# Chapter 3: Stoichiometry:

Watch **Bozeman Videos** & other videos on my website for additional help:
Big Idea 1:
- Chemical Analysis
- Conservation of Atoms
- Balancing Equations
- Symbolic Representation
- Mole

Big Idea 3:
- Stoichiometry

## 3.1 Chemical Equations

The quantitative nature of chemical formulas and reactions is called **stoichiometry**.

**STOICHIOMETRY** - Involves using relationships between reactants and/or products in a chemical reaction to determine quantitative data.

- **One of the most important skills you can learn as you embark upon AP Chemistry!**
  - Get good at this and you will do well all year. This NEVER goes away!

**Antoine Lavoisier** (1743-1794)--**law of conservation of matter**: *matter can neither be created nor destroyed*

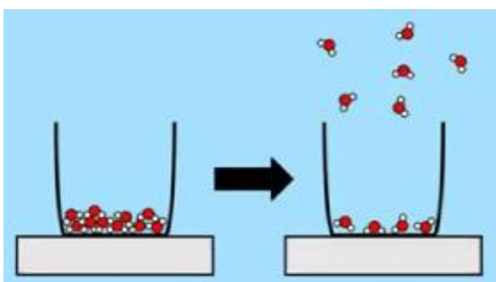
- this means "balancing equations" is all his fault!!

| Learning Objective | Essential Knowledge |
| --- | --- |
| **AP Unit 4.1**<br><br>**TRA-1.A** Identify evidence of chemical and physical changes in matter. | **TRA-1.A.1** A physical change occurs when a substance undergoes a change in properties but not a change in composition. Changes in the phase of a substance (solid, liquid, gas) or formation/ separation of mixtures of substances are common physical changes. |
| | **TRA-1.A.2** A chemical change occurs when substances are transformed into new substances, typically with different compositions. Production of heat or light, formation of a gas, formation of a precipitate, and/or color change provide possible evidence that a chemical change has occurred. |

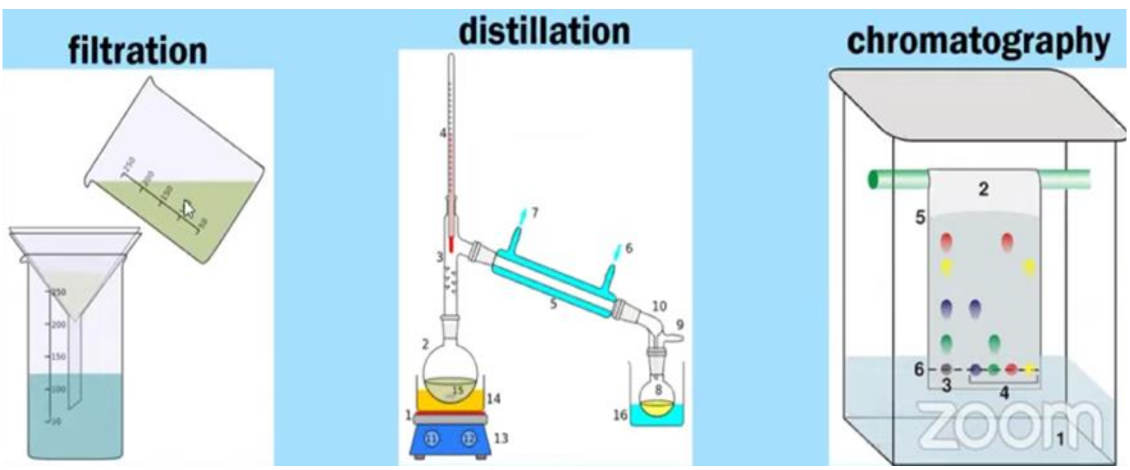
| | |
|---|---|
| **AP Unit 4.4**<br><br>**TRA-1.D** Explain the relationship between macroscopic characteristics and bond interactions for:<br>a. Chemical processes.<br>b. Physical processes. | **TRA-1.D.1** Processes that involve the breaking and/or formation of chemical bonds are typically classified as chemical processes. Processes that involve only changes in intermolecular interactions, such as phase changes, are typically classified as physical processes.<br>**TRA-1.D.2** Sometimes physical processes involve the breaking of chemical bonds. **For example, plausible arguments could be made for the dissolution of a salt in water, as either a physical or chemical process, involves breaking of ionic bonds, and the formation of ion-dipole interactions between ions and solvent.** |

**Physical/Chemical changes**

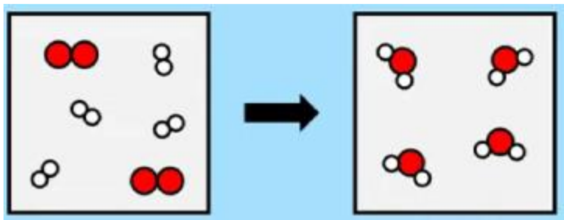
- In a **physical change -** occurs when a substance undergoes a change in properties, but not a change in composition (what it is made of).
  - **Intermolecular forces** (attraction BETWEEN molecules) are broken (more coming in Ch 11).
  - Examples are the following:
    - Phase change
      - $H_2O$ (s) → $H_2O$ (l)
      - $CH_3OH$ (l) → $CH_3OH$ (g)
      - $C_5H_{12}$ (l) → $C_5H_{12}$ (s)



- <u>Separating a Mixture</u>
  - **Filtration -** separating a solid from a heterogenous mixture
  - **Distillation -** separating a mixture of 2 substances with different boiling points
  - **Chromatography -** separate 2 or more substances of a <u>mixture</u> on the basis on how well they interact with a mobile phase and travel up the paper. Typically, the sample is suspended in the liquid or gas phase and is separated or identified based on how it flows through or around a liquid or solid phase.
  - **Dissolution** (dissolving a solid in $H_2O$)
    - The dissolving of a salt is typically considered a <u>physical change</u>. However, ionic bonds are broken, attractions between water and ion are formed, and heat is often <u>released</u>.



- A **<u>chemical change –</u>** occurs when substances are transformed into new substances, with different compositions.
    - Can be distinguished from a physical change because **<u>bonds, or INTRAMOLECULAR attraction,</u>** are broken and/or formed. (more coming in Ch 11)
    - Examples are the following:
        - $2H_2O\ (l) \rightarrow 2\ H_2\ (g) + O_2\ (g)$
        - $CO_2\ (g) + 2\ H_2\ (g) \rightarrow CH_3OH\ (g)$
        - $2\ Na\ (s) + 2\ H_2O\ (l) \rightarrow 2\ NaOH\ (aq) + H_2\ (g)$



**EVIDENCE A CHEMICAL REACTION HAS OCCURRED:**

- ✓ Formation of a **gas**(clearly not boiling). May result in an apparent change in mass.
- ✓ Formation of a **PRECIPITATE (solid forms), heat or energy** is absorbed or released
- ✓ **Color** Change

**Symbols of chemical equations**

(s) is for solids (including crystals & precipitates)

(g) is for gas

(*l*) is for pure liquids (ex: liquid water, melted compound)

(aq) is for solutions and acids (stands for aqueous)

→ "yield" symbol for reactions that proceed effectively 100% to form products.

