Many ionic compounds, when crystallized from an aqueous solution, will take up definite amounts of water as an integral part of their crystal structures. You can drive off this water of crystallization by heating the hydrated substance to convert it to its anhydrous form. Because the law of definite composition holds for crystalline hydrates, the number of moles of water driven off per mole of the anhydrous compound is a simple whole number. If the formula of the anhydrous compound is known, you can use your data to determine the formula of the hydrate.

**OBJECTIVES**

Determine that all the water has been driven from a hydrate by heating a sample to constant mass.

Use experimental data to calculate the number of moles of water released by a hydrate.

Infer the empirical formula of the hydrate from the formula of the anhydrous compound and experimental data.

**MATERIALS**

- balance, centigram
- Bunsen burner and related equipment
- crucible and cover
- desiccator
- iron ring
- magnesium sulfate, Epsom salts, hydrated crystals, MgSO₄·nH₂O
- pipe-stem triangle
- ring stand
- sparker
- spatula
- crucible tongs

⚠️ Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the locations of the emergency lab shower and the eyewash station and the procedures for using them.

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Procedure

1. Throughout the experiment, handle the crucible and cover with clean crucible tongs only. Place the crucible and cover on the triangle as shown in Figure A. Position the cover slightly tipped, leaving only a small opening for any gases to escape. Preheat the crucible and its cover until the bottom of the crucible turns red.

CAUTION The crucible and cover are very hot after each heating. Remember to handle them only with tongs.

2. Using tongs, transfer the crucible and cover to a desiccator. Allow them to cool 5 min in the desiccator. Never place a hot crucible on a balance. When the crucible and cover are cool, determine their mass to the nearest 0.01 g. Record this mass in the Data Table.

3. Using a spatula, add approximately 5 g of magnesium sulfate hydrate crystals to the crucible. Determine the mass of the covered crucible and crystals to the nearest 0.01 g. Record this mass in the Data Table.

4. Place the crucible with the magnesium sulfate hydrate on the triangle, and again position the cover so that there is a small opening. Too large an opening may allow the hydrate to spatter out of the crucible. Heat the crucible very gently with a low flame to avoid spattering any of the hydrate. Increase the temperature gradually for 2 or 3 min. Then, heat strongly, but not red-hot, for at least 5 min.

5. Using tongs, transfer the crucible, cover, and contents to the desiccator, and allow them to cool for 5 min. Then, using the same balance you used in Step 2, determine their mass. Be sure the crucible is sufficiently cool, because heat can affect your measurement. Record the mass in the Data Table.
Water of Hydration

6. Again heat the covered crucible and contents strongly for 5 min. Allow the crucible, cover, and contents to cool in the desiccator, and then use the same balance as before to determine their mass. If the last two mass measurements differ by no more than 0.01 g, you may assume that all the water has been driven off. Otherwise, repeat the heating process until the mass no longer changes. Record this constant mass in your Data Table.

DISPOSAL

7. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of the MgSO₄ in your crucible as your teacher directs. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

Data Table

<table>
<thead>
<tr>
<th>Description</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of empty crucible and cover</td>
<td>22.755</td>
</tr>
<tr>
<td>Mass of crucible, cover, and magnesium sulfate hydrate</td>
<td>27.495</td>
</tr>
<tr>
<td>Mass of crucible, cover, and anhydrous magnesium sulfate after 1st heating</td>
<td>25.348</td>
</tr>
<tr>
<td>Mass of crucible, cover, and anhydrous magnesium sulfate after last heating</td>
<td>25.064</td>
</tr>
</tbody>
</table>

Analysis

1. Organizing Data Calculate the mass of anhydrous magnesium sulfate (the residue that remained after driving off the water). Record the mass in the Calculations Table.

\[25.064 - 22.755 = 2.3099\]

2. Organizing Data Calculate the number of moles of anhydrous magnesium sulfate. Record the number of moles in the Calculations Table.

\[2.3099 \text{ g} \times \frac{1 \text{ mol}}{57.375 \text{ g}} = 0.041 \text{ mol}\]

3. Organizing Data Calculate the mass of water driven off from the hydrate. Record the mass in the Calculations Table.

\[27.495 - 25.064 = 2.431\]

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Holt ChemFile A 30 Skills Practice Experiment
4. Organizing Data Calculate the number of moles of water driven off from the hydrate. Record the number of moles in the Calculations Table.

\[ \frac{2.431 \text{ g}}{16.07 \text{ g/mol}} = 0.135 \text{ mol} \]

5. Organizing Ideas Write the equation for the reaction that occurred when you heated hydrated MgSO₄ in this experiment. Use the letter \( n \) to represent the number of moles of water driven off per mole of anhydrous magnesium sulfate.

\[ \text{MgSO}_4 \cdot n\text{H}_2\text{O}(s) \]

6. Organizing Conclusions Using your answers to Calculations Items 2, 4, and 5, determine the mole ratio of MgSO₄ to H₂O to the nearest whole number. Record the ratio in the Calculations Table.

