Stoichiometry
Exploring a Student-Friendly Method of Problem Solving

Stoichiometry comes in two forms: composition and reaction. If the relationship is between the quantities of each element in a compound it is called composition stoichiometry. If the relationship is between quantities of substances as they undergo a chemical reaction it is called reaction stoichiometry. In plain English, if you have to calculate just about anything to do with moles or chemical quantities, we collectively call those computations *stoichiometry.*

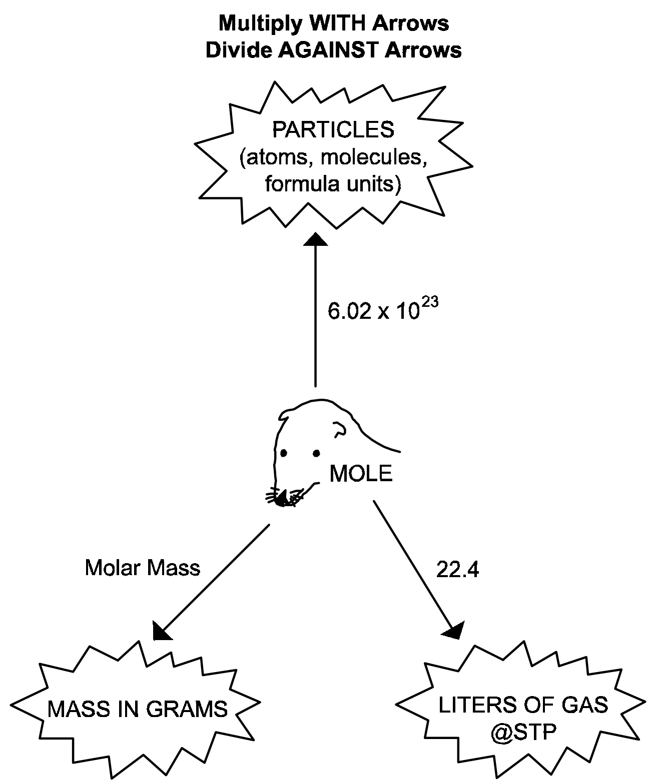
THE MOLE CONCEPT

To be successful in solving stoichiometry problems you must be familiar with the following terms and their meaning:

- **mole**—the number of C atoms in exactly 12.0 grams of ^{12}C; also a number, 6.02×10^{23} just as the word "dozen" means 12 and "couple" means 2.
- **Avogadro's number**, n—6.02×10^{23}, the number of particles in a mole of anything
- **molar mass**—the mass of one mole of particles in grams; the sum of the atomic masses in a chemical formula. When calculating be sure to multiply the atomic mass of each element by the subscript following that element. The molar mass for H_2O is 2(1.01 for hydrogen) + 16.00 for oxygen = 18.02 grams for a mole of water molecules.

THE MOLE MAP

The mole map is a quick way to calculate any quantity you desire that is related to moles. Learn to reproduce this map (in other words, memorize it) so that you may use it as a problem-solving tool when faced with stoichiometry problems.



Notice the number of moles is *always* in the center of the map which means when converting moles into any other unit you must multiply the number of moles by the conversion factor found next to the arrow. The conversion factors included on this map will allow you to convert the number of moles to either the number of particles, mass of matter in grams, or liters of a gas at standard temperature and pressure, STP (1 atm pressure and 273 K).

If you have one of the above quantities and wish to find the number of moles simply start at the spot on the map that corresponds to the quantity "given" in your problem and divide by the appropriate conversion factor. Once there, you can multiply by the appropriate conversion factor to solve for a different quantity.

Substance	MM	Number of moles	Mass in grams	Number of Particles	Liters of gas @ STP
carbon dioxide		3.0			
oxygen			64.0		
methane					6.25
nitrogen				9.50×10^{25}	

Example: Complete the following table using the mole map and a periodic table. Your teacher has the correct answers.

