

Chemistry 120

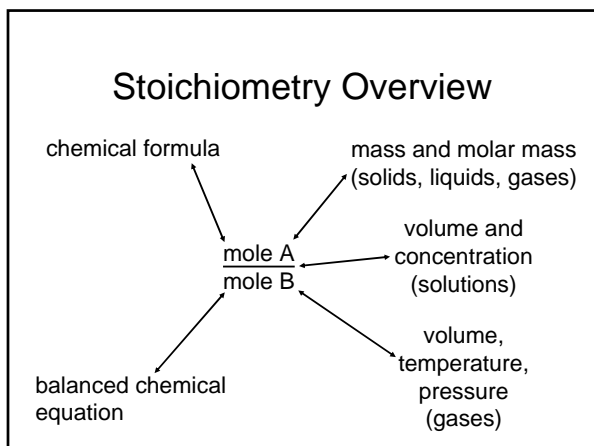
General Stoichiometry

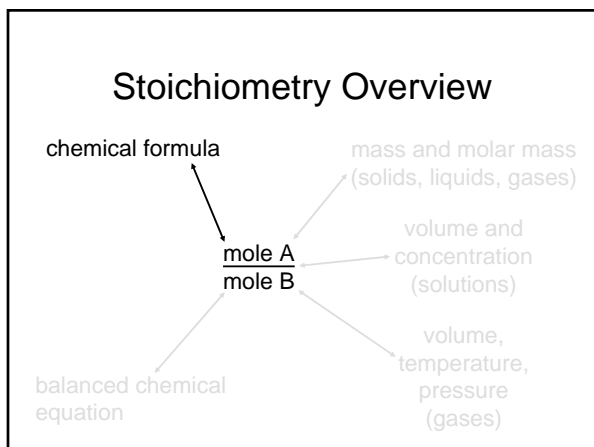
Chemical Reaction

- Transformation of a Set of Substances to another Set
 - Recipe for change is a *chemical equation*
Reactants \longrightarrow Products
- Some Reactions are named
 - Example: a *combustion reaction*
- Chemical Reactions rearrange Atoms
 - Can't measure individual particles
 - Connect atomic and laboratory scales

The Mole

- Define a *Mole* as 6.0221367×10^{23} things or $6.0221367 \times 10^{23} \text{ (things) \cdot mole}^{-1}$
 - Big dozen
- Relationship of Chemical Formula to Moles
- *Molar Mass*
 - Atomic weight for elements
 - *Molecular weight*
 - *Formula weight*





Connecting the Mole and the Chemical Formula

- Mole and Chemical Formula are basis for understanding Chemical Reactions
 - Convert mass to number of particles
- From a Chemical Formula can get
 - Molar mass of substance
 - Moles of element per mole substance
 - Calculate mass of an element in a given mass of substance (*weight percent*)

Connecting the Mole and the Chemical Formula

- How many moles of C are there in 1 mole $C_6H_{12}O_6$?
- What is the Molar Mass of $C_6H_{12}O_6$?
- How many Moles of $C_6H_{12}O_6$ are there in 45.0 g of $C_6H_{12}O_6$?
- How many Molecules of $C_6H_{12}O_6$ are there in 45.0 g of $C_6H_{12}O_6$?
- How many moles of C are there in 45.0 g $C_6H_{12}O_6$?

Percent Composition

- Can describe Compounds by *Percentage by Weight* of Elements present
 - Mass (g) of element in 100 g of compound
- Historical Context
- Modern Context
 - Determine chemical formula
 - Elemental analysis (*empirical formula*)
 - Mass spectroscopy
 - Other physical methods

Percent Composition

- What is the percent by weight of each element in $C_6H_{12}O_6$?

- Determine compound's molar mass
- Find mass of each element in 1 mole of compound.
- Divide each element's mass by compound's molar mass.
- Multiply by 100 to obtain percentages.

Percent Composition and Empirical Formulas

- An unknown orange compound is found to be 64.56% C and 5.42% H. If it contains only C, H, and Fe, what is the compound's empirical formula?

