Science-1A Lecture: Week-12, Monday, October 25, 2021
Balancing Chemical Reaction Equations

Counting Atoms

When we write the molecule for water, we write it as H_2O meaning that there are 2 hydrogen atoms attached to a single oxygen atom. If we write, $3H_2O$ we are considering 3 water molecules with a total of 6 hydrogens and 3 oxygens. Similarly, carbon dioxide CO_2 is a molecule with 1 carbon and 2 oxygens. We learned earlier that oxygen atoms in air are in pairs as O_2 molecules; individual oxygen atoms are exceedingly rare. Now, consider the **combustion** (reacting with oxygen to release energy) of the flammable gas methane CH_4. This reaction is represented by the chemical equation

$$CH_4 + 2O_2 \Rightarrow CO_2 + 2H_2O + \text{energy}$$

In a chemical reaction, no new atoms appear and none are destroyed, so if we count the different kinds of atoms in the reactants (the stuff at the left), each type must have the same count as in the products (the stuff at the right).

In this reaction, there is 1 carbon on the left, and there must still be exactly 1 carbon on the right side. The 4 hydrogens on the left must be matched by 4 hydrogens on the right. Each water molecule on the right has 2 hydrogens, and therefore the 2 water molecules on the right have the required total of 4 hydrogens. The 2 oxygen molecules each have 2 oxygen atoms so there are 4 oxygen atoms on the left. On the right side, the carbon dioxide molecule has 2 oxygen atoms, and each of the 2 water molecules has 1. That makes a total of 4 oxygen atoms on the right. The equation is balanced.

1 C, 4 H, 4 O on the left 1 C, 4 H, 4 O on the right

Balancing Chemical Combustion Equations

Lets now burn methyl alcohol CH_3OH. The obvious first try for its combustion equation is

$$CH_3OH + O_2 \Rightarrow CO_2 + H_2O + \text{energy} \quad \text{(not correct).}$$

Here, there is 1 carbon on each side and 3 oxygens on each side. That is good, but unfortunately there are 4 hydrogens on the left and only 2 on the right. We can fix that by having it produce 2 water molecules on the right. We then have

$$CH_3OH + O_2 \Rightarrow CO_2 + 2H_2O + \text{energy} \quad \text{(still not correct)}$$

Now, however, we have messed up the oxygen balance – there are 3 O's on the left, but 4 O's on the right. We need another O on the left where single O's only come with the methyl alcohol molecule. We must therefore use two methyl alcohol molecules and re-do our entire balancing procedure:

$$2CH_3OH + O_2 \Rightarrow 2CO_2 + 4H_2O + \text{energy} \quad \text{(still not correct)}$$

With 2 carbons on the left, we needed 2 carbon dioxide molecules on the right. We also now have 8 H's on the left and must match them with 4 waters on the right. We must then check the O's again. There are 4 O's on the left, but now we have 8 O's on the right. We can attain a balance if we add 2 more oxygen molecules on the left, each providing 2 more oxygen atoms for a total of 4 more:

$$2CH_3OH + 3O_2 \Rightarrow 2CO_2 + 4H_2O + \text{energy} \quad \text{(correct!)}$$

Counting them:

2 C, 8 H, 8 O on the left 2 C, 8 H, 8 O on the right

The steps to balancing these combustion reactions are as follows:

1. Start by assuming that there is only 1 molecule to be burned (CH_3OH in this example) and balance the carbons and hydrogens by putting suitable numbers before the CO_2 and H_2O molecules.
2. Count the oxygen atoms and see if they are balanced. If the oxygen atom count on the right is in excess of the number on the left by an **even** number, you just add a suitable number of O_2 molecules on the left and you are done.
3. If, however, the oxygen count on the right is in excess by an **odd** number, the number of molecules to be burned must be doubled. Then, when you re-balance the carbons and hydrogens, the oxygen excess will be an even number which can be fixed by increasing the number of oxygen molecules.
4. **Make sure that you have the smallest set of numbers that work.** For example, the equation

$$4CH_3OH + 6O_2 \Rightarrow 4CO_2 + 8H_2O + \text{energy}$$

is balanced, but should have all its coefficients divided by 2 just as you do when reducing fractions to lowest terms: 2/4 becomes 1/2.