7. Organizing Conclusions Use your answer to Calculations Item 6 to write the formula for the magnesium sulfate hydrate. Record the formula in the Calculations Table.

<table>
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<tr>
<th>Calculations Table</th>
</tr>
</thead>
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<tr>
<td>Mass of anhydrous magnesium sulfate</td>
</tr>
<tr>
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</tr>
<tr>
<td>Mass of water driven off from hydrate</td>
</tr>
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</tr>
<tr>
<td>Mole ratio of anhydrous magnesium sulfate to water</td>
</tr>
<tr>
<td>Empirical formula of the hydrate</td>
</tr>
</tbody>
</table>
Conclusions

1. **Applying Conclusions** The following results were obtained when a solid was heated by three different lab groups. In each case, the students observed that when they began to heat the solid, drops of a liquid formed on the sides of the test tube.

<table>
<thead>
<tr>
<th>Lab group</th>
<th>Mass before heating</th>
<th>Mass after heating</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1.48 g</td>
<td>1.26 g</td>
</tr>
<tr>
<td>2</td>
<td>1.64 g</td>
<td>1.40 g</td>
</tr>
<tr>
<td>3</td>
<td>2.08 g</td>
<td>1.78 g</td>
</tr>
</tbody>
</table>

a. Could the solid be a hydrate? What evidence supports your answer?

Yes, it could. The mass decreases after heating. How else do we know?

b. If, after heating, the solid has a molar mass of 208 g/mol and a formula of **XY**, what is the formula of the hydrate? Use **Lab group 1**’s results.

\[
\begin{align*}
1.48 - 1.26 &= 0.22 \\
1.26 \times \frac{1 \text{ mol}}{200} &= 0.006 \text{ mol} \times x = 1 \\
0.22 \times \frac{1}{0.006} &= 0.12 \text{ mol} H_2O
\end{align*}
\]

**XY · H_2O**

2. **Applying Conclusions** Some cracker tins include a glass vial of drying material in the lid to keep the crackers crisp. In many cases, the material is a mixture of magnesium sulfate and cobalt chloride indicator. As the magnesium sulfate absorbs moisture (MgSO₄ · H₂O + 6H₂O → MgSO₄ · 7H₂O), the indicator changes color from blue to pink (CoCl₂ · 4H₂O + 2H₂O → CoCl₂ · 6H₂O). When this drying mixture becomes totally pink, it can be restored if it is heated in an oven. What two changes are caused by the heating?

The heating rids of the H₂O and ?

---

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Holt ChemFile A 32 Skills Practice Experiment
3. **Applying Conclusions** How does the experiment you performed exemplify the law of definite composition?

The Law of definite composition is the statement that in a pure compound, the elements are always combined in fixed proportions by weight. The experiment exemplifies this by the elements being combined in fixed proportions by weight.

4. **Analyzing Methods** Why did you use the same balance each time you measured the mass of the crucible and its contents?

To reduce the amount of possible error.

5. **Predicting Outcomes** Use handbooks to investigate the properties of the compounds listed in the following table. Write answers to these questions in the table: Does the compound form a hydrate? How would you describe the appearance of the anhydrous compound? If your teacher approves, repeat the experimental procedure, using one of the listed compounds, and verify its hydrate formula. Explain any large deviation from the correct hydrate formula.

