Announcements

HOUR EXAM 1              July 18  6-7:30PM

--Want me to do recitation again?

--Skip Combustion Analysis & Isomers (p.82-83 in Principles of Chemistry Text)

See me if you don’t understand!

Chapter 3: 6, 10, 12, 14, 16, 18, 20, 23, 26, 28, 30, 32, 33, 38, 40, 41, 43, 45, 49, 51, 65, 66, 69, 71, 73, 75, 83, 85, 89, 93, 95, 3.119 (Principles of Chemistry)


Learning Objectives

1. Understand relative atomic masses, average isotopic mass.
2. Connect the dots between amu & grams & the periodic table via the mole.
3. Compute a molecular & molar mass of a substance from a formula.
4. Using factor label method to convert between grams<=>moles<=>molecules
5. % Mass To Empirical Formula and Vis-versa
6. Balancing equations and mastering stoichiometry
7. Limiting Reagent, Yields, Solution Stoichiometry

A chemical formula determines the % mass of each element in a compound.

\[
\text{n x molar mass of element} \quad \frac{\text{molar mass of compound}}{\text{x 100%}}
\]

\(n\) is the number of moles of an element in 1 mole of the compound.

\[
\%C = \left(\frac{2 \times (12.01 \text{ g})}{60.05 \text{ g}}\right) \times 100\% = 40.0\%
\]

\[
\%H = \left(\frac{4 \times (1.008 \text{ g})}{60.05 \text{ g}}\right) \times 100\% = 6.714\%
\]

\[
\%O = \left(\frac{2 \times (16.00 \text{ g})}{60.05 \text{ g}}\right) \times 100\% = 53.28\%
\]

40.0% + 6.71% + 53.2% = 100.0%

A laboratory technique called “elemental analysis” can determine the % by mass of each element in a compound. We can compute the “empirical formula” of any compound with that information.

\[
\begin{align*}
\text{Molar Mass} & \quad \text{Empirical Formula} \\
\text{% Mass Element} & \quad \text{Empirical Formula} \\
\%C = a\% & \quad C_xH_yO_z \\
\%H = b\% & \\
\%O = c\% \\
\end{align*}
\]

Empirical and Molecular Formulas

Molecular Formula

The formula of a compound as it actually exists according to experimental data. It is a multiple of the empirical formula.

<table>
<thead>
<tr>
<th>Formula</th>
<th>CH₂O</th>
<th>C₂H₄O₂</th>
<th>C₃H₆O₃</th>
<th>C₄H₈O₄</th>
<th>C₅H₁₀O₅</th>
</tr>
</thead>
<tbody>
<tr>
<td>Molar Mass</td>
<td>30.02</td>
<td>60.05</td>
<td>90.08</td>
<td>120.10</td>
<td>150.13</td>
</tr>
</tbody>
</table>

Empirical Formula

The simplest formula for a compound that gives rise to the smallest set of whole numbers of atoms.

\(\text{CH}_2\text{O}\)
Compounds that have the same % mass of its elements have the same empirical formula!

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular Formula</th>
<th>Whole-Number Multiple</th>
<th>M (g/mol)</th>
<th>Use or Function</th>
</tr>
</thead>
<tbody>
<tr>
<td>formaldehyde</td>
<td>CH₂O</td>
<td>1</td>
<td>30.03</td>
<td>disinfectant; biological preservative</td>
</tr>
<tr>
<td>acetic acid</td>
<td>C₂H₄O₂</td>
<td>2</td>
<td>60.05</td>
<td>acetate polymers; vinegar(5% soln)</td>
</tr>
<tr>
<td>lactic acid</td>
<td>C₃H₆O₃</td>
<td>3</td>
<td>90.09</td>
<td>sour milk; forms in exercising muscle</td>
</tr>
<tr>
<td>erythrose</td>
<td>C₄H₈O₄</td>
<td>4</td>
<td>120.10</td>
<td>part of sugar metabolism</td>
</tr>
<tr>
<td>ribose</td>
<td>C₅H₁₀O₅</td>
<td>5</td>
<td>150.13</td>
<td>component of nucleic acids and B₆</td>
</tr>
<tr>
<td>glucose</td>
<td>C₆H₁₂O₆</td>
<td>6</td>
<td>180.16</td>
<td>major energy source of the cell</td>
</tr>
</tbody>
</table>

All have the same % by mass:
40.0% C      6.71% H      53.3% O.

Determining the Empirical Formula from Masses of Elements (Elemental Analysis)

Elemental analysis of a sample of an ionic compound gave the following results: 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. Determine the empirical formula and the name of this compound?