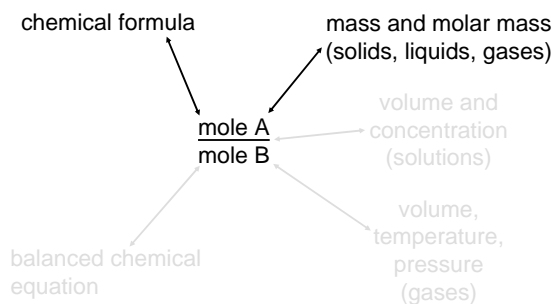
- Reverse of previous problem!
- Find mass of each element in 100 g of compound.
- Find moles of each element.
- Divide by smallest number of moles.
- Multiply by a whole number to clear fractions.

Percent Composition and Empirical Formulas

- If the molar mass of this compound is found to be $186.0 \text{ g}\cdot\text{mol}^{-1}$ by mass spectroscopy, what is the compound's chemical formula?

- Find the molar mass of the empirical formula.
- Divide molar mass by empirical molar mass.
- Result is number to multiply empirical formula by to get true molecular formula (usually a whole number).
- Multiply to obtain true molecular formula.

Stoichiometry Overview



**Determination of
H₂O in a Hydrate**

- Solid Compounds which have Water Incorporated into their Structure
 - Water can be removed by heating
- Named in Normal Way, specifying Number of Water Molecules
- Example: CaSO₄·2H₂O
Calcium sulfate dihydrate

Waters of hydration indicated with a “.”

When calculating a molar mass of a hydrate, must include the waters!

Hydrates

- Number of Waters of Hydration usually a Whole Number or Simple Fraction
 - Fractions named using *hemi-*
 - CuSO₄·2.5H₂O = copper(II) sulfate hemipentahydrate
- Compounds with no Waters of Hydration are *Anhydrous*
 - Anhydrous copper(II) sulfate = CuSO₄

Determining Waters of Hydration

- “Epsom salt”, MgSO₄·xH₂O will lose all of its waters of hydration upon heating. After heating a 1.687 g sample of MgSO₄·xH₂O, 0.824 g MgSO₄ remains. How many waters of hydration are in Epsom salt?

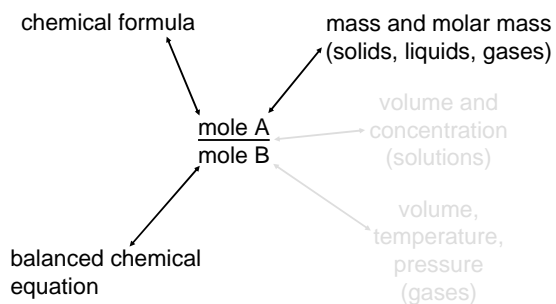
Using Mass of Products to determine Chemical Formula

- Chemical Equation is useful, but not necessary
 - Need to know relationship between products and reactant
 - Forms basis for most elemental analysis
- Worked exactly like previous Examples
- Key is Practice and keeping your Wits about You!

Using Mass of Products to determine Chemical Formula

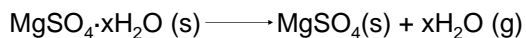
- A compound containing only Fe, C and O is burned in excess pure O_2 . Burning 1.959 g of the compound gives 0.799 g Fe_2O_3 and 2.200 g CO_2 , what is the % by mass of each element and what is the compound's empirical formula?

