Chapter 3. Chemical Stoichiometry

1. Atomic Masses (section 2.5)
Two important definitions from junior high:

percent =

fraction =

“atomic weight”

e.example of a weighted average:

Example:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>mass (amu)</th>
<th>abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{24}\text{Mg}$</td>
<td>23.9850</td>
<td>78.99%</td>
</tr>
<tr>
<td>$^{25}\text{Mg}$</td>
<td>24.9858</td>
<td>10.00%</td>
</tr>
<tr>
<td>$^{26}\text{Mg}$</td>
<td>25.9826</td>
<td>11.01%</td>
</tr>
</tbody>
</table>

Example: The atomic weight of carbon is 12.011. Given that naturally occurring carbon consists of only two isotopes, $^{12}\text{C}$ and $^{13}\text{C}$, and given that the mass of $^{13}\text{C}$ is 13.00335 amu, what is the % abundance of each isotope?

2. The Mole (section 3.1)

Definition: “the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12.” (Silberberg, Principles, p. 70)

1. “Mole” is a ___________________ word.

   1 dozen eggs =
   a couple of tacos =
   1 gross of pencils =
   1 mol of particles =
   1 mol of atoms =
   1 mol of electrons =
   1 mol of molecules =

Example: How many moles of atoms are present in $9.5 \times 10^{15}$ atoms?
Example: How many molecules are present in 2.35 mol of molecules?

2. “Mole” relates to ___________.
   
   1 mol C =
   1 mol Na =
   1 mol H₂O =

Example: How many moles of water are in 25.0 g of water?

Example: How many water molecules are in 675 g of water?

3. “Mole” concept provides ______________________ ratios.

   There are two places to get mol-to-mol ratios:
   
   1)
   2)

   Examples: H₂O

   2 C₂H₆(g) + 7 O₂(g) → 4 CO₂(g) + 6 H₂O(l)

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3. **Molar Mass (section 3.1)**

   definition:

   other terms:

   Example: What is the molar mass of water?

   What is the molar mass of ammonium phosphate?
**The molar mass of a substance can be used as a ________________ ________________ to convert from mass to moles, and from moles to mass.**

**A useful formula for molar mass is:**

Example: How many moles of water are in 9.0 g of water?

Example: How many hydrogen atoms are in 10.0 g of water?

Example: How many ions are in 15.0 g of ammonium carbonate?

4. **Balancing Chemical Equations (section 3.3)**

In order to correctly calculate quantities in chemical reactions, the balanced equation must be used.

Examples: Balance the following equations.

\[
\begin{align*}
\text{Fe}_3\text{O}_4 & \quad + \quad \text{H}_2 & \quad \rightarrow \quad \text{Fe} & \quad + \quad \text{H}_2\text{O} \\
\text{Fe}_2\text{O}_3 & \quad + \quad \text{C} & \quad \rightarrow \quad \text{Fe} & \quad + \quad \text{CO}_2 \\
\text{NH}_3 & \quad + \quad \text{O}_2 & \quad \rightarrow \quad \text{NO} & \quad + \quad \text{H}_2\text{O} \\
(\text{NH}_4)_3\text{PO}_4 & \quad + \quad \text{Mg(NO}_3)_2 & \quad \rightarrow \quad \text{Mg}_3(\text{PO}_4)_2 & \quad + \quad \text{NH}_4\text{NO}_3
\end{align*}
\]

**One important reaction to know: Combustion.**
When a hydrocarbon (a compound containing only C and H), or a compound containing only C, H, and O undergoes complete combustion in oxygen, the products are \( \text{CO}_2 \) and \( \text{H}_2\text{O} \).

Example: Write the balanced equation for the complete combustion of ethane, \( \text{C}_2\text{H}_6 \) in oxygen.

