

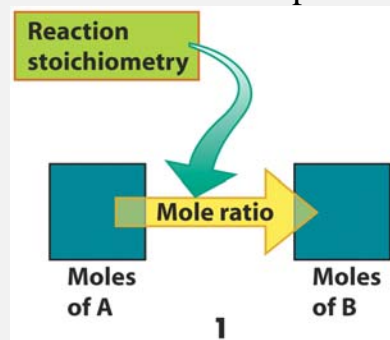
# HIGH SCHOOL CHEMISTRY REVIEW

## LECTURE 2: REACTION STOICHIOMETRY

*Chapter summary.* We just learned that simple quantitative relationships based upon the idea of the law of simple proportions could be combined with other concepts from Dalton's Atomic Theory to create a host of problems based upon the quantitative relationships between atoms in molecules. We learned to use unit factors to convert between

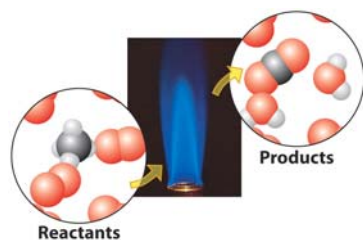
grams and moles  
moles and # of atom or molecules  
atoms<sub>A</sub> and molecule<sub>A</sub>

Now we take things just a single step further by applying these concepts of stoichiometry to CHEMICAL REACTIONS. After all, since we know there is conservation of mass in the world, then when a chemical change occurs and new compounds are formed by chemical reaction, there must be a collection of simple unit factors that describe these changes. The additional unit factor is simply asking for a stoichiometric relationship between numbers of atoms or molecules of one chemical species and another chemical species involved in a reaction.



Example. Consider the combustion of methane, CH<sub>4</sub>:

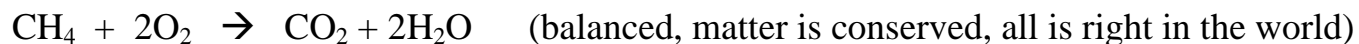
methane + oxygen → water and carbon dioxide



Writing it without concern for stoichiometry we have



But by inspection we can find a way for conserve matter in a balanced reaction by placing a coefficient of 2 in front of the diatomic oxygen and in front of the water:



From this balanced equation we can form all kinds of units factors relating molecules or atoms of reactants to molecules or atoms of products.

For example

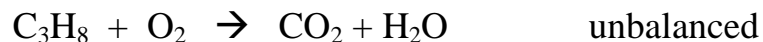
there is:            one mole of CH<sub>4</sub> for every one mole of H<sub>2</sub>O  
there are:            2 molecules of O<sub>2</sub> for every one molecule of CO<sub>2</sub>  
there are even:      four atoms of **H** in CH<sub>4</sub> for every two atoms of **O** in 2H<sub>2</sub>O

and many more, many more unit factors. This means that we can do all the stoichiometry problems from compositional stoichiometry with the added twist that we can add unit factors for reactions.

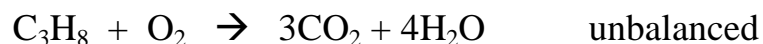
### Balancing Chemical Equations

One annoying feature of using chemical reactions is that sometimes the equations you are given are not balanced (i.e., they don't satisfy the law of conservation of mass). For example the first methane combustion equation above was not balanced. Well you can't use unit factors if don't have balanced equations, so we need some practice with balancing.

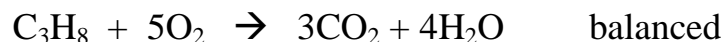
So let's try practicing with another reaction. This time propane combustion.



Coefficients are 1, 1, 1, 1 = 4 total molecules in unbalanced reaction.



Coefficients are 1, 1, 3, 4 = 9 total molecules in unbalanced reaction.



Coefficients are 1, 5, 3, 4 = 13 total molecules in a balanced reaction.

How did he do that? A few hints for balancing equations (these don't work for complicated redox problems, but they are a way to start with simpler equations.

1. First balance the compounds which contain elements that appear only once on each side of the equation. For example, in the propane combustion reaction above, note that C only appears once on the left (in  $\text{C}_3\text{H}_8$ ) and once on the right (in  $\text{CO}_2$ ) above. **BALANCE IT FIRST.** Note that oxygen atoms appear in two compounds on the right ( $\text{CO}_2$  and  $\text{H}_2\text{O}$ ), do it later.

2. Save the elements or atoms with a single element to balance last (He, H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, P<sub>4</sub>, S<sub>8</sub>,...). For example, solve for O<sub>2</sub> last in the propane problem above.

