Chapter 6 CHEMICAL REACTIONS

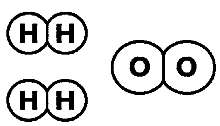
We just spent several chapters talking about chemical units, formula units, masses, etc. for one element or compound. Now we start talking about chemical reactions.

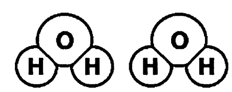
6.1 Reactions and equations

In its simplest, a chemical reaction involves a change in chemical identity. Elemental forms or compounds undergo chemical reactions and produce other elemental forms or compounds of different identities. When gasoline burns in the engine of your car, the compounds originally present in the gasoline undergo a reaction (combustion) and different compounds (mostly carbon dioxide and water) are formed in the process. The molecules originally present in your gasoline are no longer there (unless you need new rings or a new ignition system, or your emissions control system is beyond kaput, but I'll assume you're up to snuff).

Much of chemical reactions is like cooking. You know what goes in and you know what comes out (or at least what's supposed to come out). But you have to get the amounts right. The connection to cooking is actually greater than you may realize: much of cooking is chemistry. In fact, some of those chemical reactions are very complicated. The "browning" process is a very complicated chemical process involving many chemical reactions, although nobody really understands them all. Nevertheless, that's what contributes to the flavor of those cookies, pie crust, etc.

Quantitative aspects are extremely important. Chemical reactions are subject to conservation of mass. For typical chemical reactions, this is reflected in the conservation of atoms, which is itself a reflection of conservation of matter. Chemical units change. Chemical bonds can be broken and new ones can be made. Atoms can relocate and form new arrangements with new bonds to different atoms, but the total number of atoms stays the same. Let's just look at a very simple illustration. When hydrogen gas burns, it reacts with oxygen in the air to form water. The initial chemical units are diatomic H_2 and O_2 molecules (shown at left). The reaction produces molecules of water, H_2O (shown at right). The chemical units before and after are different but the number of atoms stays the same. Atoms are conserved. Matter is conserved. Mass is conserved.





Chemists use a line notation for depicting chemical reactions. The notation is called a chemical equation and it lists the compounds before and after the reaction. The notation is represented as follows, wherein A, B, C and D are generic chemical formulas of different compounds.

$$A + B \rightarrow C + D$$

This equation would mean that compound A reacts with compound B to form compound C and compound D. Compound A and compound B are called "reactants" and these are shown to the left of the arrow. Compound C and compound D are called "products" and these are shown to the right of the arrow. The arrow conveys the sense of change as the reactants react to produce products. Reactants can be given in any sequence on the left. Products can be given in any sequence on the right. We could also write the following; it says the same thing.

$$B + A \rightarrow D + C$$

Just remember that reactants must be on the left and products must be on the right.

Another word frequently used is "reagent". Sometimes it's just used for a reactant, but I'll use "reagent" as a generic term for a reactant or a product.

All equations must be "balanced". Everything must be accounted for. ALWAYS. There are two aspects of balancing an equation: atom count must be balanced and charges must be balanced. Right now, we'll stick with neutral compounds of zero charge; this means we won't need to cover charge balance until later. We DO need to do atom balance. Remember what I said above about reactions.

> CHEMICAL UNITS CHANGE. CHEMICAL BONDS CAN BE BROKEN AND NEW ONES CAN BE MADE. ATOMS CAN RELOCATE AND FORM NEW ARRANGEMENTS WITH NEW BONDS TO DIFFERENT ATOMS, BUT THE TOTAL NUMBER OF ATOMS STAYS THE SAME.

The underlined part is called atom-balance; again, since atoms are matter and matter is mass, this is also called mass-balance. The chemical equation must properly reflect this balance. Although the chemical units and identities are changing, the count of atoms of each kind is the same on the left and right sides of the equation. Atom-balance is not concerned with the identities of the chemical units. It is only concerned with total atom count.

Let's do a simple example.

Chlorine, Cl_2, (one of the diatomic halogens) reacts with zinc metal, Zn, to form zinc chloride, $ZnCl_2$, which is a colorless ionic compound. We can write this as follows.

$$Cl_2 + Zn \rightarrow ZnCl_2$$

If we count the atoms of each kind, we see that the equation is atom-balanced. The equation shows a total of two Cl's on the left and a total of two Cl's on the right. The equation shows one Zn on the left and one Zn on the right. The counts are the same left and right. The identities of the chemical units are different, but the atom counts are the same. That's atom-balance. That's mass-balance.

It is important to keep in mind the quantitative relationships which are expressed by the chemical equation. This simple equation tells us that one and only one molecule of Cl_2 reacts with one and only one formula unit of Zn to produce one and only one formula unit of $ZnCl_2$. The reactants and product are all in a one-to-one-to-one ratio. No other ratio can fit this reaction.

