Unit 5 (Intro to Chemical Quantities and Reactions) *Test Review Sheet*

Topic 1: Chemical Quantities

A **mole** is a unit that describes the quantity of 6.02×10^{23}. A mole is, therefore, a unit of quantity in the same sense that a dozen refers to the quantity of 12.

1 mole Na
6.02×10^{23}
atoms of Na

STP

1 mole CO_2
6.02×10^{23}
molecules of CO_2

Number of Particles = Given moles x (6.02×10^{23}) or **Moles** = $\dfrac{\text{Given number of particles}}{6.02 \times 10^{23}}$

The **molar mass** of a substance is the mass, in grams, of 1 mole of that substance. One mole of a substance contains 6.02×10^{23} particles (atoms, molecules, or ions) found in that substance.

For an example, water is composed of water molecules:
One mole of water contains 6.02×10^{23} (602000000000000000000000) molecules of water.

The molar mass of water, which is known to be 18 g, is the mass of 6.02×10^{23} molecules of water.

Below are the different variations of molar mass.

Atomic mass specifically refers to the mass of 1 mole of an element.
Formula mass is commonly used when referring to the mass of 1 mole of any substance.
Molecular mass is commonly used when referring to the mass of 1 mole of a molecular substance.

Regardless of the formula, the mass of one mole of a substance is the sum of the mass of all atoms in the formula. You are shown how to determine or calculate molar mass in the set below.

There a few different methods you can use to setup and calculate the formula mass of a given substance. Regardless of the method, the following three steps will be involved.

Step 1: Determine how many of each element is in the formula (*Be sure to count correctly)*
Step 2: Multiply the number of each element by the rounded atomic mass from Periodic Table
Step 3: Add up the total mass of all the elements in the formula to get the formula mass.

Error in any of these steps may result in incorrect formula mass.

Found to the top left of the element symbol on the Periodic Table

Example 1

What is the formula mass of $Al(OH)_3$?

Step 1: 1 Al 3 O 3 H
Step 2: 1(27) + 3(16) + 3(1) *setup*
Step 3: 27 + 48 + 3

Formula mass = **78 g/mole** *calculated result*

Example 2

What is the gram-formula mass of $NaNO_3 \cdot 4H_2O$?

Atoms	*Atomic Mass* x	*How Many* =	*Total Mass*
Na	23	1	23 g
N	14	1	14 g
H	1	8	8 g
O	16	7	112 g

You can also setup with a table like this

Formula mass = **157 g/mol**

Mole-Mass Relationship

The mass of one mole of a substance is the gram-formula mass of that substance. In other words, 6.02×10^{23} particles (one mole) of a given substance has a mass equal to the calculated gram-formula mass. What if there is more than one mole (more than 6.02×10^{23} particles of that substance)? It makes sense to think that a sample containing more than one mole of a substance (more than 6.02×10^{23} particles) will have a mass that is greater than the calculated gram-formula mass. Likewise, a sample containing less than one mole (fewer than 6.02×10^{23} particles) will have a mass that is less than the calculated gram-formula mass. The equation and examples below show how to calculate the mass of a substance if the number of moles of the substance is given.

$$\textbf{Moles} = \frac{\text{Given Mass}}{\text{Formula mass}}$$

What is the number of moles of zinc in a 130.8-gram sample?

Set up and solve using Table T equation above

$$\textbf{Moles} = \frac{130.8\text{ g Zn}}{65.4\text{ g/mol}} \qquad \textit{numerical setup}$$

$$\textbf{Moles} = \textbf{2.0 moles Zn} \qquad \textit{calculated result}$$

$$\textbf{Mass} = \text{Given moles} \times \text{Formula mass}$$

Table T Equation

What is the mass of 0.25 moles of O_2?

Set up and solve using the equation above

Mass = moles x Formula mass of O_2

Mass = 0.25 x 32 *numerical setup*

Mass = **8.0 g O_2** *calculated result*

Percent Composition by Mass- Use formula in Reference Table T

Use *Reference Table T* equation below to calculate percent composition.

Reference Table T equation

$$\textbf{Percent Composition} = \frac{\text{Mass of the element in a formula (part)}}{\text{Formula mass of the given formula (whole)}} \times 100$$

To determine percent composition of any element in a formula

Step 1: *Determine* the total mass of each element in the formula

Step 2: *Add up* **all** masses to get the formula mass

Step 3: *Divide* the mass of the element in question by the formula mass, then multiply by 100

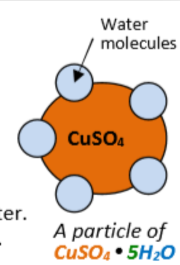
Ex: % comp by mass of Mg in magnesium hydroxide

$$\frac{\text{Mg}}{\text{Mg(OH)}_2} \quad \frac{24\text{ g}}{58\text{ g}} \times 100\% = 41.4\%$$

Hydrates are ionic compounds that contain water within their crystalline structures.

