

STOICHIOMETRY**(I) Molecular Mass:**

- The sum of the masses of the atoms in the molecular formula.
- Ex: What is the molecular mass of propionic acid ($\text{CH}_3\text{CH}_2\text{COOH}$)?

(II) Formula Mass:

- The sum of the masses of the atoms or ions present in a formula unit.
- Ex: What is the formula mass of $\text{Ca}(\text{H}_2\text{PO}_4)_2$?

(III) The Mole:

- An amount of substance that contains as many elementary entities as there are atoms in exactly 12 g of C-12.
- Avogadro's number is the actual number of entities in 1 mole of substance:
 $6.022 \times 10^{23} \text{ mol}^{-1}$

(IV) Molar Mass:

- The mass of 1 mole of a particular substance.
- Is numerically equivalent to the atomic mass or molecular mass but is expressed in grams instead of amu.
- Ex: What is the molar mass of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ (sucrose)?
- **Mole problems:**
- Ex: 1. Calculate the number of moles in 15.2 g of H_2SO_4
2. Calculate the volume occupied by 1.38 mol of carbon tetrachloride, a liquid with a density of 1.59g/mL.
3. Calculate the weighted average mass of a calcium atom in grams.
4. Calculate the number of Br_2 molecules in 125 mL of liquid bromine ($d = 3.12\text{g/mL}$).

(V) Mass Percent Composition:**A) From chemical formulas:**

- Calculate the molar mass of the compound.
- For each element in the compound, divide the element's mass by the molar mass then multiply by 100 to obtain percentage.
- Check the percentages to ensure they add up to 100%
- Ex: 1. Find the mass percent of each element in urea ($\text{CO}(\text{NH}_2)_2$)
2. How many grams of nitrogen are in 46.34 g ammonium nitrate?

B) Chemical formulas from mass percent composition:**1. Empirical formulas (the simplest formula):**

- i) Assume the total sample mass is 100.0 g and calculate the percentages to grams.
- ii) For each element, use the molar mass to calculate the number of moles.
- iii) Divide each number of moles by the smallest number to attempt to get whole numbers.
- iv) Use the whole numbers as subscripts for the formula.

Ex: The diuretic metabutamate has the mass percent composition: 51.70% C, 8.68% H, 12.06 % N, and 27.55 % O. Determine its empirical formula.

2. Molecular formula (the actual formula):

- Calculate the empirical formula.
- Calculate the formula mass for the empirical formula.
- Divide the empirical formula mass by the molecular mass giving the integral factor.
- Multiply all subscripts in the empirical formula by the integral factor to give the molecular formula.

Ex: Ethylene has a molecular mass of 28.0 u, cyclohexane has a molecular mass of 84.0 u, and 1-pentene has a molecular mass of 70.0 u. All of them have an empirical formula of CH_2 , give the molecular formula for each compound.

C) Elemental analysis: Experimental determination of Mass Percent Composition.

Ex: A 0.3629 g sample of tetrahydrocannabinol, the principal active ingredient in marijuana, is burned in oxygen to yield 1.0666 g of carbon dioxide and 0.3120 g of water.

- Calculate its mass percent composition.
- Calculate its empirical formula.

Ex: An 0.0989 g sample of an alcohol is burned in oxygen to yield 0.2160 g CO_2 and 0.1194 g H_2O . Calculate the mass percent composition and the empirical formula for the compound.

(VI) Reaction Stoichiometry:

Ex's:

- How many grams of magnesium metal are required to convert 83.6 g TiCl_4 to titanium metal?
- Upon being heated or exposed to severe mechanical shock, ammonium nitrate decomposes into nitrogen and oxygen gases and water vapor. If 75.5 g of ammonium nitrate decomposes, how many grams of nitrogen and how many grams of oxygen are produced?
- How many milliliters of liquid water should be produced by the combustion in abundant oxygen of 775 mL of octane $\text{C}_8\text{H}_{18(l)}$? Assume that all volumes are measured at 20 °C where the densities are 0.7025 g/mL for octane and 0.9982 g/mL for water.

A) Limiting reactants:

- The reactant that is completely consumed in a chemical reaction and that limits the amount of product formed.
- Calculate the expected yield using each of the given reactant amounts.
- Whichever reactant gives the least product is the limiting reactant.

Ex: a) Iron (II) sulfide reacts with hydrochloric acid to give hydrogen sulfide and iron (II) chloride. If 10.2 g HCl is added to 13.2 g FeS , how many grams of H_2S can be formed? What is the mass of the excess reactant remaining?

B) Yields of Chemical reactions:

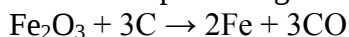
- Theoretical yield = the calculated quantity of product.
- Actual yield = the measured quantity of product.
- Percent yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

Ex: 1) In the reaction: $\text{Fe}_2\text{O}_{3(s)} + 3\text{CO}_{(g)} \rightarrow 2\text{Fe}_{(s)} + 3\text{CO}_{2(g)}$

a) If you start with 150 g of Fe_2O_3 as the limiting reagent, what is the theoretical yield of Fe?

b) If the actual yield was 87.9 g, what was the percentage yield?

2) An iron ore sample contains Fe_2O_3 and other impurities. A 752g sample of impure iron ore is heated with excess carbon producing 453 g of pure iron by the following reaction:



What is the mass percent of Fe_2O_3 in the original sample assuming that Fe_2O_3 is the only source of iron and the reaction goes to completion?

(VII) Solutions:

- Solute = the substance being dissolved (usually the lesser amount)
- Solvent = the substance doing the dissolving.

A) Molarity

- $M = \frac{\text{moles of solute}}{\text{liters of solution}}$
- Ex's: Calculate the molarity of
 - a) 3.00 mole of KI in 2.39 L of solution
 - b) 0.522 g HCl in 0.592 L of solution.
 - c) How many grams of potassium hydroxide are needed to prepare 10.0 mL of 0.100 M KOH?
 - d) A bottle of aqueous formic acid is 90.0% HCOOH by mass and has a density of 1.20 g/mL. Calculate the molarity of the solution.

B) Dilution of solutions:

- $M_1 \times V_1 = M_2 \times V_2$ (M=molarity V=volume)

Ex.'s:

- a) How many mL of a 12 M solution of HCl are required to make 250 mL of a 0.10 M solution?
- b) How many mL of 5.0 M $\text{K}_2\text{Cr}_2\text{O}_7$ must be diluted in order to prepare 250 mL of 0.10 molar solution?

C) Preparation of solutions:

- You should know and be able to describe how to prepare a solution of a particular molarity either from a solid solute or from a solution of known molarity.

Important points to remember:

- Always use the most accurate glassware available. This will usually be a volumetric flask and possibly a measuring cylinder, buret, or pipet.
- The volume is total volume (solute and solvent together). You will need to mix the solute and solvent in a volumetric flask then add water to get the correct total volume.
- You cannot assume that volumes are additive unless it is specified (50mL + 50 mL does not always equal 100 mL).
- Distilled or deionized water should always be used.
- If a strong acid is being diluted always add the acid to the water with plenty of swirling.