It is also common in chemical equations to show the phases of the various reagents. Sometimes this is important and sometimes it is not. When it is done, phases are included within parentheses for each reagent. For the above reaction, the equation with phases is written as follows.

$$Cl_2(g) + Zn(s) \rightarrow ZnCl_2(s)$$

(*g*) means gas phase. (*s*) means solid phase. There are also (*l*) which means liquid phase and (*aq*) which means aqueous phase. Aqueous phase means dissolved in water, not just floating around in water. Floaters don't count as aqueous; it must be fully dissolved in water in order to count as (*aq*).

We don't need phase information right now so I won't use it presently. We will need it in a later Chapter, so keep it in mind. Some instructors require phase information throughout; be aware of this if necessary.

Our example above is a one to one ratio of everything, but most reactions are not that simple. The equations for some reactions can be quite tedious to balance. We need to look at more details of balance and other aspects of equations.

6.2 Getting into details

Let's go through a balancing example step by step. We'll do the combustion of hydrogen. I pictured the molecules above, but let's work through the details.

The first step in any equation is to show reactants on the left and products on the right. All reactant formulas and all product formulas must be known before attempting any balancing act. The reaction in this case is between the elements H_2 and O_2 to produce water, H_2O.

(unbalanced) $$H_2 + O_2 \rightarrow H_2O$$

Once the formulas of all reagents are listed, you can start your count. Here, there are two H's on the left and two H's on the right so the H's balance. There are two O's on the left but only one O on the right, so that's no good. The equation is not balanced. You must balance it.

Balancing is not an automatic thing. It takes practice. Sometimes it's a juggling trick for the numbers. There are no universal methods for balancing every equation. Some instructors have some methods for certain kinds of reactions. These methods work fine for those reactions but not for others. As far as I'm concerned, the only universal method that works is called "eyeball balance", which means you stare at it until you get it right.

Balancing is achieved by the use of coefficients in the equation. The coefficients designate the relative ratios of the reagents. For our H_2/O_2 reaction, we can balance the O's by entering a coefficient of two on H_2O.

(unbalanced) $$H_2 + O_2 \rightarrow \underline{2}\, H_2O$$

These coefficients are just like algebra coefficients: they multiply whatever comes next. Here, the coefficient of two means that two molecules of H_2O are involved; this brings the count of O's to two on the right. Oxygens are now balanced for both sides. Unfortunately, the coefficient screwed up the H balance: now there are two H's on the left but a total of four H's on the right. Have no fear, another coefficient can handle that.

$$\underline{2}\ H_2 + O_2 \rightarrow 2\ H_2O$$

Four H's left and right. Two O's left and right. It's fully balanced. Catch the significance of the balanced equation: for this reaction, two molecules of H_2 react with one molecule of O_2 to produce two molecules of H_2O. No other ratio can fit this reaction.

IMPORTANT! ALL BALANCING IS DONE WITH COEFFICIENTS. Which means DON'T YOU DARE TOUCH THOSE SUBSCRIPTS! Some students run into troubles this way. Let me show you an example. Go back to the start of the H_2 combustion example.

(unbalanced) $$H_2 + O_2 \rightarrow H_2O$$

It would be tempting to balance the equation by just a little change in the product formula. Why not just make it H_2O_2?

$$H_2 + O_2 \rightarrow H_2O_2$$

Yes, it's balanced, but now you've got the wrong product! That's not water, it's hydrogen peroxide! Which means <u>you balanced the wrong equation</u>. Remember what I said earlier.

> THE FIRST STEP IN ANY EQUATION IS TO SHOW REACTANTS ON THE LEFT AND PRODUCTS ON THE RIGHT. ALL REACTANT FORMULAS AND ALL PRODUCT FORMULAS MUST BE KNOWN BEFORE ATTEMPTING ANY BALANCING ACT.

Once a reagent's formula is specified, you <u>cannot change subscripts</u> because that changes identity. You can only use coefficients in order to balance equations.

By the way, you may have noticed by now that you don't have to write in a coefficient when the coefficient is one. Coefficients of one are generally not shown, so these should be left out in the end.

In my examples above for Cl_2/Zn and H_2/O_2, I made a point about the meaning of the balanced equation in terms of the number of chemical units involved. I mentioned this in terms of individual molecules and formula units but the number can also be in moles, since mole is a number. We can restate the conclusions as follows.

$$Cl_2 + Zn \rightarrow ZnCl_2$$

> One mole of Cl_2 molecules reacts with one mole of Zn formula units to produce one mole of formula units of $ZnCl_2$.

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

> Two moles of H_2 molecules react with one mole of O_2 molecules to produce two moles of H_2O molecules.