$\rightleftharpoons$ *or* $\rightleftarrows$ "yield" symbol for reactions that proceed less than 100% to form products – these reactions are reversible

**Chemical equations** give a description of a chemical reaction.

- There are two parts to any equation:
    - **REACTANTS** (written to the left of the arrow)
    - **PRODUCTS** (written to the right of the arrow):

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

- There are two sets of numbers in a chemical equation:
    - BIG numbers in front = **COEFFICIENTS**
    - small numbers in the formulas = **SUBSCRIPTS**

- The subscripts give the ratio in which the atoms are found in the molecule.

Example:

- $H_2O$ means there are <u>two H **atoms**</u> for each one molecule of water.
- **2**$H_2O$ means that there are <u>two water **molecules**</u> present.

**Equations**

- Matter cannot be created or lost in any chemical reaction.
    - Therefore, we must *balance* the chemical equation.
    - When balancing a chemical equation, we adjust <u>**only the coefficients**</u>**.**
    - <u>Subscripts in a formula are ***NEVER*** changed when balancing an equation.</u>
        - Example: the reaction of methane with oxygen:

(USE ATOM INVENTORY TO BALANCE)

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

| Reactants | Products |
|---|---|
| C = 1 | C = 1 |
| H = 4 | H = 2 |
| O = 2 | O= 3 |

Hints to balance equations:

1. Balance **metals** first
2. Next balance **nonmetals**, except leave H and O
3. Balance H
4. Balance O last

$$CH_4 + \mathbf{2}O_2 \rightarrow CO_2 + \mathbf{2}H_2O$$

***AP EXAM TIPS***

**YOU SHOULD DOUBLE-CHECK YOUR ANSWER TO MAKE SURE THAT EACH CHEMICAL FORMULA HAS BEEN WRITTEN CORRECTLY AND THAT YOUR EQUATION HAS BEEN BALANCED CORRECTLY.**

**Indicating the States of Reactants and Products – <u>LEARN RULES IN CH 4</u>**

- The physical state of each reactant and product must be added to the equation:

$$CH_4(g) + 2O_2(g) \xrightarrow{\Delta} CO_2(g) + 2H_2O(l)$$

- ABOVE the reaction arrow - "Δ" is often used to indicate the addition of heat, but could also be ELECTRICITY.

**If the equations says SOLUTIONS, its aqueous (aq!!!)**

- ACID (H??) = aqueous
- Nonmetal + nonmetal = usually a gas
- INCLUDE S.O.M ON EVERYTHING!!!

**Examples**

**Write a Balanced Chemical Equation and include states of matters.**

1. Solid calcium hydroxide, $Ca(OH)_2(s)$, reacts with a solution of aqueous hydrobromic acid, $HBr(aq)$, to produce water and aqueous calcium bromide, $CaBr_2(aq)$.

2. Solutions of phosphoric acid, $H_3PO_4$, and potassium hydroxide, KOH, are combined, and the products of the reaction are water and potassium phosphate, $K_3PO_4$

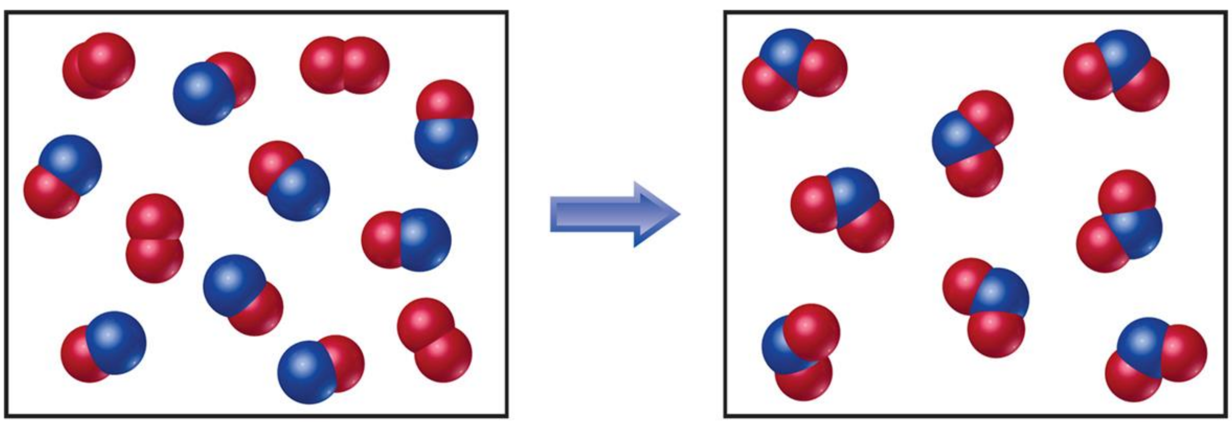
3. A sample of $C_5H_{12}O$ undergoes combustions by reacting completely with $O_2$ to produce $CO_2$ and $H_2O$.

4. Solutions of $Mg(NO_3)_2$ and $Na_3PO_4$ are combined. The products are $NaNO_3$ and $Mg_3(PO_4)_2$

| Learning Objective | Essential Knowledge |
|---|---|
| Unit 4.3<br><br>**TRA-1.C** Represent a given chemical reaction or physical process with a consistent particulate model. | **TRA-1.C.1** Balanced chemical equations in their various forms can be translated into symbolic particulate representations |

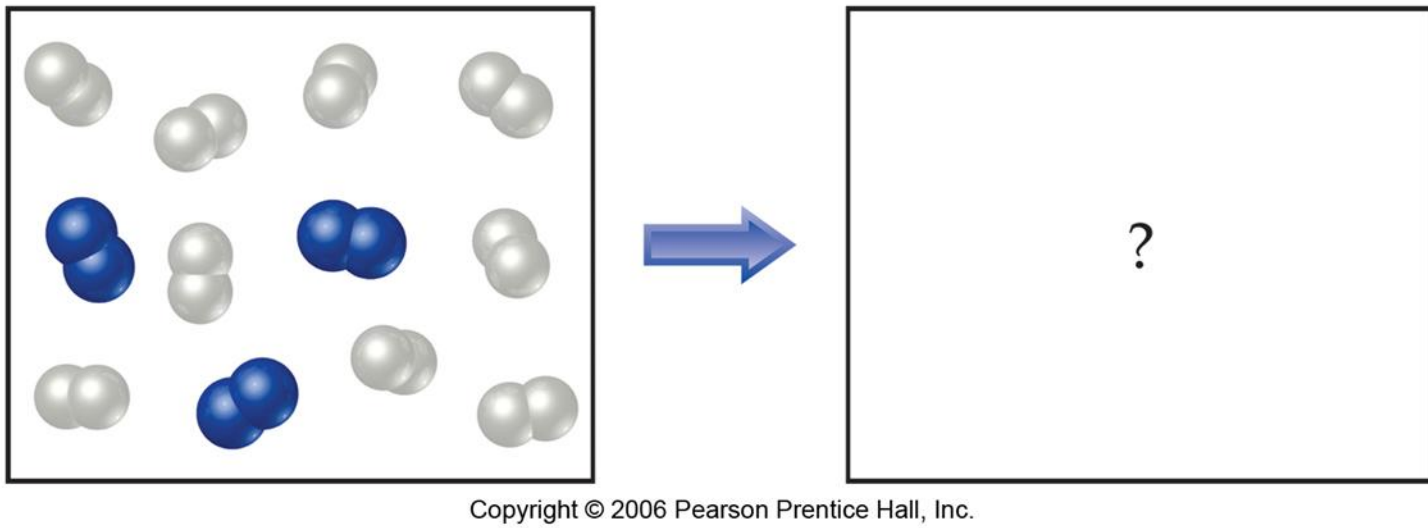
## SAMPLE EXERCISE 3.1 Interpreting and Balancing Chemical Equations