If you follow these two rules for balancing equations, solving every equation will be a snap---unless the rules don't apply. Then you are on your own and will wish you'd practiced more.

Look, everyone can balance equations. It is just a matter of how much time it takes: practice a lot and solve your exams equation in 30 seconds, practice while taking the exam, and do the problem in five minutes. It is up to you.

A little practice for right now. Try to follow the rules:

**Example 1.** Balance



Start with Na that appears only once on either side and end with Cl<sub>2</sub> or H<sub>2</sub>, which are elements:

Step one, Na are already balanced:



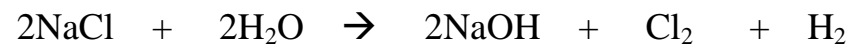
Step two, two Cl on right means need to add two on left:



Step 3, oops, now Na are not balanced, need to add two to right side



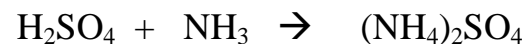
Step 4, now balance the O:



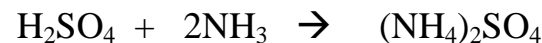
Step 5, finally, balance the H--oh, it already is:



**Example 2.** Balance:

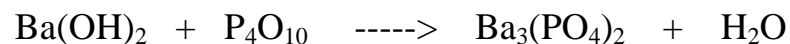


By inspection, add a 2 in front of the ammonia

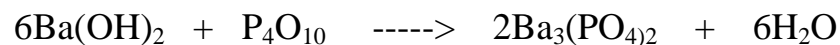


That was easy.

**Example 2.** Balance by yourself.



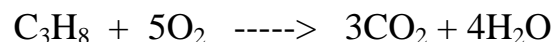
That wasn't so easy, was it? Here is the answer.



Okay, so the rules don't always work. Actually balancing chemical reactions can get very difficult when oxidation numbers change during a reaction. We'll explore more systematic ways to balance chemical reactions toward the end of CH302 when we start learning about electrochemistry. For now, assume that simple "inspection" will do the trick most of the time.

### Reaction Stoichiometry Calculations

Assuming that you have a balanced equation, you can do a whole host of problems that involve converting between grams of this and atoms of that which are conceptually identical to compositional stoichiometry problems. For example, consider how many units factors there are for

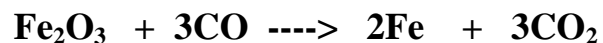


There are a bunch, like:

- 1 molecule of propane for every 3 molecules of CO<sub>2</sub>
- 3 atoms of C in propane for 4 molecules of water
- 9 atoms of hydrogen in propane for 8 atoms of hydrogen in water

and on and on? How many all together? Would you believe hundreds?

Now on to some problems. For the following problems, use the reaction



**Problem 1.** How many **CO** molecules are required to react with 25 formula units of **Fe<sub>2</sub>O<sub>3</sub>**?

$$\begin{aligned} ? \text{ CO molecules} &= (25 \text{ Fe}_2\text{O}_3 \text{ form. unit}) \left( \frac{3 \text{ CO molecules}}{1 \text{ Fe}_2\text{O}_3 \text{ form. unit}} \right) \\ &= 75 \text{ CO molecules} \end{aligned}$$

**Problem 2.** How many **Fe** atoms are produced by reaction of  $2.5 \times 10^5$  formula units of **Fe<sub>2</sub>O<sub>3</sub>** with excess **CO**?

$$\begin{aligned} ? \text{ Fe atoms} &= (2.5 \times 10^5 \text{ Fe}_2\text{O}_3) \left( \frac{2 \text{ Fe atoms}}{1 \text{ Fe}_2\text{O}_3} \right) \\ &= 5 \times 10^5 \text{ Fe atoms} \end{aligned}$$

**Problem 3.** What mass of **CO** is required to react with 146 grams of **Fe<sub>2</sub>O<sub>3</sub>**?

$$\begin{aligned} ? \text{ g CO} &= (146 \text{ g Fe}_2\text{O}_3) \left( \frac{1 \text{ mole Fe}_2\text{O}_3}{159.69 \text{ g}} \right) \left( \frac{3 \text{ mole CO}}{1 \text{ mole Fe}_2\text{O}_3} \right) \left( \frac{28 \text{ g CO}}{1 \text{ mole CO}} \right) \\ &= 76.7 \text{ g CO} \end{aligned}$$

