**Molar Mass**

- Molar mass = Mass in grams of one mole of any element, numerically equal to its atomic weight
- Molar mass of molecules can be determined from the chemical formula and molar masses of elements
  - Each H₂O molecule contains 2 H atoms and 1 O atom
  - Each mole of H₂O molecules contains 2 moles of H and 1 mole of O
  - One mole of O atoms corresponds to 15.9994 g
  - Two moles of H atoms corresponds to 2 x 1.0079 g
  - Sum = molar mass = 18.0152 g H₂O per mole

**Calculation of Molar Masses**

- Calculate the molar mass of the following
  - Magnesium nitrate, Mg(NO₃)₂
    - 1 Mg = 24.3050
    - 2 N = 2 x 14.0067 = 28.0134
    - 6 O = 6 x 15.9994 = 95.9964
    - Molar mass of Mg(NO₃)₂ = 148.3148 g
  - Calcium carbonate, CaCO₃
    - 1 Ca = 40.078
    - 1 C = 12.011
    - 3 O = 3 x 15.9994
    - Molar mass of CaCO₃ = 100.087 g
  - Iron(II) sulfate, FeSO₄
    - Molar mass of FeSO₄ = 151.909 g

**Solutions**

- Solution: a homogenous mixture in which the components are evenly distributed in each other
- Solute: the component of a solution that is dissolved in another substance
- Solvent: the medium in which a solute dissolved to form a solution
- Aqueous: any solution in which water is the solvent

**Chapter 3**

**Chapter 5**

**Solutions**

- The properties and behavior of solutions often depend not only on the type of solute but also on the concentration of the solute.
  - Concentration: the amount of solute dissolved in a given quantity of solvent or solution
    - many different concentration units
      - (%, ppm, g/L, etc)
    - often expressed as Molarity
Solution Concentrations

- **Molarity** = moles of solute per liter of solution
- Designated by a capital M (mol/L)

- 6.0 M HCl
  - 6.0 moles of HCl per liter of solution.

- 9.0 M HCl
  - 9.0 moles of HCl per liter of solution.

Solution Concentrations

Molarity can be used as a conversion factor.

- The definition of molarity contains 3 quantities:
  - Molarity = \( \frac{\text{moles of solute}}{\text{volume of solution in liters}} \)
  - If you know two of these quantities, you can find the third.

Solution Concentrations

Determine the molarity of each solution

- 2.50 L of solution containing 1.25 mol of solute

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

- 225 mL of solution containing 0.486 mole of solute

- 100. mL of solution containing 2.60 g of NaCl

Strategy:

- \( g \rightarrow \text{mol} \rightarrow \text{molarity} \)

Solution Concentrations

Example: How many moles of HCl are present in 2.5 L of 0.10 M HCl?

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

Given:
- 2.5 L of soln
- 0.10M HCl

Find:
- mol HCl

Use molarity as a conversion factor

\[
\text{Mol HCl} = 2.5 \text{ L soln} \times \frac{0.10 \text{ mol HCl}}{1 \text{ L of soln}}
\]

\[
= 0.25 \text{ mol HCl}
\]
**Solution Preparation**

Example: What volume of a 0.10 M NaOH solution is needed to provide 0.50 mol of NaOH?

Given: 0.50 mol NaOH
0.10 M NaOH
Find: vol soln

*Use M as a conversion factor*

Vol soln = \( \frac{0.50 \text{ mol NaOH}}{0.10 \text{ mol NaOH}} \times 1 \text{ L soln} = 5.0 \text{ L solution} \)

**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?*

**Solution Preparation**

Example: How many grams of CuSO\(_4\) are needed to prepare 250.0 mL of 1.00 M CuSO\(_4\)?

Given: 250.0 mL solution, 1.00 M CuSO\(_4\)
Find: g CuSO\(_4\)

*Conversion factors: Molarity, molar mass*

Strategy:

\[ \text{ml} \rightarrow \text{l} \rightarrow \text{mol} \rightarrow \text{grams} \]

\[
g \text{CuSO}_4 = \frac{250.0 \text{ ml soln}}{1000 \text{ ml}} \times 1.00 \text{ mol} \times 159.6 \text{ g CuSO}_4 \]

\[ = 39.9 \text{ g CuSO}_4 \]

**Solution Preparation**

Example: What volume of a 0.10 M NaOH solution is needed to provide 0.50 mol of NaOH?

