Name___________________________________

Math of Chemistry Part 2

Unit 6B

Vocabulary

1. subscript: a number to the right and slightly below a symbol; tells the number of atoms of that element present.
2. coefficient: a number in front of a formula in a chemical equation. Applies to the entire formula; tells how many units of the formula are present.
3. reactant: a substance present before a chemical reaction begins. Shown on the left side of a chemical equation.
4. product: a substance formed during a chemical reaction. Shown on the right side of a chemical equation.
5. law of conservation of mass: matter cannot be created or destroyed. There must be the same number of each kind of atoms present before and after a reaction. There will also be the same mass present before and after a reaction.
6. empirical formula: the simplest whole number ratio of atoms of each element in a compound. Example: the empirical formula for sugar ($C_6H_{12}O_6$) is CH_2O
7. molecular formula: shows the actual numbers of each type of atom in the molecule. Ex: CO_2
8. mole: 6.02×10^{23} of whatever objects you are counting.
9. empirical formula: formula showing the lowest whole number ratio of elements in a compound
10. molecular formula: formula showing the number of atoms of each element present in the actual compound
11. mole ratio: ratio of the coefficients in a balanced equation. For the balanced equation $2Na + Cl_2 \rightarrow 2NaCl$, the mole ratio of chlorine to sodium is 1:2. The mole ratio of NaCl to Na is 2:2.

Learning Targets:

Enduring Understanding: Chemical reactions can be predicted based on the nature of the reactants.

Essential Questions: How and why do atoms form compounds? How are the properties of a compound determined? How can you determine the quantities for chemical equations? How much of each ingredient do you need to make something? Do chemical equations conserve mass?

Objectives: ***I CAN:***

1. Identify a compound as empirical, molecular or structural
2. Read chemical equations correctly.
3. Identify missing reactants or products in a chemical reaction.
4. Balance chemical equations using the law of conservation of mass as rationale.
5. Distinguish between formula mass and gram formula mass
6. Given a chemical formula, determine the formula mass and the gram formula mass
7. **Given an empirical formula and the molar mass (gram formula mass) of a compound, determine the molecular formula of the compound.**
8. Given a chemical formula, determine the % composition by mass
9. Define "mole"
10. Given a mass of a compound, determine the number of moles of the substance present,
11. Given a number of moles of a compound, determine the mass of a compound.
12. **Use a balanced equation to determine quantitative molar relationships among the reactants and products (given a starting amount of one chemical reactant or product, calculate the matching amount of another reactant or product.**
13. **Determine the mole ratio of elements and compounds given a balanced (or unbalanced) chemical reaction.**

Calendar for unit 6A Regents Chemistry: Red (2 & 3A) and Yellow (9 &8A) classes

1/29	***30***	***31***	***2/1***	***2/2***
A	**B**	**C**	**D**	**E**
6.7: Determining empirical formulas & molecular formulas given data HW 6B Assignment 1	Mystery problem: Can you solve the case?	6.8 Determining and calculating using mole ratios HW 6B Assignment 2	Review for mini test 6B	Unit 6B mini test (includes concepts from 6a)
Midterm analysis and reflection Lab		Chemistry Work Period		Chemistry Work Period

Calendar for unit 6A Regents Chemistry: Blue (5 & 4B) and Green (7 & 8B) Classes:

1/29	***30***	***31***	***2/1***	***2/2***
A	**B**	**C**	**D**	**E**
6.7: Determining empirical formulas & molecular formulas given data HW 6B Assignment 1	Mystery problem: Can you solve the case?	6.8 Determining and calculating using mole ratios HW 6B Assignment 2	Review for mini test 6B	Unit 6B quiz (includes concepts from 6a)
	Midterm analysis and reflection Lab		Chemistry Work Period	

How are empirical and molecular formulas determined? How might this be applicable to real world scenarios (mystery)?

A. Quick review: An empirical formula is …

B. Mass spectrometers are used in chemistry labs, crime labs and forensic science to help determine the ratios (and therefore percent) of elements in a sample. Once this is determined, the scientist (YOU) can determine the empirical formula!

Let's try: A compound is analyzed and found to contain 25.9% nitrogen and 74.1% oxygen. What is the empirical formula of the compound?