Stoichiometry Overview



Chemical Reactions and the Mole

- Epsom Salt Problem could have been written as a Chemical Reaction

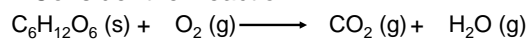


- Note Convention of Showing States
 - Solid = (s), or sometimes ↓
 - Liquid = (l)
 - Gas = (g), or sometimes ↑
 - Aqueous solution = (aq)

Chemical Reactions and the Mole

- Chemical Equations must be *Balanced*
 - Number of atoms (moles of each element) must be the same on both sides
 - Law of Conservation of Matter

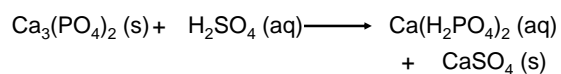
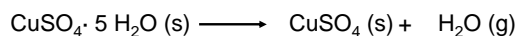
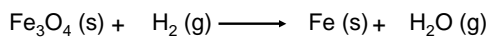
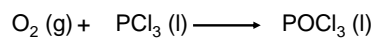
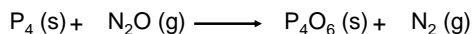
- Consider the Reaction



Balance by adding *stoichiometric coefficients*.

Do NOT change the identity of the products and/or reactants to balance the equation!

Practice

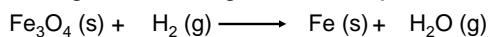


Stoichiometry

- Relationships between Quantities of Chemical Products and Reactants
 - Balanced chemical equation is a recipe for turning ingredients (reactants) into products
 - Mole is the key
- Chemical Equation tells us how many Particles and how many Moles react
 - Always convert grams to moles and back again

Using Stoichiometry

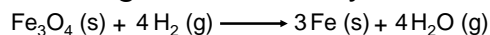
- If we have 25.00 g Fe_3O_4 and 1.00 g H_2 , how many grams of Fe can be obtained using the following chemical equation?



First step. Always check to make sure the equation is balanced.

Beware! Just because one reactant looks like it is present in excess, based on mass, don't be fooled. Only moles and stoichiometric coefficients matter.

Using Stoichiometry

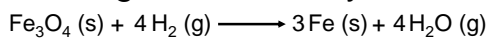


mass (g)	25.00	1.00	?	
molar mass (g/mole)	231.54	2.0158	55.847	Don't care about H_2O
moles	0.1079 ₇	0.496 ₀₈		

$$25.00 \text{ g Fe}_3\text{O}_4 \left(\frac{1 \text{ mole Fe}_3\text{O}_4}{231.54 \text{ g Fe}_3\text{O}_4} \right) = 0.1079_7 \text{ mole Fe}_3\text{O}_4$$

$$1.00 \text{ g H}_2 \left(\frac{1 \text{ mole H}_2}{2.0158 \text{ g H}_2} \right) = 0.496_{08} \text{ mole H}_2$$

Using Stoichiometry



mass (g)	25.00	1.00	?	
molar mass (g/mole)	231.54	2.0158	55.847	
moles	0.1079 ₇	0.496 ₀₈		

Limiting Reagent: the reactant which determines the amount of product formed.

Law of Conservation of Mass

Using Stoichiometry

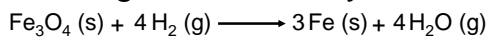
Which reactant gives least amount of product?

$$0.1079_7 \text{ mole Fe}_3\text{O}_4 \left(\frac{3 \text{ mole Fe}}{1 \text{ mole Fe}_3\text{O}_4} \right) = 0.3239_1 \text{ mole Fe}$$

$$0.496_{08} \text{ mole H}_2 \left(\frac{3 \text{ mole Fe}}{4 \text{ mole H}_2} \right) = 0.372_{06} \text{ mole H}_2$$

If all of the Fe_3O_4 reacts, get fewer moles of Fe than if all H_2 reacts. Fe_3O_4 is limiting reagent.

Using Stoichiometry



mass (g)	25.00	1.00	18.09	
molar mass (g/mole)	231.54	2.0158	55.847	
moles	0.1079 ₇	0.496 ₀₈	0.3239 ₁	

Use molar mass to convert moles Fe to grams Fe.

$$0.3239_1 \text{ mole Fe} \left(\frac{55.847 \text{ g Fe}}{1 \text{ mole Fe}} \right) = 18.09 \text{ g Fe}$$

This reaction can produce 18.09 g Fe.

Using Stoichiometry

- How much H₂, in grams, is left?

