INTRODUCTION TO CHEMISTRY
(SC-230)

UNIT 7. MASS AND MOLE

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1. Introduction to Mass and Mole

- When you buy food in the shop you do it in different ways.
    1) You may count oranges.
    2) Buy a "package of eggs" e.g., a dozen.
    3) Buy bulk food like rice by mass (weight) e.g., 20lbs of rice – imagine having to buy rice by the number of grains!!!
- Chemists use these 3 methods too;
    1) Counting.
    2) Using a package that represents a quantity.
    3) Measuring mass.

- For a chemist it is relatively easy to weigh something to get a mass.
    - For example: A sample of copper sulfate that weighs 1.43 grams

- However what if I was to ask:
    How many <u>atoms</u> are there in the above sample of 1.43 grams of copper sulfate?
    or
    How many <u>molecules</u> are there in the above sample of 1.43 grams of copper sulfate?

    - This would be more difficult as there are too many to count.

- However a chemist wants to know this. For example, they will want to know how many atoms of carbon are there in 1 molecule of acetic acid? How many molecules of acetic acid can be obtained from 1 molecule of ethanol?

- To arrive to such calculations, chemists use the **atomic mass** of given elements. By itself, the atomic mass is a "relative mass" and it is derived from a point of reference called the Atomic Mass Unit. (amu).
    - **Atomic Mass Unit (amu)—***1/12 the mass of carbon-12 isotope (which implies that 1 carbon = 12 hydrogen)*
- Referring to the periodic table we can easily find the amu for a given element which is simply **<u>the atomic mass itself</u>**.
    - Ex: Oxygen has 16.0 as an atomic mass which means 16.0 amu
    - Ex: Sulfur has 32.0 as an atomic mass which means 32.0 amu
    - Ex: Nitrogen has 14.0 as an atomic mass which means 14.0 amu

- How do we apply this to compounds and what is the link with the "molar mass".

2. Calculating Molecular Mass and Formula Mass

- **Molecular Mass =** *The sum of the atomic masses of all the atoms in a molecule of the substance*.
- **Formula Mass** = *The sum of the atomic masses of all atoms in a formula unit of the compound, whether molecular or not.*
- **[Molecular and Formula Mass =** *sum of the atomic mass of all the elements composing the compound]*

Examples:

Molecular mass of H_2O = 18 amu **(amu = atomic mass unit)**

- H_2O Atomic mass of Hydrogen = 1 (2 X 1 = 2 amu)
  Atomic mass of Oxygen = 16 (1 X 16 = 16 amu)
  Total mass 18 amu

Molecular mass of H_2SO_4 = 98 amu

- H_2SO_4 Atomic mass of Hydrogen = 1 (2 X 1 = 2 amu)
  Atomic mass of Sulfur = 32 (1 X 32 = 32 amu)
  Atomic weight of Oxygen = 16 (4 X 16 = 64 amu)
  Total weight 98 amu

PRACTICE:

What is the molecular/formula mass of the following substances?

CO_2	
CH_4	
$CuSO_4$	
H_3PO_4	
$CaCO_3$	
Chloroform ($CHCl_3$)	
Iron (III) sulfate [$Fe_2(SO_4)_3$]	
H_2O_2	
HNO_3	

3. The Mole Concept

Why the mole?

- The mole allows chemists to count atoms, ions and molecules using mass.
- You need to understand the mole to apply stoichiometry.
  - **Stoichiometry ---***the calculation of the quantities of reactants and products involved in a chemical reaction.*
  - Stochiometry is based on the chemical equation and on the relationship between <u>mass</u> and <u>moles</u>.

- Before we explain what is a mole, we first must introduce "**Avogadro's Hypothesis**"
  - **Avogadro's Hypothesis --** *Equal volumes of gases at the same temperature and pressure have equal numbers of molecules*

- This led to the Avogadro's Number [N]
  - **Avogadro's number** – is the number of atoms in 12.00 grams of Carbon-12
  - The Avogadro's number is equal to the following: **($N = 6.022 \text{ X } 10^{23}$)**

What is a mole and its link to the Avogadro's number?

