## Lecture Notes: Stoichiometry

I. What is stoichiometry?
   a. Technically, stoichiometry is the measurement of chemical quantities.
   b. However, in this course stoichiometry will usually refer to the use of a chemical equation to predict how much of some substance is produced or reacted based on the amount of some other substance that is involved in the reaction.

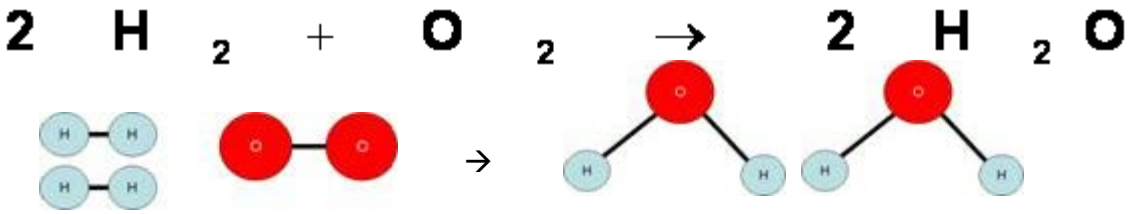
II. Why do we care about stoichiometry?
   a. This is "real" chemistry! That is, you will be able to predict how much of some chemical will be produced based on the starting amounts of the reactants. Also, you will be able to calculate how many grams of reactants will be needed to produce a given amount of some other chemical.
   b. Example: If I know how much steel I need, then how many tons of iron and carbon will be needed to produce that quantity of steel?
   c. Example: If I know how many tons of flour, eggs, milk, and sugar that I have, then how many cakes could I produce using a given recipe?

III. A review of how chemical reactions occur and the meaning of a chemical equation
   a. What is happening to atoms and molecules during a chemical reaction?
      i. Consider the unbalanced chemical equation for the reaction of hydrogen and oxygen below:

$$H_2 + O_2 \rightarrow H_2O$$

      ii. Now, consider the actual molecules reacting with each other:


   b. How is the law of conservation of matter obeyed?
      i. Clearly, neither of the two depictions above is accurate, because the number of atoms before the reaction is not equal to the number of atoms after the reaction. It is important to remember the law of conservation of matter: Atoms can neither be created nor destroyed during an ordinary chemical reaction.
      ii. Thus, in order for the law of conservation of matter to be obeyed, the reaction can only take place with the following ratio of substances:




   c. What is the meaning of the coefficients in a chemical equation?
      i. The coefficients in the equation, therefore, tell us how many molecules or moles of each substance are needed for the reaction to occur.
      ii. It is important to remember that these coefficients do NOT tell us the ratio of grams of each substance. Just because there are two moles of $H_2$ needed to react with each mole of $O_2$ does NOT mean that there are two GRAMS of hydrogen needed for every GRAM of oxygen. That would only be the case of each mole of $O_2$ weighed as much as each mole of $H_2$, and the periodic table shows that this is clearly not the case.
      iii. Thus, in order to use a chemical equation to predict the amounts of substances used in chemical reaction, we must always solve such a problem using moles as our unit of matter. Additional conversion

steps will be required if the problem does not actually supply or ask for the number of moles.

IV. Examples of simple stoichiometry problems: moles to moles.

How many moles of water can be produced from 2.88 moles of $O_2$ and excess $H_2$?

Solution:

$$2.88 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2} = \underline{5.76} \text{ mol } H_2O$$

$$2H_2 + O_2 \longrightarrow 2H_2O$$

How many moles of water can be produced from 2.88 moles of $H_2$ and excess $O_2$?

Solution:

$$2.88 \text{ mol } H_2 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2} = \underline{2.88} \text{ mol } H_2O$$

How many moles of $H_2$ are needed to produce 10.8 moles of water (assuming excess $O_2$)?

Solution:

$$10.8 \text{ mol } H_2O \times \frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O} = \underline{10.8} \text{ mol } H_2$$

V. An example of a simple stoichiometry problem: grams to moles

How many moles of water can be produced from 2.88 grams of $O_2$ and excess $H_2$?

Solution:

$$2.88 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.0 \text{ g } O_2} \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2} = \underline{0.180} \text{ mol } H_2O$$

VI. An example of a simple stoichiometry problem: moles to grams

How many grams of water can be produced from 2.88 moles of $O_2$ and excess $H_2$?

Solution:

$$2.88 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2} \times \frac{18.0 \text{ g } H_2O}{1 \text{ mol } H_2O} = \underline{104} \text{ g } H_2O$$

VII. An example of a not-so-simple stoichiometry problem: grams to grams

How many grams of oxygen are needed to produce 13.0 g of water (assuming excess hydrogen)?