Calculate how many moles of H₂ reacted with Fe₃O₄.

$$0.1079_7 \text{ mole Fe}_3\text{O}_4 \left(\frac{4 \text{ mole H}_2}{1 \text{ mole Fe}_3\text{O}_4} \right) = 0.4318_8 \text{ mole H}_2$$

Calculate how many moles of H₂ left, then grams.

$$0.496_{08} \text{ mole H}_2 - 0.4318_8 \text{ mole H}_2 = 0.064_2 \text{ mole H}_2$$

$$0.064_2 \text{ mole H}_2 \left(\frac{2.0158 \text{ g H}_2}{1 \text{ mole H}_2} \right) = 0.13 \text{ g H}_2$$

There are 0.13 g H₂ are left.

Percent Yield

- Many Reactions do not give expected Amount of desired Products
 - Mechanical losses
 - Some reactions don't go to completion
 - Side reactions
- Define *Theoretical Yield* as the Amount of Product that Stoichiometry predicts
 - The 18.09 g Fe in the previous example is the theoretical yield for the reaction

Percent Yield

- Define a *Percent Yield* as follows

$$\% \text{ Yield} = \left(\frac{\text{mass of product obtained}}{\text{theoretical yield}} \right) \times 100$$

If 15.01 g Fe were obtained in the previous reaction, the percent yield would be

$$\% \text{ Yield} = \left(\frac{15.01 \text{ g Fe}}{18.09 \text{ g Fe}} \right) \times 100 = 83\%$$

Though more significant figures may be justified, % yields are usually reported as whole numbers.

Using Percent Yield

- If 30.0 g of Fe are needed, how much Fe_3O_4 , in grams, must we start with in this reaction? Assume that Fe_3O_4 is the limiting reagent and that the reaction goes in 82.97% yield.

Modified Limiting Reagent Problems

- Do not need to determine Limiting Reagent in all Stoichiometry Problems
- These Phrases are Clues
 - *what (mass) is needed to exactly react with*
 - *what (mass) is needed to completely convert*
 - *the complete reaction of*
 - *a stoichiometric amount*
 - *(present in) excess*

Example of a Modified-Limiting Reagent Problem

- Disulfur dichloride can be prepared according to the following reaction. If you begin with 5.23 g SCl_2 and excess NaF, what is the theoretical yield of S_2Cl_2 ?



“Excess” = no need to determine limiting reagent.

NaF is in excess (there is more than is needed).
So SCl_2 is limiting reagent.

Example of a Modified-Limiting Reagent Problem



mass (g)	5.23	xs		2.29	3
molar mass (g/mole)	102.971			135.037	
moles	1			2	

Calculate mass of SCl_2 in one step.

$$5.23 \text{ g SCl}_2 \left(\frac{1 \text{ mole SCl}_2}{102.971 \text{ g SCl}_2} \right) \left(\frac{1 \text{ mole S}_2\text{Cl}_2}{3 \text{ mole SCl}_2} \right) \left(\frac{135.037 \text{ g S}_2\text{Cl}_2}{1 \text{ mole S}_2\text{Cl}_2} \right) = 2.29 \text{ g S}_2\text{Cl}_2$$

Example of a Modified-Limiting Reagent Problem

- How much NaF was needed to exactly react with the SCl_2 ?



mass (g)	5.23 g	2.84 g	3		
molar mass (g/mole)	102.971	41.988			
moles	1	2			

$$5.23 \text{ g SCl}_2 \left(\frac{1 \text{ mole SCl}_2}{102.971 \text{ g SCl}_2} \right) \left(\frac{4 \text{ mole NaF}}{3 \text{ mole SCl}_2} \right) \left(\frac{41.988 \text{ g NaF}}{1 \text{ mole S}_2\text{Cl}_2} \right) = 2.84 \text{ g NaF}$$

Summary

- Connection between Atomic World and Macroscopic World is the Mole
- Balanced Chemical Equation
 - Recipe for changing matter
 - Establishes molar relationships between products and reactants
- Key to doing Understanding Stoichiometry and doing well is Practice, Practice, Practice!