- **Mole** = the quantity of a given substance that contains as many molecules or formula units as the number of atoms in exactly 12g of carbon-12. In other words, it is "an Avogadro's number of anything.

1 mole = $6.022 \text{ X } 10^{23}$ of anything

- The term *mole*, is just a number like: like *dozen* or *pair* which refers to a particular number of things e.g.,

  A dozen eggs = 12 eggs
  A pair = 2
  A mole = 6.022×10^{23}

What is a molar mass?

- **Molar mass** – *is the numeric equivalent of the atomic mass (amu) in grams*.
  - Thus the **molar mass** is the atomic mass in grams.
- In turn, a molar mass is the mass of 6.022×10^{23} particles.

<u>Example of molar mass of elements:</u>

• Hydrogen (H)	1.01 amu;	1.01 grams	1 mol of H	$6.022 \text{ X } 10^{23}$ atoms
• Carbon (C)	12.01 amu	12.01 grams	1 mol of C	$6.022 \text{ X } 10^{23}$ atoms

<u>Example of molar mass of compounds/molecules:</u>

- ***A mol of a compound is the numeric equivalent of the atomic masses but in grams.***

- Ethanol has a molecular formula of C_2H_5OH , the molar mass of the compound will be:
  - C: (2x12=24); H: (5x1=5); O: (1x16=16); H: (1x1=1) : Total: 46amu = 46 g = 1 Mol C_2H_5OH = $6.22 \text{ X } 10^{23}$ molecules

- Carbon dioxide has a molecular formula of CO_2 ; its molar mass will be:
  - C: (1x12=12); O: (2x16=32) Total: 44amu = 44 g = 1 Mol CO_2 = $6.22 \text{ X } 10^{23}$ molecules

<u>Rules</u>:
THE MOLAR MASS (g/mol) IS NUMERICALLY EQUAL TO THE FORMULA/MOLECULAR MASS IN ATOMIC MASS UNITS (amu)

NOTE:
You need to respect the following steps:

Grams <=> moles <=> particles

CONVERSION:
Let us use the following example:

$$N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$$

1 molecule N_2	+	3 molecules H_2	$\rightarrow$ 2 molecules NH_3	**Molecular interpretation**
1 mol N_2	+	3 mols H_2	$\rightarrow$ 2 mols NH_3	**Molar interpretation**
28.0 g N_2	+	3 X 2.02g H_2	$\rightarrow$ 2 X 17.0 g NH_3	**Mass interpretation**

- Do not forget that 1 mol of anything = 6.022×10^{23} particles
    - Particles can be : (i) atoms ; (ii) ions; (iii) molecules; (iv) formula units
- With this information on hand, we can go from moles to grams and grams to moles and moles to number of particles and vice versa.

Examples:

How many atoms are in 1.5 moles of gold (Au)

1.5 mol X 6.022×10^{23} atoms / mol = 9.03×10^{23} atoms of Au

How many moles are in 29.58 g of Carbon (C)?

C = 12.01 g/mole

29.58 g X 1 mol C/12.01g = 2.46 mol of Carbon

How many atoms are in 5.21 g of sulfur (S)?

S = 32.07 g/mol

5.21 g S X 1mol S/32.07 g X 6.022×10^{23} atoms/mol = 9.78×10^{22} atoms S

How many moles are in 15 g of lithium (Li)?

Molar mass of lithium = 7 g/mol

15 ~~grams~~ X 1 mole/7 ~~grams~~ = 2.14 moles lithium

How many grams are in 2.4 moles of sulfur (S)?

S = 32.07 g/mol

2.4 ~~moles~~ X 32 grams/1 ~~mole~~ = 76.8 grams Sulfur

How many grams are in 4.5 moles of sodium fluoride (NaF)?

Molar Mass of NaF : is 23 + 19 = 42 g/mole

4.5 ~~moles~~ X 42 grams/1 ~~mole~~ = 189 grams NaF

How many moles are in 98.3 grams of aluminum hydroxide, $Al(OH)_3$?