The following diagram represents a chemical reaction in which the red spheres are oxygen atoms and the blue spheres are nitrogen atoms. **(a)** Write the chemical formulas for the reactants and products. **(b)** Write a balanced equation for the reaction. **(c)** Is the diagram consistent with the law of conservation of mass?



# PRACTICE EXERCISE

**In order to be consistent with the law of conservation of mass, how many $NH_3$ molecules should be shown in the right box of the following diagram?**

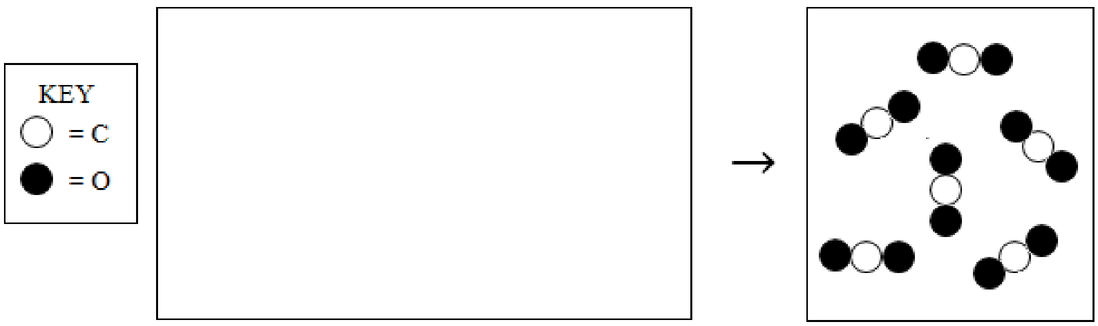


**Examples**

1. Consider the following unbalanced equation:

$$\ldots\ldots\ CO\ (g)\ +\ \ldots\ldots O_2\ (g) \rightarrow \ldots\ldots\ CO_2\ (g)$$

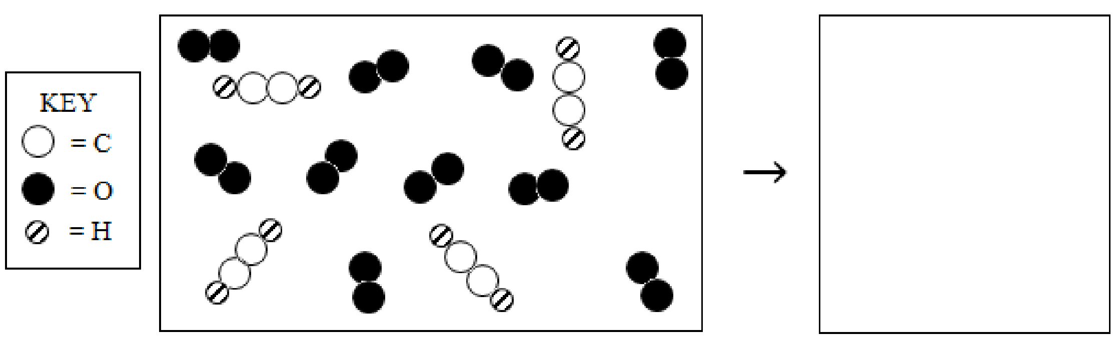
   a) Balance the chemical equation
   b) The particulate diagram represents the products of the reaction. In the other box, accurately depict the reaction mixture. Pay attention to number and types of atoms involved.




2. Consider the following unbalanced equation:

$$\ldots\ldots C_2H_2\ (g) + \ldots\ldots O_2\ (g) \rightarrow \ldots\ldots CO_2\ (g) + \ldots\ldots H_2O\ (g)$$

   a) Balance the equation
   b) The particulate diagram represents the reactants of the reaction. In the other box, accurately depict the products. Pay attention to number and types of atoms involved. Use  to represent water and  to represent carbon dioxide.



## 3.2 Some Simple Patterns of Chemical Reactivity – 5 Types of Reactions

1. C**ombination reactions -** two or or more substances react to form one product.

$$A + B \rightarrow C$$

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

2. D**ecomposition reactions -** one substance undergoes a reaction to produce two or more other substances. **Needs heat ALWAYS to break bonds… Δ not always written since it is implied that decomposition needs heat to break bonds!!**

$$A \rightarrow B + C$$

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

3. **Combustion reactions -** involve the reaction of **$O_2$**($g$) from air. (2 types of combustion)

$$A + O_2(g) \rightarrow A_xO_y$$

**or**

$$C_xH_y(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l) \quad \text{(ALWAYS PRODUCTS)}$$

**TIP FOR BALANCING COMBUSTIONREACTIONS: BALANCE CARBON FIRST, THEN HYDROGEN, THEN OXYGEN**

- **Combustion usually very exothermic**
  - a) metal + $O_2 \rightarrow M_xO_Y$ (metal oxide ; 1 molecule )
  - b) nonmetal + $O_2 \rightarrow Nm_xO_y$
  - c) $C_xH_y + O_2 \rightarrow CO_2 + H_2O$ (always produces this!)
  - d) $C_xH_yO + O_2 \rightarrow CO_2 + H_2O$ (harder to balance)

**Learn the other types of reactions in Ch.4**

## 3.3 Formula & Molecular Weights (AMU, not grams) – NEVER ROUND FROM P.T.

**Formula and Molecular Weights**

- **Formula weight** (FW) is the sum of atomic weights in an **IONIC** formula.
  - Example: FW (NaCl)
    - = 22.99 amu + 35.45 amu = 58.44 amu NaCl.