Step 1: Write the element symbol with the percent out of 100 turned into 100g **Step 2:** Divide each element by the gram formula mass to get moles (table T formula) **Step 3:** Divide each mole amount by the smaller number of moles (this is to help get whole number subscripts) **Step 4:** If subscripts are still not whole numbers, multiply each part of the ratio by the smallest whole number that will convert both subscripts to whole numbers.	N	O

C. What if this does us NO good? It's the molecular formula (typically) that results in identification of an unknown. This is also determined from a mass spectrometer. Sometimes the molecular formula is _________________________________ as its experimentally determined empirical formula, or it s a simple _____________- ____________ multiple of its empirical formula. **How can this be done?**

Let's try! Practice Problem:

Calculate the molecular formula of a compound whose molar mass is 60.0 g/mol and empirical formula is CH_4N.

STEPS TO FIND A MOLECULAR FORMULA	
Step 1: Find the mass of the empirical formula given in grams	
Step 2: Divide the molecular mass given for the compound (the LARGER mass) by the mass of the empirical formula that you just found in step 1 (the smaller mass).	
Step 3: Round your answer to the closest whole number.	
Step 4: You now have a multiplier from step 3. Multiply each subscript in the given empirical formula by this multiplier. You now have the molecular formula you wanted.	

Practice Questions:

1. Find the molecular formula of ethylene glycol, which is used as antifreeze. The molar mass is 62g/ mol and the empirical formula is CH_3O
2. A compound has an empirical formula of CH_2 and a molecular mass of 42 g/mol. Determine its molecular formula.
3. What information can you obtain from an empirical formula?
4. Calculate the percent composition of nitrogen in the compound that forms when 222.6 g N combines completely with 77.4 g O.
5. The compound methyl butanoate smells like apples. Its percent composition is 58.8% C, 9.8% H, and 31.4% O and its molar mass is 102 g/ mol. What is its empirical formula? What is its molecular formula?

Assignment 6.1**B**

Work required for credit on 1, 2, 3, 6, 7, 8.

____ 1. What is the molecular formula of a compound that has a molecular mass of 54 and the empirical formula C_2H_3?

A) C_8H_{12} B) C_6H_9 C) C_4H_6 D) C_2H_3

____ 2. A compound has the empirical formula CH_2O and a gram-formula mass of 60. grams per mole. What is the molecular formula of this compound?

A) C_3H_8O B) $C_4H_8O_4$ C) CH_2O D) $C_2H_4O_2$

____ 3. A compound has a molecular mass of 54 and an empirical formula of C_2H_3. What is the molecular formula of the compound?

A) C_5H_8 B) C_6H_{10} C) C_2H_3 D) C_4H_6

____ 4. The formula C_2H_4 can be classified as

A) a molecular formula, only
B) both a structural formula and an empirical formula
C) both a molecular formula and an empirical formula
D) a structural formula, only

Base your answers to questions **5** and **6** on the information below.

Naphthalene, a nonpolar substance that sublimes at room temperature, can be used to protect wool clothing from being eaten by moths.

5. Explain, in terms of *intermolecular forces*, why naphthalene sublimes.

6. The empirical formula for naphthalene is C_5H_4 and the molecular mass of naphthalene is 128 grams/mole. What is the molecular formula for naphthalene?

____

7. Given the compound $C_4H_{10}O_8$.

a Calculate the molar mass of the compound.

b Calculate the number of moles in 17.7 grams of the compound.

c What is the empirical formula for this compound?

8. A compound contains 46.7% nitrogen and 53.3% oxygen by mass. What is the empirical formula of the compound?

Topic 6.7 **Mole Ratios and Conversions using them. How can we convert between moles of elements and compounds in a given chemical reaction?**

Task: read the below information and fill in any missing information based on the reading.