Given: 0.50 mol NaOH
0.10 M NaOH
Find: vol soln

Vol soln = \( \frac{0.50 \text{ mol NaOH}}{0.10 \text{ mol NaOH}} \times 1 \text{ L soln} = 5.0 \text{ L solution} \)
Solution Preparation

- Describe how to prepare the following:
  - 500. mL of 1.00 M FeSO₄

**Strategy:**

\[
\text{mL} \rightarrow \text{L} \rightarrow \text{mol} \rightarrow \text{grams}
\]

\[
\text{grams} = 500.0 \text{ ml} \times \frac{1 \text{ L}}{1000 \text{ ml}} \times \frac{1.00 \text{ mol}}{1 \text{ L}} \times \frac{151.909 \text{ g}}{1 \text{ mol}}
\]

\[
= 75.9545 \text{ g}
\]

- 100. mL of 3.00 M glucose

- 250. mL of 0.100 M NaCl

Solution Preparation

- Steps involved in preparing solutions from pure solids
  - Calculate the amount of solid required
  - Weigh out the solid
  - Place in an appropriate volumetric flask
  - Fill flask about half full with water and mix.
  - Fill to the mark with water and invert to mix.

Solution Preparation

- Solutions of exact concentrations can also be prepared by diluting a more concentrated solution of the solute to the desired concentration.
Many laboratory chemicals such as acids are purchased as concentrated solutions (stock solutions).
- 12 M HCl
- 12 M H₂SO₄

More dilute solutions are prepared by taking a certain quantity of the stock solution and diluting it with water.

Solution Preparation

A given volume of a stock solution contains a specific number of moles of solute.
- 25 mL of 6.0 M HCl contains 0.15 mol HCl
  (How do you know this???)

If 25 mL of 6.0 M HCl is diluted with 25 mL of water, the number of moles of HCl present does not change.
- Still contains 0.15 mol HCl

Moles solute \( \frac{\text{before dilution}}{\text{after dilution}} \) moles solute

Although the number of moles of solute does not change, the volume of solution does change.

The concentration of the solution will change since Molarity = \( \frac{\text{mol solute}}{\text{volume solution}} \)

When a solution is diluted, the concentration of the new solution can be found using:
\[
M_c \cdot V_c = M_d \cdot V_d
\]
Where,
- \( M_c \) = concentration of concentrated solution (mol/L)
- \( V_c \) = volume of concentrated solution
- \( M_d \) = concentration of diluted solution (mol/L)
- \( V_d \) = volume of diluted solution
Solution Preparation

Example: What is the concentration of a solution prepared by diluting 25.0 mL of 6.00 M HCl to a total volume of 50.0 mL?

Given: \( V_c = 25.0 \text{ mL} \)
\( M_c = 6.00 \text{ M} \)
\( V_d = 50.0 \text{ mL} \)

Find: \( M_d \)

\[
M_c \cdot V_c = M_d \cdot V_d
\]

\[
6.00 \text{ M} \times 25.0 \text{ mL} = M_d \times 50.0 \text{ mL}
\]

\[
M_d = \frac{6.00 \text{ M} \times 25.0 \text{ mL}}{50.0 \text{ mL}} = 3.00 \text{ M}
\]

Note: \( V_c \) and \( V_d \) do not have to be in liters, but they must be in the same units.

Solution Preparation

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

Given: \( M_c = 6.00 \text{ M} \)
\( M_d = 0.250 \text{ M} \)
\( V_d = 500.0 \text{ mL} \)

Find: \( V_c \)

\[
M_c \cdot V_c = M_d \cdot V_d
\]

\[
6.00 \text{ M} \times V_c = 0.250 \text{ M} \times 500.0 \text{ mL}
\]

\[
V_c = \frac{0.250 \text{ M} \times 500.0 \text{ mL}}{6.00 \text{ M}}
\]

Concentrations of Acids

➢ The pH scale

\[
\text{pH} = -\log [\text{H}^+]
\]