<table>
<thead>
<tr>
<th>Name of Compound</th>
<th>Water of crystallization?</th>
<th>Appearance of anhydrous compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium carbonate</td>
<td>Yes</td>
<td>White Powder</td>
</tr>
<tr>
<td>Sodium sulfate</td>
<td>Yes</td>
<td>White powder</td>
</tr>
<tr>
<td>Sodium aluminum sulfate</td>
<td>Yes</td>
<td>White Crystal</td>
</tr>
<tr>
<td>Potassium chloride</td>
<td>No</td>
<td></td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>Yes</td>
<td>White Crystal</td>
</tr>
<tr>
<td>Copper sulfate</td>
<td>Yes</td>
<td>White Powder</td>
</tr>
</tbody>
</table>


Many ionic compounds, when crystallized from an aqueous solution, will take up definite amounts of water as an integral part of their crystal structures. You can drive off this water of crystallization by heating the hydrated substance to convert it to its anhydrous form. Because the law of definite composition holds for crystalline hydrates, the number of moles of water driven off per mole of the anhydrous compound is a simple whole number. If the formula of the anhydrous compound is known, you can use your data to determine the formula of the hydrate.

**OBJECTIVES**

**Determine** that all the water has been driven from a hydrate by heating a sample to constant mass.

**Use** experimental data to calculate the number of moles of water released by a hydrate.

**Infer** the empirical formula of the hydrate from the formula of the anhydrous compound and experimental data.

**MATERIALS**

- balance, centigram
- Bunsen burner and related equipment
- crucible and cover
- desiccator
- iron ring
- magnesium sulfate, Epsom salts, hydrated crystals, MgSO₄ \( \cdot n\text{H}_2\text{O} \)
- pipe-stem triangle
- ring stand
- sparker
- spatula
- crucible tongs

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Procedure

1. Throughout the experiment, handle the crucible and cover with clean crucible tongs only. Place the crucible and cover on the triangle as shown in Figure A. Position the cover slightly tipped, leaving only a small opening for any gases to escape. Preheat the crucible and its cover until the bottom of the crucible turns red.

   CAUTION The crucible and cover are very hot after each heating. Remember to handle them only with tongs.

2. Using tongs, transfer the crucible and cover to a desiccator. Allow them to cool 5 min in the desiccator. Never place a hot crucible on a balance. When the crucible and cover are cool, determine their mass to the nearest 0.01 g. Record this mass in the Data Table.

3. Using a spatula, add approximately 5 g of magnesium sulfate hydrate crystals to the crucible. Determine the mass of the covered crucible and crystals to the nearest 0.01 g. Record this mass in the Data Table.

4. Place the crucible with the magnesium sulfate hydrate on the triangle, and again position the cover so that there is a small opening. Too large an opening may allow the hydrate to spatter out of the crucible. Heat the crucible very gently with a low flame to avoid spattering any of the hydrate. Increase the temperature gradually for 2 or 3 min. Then, heat strongly, but not red-hot, for at least 5 min.

5. Using tongs, transfer the crucible, cover, and contents to the desiccator, and allow them to cool for 5 min. Then, using the same balance you used in Step 2, determine their mass. Be sure the crucible is sufficiently cool, because heat can affect your measurement. Record the mass in the Data Table.
6. Again heat the covered crucible and contents strongly for 5 min. Allow the crucible, cover, and contents to cool in the desiccator, and then use the same balance as before to determine their mass. If the last two mass measurements differ by no more than 0.01 g, you may assume that all the water has been driven off. Otherwise, repeat the heating process until the mass no longer changes. Record this constant mass in your Data Table.

DISPOSAL

7. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of the MgSO₄ in your crucible as your teacher directs. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

<table>
<thead>
<tr>
<th>Data Table</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of empty crucible and cover</td>
<td>13.647 g</td>
</tr>
<tr>
<td>Mass of crucible, cover, and magnesium sulfate hydrate</td>
<td>18.518 g</td>
</tr>
<tr>
<td>Mass of crucible, cover, and anhydrous magnesium sulfate after 1st heating</td>
<td>16.343 g</td>
</tr>
<tr>
<td>Mass of crucible, cover, and anhydrous magnesium sulfate after last heating</td>
<td>16.239 g</td>
</tr>
</tbody>
</table>

Analysis

1. **Organizing Data** Calculate the mass of anhydrous magnesium sulfate (the residue that remained after driving off the water). Record the mass in the Calculations Table.

\[
\begin{align*}
\text{Mass of crucible, cover, and anhydrous magnesium sulfate} & \quad 16.343 g \\
\text{Mass of empty crucible and cover} & \quad 13.647 g \\
\hline
\text{Difference} & \quad 2.592 g
\end{align*}
\]

2. **Organizing Data** Calculate the number of moles of anhydrous magnesium sulfate. Record the number of moles in the Calculations Table.