The mole interpretation will be the one which we use the most.

Let me make some final points about coefficients. There are occasions where fractional coefficients are useful. Let's restart the H_2/O_2 example.

(unbalanced) $$H_2 + O_2 \rightarrow H_2O$$

The initial problem was that the O's were not balanced. Above, we used a coefficient of two on the H_2O. Another option is to use a fractional coefficient on the O_2.

$$H_2 + \frac{1}{2}\ O_2 \rightarrow H_2O$$

This also achieves balance: two H's on both sides and one O on both sides. The mole interpretation still applies: one mole of H_2 molecules reacts with one-half mole of O_2 molecules to produce one mole of H_2O molecules. The molecule interpretation, however, cannot be taken literally since one-half molecule of O_2 has no meaning. Nevertheless, the counts do work out and the equation is still balanced.

Fractional coefficients can be troublesome for the types of calculations which we encounter in later Chapters. You can always clear out the fractional coefficients by multiplying through the entire equation by some number. Notice that if you take the above equation

$$H_2 + \frac{1}{2} O_2 \rightarrow H_2O$$

and multiply everything by two, then you get

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

which is what we had before. This is totally legal. So go ahead and use fractions to get the balance, but then clear them in the end by multiplying them out. In most cases, coefficients should be in the smallest, whole numbers. Some instructors ban fractional coefficients completely, so be aware of this if it applies for you. Ironically, there are certain definitions of equations which require fractional coefficients, but we won't need these until Chapter 19.

Speaking of multiplying coefficients, don't overdo it. I said coefficients should be in the smallest, whole numbers. The following is wrong.

$$4\ H_2 + 2\ O_2 \rightarrow 4\ H_2O$$

Here, you would divide through by two in order to obtain the smallest, whole numbers for the coefficients.

Let's do something different.

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Example 1. Consider the reaction of lithium carbonate with hydrogen bromide, which produces lithium bromide, carbon dioxide and water. Balance the equation for that reaction.

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What's first? Well, you need some formulas.

Reactants	lithium carbonate, Li_2CO_3	Products	lithium bromide, LiBr
	hydrogen bromide, HBr		carbon dioxide, CO_2
			water, H_2O

Now set up the equation: reactants go on the left, products go on the right. Draw in your arrow. Leave space for coefficients.

(unbalanced) $\_\ Li_2CO_3 + \_\ HBr \rightarrow \_\ LiBr + \_\ CO_2 + \_\ H_2O$

There are five coefficients to worry about. You can actually start with the count for Li or for C or for O or for H or for Br and go in any sequence. I'll just go in the sequence as written.

‣ Li: There are two Li's on the left and only one Li on the right. That's no good. Enter a coefficient of two on LiBr in order to balance lithiums.

(unbalanced) $\_\ Li_2CO_3 + \_\ HBr \rightarrow \underline{2}\ LiBr + \_\ CO_2 + \_\ H_2O$

Now, here's a catch: as soon as you entered that coefficient, you have actually committed to two coefficients. By entering 2 on LiBr, you have committed to a coefficient of one for Li_2CO_3. This is a subtle but important point.

(unbalanced) $\underline{1}\ Li_2CO_3 + \_\ HBr \rightarrow 2\ LiBr + \_\ CO_2 + \_\ H_2O$

Feel free to enter ones as you balance. This can help you keep track of which coefficients are done and which ones need help.

‣ C: There is one C on the left, and that number has been fixed by the coefficient of one. You need one C on the right. Give CO_2 a coefficient of one.

(unbalanced) $1\ Li_2CO_3 + \_\ HBr \rightarrow 2\ LiBr + \underline{1}\ CO_2 + \_\ H_2O$

‣ O: There are three O's on the left, as set by the coefficient for Li_2CO_3. You must have three O's on the right. Two have already been fixed by the coefficient for CO_2. For the third O, give a coefficient of one to H_2O.

(unbalanced) $1\ Li_2CO_3 + \_\ HBr \rightarrow 2\ LiBr + 1\ CO_2 + \underline{1}\ H_2O$

‣ H: The number of H's on the left hasn't been entered yet, but the number on the right has been fixed at two. You need a two for the HBr.

$$1\ Li_2CO_3 + \underline{2}\ HBr \rightarrow 2\ LiBr + 1\ CO_2 + 1\ H_2O$$

▸ Br: There are two Br's on the left and two Br's on the right, as determined by the coefficients so far. Fine. Leave them. Everybody's happy. Like I said earlier, though, coefficients of one don't have to be shown, so we can clear them in the end.

$$Li_2CO_3 + 2\ HBr \rightarrow 2\ LiBr + CO_2 + H_2O$$

That's it.

