Stoichiometry

The branch of chemistry dealing with mass relationships of elements within compounds and among reactants and products in chemical reactions.

Atomic Mass: The mass in atomic mass units of an element

Molecular Mass: The mass in atomic mass units of a molecule

Formula Mass: The mass in atomic mass units of an ionic compound

Na - 23.0 X 1 = 23.0
Cl - 35.5 X 1=35.5

58.5 amu

C - 12 X 1 = 12
O - 16 X 2 = 32
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44 amu

Moles

- The quantity of matter containing Avogadro's number of particles
- 6.022 X 10^{23} particles
- Particles may include:
  - subatomic particles
  - ions
  - atoms
  - molecules

Molar Mass: The mass of one mole of a substance in grams

Stoichiometric Calculations

Calculating moles from grams

Divide grams by molar mass

Calculating grams from moles

Multiply moles by molar mass

Solutions: A homogeneous mixture of two or more substances

Components – Two parts consisting of the Solvent and Solute

Concentration – General terms consist of dilute or concentrated. Specific term describing the concentration is molarity (M). M = moles of solute / L of solution

Example: White vinegar is a solution of acetic acid, CH_3CO_2H, in water. Vinegar, with an acidity of 5.00%, contains 50.4 g of acetic acid in 1.00 L of vinegar. Determine the concentration in moles per liter.

First step -- Convert grams of CH3CO2H to moles:

$$mol\ CH_3CO_2H = 50.4g\ CH_3CO_2H \times \frac{1\ mol\ CH_3CO_2H}{60.0\ gCH_3CO_2H} = 0.839\ mol\ CH_3CO_2H$$

Second step --- Convert moles to molarity

$$M\ CH_3CO_2H = \frac{0.839\ mol\ CH_3CO_2H}{1\ L\ solution} = 0.839\ M\ CH_3CO_2H$$

Titration: A process for determining the concentration of a solution by allowing a carefully measured volume to react with a solution of a second substance whose concentration is known.

Percent Composition

$$\%X = \frac{mass\ X}{mass\ sample} \times 100\%$$

Empirical Formula

$$mol\ X = \frac{mass\ X}{molar\ mass}$$

Example: CO_2

First step:

$$C: 1 \times 12.0\ amu = 12.0\ amu$$

$$\underline{O: 2 \times 16.0\ amu = 32.0\ amu}$$

44.0 amu

Second step:

$$C: (12.0\ amu\ \div 44.0\ amu)\ \times 100\% =\ 27.3\ \%$$

$$O: (32.0\ amu\ \div 44.0\ amu)\ \times 100\% = 72.7\ \%$$

Types of data used in calculating the empirical formula

- Percentage Composition Data -- utilizes data as percentages of substance
- Relative Mass Data -- utilizes data as grams of substance

Determination of Molecular Formula

(empirical formula)$_x$ = molecular formula $\quad x = \frac{molecular\ mass}{empirical\ formula\ mass}$

Example 1: A compound of carbon and hydrogen contains 92.3 % C and has a molecular mass of 78.1 amu. Determine its molecular formula.

$$C: 92.3\ \%\ \Rightarrow 92.3\ g\ \times \frac{1\ mol}{12.0\ g} = 7.69\ mol$$

$$H: 7.7\ \%\ \Rightarrow 7.7\ g\ \times \frac{1\ mol}{1.01\ g} = 7.62\ mol$$

Then: $C: 7.69\ mol\ \div 7.62 = 1$ and H: $7.62\ mol\ \div 7.62 = 1$ So… CH

But… molecular mass is 78.1 amu So…$x = \frac{78.1\ amu}{formula\ mass} = \frac{78.1\ amu}{13.0\ amu}$ =6.00

Finally… $(CH)_x = (CH)_6 = C_6H_6$

Example 2: Determine the molecular formula of a compound of 43.6% phosphorus and 56.4%.

$$P: 43.6\ \%\ \Rightarrow 43.6\ g\ \times \frac{1\ mol}{31.0\ g} = 1.41\ mol$$

$$O: 56.4\ \%\ \Rightarrow 56.4\ g\ \times \frac{1\ mol}{16.0\ g} = 3.52\ mol$$

Then: P: $1.41\ mol\ \div 1.41 = 1$ and O: $3.52\ mol\ \div 1.41 = 2.5$

Since the ratios are not integer (almost integers) then they must be multiplied by an common factor that will give an integer.