**Problem 4.** What mass of **CO<sub>2</sub>** can be produced by the reaction of 0.540 moles of **Fe<sub>2</sub>O<sub>3</sub>** with excess **CO**?

$$\begin{aligned}
 ? \text{ g CO}_2 &= (0.54 \text{ mole Fe}_2\text{O}_3) \left( \frac{3 \text{ mole CO}_2}{1 \text{ mole Fe}_2\text{O}_3} \right) \left( \frac{44 \text{ g CO}_2}{1 \text{ mole CO}_2} \right) \\
 &= 71.3 \text{ g CO}_2
 \end{aligned}$$

**Problems 5.** What mass of **Fe<sub>2</sub>O<sub>3</sub>** reacts with excess **CO** if the reaction produces 8.65 grams of **CO<sub>2</sub>**.

$$\begin{aligned}
 ? \text{ g Fe}_2\text{O}_3 &= (8.65 \text{ g CO}_2) \left( \frac{1 \text{ mole CO}_2}{44 \text{ g CO}_2} \right) \left( \frac{1 \text{ mole Fe}_2\text{O}_3}{3 \text{ mole CO}_2} \right) \left( \frac{159.69 \text{ g Fe}_2\text{O}_3}{1 \text{ mole Fe}_2\text{O}_3} \right) \\
 &= 10.6 \text{ g Fe}_2\text{O}_3
 \end{aligned}$$

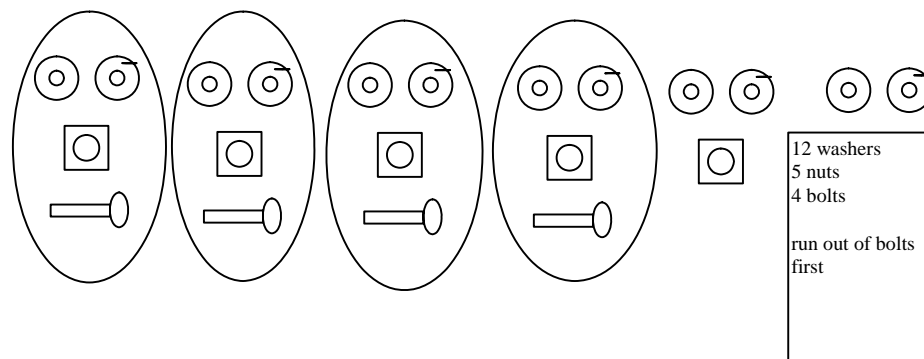
And now some specialty problems involving reaction stoichiometry.

### Limiting Reagents

You'll notice that in several problems I referred to "excess of this reagent", or "left over that reagent". In every reaction we will either have a situation in which a **STOICHIOMETRIC** amount of reagents react (for example, **EXACTLY** three times as many **CO** as **Fe<sub>2</sub>O<sub>3</sub>**) or we will have one of the reagents left over. Practically speaking, we can never hope to use such accurate measures, so we will always have **ONE REAGENT WE RUN OUT OF FIRST IN A CHEMICAL**

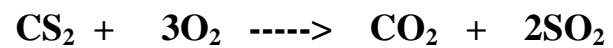
**REACTION.** This is known as the **LIMITING REAGENT**.

Example: I want to assemble a gadget that requires one nut, one bolt and two washers for every hole. I have in my garage a bucket filled with 12 washers, 4 bolts and five nuts. What is the **LIMITING SMALL METAL OBJECT**?



Of course, rather than small metal objects, I can use chemical reagents.

**Problem 6.** What is the maximum mass of  $\text{SO}_2$  that can be produced by the reaction of 95.6 grams of  $\text{CS}_2$  with 110 grams of  $\text{O}_2$ ?





First, find limiting reagent

$$? \text{ moles CS}_2 = 95.6 \text{ g CS}_2 \left( \frac{1 \text{ mole CS}_2}{76 \text{ g CS}_2} \right) = 1.26 \text{ moles CS}_2$$

$$? \text{ moles O}_2 = 110 \text{ g O}_2 \left( \frac{1 \text{ mole O}_2}{32 \text{ g O}_2} \right) = 3.44 \text{ moles O}_2$$

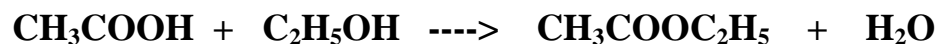
But we need 3 moles O<sub>2</sub> for every mole CS<sub>2</sub>. So if we have 1.26 mole CS<sub>2</sub>, then we need 3 X 1.26 = 3.77 moles O<sub>2</sub>. We have not quite enough O<sub>2</sub>, so it is the limiting reagent.