Now that you understand how to use the mole map, you must realize that being successful in solving stoichiometry problems requires that you have demonstrated proficiency in the following individual skills:

- *Correctly* writing chemical formulas—this requires knowledge of your polyatomic ions and being able to use the periodic table to deduce what you have not had to memorize.
- *Correctly* calculating molar masses from a chemical formula.
- *Correctly* balancing chemical equations.

SOLVING STOICHIOMETRY PROBLEMS

This method will use a table to help you organize your thoughts and focus on the task at hand. Each time you embark upon solving a reaction stoichiometry problem, reproduce the row headings found below and place them on the left side of your paper as shown:

Molar Mass:
Balanced Eq'n:
mole:mole
 # moles:
amount:

Once you have a clean template written on your paper, work through Example 2 and add the appropriate information to you table.

Example 2: What mass of oxygen will react with 96.1 grams of propane? (Notice the balanced equation was not provided, you will have to supply the chemical formulas!)

1. Write a chemical equation. Double check to be sure you have correct chemical formulas!

Molar Mass:								
Balanced Eq'n	C_3H_8	+	O_2	$\rightarrow$	CO_2	+	H_2O	
mole:mole								
# moles								
amount								

2. Calculate the molar masses and write them in parentheses above the formulas—soon you'll figure out you don't have to do this for every reactant and product, just those your question uses.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	O_2	→	CO_2	+	H_2O
mole:mole							
# moles							
amount							

3. Balance the chemical equation using coefficients. Do not attempt to change the subscripts!

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole							
# moles							
amount							

4. Look at the coefficients in the balanced equation. These numbers ARE the **mole:mole** ratio. Re-write them under the balanced equation so that they are easy to see.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles							
amount							

5. Next, re-read the problem and look for some information about an amount of one of the substances—in this example it is 96.1 g of propane.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles							
amount	96.1 grams						

6. Use the mole map to calculate the number of moles of something, anything! Start at 96.1 grams, divide (against the arrow) by molar mass to get the # moles of propane.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	→	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18		10.9		6.54		8.72
amount	$\frac{96.1\text{g}}{44.11\text{g/mol}} = 2.18\text{mol}$						

7. USE the mole: mole ratio to find the moles of EVERYTHING! If 1 = 2.18 mol propane, then moles oxygen equals 5(2.18) mol or 10.9 moles etc…. (IF you found the number of moles for something that has a coefficient other than "1", just divide by the coefficient it has and multiply by the new coefficient.) Leave everything in your calculator—we only rounded to save space!

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	$\rightarrow$	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18		10.9		6.54		8.72
amount	9.16 g						

8. Re-read the problem to determine which amount you are seeking. First, we'll find the mass of oxygen required since that's what the problem asked. 10.9 moles x 32.00 g/mol = 349 g of oxygen.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	$\rightarrow$	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18		10.9		6.54		8.72
amount	9.16 g		349 g		146 L		

What if part (b) of Example 2 asked for liters of CO_2 at STP? Simply take the number of moles of CO_2 from the table and use the mole map to solve for the liters of gas at STP. Start in the middle of the map with 6.54 moles and multiply by 22.4 L/mol to determine that 146 L of CO_2 gas is produced at STP.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	$\rightarrow$	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18 mol		10.9		6.54		8.72
amount	9.16 g		349 g				

What if part (c) asked how many water molecules are produced? Simply multiply the number of moles of water 8.72 mol by 6.02 X 10^{23} to determine that 5.25 x 10^{24} molecules of water were produced.

Molar Mass:	(44.11)		(32.00)		(44.01)		(18.02)
Balanced Eq'n	C_3H_8	+	5 O_2	$\rightarrow$	3 CO_2	+	4 H_2O
mole:mole	**1**		**5**		**3**		**4**
# moles	2.18		10.9		6.54		8.72
amount	96.1 g						5.25 x 10^{24}

CALCULATIONS INVOLVING A LIMITING REACTANT

Ever notice how hot dogs are sold in packages of 10 and the buns come in packages of 8? The bun is the limiting reactant and limits hot dog production to 8 when one package of each is purchased. The limiting reactant (or reagent) is the one that runs out first in the chemical reaction.