Solution:

$$13.0\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.0\text{ g } H_2O} \times \frac{1\text{ mol } O_2}{2\text{ mol } H_2O} \times \frac{32.0\text{ g } O_2}{1\text{ mol } O_2} = \underline{11.6}\text{ g } O_2$$

VIII. An example of a simple stoichiometry problem: moles to liters of a gas

How many liters of $H_2$ gas (at STP) are needed to completely react with 5.00 mol of $O_2$ gas?

$$5.00\text{ mol } O_2 \times \frac{2\text{ mol } H_2}{1\text{ mol } O_2} \times \frac{22.4\text{ L } H_2}{1\text{ mol } H_2} = \underline{224}\text{ L } H_2$$

IX. An example of a not-so-simple stoichiometry problem: liters of a gas to grams

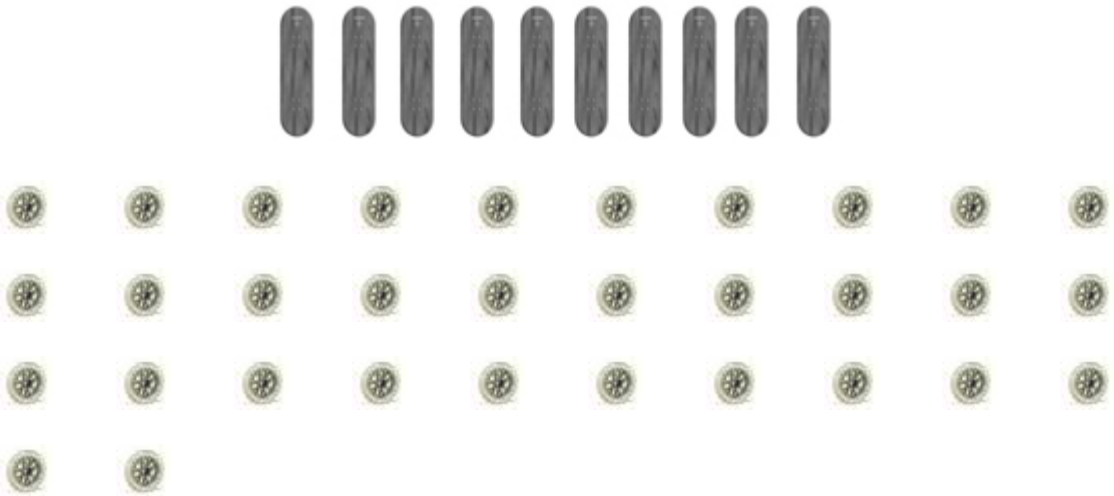
How many liters of $H_2$ gas (at STP) are needed to completely react with 5.00 g of $O_2$ gas?

$$5.00\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{2\text{ mol } H_2}{1\text{ mol } O_2} \times \frac{22.4\text{ L } H_2}{1\text{ mol } H_2} = \underline{7.00}\text{ L } H_2$$

X. Limiting and Excess Reagents.

a. Up until now, we have assumed that there is enough or even too much of all of the reactants in order for the reaction to take place. For instance, in the examples above, we used the chemical reaction between oxygen and hydrogen to produce water. In each case, if the amount of hydrogen was specified, then it was noted that there was more than enough ("excess") oxygen available to react with the hydrogen.

b. But . . . what if there weren't enough of the second chemical to react with all of the first chemical? Imagine this: you are making skateboards for your skateboard company. What happens if you are left with 10 decks and 32 wheels? How many complete skateboards can you make (assume that you have more than enough trucks, screws, etc)?



c. Well, because there are 4 wheels per each skate deck, you could figure this problem out two different ways. Both ways will give you the same answer.

i. Method one: assume that you use all of the skate decks and figure out how many wheels you need in

order to use up all of those decks.

$$10\ decks\ \mathrm{x}\frac{4\ wheels}{1\ deck} = 40\ wheels\ needed$$

Because you do not have 40 wheels available (remember: you only have 32), we call the wheels the **limiting reagent.** On the other hand, you will have a bunch of decks left over after you have used up all (or most) of those wheels. The skate decks are the **excess reagent.** Can you figure out how many unused skate decks will be left over?

ii. Method two: assume that you use all of the wheels and figure out how many skate decks you need in order to use up all of those wheels.

$$32\ wheels\ \mathrm{x}\frac{1\ deck}{4\ wheels} = 8\ decks\ needed$$

So, do we actually have 8 decks available? Yes we do! In fact, we have 10 decks available, which more than is needed. You can see again that the skate decks are available in excess. **The skate decks are still the excess reagent. Because all of the wheels get used up, they are the limiting reagent, just as we calculated using the first method**.