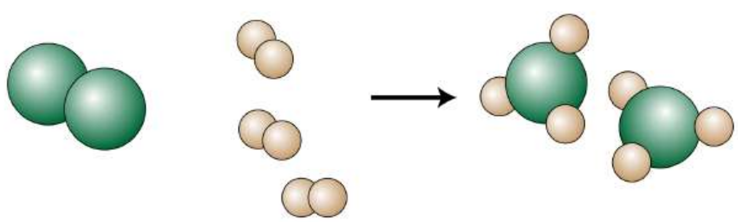
A. You want to add some sections to the porch seen to the right. Before you go to the hardware store to buy lumber, you need to determine the unit composition (the material between two large uprights). You count how many posts, how many boards, how many rails - then you decide how many sections you want to add before you calculate the amount of building material needed for your porch expansion. (note- this is just information)

B. Chemical equations express the amounts of reactants and products in a reaction. The coefficients of a balanced equation can represent either the number of molecules or the number of moles of each substance. The production of ammonia (NH3) from nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

Using the text above, what helps to represent the number of molecules or moles of each substance?______________

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

The balanced equation can be analyzed in several ways, as shown in **Figure** 1.2.



1 molecule $\mathbf{N_2}$	3 molecules $\mathbf{H_2}$	2 molecules $\mathbf{NH_3}$
1 mol $\mathbf{N_2}$	3 mol $\mathbf{H_2}$	2 mol $\mathbf{NH_3}$
28.02 g $\mathbf{N_2}$	3 x 2.02 g = 6.06 g $\mathbf{H_2}$	2 x 17.04 g = 34.08 g $\mathbf{NH_3}$

FIGURE 1.2
This representation of the production of ammonia from nitrogen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.

We see that 1 molecule of nitrogen reacts with 3 molecules of nitrogen to form 2 molecules of ammonia. **Put a box around these values in the above picture.** This is the smallest possible relative amounts of the reactants and products. To consider larger relative amounts, each coefficient can be multiplied by the same number. For example, 10 molecules of nitrogen would react with 30 molecules of hydrogen to produce 20 molecules of ammonia.

The most useful quantity for counting particles is the mole. So if each coefficient is multiplied by a mole, the balanced chemical equation tells us that 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia. This is the conventional way to interpret any balanced chemical equation. **Place a star next to these values in the above picture.**

Finally, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. 1 mol of nitrogen has a mass of 28.02 g, while 3 mol of hydrogen has a mass of 6.06 g, and 2 mol of ammonia has a mass of 34.08 g. **Place an equal sign in the correct location in the above picture to show that this statement is true.**

Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved. **Are the number of molecules conserved in the above equation (Y or N)?________________.**

A mole ratio is a conversion factor that relates the amounts in moles of any two substances in a chemical reaction. The numbers in a conversion factor come from the coefficients of the balanced chemical equation. The following six mole ratios can be written for the ammonia forming reaction above.

1 mol N_2:3 mol H_2	or	3 mol H_2:1 mol N_2
1 mol N_2:2 mol NH_3	or	2 mol NH_3:1 mol N_2
3 mol H_2:2 mol NH_3	or	2 mol NH_3:3 mol H_2

Note that each of these can be changed into a true "fraction" or used as a ratio as seen above. The Regents will ask you to determine mole ratios (usually with choices similar to above) AND convert using these ratios where you will need to change them to fractions.

Mole ratios summary:

- They are used to convert moles of one substance to _____________ of another substance in a reaction.
- This is determined by the values of the ________________________ in a BALANCED chemical equation.

Sample Problem: Mole Ratio

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen given the balanced equation?

$$3H_2 + N_2 \rightarrow 2NH_3$$

Step 1: List the known quantities and plan the problem.

Known

- given: H_2 = 4.20 mol

Unknown

- mol of NH_3

The conversion is from mol $H_2 \rightarrow NH_3$. The problem states that there is an excess of nitrogen, so we do not need to be concerned with any mole ratio involving N2. Underline these two substances in the chemical reaction and put your known under the "known" and an "x" under the "unknown."

$$\underset{4.20}{\underline{3H_2}} + N_2 \rightarrow \underset{x}{\underline{2NH_3}}$$

Step 2: Setup your ratios and solve.

$$\frac{3\ mol\ H_2}{4.20\ mol\ H_2} = \frac{2\ mol\ NH_3}{x\ mol\ NH_3}$$

Cross multiply and solve for x

3x=(4.20)(2)

$$\frac{3x}{3} = \frac{8.40}{3}$$

X=2.80 mol NH_3

Step 3: Think about your result. The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

Summary statement: Mole to mole calculations are one step calculations that ARE EASY using:

- ________________chemical equation AND - ____________________ ratios.