Sometimes you'll start a balancing problem in some sequence for the elements and then get stuck. That's not unusual. One useful point to keep in mind is that <u>if the balance is not working out, start over but use a different sequence</u> for the elements.

I want to end this section with two points to note. One long, one short.

First, chemical equations only involve the before and after. They don't say a thing about in-between. They don't say a thing about HOW. How did the atoms form the new compounds? How did the bonds break? How did the new ones form? Let's reconsider the H_2/O_2 reaction. Go back and take another look at the picture of the molecules which I presented in the beginning of the Chapter. Notice the bonds involved. There is a covalent bond in the H_2 molecule which connects the two H's. There is a covalent bond in the O_2 molecule which connects the two O's. During the reaction, these bonds are broken. How? There are two covalent bonds in each H_2O molecule which connects the O to the H's. How did these form? What was the actual sequence of events? Let's consider a more complicated case, our example with Li_2CO_3 and HBr. This reaction involves breaking ionic bonds and covalent bonds and making new ones of each type. How does all this happen? What is the actual sequence of events?

All of these questions are very important. They address the critical, fundamental notions of HOW chemical units come together and undergo chemical change to form new products. Think about the following. I mentioned hydrogen burns violently and even explosively, but if you combine pure hydrogen gas and pure oxygen gas in a container, they could sit there for years and years and do nothing. Put a match to it, however, and it blows skyhigh. You know from your life's experiences that you have to start a fire with a fire. All of these things have to do with HOW a reaction actually occurs.

(Actually, in some cases, you can start a fire without a spark or without any heat whatsoever. Pyrophoric compounds will start on fire just by themselves when they are exposed to air, just at normal temperature. They're a real nuisance to work with.)

We call the HOW of a reaction its "mechanism", meaning the "mechanism" by which the old bonds are broken and the new bonds are made. There are many aspects to mechanism. It depends on the atoms within the chemical units which are involved. It depends on the shapes of those chemical units. It depends on the bonds which are involved. It depends on the energy which is available. It depends on a lot of things which we haven't gotten to yet. So for now, I won't dwell further on mechanisms. After we've done some of these other aspects, we'll talk quite a bit about mechanisms starting in Chapter 48. This is all part of the Grand Puzzle.

The second point before closing this section is that chemical equations share many similarities to algebraic equations. Each chemical formula represents an *x*, *y* or *z*. Coefficients are used to make things equal or balanced. The arrow is equivalent to the equal sign. And, just like an algebra problem, THE BEST WAY TO LEARN IT IS TO DO IT. Do a bunch of problems. Then do another bunch. The examples above were easy. Some of them get much worse.

Practice, practice, practice.

6.3 Kinds of reactions: combustion

Chemical reactions can be classified in many cases by the type of reaction which they represent. We will encounter several types as we proceed through the upcoming chapters. There's one which I am going to elaborate upon at this time. It's combustion. This is a normal everyday word and it carries the normal everyday meaning. Well, almost. We'll get a bit more specific about the term.

Combustion reactions are common and are associated with burning. You are familiar with the term: the combustion of gasoline in a car engine, or the burning of wood, or etc. But we also speak of burning fats and sugars in our bodies. Have you ever gotten a grease fire on a grill? I do that often, although not deliberately. Grease burns on the grill and produces carbon dioxide and water. When you eat grease in your food, your body will also burn it to carbon dioxide and water. It's the same overall reaction. Within your body, you burn greases, you burn fats and you burn sugars, but you're not on fire. Thus, the term "combustion" does not require a flame. Although a flame is not always involved, something else is:

oxygen, O_2. The combustion of gasoline in your car engine is a reaction with the O_2 from the air. The burning of wood is also a reaction with O_2 from the air. The primary reason that you even breathe is to provide O_2 for the combustion reactions in your cells. So we shall define a combustion reaction as any reaction between O_2 and one other reactant. This keeps things simple, although there are exceptions.

We mentioned the combustion of hydrogen above.

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

Hydrogen burns very well. Explosively, if not controlled. It's been used in a lot of rockets. Interestingly, hydrogen combustion may also be coming to an application near you, although it won't involve a rocket engine. Instead, combustion will occur within a "fuel cell" which produces electrical energy directly. It's still the same reaction as above, but there's no flame. Such devices are already in use in various roles, such as backup/emergency electrical generators and generators in remote locations far from power lines. Someday they may be widespread in vehicles. We'll discuss these a bit more in Chapter 64.