- **Molecular weight** (MW) is the sum of the atomic weights in a **MOLECULAR** formula.
  - Example: MW ($C_6H_{12}O_6$)
    - = 6(12.011 amu) + 12 (1.01 amu) + 6 (16.00 amu) = 180.19 amu $C_6H_{12}O_6$

**Molar Mass**

- The mass in grams of 1 mole of substance. (molar mass = 1 mole).
  - **Molar mass has units of g/mol**

- The molar mass of a molecule is the <u>sum of the masses of all atoms in the formula:</u>

  Example: The molar mass of $N_2$ = 2 x (molar mass of N) = 28.02 g/mol

**(Don't round masses on P.T.)**

**Formula Weight (amu) = Molecular Weight (amu) = Molar Mass (g/mol)**
**All calculated the same way!!!**

## **Percentage Composition from Formulas (Also been called Percent by Mass)**

In chemistry, percentage composition is based on **mass**, not on numbers of atoms present. We are going to find the percentage by mass of certain elements in a compound by comparing the mass of the element to the mass of the entire compound.

**Law of constant composition:** A sample of a **pure substance** will have a constant **percent by mass** for each element present in the substance, regardless of the size of the sample.

- Since this is an **intensive property** (constant regardless of the sample size), we will assume a total of one mole, therefore a total mass equivalent to the **molar mass** of the substance.

**Percentage composition** is obtained by dividing the mass of each element by the molar mass of the compound and multiplying by 100.

$$\%\ \text{element} = \frac{(\text{total mass of element in formula})}{(\text{molar mass of compound})} x100$$

**CHECK YOUR WORK - ALL PERCENTS MUST EQUAL 100% (+/- 0.5)**

**Examples**

1. What is the complete percentage composition (by mass) of all elements in potassium carbonate?

2. What is the % of oxygen in barium chlorate hexahydrate?

## 3.4 Avogadro's Number and The Mole

| Learning Objective | Essential Knowledge |
|---|---|
| **Unit 1.1**<br><br>**SPQ-1.A** Calculate quantities of a substance or its relative number of particles using dimensional analysis and the mole concept. | **SPQ-1.A.1 One cannot count particles directly while performing laboratory work. Thus, there must be a connection between the masses of substances reacting and the actual number of particles undergoing chemical changes.** |
| | **SPQ-1.A.2** Avogadro's number ($N_A = 6.022 \times 10^{23}$ $mol^{-1}$) provides the connection between the number of moles in a pure sample of a substance and the number of constituent particles (or formula units) of that substance. |
| | **SPQ-1.A.3** Expressing the mass of an individual atom or molecule in atomic mass units (amu) is useful because the average mass in amu of one particle (atom or molecule) or formula unit of a substance will always be numerically equal to the molar mass of that substance in grams. Thus, there is a quantitative connection between the mass of a substance and the number of particles that the substance contains.<br><br>EQN: $n\ (moles) = \frac{\text{m (mass)}}{\text{(M) Molar mass}}$ |

**Linking the world of atoms and molecules to the scale of a laboratory experiment!**

**The**  **is the of**  **Chemistry!!!**

**Moles (mol) =** standard number of particles (atoms, ions, molecules, formula units) and can be defined as, the amount of any substance that contains the same number of particles.

| |
|---|
| **1 mole** of something = $6.022 \times 10^{23}$ atoms, ions, molecules, or formula units. **(Avogadro's #)** |

**Representative particles:**

1. Element (except diatomics) = **atoms** (the smallest particle of an element which will retain the properties of that element.)
2. Diatomic elements = **Molecules**
3. Molecular (Covalent = non-metals only) compound = **Molecules**
4. Ionic compound (metal with nonmetal or polyatomic ion) = **Formula unit**

**Molar Mass is a general term that we use for the following three terms:**

**Atomic Mass Units** (amu) is used for **elements** and is found on the periodic table.

1 **atom** of Oxygen = 16.00amu

1 **mole** of Oxygen atoms = 16.00 g

**Molecular Mass** (grams) is used for **molecular (covalent) compounds** that exist as separate distinct molecules.

**Formula Mass** (grams) is used for **ionic compounds** since they do not exist as molecules but rather as crystals.

**Molar Mass From Data.**

$$MM = \frac{mass\ (g)}{moles\ (mol)}$$

**Examples**

1. If 47.4 grams of an ionic compound are known to be 1.185 moles of that compound, what is the **formula mass** of the compound?

**2.** If 256 grams of a pure monatomic element (meaning the element is NOT diatomic) is known to be 8.26 moles of that element, what **element** are we talking about**?**

**3.** How many **carbon atoms** are there in 16.0 grams of glucose, $C_6H_{12}O_6$?

**4.** If a sample of aluminum sulfite is known to contain 4.55 X $10^{23}$ **atoms of aluminum**, what is the mass of the entire sample of aluminum sulfite in grams?

**5.** How many **nitrogen atoms** are there in 45.8 grams of ammonium nitride?

**6.** How many **grams of oxygen** are present in 156.8 grams of calcium phosphate?

**7.** How many **gallium ions** are there in 65.0 grams of gallium sulfate?

**College board YouTube AP Chemistry: 1.1-1.4 Moles, Mass Spectrometry, Elemental Composition, and Mixtures**

WATCH THIS VIDEO BY SEAN BYRNE: Starting ~ 41 minutes

| Learning Objective | Essential Knowledge |
|---|---|
| **Unit 1.4**<br><br>**SPQ-2.B** Explain the quantitative relationship between the elemental composition by mass and the composition of substances in a mixture. | **SPQ-2.B.1** While pure substances contain molecules or formula units of a single type, mixtures contain molecules or formula units of two or more types, whose relative proportions can vary.<br><br>**SPQ-2.B.2** Elemental analysis can be used to determine the relative numbers of atoms in a substance and to determine its purity. |

**Example: A sample of carbonate rock is a mixture of CaCO3 and MgCO3. The rock is analyzed in a laboratory, and the results are recorded in the table below. Calculate the mole ratio of Ca to Mg in the rocks?**

| **Total Mass of sample** | **Mass of C in sample** | **Mass of Mg in sample** | **Mass of Ca in sample** | **Molar mass of C** | **Molar Mass of Mg** | **Molar Mass of Ca** |
|---|---|---|---|---|---|---|
| **98.5g** | **12.0g** | **2.4g** | **36.1g** | **12.0 g/mol** | **24.3 g/mol** | **40.1 g/mol** |

## 3.5 Empirical Formulas from Analyses

| Learning Objective | Essential Knowledge |
|---|---|
| **Unit 1.3**<br><br>**SPQ-2.A** Explain the quantitative relationship between the elemental composition by mass and the empirical formula of a pure substance. | **SPQ-2.A.1** Some pure substances are composed of individual molecules, while others consist of atoms or ions held together in fixed proportions as described by a formula unit. |
| | **SPQ-2.A.2** According to the law of definite proportions, the ratio of the masses of the constituent elements in any pure sample of that compound is always the same. |
| | **SPQ-2.A.3** The chemical formula that lists the lowest whole number ratio of atoms of the elements in a compound is the empirical formula. |

The **Law of Definite Proportions** states that compounds always contain the same relative mass of constituents, regardless of the source. What this means is the subscripts in a chemical formula are always the same for any given compound or molecule.