XI. You will be expected to evaluate a problem in which the amounts of more than one reactant are given and determine the limiting reagent. Then, there will still remain for you the additional "normal" problem of solving the stoichiometry problem. To summarize: when you solve a stoichiometry problem, **first you must determine which substance is the limiting reagent.** Many times this part of the problem is solved for you, making the problem a whole lot faster to solve. (After you figure out which substance is the limiting reagent, you should temporarily **ignore the excess reagent** unless you are asked a specific question about it.) **Second, solve the stoichiometry problem as you normally would, using the limiting reagent.**

a. Example problem: Consider the following chemical reaction between lead (II) nitrate and sodium iodide.

$$\mathrm{Pb(NO_3)_2}\ (aq) + \ \mathrm{2NaI}\ (aq) \rightarrow \mathrm{PbI_2}(s) + \ \mathrm{2NaNO_3}\ (aq)$$

If 2.00 mol of $Pb(NO_3)_2$ is reacted with 3.00 mol of NaI, determine

i. Which substance is the limiting reagent?
ii. Which is the excess reagent?
iii. How many moles of the $PbI_2$ product are produced? How many grams?
iv. How many moles of the excess reagent remain after the reaction is complete?

Solution:

(i) $2.00 \text{ mol } Pb(NO_3)_2 \times \frac{2 \text{ mol NaI}}{1 \text{ mol } Pb(NO_3)_2} = \underline{4.00 \text{ mol NaI}}$ needed.

There are only 3.00 mol NaI available, though, so NaI is the limiting reagent.

(ii) Thus, the $Pb(NO_3)_2$ is the excess reagent.

(iii) Use the limiting reagent.

$3.00 \text{ mol NaI} \times \frac{1 \text{ mol } PbI_2}{2 \text{ mol NaI}} = \underline{1.50} \text{ mol } PbI_2$ produced.

$1.50 \text{ mol } PbI_2 \times \frac{461 \text{ g mol } PbI_2}{1 \text{ mol } PbI_2} = \underline{691.5} \text{ g } PbI_2$

From periodic table!

(iv) $1.50 \text{ mol } PbI_2 \times \frac{1 \text{ mol } Pb(NO_3)_2}{1 \text{ mol } PbI_2} = \underline{1.50 \text{ mol } Pb(NO_3)_2}$ used up.

2.00 mol available
−1.50 mol used
0.50 mol left over

XII. Calculating percent yield

a. "Yield": generally, how much of a substance is made. We speak of two types of yields . . .
b. Theoretical yield: how much of a substance *should* be produced in a chemical reaction
c. Actual yield: the amount of a substance that is *actually* produced in a chemical reaction
d. % Yield

i. $\% \, Yield = \frac{actual\ yield}{theoretical\ yield} \times 100$
ii. In other words, to calculate the % yield of a reaction, take the amount that you *did* produce in the reaction and divide by what you *should have* produced in the reaction.
iii. If the reaction proceeds perfectly (which never happens), you will produce the exact amount that is theoretically possible, and you will get a 100% yield. (We had been assuming this "perfection" to always be the case up until now, just to keep things simple.)
iv. A super-duper simple example of % yield.
1. You decide to make some cookies using a cookie mix from the store. The package says that the whole box of cookie mix should produce ("yield") 2 dozen cookies. However, when you make the cookies, you only get 18 cookies (probably because you kept on eating the raw cookie batter while you were waiting for the oven to heat up, you little piggy).
2. Let's calculate your % yield.

$$\% \, Yield = \frac{actual\ yield}{theoretical\ yield} \times 100$$

$$\% \, Yield = \frac{18 \text{ cookies}}{24 \text{ cookies}} \times 100$$

$$\% \, Yield = 0.75 \times 100$$

$$\% \, Yield = 75\%$$

v. Another example of calculating % yield, this time involving a more boring chemical reaction.