- pH of vinegar = \(-\log (1.6 \times 10^{-3} \text{ M}) = -(-2.80) = 2.80\)
- pH of pure water = \(-\log (1.0 \times 10^{-7} \text{ M}) = -(-7.00) = 7.00\)
- pH of blood = \(-\log (4.0 \times 10^{-8} \text{ M}) = -(-7.40) = 7.40\)
- pH of ammonia = \(-\log (1.0 \times 10^{-11} \text{ M}) = -(-11.00) = 11.00\)

➢ Lower the concentration of \( \text{H}^+ \), higher the pH

acidic substance, \( \text{pH} < 7 \)
Basic substance, \( \text{pH} > 7 \)
neutral, \( \text{pH} = 7 \)
On serial dilution of acid solution, the pH increases because the concentration of \( H^+ \) ions decreases with dilution:

<table>
<thead>
<tr>
<th>Concentration</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.1M HCl</td>
<td>1</td>
</tr>
<tr>
<td>0.01 M HCl</td>
<td>2</td>
</tr>
<tr>
<td>0.001 M HCl</td>
<td>3</td>
</tr>
</tbody>
</table>

The \( H^+ \) concentration of a solution of known pH can be calculated using the following equation:

\[
[H^+] = 10^{-pH}
\]

Calculate pH of 0.0065 M HCl solution.

Calculate the concentration of \( H^+ \) ion in a solution of pH 7.5.

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?

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?
Solution Preparation Review

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

Conversion Factors

\[
\text{Number of particles} \quad \text{Moles} \quad \text{Mass}
\]

- Avogadro’s number: \(6.022 \times 10^{23}\)
- Molar mass

Molar Conversions

- Determine the following:
  - The moles of potassium atoms in a 50.0 g sample
  \[
  \text{Grams} \times \frac{1 \text{ mol}}{\text{grams}} = \text{moles}
  \]
  - The mass of Mg in a 1.82 mole sample
  \[
  \text{Moles} \times \frac{\text{grams}}{1 \text{ mol}} = \text{grams}
  \]

- The moles of FeCl\(_3\) in a 50.0 g sample

- The mass of MgCl\(_2\) in a 2.75 mole sample
Molar Conversions

- glucose = C₆H₁₂O₆ (Molar mass = 180.1 g)
- How many moles of glucose are contained in 70.0 g of C₆H₁₂O₆?
- How many moles of oxygen are contained in 70.0 g of C₆H₁₂O₆?
- How many grams of H are contained in 70.0 g of C₆H₁₂O₆?

\[
\text{Moles} = \frac{\text{grams}}{\text{molar mass}} \times 1 \text{ mol grams} = \text{moles}
\]

Molar Ratios

- The relative number of moles of each element in a substance can be used as a conversion factor called the molar ratio.
- Molar ratio = \( \frac{\text{moles element A}}{\text{mole of substance}} \)
  or
- Molar ratio = \( \frac{\text{moles element A}}{\text{moles element B}} \)

Molar Ratios

- \( \text{C}_6\text{H}_{12}\text{O}_6 \):
  - Molar Ratio = \( \frac{6 \text{ moles of C}}{1 \text{ mole } \text{C}_6\text{H}_{12}\text{O}_6} \)
  - Molar Ratio = \( \frac{12 \text{ moles H}}{1 \text{ mole } \text{C}_6\text{H}_{12}\text{O}_6} \)
  - Molar Ratio = \( \frac{6 \text{ moles O}}{1 \text{ mole } \text{C}_6\text{H}_{12}\text{O}_6} \)
  - Molar Ratio = \( \frac{6 \text{ moles C}}{12 \text{ moles H}} = \frac{1 \text{ mol C}}{2 \text{ moles H}} \)
Molar Conversions

➢ Molar ratios can be used to determine the number of moles of a particular element in a given substance.