Plan of attack: First, you need to recognize that you ***need*** a plan of attack! Your clue is that you are given TWO amounts of matter that react.

Next, when in doubt…find the number of moles. Set up your table like before, only now you will have TWO amounts and thus TWO numbers of moles to get you started. Simply divide each number of moles by its coefficient in the balanced equation and select the smaller number of moles since it is the limiting reagent. Be sure to cross out the larger value of moles and recalculate its value based on the limiting reagent.

The Haber process synthesizes ammonia, often used for fertilizer production, from the reaction of nitrogen and hydrogen gases. The nitrogen gas is drawn from the atmosphere while the hydrogen gas is obtained from the reaction of methane with water vapor. It is carried out at 250 atm and 400°C in the presence of a catalyst. This process has ultimately saved millions from starvation. The reaction is below:

Molar Mass:	(28.02)		(2.02)		(17.04)
Balanced Eq'n	N_2	+	3 H_2	$\rightarrow$	2 NH_3
mole:mole	**1**		**3**		**2**
# moles					
amount					

Suppose 25.0 kg of nitrogen reacts with 5.00 kg of hydrogen to form ammonia. (a) What mass of ammonia can be produced? (b) Which reactant is the limiting reactant? (c) What is the mass of the reactant that is in excess?

Insert the masses in the amount row and find the number of moles of BOTH!

Molar Mass:	(28.02)		(2.02)		(17.04)
Balanced Eq'n	N_2	+	3 H_2	$\rightarrow$	2 NH_3
mole:mole	**1**		**3**		**2**
# moles	892 moles		2,475 moles		
amount	25,000 g		5,000 g		

Next, divide each number of moles by its coefficient. Compare these mole values and select the smaller number of moles as your limiting reactant. N_2 has a coefficient of one already, so just focus on the 892 moles present. For H_2, your focus becomes 2,475 ÷ 3 = 825 moles, which is smaller than the number of N_2 moles, so H_2 is your limiting reagent.

Now you are ready to modify the table: Cross out 25,000 g and 892 moles of N_2 since they will not limit the reaction and replace it with the 825 moles of N_2 consumed (since nitrogen has a coefficient of one). If needed, you could then calculate the actual mass of N_2 that reacts.

Molar Mass:	(28.02)	(2.02)	(17.04)
Balanced Eq'n	N_2 +	$3\ H_2$ →	$2\ NH_3$
mole:mole	**1**	**3**	**2**
# moles	825 mol used ↑ ~~892~~ moles	∴ 2,475 moles used	∴ 1,650 mol produced
amount	∴ 825 mol x 28.02 = 23,117 g used ~~25,000 g~~ 1,883 g excess!!	5,000 g	1,650 mol (17.04) = 28,116 g produced

Here's the question again, let's clean up any significant digit issues:

Suppose 25.0 kg of nitrogen reacts with 5.00 kg of hydrogen to form ammonia. (3 sig. dig. limit)

(a) What mass of ammonia can be produced? 28,100 g produced = 28.1 kg (it is always polite to respond in the mass unit given).

(b) Which reactant is the limiting reactant? hydrogen—once that's established, forget about how much nitrogen you were given and base the calculation on hydrogen's amount.

(c) What is the mass of the reactant that is in excess? 1,883 g = 1.88 kg excess nitrogen.

PURPOSE

In this activity you will explore a different method of solving stoichiometry problems. At the end of this activity you should be able to determine which method of problem solving allows you to work the fastest and most accurately.

MATERIALS

calculator　　　　paper and pencil

PROCEDURE

1. Solve the problems found on your student answer pages. Use your own paper but staple your student answer page to the front of your notebook paper to turn in to your teacher.

