1.1 Molar Mass

- **Molar mass** =

- Molar mass of molecules can be determined from the chemical formula and molar masses of elements
  - Each H₂O molecule
  - Each mole of H₂O molecules contains …
  - One mole of O atoms corresponds to ____ g
  - Two moles of H atoms corresponds to ____ g
  - Sum = molar mass = ____ g H₂O per mole

1.2 Molar Mass

- Calculate the molar mass of the following
  - Magnesium nitrate
  - Calcium carbonate
  - Iron(II) sulfate

1.3 Conversion Factors

<table>
<thead>
<tr>
<th>Number of particles</th>
<th>Moles</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Avogadro’s number</td>
<td>6.022 x 10²³</td>
<td>Molar mass</td>
</tr>
</tbody>
</table>

1.4 Molar Conversions

- Review:
  - The moles of potassium atoms in a 50.0 g sample
  - The mass of Mg in a 1.82 mole sample
1.5 Molar Conversions

- glucose = C₆H₁₂O₆
- How many moles of glucose are contained in 70.0 g of C₆H₁₂O₆?
- How many moles of H are contained in 70.0 g of C₆H₁₂O₆?
- How many grams of H are contained in 70.0 g of C₆H₁₂O₆?

1.6 Molar Conversions

- How many moles of N are contained in 70.0 g of C₆H₅NO₃?
- How many grams of N are contained in 70.0 g of C₆H₅NO₃?
- How many atoms of O are contained in 70.0 g of C₆H₅NO₃?

1.7 Mass Percentage

- What is the mass percentage of each element in urea (CH₄N₂O)?

1.8 Empirical Formulas

- A sample of nicotine has the following mass composition: 74.03% C; 8.70% H; 17.27% N
- What is the empirical formula of nicotine?
1.9 Using Empirical Formulas

- A sample of nicotine has the following mass composition: 74.03% C; 8.70% H; 17.27% N
- The molar mass of nicotine is 162.23 g
- What is the molecular formula of nicotine?

1.10 Using Mass Percent Data

- Ethylene glycol analyzes as 38.70 %C, and 9.74 %H with the remainder being oxygen.
- Determine the empirical formula of ethylene glycol

2.1 Solutions

- Solute
- Solvent
- Solution
- Aqueous

2.2 Solution Concentrations

- **Molarity** = moles of solute per liter of solution
  - Designated by a capital **M** (mol/L)
- Determine the molarity of each solution
  - 2.50 L of solution containing 1.25 mol of solute
  - 225 mL of solution containing 0.486 mole of solute
  - 100. mL of solution containing 2.60 g of NaCl
2.3 Solution Preparation

- Solutions of exact concentrations are prepared by dissolving the proper amount of solute in the correct amount of solvent – to give the desired final volume
- Determine the proper amount of solute

- How is the final volume measured accurately?

2.4 Solution Preparation

- Describe how to prepare the following:
  - 500. mL of 1.00 M FeSO\textsubscript{4}
  - 100. mL of 3.00 M glucose
  - 250. mL of 0.100 M NaCl

2.5 Solution Preparation

- Solutions of exact concentrations can also be prepared by diluting a more concentrated solution of the solute to the desired concentration

- Suppose you want to make 250. mL of 0.100 M NaCl, starting with 1.00 M NaCl. How much of the 1.00 M NaCl would you have to use?

2.6 Diluting a Solution

- First determine the number of moles of solute that will be contained in the final solution (in the 250. mL of 0.100 M NaCl)
  \[
  (0.250 \text{ L}) \times \frac{0.100 \text{ mol NaCl}}{\text{L}} = 0.0250 \text{ mol NaCl}
  \]
- Next determine what volume of the more concentrated solution will contain the same number of moles of solute
- Since the more concentrated solution is 1.00 M, the volume is calculated as follows:
  \[
  (0.0250 \text{ mol NaCl}) \times \frac{\text{L}}{1.00 \text{ mol NaCl}} = 0.0250 \text{ L}
  \]
2.7 Diluting a Solution