\[
\begin{align*}
\text{Molar mass of MgSO}_4 & \quad 84.32 g/mol \\
\text{Moles of MgSO}_4 & = \frac{2.592 g \times 1 \text{ mol}}{120.349 g}\quad \approx 0.021 \text{ mol}
\end{align*}
\]

3. **Organizing Data** Calculate the mass of water driven off from the hydrate. Record the mass in the Calculations Table.

\[
\begin{align*}
\text{Mass of crucible, cover, and anhydrous magnesium sulfate} & \quad 16.349 g \\
\text{Mass of empty crucible and cover} & \quad 10.239 g \\
\hline
\text{Difference} & \quad 6.110 g
\end{align*}
\]

\[
2.279 g \times \text{H}_2\text{O}
\]
**Water of Hydration continued**

4. **Organizing Data** Calculate the number of moles of water driven off from the hydrate. Record the number of moles in the Calculations Table.

\[
\begin{align*}
\text{K: } & 2.279 \text{ g } H_2O \\ \text{W: } & 1 \text{ mol } H_2O \\ \text{moles of } H_2O & = \frac{2.279 \text{ g } H_2O}{18.02 \text{ g } H_2O} \\
& = 0.126 \text{ mol } H_2O
\end{align*}
\]

5. **Organizing Ideas** Write the equation for the reaction that occurred when you heated hydrated MgSO₄ in this experiment. Use the letter \( n \) to represent the number of moles of water driven off per mole of anhydrous magnesium sulfate.

\[
\text{MgSO}_4 \cdot nH_2O \overset{\text{heat}}{\longrightarrow} \text{MgSO}_4 + nH_2O(g)
\]

6. **Organizing Conclusions** Using your answers to Calculations Items 2, 4, and 5, determine the mole ratio of MgSO₄ to H₂O to the nearest whole number. Record the ratio in the Calculations Table.

\[
\frac{0.21}{0.21} = \frac{1.26}{0.21} \rightarrow \text{MgSO}_4 : H_2O \rightarrow 1 : 6
\]

7. **Organizing Conclusions** Use your answer to Calculations Item 6 to write the formula for the magnesium sulfate hydrate. Record the formula in the Calculations Table.

\[
\text{MgSO}_4 \cdot 6H_2O
\]

<table>
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</tr>
<tr>
<td>Moles of water driven off from hydrate</td>
<td>0.126 g</td>
</tr>
<tr>
<td>Mole ratio of anhydrous magnesium sulfate to water</td>
<td>1 : 6</td>
</tr>
<tr>
<td>Empirical formula of the hydrate</td>
<td>MgSO₄ · 6H₂O</td>
</tr>
</tbody>
</table>

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Holt ChemFile A 31 Skills Practice Experiment
Water of Hydration continued

Conclusions

1. Applying Conclusions The following results were obtained when a solid was heated by three different lab groups. In each case, the students observed that when they began to heat the solid, drops of a liquid formed on the sides of the test tube.

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<tr>
<th>Lab group</th>
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a. Could the solid be a hydrate? What evidence supports your answer?

b. If, after heating, the solid has a molar mass of 208 g/mol and a formula of XY, what is the formula of the hydrate? Use Lab group 1’s results.

2. Applying Conclusions Some cracker tins include a glass vial of drying material in the lid to keep the crackers crisp. In many cases, the material is a mixture of magnesium sulfate and cobalt chloride indicator. As the magnesium sulfate absorbs moisture (MgSO₄ · H₂O + 6H₂O → MgSO₄ · 7H₂O), the indicator changes color from blue to pink (CoCl₂ · 4H₂O + 2H₂O → CoCl₂ · 6H₂O). When this drying mixture becomes totally pink, it can be restored if it is heated in an oven. What two changes are caused by the heating?
3. Applying Conclusions  How does the experiment you performed exemplify the law of definite composition?

This experiment shows the law of definite composition because it has to do with percent compositions, and the magnesium sulfate was separated from the water, which was definitely a part of the magnesium before hand.

4. Analyzing Methods  Why did you use the same balance each time you measured the mass of the crucible and its contents?

Because changing balances increases the chance of error. One balance has a consistent error, so its better to stay the same.

5. Predicting Outcomes  Use handbooks to investigate the properties of the compounds listed in the following table. Write answers to these questions in the table: Does the compound form a hydrate? How would you describe the appearance of the anhydrous compound? If your teacher approves, repeat the experimental procedure, using one of the listed compounds, and verify its hydrate formula. Explain any large deviation from the correct hydrate formula.