- So, basically, any sample of pure water contains 11.19% hydrogen and 88.81% oxygen by mass….no matter how much water is in the container!

- **Empirical formula:** lowest whole number ratio of atoms in a compound or molecule. (Example $CH_2O$ or $H_2O$)
- **Molecular formula:** larger whole number ratio of atoms in a compound or molecule
  - Each subscript in the empirical formula is multiplied by the EF to determine the molecular formula. (Example $C_6H_{12}O_6$)

**STEPS to solve empirical formula problems:**

1. **PERCENT TO MASS** – If you start with percent's, first check that all elements equal 100%.
   a. If they do (most of time it will), then you state "100% = 100g)
   b. Since % composition is an intensive property, it is independent of the amount present. Therefore, the easiest way to convert % to mass is to assume 100 grams. (example: 42% of 100 g = 42 g!)
   c. **If you are given MASS**, then move to step #2. No need to mess with percents.

2. **MASS TO MOLE -** Conversion! Be careful with sig figs to maintain your accuracy.

3. **DIVIDE BY SMALLEST** - Divide each elements moles by the least moles. The subscripts in an empirical formula represents the smallest WHOLE NUMBER ratio. By finding the smallest value of moles and dividing all of them by that number, we assure that our smallest possible subscript is one.

4. **MULTIPLY 'TIL WHOLE** - If all of the results in #3 are whole numbers, simply use these whole numbers as subscripts and write the empirical formula. If all of the results in #3 are NOT whole numbers, you must multiply ALL of the results in #3 by some number which will make everything a whole number. **"All for one, one for all! Whatever you do for one element, do for all elements"**

   If the decimal portion is:
   - .2 or .8 then multiply by 5
   - .25 or .75 then multiply by 4
   - .33 or .67 then multiply by 3
   - .5 then multiply by 2

**Examples**

1. A 200 gram sample of compound which contains only carbon, hydrogen and oxygen is found to contain 94.74 grams of carbon, 21.05 grams of hydrogen and 84.21 grams of oxygen. What is the empirical formula of the compound?

**Molecular Formula vs. Empirical Formula**

- The empirical formula (relative ratio of elements in the molecule) may not be the molecular formula (actual ratio of elements in the molecule).
- Example: ascorbic acid (vitamin C) has the empirical formula $C_3H_4O_3$.
- The molecular formula is $C_6H_8O_6$.

**STEPS to solve molecular formula problems:**

1. **Find the molar mass of the empirical formula, $MM_{EF}$.**
2. **Divide the given molecular mass of the entire molecule (MM molecule) by the molecular mass of the empirical formula ($MM_{EF}$). The result should be a whole number or very close to a whole number (If not, then you messed up somewhere along the way). This number indicates the number of empirical formula units (EFU) in the molecular formula.**
3. **Use this whole number result to multiply ALL of the subscripts of the empirical formula by and write the molecular formula.**

$$\#EFU = \frac{MM\ molecule}{MM\ of\ EF}$$

2. If the empirical formula for a carbohydrate is $CH_2O$ and its molecular mass is known to be 240 g/mole, what is the molecular formula for the carbohydrate?

3. Caffeine (MM = 194.19) was completely decomposed and found to contain 49.48% carbon, 5.20 % hydrogen, 28.86 % nitrogen, and the rest oxygen. Determine the empirical and molecular formulas for caffeine.

**College board YouTube AP Chemistry: 1.1-1.4 Moles, Mass Spectrometry, Elemental Composition, and Mixtures**

WATCH THIS VIDEO BY SEAN BYRNE

Answer shown in AP You tube video 1.5-1.8 (~8 minutes)

4. A 31 g sample of a compound that contains only the elements C, H and N is completely burned in $O_2$ to produce 44.0 g of $CO_2$, 45.0 g of $H_2O$, and 92.0 g of $NO_2$. Determine the empirical formula of the compound.

***Empirical Formulas – See video on YouTube under Chemistry with Doc Dena***
***Video: Empirical Formulas Combustion L2***

5. Oxalic acid is used in the paint, cosmetics, and ceramics industries and is found in many plants and vegetables. It contains only the elements C, H, and O. If 0.513 g of the acid is burned in oxygen, 0.501 g of $CO_2$ and 0.103 g of $H_2O$ result. (a) What is the empirical formula of oxalic acid? (b) If the molar mass of the acid is 90.04 g/mol, what is the molecular formula?

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

***Hydrate Formulas – See video on YouTube under Chemistry with Doc Dena***
***Video: Hydrate Formulas L1***

6. Sodium thiosulfate is used in photography as a fixer to dissolve silver halides. A 16.59 g sample of hydrated sodium thiosulfate ($Na_2S_2O_3 \bullet x\ H_2O$) is heated to drive off the water. The dry sample has a mass of 10.59 g of sodium thiosulfate. What is the formula for the hydrate?

## 3.6 Quantitative Information from Balanced Equations

| Learning Objective | Essential Knowledge |
|---|---|
| Unit 4.5<br><br>**SPQ-4.A** Explain changes in the amounts of reactants and products based on the balanced reaction equation for a chemical process. | **SPQ-4.A.1** Because atoms must be conserved during a chemical process, it is possible to calculate product amounts by using known reactant amounts, or to calculate reactant amounts given known product amounts. |
| | **SPQ-4.A.2** Coefficients of balanced chemical equations contain information regarding the proportionality of the amounts of substances involved in the reaction. These values can be used in chemical calculations involving the mole concept. |
| | **SPQ-4.A.3** Stoichiometric calculations can be combined with the ideal gas law and calculations involving molarity to quantitatively study gases and solutions. |

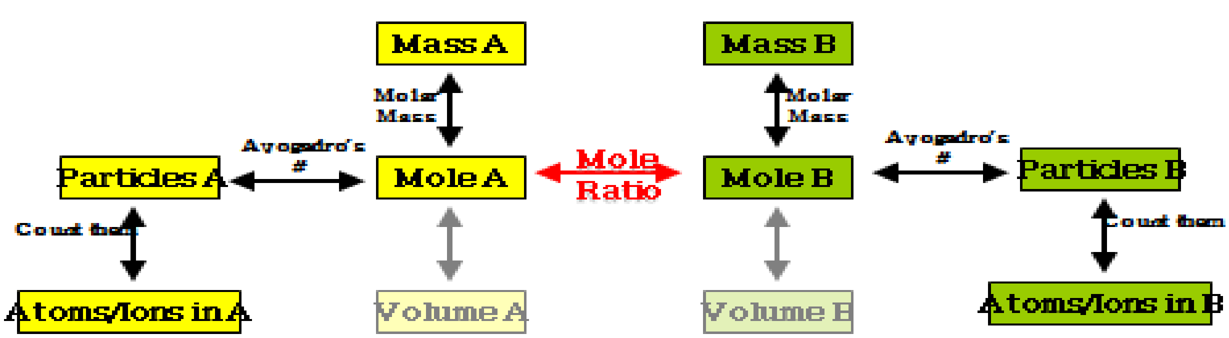
**Mole ratios** is the only way to convert between reactants and products in a reaction.
- The number of grams of reactant **cannot be *directly*** related to the number of grams of product. (can't just go grams to grams. Always have to use moles)

The balancing coefficients indicate the ratio of **MOLES** involved in the reaction.