- In other words, taking 25.0 mL (0.0250 L) of 1.00 M NaCl and diluting it to 250. mL with pure water will produce a 0.100 M solution.
- Moles of NaCl are same in both solutions.
- Moles of NaCl don't change by upon dilution.

\[(0.0250 \text{ L} \cdot 1.00 \text{ M}) = (0.250 \text{ L} \cdot 0.100 \text{ M})\]

\[0.0250 \text{ moles} = 0.0250 \text{ moles}\]

\[M_c \cdot V_c = \text{moles} = M_d \cdot V_d\]

2.8 Solution Preparation

\[M_c \cdot V_c = \text{moles} = M_d \cdot V_d\]

- Describe how to prepare 500. mL of 0.250 M NaOH solution using a 6.00 M NaOH solution.

- What volume of 2.30 M NaCl should be diluted to give 250. mL of a 0.90 M solution?

2.9 Solution Preparation Review

- If you dissolve 9.68 g of potassium chloride in 1.50 L, what is the final molar concentration?

- How many grams of sodium sulfate are contained in (dissolved in) 45.0 mL of 3.00 M solution?

2.10 Solution Preparation Review

- What volume of 8.00 M sulfuric acid should be diluted to produce 0.500 L of 0.250 M solution?

- What's the pH of the final solution?
2.11 Solution Preparation Review

- What volume of 2.06 M potassium permanganate contains 322 g of the solute?

2.12 Solution Preparation Review

- How would you prepare 100. mL of a nitric acid solution with pH = 2.89? You have 0.0100 M nitric acid, a graduated cylinder and a 100. mL volumetric flask to work with. (hint: is nitric acid a strong acid?)

3.1 Moles and Chemical Reactions

- We have used the mole concept to calculate mass relationships in chemical formulas
  - Molar mass of ethanol (C$_2$H$_5$OH)?

- Mass percentage of carbon in ethanol?

3.2 Stoichiometry

- We can also use the mole concept to calculate mass relationships in chemical reactions

- **Stoichiometry** is the study of mole and mass relationships
  - It requires a balanced equation
  - Chemical equation relates moles of reactants to moles of products
  - The equation **DOES NOT directly** relate the masses of reactants and products
  - First set up the correct balanced equation (in terms of moles), then worry about converting to mass
3.3 Moles and Chemical Reactions

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \]

- The following statements are true:
  - 1 CH\(_4\) molecule + 2 O\(_2\) molecules → 1 CO\(_2\) molecule + 2 H\(_2\)O molecules
  - 6.022 x 10\(^{23}\) CH\(_4\) molecules + 2(6.022 x 10\(^{23}\)) O\(_2\) molecules → 6.022 x 10\(^{23}\) CO\(_2\) molecules + 2(6.022 x 10\(^{23}\)) H\(_2\)O molecules

- In other words…
  - 1 mole of CH\(_4\) + 2 moles of O\(_2\) → 1 mole of CO\(_2\) + 2 moles of H\(_2\)O

What about masses?

3.5 Stoichiometry

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \]

- Answer the following questions:
  - How many moles of O\(_2\) are required to react with 1.72 moles of CH\(_4\)?
  - How many grams of H\(_2\)O will form when 1.09 moles of CH\(_4\) react with excess O\(_2\)?
  - How many grams of O\(_2\) must react with excess CH\(_4\) to produce 8.42 grams of CO\(_2\)?

3.4 Moles and Chemical Reactions

- To make use of the chemical equation
  1. Convert mass of the known substance A to moles (if necessary)
  2. Use the coefficients in the balanced chemical equation to relate the moles of substance A to the moles of substance B
  3. Convert moles of B to mass (if necessary)

3.6 Stoichiometry

- For the following reaction:
  \[ \text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g) \]
  - How many moles of NH\(_3\) would form if 2.11 moles of N\(_2\) reacted with excess H\(_2\)?
  - How many grams of N\(_2\) are required to react completely with 9.47 grams of H\(_2\)?
3.7 Stoichiometry

- Metallic iron reacts with oxygen to form iron(III) oxide
  - Balanced eqn:

  \[ \text{Fe} + \frac{3}{2} \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3 \]

- What mass of iron needed to produce 5.00 g of product?
- What mass of oxygen is required?

