CHM 130LL: Mole Relationships

Introduction

Moles

A mole, known as the "chemist's dozen," relates the submicroscopic world of atoms and molecules to the much larger world of weighable amounts of chemicals. Just like a dozen is always 12 items, a mole is always 6.02×10^{23} items. Atoms and molecules are so small that it takes a very large number of them to have enough chemical to weigh out or use in a reaction. To obtain the weight of a mole of atoms, you simply need to look up the ***molar mass*** of the appropriate element on the periodic table. For example, a mole of C atoms has a mass of 12.01 grams while one mole of Cu atoms has a mass of 63.55 grams. One "mole" always contains the same number of atoms, $\mathbf{6.02 \times 10^{23}}$ (known as ***Avogadro's number***).

Mole Calculations

Most chemical calculations do not require the use of Avogadro's number, only an understanding of its concept. We can see this by again using the "dozen" analogy. A person does not need to use the number "12" if he is buying something that comes in dozens. For example, if he wanted to buy 3 dozen eggs that cost \$3 per dozen, he will need to pay 3×\$3.00 or \$9.00.

$$3 \text{ dozen eggs } \times \frac{\$3.00}{\text{dozen eggs}} = \$9.00$$

Likewise, a chemist does not need to use Avogadro's number to determine the mass of carbon to weigh out a given number of moles in the lab. One simply needs to know the mass of one mole of carbon (12.01 grams), or the molar mass of carbon, $\frac{12.01 \text{ g C}}{\text{mole C}}$, to determine the mass of carbon required to get 3.00 moles of C atoms, as shown below:

Example 1: Calculate the mass of carbon needed to get 3.00 moles of carbon.

$$3.00 \text{ moles C} \times \frac{12.01 \text{ g C}}{\text{mole C}} = 36.0 \text{ g C}$$

One can also calculate how many moles of carbon are contained in a certain mass of carbon. This is analogous to a shopper determining how many dozen eggs she can get given \$15.

$$\$15 \times \frac{\text{dozen eggs}}{\$3} = 5 \text{ dozen eggs}$$

Likewise, if a chemist has 15.0 grams of carbon, the chemist can determine the number of moles of carbon present in the following way:

Example 2: How many *moles* of carbon are present in 15.0 grams of carbon?

$$15.0 \text{ g C} \times \frac{\text{mole C}}{12.01 \text{ g C}} = 1.25 \text{ moles C}$$

The examples above use the element carbon, which exists as individual atoms, but compounds generally exist as molecules or ionic compounds. For example, water exists as H_2O molecules, while table salt, NaCl, exists as a three-dimensional network of Na^+ and Cl^- ions. To determine the mass of one mole of a compound, one must multiply the molar mass of each element by the number of atoms present for the element, then add all the masses of the elements to get the total molar mass for the compound.

For example, one can calculate the molar mass (MM) for $Ca(NO_3)_2$ as follows:

$$\begin{aligned} \text{MM of } Ca(NO_3)_2 &= \text{MM of Ca} + 2(\text{MM of N}) + 6(\text{MM of O}) \\ &= (40.08\ \frac{g}{mol}) + 2(14.01\ \frac{g}{mol}) + 6(16.00\ \frac{g}{mol}) \\ &= 164.10\ \frac{g}{mol}\ Ca(NO_3)_2 \end{aligned}$$

Balanced Chemical Equations and Moles

Balanced chemical equations give the relationship between the number of molecules in the reactants and products. For example, the balanced equation

$$\mathbf{3\ H_2\,(g) + N_2\,(g) \rightarrow 2\ NH_3\,(g)}$$

tells us that 3 molecules of H_2 react with 1 molecule of N_2 to produce 2 molecules of NH_3. We can also "read" the balanced equation in terms of moles. That is, 3 moles of H_2 react with 1 mole of N_2 to produce 2 moles of NH_3. From this statement, we can always relate the moles of any reactant or product to the moles of any other reactant or product using a mole-to-mole ratio obtained from the equation. For example, if we want to know how many moles of H_2 it takes to react with 2.00 moles of N_2, we would simply use the mole ratio between H_2 and N_2 from the equation above, being sure to set up the unit factor to cancel moles of N_2 and leave moles of H_2. The problem is set up below:

Example: How many moles of H_2 are required to react with 2.00 moles of N_2?

$$2.00 \text{ moles } N_2 \times \frac{3 \text{ moles } H_2}{1 \text{ mole } N_2} = 6.00 \text{ moles } H_2$$