**In order to use the mole ratio, you must have a BALANCED equation!!!!!! :]**

grams A → moles A → moles B → grams B

## Stoichiometry Conversion Chart




Step 1: Make sure you have a balanced equation!

Step 2: Locate the known (A) and unknown (B).

Step 3: Follow the chart!

**College board YouTube AP Chemistry: 1.1-1.4 Moles, Mass Spectrometry, Elemental Composition, and Mixtures**

WATCH [THIS VIDEO] BY SEAN BYRNE: Starting ~25minutes

A student is given 50.0 mL of a solution of $Na_2CO_3$ of unknown concentration. To determine the concentration of the solution, the student mixes the solution with excess 1.0 M $Ca(NO_3)_2$ (*aq*), causing a precipitate to form. The balanced equation for the reaction is shown below

$$Na_2CO_3(aq) + Ca(NO_3)_2(aq) \longrightarrow 2\ NaNO_3(aq) + CaCO_3(s)$$

The student filters and dries the precipitate of $CaCO_3$ (molar mass 100.1 g/mol) and records the data in the table below.

| | |
|---|---|
| Volume of $Na_2CO_3$ solution | 50.0 mL |
| Volume of 1.0 M $Ca(NO_3)_2$ added | 100.0 mL |
| Mass of $CaCO_3$ precipitate collected | 0.93 g |

Determine the number of moles of $Na_2CO_3$ in the original 50.0 mL of solution.

**College board YouTube AP Chemistry: 4.5-4.9 Stoichiometry, Titration, Acid-Base Reactions, and Redox Reactions** Starting ~11 minutes
WATCH THIS VIDEO BY MICHAEL FARABAUGH

**STOICHIOMETRY -** the following unbalanced chemical equation and question will hit on the following topics

$$\mathbf{Al}(s) + \mathbf{HCl}(aq) \longrightarrow \mathbf{AlCl_3}\ (aq) + \mathbf{H_2}(g)$$

**a. moles to moles**

Al(*s*) and HCl(*aq*) react together according to the chemical equation shown above. If 0.36 mol of $AlCl_3$ is produced in this reaction, how many moles of $H_2$ are also produced? (Answer: 0.54 mol $H_2$)

**b. grams to grams**

Al(*s*) and HCl(*aq*) react together according to the chemical equation shown above. How many grams of Al are required to produce 75 grams of $H_2$? Assume that HCl is added in excess. (Answer: 670g Al)

**Examples**

***Mass to Mass Stoich – See video on YouTube under Chemistry with Doc Dena Video: Mass to Mass Stoich L1***

**1.** How many grams of sodium chloride could **theoretically** be formed (also called the **THEORETICAL YIELD or MAXIMUM YIELD** – its all you can produce with what you are GIVEN**)** in a synthesis reaction in which 30.00 grams of chlorine gas are reacted with excess solid sodium?

**Reaction Stoich with liquids!!!**
**Use density (g/mL), NOT MOLAR VOLUME since molar volume can only be used for gases (Learning this in Ch 10)**

***RXN Stoich and Density – See video on YouTube under Chemistry with Doc Dena Video: RXN Stoich and Density L2***

2. How many liters of liquid bromine are needed to react completely with 400.0 grams of calcium iodide? The density of liquid bromine is 3.12 g/mL.

3. If 46.7 mL of mercury are added to <u>excess silver nitrate</u>, how many grams of mercury (I) nitrate ($Hg_2(NO_3)_2$) could theoretically be formed? The density of mercury is 13.6 g/mL.

## 3.7 Limiting Reactants – Muy Importante!!!

Ever notice how hot dogs are sold in packages of 10 while the buns come in packages of 8?? The bun is the limiting reactant and limits the hot dog production to 8 as well! The limiting reactant [or reagent] is the one consumed most entirely in the chemical reaction.

**Plan of attack in Chemistry**: One reactant is completely used up. Once it is consumed, it is converted into product and then stops the reaction from continuing. This reactant that gets used up first is called the **limiting reagent** and it is what determines the quantities of products that form.

So, the other reactant(s) are an **excess reagent**, which means that after the reaction has stopped you still have reactant left over that isn't being used (waste).

The **excess reactant** remains because there is nothing with which it can react........The reactant that produces the less amount of product is the **limiting reagent** IN OTHER WORDS; WHO STOPS THE REACTION/ RUNS OUT FIRST

**STEPS FOR L.R**

1) If have grams of reactant 1, go to grams of reactant 2 ( same goes for moles)
2) Determine what you need vs. what you have
3) Every calculation starts with a limiting reactant…..why!!!

**Examples**

**1.** The two non-metals, sulfur and chlorine, react according to the equation below.

$$S(s) + 3Cl_2(g) \rightarrow SCl_6(l)$$

If 202 g of Sulfur are allowed to react with 303 g of $Cl_2$ in the reaction above, which is the limiting reactant, how much product will be produced and what mass of the excess reactant will be left over?

**College board YouTube AP Chemistry: 4.5-4.9 Stoichiometry, Titration, Acid-Base Reactions, and Redox Reactions**

WATCH THIS VIDEO BY MICHAEL FARABAUGH (~20min)

**2.**

$$2\ Al(s) + 6\ HCl(aq) \longrightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$$

**limiting reactant and theoretical yield**

125 g of Al(*s*) reacts with 2.50 L of 3.20 *M*(*aq*) according to the chemical equation shown above.

❑ Which chemical (Al or HCl), is the limiting reactant?

❑ What is the theoretical yield of $H_2$ in units of grams?