<table>
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</tr>
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<td></td>
<td></td>
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<td>Copper sulfate</td>
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definite amounts of water as an integral part of their crystal structures. You can
drive off this water of crystallization by heating the hydrated substance to con-
vert it to its anhydrous form. Because the law of definite composition holds for
crystalline hydrates, the number of moles of water driven off per mole of the
anhydrous compound is a simple whole number. If the formula of the anhydrous
compound is known, you can use your data to determine the formula of the
hydrate.

**OBJECTIVES**

**Determine** that all the water has been driven from a hydrate by heating a sample
to constant mass.

**Use** experimental data to calculate the number of moles of water released by a
hydrate.

**Infer** the empirical formula of the hydrate from the formula of the anhydrous
compound and experimental data.

**MATERIALS**

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- magnesium sulfate, Epsom salts,
  hydrated crystals, MgSO₄·nH₂O
- pipe-stem triangle
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- spatula
- crucible tongs

⚠ ⚠ **Always wear safety goggles and a lab apron to protect your eyes**

and clothing. If you get a chemical in your eyes, immediately flush
the chemical out at the eyewash station while calling to your teacher. Know the
locations of the emergency lab shower and the eyewash station and the procedures
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Water of Hydration continued

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Procedure

1. Throughout the experiment, handle the crucible and cover with clean crucible tongs only. Place the crucible and cover on the triangle as shown in Figure A. Position the cover slightly tipped, leaving only a small opening for any gases to escape. Preheat the crucible and its cover until the bottom of the crucible turns red.

   CAUTION The crucible and cover are very hot after each heating. Remember to handle them only with tongs.

2. Using tongs, transfer the crucible and cover to a desiccator. Allow them to cool 5 min in the desiccator. Never place a hot crucible on a balance. When the crucible and cover are cool, determine their mass to the nearest 0.01 g. Record this mass in the Data Table.

3. Using a spatula, add approximately 5 g of magnesium sulfate hydrate crystals to the crucible. Determine the mass of the covered crucible and crystals to the nearest 0.01 g. Record this mass in the Data Table.

4. Place the crucible with the magnesium sulfate hydrate on the triangle, and again position the cover so that there is a small opening. Too large an opening may allow the hydrate to spatter out of the crucible. Heat the crucible very gently with a low flame to avoid spattering any of the hydrate. Increase the temperature gradually for 2 or 3 min. Then, heat strongly, but not red-hot, for at least 5 min.

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6. Again heat the covered crucible and contents strongly for 5 min. Allow the crucible, cover, and contents to cool in the desiccator, and then use the same balance as before to determine their mass. If the last two mass measurements differ by no more than 0.01 g, you may assume that all the water has been driven off. Otherwise, repeat the heating process until the mass no longer changes. Record this constant mass in your Data Table.

DISPOSAL

7. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of the MgSO₄ in your crucible as your teacher directs. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

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<tr>
<td>Mass of crucible, cover, and anhydrous magnesium sulfate after last heating</td>
</tr>
</tbody>
</table>

Analysis

1. **Organizing Data** Calculate the mass of anhydrous magnesium sulfate (the residue that remained after driving off the water). Record the mass in the Calculations Table.

   \[18.925 - 16.547 = 2.378\]

2. **Organizing Data** Calculate the number of moles of anhydrous magnesium sulfate. Record the number of moles in the Calculations Table.

   \[
   \begin{align*}
   \text{Mg}_2\text{SO}_4 &\rightarrow 2\text{MgS}O_4 + 32.00 + 4(15.999) = 120.361 \\
   0.01 &\text{ mol} & \times \frac{120.361}{32.00} & = 0.370 \text{ mol} \\
   \end{align*}
   \]

3. **Organizing Data** Calculate the mass of water driven off from the hydrate. Record the mass in the Calculations Table.

   \[18.925 - 16.547 = 2.378\]
Water of Hydration continued

4. Organizing Data Calculate the number of moles of water driven off from the hydrate. Record the number of moles in the Calculations Table.

\[
\begin{align*}
2.378 \text{ g} & \div 18.015 \text{ g/mol} = 0.132 \\
& \div 1.808 \text{ g/mol} = 18.025 \text{ g per mol}
\end{align*}
\]

