + ____H_2SO_4 ---> ____SO_2 + ____$CuSO_4$

+____H_2O

6-11. Consider the reaction for the burning (reaction with oxygen, O_2) of methane (CH_4), the major component of natural gas:

CH_4 + $2O_2$ ---> CO_2 + $2H_2O$

a. How many moles of oxygen are required to react with 0.200 moles of CH_4?

b. How many grams of carbon dioxide are produced by the reaction of 8.00 grams of CH_4 with an adequate amount of oxygen?

6-12. In the reaction below, used to produce ammonia, NH_3, how many moles of ammonia are produced from 1.00 mole of hydrogen, H_2?

N_2 + 3 H_2 --> 2 NH_3

6-13. The reaction below describes the reaction of iron metal with oxygen from the air to form rust (iron oxide).

$$3\ Fe \quad + \quad 2O_2 \quad \text{-->} \quad Fe_3O_4$$

a. How many moles of iron react to form 1.00 mole of iron oxide?

b. How many grams of iron react with 16.0 grams of oxygen?

6-14. In the following reaction octane (C_8H_{18}), a major component of gasoline, is burned as fuel in an auto engine.

$$2\ C_8H_{18} \quad + \quad 25\ O_2 \quad \text{-->} \quad 16\ CO_2 + \quad 18\ H_2O$$

a. How many moles of carbon dioxide are produced from the reaction of 1.00 mole of octane?

b. How many grams of carbon dioxide are produced from the reaction of 1.00 gram of octane?

............MORE PROBLEMS TO SOLVE USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 6

6-15. How many moles of ammonium sulfate in 100 grams of ammonium sulfate, $(NH_4)_2SO_4$?

6-17. Balance the following equations:

a. $NO + O_3 \text{ --> } NO_2 + O_2$

b. $Al_2O_3 \text{ --> } Al + O_2$

c. $Ca(OH)_2 \text{ --> } CaO + H_2O$

6-18. When hydrochloric acid reacts with baking soda, the products are water, sodium chloride, and bubbles of carbon dioxide gas.

a. Balance the following chemical equation which describes these changes.

$HCl + NaHCO_3 \text{ --> } NaCl + H_2O + CO_2$

b. Balance the following equation which describes what happens when sulfuric acid is added instead.

$H_2SO_4 + NaHCO_3 \text{ --> } Na_2SO_4 + H_2O + CO_2$

6-19. In the following reaction (octane is burned in an auto engine):

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

a. How many moles of carbon dioxide are produced from the reaction of 60.0 moles of octane?

b. How many grams of carbon dioxide are produced from the reaction of 500 grams of octane?

c. You might not have realized that an auto engine produces water, but from the chemical equation you can see that it does. In cold weather this water is visible as it condenses on the cold tailpipe before the exhaust system warms up. How many grams of water are produced when 100 grams of octane are burned?