2. Be sure to show all work and pay close attention to the proper use of significant digits and units.

Name ______________________

Period______________________

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CONCLUSION QUESTIONS

1. Automotive air bags inflate when a sample of NaN_3 is very rapidly decomposed. (a) What mass of NaN_3 is required to produce 375 L of nitrogen gas at STP? (b) How many sodium atoms are produced when 375 L of nitrogen gas at STP is produced? (c) What is the mass of the sodium atoms produced?

2. When ignited or subjected to sudden impact, nitroglycerine decomposes very rapidly and exothermically:

$$4C_3H_5N_3O_9\,(l) \rightarrow 6N_2(g) + 12CO_2(g) + 10H_2O\,(g) + O_2(g) + \text{energy}$$

Pure nitroglycerine is quite dangerous. Alfred Nobel discovered in 1867 that if nitroglycerine was absorbed into porous silica, it can be handled quite safely. He named this new mixture dynamite. This brought him a great fortune which he used to establish the Nobel Prizes given today. (a) The density of nitroglycerine is 1.60 g/ml. What is the volume in liters of a mole of liquid nitroglycerine? (b) How many moles of water vapor are formed when one mole of nitroglycerine detonates and decomposes completely? (c) What is the total volume of the gases produced when one mole of liquid nitroglycerine is completely vaporized? (d) By what factor does the volume increase once the liquid nitrogen sample has been completely vaporized?

3. On April 6, 1938, at DuPont's Jackson Laboratory in New Jersey, DuPont chemist, Dr. Roy J. Plunkett, was working with gases related to Freon® refrigerants, another DuPont product. Upon checking a frozen, compressed sample of tetrafluoroethylene, he and his associates discovered that the sample had polymerized spontaneously into a white, waxy solid to form polytetrafluoroethylene (PTFE). This was not expected and a bit of a disaster. It would set their work back by many days. As it turns out PTFE is inert to virtually all chemicals and is considered the most slippery material in existence. Thus, Teflon® was born. Teflon® is a polymer. The synthesis reaction for a fragment of Teflon® is given below:

$$C_2F_4 \rightarrow C_{50}F_{200}$$

(a) What mass of reactant is required to form 575 grams of Teflon®? (b) A skillet is massed before the application of a Teflon® coating. Its initial mass is 235 g. The coating is applied and its final mass is determined to be 269g. How many C_2F_4 particles were contained in the Teflon® coating applied to the skillet? (c) How many fluorine atoms were contained in the coating?

4. NASA has engineered reusable booster rockets for the U.S. space shuttle program. These booster rockets use a mixture of powdered aluminum metal and solid ammonium perchlorate for fuel. Solid aluminum oxide is formed as a product along with solid ammonium chloride. It is the formation of nitrogen monoxide gas and water vapor from the two solid reactants that gives the shuttle its boost. (a) What mass of ammonium perchlorate should be used in the fuel mixture for every kilogram of aluminum? (b) What is the total number of gas moles produced? (c) How many liters of water vapor are produced per kilogram of aluminum in the fuel mixture? (d) How many total liters of gas at STP are produced when 100.0 grams of powdered aluminum reacts with 100.0 grams of ammonium chloride.

5. Another DuPont product, Kevlar®, was originally developed as a replacement for steel in radial tires. Kevlar® is now used in a wide range of applications including bullet proof vests since it is twenty times stronger than steel. Kevlar® was first synthesized in 1964 by Stephanie Kwolek at the DuPont laboratories in Wilmington, Delaware. A monomer of Kevlar is $C_{14}H_{10}N_2O_2$. (a) Write the chemical reaction that would synthesize a monomer of Kevlar® from its elements. (b) If 500.0 g of both nitrogen gas and hydrogen gas at STP were reacted with excess carbon and oxygen gas, how many kilograms of Kevlar® could be produced? (c) What mass of excess reagent remains unreacted? (d) How many Kevlar® monomer units are produced as a result?