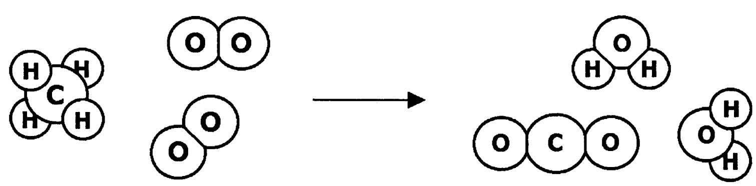
Some metals can burn. You may not know this but you've probably seen it. Magnesium burns with a brilliant white flame and small particles of Mg metal are in many fireworks to produce the white sparks. The combustion reaction can be written as the following equation.

$$2\,Mg + O_2 \rightarrow 2\,MgO$$

Many compounds also do combustion. For example, carbon compounds are very good at this. Most fuels in use today, including coal, gas and gasoline, are mixtures of carbon compounds. One of the simplest carbon compounds is methane, CH_4, which is a colorless and odorless gas. (I mentioned this one and others near the end of Chapter 4.) Methane is one of the components in natural gas. It's also called marsh gas since it is produced by natural decay processes in swamps. Micro-organisms in the guts of animals also produce methane as they decompose food passing through the digestive tract; the methane mixes with other gases and vents either forward or rearward. (The technical terms are eructation and flatus. The nontechnical terms are belches and farts.) Methane is a greenhouse gas which makes it an environmental concern. Releases of methane to the environment are carefully studied, even when Nature is doing the release. As an example, millions of tons of CH_4 can be produced and released annually just from cattle. As a fuel, the combustion reaction for methane is the following.

$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$$

Let me throw in another picture of the molecular changes which are involved in chemical reactions. Such pictures are better at emphasizing how much the chemical units change. For this reaction, we start with a methane molecule and two diatomic O_2 molecules. We form one carbon dioxide molecule and two water molecules. As we talk about this and any kind of reaction, you should keep in mind this idea of chemical change. The balanced chemical equation is a convenient method of conveying the basic information but the real picture involves all of these changes in the chemical units.



OK, let's do a balancing Example with combustion.

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Example 2. Isopropyl alcohol, C_3H_8O, is the stuff in rubbing alcohol. Read the label. The rubbing alcohol sold in stores typically runs ~70% isopropyl alcohol and the rest is water; higher percent amounts are also available. Isopropyl alcohol burns to produce carbon dioxide, CO_2, and water, H_2O. Write the balanced equation for the combustion.

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Line up your reactants and products.

(unbalanced) $\_\,C_3H_8O + \_\,O_2 \rightarrow \_\,CO_2 + \_\,H_2O$

Here is one clue for balancing combustions which works well: do O's last. Now, give it a whirl.

‣ C: As currently written, there are three C's on the left and one C on the right. Enter a coefficient of three on CO_2 in order to balance all carbons on both sides.

(unbalanced) $\_\ C_3H_8O\ +\ \_\ O_2\ \rightarrow\ \underline{3}\ CO_2\ +\ \_\ H_2O$

By entering the three on CO_2, you have taken a coefficient of one for C_3H_8O.

(unbalanced) $\underline{1}\ C_3H_8O\ +\ \_\ O_2\ \rightarrow\ 3\ CO_2\ +\ \_\ H_2O$

‣ H: There are eight H's on the left by the assigned coefficient for C_3H_8O. We need eight H's on the right. Enter a coefficient of four on H_2O in order to balance the H's.

(unbalanced) $1\ C_3H_8O\ +\ \_\ O_2\ \rightarrow\ 3\ CO_2\ +\ \underline{4}\ H_2O$

‣ O: Notice that the right side has all coefficients assigned. That fixes the number of O's at a total of ten on the right. We need a total of ten on the left. One O on the left has been fixed by the coefficient for C_3H_8O. We need nine more O's on the left, and they must be provided by the coefficient of O_2. We can do this with a fractional coefficient.

$$1\ C_3H_8O\ +\ \frac{9}{2}O_2\ \rightarrow\ 3\ CO_2\ +\ 4\ H_2O$$

That's balanced. Before we call it done, let's clear the fraction by multiplying through by two.

$$2\ C_3H_8O\ +\ 9\ O_2\ \rightarrow\ 6\ CO_2\ +\ 8\ H_2O$$

Whenever you multiply through, every coefficient gets multiplied. It's just like the connection to algebra equations which I mentioned earlier. Be careful of this. Messing this up is a common error by students because they forget to multiply every coefficient.

Here's one last point from this Example. If O's are in the reactant which is combusting, then those O's count the same as the O's from O_2. They all just scramble together. In the isopropyl alcohol example, the original O in C_3H_8O ends up either in the CO_2 or in the H_2O and it doesn't matter which. Only the count matters for balance.

Since combustion reactions are such a common experience and since there are patterns to the products which can form, it is frequently useful (and frequently required by instructors) for you to know a bit more about them. Remember our working definition: a combustion is any reaction between O_2 and one other reactant. Here are common guidelines for combustion reactions.

