Quantitative Chemistry 3.1

Molecular Weight
Mass %

Formula and Molecular Weights
The sum of the atomic masses of the atoms in a chemical formula

Formula weight – Usually refers to the basic unit in a network solid

Molecular weight – Is used to refer to molecules that can exist independently

Formula Weight of Magnesium Hydroxide
Mg(OH)₂

1 Mg atom 1(24.31 amu) = 24.31 amu
2 O atoms 2(16.00 amu) = 32.00 amu
2 H atoms 2(1.01 amu) = 2.02 amu

Formula Mass = 58.33 amu

Mass Percent
• The percentage by mass of an element in a substance.

\[
\text{mass % element} = \left( \frac{\text{# of atoms of element}}{} \right) \left( \frac{\text{element's atomic mass}}{} \right) \times 100
\]

\[
(fomula \ weight \ of \ compound)
\]

Ex1) Mass % of Elements in Ilmenite (FeTiO₃)
Ex1) Ilmenite is a natural source of titanium. Find the mass percent of each element in the mineral Ilmenite, FeTiO₃.

Ex1) Mass % of Elements in Ilmenite (FeTiO₃)

Mass Percent of Iron
Ex1) Mass % of Elements in Ilmenite (FeTiO$_3$)

Mass Percent of Titanium

Ex1) Mass % of Elements in Ilmenite (FeTiO$_3$)

Mass Percent of Oxygen

Finding the Mass % of Water in a Hydrate

- The same procedure can be used to find the mass % of water in a hydrate.
- As we are concerned with water, not the individual elements, the formula becomes:

\[
\text{mass % H}_2\text{O} = \frac{\# \text{ of water molecules}(MW_{H_2O})}{\text{formula weight of compound}} \times 100
\]

Ex2) Find the Mass % of Water in a Hydrate

Ex2) Find the mass % of water in AlCl$_3$·6H$_2$O.

Ex3) Mass Percent

Ex3) A sample of calcium carbonate (CaCO$_3$) is known to contain some impurities. It is found that Ca makes up 11% of the entire mass of the sample. All of the Ca comes from the CaCO$_3$ compound. Find the mass percent of CaCO$_3$ in the sample.

Ex3) Mass Percent

Set up a ratio and solve
Quantitative Chemistry 3.2

The Mole
Converting from Mass to Moles
Converting from Moles to Mass
Converting from Moles to Atoms

Avogadro’s Number ($N_A$)

$6.022 \times 10^{23}$ mol$^{-1}$

The Mole

1 mol Ne atoms = $6.022 \times 10^{23}$ Ne atoms = 20.18 g

1 mol Cl$^-$ ions = $6.022 \times 10^{23}$ Cl$^-$ ions = 35.45 g

Molar Mass of CO$_2$

Molecular Weight = 12.01 amu + 2(16.00 amu)
= 44.01 amu/CO$_2$ molecule
= 44.01 g/mol CO$_2$

1 mol CO$_2$ molecules = $6.022 \times 10^{23}$ CO$_2$ molecules
= 44.01 g

Ex) Converting Grams to Moles

Ex) How many moles of NaCl are there in 48 g of NaCl?

Ex) Converting Moles to Grams

Ex) How grams of C$_2$H$_6$ are there in 18.7 moles of C$_2$H$_6$?
Ex) Converting from Grams to Molecules

Ex) How many water molecules are contained within a 56 g sample of water?

Ex) Converting from Grams to Atoms

Ex) How many hydrogen atoms are contained within a 56 g sample of water?
Stoichiometry 3.3

Mole to Mole Problems

Ex1) Mole to Mole Problems
Ex1) How many moles of ClF$_{3(g)}$ can be produced from 8.0 moles of Cl$_{2(g)}$ in the presence of excess F$_{2(g)}$?

Cl$_{2(g)}$ + 3 F$_{2(g)}$ → 2 ClF$_{3(g)}$

Ex2) Mole to Mole Problems
Ex2) How many moles of ClF$_{3(g)}$ can be produced from 6.79 moles of F$_{2(g)}$ in the presence of excess Cl$_{2(g)}$?

Cl$_{2(g)}$ + 3 F$_{2(g)}$ → 2 ClF$_{3(g)}$

Ex3) Mole to Mole Problems
Ex3) How many moles of F$_{2(g)}$ are need to produce 72.4 moles of ClF$_{3(g)}$ in the presence of excess Cl$_{2(g)}$?

Cl$_{2(g)}$ + 3 F$_{2(g)}$ → 2 ClF$_{3(g)}$

Ex4) Mole to Mole Problems
Ex4) Copper (I) sulfide is heated in the presence of oxygen gas to produce copper (I) oxide powder and sulfur dioxide gas. How many moles of O$_{2(g)}$ are needed to produce 5.82 moles of copper (I) oxide?

Step 1. Write the balanced equation

Step 2. Solve for moles of O$_{2}$

Mole Ratios

- The stoichiometric coefficients in a balanced chemical equation provide definite whole number ratios between the species involved in the reaction.

e.g.) \[ CH_4 + 2 O_2 \rightarrow CO_2 + 2 H_2O \]

1 CH$_4$ reacts with 2 O$_2$ to produce 1 CO$_2$ and 2 H$_2$O

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Stoichiometry 3.4

Mass to Mass Problems

Mass to Mass Stoichiometry

- As compounds are measured in grams, not moles, one normally uses reaction stoichiometry to determine the mass of a product that will be produced by a certain mass of a reactant, or vice versa.
- The series of calculations that must be performed is as follows:
  \[
  \text{mass}_A \rightarrow \text{moles}_A \rightarrow \text{moles}_B \rightarrow \text{mass}_B
  \]

Ex1) Predicting Mass of Products

Ex1) What mass of water is produced when a car burns 246.4 g of methane?