What if we have more than 6.00 moles of H_2 available to react with the 2.00 moles of N_2? The extra H_2 simply does not react. So if we mix our 2.00 moles of N_2 with 10.00 moles of H_2, we would have an extra 4.00 moles of H_2 left over at the end of the reaction. In this case, chemists refer to H_2 as the **reactant in excess** (leftover), and N_2 is the **limiting reactant** (because it *limits* the amount of product formed). We could then calculate the number of moles of NH_3 formed, as shown below:

$$2.00 \text{ moles } N_2 \times \frac{2 \text{ moles } NH_3}{1 \text{ moles } N_2} = 4.00 \text{ moles } NH_3$$

Now we can combine molar mass and mole calculations to determine the mass of product that can be produced given the mass of reactants used. First we use molar mass to convert the mass for each reactant to moles, then use the appropriate mole-to-mole ratio, then finally we use the molar mass of the product to convert from moles to the mass of the product. Since we solve for the same product, we simply compare the two numbers. Whichever number is smaller is the amount of product made because as soon as that amount of product is made, the reactant is completely used up, and no more product can be made.

Example: What mass of ammonia can be produced when 10.0 g of H_2 reacts with 10.0 g of N_2?

$$10.0\ \text{g}\ N_2 \times \frac{\text{mole}\ N_2}{28.02\ \text{g}\ N_2} \times \frac{2\ \text{moles}\ NH_3}{1\ \text{mole}\ N_2} \times \frac{17.04\ \text{g}\ NH_3}{\text{mole}\ NH_3} = \mathbf{12.2\ g\ NH_3}$$

$$10.0\ \text{g}\ H_2 \times \frac{\text{mole}\ H_2}{2.02\ \text{g}\ H_2} \times \frac{2\ \text{moles}\ NH_3}{3\ \text{mole}\ H_2} \times \frac{17.04\ \text{g}\ NH_3}{\text{mole}\ NH_3} = 56.2\ \text{g}\ NH_3$$

In this example, the N_2 produced the *smaller amount of NH_3, 12.2 g, so that is the amount of product formed*. Once 12.2 g of NH_3 are produced, the N_2 is completely used up, so *N_2 is the limiting reactant*. Since there is enough H_2 to make 56.2 g of NH_3, once 12.2 g of NH_3 are produced, there is still some H_2 remaining, so *H_2 is the reactant in excess*.

In this experiment, you will prepare and mix solutions of lead(II) nitrate and potassium iodide to form a precipitate, $PbI_2(s)$. The equation for the reaction is:

$$\mathbf{Pb(NO_3)_2\,(aq) + 2\,KI\,(aq) \rightarrow PbI_2\,(s) + 2\,KNO_3\,(aq)}$$

You and your partner will obtain a card indicating the mass of each reactant, $Pb(NO_3)_2$ and KI, to weigh out. You will calculate the mass of precipitate that should be produced, and thus, determine which reactant is the limiting reactant and which is the reactant in excess. Then you will carry out the experiment to check your predictions.

Procedure

Caution: Lead (II) nitrate is toxic. DO NOT INHALE ANY $Pb(NO_3)_2$ POWDER! Avoid contact with skin, eyes, and clothing. If the solid or solution contacts your skin, rinse completely with water immediately.

A. Weigh Out the Reactants

1. Check your card to determine the mass of each reactant to weigh out. Record your card number on your Lab Report. Because you will use a beaker to weigh out the solids, the stockroom staff has provided ***examples*** of 0.10 grams, 0.20 grams, and 0.30 grams of $Pb(NO_3)_2$ and KI in 50-mL beakers to help you estimate the amount of reactant you need.
   PLEASE DO NOT USE THESE SAMPLES!
2. To weigh each compound, put a clean, dry, 50-mL beaker on the electronic balance, and tare (zero) the balance. Remove the beaker from the balance *without* zeroing the balance. Add the approximate amount of solid you will need, and place the beaker back on the balance. If you are within 0.02 grams of the amount on the card, record the mass of your sample (to 4 decimal places). If you are not within this amount, remove the beaker from the balance, and use a clean, dry spatula to add or remove a little solid, and try again. **Do not place any chemicals back in the reagent bottle!** Instead, dispose of excess reagent in the beaker labeled “Waste” in the box near the balances. Be sure to ***record the mass of your sample to 4 decimal places***.

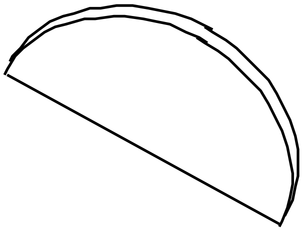
B. Carry Out the Reaction

Add about 10-15 mL of deionized water to each reactant to dissolve completely, stirring with a glass stirring rod. Then mix the two solutions in one beaker, again stirring with the glass rod to ensure complete mixing. Use a stream of deionized water from your wash bottle to wash any remaining reagent from its original beaker into your mixture. Let the precipitate sit for 5 minutes.