Equations Actually can go Both Ways

The equations above, as well as most that we will talk about in this course, release energy. Once ignited, they proceed from reactants to products while releasing energy. Chemistry, however, really goes in both directions. This is easy to understand with dissolving salt discussed in the October 13 lecture notes at *https://yosemitefoothills.com/Science-1A/OnlineLectureAndLabNotes/Week-10-Lab-Wednesday-October-13-2021.pdf* . That reaction is best written with a double arrow (<=>).

$$\mathrm{NaCl(s)} \Leftrightarrow \mathrm{Na^+(aq)} + \mathrm{Cl^-(aq)}$$

When salt is added to pure water, it dissolves and the reaction proceeds from left to right, but as water evaporates from a saturated salt solution, the reaction goes from right to left and salt crystals form. Reaction directions depend on the concentrations of the reactants and products – excessive reactants cause more products to be produced, but excessive products can reverse the reaction. The double arrow makes the meanings of the terms “reactants” and “products” ambiguous. The direction of reaction is actually dependent on the concentrations of molecules at the left and right sides as well as on applied energy.

Our moon explorers obtained water and electricity from the reaction

$$2\,\mathrm{H_2(g)} + \mathrm{O_2(g)} \Rightarrow 2\,\mathrm{H_2O(l)} + \text{electricity}$$

This, however, was not done in an explosion (except for an unfortunate accident on Apollo 13), but rather by using high pressure (more concentrated) gasses on the left to cause a controlled amount of water and electricity to be produced. Devices for doing this are called **fuel cells**. The high pressure of the gas causes its molecules to be in such close proximity that they can be made to react in a non-explosive manner.

When we used electrolysis to collected H_2 and O_2 gas in a Baggie, they did not react while we added electricity to slowly break apart the water. A spark was necessary to trigger the chain reaction.

Examples and Practice

You are now ready to check your understanding by looking at the chemical reaction equations shown at *https://yosemitefoothills.com/Science-1A/Handouts/Week-11/BalancingChemicalEquations.pdf*

In each case, pretend that the number of molecules is missing and that you must figure out what they are. That is how they are presented on Quizzes and Tests. For example, with Octane, pretend you are starting with

$$\underline{\quad\ }\ \mathrm{C_8H_{18}} + \underline{\quad\ }\ \mathrm{O_2} \Rightarrow \underline{\quad\ }\ \mathrm{CO_2} + \underline{\quad\ }\ \mathrm{H_2O} + \text{energy}$$

You should always start at the left side by assuming a 1 for the number of hydrocarbons and 1 for the number of oxygen molecules. Then, figure out how many CO_2 and H_2O molecules must be produced. Finally, figure out how many more oxygen atoms are on the right than on the left. If that is an even number, you just need to increase the number of O_2 molecules on the left to balance them. But if that is an odd number, you will need to change the number of hydrocarbons to a 2 and start over again like we did above when burning methyl alcohol.

It is easy to get mixed up doing this. So after you think you have it, you must check by counting atoms on each side. For this octane example, the final equation is

$$2\,\mathrm{C_8H_{18}} + 25\,\mathrm{O_2} \Rightarrow 16\,\mathrm{CO_2} + 18\,\mathrm{H_2O} + \text{energy}$$

where we have 2x8=16 C’s, 2x18=36 H’s, and 25x2=50 O’s on the left and 16x1=16 C’s, 18x2=36 H’s, and 16x2+18x1=50 O’s on the right. **Be sure to also make sure your coefficients are the smallest set of numbers that work.**

The two reactions at the bottom of that page of combustion examples are different – they produce carbon monoxide CO, a deadly gas, instead of carbon dioxide CO_2. When something is burned without sufficient oxygen, carbon monoxide is produced. Many deaths are caused by cooking or heating with gas in a poorly ventilated room. The balancing procedure for these equations is the same, but you must be sure to remember that the reaction is producing CO instead of CO_2.