**College board YouTube AP Chemistry: 1.1-1.4 Moles, Mass Spectrometry, Elemental Composition, and Mixtures**

WATCH THIS VIDEO BY SEAN BYRNE

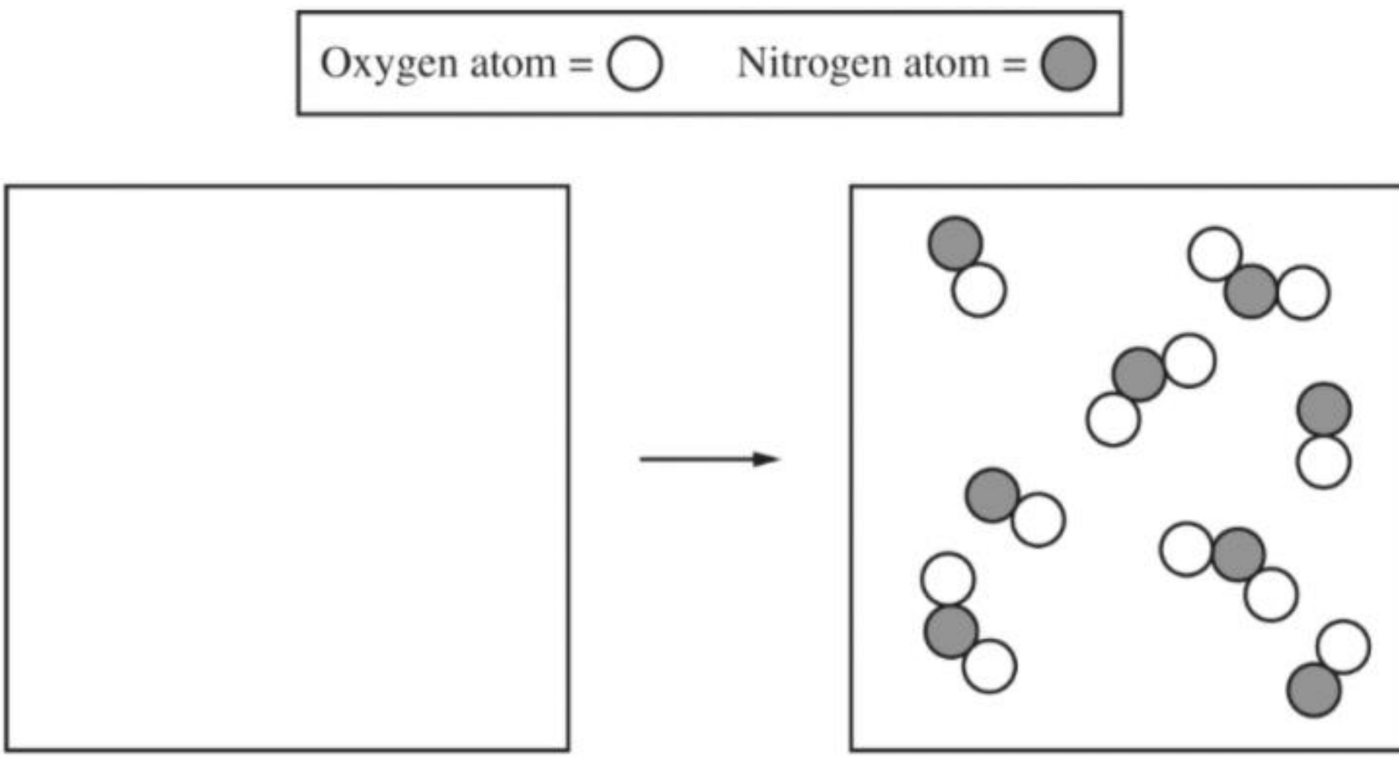
Answer shown in AP You tube video 1.5-1.8 (~11 minutes)

**3. 2018 FRQ**

$$2\ NO(g) + O_2(g) \longrightarrow 2\ NO_2(g)$$

A student investigates the reactions of nitrogen oxides. One of the reactions in the investigation requires an equimolar mixture of $NO(g)$ and $NO_2(g)$, which the student produces by using the reaction represented above.

(a) The particle-level representation of the equimolar mixture of $NO(g)$ and $NO_2(g)$ in the flask at the completion of the reaction between $NO(g)$ and $O_2(g)$ is shown below in the box on the right. In the box below on the left , draw the particle-level representation of the reactant mixture $NO(g)$ and $O_2(g)$ that would yield the product mixture shown in the box on the right. In your drawing, represent oxygen atoms and nitrogen atoms as indicated below.

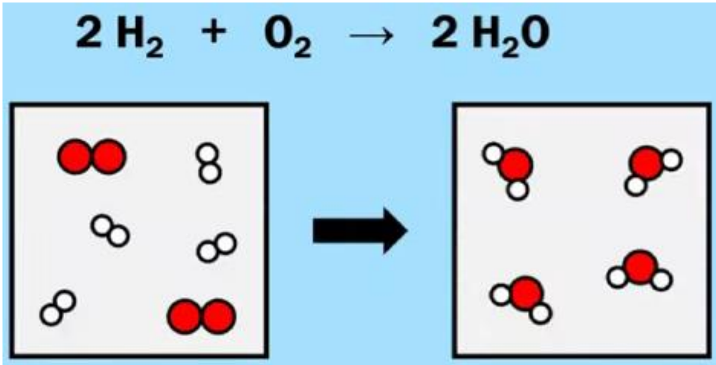



Reactant Mixture　　　　Product Mixture

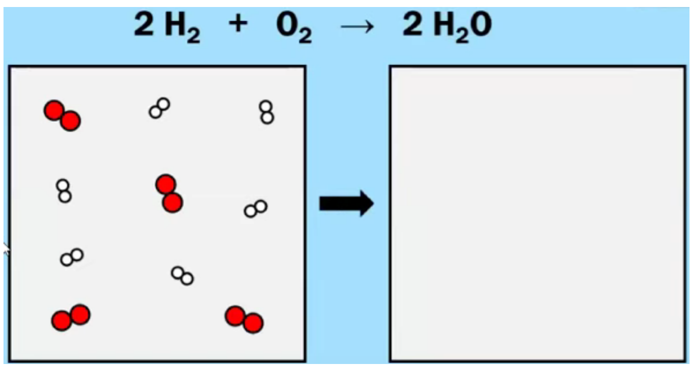
**College board YouTube AP Chemistry: 4.1-4.4**

WATCH THIS VIDEO BY MICHAEL FARABAUGH (~39min)

