Unit 6: Reactions & Stoichiometry

You may write notes on one side of a 3x5 card. You'd better bring a calculator for this test!!

I. Balancing!
   A. The goal is to have the same number of ATOMS of each element on both sides of the arrow.
      $3\ Ba(NO_3)_2 + 2\ AlPO_4$ ----> $1\ Ba_3(PO_4)_2 + 2\ Al(NO_3)_3$
      1. *Coefficients* are the normal sized numbers that you used to balance the reactions.
      2. NEVER change the *subscripts* when balancing a reaction—that changes the identity of the compounds. You're not allowed to do that.
      3. The arrow is chemistry's version of an equal sign
         a. If it isn't actually equal, something went wrong back at the ranch
   B. Chemicals on the left of the arrow are "reactants"
      1. That's because they react.
   C. Chemicals on the right of the arrow are "products"
      1. That's because they are produced

II. Stoichiometry
   A. This is a really fancy word that means "mole-to-mole relationships"
   B. The relationship is defined by the coefficients of a balanced equation.
      $3\ Ba(NO_3)_2 + 2\ AlPO_4$ ----> $1\ Ba_3(PO_4)_2 + 2\ Al(NO_3)_3$
      1. Given the equation above, 3 moles of $Ba(NO_3)_2$ can produce 1 mole of $Ba_3(PO_4)_2$.
         a. So 9 moles of $Ba(NO_3)_2$ can produce 3 mole of $Ba_3(PO_4)_2$
         b. Or 7.1 moles of $Ba(NO_3)_2$ can produce 2.37 mole of $Ba_3(PO_4)_2$
      2. The ratios in a balanced equation always hold true for numbers of *moles*.
      3. To calculate a ratio between any two chemicals in a balanced equation . . .
         a. Basically you multiply by one chunk of the equation and divide by another so that your units change to the ones you want.
         b. Magic formula:

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 # moles given | (compound wanted)  = # moles of the compound wanted
                | (compound given)
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         a. So for the question: How much $Al(NO_3)_3$ is produced if 35.4 moles of $Ba(NO_3)_2$ react?
            (1) # moles given = 35.4 moles $Ba(NO_3)_2$
            (2) The "given" was $Ba(NO_3)_2$, so I will put the part of the reaction with that formula on the bottom of my conversion ($3\ Ba(NO_3)_2$) so that the barium nitrates cancel out.
            (3) The "wanted" is $Al(NO_3)_3$, so I will put the part of the reaction with that formula on the top of my conversion ($2\ Al(NO_3)_3$). This will make my answer come out in moles of aluminum nitrate.
         a. The math looks like this:

$$\frac{35.4 \text{ moles } Ba(NO_3)_2 \quad | \quad 2\ Al(NO_3)_3}{| \quad 3\ Ba(NO_3)_2} = (35.4 \times 2) / 3 = 23.6 \text{ mol } Al(NO_3)_3$$

   C. It is called a MOLE RATIO because it has to be in MOLES
      1. If you're given (or asked for) some unit that isn't moles, you have to convert!
         a. We've learned several things we can convert to moles. (grams, liters of solution, liters of gas at STP, particles…)
         b. Let's review those, shall we?
      2. Mass (gram) conversions!
         a. Grams can be changed to moles (and back) using molar mass (formula mass):
         b. Molar mass is the number of grams one mole of a substance weighs.
            (1) It is always in units of grams/mole
            (2) Sometimes this is called:
               (a) Formula Mass
               (b) Formula Weight

(c) Molecular Mass
(d) Molecular Weight
(e) Molar Mass
(f) Atomic Weight

(3) They all mean pretty much the same thing.

c. The molar mass of a compound is found by adding the atomic masses of the elements in the compound

(1) The atomic mass is the number at the bottom of the square, on the periodic table on the wall.
(a) Example: for Helium (He), it is 4.00260
(b) If you just use 4.00, you'll be okay

(2) Make sure you round correctly
(a) For Sodium (Na) the atomic mass is 22.98977
(b) That rounds to 22.99, **not** 22!!

(3) H_2O
(a) Contains:
i) 2 H's
a) H has a mass of 1.01 g/mol
(b) 1 O
a) O has a mass of 16.00 g/mol
(c) 2(1.01 g/mol) + 16.00 g/mol = 18.02 g/mol
(d) That means 18.02 grams of water = 1 mol of H_2O

(4) Try this one: $Al(ClO_4)_3$ (aluminum perchlorate)
(a) Contains:
i) ____ Al
a) mass:
ii) ____ Cl
a) mass:
iii) ____ O
a) mass:
(b) ____(_____) + ____(_____) + ____(_____) =

d. If you are changing from grams to moles, the formula mass (molar mass) is your conversion factor.