**SOLUTION:**

\[ \begin{align*}
2.82 \text{ g Na} & \Rightarrow \frac{\text{mol Na}}{22.99 \text{ g Na}} = 0.123 \text{ mol Na} \\
4.35 \text{ g Cl} & \Rightarrow \frac{\text{mol Cl}}{35.45 \text{ g Cl}} = 0.123 \text{ mol Cl} \\
7.83 \text{ g O} & \Rightarrow \frac{\text{mol O}}{16.00 \text{ g O}} = 0.489 \text{ mol O} \\
\end{align*} \]

\[ \text{Na}_1 \text{Cl}_1 \text{O}_{3.98} \Rightarrow \text{NaClO}_4 \]

\( \text{NaClO}_4 \) is sodium perchlorate.

---

Determining the Empirical Formula from Mass % Data

**Problem Solving Method**

1. **Mass percent**
2. **Moles of each element**
3. **Mole ratios of elements**
4. **Empirical formula**
5. **Molecular formula**

**Step 1:** Determine the mass of each element. Assume a 100 gram sample and use the given data we have:
- C 62.58 g
- H 9.63 g
- O 27.79 g

**Step 2:** Convert masses to amounts in moles.
\[ \begin{align*}
n_C & = 62.58 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 5.210 \text{ mol C} \\
n_H & = 9.63 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.55 \text{ mol H} \\
n_O & = 27.79 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 1.737 \text{ mol O} \\
\end{align*} \]

**Step 3:** Write a tentative empirical formula. \( C_{5.21H_{9.55}O_{1.74}} \)

**Step 4:** Convert to small whole numbers.
Divide by smallest number of moles \( C_{2.99H_{5.46}O} \)

**Step 5:** Convert to a small whole number ratio. Multiply by 2 to get \( C_{5.98H_{10.99}O_2} \)

The empirical formula is \( C_{5.98H_{10.99}O_2} \)

**Empirical formula mass** = 6(12.01) + 1.008(11) + 2 (16.00) = 115 u.

**Step 6:** Now using the empirical formula mass and molecular mass together determine the molecular formula. Empirical formula mass is 115 u.
Molecular formula mass is 230 u.
\[ n = \text{Molecular mass/empirical mass} = \frac{230 \text{ amu}}{115 \text{ amu}} = 2 \]

The molecular formula is \( C_{12H_{22}O_4} \)
Determining a Molecular Formula from Elemental Analysis and Molar Mass

During physical activity, lactic acid (\(M = 90.08 \text{ g/mol}\)) forms in muscle tissue and is responsible for muscle soreness at fatigue. Elemental analysis shows that this compound contains 40.0 mass% C, 6.71 mass% H, and 53.3 mass% O.

(a) Determine the empirical formula of lactic acid.

(b) Determine the molecular formula.

Chemical equations are symbolic representations of “what happens” in a chemical reaction.

\[
\begin{align*}
\text{Reactants} & \quad \text{“Yields”} \quad \text{Products} \\
\text{CH}_4(g) + 2\text{O}_2(g) & \quad \rightarrow \quad \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \\
\end{align*}
\]

- Balanced equations = Conservation of Mass!
- Balanced equations = Correct Stoichiometry!
- Phases specified: (s) for solids, (l) for liquids, (g) for gases, (aq) for aqueous.

Only balanced chemical equations contain useful stoichiometric conversion factors.

\[
\begin{align*}
\text{NH}_3 + \text{O}_2 & \quad \rightarrow \quad \text{NO} + \text{H}_2 \\
\text{Not Balanced} \quad & \quad 1 \text{ moles NH}_3 \neq 1 \text{ moles O}_2 \\
6\text{NH}_3 + 3\text{O}_2 & \quad \rightarrow \quad 6\text{NO} + 9\text{H}_2 \quad \text{Balance first!} \\
\end{align*}
\]

Correct Stoichiometric Conversion Factors

\[
\begin{align*}
6 \text{ mol NH}_3 = 3 \text{ mol O}_2 & \quad 6 \text{ mol NH}_3 = 6 \text{ mol NO} \\
6 \text{ mol NH}_3 = 9 \text{ mol H}_2 & \quad 3 \text{ mol O}_2 = 6 \text{ mol NO} \\
3 \text{ mol O}_2 = 9 \text{ mol H}_2 & \quad 6 \text{ mol NO} = 9 \text{ mol H}_2 \\
\end{align*}
\]

We have to know how to balance equations! It’s trial and error!