3.8 Stoichiometry

- The steel industry relies on the reaction between iron(III) oxide and carbon to produce iron and CO$_2$.
  - Balanced eqn:

  \[ \text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 2\text{Fe} + 3\text{CO}_2 \]

- What mass of iron can be obtained from 454 g of iron(III) oxide?
- What mass of carbon is required to react with 454 g of iron(III) oxide?

3.9 Stoichiometry

- Ethane (C$_2$H$_6$) burns in oxygen to form CO$_2$ and water
  - Balanced eqn:

  \[ 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]

- What mass of ethane is required to produce 100. g of water?
- What mass of CO$_2$ is formed along with the 100. g of water?

3.10 Stoichiometry

- Cisplatin, Pt(NH$_3$)$_2$Cl$_2$, is a potent anti-cancer drug
  - Balanced eqn:

  \[ \text{K}_2\text{PtCl}_4 + 2\text{NH}_3 \rightarrow \text{Pt(NH}_3)_2\text{Cl}_2 + 2\text{KCl} \]

- What mass of ammonia is required to react with 16.0 g of K$_2$PtCl$_4$?
- What mass of cisplatin will form?
4.1 The Limiting Reactant

- Reactions proceed only as long as all of the reactants are present
- A car engine will run as long as both air (oxygen) and gasoline are present
- When one is absent (ever run out of gas?), the car engine stops
- In chemical reactions, very often one reactant is used up before the other
- The one that is used up is called the limiting reactant
- Once it is gone, the reaction stops
- It limits or determines the amount of product that can form

4.2 The Limiting Reactant

- Hydrogen and fluorine react according to the equation:
  \[ H_2 + F_2 \rightarrow 2 \text{HF} \]
- What happens when fluorine is consumed?
  1.
  2.
- Which reactant limits the reaction? Which reactant is “limiting”??
CH₄ + 2 O₂ → CO₂ + 2 H₂O

Now suppose that the 20.0 grams of CH₄ reacts with the 75.0 grams of O₂.
What mass of the unused reactant remains?

H₂ + C → CH₄

Suppose that 20.0 g of H₂ reacts with 40.0 g of carbon.
Which reactant is limiting?
What mass of CH₄ could form when combining these reactants?
What mass of the unused reactant is left over?

We have been calculating how much product would form for various chemical reactions.
The amount we calculate is usually higher than what we would obtain in lab.
This is not a violation of the law of conservation of matter.
Why are the two values different?
5.2 Reaction Yield

- The mass of product obtained in the reaction is the **actual yield**
- The maximum amount possible, calculated from the limiting reagent principle, is the **theoretical yield**
- The **percent yield** is defined as:

\[
\text{percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100
\]

5.3 Reaction Yield

- A chemist isolates 17.43 g of product from a reaction that has a calculated theoretical yield of 21.34 g. What is the percent yield?

5.4 Reaction Yield

- Lime (CaO) is produced by heating CaCO₃ according to the following equation:

\[
\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)
\]

- Suppose we heat 510 grams of CaCO₃ and isolate 235 grams of CaO. What is the percent yield?

5.5 No calculators allowed

- How many moles of ammonia could form by reacting 6.0 moles of hydrogen with 6.0 moles of nitrogen? The balanced equation is shown below.

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

- How many moles of carbon dioxide could form by reacting 4.0 moles of ethylene (C₂H₄) with 4.0 moles of water? The balanced equation is shown below.

\[
\text{C}_2\text{H}_4(g) + 4 \text{H}_2\text{O}(g) \rightarrow 2 \text{CO}_2(g) + 6 \text{H}_2(g)
\]
6.1 Combustion Analysis

- A 1.125 g sample of a hydrocarbon was burned to produce 3.447 g of CO₂ and 1.647 g of H₂O. Determine the empirical formula of the hydrocarbon (CₓHᵧ).

6.2 Combustion Analysis

- Thiols are foul-smelling hydrocarbons that contain sulfur. A 2.500 g sample of a thiol was burned to produce 4.880 g of CO₂, 2.497 g of H₂O, and 1.776 g of SO₂. Determine the empirical formula of the thiol (CₓHᵧSₓ).