➢ How many moles of oxygen are contained in 70.0 g of \( \text{C}_6\text{H}_{12}\text{O}_6 \)?

\[
\text{Moles of O} = 70.0 \text{ g of C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.1 \text{ g}} \times 6 \text{ mol O/mol C}_6\text{H}_{12}\text{O}_6
\]

\[= 2.33 \text{ mol O}\]

Molar Conversions

➢ Once the number of moles of a substance present is known, we can use:
  – Molar mass to find the number of grams
  – Avogadro’s number to find the number of atoms, ions, or molecules

➢ How many grams of H are contained in 70.0 g of \( \text{C}_6\text{H}_{12}\text{O}_6 \)?
Molar Conversions

Given: 70.0 g of $C_6H_{12}O_6$
Find: gram H

$$g\ H = \frac{70.0\ g\ C_6H_{12}O_6 \times 1\ mol\ C_6H_{12}O_6 \times 12\ mol\ H}{180.1\ g\ mol\ C_6H_{12}O_6 \times 1.0079\ g\ mol\ H}$$

$$= 4.70\ g\ H$$

Molar Conversions

- How many moles of nitrogen are contained in 70.0 g of $C_6H_5NO_3$?
- How many grams of oxygen are contained in 1.5 moles of $C_6H_5NO_3$?
- How many atoms of C are contained in 70.0 g of $C_6H_5NO_3$?

Molar Conversions

- How many moles of $H^+$ ions are present in 2.5 moles of $H_2SO_4$?

Given: 2.5 moles $H_2SO_4$
Find: moles $H^+$

Conversion factor: molar ratio

$$\text{Mol H}^+ = \frac{2.5\ mol\ H_2SO_4 \times 2\ mol\ H^+}{1\ mol\ H_2SO_4}$$

$$\text{Mol H}^+ = 5.0\ mol\ H^+$$

Empirical and Molecular Formulas

- **Empirical formula**: smallest whole-number ratio of atoms present in a compound
- **Molecular formula**: actual number of each type of atom present in a given compound

Many molecular compounds have different empirical and molecular formulas:

<table>
<thead>
<tr>
<th>Molecular Formula</th>
<th>Empirical Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>$C_6H_6$</td>
<td>CH</td>
</tr>
<tr>
<td>$C_6H_{12}O_6$</td>
<td>$CH_2O$</td>
</tr>
<tr>
<td>$N_2H_4$</td>
<td>NH$_2$</td>
</tr>
<tr>
<td>$N_2O_4$</td>
<td>NO$_2$</td>
</tr>
</tbody>
</table>
Percent Composition

- Empirical formulas are generally obtained by determining the percent composition of a compound:

  - Percent composition:
    - the percentage of the mass contributed by each element in a substance
  

  \[ \% \text{ Element } X = \left( \frac{\# \text{ atoms of } X}{\text{FW of compound}} \right) \times 100\% \]

**Formula Weight (FW):** the sum of the atomic weights of all of the atoms in a chemical formula

*Note: Molar mass is a mass in grams that is numerically the same as the formula weight*

---

Mass Percentage

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

  \[
  \text{FW} = 12.011 \text{ amu} + 4(1.0079 \text{ amu}) + 2(14.0067 \text{ amu}) + 1(15.9994 \text{ amu}) \\
  = 60.0554 \text{ amu}
  \]

  \[
  \% \text{ C} = \frac{1(12.011 \text{ amu})}{60.0554 \text{ amu}} \times 100\% = 20.00\% \text{ C}
  \]

  \[
  \% \text{ H} = \frac{4(1.0079 \text{ amu})}{60.0554 \text{ amu}} \times 100\% = 6.71\% \text{ H}
  \]

  \[
  \% \text{ N} = \frac{2(14.0067 \text{ amu})}{60.0554 \text{ amu}} \times 100\% = 46.65\% \text{ C}
  \]

  \[
  \% \text{ O} = \frac{1(15.9994 \text{ amu})}{60.0554 \text{ amu}} \times 100\% = 26.64\% \text{ H}
  \]

---

Calculating Empirical Formulas

- Percent composition data is commonly used to determine the empirical formula of a compound.