**4. Using Particle Diagrams to Represent Reactions**




**Particle Diagrams can also be used to illustrate the concept of limiting reactants and theoretical yield.**

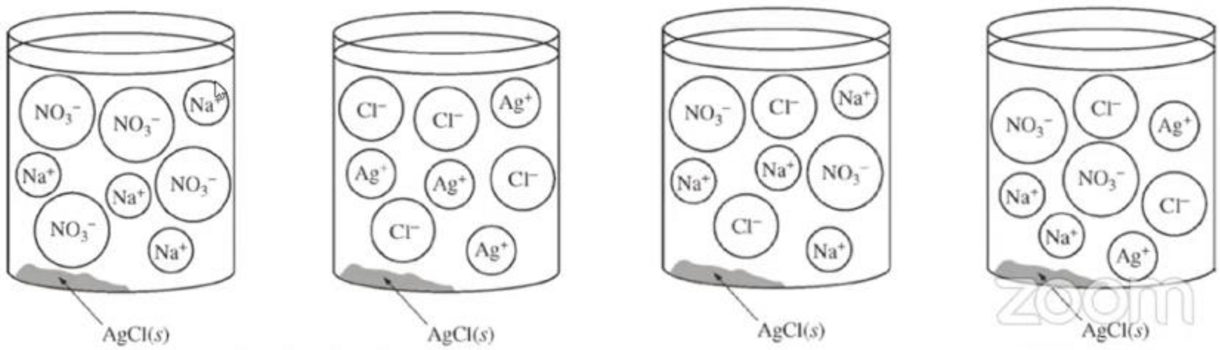


**College board YouTube AP Chemistry: 4.1-4.4 (~41.3min)**

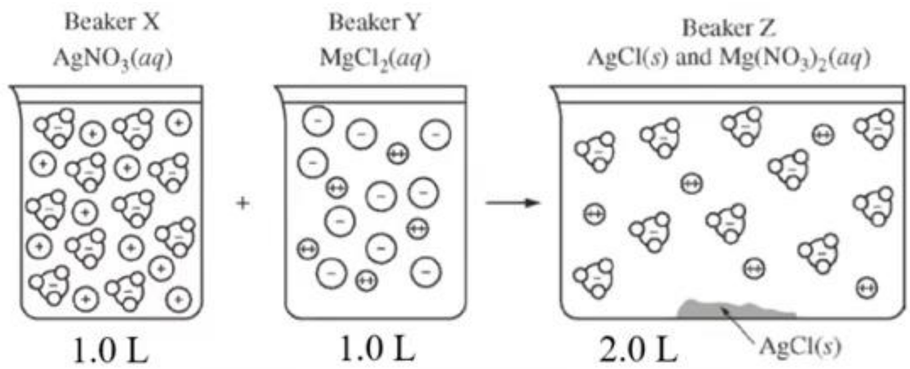
**Guided Practice - Using Particle Diagrams to Represent Reactions (Continued)**

$$\mathbf{AgNO_3\,(\mathit{aq}) + NaCl(\mathit{aq}) \longrightarrow NaNO_3(\mathit{aq}) + AgCl(\mathit{s})}$$

**5.** A dilute solution of $AgNO_3(aq)$ is combined with excess $NaCl(aq)$ to form $AgCl(s)$. Which of these diagrams best represents the ions that are present in significant concentrations in the solution? Justify your answer.



**College board YouTube AP Chemistry: 4.1-4.4 (~44.3min)**




**2 $AgNO_3$(*aq*) + $MgCl_2$(*aq*) ⟶ 2 AgCl(*s*) + $Mg(NO_3)_2$(*aq*)**

**6.** After examining the particle diagram shown above, a student made the claim that the concentration of $AgNO_3$(*aq*) is the same as the concentration of $MgCl_2$(*aq*).

Do you agree with the student's claim? Justify your answer based on the information in the particle diagram.

**7.** After examining the particle diagram shown above, a student made the claim that the value of [$Mg^{2+}$] in Beaker Y is twice as much as the value of [$Mg^{2+}$] in Beaker Z.

Do you agree with the student's claim? Justify your answer based on the information in the particle diagram.

**College board YouTube AP Chemistry: 4.1-4.4 (~48.3min)**

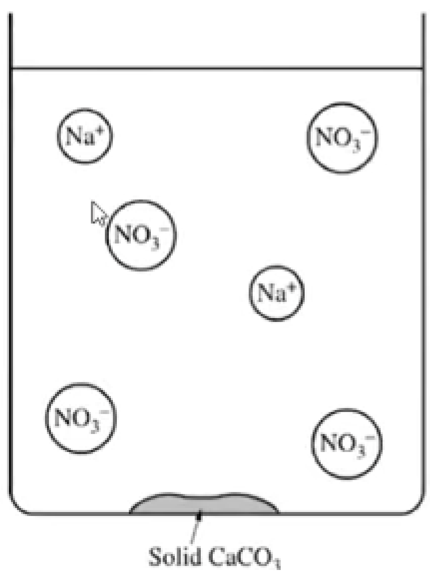
**8. Guided Practice - Using Particle Diagrams to Represent Reactions (Continued)**

A student is given 50.0 mL of a solution of $Na_2CO_3$ of unknown concentration. To determine the concentration of the solution, the student mixes the solution with excess 1.0 M $Ca(NO_3)_2$ (*aq*), causing a precipitate to form. The balanced equation for the reaction is shown below

$$\mathbf{Na_2CO_3(\mathit{aq}) + Ca(NO_3)_2(\mathit{aq}) \longrightarrow 2\ NaNO_3\ (\mathit{aq}) + CaCO_3(\mathit{s})}$$

(a) Write the net ionic equation for the reaction that occurs when the solutions of $Na_2CO_3$ and $Ca(NO_3)_2$ are mixed.

(b) The diagram below is incomplete. Draw in the species needed to accurately represent the major ionic species remaining in the solution after the reaction has been completed.



**Theoretical, Actual, & Percent Yields**

- The amount of <u>product you should have at the end</u> of the reaction (no errors) - **THEORETICAL yield**.
- The amount of product you <u>actually recovered</u> in the reaction - **ACTUAL yield**
- The **PERCENT yield** relates the actual yield to the theoretical yield:
- If your lab technique is perfect, you should get 100% percent yield ☺
- Bad data if it is less than 85%.
- You can't have over 100% - IMPOSSIBLE!!! YOU CAN'T CREATE MATTER!!!

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

The "**experimental**" or "**actual**" is how much you actually made. This value is determined in lab.

You will not always make all you were supposed to in a reaction! The % yield is the % that you got of what you were supposed to get. It is sometimes called the % efficiency.

$$\%\ \boldsymbol{yield} = \frac{\boldsymbol{experimental}}{\boldsymbol{theoretical}\ (\boldsymbol{maximum})} \times \mathbf{100}$$

The "**theoretical**" is the value that is determined from stoichiometry (a calculation). It is how much you should have gotten if everything went perfectly with your reaction.

**9.** If 32.00 grams of methane at STP are burned in a combustion reaction and 30.00 grams of water are experimentally produced, what is the % yield (or efficiency) of the reaction?

**10.** The reaction of combustion of octane in an engine is known to be only 84.3% efficient. How many grams of water would you actually expect to form if 95.00 grams of liquid octane are burned?

***Percent Yield – See video on YouTube under Chemistry with Doc Dena Video: Stoichiometry limiting 1***

**11.** 35.60 grams of zinc are reacted with 100.00 grams of iron(III) sulfate.

$$\mathbf{3Zn \quad + \quad Fe_2(SO_4)_3 \rightarrow \quad 3ZnSO_4 \quad + \quad 2Fe}$$

a. What is the limiting reactant? ___________________________________

b. What is the excess reactant? ___________________________________

c. How many grams of the zinc will be left over? _______________________

d. How many grams of the iron(III) sulfate will be left over? _______________

e. What is the theoretical yield (maximum yield) of pure iron in grams? _________________