Consider the chemical reaction between lead (II) nitrate and sodium iodide:

$$Pb(NO_3)_2\,(aq) + 2NaI\,(aq) \rightarrow PbI_2\,(s) + 2NaNO_3\,(aq)$$

As an engineer, you have found that the % yield of this reaction is 92.0% under the particular conditions of your factory. You need to determine how many mol of $PbI_2$ can be produced from 6.00 moles of $Pb(NO_3)_2$ and excess NaI. Calculate the following: (a) the theoretical yield of the reaction and (b) the actual yield of the reaction.
Solution:

$$6.00 \text{ mol } Pb(NO_3)_2 \times \frac{1 \text{ mol } PbI_2}{1 \text{ mol } Pb(NO_3)_2} = 6.00 \text{ mol } PbI_2$$

theoretical yield.

actual yield = only 92% of theoretical yield.

$$= (0.92)(6.00 \text{ mol}) = 5.52 \text{ mol } PbI_2$$

Or, if you prefer, $\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

$$92\% = \frac{\text{actual yield}}{6.00} \times 100$$

$$0.92 = \frac{\text{actual yield}}{6.00}$$

$$5.52 \text{ moles} = \text{actual yield}$$

vi. Chemical reactions, in reality, are complicated things. Example: consider the reaction that takes place when you burn gasoline in your car's engine.

1. We know that you can get a lot more energy from this reaction if you drive slowly. (Try it: calculate how many miles you can drive on a tank of gas by driving 35 mile per hour; see how this fuel economy changes when you drive 65mph!) The fuel burns more efficiently at low speeds, and less efficiently at high speeds. What is going on?
2. The answer is: competing side reactions and incomplete combustion. In theory, all of your fuel in the gas tank should be converted to $CO_2$ and $H_2O$ once it reacts with $O_2$. In reality, there will be a significant amount of CO produced, which means that you have not completely oxidized the fuel and have therefore not derived the maximum amount of work from that tank of gas. You paid for the gasoline, but you have not made it do all of what it could do (move the pistons and your tires to get you down the street). It is easier for your car's engine to match up fuel with oxygen when the reaction is proceeding more slowly. Also, the faster a car goes, the harder it is for the cart to push through the air around it. (Try this at the beach: see how hard it is to walk through waist-deep water. Now try to *run* through waist-deep water. Which one do you think burns more energy?)
3. If your car isn't tuned up very well, there could even be some unburned fuel making it through your exhaust system. Modern cars have a lot of computer circuitry that is used to continuously monitor the engine's performance. Older cars used a relatively inefficient device to mix the fuel

and oxygen called a carburetor; newer cars use electron fuel injection.

4. By the way, in case you really care about saving money on fuel, it's pretty easy to do.
    a. First, inflate your tires to the correct pressure. This is easy and usually free.
    b. Buy the right type of fuel for your car. My car is designed to take the cheapest type of fuel ("87" grade), and the more expensive stuff won't make it run any faster.
    c. Friction is the enemy. Change your oil and oil filter every 3,000 miles so that your engine parts sli-i-i-de past each other easily. This might cost you 20 or 30 bucks every 3 months.
    d. Let the air in: change your air filter. This might cost your 40 or 50 bucks every 6 or 9 months.
    e. Get your car serviced regularly. The most expensive maintenance that you have done on your car, while important to the safety of your vehicle, is not necessarily the stuff that affects gas mileage.

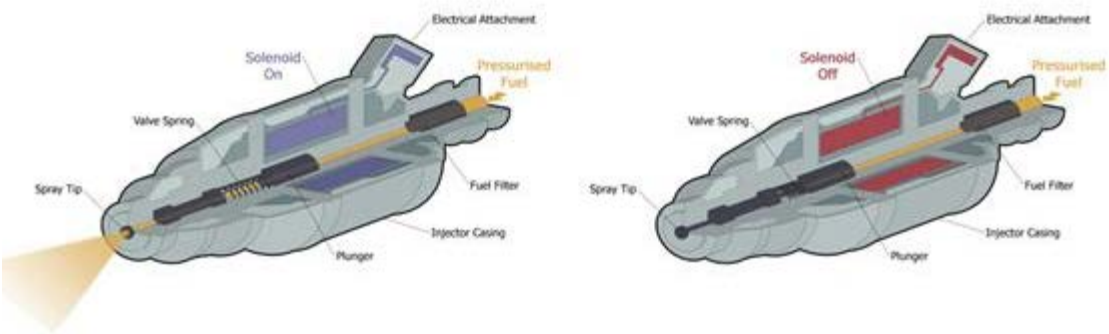



**Figure 1: A diagram of an automobile's electronic fuel injector that I copied from Wikipedia. This is the thing that mixes oxygen and fuel as efficiently as possible when you are driving. That is, it tries to make sure that you get the highest % yield from the combustion reaction in your car by trying to match every gasoline molecule with the perfect number of oxygen molecules. It works by using sensors to monitor what is going into the engine and what is coming out of the engine. The computer uses this information from the sensors to adjust the air-fuel mixture.**