1. O_2 is one of two reactants.
2. If the other reactant is elemental carbon or a carbon compound, then carbon dioxide (CO_2) will be one of the products.
3. If the other reactant is elemental hydrogen or a hydrogen compound, then water (H_2O) will be one of the products.
4. If the other reactant is an elemental metal or a metal compound, then a metal oxide is a product. This will take an ionic formula. If the metal cation is of the constant charge type, then the formula of the product can be determined from the charge. If the metal cation is of the variable charge type, then more information is needed in order to know which product is intended.

We can see these four guidelines in the combustion equations which I wrote above. All of them had O_2 as a reactant (#1). The equation for the combustion of methane (CH_4) and for the combustion of isopropyl alcohol (C_3H_8O) both had CO_2 as a product (#2). The equation for the combustion of elemental hydrogen (H_2), for the combustion of methane (CH_4) and for the combustion of isopropyl alcohol (C_3H_8O) all had H_2O as a product (#3). The equation for the combustion of magnesium had magnesium oxide (MgO) as a product (#4).

Just like every set of guidelines, they're called guidelines and they're not called rules for a reason. Exceptions are possible.

Let's do a new Example of combustion starting completely from scratch. I'll pick something unusual: C_2H_6Zn.

Example 3. Write a balanced equation for the combustion of C_2H_6Zn.

First step: put reactants on the left, products on the right. So far we've got C_2H_6Zn as a reactant.

(unbalanced) $C_2H_6Zn\ \rightarrow$

Since it's combustion, we also know O_2 is a reactant.

(unbalanced) $C_2H_6Zn + O_2 \rightarrow$

According to the guidelines above, carbon dioxide (CO_2), water (H_2O) and a metal oxide compound for zinc will be the products. Zinc forms a constant charge cation (Zn^{2+}), so the product will be ZnO, zinc oxide. Now we have the formulas for everything involved. Write them into the equation.

(unbalanced) $\_\ C_2H_6Zn + \_\ O_2 \rightarrow \_\ CO_2 + \_\ H_2O + \_\ ZnO$

Notice that there are five coefficients to worry about.

‣ C: As initially written, there are two C's on the left and one C on the right. Double the CO_2.

(unbalanced) $\_\ C_2H_6Zn + \_\ O_2 \rightarrow \underline{2}\ CO_2 + \_\ H_2O + \_\ ZnO$

By this move, you have taken one to be the coefficient on C_2H_6Zn.

(unbalanced) $\underline{1}\ C_2H_6Zn + \_\ O_2 \rightarrow 2\ CO_2 + \_\ H_2O + \_\ ZnO$

‣ H: There are six H's on the left as set by the coefficient for C_2H_6Zn. You need six H's on the right. Triple the H_2O.

(unbalanced) $1\ C_2H_6Zn + \_\ O_2 \rightarrow 2\ CO_2 + \underline{3}\ H_2O + \_\ ZnO$

‣ Zn: There's one Zn on the left; we need one Zn on the right. Set ZnO at one on the right.

(unbalanced) $1\ C_2H_6Zn + \_\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O + \underline{1}\ ZnO$

‣ O: Only one coefficient remains and this one must balance the O's. There are eight O's total on the right, so we need eight on the left. Give O_2 a coefficient of four.

$$1\ C_2H_6Zn + \underline{4}\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O + 1\ ZnO$$

Clear your ones.

$$C_2H_6Zn + 4\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O + ZnO$$

Done.

This ends our discussion for now for setting up and balancing equations. Now we turn to a different matter regarding chemical equations.

6.4 Handling multi-step processes

A multi-step process is one which can be broken down into two or more steps. These are very common in biology and also in industry. Earlier I mentioned the combustion of fats and sugars in your body. This does not produce a flame because Nature has designed an incredibly nifty system to do it in steps, releasing parts of the energy at a time. The combustion of one fat molecule or one sugar molecule in your body could involve dozens of steps.

When dealing with a multi-step process, you can deal with one of the steps at a time or you can deal with the grand overall result. This means that you can write balanced equations for each of the steps or you can write an "overall" equation. The overall equation can be derived by adding all of the individual equations. This brings up an important point about chemical equations: <u>chemical equations are additive</u>. This is a very useful point here and it will prove very important in Chapter 19 also. Notice that this is another connection to algebra equations of *x*'s, *y*'s and *z*'s. If you have two algebra equations, these can be added together to produce an overall equation, cancelling things that appear on both sides. Let's look at an algebraic example.

Given this equation: $x + 3y = 7 + 4z^2$

and this equation: $2x + z^2 = 8 + y$

add them together to get: $3x + 3y + z^2 = 15 + 4z^2 + y$

which simplifies to: $3x + 2y = 15 + 3z^2$

You can do the same thing with chemical equations: you can add them together to get a sum, and you would cancel things on both sides. Let me show you how to apply this.