Names and formulas of three common hydrates are given below:

$CaCl_2 \bullet 2H_2O$	Calcium chloride dihydrate	(**di**hydrate means **2**H_2O)
$MgSO_4 \bullet 7H_2O$	Magnesium sulfate heptahydrate	(**hepta**hydrate means **7**H_2O)
$CuSO_4 \bullet 5H_2O$	Copper(II) sulfate pentahydrate	(**penta**hydrate means **5**H_2O)



Gram-formula mass (mass of 1 mole) of a hydrate is due in parts to the mass of the water. **Percent composition** of water in a hydrate can be calculated using the equation below.

$$\textbf{Percent H}_2\textbf{O in a hydrate} = \frac{\text{Total mass of H}_2\text{O (part)}}{\text{Formula mass of hydrate (whole)}} \times 100$$

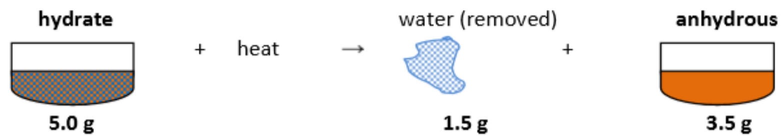
Reminder for Hydrates: $CuSO_4 * 5H_2O$. The dot does NOT mean multiplication. It means that for every copper sulfate, there are 5 water molecules. The gram formula mass of all the atoms in the hydrate are added.

In Lab

What is the percent composition of water in the hydrate $CuSO_4 \bullet 5H_2O$?
Answer: (90g/250g) x 100 = 36%

Percent Composition of Hydrates from Experimental Data

A **hydrate** can be heated to remove its water (by evaporation) in a laboratory experiment.
Anhydrous solid is what remains after a hydrate is heated to remove the water. If the masses of the hydrate and anhydrous solids are known, the mass of water that was in the hydrate can be determined. From this mass, the percent of water in the hydrate can be calculated.



$$\textbf{\% Water} = \frac{\text{Mass of water}}{\text{Mass of hydrate}} \times 100$$

$$\%\ \text{water} = \frac{1.5\ \text{g}}{5.0\ \text{g}} \times 100 = \mathbf{30\%}$$

Empirical/Molecular Formula

A **molecular formula** of a substance shows the true composition of a substance. For an example, water has a molecular formula of H_2O. This formula shows the true composition of water. The Molecular mass (mass of 1 mole) of water (18 grams) is the mass calculated from this formula.
The **empirical formula** of a substance shows atoms in a formula in their lowest ratio.

Name	Empirical Formula	Molecular Formula
Water	H_2O	H_2O
Hydrogen peroxide	HO	H_2O_2
Glucose	CH_2O	$C_6H_{12}O_6$
Dinitrogen monoxide	N_2O	N_2O
Caffeine	$C_4H_5N_2O$	$C_8H_{10}N_4O_2$

*If the Largest common multiple in a molecular formula is 1 (The formula cannot be further reduced), the molecular formula and the empirical formula are the same for that compound.

If the empirical formula and molecular mass of a substance are known, the molecular formula of the substance can be determined by following the steps below.

To determine molecular formula
Step 1: Determine the mass of the empirical formula
Step 2: Determine how many units of the empirical formula there are by dividing the given molecular mass by the calculated empirical mass from step 1
Step 3: Determine the molecular formula by multiplying each subscript of the empirical formula by step 2 answer.

Example 1

A substance has a molecular mass of 116 g and an empirical formula of C_2H_5. What is the molecular formula of this substance?

Step 1: Mass of C_2H_5 = 2 C + 5 H = 2(12) + 5(1) = 29 g

Step 2: $\dfrac{\text{molecular mass}}{\text{Empirical mass}} = \dfrac{116\ \text{g}}{29\ \text{g}} = \boldsymbol{4}$

Step 3: Molecular formula = ***4***(C_2H_5) = $\boxed{C_8H_{20}}$

Topic 2: Chemical Reactions

Chemical equation uses symbols to show changes in chemical compositions of substances during a chemical change. A **chemical reaction** is a way by which chemical changes occur.

Reactants are the starting substances that will go through the chemical change.
Reactants are shown to the *left* of the arrow in equations.

Products are the substances that remain after a change has occurred.
Products are shown to the *right* of the arrow in equations.

An **arrow** separates reactants from products, and can be read as "yields" or "produces."

