Chem 103, Section F0F
Unit IV - Stoichiometry of Formulas and Equations
Lecture 11

- The concept of a mole, which is a very large group of atoms or molecules
- Determining the formulas for a compound

Lecture 11 - Stoichiometry

Reading in Silberberg

- Chapter 3, Section 1 *The Mole*
- Chapter 3, Section 2 *Determining the Formula of an Unknown Compound*

Lecture 11 - Introduction

Stoichiometry is the study of the quantitative aspects of chemical formulas and chemical reactions.

- Using the tools of stoichiometry, you can predict the quantities of reactants and products that can be consumed or produced in a chemical reaction.
- These calculations will require working with *chemical formulas* and *balanced chemical reactions*.

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$$

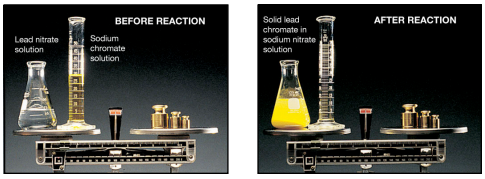
chemical formula for methane

balanced chemical equation for the combustion of methane

Lecture 11 - The Mole

We usually quantify objects either by counting them or weighing them.

- For liquids we also measure volumes.



Lecture 11 - The Mole

In chemistry we often need to deal with numbers of molecules

- For example, when working with balanced chemical equations.
  - For the combustion of methane
    - 1 molecule of methane reacts with
    - 2 molecules of oxygen to produce
    - 1 molecule of carbon dioxide and
    - 2 molecules of water.

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$$

balanced chemical equation for the combustion of methane

Lecture 11 - The Mole

The problem is, molecules are too small to count in the lab.

The concept of the mole allows us to count molecules by weighing them.

Lecture 11 - The Mole

Because any pure substance contains either identical atoms, molecules (covalent compounds) or formula units (ionic compounds) (see Dalton's postulates)

- a given mass of a pure substance will always contain the same number of either identical atoms, molecules (covalent compounds) or formula units (ionic compounds).
- For example, 10 g of methane contains 3.754 x 10^{23} methane molecules.
  - How did I know this?

Lecture 11 - The Mole

A mole (mol) is defined as

- The amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12.
- This number is called **Avagodro's Number** and is equal to 6.022 x 10^{23}

Alert!!! Alert!!!
Tomorrow is Mole Day

For example
- 1 mol of c n-12 atoms
- 1 mol of H_2O contains 6.022 x 10^{23} water molecules.
- 1 mol of NaCl 6.022 x 10^{23} NaCl formula units.

Lecture 11 - The Mole

A mole is more than just a number; in chemistry it has a special meaning:

- 1 mol of atoms for an element has a mass in grams that is *numerically equivalent* to the average mass of an atom of the element in amu's (atomic mass units).
  - 1 atom of C has an average mass of 12.01 amu
    - 1 mol of C atoms has a mass of 12.01 g
  - 1 atom of H has an average mass of 1.008 amu
    - 1 mol of H atoms has a mass of 1.008 g

Lecture 11 - The Mole

The mass of 1 mol of a substance made up of molecules or formula units can be calculating using the chemical formula to determine the numbers for each type of atom in a molecule of a substance:

- For example, 1 mol of methane (CH_4) has a mass of

$$\begin{aligned} 1 \times & \ 12.01\,\text{g of carbon (C)} \\ +\,4 \times & \ 1.008\,\text{g of hydrogen (H)} \\ = & \ 16.042\,\text{g of } CH_4 \end{aligned}$$

Lecture 11 - The Mole

Molar mass is defined as

- *the mass of mol of atoms, molecules or formula units of a substance*.
  - The units are g/mol
- For example, the molar mass of methane 16.042 g/mol

Lecture 11 - The Mole

Table 3.1 Summary of Mass Terminology*

Term	Definition	Unit
Isotopic mass	Mass of an isotope of an element	amu
Atomic mass (also called atomic weight)	Average of the masses of the naturally occurring isotopes of an element weighted according to their abundance	amu
Molecular (or formula) mass (also called molecular weight)	Sum of the atomic masses of the atoms (or ions) in a molecule (or formula unit)	amu
Molar mass ($\mathcal{M}$) (also called gram-molecular weight)	Mass of 1 mole of chemical entities (atoms, ions, molecules, formula units)	g/mol

*All terms based on the ^{12}C standard: 1 atomic mass unit = $\frac{1}{12}$ mass of one ^{12}C atom.

Lecture 11 - Clicker Question 1

Calculate the molar mass of BF_3

The molar mass of BF_3 is

A) 57.81 g/mol
B) 48.21 g/mol
C) 67.81 g/mol
D) 29.81 g/mol

Lecture 11 - The Mole

Calculating molar masses

- Elements
  - Determine whether the element exists as either individual atoms (metals and noble gases), or as molecules (nonmetals).
  - For individual atoms, the molar mass is numerically equal to the molecular mass of an atom in amu
  - For molecules multiply the molecular mass of one atom by the number of atoms in a molecule).