Example: Write the balanced equation for the complete combustion of ethanol, \( \text{CH}_3\text{CH}_2\text{OH} \), in oxygen.
5. **Reaction Stoichiometry (section 3.4)**
Stoichiometry involves using the coefficients in a balanced chemical equation to determine quantities in chemical reactions.

\[ 3\text{H}_2(g) + \text{N}_2(g) \rightarrow 2\text{NH}_3(g) \]

*means*

Thus, we derive **mole-to-mole ratios**:

**Example:** How much ammonia can be formed if 12.0 g of H\(_2\) reacts with excess nitrogen?

**Example:** What mass of oxygen gas is required for the complete combustion of 29.0 g of butane, C\(_4\)H\(_{10}\)?

**Example:** What mass of oxygen gas is required for the complete combustion of 10.0 g of cyclopentanol, C\(_5\)H\(_9\)OH?

6. **Limiting Reactant (section 3.4)**
It is very rare for the reactants in a chemical reaction to be mixed in the exact proportions required by the equation. Usually, one reactant will be used up (the **limiting reactant** or **limiting reagent**) while the other reactants are called **excess reactants** (or **excess reagents**).

One way to determine the limiting reactant: Set up two cases. In each case, assume that one reactant is the limiting reactant, and calculate the mass of product. The reactant yielding the least product is the limiting reactant.

**Bologna Sandwich Example:** If you have 50 slices of bread and 50 slices of bologna, how many sandwiches can you make? If you make 17 sandwiches, what is your percent yield?

**Example:** If 28 g of nitrogen reacts with 18 g of hydrogen, what is the theoretical yield of ammonia?
7. Percent Yield (section 3.4)

In most cases, the amount of product collected at the end of a reaction (the actual yield) is less than the yield that we predict we should get using reaction stoichiometry (the theoretical yield).

\[
\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]

Example: If 16 g of hydrogen reacts with 16 g of oxygen and 9 g of water is produced, what is the percent yield?

8. Percent Composition of Compounds (section 3.1)

Junior High definition of percent:

Example: What is the weight percent of nitrogen in ammonium phosphate?

9. Determining the formula of a Compound (section 3.2)

A molecular formula shows:

An empirical formula shows:

Example: What is the empirical formula of benzene?

Two keys to solving empirical formula problems:

1. 

2. 

Example: What is the empirical formula of a compound which consists of 52.13% C, 13.14% H and 34.73% O by weight?
If you are given the percent composition and the molar mass, the problem is easier. Instead of starting with 100 g of unknown, start with __________________________.

Example: A certain compound consists of 38.70% C, 9.74% H, and 51.56% O by mass. The molar mass of this compound is 62.07 g/mol. What is the molecular formula of this compound?

10. Combustion Analysis (Section 3.2)

In combustion analysis, all of the carbon in the unknown is converted to _______________

and all of the hydrogen in the unknown is converted to ________________.

Example: An unknown compound consists of carbon, hydrogen, and oxygen. When a 1.125 g sample of the unknown was analyzed by combustion analysis, 1.649 g of CO₂ and 0.675 g of H₂O were formed. If the molar mass of this unknown is 180.2 g/mol, what is the molecular formula?
11. **Molarity and Solution Stoichiometry (Section 3.5)**

*Molarity* \((M)\) is used to describe solution concentration because it is very easy to determine.

\[
\text{Molarity} = \frac{\text{number of moles of solute}}{\text{volume of solution}}
\]

This quickly gives us the number of moles of solute:

\[
\text{# mol solute} = \frac{n}{V}
\]

Three things to know about molarity:

1. \(M \times V = \) _________________________
2. When I see “M”, I think _______________________
3. Molarity can be used as a ________________ ______________ in stoichiometry problems.

Solution by dilution:

**Problems:**

1. What is the concentration of a solution of HCl that was prepared by diluting 5.00 mL of 12.0 M HCl to 150.0 mL?

2. If 5.85 g of sodium chloride (NaCl, molar mass = 58.5 g/mol) is dissolved in 100 mL of solution, what is the concentration of the solution expressed in molarity?

3. How many moles of HCl are present in 320 mL of 0.250 M HCl?

4. Consider the following reaction:

\[
\text{MgCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(aq) + \text{H}_2\text{O(l)} + \text{CO}_2(g)
\]

How many L of 0.100 M HCl are required to completely react with 13.0 g of MgCO\(_3\)?