Unit 5: The Mole

Mrs. Snyder

Honors Chemistry

Unit Learning Objectives:

By the end of the unit students will be able to…

- Define the quantity of one mole by stating Avogadro's number.
- Determine the number of atoms of each element in a given compound from the chemical formula.
- Define molar mass.
- Determine the molar mass for a given element/compound with the appropriate units.
- Perform conversions between moles, mass, and atoms/molecules and solve problems involving these quantities, giving answers with the appropriate units and significant figures.
- Name and write formulas for hydrates
- Determine the percent composition for a compound from the chemical formula. Determine the percent composition of water in a hydrate.
- Determine the empirical formula and the molecular formula for a compound from the percent composition of mass information.
- Identify the reacts and products of a chemical reaction.
- State the Law of Conservation of Mass
- Balance chemical equations beginning with either chemical names or formulas
- Classify reactions as synthesis, decomposition, single replacement, double replacement, neutralization, or combustion.
- Predict the products of a reaction given the formulas or names of reactants.
- Label the state of a substance in a chemical reaction as solid, liquid, gas, or solution (aq).

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November 5 Mole Conversions 2 Step Problems	6 **Lab: Empirical Formula (Hydrates Lab)**	7 Hydrates/Percent Composition Finish Lab	8 Empirical/Molecular Formula	9 **Quiz: Mole Conversions**
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It is important to be able to determine the number of atoms of each element in a compound. The number of atoms of each element can be found from the subscripts in the chemical formula. If there are parenthesis in the formula, the number of atoms of an element is calculated by multiplying the number outside the parenthesis by the subscript for that element.

Example: how many atoms of each element are in the molecule?

1. $CaCO_3$
2. $Mg_3(PO_4)_2$
3. $Fe(C_2H_3O_2)_2$

However, chemists do not work with such small amounts of atoms/molecules all the time. They work with much larger numbers and therefore chemists count in terms of the MOLE. In your OWL Candy Activity you learned about the MOLE.

What is the mole?

What is Avogadro 's number?

Chemists study chemical reactions, therefore it is important for chemists to be able to express how much of a given substance is present at both the start and end of a reaction. For this reason chemists have to be able to calculate in terms of

- Moles
- Mass
- Number of particles (number of atoms/molecules)

Chemists also need to be able to convert between any of the above three.

To convert between moles and particles use ***Avogadro's number***

Avogadro's number can be written as two different conversion fractions:

Note: Atoms are used to express the number of particles for substances with only one atom of one element. (ex. Cr, He, Fe, etc.),molecules are used to express the number of particles for substances with more than one atom, and /or more than one element (H_2O, F_2 etc.)

Example: How many atoms of carbon are in 0.033 moles of carbon?

Example: How many moles of $C_{12}H_{22}O_{11}$ contain 3.4 x 10^{22} molecules?

Practice: Complete each of the following examples.

1. How many atoms are in 0.200 mol of Au?

2. How many atoms are in 0.45 mol of carbon?

3. How many mol are in 7.8 x 10^{23} molecules of NaCl?

Molar Mass

Molar Mass is the mass, in grams, of 1 mole of an element or compound. (Units= g/mol)

Molar mass for elements can be found below the element on the periodic table (molar mass for elements is equivalent to atomic mass)

Example:

Element	Molar Mass (g/mol)
Sulfur	
Zinc	
Mercury	
uranium	

The molar mass of a compound can be determined by adding the molar mass of all of the elements in the compound. Taking into account the number of atoms of each element. Do not round the calculated values for molar mass.

Example

1. What is the molar mass of NaCl?

2. What is the molar mass of Na_2O?

3. What is the molar mass of sodium sulfate?

4. What is the molar mass of fluorine?

Molar Mass Conversions

*To convert between mass and moles, use **MOLAR MASS***

Example: What is the mass of 1.5 moles of chromium?

Example: How many moles are in 55.0 g of carbon dioxide?

*To convert between mass and atoms/molecules, **the moles must first be calculated**.*
(Two Step Conversions)

Example: What is the mass of 8.25×10^{22} molecules of NaF?

Example: How many molecules are contained in 0.850 kg of K_3PO_4?2215

Example: How many kg are in 3.21 $x10^{24}$ molecules of strontium chloride?

Practice: Mole conversions

1. How much does 4.04 x 10^{24} molecules of potassium chlorate weigh?

2. How many molecules are in 29 g of dinitrogen trioxide? How many atoms are there of each element?

Mole Conversion Summary Table

- To convert between mass and moles use ________________________________
- To convert between moles and atoms/molecules use _______________________
- To convert between mass and atoms/molecules use ________________________ and _____________________________. These are _______________________________.

HYDRATES

-
-
-
-
-

Example: Name each hydrate.

1. $Ba(OH)_2 \cdot 8H_2O$

2. $MnCl_2$

Example: Write the formula for each hydrate

1. Lithium nitrate trihydrate

2. Nickel (II) sulphate hexahydrate

Percent Composition

How to determine percent composition

1. ______________________________
2. ______________________________
3. ______________________________
4. ______________________________
5. ______________________________

Example: What is the percent composition of each element in $SrCl_2$?

Practice:

1. What is the percent composition of each element in CaC_2O_4?

2. What is the percent composition of each element in $Na_2SO_4 \cdot 10H_2O$?
   * Note: The coefficient in front of water multiplies the subscript for both hydrogen and oxygen to determine the number of atoms of each of these elements.