Another use for mole ratios as mole are represented by SUBSCRIPTS within chemical formuals

1. What is the mole ratio of H to N in the ammonia molecule? (Note this is another type of question that is asking about *inside* the chemical formula of NH_3).

2. The formula for ethanol is CH_3CH_2OH. What is the mole ratio of H to C in this molecule? Do a check… did you reduce it to the LOWEST ratio possible?

Using the knowledge you gained and the samples on the previous page, answer the following questions.

___NO(g) +___O_2(g) → ___NO_2(g)

_________N_________

_________O_________

1. Balance the equation above. (hint- you want 2 "N" on both sides)
2. What is the mole ratio of O_2 to NO_2? Be sure you put it into the correct order!
3. Use the balanced equation to determine the amount of NO_2 produced when 22.4 moles of O_2 reacts. Show your work (ratios and cross multiplying).

Teacher checkpoint

Practice with mole ratios and conversions:

Use the equation to answer the following questions: $C_4H_8 + 6\ O_2 \rightarrow 4H_2O + 4CO_2$

1. What is the mole ratio of carbon dioxide to oxygen in this reaction?
2. What is the mole ratio of butane (C_4H_8) to water in this reaction?
3. If you began this reaction with 5 moles of butane, how many moles of CO_2 would be produced? **Show your work!**

Use the equation: $3H_2 + N_2 \rightarrow 2NH_3$to answer the following questions.

4. How many moles of hydrogen are needed to react with 1 mole of nitrogen?
5. If you wanted to produce 50 moles of ammonia, how many mole of nitrogen would you need? **Show your work!**
6. To make the same 50 moles of ammonia, how many moles of hydrogen would you need? **Show your work!**
7. Which of your answers (5 or 6) would LIMIT the amount of ammonia you could actually make? This is considered your "limiting reactant."

Teacher checkpoint

Assignment 6.2**B**

Work required for credit on 2, 4, 5, 9b

____ 1. Given the balanced equation representing a reaction:

$Al_2(SO_4)_3 + 6NaOH \rightarrow 2Al(OH)_3 + 3Na_2SO_4$

The mole ratio of $NaOH$ to $Al(OH)_3$ is

A) 3:1 B) 3:7 C) 1:1 D) 1:3

____ 2. Given the balanced equation representing a reaction:
$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

What is the total number of moles of $O_2(g)$ required for the complete combustion of 1.5 moles of $C_3H_8(g)$?

A) .30 mol B) 1.5 mol
C) 4.5 mol D) 7.5 mol

____ 3. Given the balanced equation representing a reaction:
$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$
What is the mole ratio of $CO(g)$ to $CO_2(g)$ in this reaction?

A) 3:2 B) 1:1 C) 2:1 D) 1:2

____ 4. Given the balanced equation:

$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$
What is the total number of moles of $O_2(g)$ that must react completely with 5.00 moles of C_4H_{10} (g)?

A) 10.0 B) 20.0 C) 32.5 D) 26.5

____ 5. Given the balanced equation representing a reaction:

$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
What is the *minimum* number of moles of O_2 that are needed to completely react with 16 moles of NH_3?

A) 80. mol B) 16 mol
C) 64 mol D) 20. mol

____ 6. Which quantity of particles is correctly represented by the formula H_2SO_4?

A) 1.0 mole of ions
B) 1.0 mole of molecules
C) 6.0×10^{23} ions
D) 6.0×10^{23} atoms

____ 7. What is the total number of moles of atoms in one mole of $(NH_4)_2SO_4$?

A) 15 B) 14 C) 10 D) 11

____ 8. What is the total number of moles of oxygen atoms in 1 mole of ozone? (Molecular mass = 48)

A) 1 B) 2 C) 3 D) 4

9. Carbon disulfide is an important industrial solvent. It is prepared by the reaction shown below.

__C + __ SO_2 → __ CS_2 + _4_ CO

A. Balance the equation above. One coefficient has been done for you.

B. Calculate how many moles of CS_2 form when 6.3 moles of carbon react. Show all work.

Use this space to work out problems from your unit 6A test so you

a) know what you got wrong

b) don't make the same mistake again