  **Empirical formula:** smallest whole-number ratio of atoms present in a compound

  **Percent composition:** the percentage of the mass contributed by each element in a substance

- Four steps:
  - percent to mass
  - mass to mole
  - divide by smallest
  - multiple 'til whole
Calculating Empirical Formulas

Step 1: % to mass
If you assume 100.00 g of substance, then
74.03 % C \( \Rightarrow \) 74.03 g C
8.70 % H \( \Rightarrow \) 8.70 g H
17.27 % N \( \Rightarrow \) 17.27 g N

Step 2: Mass to moles
# moles C = \( \frac{74.03 \text{ g C}}{12.01 \text{ g C}} \times 1 \text{ mole C} \) = 6.164 mol C
# moles H = \( \frac{8.70 \text{ g H}}{1.008 \text{ g H}} \times 1 \text{ mole H} \) = 8.631 mol H
# moles N = \( \frac{17.27 \text{ g N}}{14.00 \text{ g N}} \times 1 \text{ mole N} \) = 1.234 mol N

Step 3: Divide by smallest
(\text{this gives the molar ratio of the elements})
C: \( \frac{6.164 \text{ moles C}}{1.234 \text{ moles N}} = 4.995 \)
H: \( \frac{8.631 \text{ mol H}}{1.234 \text{ mol N}} = 6.994 \)
N: \( \frac{1.234 \text{ mol N}}{1.234 \text{ mol N}} = 1.000 \)

Step 4: Multiply ‘til Whole
(\text{this step is necessary only when step 3 does not give whole number molar ratios})
C: \( \frac{6.164 \text{ moles C}}{1.234 \text{ moles C}} \times 1 = 5.000 \text{ mol C} \)
H: \( \frac{8.631 \text{ mol H}}{1.234 \text{ mol C}} \times 1 = 7.000 \text{ mol H} \)
N: \( \frac{1.234 \text{ mol N}}{1.234 \text{ mol C}} \times 1 = 1.000 \text{ mol N} \)
Calculating Empirical Formulas

Use these numbers to write the empirical formula:

\[
\begin{align*}
5.000 \text{ mol C} \\
7.000 \text{ mol H} \\
1.000 \text{ mol N}
\end{align*}
\]

\(\text{C}_3\text{H}_7\text{N}\)

Calculating Empirical Formulas

An iron compound contains 69.943 % Fe and 30.057 % O. Calculate its empirical formula.

Step 1: % to mass

- 69.943 % Fe \(\rightarrow\) 69.943 g Fe
- 30.057 % O \(\rightarrow\) 30.057 g O

Step 2: Mass to Moles

\[
\begin{align*}
\text{mol Fe} &= \frac{69.943 \text{ g Fe}}{55.845 \text{ g Fe}} \times \text{mol Fe} = 1.2524 \text{ mol Fe} \\
\text{mol O} &= \frac{30.057 \text{ g O}}{15.9994 \text{ g O}} \times \text{mol O} = 1.8786 \text{ mol O}
\end{align*}
\]

Calculating Empirical Formulas

Step 3: Divide by smallest

\[
\begin{align*}
\text{O}: \quad \frac{1.8786 \text{ mol O}}{1.2524 \text{ mol Fe}} &= 1.5000 \\
\text{Fe}: \quad 1.2524 \text{ mol Fe} &= 1.0000
\end{align*}
\]
Calculating Empirical Formulas

- **Step 4:** Multiply ‘til whole

FeO\text{1.5} doesn’t make sense.

\[
\begin{align*}
1.5000 \text{ mol O} \times 2 &= 3.000 \text{ mol O} \\
1.0000 \text{ mol Fe} \times 2 &= 2.000 \text{ mol Fe}
\end{align*}
\]

Fe\text{2O}_3

---

Using Empirical Formulas to Find Molecular Formulas

- **Steps:**
  - Find the empirical formula
  - Calculate the formula weight for the empirical formula.
  - Get the whole number ratio between MW and FW
  - Multiply the subscripts in the empirical formula by the whole number ratio.

Using Empirical Formulas to Find Molecular Formulas

- **Step 1:**
  - Empirical formula of nicotine: C_{5}H_{7}N

- **Step 2:**
  - Formula Weight = 81.06

Molar mass = Molecular weight = 162.23

- **Step 3:**
  - Ratio = \frac{MW}{FW} = \frac{162.23}{81.06} = 2.00

- **Step 4:**
  - Molecular formula: C_{(5x2)}H_{(7x2)}N_{(1x2)} = C_{10}H_{14}N_{2}
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.

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.