$$? \text{ g SO}_2 = (3.44 \text{ moles O}_2) \left( \frac{2 \text{ mole SO}_2}{3 \text{ moles O}_2} \right) \left( \frac{64 \text{ g SO}_2}{1 \text{ mole SO}_2} \right) = 146.8 \text{ g SO}_2$$

### Percent Yield in a Reaction

Another concept of interest is the PERCENT YIELD. Ideally, if you react every single molecule in one mole of CS<sub>2</sub> in the above reaction, you can make 2 moles of SO<sub>2</sub>. This would correspond to a 100% yield of product. Sadly, life doesn't work that way. If instead we ended up with only one mole of SO<sub>2</sub>, we would say that we had achieved only a 50% yield. When you get to organic chemistry lab, you will learn that your chances for medical school might hang in the balance on how high a percent yield you can get from your starting material.

**Problem 7.** A 10.0 gram sample of ethanol, C<sub>2</sub>H<sub>5</sub>OH, was boiled with acetic acid, CH<sub>3</sub>COOH, to produce a 14.8 gram yield of ethyl acetate, CH<sub>3</sub>COOC<sub>2</sub>H<sub>5</sub>. What is the percent yield?



$$? \text{ moles C}_2\text{H}_5\text{OH} = 10.0 \text{ g} \left( \frac{1 \text{ mole}}{46 \text{ g C}_2\text{H}_5\text{OH}} \right) = 0.22 \text{ moles of reagent}$$

$$? \text{ moles CH}_3\text{COOC}_2\text{H}_5 = 14.8 \text{ g} \left( \frac{1 \text{ mole}}{88 \text{ g CH}_3\text{COOC}_2\text{H}_5} \right) = 0.17 \text{ moles of product}$$

formed in 1:1 mole ratio so

$$\% \text{ yield} = \frac{0.17 \text{ mole product}}{0.22 \text{ mole reagent}} = 78\%$$

### **Solution Stoichiometry—because most interesting stuff happens in water.**

What follows is a discussion of how to determine the concentration of a solution and how to determine the concentration after the reaction of a solution. How important is this material? Well most chemical reactions of use to human beings occur in solution (specifically water) where the reactant molecules of interest can scurry around in the solvent looking for other reagents. This makes for a heck of a lot more exciting reaction chemistry than if you placed two rocks (solids) on top of each other and waited for the reaction to occur. So learning about compositional stoichiometry and reaction stoichiometry in solutions is a very important thing.

A couple of definitions before we start: **SOLUTE** is the compound of interest that you put into a solution and **SOLVENT** is the non-reactive excess liquid to which you add the solute.

So if you dump salt in water, you are dumping a **SOLUTE** into the **SOLVENT**.

#### **Solution Calculations**

Once you add salt to the water, you might want to know how salty the water was. Is it **REALLY SALTY**, or **REALLY, REALLY SALTY**? Clearly we need some **QUANTITATIVE** measure of solute concentration. There are many ways to do this. I give you the two most common.

#### **Percent by mass**

This measure of concentration is used because it follows simply from doing a measurement on a balance.

Example:

add 10 grams of salt to 90 grams of water, then

the solution is 10% by weight, salt, and 90 % by weight water.

(Remember we found out the Coca-Cola is a 10% by weight sugar in water.

**Problem 8.** What mass of NaOH is required to prepare a 250 gram solution that is 8% NaOH?

$$8\% \text{ NaOH} = \frac{8 \text{ g NaOH}}{92 \text{ g H}_2\text{O}} \text{ or } \frac{8 \text{ g NaOH}}{100 \text{ g solution}}$$

$$\begin{aligned} ? \text{ mass NaOH} &= 250 \text{ g soln} \left( \frac{8 \text{ g NaOH}}{100 \text{ g soln}} \right) \\ &= 20 \text{ g NaOH} \end{aligned}$$

Of course, most often we weigh the solute and we measure the solvent (a liquid) by volume, for example, using a graduated cylinder. Thus, we have to use DENSITY to convert from volume to weight.

**Problem 9.** Calculate the mass of NaOH in 300 ml of solution that is 8% NaOH if the density of solution is 1.09 g/ml.

$$? \text{ g NaOH} = 300 \text{ ml soln} \left( \frac{1.09 \text{ g}}{1 \text{ mL}} \right) \left( \frac{8 \text{ g NaOH}}{100 \text{ g soln}} \right) = 26 \text{ g NaOH}$$

## Molarity

By far the most common method for expressing concentration is not on a mass basis but on a MOLE basis. Thus the MOLARITY is defined as

$$\frac{\text{moles}}{\text{liter}} = \frac{\text{grams / molecular weight}}{\text{liter}}$$

If you stay in science, you will make about 200,000 different molarity solutions of things, so you might as well learn how to do it now.