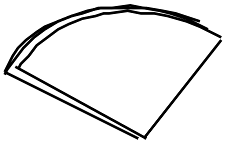
C. Collect and Dry the Precipitate

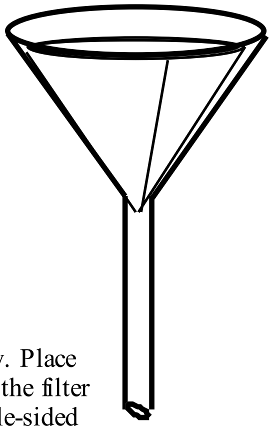
1. Fold a piece of filter paper twice (as shown below). Weigh and record the mass of the filter paper.

Step 1: Fold filter paper in half and crease lightly



Step 2: Fold again into quarters.





Step 3: Lift up one layer of filter paper, leaving 3 layers below. Place the filter paper cone into the funnel. Press the edges of the filter paper against the sides of the funnel, and wet the single-sided edge with deionized water, so the paper sticks to the funnel..

2. Place the filter paper into the funnel, and place a clean flask under the funnel to catch the filtrate. Carefully filter the precipitate without overfilling. **Collect and save the filtrate.**

   Note: Your filtrate should be clear or yellow in color but with no particles. If you have particles in your filtrate, heat the filtrate and remaining precipitate for 5 minutes on a hot plate at the medium setting, then cool before filtering into the same filter paper.

3. Use a stream of water from your wash bottle and your rubber policeman to transfer all of the precipitate from the beaker to your filter paper. After most of the filtrate drips out, wash the precipitate several times with deionized water from your wash bottle. When the contents of the filter no longer drip, remove the paper from the funnel, spread it out, and place it on paper towels. Write your initials on the towels, then place the precipitate in the oven to dry.

Caution: Lead (II) iodide, PbI_2, is toxic. DO NOT INHALE ANY PbI_2 POWDER! Avoid contact with skin, eyes, and clothing. If the solid or solution contacts your skin, rinse completely with water immediately.

4. ***Note: After heating, the precipitate will be very light and can easily blow away. Place your watchglass over the precipitate on the filter paper when you remove it from the oven to prevent the precipitate from blowing away.*** Carefully remove your precipitate and filter paper from the oven after about 20-30 minutes, and allow it to cool to room temperature. Place your hand above the precipitate to verify that no heat is emanating from it before

weighing.

D. Calculating the Theoretical Yield and Determining the Limiting Reactant and the Reactant in Excess

1. While your precipitate is drying in the oven, calculate the molar mass for each of the following compounds: $Pb(NO_3)_2$, KI, and PbI_2 (each to 2 decimal places).
2. Calculate the mass of PbI_2 that can be produced with the amount of $Pb(NO_3)_2$ used, then calculate the mass of PbI_2 that can be produced with the amount of KI used.
3. Compare the mass of PbI_2 that can be produced from both reactants; the lower mass of PbI_2 is the theoretical yield for the precipitate. Indicate if $Pb(NO_3)_2$ or KI is the limiting reactant (or limiting reagent) and the other reactant as the reactant in excess (excess reagent).

E. Check Your Calculations

a. Indicate the limiting reactant and the reactant in excess based on your calculations.

b. Considering the limiting reactant and the reactant in excess, indicate which reactant, $Pb(NO_3)_2$ or KI, must be present in the filtrate (the solution that went through the filter paper).

c. Given the reactant you indicated must be present in the filtrate solution, predict if a precipitate will form if you add more $Pb(NO_3)_2(aq)$ to your filtrate and explain why. Similarly, predict if a precipitate will form if you add more $KI(aq)$ to your filtrate and explain why.

d. Next, test out your predictions. Use a disposable pipet to transfer a little of your filtrate solution into each of two test tubes. Solutions of KI and $Pb(NO_3)_2$ have been prepared in dropper bottles for you to use. Add a few drops of the prepared $Pb(NO_3)_2$ solution to one of the test tubes, and record your observations. Next, add a few drops of the prepared KI solution to the other test tube, and record your observations.

e. Indicate if your experimental observations agree with your predicted results. If they did not agree, double-check your calculations to make sure you have the correct limiting reactant and reactant in excess.