(1) The molar mass should be below the line (multiply by 1 mol and divide by the molar mass)

(2) 100 grams of H_2O = ? moles of H_2O
(a) The first thing I need to do is find the molar mass of H_2O (18.02 g/mol).
(b) Now I can use 18.02 g H_2O = 1 mol H_2O to convert to mol H_2O

$$\frac{100 \text{ g } H_2O \mid 1 \text{ mol } H_2O}{\mid 18.02 \text{ g } H_2O} = (100 \times 1) / 18.02 = 5.5 \text{ mol } H_2O$$

e. If you are changing from moles to grams

(1) The molar mass should be above the line (multiply by the molar mass and divide by 1 mol)

(2) 2.5 moles of H_2O = ? grams of H_2O
(a) The first thing I need to do is find the molar mass of H_2O (18.02 g/mol).
(b) Now I can use 18.02 g H_2O = 1 mol H_2O to convert mol H_2O to grams H_2O

$$\frac{2.5 \text{ mol } H_2O \mid 18.02 \text{ g } H_2O}{\mid 1 \text{ mol } H_2O} = (2.5 \times 18.02) / 1 = 45.05 \text{ g } H_2O$$

3. Particle conversions!

a. Avogadro's Number = 6.02×10^{23}

(1) This is the number of particles in a mole of a substance

(2) In chemistry, Avogadro's number is always used to count ***particles.***

(3) *Particles* can mean:
(a) atoms
(b) molecules
(c) formula units

b. $6.02x10^{23}$ particles = 1 mol
   (1) This can be used to convert from particles to moles:
      (a) If you have $2.71x10^{24}$ atoms, how many moles do you have?
      (b) $\frac{2.71x10^{24} \text{ atoms}}{6.022x10^{23} \text{ particles}} \cdot \frac{1 \text{ mol}}{} = (2.71x10^{24} \text{ x } 1)/(6.022x10^{23}) = 4.50 \text{ mol}$
      (c) The units are okay because "atoms" and "particles" cancel one another out—they are secretly the same thing.
   (2) It can be used to convert from moles to particles:
      (a) If you have 2.35 moles of H_2O, how many water molecules are there?
      (b) $\frac{2.35 \text{ mol } H_2O}{1 \text{ mol}} \cdot \frac{6.02x10^{23} \text{ particles}}{} = (2.35 \text{ x } 6.02x10^{23})/1 = 1.42x10^{24}$ molecules
   (3) The units are okay because "molecules" and "particles" are secretly the same thing.
c. So basically:
   (1) Particles = mol x Avogadro;
   (2) Moles = particles/Avogadro;

4. Concentration in Molarity (M)
   a. Usually, you don't use dry chemical for a reaction. You dissolve it in water first (because that gives it wayyy more surface area, and then it can react faster).
      (1) If something is dissolved in water, we use the concentration of the solution to solve for moles
      (2) We like to measure concentration in molarity (M)
         (a) Molarity = $\text{moles}_{\text{solute}}/\text{Litres}_{\text{solution}}$
         (b) A molarity is the conversion factor between moles and liters
            i) 3.0 M means 3.0 moles = 1 L
            ii) 0.25 M means 0.25 moles = 1L
      (3) Make sure the volume is in liters.
         (a) mL/1000 = L
         (b) Or move the decimal 3 places.
            i) 25 mL = 0.025 L
            ii) 130 mL = 0.130 L
   b. So basically:
      (1) Moles = Liters x Molarity
      (2) Liters = moles/Molarity
      (3) (If you are asked "what volume" of something, that means solve for LITERS.)

5. Concentration in molality (*m*)
   a. Again, we don't generally use dry chemical for reactions, because chemical dissolved in solution reacts much more quickly. Molality is a MASS-based concentration system.
      (1) Concentration in molality (*m*)
         (a) Molality = $\text{moles}_{\text{solute}}/\text{kilogram}_{\text{solvent}}$
            i) Per kg solvent!!
            ii) Solvent is the liquid we dissolved the solute in.
         (b) A molality is the conversion factor between moles and kg of solvent
            i) 3.0 *m* means 3.0 moles = 1 kg solvent
            ii) 0.25 *m* means 0.25 moles = 1 kg solvent
      (2) Make sure the solvent mass is in kilograms.
         (a) g/1000 = kg
         (b) Or move the decimal 3 places.
            i) 25 g = 0.025 kg
            ii) 130 g = 0.130 kg
   b. So basically:
      (1) Moles = kg x molality
      (2) Kilograms of solvent = moles/molality
      (3) (If you are asked "what mass of solvent" of something, that means solve for KILOGRAMS.)