So… P: 1 X 2 = 2 and O: 2.5 X 2 = 5 Therefore… P_2O_5

Types of Equation Based Calculations

Mole to Mole
Mole to Mass
Mass to Mole
Mass to Mass

Procedure

- Write a balanced chemical equation
- Determine what is given and what is to be determined
- Determine the molar ratios
- Use dimensional analysis to:
  - convert mass to moles
  - compare moles
  - convert moles to mass

An Example: What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g (1 L) of octane, C_8H_{18}, one of the principal components of gasoline?

$$\begin{array}{cc} \boxed{702\ g} & \boxed{x\ g} \\ 2C_8H_{18} & +\ 25O_2 \rightarrow 16CO_2 + 18HO \\ \boxed{2\ mol} & \boxed{25\ mol} \end{array}$$

$$x\ g\ O_2 = 702\ g\ C_8H_{18} \times \frac{1\ mol\ C_8H_{18}}{114.23\ g\ C_8H_{18}} \times \frac{25\ mol\ O_2}{2\ mol\ C_8H_{18}} \times \frac{32.00\ g\ O_2}{1\ mol\ O_2} = 2.46 \times 10^3\ g\ O_2$$

Another Example: What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the reaction of magnesium chloride, $MgCl_2$, with NaOH?

$\boxed{\text{x g}}$ $\boxed{\text{16 g}}$

$MgCl_2 + 2NaOH \rightarrow Mg(OH)_2 + 2NaOH$

$\boxed{\text{2 mol}}$ $\boxed{\text{1 mol}}$

$$x\ g\ NaOH = 16\ g\ Mg(OH)_2 \times \frac{1\ mol\ Mg(OH)_2}{58.3\ g\ Mg(OH)_2} \times \frac{2\ mol\ NaOH}{1\ mol\ Mg(OH)_2} \times \frac{40.0\ g\ NaOH}{1\ mol\ NaOH} = 22\ g\ NaOH$$

Percentage Yield $$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100\%$$

A Percent Yield Example: A general chemistry student, preparing copper metal by the reaction of 1.274 g of copper (II) sulfate with zinc metal, obtained a yield of 0.392 g of copper. What was the percent yield?

$\boxed{\text{1.276 g}}$ $\boxed{\text{x g}}$

$CuSO_4 + Zn \rightarrow Cu + ZnSO_4$

$\boxed{\text{1 mol}}$ $\boxed{\text{1 mol}}$

$$x\ g\ Cu = 1.274\ g\ Cu(SO)_4 \times \frac{1\ mol\ Cu(SO)_4}{159.6\ g\ Cu(SO)_4} \times \frac{1\ mol\ Cu}{1\ mol\ Cu(SO)_4} \times \frac{63.54\ g\ Cu}{1\ mol\ Cu} = 0.5072\ g\ Cu$$

Since there are 0.5072 g Cu, % yield can now be calculated from the formula

$$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100\% = \frac{0.392\ g}{0.5072\ g} \times 100\% = 77.3\%$$

Limiting Reagents

- Calculate a mass to mass stoichiometric problem using one of the two givens.
- Calculate the same problem, but using the second of the two givens.
- Determine which one limits the other
- This is the limiting reagent.

Reactions with Limiting Amounts of Reactants

- Most reactions do not use exactly the right proportions of reactants.
- Many reactions use an excess amount of one reactant.
- Whenever the ratios of reactant molecules used in an experiment are different from those give by the coefficients of the balanced equation, a surplus of one reactant remains.
- The extent to which a chemical reaction takes place depends on the reactant that is present in the limiting amount – The Limiting Reactant

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