\[
\begin{align*}
\_\text{Na}_3\text{PO}_4(\text{aq}) + \_\text{HCl}(\text{aq}) & \quad \rightarrow \_\text{H}_2\text{PO}_4(\text{aq}) + \_\text{NaCl}(\text{aq}) \\
\_\text{Ba(OH)}_2(\text{aq}) + \_\text{HCl}(\text{aq}) & \quad \rightarrow \_\text{H}_2\text{O(l)} + \_\text{BaCl}_2(\text{aq}) \\
\_\text{CrH}_4 + \_\text{O}_2 & \quad \rightarrow \_\text{H}_2\text{O(l)} + \_\text{CO}_2(\text{g}) \\
\end{align*}
\]
Solutions

\[ \text{Na}_3\text{PO}_4(\text{aq}) + 3\text{HCl}(\text{aq}) \rightarrow \text{H}_3\text{PO}_4(\text{aq}) + 3\text{NaCl}(\text{aq}) \]

\[ \text{Ba(OH)}_2(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)} + \text{BaCl}_2(\text{aq}) \]

\[ 2\text{C}_7\text{H}_{14} + 21\text{O}_2 \rightarrow 14\text{H}_2\text{O}(l) + 14\text{CO}_2(g) \]

Some are tougher than others......practice!

Example: Ethane, \( \text{C}_2\text{H}_6 \), reacts (is combusted are key words) with \( \text{O}_2 \) to form \( \text{CO}_2 \) and \( \text{H}_2\text{O} \). Write a balanced equation for this reaction.

1. Write the correct formula(s) for reactants and products.
   \[ \text{C}_2\text{H}_6 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

2. Start by balancing those elements that appear in only one reactant and one product.
   \[ \begin{array}{c}
   \text{C}_2\text{H}_6 + \text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} \\
   \text{6 hydrogen} \quad \text{2 hydrogen} \quad \text{multiply H}_2\text{O} \quad \text{by 3}
   \end{array} \]

3. Balance those elements that appear in two or more reactants or products.
   \[ \begin{array}{c}
   \text{C}_2\text{H}_6 + \frac{7}{2}\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} \\
   \text{2 O on left} \quad \text{4 O} \quad \text{3 O} \quad \text{= 7 oxygen on right} \\
   \text{multiply O}_2 \quad \text{by} \quad \frac{7}{2}
   \end{array} \]

4. Check to make sure that you have the same number of each type of atom on both sides of the equation.
   \[ 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]

To learn how to balance equations quickly for Exams you have to practice! Try it.

\[ \text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe(s)} + 3\text{CO}_2(\text{g}) \]

\[ \text{Na}_2\text{SO}_4(\text{s}) + \text{C(\text{s})} \rightarrow \text{Na}_2\text{S(s)} + \text{CO}_2(\text{g}) \]

\[ \text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2(\text{g}) \]

\[ \text{Mg} + \text{H}_2\text{O(l)} \rightarrow \text{Mg(OH)}_2(\text{s}) + \text{NH}_3(\text{g}) \]

\[ \text{H}_2\text{S(g)} + \text{SO}_2(\text{g}) \rightarrow \text{S(s)} + \text{H}_2\text{O(l)} \]

A balanced chemical equation contains all the information we need to do many calculations!

There is a lot of information here!

- 1 mol \( \text{Fe}_2\text{O}_3 \) \( \rightarrow \) 3 mol CO
- 1 mol \( \text{Fe}_2\text{O}_3 \) \( \rightarrow \) 3 mol \( \text{CO}_2 \)
- 3 mol \( \text{CO} \) \( \rightarrow \) 2 mol Fe
- 3 mol \( \text{CO} \) \( \rightarrow \) 3 mol \( \text{CO}_2 \)

\[ \text{not balanced} = \text{NOT RIGHT!} \]

\[ \text{Fe}_2\text{O}_3(\text{s}) + \text{CO(g)} \rightarrow \text{Fe(s)} + \text{CO}_2(\text{g}) \]
Steps to mastering stoichiometry!

1. Always write a balanced chemical equation.
2. Work in moles—not masses....we need to count?
3. Use dimensional analysis correctly.

<table>
<thead>
<tr>
<th>Grams of</th>
<th>Molar Mass</th>
<th>Grams of</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reactant</td>
<td>balanced equation</td>
<td>Product</td>
</tr>
<tr>
<td>Moles of Reactant</td>
<td>balanced equation</td>
<td>Moles of Product</td>
</tr>
</tbody>
</table>

Iron III oxide reacts with carbon monoxide as shown below. How many CO molecules are required to react with 25 formula units of Fe₂O₃ as shown below in the balanced equation?

Fe₂O₃(s) + CO(g) → Fe(s) + CO₂(g)

How many CO molecules are required to react with 25 formula units of Fe₂O₃ as shown below in the balanced equation?