6. Molarity vs. Molality:
   a. Here's a hint for keeping your concentrations straight:

(1) Molarity is a CAPITAL M.
   (a) It is based on liters, which have the symbol CAPITAL L
(2) Molality is a lowercase *m*
   (a) It is based on kilograms, which have the symbol lowercase kg

D. The **real stoichiometry** (two and three-step problems like we did on recent worksheets & Skyward assignments) comes when you have to COMBINE a mole-ratio conversion with one (or more) of the mole conversions above. For example:

$Al_2(SO_4)_3$ + $Ba(OH)_2$ ---> $BaSO_4$ + $Al(OH)_3$

1. 10 grams of $Ba(OH)_2$ react. How many grams of $Al_2(SO_4)_3$ react?
   a. What was I given? 10 grams of $Ba(OH)_2$
   b. What am I solving for? grams of $Al_2(SO_4)_3$
   c. The two formulas don't match: $Ba(OH)_2 \neq Al_2(SO_4)_3$ That means I will do a mole-to-mole ratio to change compounds.
      (1) You know you have 10 **g** of $Ba(OH)_2$
      (2) You have to change this into moles in order to use a mole ratio (Mole ratios don't work on masses. Only on moles. *I can't do a mole ratio until I'm in moles*.) I can solve for moles using a *molar mass* conversion:
         (a) Grams on the bottom (so they drop out), 1 mol on top
         (b) You have to find the molar mass of $Ba(OH)_2$ to do this
            i) 1(137.3) + 2(16.0) + 2(1.0) = 171.3 g/mol

$$\frac{10\text{ g Ba(OH)}_2}{} \left| \frac{1\text{ mol}}{171.3\text{ g}} \right. = 0.058 \text{ mol Ba(OH)}_2$$

      (3) Now that I have moles, I can do my *mole* ratio. Since a mole ratio comes from a balanced chemical reaction, I'm going to need to balance the reaction I was given before I try to change $Ba(OH)_2$ to $Al_2(SO_4)_3$
         (a) So I balance it.
         (b) $1Al_2(SO_4)_3 + 3Ba(OH)_2$ ----> $2Al(OH)_3 + 3BaSO_4$
      (4) Now I can use the mole ratio

$$\frac{0.058\text{ moles Ba(OH)}_2}{} \left| \frac{1\text{ Al}_2\text{(SO}_4\text{)}_3}{3\text{ Ba(OH)}_2} \right. = 0.019 \text{ mol Al}_2(SO_4)_3$$

      (5) The question asked for ***grams*** of $Al_2(SO_4)_3$, so 0.019 isn't my final answer (because it is in moles, not grams) so now I need to do a mol-to-gram conversion
         (a) First I need the molar mass of $Al_2(SO_4)_3$
            i) 2(27.0) + 3(32.1) + 12(16.0) = 342.3 g/mol
         (b) Now I can convert to grams

$$\frac{0.019\text{ mol Al}_2\text{(SO}_4\text{)}_3}{} \left| \frac{342.3\text{ g}}{1\text{ mol}} \right. = 6.50 \text{ g Al}_2(SO_4)_3$$

         (c) I can actually do all three steps on one line if I want to:

$$\frac{10\text{ g Ba(OH)}_2}{} \left| \frac{1\text{ mol}}{171.3\text{ g}} \right| \frac{1\text{ Al}_2\text{(SO}_4\text{)}_3}{3\text{ Ba(OH)}_2} \left| \frac{342.3\text{ g}}{1\text{ mol}} \right. = 6.50\text{ g Al}_2\text{(SO}_4\text{)}_3$$

         (d)
2. Try this yourself! Follow the steps above for the following:

3. How many grams of HNO_3 react with 15 grams of $CaCO_3$?
   $CaCO_3 + HNO_3$ ----> $H_2CO_3 + Ca(NO_3)_2$ *(This reaction is not yet balanced.)*

E. In summary:
1. Stoichiometry problems are based on the MOLE RATIO (which comes from a balanced chemical equation) because that is how we relate two different chemical compounds.

2. But you can't do a mole ratio unless you have MOLES.
3. So these problems are frequently **three steps**:
   a. Change some silly unit into moles *(If the question GIVES you moles to start with, you get to skip this step.)*
   b. Do a mole ratio to change the given compound into the compound you're solving for.
   c. Change moles into some silly non-mole unit. *(If the question ASKS you for moles in the answer, you get to skip this step.)*

