CHEMICAL REACTIONS—HOW MOLECULES ARE FORMED, BROKEN DOWN, AND REFORMED

In the course of a chemical reaction, the bonds that hold together the atoms that make up the reactants break. The free atoms then form new bonds with one another to form new molecules—the products of the reaction. Take a look at the chemical reaction below:

$$C_3H_8(g) + 5O_2(g) \rightarrow 4H_2O(l) + 3CO_2(g)$$

What does this equation tell us? Well, two things. On a molecular level, this equation says that 1 molecule of propane (C_3H_8) and 5 molecules of oxygen react to form 4 molecules of water and 3 molecules of carbon dioxide. It also tells us that 1 mole of C_3H_8 reacts with 5 moles of O_2 to form 4 moles of H_2O and 3 moles of CO_2. This equation also indicates the state of each reactant and product: (*s*) means solid, (*l*) means liquid, and (*g*) means gas.

Note that when a substance is dissolved in water to create a solution, it is in the aqueous phase, represented with (*aq*). This is not the same thing as the liquid phase. If you were to take table salt (NaCl (*s*)) and heat it until it melted, it would become NaCl (*l*). However, if you took that table salt and dissolved it into a glass of water, it would create NaCl (*aq*). Much of chemistry is solution-based, and frequently compounds will be dissolved in water to form aqueous solutions before any reactions take place.

Chemical Equations Must Be Balanced

Look again at the chemical equation we just presented:

$$C_3H_8(g) + 5O_2(g) \rightarrow 4H_2O(l) + 3CO_2(g)$$

This is a balanced equation. How can you tell? For each element on the left side of the equation, multiply the molecular coefficient by the element's subscript. (Any number that doesn't appear is assumed to be 1.) For oxygen, in $5O_2$, there are 5 × 2 = 10 oxygen atoms. Now do the same for the right side of the equation. In $4H_2O$, there are 4 × 1 = 4 oxygen atoms. In $3CO_2$, there are 3 × 2 = 6 oxygen atoms. So there are 4 + 6 = 10 oxygen atoms on the right side, and since there are also 10 oxygen atoms on the left side, oxygen is balanced. Now check to see that carbon and hydrogen are also balanced. They are. There are 3 carbons on the left and 3 carbons on the right. There are 8 hydrogens on the left and 8 hydrogens on the right. For each element in a balanced equation, the total number of atoms on the left must equal the total number of atoms on the right.

Balancing Equation Steps:

1. First, start with the smallest answer choice, and plug it in front of what you want to solve for.
2. Next, turn to the other side of the equation and fill in coefficients based on what you started with.
3. If you can't find whole number coefficients so all the atoms on either side of the equation add up, turn to the next smallest answer choice and repeat.
4. If you can balance the equation, check to make sure there is no common factor for all of the coefficients. For example, 4, 6, and 10 have a common factor of 2. If there is a common factor, divide the coefficients, and you have your answer. If there isn't, you already have your answer.

On the SAT Chemistry Subject Test, you may see up to five questions that will show you unbalanced equations and ask you to balance them. Here's what those will look like.

29. $\ldots C_2H_4(g) + \ldots O_2(g) \rightarrow \ldots CO_2(g) + \ldots H_2O(l)$

When the equation above is balanced and all coefficients are reduced to lowest whole-number terms, which of the following would be the coefficient for CO_2 ?

(A) 1

(B) 2

(C) 4
(D) 5
(E) 6

Fortunately, these questions are easy to answer if you use the “plug-in” balancing strategy. Start with choice A, and put the number 1 in front of CO_2. (Don’t be afraid to write in your test booklet as much as you want. Your test booklet belongs to you, and nobody cares what you write in it; only your answer sheet is scored.) By adding a 1 in front of CO_2, we end up with 1 carbon on the right. Yet, we have at least 2 carbons on the left, so we know that 1 is not the answer.

Let’s try choice B. Put a 2 in front of CO_2, and see what happens. We have 2 carbons on the right and 2 on the left. Good. We have 4 hydrogens on the left, so let’s try putting a 2 in front of the H_2O. That gives us a total of 6 oxygens on the right and 2 on the left. Let’s put a 3 in front of O_2 on the left, so we have a total of 6 oxygens on the left.

Now we have 2 carbons on the right and the left, 4 hydrogens on the right and the left, and 6 oxygens on the right and the left. The equation is balanced, and B is correct. So, every time a question asks you to balance an equation, use the plugging-in strategy. It can’t fail.

STOICHIOMETRY

Sometimes SAT Chemistry questions will ask you to determine how much product is formed or reactant is consumed in the course of a chemical reaction. These are stoichiometry questions. When you begin a stoichiometry problem, always remember to work from a balanced equation. The coefficients in front of each species indicate the molar ratio between the species. Consider the reaction between ammonia gas and oxygen, which yields nitrogen monoxide and water.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

This equation tells us that

- For every 4 moles of ammonia consumed, 5 moles of oxygen are also consumed.
- For every 5 moles of oxygen consumed, 6 moles of water are produced.
- For every 4 moles of ammonia consumed, 4 moles of nitrogen monoxide are produced. In other words, the molar ratio of ammonia consumption to nitrogen monoxide production is 1:1.

How do you put these molar ratios to use? Take a look: If 2 moles of ammonia are consumed, how many moles of water are produced?

From the balanced equation, we see that for every 4 moles of ammonia that are consumed, 6 moles of water are produced. So the ratio of ammonia to water is 4:6 or 2:3. So 2 moles of ammonia will react completely to produce 3 moles of water.

Remember that reactants combine according to their mole ratio. So if we combine two reactants that are not in their correct stoichiometric mole ratio, one reactant will be consumed first. When this happens, the excess reactant will remain unreacted.

Consider this question: If 34 g of ammonia and 32 g of oxygen are combined, how many grams of nitrogen monoxide will be produced?

The formula weight of NH_3 is 17 amu; 1 mole of NH_3 has a mass of 17 g and 2 moles have a mass of 34 g. This means that we have 2 moles of ammonia. How many moles of oxygen do we have? Since the formula weight of O_2 is 32 amu, 1 mole of O_2 has a mass of 32 g. So our actual mole ratio of ammonia to oxygen is 2:1. However, from the balanced equation given earlier in this section, we see that the ratio of ammonia consumption to oxygen consumption is 4:5. In keeping with this stoichiometric ratio, 2 moles of ammonia would react with 2.5 moles of oxygen—but we have only 1 mole of oxygen. We will run out of oxygen before all 2 moles of ammonia have reacted, and so oxygen acts as the **limiting reagent**. Once the limiting reagent has been consumed, the reaction will no longer proceed. So we know that 1 mole of oxygen actually reacts, and since 4 moles of nitrogen monoxide are produced for every 5 moles of oxygen that react, 4/5 mole, or 0.8 mole, of nitrogen monoxide is produced when 1 mole of oxygen reacts. Finally, the formula weight of NO is 30 amu, so 1 mole of NO has a mass of 30 g. The mass of 0.8 mole of NO is (0.8 mole)(30 g/mole), or 24 g.

Close, But NO Cigar

It's rare in a laboratory to get 100% of the expected product, as a variety of experimental conditions can cause a loss. The percent yield in a reaction describes the percentage of the product that was successfully created, which is written as $\frac{\text{actual}}{\text{expected}}$. So in this example, if only 20 g of NO were produced, the percent yield would be $\frac{20\text{ g}}{24\text{ g}} = \frac{5}{6} = 83\%$.