Unit 3 Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Chemical Equations Date \_\_\_\_\_\_\_\_\_\_\_\_ Block \_\_\_\_

& Stoichiometry

Unit 3C – Stoichiometry

### Knowledge/Understanding Goals:

* Molar Ratios and Stoichiometric Relationships
* The role of limiting reagents in chemical reactions

### Skills:

1. Write a proper chemical equation if given the chemical names.
2. Use stoichiometric relationships to predict theoretical yields or amount of reactant needed.
3. Identify limiting reagents.

### Notes:

Stoichiometry: The molar measurement of how much of each reactant is used and how much of each product is produced in a chemical reaction.

For example, in the chemical reaction:

3 CaCl2 + 2 Na3PO4 🡪 6 NaCl + Ca3(PO4)2

3 molecules of CaCl2 would produce 1 molecule of Ca3(PO4)2. Because a mole is always the same number of molecules, this means 3 moles of CaCl2 produces 1 mole of Ca3(PO4)2.

Stoichiometry is simply the process of using the **coefficients** in a balanced chemical equation to **convert** from **moles** of one compound to moles of another.

In the equation below, we can use ***any pair*** of coefficients to make a conversion factor. There are six possible conversion factors you could get from the equation:

3CaCl2 + 2Na3PO4 🡪 6NaCl + Ca3(PO4)2

3 mol CaCl2 = 6 mol NaCl

3 CaCl2 + 2 Na3PO4 🡪 6 NaCl + Ca3(PO4)2

2 mol Na3PO4 = 6 mol NaCl  
  
 3 mol CaCl2 = 1 mol Ca3(PO4)2

Three of the conversion factors are shown above. The other three (not shown) are:

3 mol CaCl2 = 2 mol Na3PO4  
6 mol NaCl = 1 mol Ca3(PO4)2  
2 mol Na3PO4 = 1 mol Ca3(PO4)2

In order to solve any stoichiometry problem, simply find the two compounds you need to connect and write the conversion factor using the molar ratio (coefficients) between the two compounds.

Ex: How much CO2 and how much H2O would be produced, and how much O2 would be required to burn 1.3 mol of propane (C3H8) according to the following equation?

C3H8 + 5 O2 🡪 3 CO2 + 4 H2O



\*Note that you must be in ***moles*** prior to applying the molar ratio or your units will not cancel

Ex 2: If you add a solution containing 1.75 mol of CaCl2 to a beaker of AgNO3 solution, how much solid AgCl would be formed? The chemical equation is:

2 AgNO3 (aq) + CaCl2 (aq) 🡪 2 AgCl (ppt) + Ca(NO3)2 (aq)



Limiting Reagents

Very rarely do you have the exact number of moles required to allow a reaction to go to completion (use up all of both reactants). In the above example for silver nitrate and calcium chloride, if you don’t have exactly 2 moles of AgNO3 per every 1 mole CaCl2, one of the reactants will be left over in excess.

The reactant that you run out of is called the limiting reactant (or limiting reagent) because running out of it is what limits how much product you can make, as it stops the reaction.

**Limiting Reagent**: the reactant that is not present in a high enough quantity to

completely react the other reactants.

* Stops a reaction as soon as all of this reagent is reacted.

**Excess Reagent**: the reactant that is present in a high enough quantity to

completely react the other reactants and have some reagent left over.

* Does not limit the reaction, which will continue as long as the other reactants are still present.

Tips for identifying limiting reactant problems:

1. If a scenario directly states that one reactant is present “in excess” or is “concentrated”, it is not a limiting reagent problem. You already know that the reactant not present in excess is limiting.
2. If you are given the amount of both reactants mixed in a reaction, it is a limiting reagent problem.

You can determine which reactant is limiting by comparing:

*how much of the reactant you* ***have***

***vs***

*how much would be* ***needed*** *to completely use up all of the other reactant*

Ex: Consider the following reaction:

6[](http://images.google.com/imgres?imgurl=http://www1.istockphoto.com/file_thumbview_approve/137551/2/istockphoto_137551-blank-soda-can.jpg&imgrefurl=http://www.istockphoto.com/file_closeup/image_composition_and_techniques/isolated_images/isolated_on_white/137551_blank_soda_can.php?id=137551&h=380&w=287&sz=10&hl=en&start=2&um=1&usg=__r45K1jPbQmWmiMAee3MjpVp0y9g=&tbnid=GdczgeM90ScZTM:&tbnh=123&tbnw=93&prev=/images?q=soda+can&um=1&hl=en)+ [](http://www.alltooflat.com/pranks/myths/ring/ring05.jpg) 🡪 

Suppose you have 51 cans and four six-pack rings. There are two possibilities:

1. We use up all of the cans. (Situation A)
2. We use up all of the six-pack rings. (Situation B)

|  |  |  |
| --- | --- | --- |
| **Situation** | **Cans** | **Six-Pack Rings** |
| A | **Have 51** | Would need 6.5 |
| B | Need 24 | **Have 4** |

As you can see, situation A can’t happen, because we would need too many six-pack rings. Situation B can happen, so it does.

Therefore, we make 4 six-packs, and then we run out of six-pack rings. This means six-pack rings are the limiting reactant, and we use all of them. We used up 24 cans (the non-limiting reactant), which means we had 27 cans left over.

How to ID the limiting reagent:

\*\*To begin, you must know:\*\*

1. How much of each reactant you have (grams, liters, or moles)
2. Have a complete balanced chemical equation written for the reaction

Ex: 12.8 g of copper (II) fluoride is reacted with 10.72 g of aluminum nitrate to produce copper (II) nitrate and aluminum fluoride.