Ex2) Predicting Mass of Products

Ex2) How many grams of CO₂ and Fe are produced when 114 g of carbon monoxide gas is added to a vessel containing excess hot iron (III) oxide?

Step 1. Write Balanced Chemical Equation

Step 2. Find Masses of CO₂ and Fe

Ex3) Predicting Mass of Reactants

Ex3) What mass of sodium bicarbonate is needed to produce 32 g of Na₂CO₃?

\[
2 \text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]
Stoichiometry 3.5

Limiting Reactant
- The reactant that is used up limits how far the reaction will proceed.

Excess Reactant
- The reactant that is leftover when the reaction is complete.

Ex1) Limiting Reactant Problem
(a) What is the limiting reactant when 28 g of Glucose reacts with 14 g of Oxygen gas?
(b) What mass of CO₂ is produced?

Step 1. Write a Balanced Chemical Equation

\[ C_6H_{12}O_6 + 6 \text{ O}_2 \rightarrow 6 \text{ CO}_2 + 6 \text{ H}_2\text{O} \]

Ex1) Limiting Reactant Problem (cont.)
(a) What is the limiting reactant when 28 g of Glucose reacts with 14 g of Oxygen gas?
(b) What mass of CO₂ is produced?

Step 2. Find the mass of CO₂ that would be produced by each reactant.

\[ C_6H_{12}O_6 + 6 \text{ O}_2 \rightarrow 6 \text{ CO}_2 + 6 \text{ H}_2\text{O} \]
Ex1) Limiting Reactant Problem (cont.)

Step 3. Compare the two masses produced. The reactant that produced the smallest quantity of product is the limiting reactant.

19 g < 41 g
(a) Thus, O₂ is the limiting reactant.
(b) 19 g of CO₂ is produced in theory.

Ex2) Percent Yield

Ex2) Find the percent yield if only 15 grams of CO₂ were produced in the previous problem.

\[ \%\text{Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 \]

Ex3) Limiting Reactant

Ex3) (a) Find the limiting reactant 155 g of N₂H₄(l) react with 175 g N₂O₄(g).
(b) What mass of H₂O(g) is produced?

\[ 2 \text{N₂H₄(l)} + \text{N₂O₄(g)} \rightarrow 3 \text{N₂(g)} + 4 \text{H₂O(g)} \]

Ex3) Limiting Reactant (cont.)

Step 1. Find the mass of H₂O that could be produced by each reactant.

\[ 2 \text{N₂H₄(l)} + \text{N₂O₄(g)} \rightarrow 3 \text{N₂(g)} + 4 \text{H₂O(g)} \]

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Ex3) Limiting Reactant (cont.)

Step 3. Compare the two masses produced. The reactant that produced the smallest quantity of product is the limiting reactant.

137 g < 174 g
(a) Thus, N₂O₄ is the limiting reactant.
(b) 137 g of H₂O is produced in theory.

Ex4) Percent Yield

Ex2) Find the percent yield if only 129 grams of H₂O were produced in the previous problem.
Stoichiometry 3.6

Empirical and Molecular Formulas

Molecular and Empirical Formulas

Molecular Formulas

- Chemical formulas that provide the actual number of each type of atom in molecule.

Empirical Formulas

- Chemical formulas that provide the relative number of each type of atom in molecule.
  (a ratio in simplest form)

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<th>Empirical Formula</th>
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<td>Dextrose</td>
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<tr>
<td>Hydrazine</td>
<td>N₂H₄</td>
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</tr>
</tbody>
</table>

Ex1) Empirical Formula

Ex1) A sample of caffeine was found to contain 49.5% Carbon, 28.9% Nitrogen, 16.5% Oxygen, and 5.1% Hydrogen by mass.
Find the empirical formula for caffeine.

Ex1) Empirical Formula (cont.)

Step 1. Assume you have a 160g sample and convert each element into moles.
Ex1) Empirical Formula (cont.)

Step 2. Divide the number of moles of each element by the smallest value for moles calculated in the previous step.

Ex2) Molecular Formula

Ex2) The molar mass of caffeine is 194.2 g/mol. Find the molecular formula for caffeine.

Empirical Formula_{caffeine} = C_{4}H_{N_{2}}O

\[
\text{Molar Mass} = \text{Multiplier for Empirical Formula} \times \frac{\text{Empirical Formula Mass}}{12} = \text{Empirical Formula Mass}
\]

Ex3) Combustion Analysis

Ex3) A 2.04 g sample containing C, H, and O underwent combustion analysis. 4.49 g of CO₂ and 2.45 g of H₂O were produced.

Find the Empirical Formula.

Your Thinking...

1) Use stoichiometry to find the grams of Carbon and Hydrogen produced. Subtract the total mass from the sum of the masses of carbon and hydrogen to find the mass of oxygen.

2) Find moles of carbon, hydrogen, and oxygen

3) Find the empirical formula

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Ex3) Combustion Analysis (cont.)
Step 2. Find moles of carbon, hydrogen, and oxygen

Ex3) Combustion Analysis (cont.)
Step 3. Divide the number of moles of each element by the smallest value of moles calculated.

Ex4) Molecular Formula
Ex4) Further experiments showed that molar mass of the hydrocarbon from example 3 is 60.11 g/mol. What is the molecular formula of this compound?