Let's talk about how hydrazine, N_2H_4, is made. Hydrazine is a very important compound by itself and it's also important for making other compounds; despite this, it is not a compound most people typically hear about. Hydrazine is used to make pharmaceutical and agricultural compounds. It's used in boiler water to prevent corrosion. It or compounds made from it are used in rocket engines. Thus, the stuff gets around, whether you hear about it or not.

The industrial preparation of hydrazine is a multi-step chemical process. The overall process uses ammonia, NH_3, and sodium hypochlorite, NaClO, as reactants. I mentioned ammonia near the end of Chapter 4 and how it's the stuff sold in stores dissolved in water. You're probably also familiar with solutions of sodium hypochlorite although you may not know it. It's the active ingredient in the typical, liquid laundry bleaches. Read it on the label. No, it's not the powdered bleaches; it's the one they also call "chlorine bleach". NaClO is what does the job. It's strong stuff.

Needless to say, big industry does not go to the grocery store to get their ingredients, although the compounds in the solutions are the same ones. The overall balanced equation for making hydrazine industrially is the following.

$$2\ NH_3 + NaClO \rightarrow N_2H_4 + NaCl + H_2O$$

We can break this overall process into three steps. In the first step, one ammonia molecule and one sodium hypochlorite formula unit react.

‣ Step 1 $$NH_3 + NaClO \rightarrow ClNH_2 + NaOH$$

The products are $ClNH_2$, which is called "chloramine", and sodium hydroxide. Chloramine is nasty stuff. Can kill quickly. In industry, they take suitable precautions for handling this. Unfortunately, homeowners don't always take the necessary precautions. How does this apply to homeowners? Cleaning. Ammonia is used for household cleaning. Even if you're not using a bottle called ammonia, a lot of household cleaners still have ammonia in them. Read the label. Chlorine bleach contains sodium hypochlorite. Bleach is a powerful cleaner so some people think they can add bleach and ammonia and get something even more powerful. Maybe you've heard the warnings: never mix bleach and ammonia products. Chloramine is why. Mixing bleach and ammonia products in your bucket makes chloramine and other related compounds which can damage your lungs or even drop you dead. Yes, it has killed numerous people in the past. This is real. Don't even think about doing it. Is your toilet bowl worth dying for?

Ironically, you may be drinking tap water which contains chloramine in small amounts. It's been used for many years as a disinfecting agent for some municipal water supplies. Chlorine has also been used in water supplies, and there are pros and cons in the selection of either.

Let's get back to hydrazine.

In Step 2, the chloramine molecule produced in Step 1 reacts with another molecule of ammonia to make one formula unit of N_2H_5Cl. This compound is ionic: it has a polyatomic cation, $N_2H_5^+$, and a chloride anion, Cl^-.

‣ Step 2 $$ClNH_2 + NH_3 \rightarrow N_2H_5Cl$$

The other product from Step 1 was sodium hydroxide, NaOH. It reacts with the N_2H_5Cl from Step 2.

‣ Step 3 $$N_2H_5Cl + NaOH \rightarrow N_2H_4 + NaCl + H_2O$$

Those are our three individual steps. These all occur while the reaction proceeds. They all happen in the same solution in the same pot. (This is not true of all multi-step processes. Some require separate solutions in separate containers for the different steps.) Now, let's add the three steps together to get an overall equation for the process as a whole. All reagents to the left of the arrows add together. All reagents to the right of the arrows add together.

Step 1	$NH_3 + NaClO$	$\rightarrow$	$ClNH_2 + NaOH$
Step 2	$ClNH_2 + NH_3$	$\rightarrow$	N_2H_5Cl
Step 3	$N_2H_5Cl + NaOH$	$\rightarrow$	$N_2H_4 + NaCl + H_2O$
‣ Sum	$NH_3 + NaClO + ClNH_2 + NH_3 + N_2H_5Cl + NaOH$	$\rightarrow$	$ClNH_2 + NaOH + N_2H_5Cl + N_2H_4 + NaCl + H_2O$

Cancel out anything which is on both the left and right sides. That includes $ClNH_2$, N_2H_5Cl and NaOH. Go ahead, cross them out on both sides. The equation simplifies to

$$2\ NH_3 + NaClO \rightarrow N_2H_4 + NaCl + H_2O$$

which is the exact same overall equation which I gave earlier above.

By the way, anything which exactly cancels out of this additivity process is called an intermediate. Intermediates are produced in one step but they react in the same amount in a subsequent step. As long as they are produced and reacted in the same amount, then they cancel out of the final, overall equation. For our example here, $ClNH_2$, N_2H_5Cl and NaOH are all intermediates.