Fe₂O₃(s) + 3CO(g) → 2Fe(s) + 3CO₂(g)

How many CO molecules are required to react with 25 formula units of Fe₂O₃ as shown below in the balanced equation?

Fe₂O₃(s) + 3CO(g) → 2Fe(s) + 3CO₂(g)

How many iron atoms can be produced by the reaction of 2.50 x 10⁵ formula units of iron (III) oxide with excess carbon monoxide?

Fe₂O₃(s) + 3CO(g) → 2Fe(s) + 3CO₂(g)

# Fe atoms =

= 2.50 x 10⁵ fu Fe₂O₃ x \( \frac{2 \text{ Fe atoms}}{1 \text{ fu Fe₂O₃}} \) = 5.00 x 10⁵ Fe atoms

What mass of CO is required to react with 146 g of iron (III) oxide?

Fe₂O₃(s) + 3CO(g) → 2Fe(s) + 3CO₂(g)

? g CO = 146 g Fe₂O₃ x \( \frac{1 \text{ mol Fe₂O₃}}{159.7 \text{ g Fe₂O₃}} \) x \( \frac{3 \text{ mol CO}}{1 \text{ mol Fe₂O₃}} \) x \( \frac{28.0 \text{ g CO}}{1 \text{ mol CO}} \) = 76.8 g CO
What mass (in grams) of iron (III) oxide reacted with excess carbon monoxide if the carbon dioxide produced by the reaction had a mass of 8.65 grams?

Fe₂O₃ (s) + 3CO(g) → 2Fe(s) + 3CO₂(g)

? g Fe₂O₃ = 8.65 g CO₂ × \frac{1 \text{ mol CO₂}}{44.0 \text{ g CO₂}} × \frac{1 \text{ mol Fe₂O₃}}{3 \text{ mol CO₂}} × \frac{159.7 \text{ g Fe₂O₃}}{1 \text{ mol Fe₂O₃}} = 10.5 \text{ g Fe₂O₃}

What mass of carbon dioxide can be produced by the reaction of 0.540 mole of iron (III) oxide with excess carbon monoxide?

Fe₂O₃ (s) + 3CO(g) → 2Fe(s) + 3CO₂(g)

? g CO₂ = 0.540 \text{ mol Fe₂O₃} × \frac{3 \text{ mol CO₂}}{1 \text{ mol Fe₂O₃}} × \frac{44.0 \text{ g CO₂}}{1 \text{ mol CO₂}} = 71.3 \text{ g CO₂}

The **limiting reagent** is the reactant that runs out (is consumed) and determines the quantity or amount of product that can be formed.

We need to “count” the number of reactants to figure out what will limit the amount of cars produced— it’s not how much they weigh!

Do You Understand Limiting Reagent II?

If 25.0 g CH₄ is combusted with 40.0 g O₂, which reactant is the limiting reactant?

Notice the absence of “excess” and there are two reactant masses in the problem = limiting reagent!

There are two ways you can calculate the answer.

**Method 1:** For both reactants, use balanced equation & implied stoichiometric factors and compute the amount of any product formed (I prefer and teach this method!)

**Method 2:** Pick one of the reactant and compute how much of the other reactant you need and compare with
• If 25.0 g CH\textsubscript{4} is combusted with 40.0 g O\textsubscript{2}, which reactant is the limiting reactant?

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

**Method 1** (Calculating the # moles of product formed by each reactant to determine which reactant makes the least amount)

1. Let’s use the amount of CO\textsubscript{2} formed as our “yardstick” of how much product can be made (we could chose H\textsubscript{2}O).

\[
\begin{align*}
\text{mol CO}_2 &= 25.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} = 1.559 \text{ mol CO}_2 \\
\text{mol CO}_2 &= 40.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol O}_2} = 0.0625 \text{ mol CO}_2
\end{align*}
\]

O\textsubscript{2} must be the limiting reagent as the amount of CO\textsubscript{2} produced is the least amount of product!

**Method 11** (Directly comparing amounts of reactants given in the problem to which is the limiting reagent--less steps)

\[
\begin{align*}
g \text{ O}_2 \text{ needed} &= 25.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{2 \text{ mol O}_2}{1 \text{ mol CH}_4} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 99.75 \text{ g O}_2
\end{align*}
\]

O\textsubscript{2} is the limiting reagent as we need 99.75 g of it but we are are only given 40.0 g O\textsubscript{2}! Thus, the amount of product that can be formed is determined by the amount of O\textsubscript{2} not by the amount of methane, CH\textsubscript{4}.

• If 25.0 g CH\textsubscript{4} is combusted with 40.0 g O\textsubscript{2}, which reactant is the limiting reactant?

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