Lecture 11 - The Mole

Calculating molar masses

- Compounds
  - The molar mass is the sum of themolar mases of the atoms of the elements in the chemical formula.
  - For example, glucose ($C_6H_{12}O_6$):

Table 3.2 Information Contained in the Chemical Formula of Glucose, $C_6H_{12}O_6$ ($\mathcal{M}$ = 180.16 g/mol)

	Carbon (C)	Hydrogen (H)	Oxygen (O)
Atoms/molecule of compound	6 atoms	12 atoms	6 atoms
Moles of atoms/mole of compound	6 mol of atoms	12 mol of atoms	6 mol of atoms
Atoms/mole of compound	$6(6.022\times10^{23})$ atoms	$12(6.022\times10^{23})$ atoms	$6(6.022\times10^{23})$ atoms
Mass/molecule of compound	6(12.01 amu) = 72.06 amu	12(1.008 amu) = 12.10 amu	6(16.00 amu) = 96.00 amu
Mass/mole of compound	72.06 g	12.10 g	96.00 g

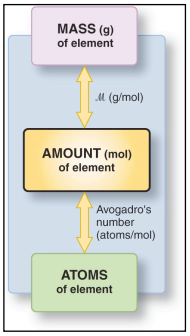
Lecture 11 - The Mole

The molar mass, along with Avogadro's number and the chemical formula, allows us to interconvert between

- the mass (g)
- number of mols of a substance
- number of atoms, molecules or formula units.

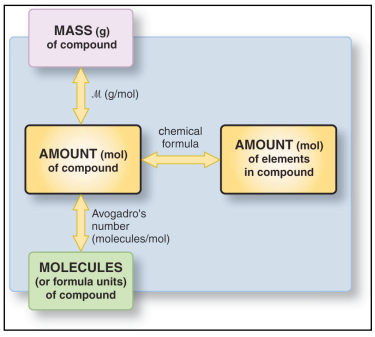
Lecture 11 - The Mole

- For elements:



Lecture 11 - The Mole

- For compounds:



Lecture 11 - The Mole

Mass Percent from the Chemical Formula

- The molecular or formula unit for a compound can be used to calculate the mass percent of an element in a quantity of a substance.

Lecture 11 - Question 2

Calculate the Mass % of sodium in sodium carbonate.

Lecture 11 - Formula of an Unknown Compound

Empirical Formulas

- There ways that a chemist can decompose a compound and determine experimentally (empirically) the mass percent of its constituent elements
- From these data, an **empirical formula** for a substance can be determined.
- The empirical formula gives the whole number ratio of the elements in a substance.
  - For an ionic compound, the formula unit and the empirical formula are the same thing.
  - For a covalent compound, the molecular formula is either equal to the empirical formula, or is an integer multiple of the empirical formula.

Lecture 11 - Formula of an Unknown Compound

The empirical formula is found by

- First dividing the the moles for each by the element with the smallest number of moles,
- multiplying these ratios by increasing integers: 1, 2, 3, ... until a near integer is obtained for all of the elements.

These integers are the subscripts for each element in the empirical formula.

- This is probably best demonstrated by example.

Lecture 11 - Question 3

Find the empirical formula of the following compound:

9.903 g of phosphorus (P) combines with 6.99 g of bromine (Br).

Lecture 11 - Formula of an Unknown Compound

Molecular Formulas

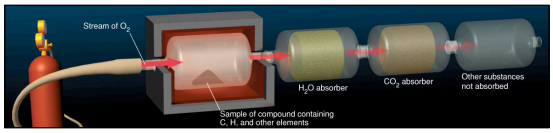
- The molar mass of a substance can be combined with the empirical formula to determine the molecular formula or formula unit for a substance.
  - Divide the molar mass by empirical molar mass
    - This should result in an integer
  - Multiply the subscripts in the empirical formula by this integer to get the molecular formula.

As an example, the empirical and molecular formulas can be determined for hydrocarbons by doing a combustion analysis.

Lecture 11 - Formula of an Unknown Compound

For example

- The empirical and molecular formulas can be determined for hydrocarbons by doing a combustion analysis.



Lecture 11 - Question 4

A dry-cleaning solvent ($\mathcal{M}$ = 146.99 g/mol) that contains C, H, and Cl is suspected to be a cancer-causing agent. When a 0.250-g sample was studied by combustion analysis, 0.451 g of CO_2 and 0.0617 g of H_2O formed.

Find the molecular formula for this solvent

Lecture 11 - Formula of an Unknown Compound

Chemical Formulas and Molecular Structures

- Some times different substances can share the same molecular formulas.
  - This is particularly prevalent with organic molecules, where the same elements can be connected in a multitude of ways.
  - Such substances are called **constitutional isomers**.

Table 3.4 Two Pairs of Constitutional Isomers

	C_4H_{10}		C_2H_6O	
Property	**Butane**	**2-Methylpropane**	**Ethanol**	**Dimethyl Ether**
$\mathcal{M}$ (g/mol)	58.12	58.12	46.07	46.07
Boiling point	−0.5°C	−11.6°C	78.5°C	−25°C
Density (at 20°C)	0.579 g/mL (gas)	0.549 g/mL	0.789 g/mL (liquid)	0.00195 g/mL (gas)
Structural formula				
Space-filling model				

Unit IV - Up Next

Unit IV - Chemical Bookkeeping: Stoichiometry

- Calculating the amounts of reactants consumed and the products formed in a chemical reaction
- Stoichiometry in solutions

The End