Molar mass of aluminum hydroxide is 27 + (3 x 16) + (3 x 1) = 78 g/mole

98.3 ~~grams~~ X 1 mole/78 ~~grams~~ = 1.26 moles $Al(OH)_3$

RECAPITULATION

1 mole = 6.02×10^{23} particles
1 mole = molar mass (could be atomic mass from periodic table or molecular mass)

SUPPLEMENTAL MATERIAL – Mr. Causey YouTube Videos:

Understanding the Use of the Mole in Stoichiometry
With Mr. Causey
Reference: https://www.youtube.com/watch?v=yIwFqIt3NQo

Today we are going to talk about the Mole. What is important about the mole is that you learn what it can do. You need to understand the mole because it is:

- A big part of what we do in stoichiometry and much of the advanced chemistry

You will need:

- A Periodic Table
- A Calculator

Assumptions:

- You are familiar with the periodic table
- You know how to use a calculator
- You can perform unit conversion
- You are familiar with atomic masses

Why the mole?

- The mole allows chemists to count atoms, ions, and molecules using mass.
- In other words, via the mass, one can calculate the number of atoms or molecules. This is because, atoms and molecules are really small and they cannot be counted individually.

What is a mole?

- A mole is an Avogadro's number of anything.
- 1 mole = Avogadro's number = 6.022×10^{23} of anything
- One mole is just a number like:
  - 1 dozen = 12
  - 1 pair = 2
  - 1 mole = 6.022×10^{23} (it is a huge number because atoms, molecules… are very small)

What is a molar mass?

- A mole of an element is the numeric equivalent of the atomic mass but in grams.
  - So the molar mass is the atomic mass in grams
  - A molar mass is the mass of 6.022×10^{23} particles (no matter what the particle is – ions; atoms; molecules; formula units)

Example:

Hydrogen atom

H = 1.01 amu 1 mole H = 1.01 g

Carbon atom

C = 12.01 amu 1 mole C = 12.01 g

Molar mass of a compound

- A mole of a compound is the numeric equivalent of the atomic mass of the compound but in grams.
  - So the molar mass is the atomic mass in grams

Example:

Water (H_2O) = (2X1.01 = 2.02 amu) + (1 x 16.0 = 16amu) = 18.02 amu ➔ 18.02 g
Therefore: 1 mole of Water = 18.02 g

Carbon dioxide (CO_2) = (1x12=12 amu) + (2 x 16 = 32 amu) = 44 amu ➔ 44 g
Therefore: 1 mole of Carbon dioxide = 44 g

Let's look at Conversions:
1 mole / 6.022 X 10^{23} particles (atoms; ions; molecules; formula units…) or
6.022 X 10^{23} particles / 1 mole

Example

How many atoms are in 1.5 moles of gold (Au)?

1.5 ~~moles~~ X 6.022 x 10^{23} atoms / 1 ~~mole~~ = 9.03 X 10^{23} atoms of Au

How many moles are in 29.58 g of Carbon (C)?

29.58 ~~g C~~ X 1 mole C/ 12.01 ~~g C~~ = 2.46 moles of C

NOTE:
You need to follow the following steps:

Grams <=> moles <=> particles

Example

How many atoms are in 5.21 g of sulfur (S)?

5.21 ~~g S~~ X 1 ~~mole S~~ / 32.07 ~~g S~~ X 6.022 x 10^{23} atoms / 1 ~~mole S~~ = 9.78 x 1023 atoms S

RECAP:

- A mole is just a number like a dozen which is 12
- A mole = 6.022 x 10^{23} of anything
- Molar mass is the numeric equivalent of the atomic mass in grams.

OTHER YOUTUBE VIDEOS BY Mr. Causey

Avogadro's Number, the Mole Concept and How to Use the Mole – Chemistry (by Mr. Causey)
https://www.youtube.com/watch?v=wrGhyhgwODo

How to Calculate Mole Ratio and Calculate Molar Mass – Chemistry (by Mr. Causey)
https://www.youtube.com/watch?v=rwhJklbK8R0

How to Change Grams to Moles to Particles with Molar Mass – Chemistry (by Mr. Causey)
https://www.youtube.com/watch?v=dql8u81NPCI