A **coefficient** is a whole number in front of a substances to show the number of moles of the substance taking part in the reaction. Moles is a unit of quantity used to express amount of a substance .

If a **special condition** is required for the reaction (i.e. a catalyst, light, heat), it is placed above the yield arrow in the chemical equation

Phases (s), (l), (g), or (aq)- are typically included in the chemical formula for the substances

An example of a chemical equation is shown below.

Reactants | **Catalyst, Light, Heat, Etc.** | **Products**

$$Zn(s) + 2HCl(aq) \xrightarrow[\text{"Yields" or "Produces"}]{\text{Catalyst, Light, Heat, Etc.}} ZnCl_2(aq) + H_2(g)$$

Coefficient of 2: indicates that there are 2 moles of HCl

Coefficient of 2 or more is always written in front of the substance in equations.

Coefficient of 1: Indicates that there is 1 mole of $ZnCl_2$.

A coefficient of 1 is never written in front of the substance.

Balancing Chemical Equations

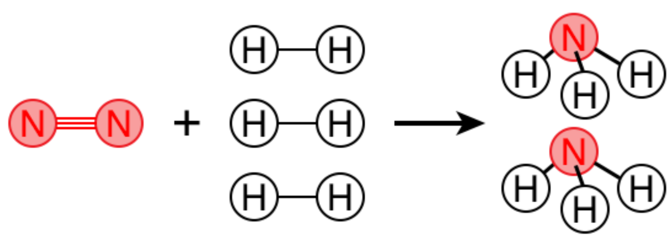
Law of Conservation states that during a chemical reaction neither atoms, mass, charges nor energy are created or destroyed. This means that during a chemical reaction, atoms, mass, charges, and energy are conserved so that their amounts are the same before and after the reaction.

A **balanced chemical** equation is a way of showing conservation in chemical reactions.

In the balanced equation below, atoms on the reactant and product sides are counted to show conservation.

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

2 N		2 N
6 H		6 H
atoms of reactants (before reaction)	=	**atoms of product** (after reaction)



The equation above is balanced because the correct combination of the *whole-number coefficients* allows the number of atoms on both sides of the equation to be equal. In other words, the substances in the reaction are represented in the correct proportion.

Note: When counting atoms in equations, always multiply the coefficient by the subscript.

Example: $2NH_3$

coefficient x subscript = 6 H atoms

Always write out the skeleton equation first, and then balance.

Ex:

Ammonia gas reacts with oxygen gas to produce gaseous nitrogen monoxide (nitric oxide) and gaseous water.

Skeleton Equation:

$NH_{3(g)} + O_{2(g)} \rightarrow NO_{(g)} + H_2O_{(g)}$

First **Last**

$2NH_{3(g)} + O_{2(g)} \rightarrow NO_{(g)} + 3H_2O_{(g)}$

Adjust the N on the product side now

$2NH_{3(g)} + O_{2(g)} \rightarrow 2NO_{(g)} + 3H_2O_{(g)}$

Since there are 2 O on the reactant side, and 5 O on the product side, we can balance them by multiplying the O_2 on the reactant side by 2.5, and then multiplying all of the coefficients by 2.

$(2NH_{3(g)} + 2.5\ O_{2(g)} \rightarrow 2NO_{(g)} + 3H_2O_{(g)})$ x2

$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$

"FML" Strategy for Balancing Chemical Equations

1. Figure out which elements to balance First, Middle, and Last. Always start by deciding which elements to balance last.
   - **Last**: any element that appears by itself (anywhere in the equation)
   - **First**: elements that appear in only one substance on each side.
   - **Middle**: every element that's not already last or first.
2. Start with any element on the "First" list. Add coefficients to make it balanced.
3. Pick another element. (Work your way through the "First," then "Middle," then "Last" lists.)
4. Repeat step #3 until everything is balanced.

Notes:

- If the same polyatomic ion is present on both sides of the equation, treat it like a "single unit" when balancing.
- If you end up with a fraction (i.e. ½, write it in temporarily, then multiply *all* of your coefficients by the denominator of the fraction to get back to whole numbers.

Stoichiometry – Mole Ratios

A chemical equation shows changes that are taking place in a chemical reaction. A balanced chemical equation is a recipe for changing one or more chemical substances to different substances.

Consider the balanced equation: $N_2 + 3H_2 \rightarrow 2NH_3$

This balanced equation reads as a recipe in the following way:

1 mole of nitrogen (N_2) is combined (or reacted) with **3** moles of hydrogen (**$3H_2$**) to yield (or make) **2** moles of ammonia (**2**NH_3).