5. Organizing Ideas Write the equation for the reaction that occurred when you heated hydrated MgSO₄ in this experiment. Use the letter \( n \) to represent the number of moles of water driven off per mole of anhydrous magnesium sulfate.

\[
\text{MgSO}_4 \cdot n \text{H}_2\text{O} \rightarrow \text{MgSO}_4 + n \text{H}_2\text{O}(g)
\]

6. Organizing Conclusions Using your answers to Calculations Items 2, 4, and 5, determine the mole ratio of MgSO₄ to H₂O to the nearest whole number. Record the ratio in the Calculations Table.

\[
\frac{0.02}{0.02} = \frac{13.2}{0.02} = 1:7
\]

7. Organizing Conclusions Use your answer to Calculations Item 6 to write the formula for the magnesium sulfate hydrate. Record the formula in the Calculations Table.

\[
\text{MgSO}_4 \cdot 7\text{H}_2\text{O} \rightarrow \text{MgSO}_4 + 7\text{H}_2\text{O}
\]

<table>
<thead>
<tr>
<th>Calculations Table</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of anhydrous magnesium sulfate</td>
</tr>
<tr>
<td>2.372 g</td>
</tr>
<tr>
<td>Moles of anhydrous magnesium sulfate</td>
</tr>
<tr>
<td>0.020 mol</td>
</tr>
<tr>
<td>Mass of water driven off from hydrate</td>
</tr>
<tr>
<td>2.378 g</td>
</tr>
<tr>
<td>Moles of water driven off from hydrate</td>
</tr>
<tr>
<td>0.132</td>
</tr>
<tr>
<td>Mole ratio of anhydrous magnesium sulfate to water</td>
</tr>
<tr>
<td>1:7</td>
</tr>
<tr>
<td>Empirical formula of the hydrate</td>
</tr>
</tbody>
</table>

Percent Error

\[
\frac{|7 - 6.00\%|}{7} \times 100 = 5.7\%
\]
Conclusions

1. Applying Conclusions  The following results were obtained when a solid was heated by three different lab groups. In each case, the students observed that when they began to heat the solid, drops of a liquid formed on the sides of the test tube.

<table>
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<tr>
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<th>Mass before heating</th>
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<tr>
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</tr>
<tr>
<td>3</td>
<td>2.08 g</td>
<td>1.78 g</td>
</tr>
</tbody>
</table>

a. Could the solid be a hydrate? What evidence supports your answer?

Yes, when it was heated, the mass decreased which suggests that a substance is being evaporated.

b. If, after heating, the solid has a molar mass of 208 g/mol and a formula of XY, what is the formula of the hydrate? Use Lab group 1’s results.

\[ \text{XY} \times \text{H}_2\text{O} \]

\[ \frac{1.48 \text{ g}}{1 \text{ mol}} = \frac{0.07 \text{ mol}}{0.012 \text{ mol H}_2\text{O}} \]
\[ \frac{2.08 \text{ g}}{1 \text{ mol}} = \frac{0.09 \text{ mol}}{0.012 \text{ mol H}_2\text{O}} \]

2. Applying Conclusions  Some cracker tins include a glass vial of drying material in the lid to keep the crackers crisp. In many cases, the material is a mixture of magnesium sulfate and cobalt chloride indicator. As the magnesium sulfate absorbs moisture (\( \text{MgSO}_4 \cdot \text{H}_2\text{O} + 6\text{H}_2\text{O} \rightarrow \text{MgSO}_4 \cdot 7\text{H}_2\text{O} \)), the indicator changes color from blue to pink (\( \text{CoCl}_2 \cdot 4\text{H}_2\text{O} + 2\text{H}_2\text{O} \rightarrow \text{CoCl}_2 \cdot 6\text{H}_2\text{O} \)). When this drying mixture becomes totally pink, it can be restored if it is heated in an oven. What two changes are caused by the heating?

Chemical and physical properties

Chemical - the amount of water per element

Physical - color change.
3. **Applying Conclusions** How does the experiment you performed exemplify the law of definite composition?

   *This states that there can't be part of an atom and it shows the ratio of water molecules per molecule of the Na₂SO₄ substance.*

4. **Analyzing Methods** Why did you use the same balance each time you measured the mass of the crucible and its contents?

