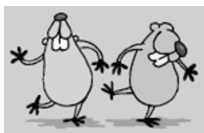
Stoichiometry & the Mole



The Mole

- __________ - SI base unit used to measure the amount of a substance.
- A mole of anything contains __________ representative particles.

Mole – Representative Particle Calculations

- Calculate the number of atoms in 3.50 moles of copper

Mole – Mass Relationship

- __________ - mass in grams of one mole of any pure substance

Calculating Molecular Mass

- What is the molecular mass of $(C_3H_5)_2S$?

Mole – Mass Calculations

- How many moles of $Ca(OH)_2$ are in 325 grams?

Mass – Particle Conversions

- How many atoms of gold are in 25.0 g of gold?

Mass – Mass Relationships

- Ammonium nitrate decomposes into dinitrogen monoxide gas and water. Determine that amount of water produced if 25.0 g of ammonium nitrate decomposes.

Limiting Reactants

- A chemical reaction will stop when you run out of one of your products
- __________– limits the extent of the reaction.
  - Determines the amount of product that is formed.
  - It runs out first
- __________– left over reactant

Example

- $S_8 + 4Cl_2 \rightarrow 4S_2Cl_2$
- 200.00 g of S_8 and 100.00 g of Cl_2 are combined in a flask. How much S_2Cl_2 will you get?

Other questions

- What was the limiting reactant?
- What was the excess reactant?
- How much excess did we have left over after the reaction was completed?

% Yield

- % = (part / whole) x 100

% Yield

- ____________________– the amount of product you should get if the experiment went perfectly.
  - You get this number from stoichiometry
- ____________________– this is what you actually got in the lab.
  - You measure this on a balance
- ____________________– how close you were to the correct answer
  - % yield = (actual / theoretical) x 100

% Yield Example

- $K_2CrO_4 + 2AgNO_3 \rightarrow Ag_2CrO_4 + 2KNO_3$
- What is the theoretical yield of Ag_2CrO_4 formed from 0.500 g $AgNO_3$?
- What is the % yield if 0.455 g is actually formed?

% Composition

- % = (part / whole) x 100
- Calculate the % Composition of iron (III) oxide

Empirical & Molecular Formulas

- ____________ – the smallest whole number ratio of elements
- ____________– the true number of elements in a compound

Steps for Calculating the Empirical Formula

1. List your givens
2. Change % to grams
3. Change grams to moles
4. Divide everything by the smallest number of moles
5. Write your formula

Empirical Formula Problem

- Calculate the empirical formula for a compound containing 48.64 g C, 8.16 g H, and 43.20 g O.

Steps for Calculating Molecular Formula

1. Calculate the empirical formula
2. Get the molecular mass of the empirical formula that you just determined
3. Divide the experimentally determined molecular mass (given) by the molecular mass of the empirical formula
4. You will get a whole number
5. Multiply everything in the empirical formula by this number

Molecular Formula Problem

- Calculate the molecular formula of a compound containing 40.68%C, 5.08%H, and 54.25%O with an experimentally determined molecular weight of 118.1 g/mol

Empirical Formula with Combustion Data Steps

1. Convert g CO_2 to g C
2. Convert g H_2O to g H
3. Subtract to get g of other element
4. Work Empirical Formula problem as usual

Example

- A compound is comprised of carbon, hydrogen, and nitrogen. When 0.1156 g of this compound is reacted with oxygen, 0.1638 g of CO_2 and 0.1676 g of water are collected. Assuming that all of the carbon in the compound is converted into CO_2, determine the empirical formula of the compound.

Example

One more example

- A 0.821 g hydrocarbon sample was combusted to yield 1.866 g CO_2 and 0.7639 g H_2O. The molecular mass was determined to be 116 g/mol. Determine the empirical and molecular formulas of the compound.