Determination of the Empirical Formula of Magnesium Oxide

Experiment 2

Goal:
- To study stoichiometric relationships governing mass and amount in chemical formulas.

Method:
- Measure masses before and after the oxidation reaction of Mg metal and O$_2$ gas
- Use masses to calculate the experimental empirical formula of magnesium oxide
- Compare the experimental empirical formula to the theoretical empirical formula
Practical Chemistry

Atoms / molecules: hard to count
Formulas / equations: written based on numerical relationships (not masses)

By using mass relationship to moles:

- Make something “unmeasurable” measurable.
- Relate numbers of atoms and molecules to “real world” values (masses).

Overview of Concepts

Law of conservation of mass
- Total mass of products must equal total mass of reactants

Law of constant composition
- Any portion of compound has same ratio of masses as elements in compound (constant mass ratio)

Molecular composition expressed 3 ways
- Mass of each element per mole of compound
- Mass of each element present relative to total mass (mass %)
- # each type of atom of per molecule/formula unit (formula)
Empirical and Molecular Formulas

Empirical Formula
- “Simplest” formula
- Agrees with elemental analysis
- Smallest set of whole # ratio of atoms

Molecular Formula
- May be same as empirical “Na$_2$Cl$_2$” = NaCl
- May be multiple of empirical H$_2$O$_2$ ≠ HO

<table>
<thead>
<tr>
<th>Name</th>
<th>Actual Formula</th>
<th>Multiple</th>
<th>Molar mass</th>
<th>Use or function</th>
</tr>
</thead>
<tbody>
<tr>
<td>Formaldehyde</td>
<td>CH$_2$O</td>
<td>1</td>
<td>30</td>
<td>Disinfectant, bio preservative</td>
</tr>
<tr>
<td>Acetic Acid</td>
<td>C$_2$H$_4$O$_2$</td>
<td>2</td>
<td>60</td>
<td>Polymers, vinegar</td>
</tr>
<tr>
<td>Lactic Acid</td>
<td>C$_3$H$_6$O$_3$</td>
<td>3</td>
<td>90</td>
<td>Sours milk, product of exercise</td>
</tr>
<tr>
<td>Erythrose</td>
<td>C$_3$H$_6$O$_3$</td>
<td>4</td>
<td>120</td>
<td>Forms during sugar metabolism</td>
</tr>
<tr>
<td>Ribose</td>
<td>C$_3$H$_6$O$_5$</td>
<td>5</td>
<td>150</td>
<td>In many nucleic acids &amp; vit B$_{12}$</td>
</tr>
<tr>
<td>Glucose</td>
<td>C$<em>6$H$</em>{12}$O$_6$</td>
<td>6</td>
<td>180</td>
<td>Nutrient for energy</td>
</tr>
</tbody>
</table>
Reaction Summary

Heating Mg metal in air:
1) $\text{Mg}_{(s)} + \text{N}_2(g) + \text{O}_2(g) \rightarrow \text{MgO}_{(s)} + \text{Mg}_3\text{N}_2_{(s)}$

Converting $\text{Mg}_3\text{N}_2$ (add $\text{H}_2\text{O}$):
2) $\text{MgO}_{(s)} + \text{Mg}_3\text{N}_2_{(s)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{MgO}_{(s)} + \text{Mg(OH)}_2_{(s)} + \text{NH}_3_{(g)}$

Heating (conversion to magnesium oxide):
3) $\text{MgO}_{(s)} + \text{Mg(OH)}_2(s) \rightarrow \text{Mg}_x\text{O}_y(s) + \text{H}_2\text{O}_{(g)}$

Crucible Set-up

Assume everything is hot once flames are on
Procedure

Heat empty crucible + lid
- **Mass** cool crucible + lid

Obtain ~0.3 g Mg ribbon
- **Mass** crucible + lid + Mg
Place lid with ~0.5cm gap

Heat gently ~1 min; strongly ~10 min
Cool; add ~ 1mL distilled H₂O

Heat gently ~2-3 min; strongly ~5 min
- **Mass** cool crucible + lid

Mass balance and moles

**Final product:** MgₓOᵧ

- Mass Mg:
  \[ m_{Mg} = m_{cr+lid+Mg} - m_{cr+lid} \]

- Mass O:
  \[ m_{O} = m_{Mg,O} - m_{Mg} \]

**Empirical formula:**

- lowest whole-# ratio Mg:O
  \[ MgₓOᵧ \]

  \[ mol_{Mg} = \frac{mass_{Mg}}{MM_{Mg}} \]
  \[ mol_{O} = \frac{mass_{O}}{MM_{O}} \]
Example Data

<table>
<thead>
<tr>
<th>Element or Compound</th>
<th>Molar Mass (g/mol)</th>
<th>%Mg (m/m)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg</td>
<td>24.31</td>
<td>---</td>
</tr>
<tr>
<td>O</td>
<td>15.98</td>
<td>---</td>
</tr>
<tr>
<td>N</td>
<td>14.01</td>
<td>---</td>
</tr>
<tr>
<td>H</td>
<td>1.00</td>
<td>---</td>
</tr>
<tr>
<td>MgO</td>
<td>40.31</td>
<td>60.31</td>
</tr>
<tr>
<td>Mg(OH)₂</td>
<td>58.33</td>
<td>41.68</td>
</tr>
<tr>
<td>Mg₃N₂</td>
<td>100.95</td>
<td>72.34</td>
</tr>
</tbody>
</table>

Using 0.3000 g Mg

If product is:  

<table>
<thead>
<tr>
<th>Theoretical yield (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>MgO</td>
</tr>
<tr>
<td>Mg(OH)₂</td>
</tr>
<tr>
<td>Mg₃N₂</td>
</tr>
</tbody>
</table>