   *So that there would be the same amount of error in each measurement.*

5. **Predicting Outcomes** Use handbooks to investigate the properties of the compounds listed in the following table. Write answers to these questions in the table: Does the compound form a hydrate? How would you describe the appearance of the anhydrous compound? If your teacher approves, repeat the experimental procedure, using one of the listed compounds, and verify its hydrate formula. Explain any large deviation from the correct hydrate formula.

<table>
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<td>10</td>
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<td>12</td>
<td>white</td>
</tr>
<tr>
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<td>5</td>
<td>white</td>
</tr>
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</tr>
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<td>5</td>
<td>grey-white</td>
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Many ionic compounds, when crystallized from an aqueous solution, will take up definite amounts of water as an integral part of their crystal structures. You can drive off this water of crystallization by heating the hydrated substance to convert it to its anhydrous form. Because the law of definite composition holds for crystalline hydrates, the number of moles of water driven off per mole of the anhydrous compound is a simple whole number. If the formula of the anhydrous compound is known, you can use your data to determine the formula of the hydrate.

**OBJECTIVES**

**Determine** that all the water has been driven from a hydrate by heating a sample to constant mass.

**Use** experimental data to calculate the number of moles of water released by a hydrate.

**Infer** the empirical formula of the hydrate from the formula of the anhydrous compound and experimental data.

**MATERIALS**

- balance, centigram
- Bunsen burner and related equipment
- crucible and cover
- desiccator
- iron ring
- magnesium sulfate, Epsom salts, hydrated crystals, MgSO₄ ⋅ nH₂O
- pipe-stem triangle
- ring stand
- sparker
- spatula
- crucible tongs

Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the locations of the emergency lab shower and the eyewash station and the procedures for using them.

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Procedure

1. Throughout the experiment, handle the crucible and cover with clean Crucible only. Place the crucible and cover on the triangle as shown in Figure A. Position the cover slightly tipped, leaving only a small opening for any gases to escape. Preheat the crucible and its cover until the bottom of the crucible turns red.

   CAUTION The crucible and cover are very hot after each heating. Remember to handle them only with tongs.

2. Using tongs, transfer the crucible and cover to a desiccator. Allow them to cool 5 min in the desiccator. Never place a hot crucible on a balance. When the crucible and cover are cool, determine their mass to the nearest 0.01 g. Record this mass in the Data Table.

3. Using a spatula, add approximately 5 g of magnesium sulfate hydrate crystals to the crucible. Determine the mass of the covered crucible and crystals to the nearest 0.01 g. Record this mass in the Data Table.

4. Place the crucible with the magnesium sulfate hydrate on the triangle, and again position the cover so that there is a small opening. Too large an opening may allow the hydrate to spatter out of the crucible. Heat the crucible very gently with a low flame to avoid spattering any of the hydrate. Increase the temperature gradually for 2 or 3 min. Then, heat strongly, but not red-hot, for at least 5 min.

5. Using tongs, transfer the crucible, cover, and contents to the desiccator, and allow them to cool for 5 min. Then, using the same balance you used in Step 2, determine their mass. Be sure the crucible is sufficiently cool, because heat can affect your measurement. Record the mass in the Data Table.
Water of Hydration continued

6. Again heat the covered crucible and contents strongly for 5 min. Allow the crucible, cover, and contents to cool in the desiccator, and then use the same balance as before to determine their mass. If the last two mass measurements differ by no more than 0.01 g, you may assume that all the water has been driven off. Otherwise, repeat the heating process until the mass no longer changes. Record this constant mass in your Data Table.

DISPOSAL

7. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of the MgSO₄ in your crucible as your teacher directs. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

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Analysis

1. Organizing Data Calculate the mass of anhydrous magnesium sulfate (the residue that remained after driving off the water). Record the mass in the Calculations Table.

\[ 16.751 \text{ g} - 14.297 \text{ g} = 2.454 \text{ g} \]

2. Organizing Data Calculate the number of moles of anhydrous magnesium sulfate. Record the number of moles in the Calculations Table.

\[ \text{MgSO}_4 = 120.361 \text{ g} \]
\[ H_2O = 18.012 \text{ g} \]
\[ \frac{2.454 \text{ g}}{120.361 \text{ g}} = 0.02039 \text{ mol of MgSO}_4 \]

3. Organizing Data Calculate the mass of water driven off from the hydrate. Record the mass in the Calculations Table.

\[ 19.182 \text{ g} - 16.751 \text{ g} = 2.431 \text{ g} \]
4. **Organizing Data** Calculate the number of moles of water driven off from the hydrate. Record the number of moles in the Calculations Table.