F. Percent Yield

After the filter paper containing the PbI_2 precipitate has come to room temperature, weigh the PbI_2 along with the paper, and record the mass. Subtract the initial mass of the filter paper alone to determine the actual mass of PbI_2 precipitate obtained. **Do not dispose of the precipitate until after you finish all of your calculations.**

Calculate the percent yield for your PbI_2 precipitate using the following formula:

$$\textbf{percent yield} = \frac{\textbf{actual yield [from experiment]}}{\textbf{theoretical yield [from calculations]}} \times 100\%$$

A percent yield greater than 100% indicates the precipitate may still be wet and must be reheated for at least 10 minutes. After reheating, allow the precipitate to cool for 10 minutes, then reweigh the precipitate and filter paper, and recalculate your percent yield.

NOTE: ***All waste—all precipitates and excess solutions with KI and $Pb(NO_3)_2$—for this experiment should be disposed of in the waste containers in the hood.***

CHM 130LL:
Mole Relationships

Name: ________________________

Partner: ________________________

Section Number: ________________

LAB REPORT

DATA

Card Number: ________

	Assigned (from card)	Actual (weighed out)
Mass of $Pb(NO_3)_2$:	__________________	__________________
Mass of KI:	__________________	__________________

Mass of filter paper: ____________________

Show calculations for the molar mass of each compound below:

$Pb(NO_3)_2$: ________________ KI: ________________ PbI_2: _______________

$$\mathbf{Pb(NO_3)_2\,(aq) + 2\,KI\,(aq) \rightarrow PbI_2\,(s) + 2\,KNO_3\,(aq)}$$

a. Calculate the mass of PbI_2*(s)* that can be produced with the ***actual*** mass of $Pb(NO_3)_2$ used.

b. Calculate the mass of PbI_2*(s)* that can be produced with the ***actual*** mass of KI used.

Calculated mass of lead (II) iodide produced (theoretical yield): ____________________

limiting reactant = _________________ reactant in excess = __________________

Check Your Calculations

a. Refer to the bottom of page 7, and identify the following:

limiting reactant = __________________ reactant in excess = ___________________

b. Based on your answers in part a, what reactant remains and must still be present in the filtrate? (Circle one) $Pb(NO_3)_2$ KI

c. Based on your answer in part b, will a precipitate form if you add more $Pb(NO_3)_2$*(aq)* to your filtrate? Explain.

Based on your answer in part b, will a precipitate form if you add more KI*(aq)* to your filtrate? Explain.

d. Record your experimental observations below:

When $Pb(NO_3)_2$*(aq)* is added to the filtrate	
When KI*(aq)* is added to the filtrate	

e. Did your results agree with your predictions in part c? (Circle one) Yes No

Note: ***If your results did not agree with your predictions, double-check your calculations for errors.***

Actual Yield

Mass of PbI_2 precipitate + filter paper (after heating) __________________

Mass of filter paper (from DATA section on p. 7) __________________

Mass of PbI_2 precipitate (actual yield) __________________

Percent Yield

Calculate the percent yield below: (Show work!)

Percent yield: __________________

Question: Two students were rushing to finish the experiment and left their precipitate in the oven for only 10 minutes. Consequently, their calculated percent yield was 105.1%. Explain how this was possible.

Stoichiometry Problems

For the problems below, show all work clearly, and express final answers with the correct units and the correct number of significant figures.

1. The fertilizer ammonium sulfate, $(NH_4)_2SO_4$, can be prepared by the reaction of ammonia, NH_3, with sulfuric acid:

$$2\ NH_3\ (g) \ + \ H_2SO_4\ (aq) \ \rightarrow \ (NH_4)_2SO_4\ (aq)$$

Calculate the mass of ammonia required to produce 125 g of ammonium sulfate.

2. Barium chloride reacts with potassium phosphate as follows:

$$3\ BaCl_2\ (aq) \ + \ 2\ K_3PO_4\ (aq) \ \rightarrow \ Ba_3(PO_4)_2\ (s) \ + \ 6\ KCl\ (aq)$$

   a. Which substance is the precipitate in the reaction above? __________________.

   b. Calculate the mass of precipitate that forms when 50.0 g of barium chloride reacts completely with excess potassium phosphate.

3. The rating used for gasoline generally indicates the percentage of octane, C_8H_{18}, present; for example, regular gasoline has a rating of 87 because it is 87% octane. The equation for the reaction shows how burning octane produces steam and carbon dioxide gas, a greenhouse gas contributing to global warming:

$$2\ C_8H_{18}\ (l) \ + \ 25\ O_2\ (g) \ \rightarrow \ 18\ H_2O\ (g) \ + \ 16\ CO_2\ (g)$$

Assuming complete combustion to products, calculate the mass of carbon dioxide gas produced when 100.0 g of octane burns in excess oxygen present in the air.