Balancing Combustion Reactions

Students often stress out over the process of balancing combustion reactions. Often they will find that they are close to solving the problem, and realize that they need to start over. Yet, combustion reactions are in fact quite easy to solve once simple techniques are learned.

To overview, combustion reactions are generally the reactions of some with oxygen. In words, it is either referred to as "combustion" or "burning." For this tutorial, combustion will refer to the reaction of organic materials (containing at minimum carbon and hydrogen) and oxygen. In addition, combustion carries some hidden meanings with it. The word can also indicate what the products are going to be. For example, if something is designated as a "complete combustion," the products are automatically assumed to be carbon dioxide and water and energy. This looks like the following

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

However, not all combustions are perfect (complete). In addition to the products listed above, other products are produced, particularly carbon monoxide. For the purposes of this tutorial, we will focus only on complete combustions

Process for Balancing Reactions

For reactions containing only carbon, hydrogen and oxygen

1. Balance the carbon atoms
2. Balance the hydrogen atoms
3. Balance the oxygen atoms

We will start off with an easy example: the combustion of ethane. This reaction looks like

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

So for the first step, we need to balance the carbons. There are two carbons on the left side, and only one on the right. Thus, if we multiply the CO_2 by two, both sides of the equation will be balanced with respect to carbon.

Step 1: $C_2H_6(g) + O_2(g) \rightarrow 2CO_2(g) + H_2O$

Next we will focus on the hydrogen atoms. There are 6 hydrogen atoms on the left side, and only two on the right side (in water). Multiplying the water by 3 will balance each side with six total hydrogen atoms. We have then

Step 2: $C_2H_6(g) + O_2(g) \rightarrow 2CO_2(g) + 6H_2O$

Finally, focusing on the oxygen, we find that there are two ***monoatomic*** oxygen atoms on the left side of the equation and 10 on the right (two oxygen atoms in the carbon dioxide, and this molecule was multiplied by two, giving a total of four oxygen atoms; and six oxygen atoms in the water). Multiplying the left side by five will give a total of 10 oxygen on each side of the equation

Step 3 - $C_2H_6(g)+5O_2(g)\rightarrow 2CO_2(g)+6H_2O$

This reaction is balance. Of course, this was pretty straight forward. What happens if we end up with a fractional number of molecules? Let's examine the combustion of ethanol

$$C_2H_5OH\,(g)+O_2(g)\rightarrow CO_2(g)+H_2O$$

This reaction is almost identical to the combustion of ethane, except the organic molecule now contains on oxygen. Steps 1 and 2 are the same as before, so we initially have

Step 1 & 2 - $C_2H_5OH(l)+O_2(g)\rightarrow 2CO_2(g)+6H_2O$

Focusing on the oxygen atoms, there are again 10 oxygen atoms on the right side, but this time there are three on the left (one in ethanol and two in diatomic oxygen). The only number that we are allowed to change is the number of moles of O_2 . In other words, the ten oxygen atoms on the right hand side of the equation are fixed. Therefore, we must solve an equation such that the total oxygen atoms from the left add up to those on the right (ie 10).

$$1Oxygen_{ethanol}+x\cdot Oxygen_{diatomic\ oxygen}=10\,Oxygen$$

Or simply

$$1+x=10$$

However, the oxygen molecule is a diatomic, so we must rewrite our equation to reflect this

$$1+2x=10$$

The 2x shows that there are two monoatomic oxygen atoms in the diatomic oxygen molecule

Solving this equation is easy, giving us x = 9 / 2. Our final equation turns out to be

Step 3 - $C_2H_5OH(l)+\frac{9}{2}O_2(g)\rightarrow 2CO_2(g)+6H_2O$

Though this is technically balanced correctly, it does not reflect the standard practice of putting chemical equations into a form with integer coefficients. This problem often arises from organic species that have an *odd number of oxygen atoms* in the molecule. With ethanol, there is only one oxygen atom. There are a couple of ways to get around this problem.

1. The harder way – Start over and multiply the organic species by two
    a. However, this does not always work out, especially if there is another product involved other than carbon dioxide and water. This will make balancing even more difficult. We will not focus on this here as it is essentially guessing and checking

2. The easy way – Multiply the entire equation by the denominator of the oxygen fraction
    a. Everything that we have calculated thus far has been correct
    b. The trick is to treat the chemical equation like a mathematical equation

So, taking where we left off with our ethanol combustion

$$C_2H_5OH(l)+\frac{9}{2}O_2(g)\rightarrow 2CO_2(g)+6H_2O$$

The denominator of the oxygen coefficient is 2. So, if we multiply the entire equation by two, we have

$$\left[C_2H_5OH(l)+\frac{9}{2}O_2(g)\rightarrow 2CO_2(g)+6H_2O\right]\cdot 2$$
$$=2C_2H_5OH(l)+9O_2(g)\rightarrow 4CO_2(g)+12H_2O$$

This is the final and correct answer for the combustion of ethanol. Notice that all of the coefficients are reduced and integers.

To illustrate how this simple method makes the whole process easier, lets examine the hypothetical complete combustion of okadaic acid, or $C_{44}H_{68}O_{13}$, with carbon monoxide as a by-product. We have then

$$C_{44}H_{68}O_{13}(s)+O_2(g)\rightarrow CO_2(g)+H_2O$$

Step 1 – Balance the carbon atoms

$$C_{44}H_{68}O_{13}(s)+O_2(g)\rightarrow 44CO_2(g)+H_2O$$

Step 2 – Balance the hydrogen atoms

$$C_{44}H_{68}O_{13}(s)+O_2(g)\rightarrow 44CO_2(g)+34H_2O$$

Step 3 – Balance the oxygen atoms

- There are 122 oxygen atoms s on the right hand side (88 in CO2 and 34 in Water)
- There are 13 invariable oxygen atoms (in the okadaic acid) and two variable oxygen atoms in the diatomic oxygen. Thus, our equation looks like
  $13+2x=122$

- Solving for x gives $x = 109 / 2$

Our equation then looks like

$$C_{44}H_{68}O_{13}(s)+\frac{109}{2}O_2(g)\rightarrow 44CO_2(g)+34H_2O$$

Multiplying it by two gives the final equation

$$2C_{44}H_{68}O_{13}(s)+109O_2(g)\rightarrow 88CO_2(g)+68H_2O$$

Though the methods demonstrated here have only been for organic species, they can also be applied to other species, such as sulfur whose reactions will not always give oxygen coefficients as fractions of 2. Regardless, after solving for all other elements, multiplying the equation by the oxygen coefficient denominator will yield the correct, reduced and balanced chemical equation for the particular reaction