III. Yield

A. "Yield" refers to the products of a reaction.
1. This is why recipes sometimes say "Yield: 12 servings"—the food is the *product* of the recipe. (Cooking, especially baking, is based on chemical reactions that work in ratios, just like stoichiometry.)
2. "Theoretical yield" specifically means "the maximum amount of product that can be produced."
   a. Generally, this is referring to one specific product, not all the products at once.
   b. Theoretical yield is the amount of product that stoichiometric math predicts.
      (1) In every stoichiometry problem in which you solved for a product, your answer was its "theoretical yield."
      (2) For example: "In the equation $2Na + 2H_2O \rightarrow 2NaOH + H_2$, how many grams of sodium hydroxide are produced from 3.0 mol of sodium with an excess of water?"
         (a) The answer is a PRODUCT (NaOH and H_2 are the products of this reaction), so this question is asking for a theoretical yield.
         (b) It could have been written this way: "In the equation $2Na + 2H_2O \rightarrow 2NaOH + H_2$, **what is the theoretical yield of sodium hydroxide** when 3.0 mol of sodium react with an excess of water?"
         (c) (Theoretical yield is always given in grams unless otherwise specified in the problem.)
         (d) Theoretical yield is the maximum amount of product possible, because it is based on a mathematical calculation that assumes EVERY particle reacts as expected and is collected successfully.
3. "Actual yield" is the amount of product collected in a real-life situation.
   a. It is almost impossible for actual yield to equal theoretical yield, because real life isn't perfect:
      (1) Some reactant might have been spilled, or not transferred to the reaction vessel.
      (2) Some product might not have been recovered from the reaction vessel.
      (3) Some reagent might have been lost if the reaction bubbled up over the edges of the container.
      (4) Some reactants might not have reacted. Not all substances react easily or immediately. Sometimes they just sit there.
      (5) Sometimes products undergo a *second* reaction, and turn into something else. (Particularly if the reaction is being heated—if the product starts to burn, a second reaction is taking place.)
   b. Actual yield *should* always be less than or equal to theoretical yield—because the law of conservation of matter means you can't create extra product molecules out of thin air.
      (1) Occasionally the *measured* yield IS larger than the theoretical yield.
         (a) This does NOT mean that the Law of Conservation of Matter is broken.
         (b) Atoms are NOT getting created out of thin air.
      (2) If the yield you measure is larger than the theoretical yield, it means that some part of what you measured is not product you thought it was.
         (a) If the product was meant to be dehydrated, it may still contain some water molecules—which are not the product you were trying to weight.
         (b) If the product has burned at all, it now contains extra oxygen—the burning process adds oxygen to substances.
         (c) If the product was supposed to be separated from another product (e.g., by precipitation and filtration) it is possible that not all of the other product was actually removed—which means part of the weight you measure is that other substance.
      (3) We still call these "actual yield," but when the actual yield is greater than the theoretical yield, it is always the case that either:
         (a) The actual yield we've measured contains something in addition to the product we intended.

(b) -OR-
(c) Bad math has occurred.

B. Percent Yield
1. Percent Yield is a way to calculate the "success" of a reaction.
   a. It is calculated the same way all percentages are calculated:
      (1) (# you got / total possible) x 100% = percentage
   b. For Percent Yield:
      (1) "Total possible" is the theoretical yield—because that is the largest amount of product possible, according to the Law of Conservation of Matter.
      (2) "# you got" is the actual yield—because that is the amount of product (or what you *hope* is product) that you were able to recover from the reaction.
      (3) Multiply by 100% at the end to turn the number into a percentage, instead of a decimal.
2. Some chemical reactions are naturally more "successful" than others.
   a. The reactions we do in labs in this class are ones that naturally have high yields—because we want to get reasonable amounts of product without too much trouble.
   b. In commercial chemistry, some reactions have extremely low percent yields (less than 1%) because the reactions are very complicated, and it is difficult to force all the reactants to do the hoped-for reaction.
      (1) This is one reason that some medications are much more expensive than others. If only 1% of the reactants turn into the correct product, it takes an enormous amount of time (and reactant) to produce large amounts of product.

C. Excess reactant
1. Stoichiometry problems frequently give a starting amount for only one of the reactants.
   a. In order to calculate yield based on this, the other reactants must be present as well.
      (1) If you want to make cookies and you "only have 3 eggs" then you cannot make any cookies at all—because you need other reactants/ingredients (flour, sugar, baking soda, etc)
      (2) If you want to make cookies and you only have 3 eggs, but you have plenty of everything else ("excess" of the other reactants) then the number of eggs is what determines the number of cookies you can make. (For a recipe that takes 2 eggs, you can make 1.5 times the recipe if you have 3 eggs.)
2. So why is would a problem state that there is an excess of the other reactant(s)?
   a. Because you need ALL the reactants to do the reaction—not just one of them.
   b. And saying there is "excess" means there is enough that you won't run out.
   c. Insufficient reactant would limit the theoretical yield. Excess means you can do a mole ratio and get a valid answer.