A balanced chemical equation shows substances that are reacting and are being produced, as well as mole proportions (or mole ratios) of the substances in the reaction. Since mole ratios of substances in a given reaction is fixed, knowing the mole ratio of the reactants will allow you to make any amounts of the products by combining more or less of the reactants in the same ratio as given in the balanced equation.

To determine mole ratio of substances in an equation:

Step 1: Indicate the coefficients of the substances

Step 2: If the coefficients you indicated are reducible, be sure to reduce them by the Greatest Common Factor (GCF).

In the equation:

$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

Mole ratios (proportions) of substances in the equation are listed below .

Mole ratio of NH_3 to O_2 is **4 : 5**

Mole ratio of NH_3 to NO is **1 : 1**
(reduced from 4 : 4 by GCF of 4)

Mole ratio of NO to H_2O is **2 : 3**
(reduced from 4 : 6 by GCF of 2)

The example below shows you how to set up and solve mole-to-mole problems

Given the balanced equation below:

$2C + 3H_2 \rightarrow C_2H_6$

How many moles of C are needed to react with 12 moles of H_2?

Setup and solve using mole proportion

2 C	+	**3**H_2 → C_2H_6		*Re-write equation*
X		12		*Write moles from question*
$\frac{2}{X}$	=	$\frac{3}{12}$		*Setup by writing proportion*
3X	=	24		*Cross multiply*
X	**=**	**8 moles of C**		*Solve for X (calculated result)*

Note- Since scientists often measure the mass of substances in grams, not moles, it is typical to be given the starting mass of a substance in grams, and must first convert to moles before using the mole ratio.

Ex: Given the balanced equation below.

$2H_2 + O_2 \rightarrow 2H_2O$

If you start the reaction with 8 grams of O_2, how many moles of H_2O can you produce? How many grams of H_2O is this?

Step 1: $\boxed{\textbf{Moles} = \frac{\text{Given Mass}}{\text{Formula mass}}}$

$\frac{8g}{32\ g/mole}$ **= .25 moles O_2 reacted**

Step 2: $2H_2 + O_2 \rightarrow 2H_2O$

.25 x

$$\frac{1}{.25} = \frac{2}{x}$$

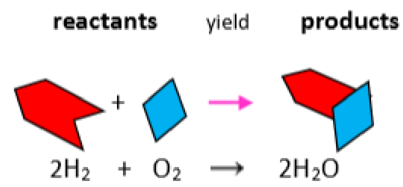
X= .5 moles H_2O produced

Step 3: $\boxed{\textbf{Mass} = \text{Given moles} \times \text{Formula mass}}$

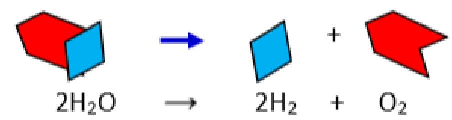
Table T Equation

.5 moles H_2O x 18 g/mole = 9 grams of H_2O produced

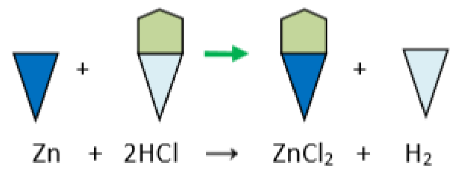
5 Main Types of Reactions



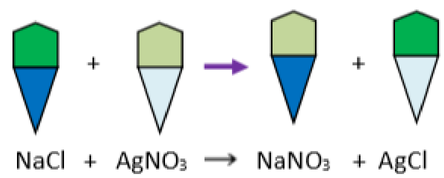
Synthesis reactions always involve two or more substances as reactants. During a synthesis reaction, the reactants combine to form one product.



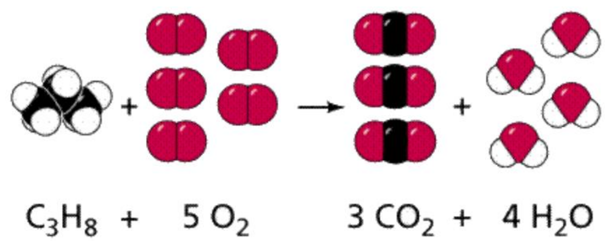
Decomposition reactions always involve a single substance as a reactant. During a decomposition reaction, the reactant breaks down (or decomposes) into two or more products.



Single replacement reactions likely involve a compound and a free element as reactants. During a single replacement reaction, the free element replaces one of the elements in the compound to form a different compound and a different free element. This reaction occurs spontaneously when the free element reactant is more reactive than the similar element in the compound it is replacing.



Double replacement reactions usually involve two aqueous solutions. During a double replacement reaction, the ions of the solutions switch with each other.



Combustion reactions typically involve the burning of an organic substance in the presence of oxygen. Water and carbon dioxide are usually the products in most combustion reactions