Chemical Equation:

1. Convert both mass values for the reagents into moles. Pick either reactant mole value to start with.
   * This will show you how much of each reactant you have available. Remember, we can directly compare grams, so we must convert moles first.

12.8g CuF2

10.72g Al(NO3)3

1. Use the molar ratio between the reactants to convert into moles of the other reactant.
   * This will show us how many moles of the other reactant we would need to completely react the first reactant.
2. Compare the calculated number of moles needed to the number of moles you have (calculated in step 1)
   * If you have enough, the other reactant is limiting.
   * If you do not have as much as you need, that reactant is limiting.

Theoretical Yield

Since limiting reactants stop the reaction as soon as they are used up, they determine the amount of product we can theoretically expect to make in a reaction **(*Theoretical Yield*)**.

To figure out how much product (Aluminum Fluoride) you can theoretically make:

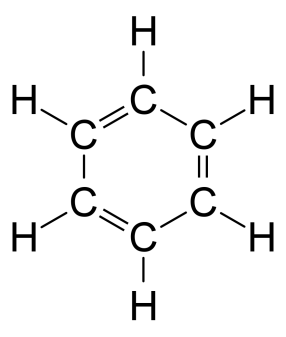
1. Identify the limiting reagent and how much of it you have in moles. (Carried over from the last example)
   * Use this molar value because the reaction will stop when we completely react all of that reactant.
2. Use the molar ratio between the limiting reagent and the product of interest (AlF3) to calculate how many moles of that product could be made.
3. Convert the moles of product to grams.
4. Compare this gram value to the mass of product actually collected in lab to find the ***Percent Yield***.

Percent Composition

**Percent Composition**: the percentage by mass of each element in a compound.

**Empirical formula**: the formula you would have for a compound if you reduced all of the

subscripts to their lowest terms. *E.g.*, the empirical formula for both acetylene (C2H2) and benzene (C6H6) is CH, but the two are very different compounds:



https://upload.wikimedia.org/wikipedia/commons/thumb/e/e3/Acetylene-2D.svg/2000px-Acetylene-2D.svg.png

You may recall that we always use empirical formulas to represent ionic compounds.

Determining Percent Composition from the Chemical Formula

To calculate the percentage by mass of any element in a chemical formula, simply calculate the molar mass of the number of atoms of that element in the compound, and divide it by the molar mass of the compound.

For example, determine the percent composition (by mass) in N2O5:

1. We can start with either atom, so I’ll arbitrarily pick nitrogen. The molar mass of the 2 nitrogen atoms in the formula is:

2 × 14.01 = 28.02

1. The molar mass of the compound is the mass of the 2 nitrogen atoms plus the mass of the 5 oxygen atoms, which is:

(2 × 14.01) + (5 × 16.00) = 28.02 + 80.00 = 108.02 g/mol

1.  nitrogen (by mass).
2. We could calculate the percentage of oxygen the same way. However, because oxygen is the only other element and the percentages must add up to 100%, it’s easier to just subtract from 100:

100 − 25.94 = 74.06% oxygen (by mass)

Determining the Empirical Formula from Percent Composition Data

1. Write the formula, but use the grams of each element as the subscripts. (If you have percentages, change % to grams.)
2. For each element in the compound, convert grams to moles.
3. Simplify the formula so that the subscripts are simple, whole numbers. (You may round the subscripts off by 5% or less.)

\*\*Note: Be able to recognize common factors.

ie: 1.2, 1.4, 1.6 = may have to multiply by a factor of 5

1.25, 1.75 = may have to multiply by a factor of 4

1.33, 1.66 = may have to multiply by a factor of 3

1.5 = may have to multiply by a factor of 2

Sample problem: a 10 g sample of a hydrocarbon is analyzed and found to contain 8.56 g of carbon and the 1.44 g of hydrogen. What is the empirical formula of this compound?

1. Write the formula as C8.56 g H1.44 g
2. Convert grams to moles:  
   C:      H:   
   Therefore, the empirical formula for this compound has the same ratio as C0.713H1.429
3. Convert the subscripts to simple whole numbers. The easiest way to do this is to divide them all by the smallest one and see what happens.

 which we can round off to CH2

If the problem gives percentages instead of actual mass, you have two options.

1. Pretend the percentages are out of 100 g total. *E.g.*, if you had a compound containing 25.3% nitrogen, you would use 25.3 g of nitrogen in your calculations.
2. If you are given the molar mass of the compound, multiply the percentages by the molar mass and use that number of grams. This way, the subscripts you get when you convert grams to moles will be the actual chemical formula.

Again: Don’t round your fractional subscripts by more than about 5%. If you have something like N1O2.5, double all of the subscripts to get N2O5. (This means you need to be able to recognize the decimal equivalents for simple fractions, such as,  etc.)

Empirical vs. Molecular Formulas

If you know the molar mass of the compound, you can use it to get from the empirical formula to the molecular formula by dividing the molar mass by the empirical mass.

For example, suppose the molar mass of the hydrocarbon in the previous problem was known to be .

The molar mass of the empirical formula (CH2) would be:

(1 x 12.011) + (2 x 1.008) = 14.027.

The actual molar mass of 42.08 is 3 times as much as the empirical formula mass (42.08 ÷ 14.027 = 3), so the actual formula must have 3 times as much of everything. This means the actual formula is C3H6.

Stoich Khan Academy Videos:

Stoichiometry: <http://bit.ly/18MLezd>

Stoich Ex Problem #1: <http://bit.ly/15NPAUH>

Limiting Reagent problem: <http://bit.ly/15TtezZ>

Empirical and Molecular Formulas: <http://bit.ly/15h8fvB>