3. What is the percent composition of water in the compound?

Review: Percent Composition and Hydrates

1. Name the following hydrate: $Mg(NO_3)_2 \cdot 6H_2O$

2. Calculate the percent composition of each element in the hydrate

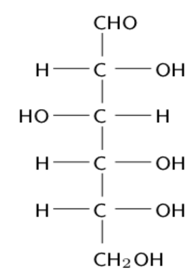
3. Calculate the percent composition for water in the hydrate

Empirical and Molecular Formulas:

Empirical Formula ____

Molecular Formula____

Example: Glucose has the *molecular formula* $C_6H_{12}O_6$ (since this is the number of atoms if each element found in its structure) but can also be reduced (by dividing by six) to its *empirical formula* CH_2O.



Note: The molecular formula is always a multiple of the empirical formula. As a result, the molar mass of the molecular formula (the molecular weight) is always a multiple of the molar mass of the empirical formula (the empirical weight).

	Empirical	**Molecular**
Formula	CH_2O	$C_6H_{12}O_6$
Molar Mass	30.026 g/mol	180.156 g/mol

Determination of Empirical formulas
The empirical formula is determined from the mass of percent composition for a compound.

Steps to Determine the Empirical formula:

1. Create a table with a row for each element and columns for mass, moles, and ratio.
2. Convert the mass of each element to moles.
3. Find the ratio of moles of each element by dividing the moles of each element by the smallest number of moles.
4. Write the empirical formula. The ratio of moles indicate the number of atoms of each element (IE the subscripts) in the compound.

Example: Determine the empirical formula if 3.50 g of a compound is found to contain 2.21 g of manganese and 1.29 g of oxygen.

Example: Determine the empirical formula for a compound that contains 65.2 % scandium and 34.8 % oxygen.

Chemical formulas must be given with whole numbers. If one of the ratios does not turn out to be a whole number (generally within ~0.05), they must all be multiplied by the same number so that they are all whole numbers.

Practice: Determine the empirical formula of a compound that contains 81.7 % carbon and 18.3 % hydrogen.

Determination of Molecular Formula:
Molecular formulas can be determined from the empirical formula and the molecular weight (the molar mass of the molecular formula).

Steps to determine the molecular formula:

1. Determine the empirical formula
2. Determine the molar mass for the empirical formula (the empirical weight)
3. Determine the value of the given molecular weight divided by the empirical weight. This value indicates what to multiply the empirical formula by to obtain the molecular formula. This value will always be a whole number.

Example: Paradichlorobenzene is the chemical found in moth balls. This compound consists of 49.0 % carbon, 2.80 % hydrogen, and 48.2 % chlorine. The molecular weight is 146.992 g/mol. Determine the molecular formula.

Example: Hydralazine is a pharmaceutical used to treat hypertension (high blood pressure). A 3.00 g sample contains 1.80 g of carbon, 0.15 g of hydrogen, and the remainder is nitrogen. The molecular weight is 160.184 g/mol. Determine the molecular formula

Practice: Xylose is a type of sugar found in the artificial sweetener called xylitol. Xylose conatins only carbon, hydrogen, and oxygen. A 2.000 g sample of xylose is found to contain 0.800 g of carbon, 0.134 g of hydrogen, and the remainder is oxygen. The molecular weight of xylose is 150.13 g/mol. Determine the molecular formula.

Chemical Reactions:

A Chemical Reaction is :

$$Reactants \rightarrow Products$$

Types of Chemical Reactions:

1- Synthesis

2- Decomposition

3- Dissociation

4- Combustion

5- Single Replacement

6- Double Replacement

7- Neutralization

In a chemical reaction the number of atoms of each element on the reactants side must be equal to the number of atoms of each element on the product side in a BALANCED chemical equation. This follows the LAW OF CONSERVATION OF MASS.

The state of a substance in a chemical reaction is given using the following symbols:

- (s) ____
- (l) ____
- (g) ____
- (aq)____

Balancing Chemical Reactions:

When the number of atoms of each element on the reactants side is not equal to the number of atoms of each elements on the product side, then the equation must be balanced.

Rules for Balancing Chemical Reactions:

1. ____
2. ____
3. ____
4. ____
5. ____

Examples: Balance the following chemical reactions.

1. Zinc oxide → zinc + oxygen

2. Aluminum + iron (III) oxide → aluminum oxide + iron

3. Calcium nitrate + sodium phosphate → calcium phosphate + sodium nitrate
   (for balancing polyatomic ions, treat them as a single unit!)

Predicting Products in Chemical Reactions:

To predict the products of a reaction:

1. Determine the type of reaction. (For **Reaction Types**: *See Jig Saw Activity*)
2. Then use the chart in your Jig Saw activity to help you determine the products for the chemical reaction.
3. Balance the chemical reaction.

Practice: Classify each reaction, predict the products, and then balance the chemical reaction.

1. $Mg + O_2 \rightarrow$

2. $C_9H_{20} + O_2 \rightarrow$

3. $HCl + Sr(OH)_2 \rightarrow$

4. $NCl_3 \rightarrow$

5. $KCl + Pb(NO_3)_2 \rightarrow$

6. $ZnCl_2 + Na \rightarrow$

7. $CaBr_2 + O_2 \rightarrow$