**Problem 10.** Calculate the molarity of a solution that contains 12.5 grams of H<sub>2</sub>SO<sub>4</sub> in 1.75 liters of solution.

$$\text{Molarity H}_2\text{SO}_4 \left( \frac{\text{moles}}{\text{Liter}} \right) = \frac{\frac{12.5 \text{ g}}{98 \text{ g / mole}}}{1.75 \text{ L}} = 0.07 \text{ M}$$

**Problem 11.** Calculate the mass of Ca(NO<sub>3</sub>)<sub>2</sub> required to prepare 3.5 liters of a 0.8 M solution.

$$? \text{ g Ca(NO}_3)_2 = \left( \frac{0.8 \text{ mole}}{1 \text{ L}} \right) (3.5 \text{ L}) \left( \frac{128 \text{ g}}{\text{mole}} \right) = 358 \text{ g}$$

**TIME OUT** For SOLUTION REACTION EQUATIONS Here is something important to know. For the rest of the chapter as we learn about neutralizations, or dilutions, or titrations, we will use the following startling facts:

$$\text{Molarity (M)} = \frac{\text{moles}}{\text{Volume (V)}}$$

So on rearranging the equation we have

$$\text{moles} = (\text{Molarity})(\text{Volume}) = MV$$

But you know that stoichiometry tells us that

moles **a** are related in simple proportions to moles **b**

So we can make all kinds of useful equations:

$$\text{moles } \mathbf{a} = M_b V_b \text{ or } M_a V_a = \text{moles } \mathbf{b}$$

or even

$$M_a V_a = M_b V_b$$

But there is more. Since

$$\text{moles} = \text{grams/molecular weight (MW)}$$

then we have a bunch of other possible equations:

$$\text{moles } \mathbf{a} = \text{grams } \mathbf{a}/\text{MW}_a = M_b V_b = \text{moles } \mathbf{b}$$

or any other similar combination.

These are the simple equations you use to do DILUTION problems and TITRATION problems and SOLUTION REACTIONS.

### A Dilution Calculation

**Problems 12.** . What volume of 18.0 M H<sub>2</sub>SO<sub>4</sub> is required to prepare 2.5 liters of 2.4 M H<sub>2</sub>SO<sub>4</sub>.

moles H<sub>2</sub>SO<sub>4</sub> before = moles H<sub>2</sub>SO<sub>4</sub> after

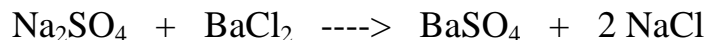
$$M_b V_b = M_a V_a$$

$$(18 \text{ M})(? V_b) = (2.4 \text{ M})(2.5 \text{ L})$$

$$V_b = 0.33 \text{ L}$$

### A Chemical Reaction in Solution.

**Problem 13.** What volume of 0.5 M BaCl<sub>2</sub> solution is required to complete a reaction with 4.32 grams of Na<sub>2</sub>SO<sub>4</sub>.



moles BaCl<sub>2</sub> = moles Na<sub>2</sub>SO<sub>4</sub>

$$M_a V_a = \frac{g_b}{MW_b}$$

$$(0.5 \text{ M})(? \text{ L BaCl}_2) = \frac{4.32 \text{ g}}{142 \text{ g/mole}}$$

$$V_{\text{BaCl}_2} = 0.61 \text{ mL} = 61 \text{ mL}$$

### Titration

If you have had a chemistry class, you probably have done a titration. A titration is nothing more than using:

- The known concentration of something
- The stoichiometry of a reaction
- One of the titration equations like  $M_1 V_1 = M_2 V_2$

to find out **the concentration of an unknown solution.**

Let's see how it works.

**Problem 14.** What is the molarity of a KOH solution if 38.7 mL of KOH is titrated with 43.2 mL of a 0.223 M HCl solution?

$$\text{moles}_b = \text{moles}_a$$

$$M_b V_b = M_a V_a$$

$$(? M_b)(38.7 \text{ mL}) = (0.223 \text{ M HCl})(43.2 \text{ mL HCl})$$

$$M_{\text{KOH}} = 0.25 \text{ M}$$