"Mg-ideal"  
"Mg-poor"  
"Mg-rich"

Theoretical

<table>
<thead>
<tr>
<th>MgO (g)</th>
<th>0.3000</th>
</tr>
</thead>
<tbody>
<tr>
<td>MgO (mol)</td>
<td>0.0123</td>
</tr>
<tr>
<td>MgO (g/mol)</td>
<td>0.4974</td>
</tr>
</tbody>
</table>

Experimental

<table>
<thead>
<tr>
<th>MgO (g)</th>
<th>0.3000</th>
</tr>
</thead>
<tbody>
<tr>
<td>MgO (mol)</td>
<td>0.0123</td>
</tr>
<tr>
<td>MgO (g/mol)</td>
<td>0.4652</td>
</tr>
</tbody>
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<th>MgO (g/mol)</th>
<th>0.4652</th>
</tr>
</thead>
<tbody>
<tr>
<td>MgO (mol)</td>
<td>0.0103</td>
</tr>
<tr>
<td>MgO (g/mol)</td>
<td>93.5</td>
</tr>
</tbody>
</table>

Share your results

Report your empirical formula on the board
Record class formulas

Discuss sources of error (include in discussion)

- Does your formula match expected one? MgO
- Does class data generally show expected results?
- Effects of:
  - Incomplete conversion to MgO → Mg, Mg₃N₂, Mg(OH)₂, …
  - Product losses – likely source of error?
  - Residual water – likely source of error?
- Brainstorm: Keep logical; discard unlikely
Data/Results

Result summary / sample calculations:

1. mass of Mg metal used
2. theoretical yield of MgO: \( \text{Mg}_{(s)} + \frac{1}{2}\text{O}_{2(g)} \rightarrow \text{MgO}_{(s)} \)
3. mass of \( \text{Mg}_x\text{O}_y \) produced
4. mass of O incorporated into product
5. moles of Mg and O in \( \text{Mg}_x\text{O}_y \)
6. empirical formula of the oxide (individual; class)
7. percent by mass of Mg and O in \( \text{Mg}_x\text{O}_y \)
8. percent yield of the reaction: \( \text{Mg} + \frac{1}{2}\text{O}_2 \rightarrow \text{MgO} \)

Discussion

Compare experimental empirical formula to theoretical empirical formula

Primary sources of error/deviation

Effects of factors such as:
- incomplete conversion of \( \text{Mg}_3\text{N}_2 \) to MgO, or
- residual \( \text{Mg(OH)}_2 \) in the product?

Does this method a valid way to determine empirical formula of metal oxides?
Report

Abstract
Results
Sample calculations
Discussion
Review questions

Example #1
2.448 g of an iron oxide was analyzed and found to contain 1.771 g Fe and 0.6766 g O. Calculate the empirical formula.

1. Calculate no. moles each

\[
\begin{align*}
1.771 \text{g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{g Fe}} &= 0.03171 \text{ mol Fe} \\
0.6776 \text{g O} \times \frac{1 \text{ mol O}}{16.00 \text{g O}} &= 0.04229 \text{ mol O}
\end{align*}
\]

2. Insert moles as subscripts

Fe_{0.03171}O_{0.04229}

3. Divide by smallest no. moles

\[
\frac{0.03171}{0.03171} = \frac{0.04229}{0.03171}
\]

4. Multiply by 2,3…until whole #s

Fe_2 O_1 = Fe_3 O_4

Formula: Fe_3 O_4
### Example #2

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

1. Calculate no. moles each

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass (g)</th>
<th>Moles Calculation</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>2.82</td>
<td>( \frac{2.82g}{23.1g/mol} \times 1 ) mol Na</td>
<td>0.1221 mol Na</td>
</tr>
<tr>
<td>Cl</td>
<td>4.35</td>
<td>( \frac{4.35g}{35.45g/mol} \times 1 ) mol Cl</td>
<td>0.1227 mol Cl</td>
</tr>
<tr>
<td>O</td>
<td>7.83</td>
<td>( \frac{7.83g}{16.00g/mol} \times 1 ) mol O</td>
<td>0.4894 mol O</td>
</tr>
</tbody>
</table>

2. Insert moles as subscripts

3. Divide by smallest no. moles

<table>
<thead>
<tr>
<th>Element</th>
<th>Na</th>
<th>Cl</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Moles</td>
<td>0.1221</td>
<td>0.1227</td>
<td>0.4894</td>
</tr>
</tbody>
</table>

4. Multiply by until whole #s

**Formula** \( \text{NaClO}_4 \)

### Example #3

Lactic acid (90.08g/mol) forms in muscle tissue and causes muscle soreness. It is 40.0 mass% C, 6.71 mass% H, and 53.3 mass% O. Find the **empirical** formula.

1. Calculate no. moles each

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass (g/mol)</th>
<th>Moles Calculation</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>40.0</td>
<td>( \frac{40.0g}{12.01g/mol} \times 1 ) mol C</td>
<td>3.331 mol C</td>
</tr>
<tr>
<td>H</td>
<td>6.71</td>
<td>( \frac{6.71g}{1.008g/mol} \times 1 ) mol H</td>
<td>6.657 mol H</td>
</tr>
<tr>
<td>O</td>
<td>53.3</td>
<td>( \frac{53.3g}{16.00g/mol} \times 1 ) mol O</td>
<td>3.331 mol O</td>
</tr>
</tbody>
</table>

2. Insert moles as subscripts

3. Divide by smallest no. moles

4. Multiply by until whole #s

**Empirical** \( \text{CH}_2\text{O} \)
Empirical Formula

Summary: Mass, Moles, Numbers