This is a simple illustration of the additivity of reactions. What makes this a simple illustration is the fact that all coefficients are one in all of the equations which are involved. Let me now show a twist which can come into play.

At high temperature, silver sulfate reacts to form elemental silver metal, SO_3 and O_2. We can break this overall process into two Steps, with Ag_2O as an intermediate.

- Step 1 $$Ag_2SO_4 \rightarrow Ag_2O + SO_3$$
- Step 2 $$2\ Ag_2O \rightarrow 4\ Ag + O_2$$

If you go ahead and add these two Steps together, you get

$$Ag_2SO_4 + 2\ Ag_2O \rightarrow Ag_2O + SO_3 + 4\ Ag + O_2$$

which simplifies to

$$Ag_2SO_4 + Ag_2O \rightarrow SO_3 + 4\ Ag + O_2$$

but that does not match the description of the process. In order for Ag_2O to be an intermediate, it must cancel out overall.

In order for Ag_2O to cancel out, it must be produced and reacted in the same amount. One mole is produced in Step 1 but two moles are reacted in Step 2. We can get these amounts to equal by doubling Step 1, and adding this equation to Step 2.

Double Step 1 $$2\ Ag_2SO_4 \rightarrow 2\ Ag_2O + 2\ SO_3$$
Step 2 $$2\ Ag_2O \rightarrow 4\ Ag + O_2$$

- Sum $$2\ Ag_2SO_4 + 2\ Ag_2O \rightarrow 2\ Ag_2O + 2\ SO_3 + 4\ Ag + O_2$$

This simplifies to

$$2\ Ag_2SO_4 \rightarrow 2\ SO_3 + 4\ Ag + O_2$$

and that is our final equation for the overall process. As you can see, this example differs a bit from the hydrazine example and these serve to illustrate some of the variations which are possible.

The additivity of reactions is a natural consequence of the conservation of matter, the conservation of mass, the conservation of atoms. All atoms are conserved, regardless of the number of steps. While additivity is introduced here with respect to conservation of matter, it also has a role in the conservation of energy. We'll see that relationship later in Chapter 19.

Problems

1. Balance the following equations.
   a. B_2H_6 + HF → BF_3 + H_2
   b. PCl_5 + H_2O → H_3PO_4 + HCl
   c. Cl_2SO + H_2O → SO_2 + HCl
   d. K_3AsO_4 + $CaCO_3$ → $Ca_3(AsO_4)_2$ + K_2CO_3

2. Balance the following equations.
   a. F_2CO + $SbCl_5$ → Cl_2CO + SbF_5
   b. Fe_2O_3 + HBr → $FeBr_3$ + H_2O
   c. H_2S + O_2 → H_2O + SO_2
   d. $AgNO_3$ + $H_2C_2O_4$ → $Ag_2C_2O_4$ + HNO_3

3. Balance the following equations.

   a. Mg_2SiO_4 + HF → MgF_2 + H_2O + SiF_4

   b. C_2H_3Cl + O_2 → CO_2 + HCl + H_2O

   c. PbI_2 + $NaHSO_4$ → $PbSO_4$ + NaI + H_2SO_4

   d. $K_2Mo_2O_7$ + K_2S → K_2MoS_4 + K_2O

4. Write the balanced equation for each of the following.

   a. Aluminum chloride reacts with sodium phosphate to produce aluminum phosphate and sodium chloride.

   b. Cobalt(II) acetate reacts with ammonium sulfide to produce cobalt(II) sulfide and ammonium acetate.

   c. Chlorine trifluoride reacts with sodium hydroxide to produce sodium chlorite, sodium fluoride and water.

5. Write the balanced equation for the combustion of C_3H_9Al.

6. Write the balanced equation for the combustion of table sugar, $C_{12}H_{22}O_{11}$.

7. Elemental chlorine reacts with sodium hydroxide in hot aqueous solution to produce sodium chloride, sodium chlorate and water. This overall reaction can be broken down into two steps.

   Step 1 $Cl_2 + 2\ NaOH \rightarrow NaCl + NaClO + H_2O$

   Step 2 $3\ NaClO \rightarrow 2\ NaCl + NaClO_3$

   Given that sodium hypochlorite is an intermediate, derive the equation for the overall reaction.

8. The reaction of calcium carbonate and hydrogen chloride can be broken into two steps. In the first step, calcium carbonate reacts with hydrogen chloride to produce calcium chloride and H_2CO_3. In the second step, H_2CO_3 reacts completely, by itself, to form carbon dioxide and water. Write the balanced equation for each step and for the overall process. Identify any intermediate(s) which is (are) involved.