\[
\frac{2401 \text{ g}}{1 \text{ mol}} = 0.1332 \text{ mol of H}_2\text{O}
\]

5. **Organizing Ideas** Write the equation for the reaction that occurred when you heated hydrated MgSO\(_4\) in this experiment. Use the letter \(n\) to represent the number of moles of water driven off per mole of anhydrous magnesium sulfate.

\[
\text{MgSO}_4 \cdot \text{H}_2\text{O} \xrightarrow{\text{heated}} \text{MgSO}_4 + n\text{H}_2\text{O}
\]

6. **Organizing Conclusions** Using your answers to Calculations Items 2, 4, and 5, determine the mole ratio of MgSO\(_4\) to H\(_2\)O to the nearest whole number. Record the ratio in the Calculations Table.

\[
\frac{0.02 \text{ mol}}{0.13 \text{ mol}} = 1.65
\]

7. **Organizing Conclusions** Use your answer to Calculations Item 6 to write the formula for the magnesium sulfate hydrate. Record the formula in the Calculations Table.

\[
\text{MgSO}_4 \cdot 7\text{H}_2\text{O}
\]

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<tr>
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</tr>
<tr>
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</tr>
<tr>
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Water of Hydration continued

Conclusions

1. Applying Conclusions The following results were obtained when a solid was heated by three different lab groups. In each case, the students observed that when they began to heat the solid, drops of a liquid formed on the sides of the test tube.

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a. Could the solid be a hydrate? What evidence supports your answer?

b. If, after heating, the solid has a molar mass of 208 g/mol and a formula of $XY$, what is the formula of the hydrate? Use Lab group 1’s results.

2. Applying Conclusions Some cracker tins include a glass vial of drying material in the lid to keep the crackers crisp. In many cases, the material is a mixture of magnesium sulfate and cobalt chloride indicator. As the magnesium sulfate absorbs moisture ($\text{MgSO}_4 \cdot \text{H}_2\text{O} + 6\text{H}_2\text{O} \rightarrow \text{MgSO}_4 \cdot 7\text{H}_2\text{O}$), the indicator changes color from blue to pink ($\text{CoCl}_2 \cdot 4\text{H}_2\text{O} + 2\text{H}_2\text{O} \rightarrow \text{CoCl}_2 \cdot 6\text{H}_2\text{O}$). When this drying mixture becomes totally pink, it can be restored if it is heated in an oven. What two changes are caused by the heating?
3. Applying Conclusions How does the experiment you performed exemplify the law of definite composition?

Magnesium sulfate has every seven water because we have selected the difference between water and magnesium sulfate that then a ratio to find that every MgSO4 will have 7H2O.

4. Analyzing Methods Why did you use the same balance each time you measured the mass of the crucible and its contents?

To show its accuracy of its contents.

5. Predicting Outcomes Use handbooks to investigate the properties of the compounds listed in the following table. Write answers to these questions in the table: Does the compound form a hydrate? How would you describe the appearance of the anhydrous compound? If your teacher approves, repeat the experimental procedure, using one of the listed compounds, and verify its hydrate formula. Explain any large deviation from the correct hydrate formula.

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<td>Sodium carbonate</td>
<td>yes</td>
<td>2Na2O + H2O = Na2CO3 + 3H2O</td>
</tr>
<tr>
<td>Sodium sulfate</td>
<td>yes</td>
<td>Na2O + H2SO4 = Na2SO4 + H2O</td>
</tr>
<tr>
<td>Sodium aluminum sulfate</td>
<td>no</td>
<td></td>
</tr>
<tr>
<td>Potassium chloride</td>
<td>yes</td>
<td>K2O + HCl = KCl + H2O</td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>yes</td>
<td>Mg(OH)2 + 2HCl = MgCl2 + 2H2O</td>
</tr>
<tr>
<td>Copper sulfate</td>
<td>yes</td>
<td>Cu(OH)2 + H2SO4 = CuSO4 + 2H2O</td>
</tr>
</tbody>
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Many ionic compounds, when crystallized from an aqueous solution, will take up definite amounts of water as an integral part of their crystal structures. You can drive off this water of crystallization by heating the hydrated substance to convert it to its anhydrous form. Because the law of definite composition holds for crystalline hydrates, the number of moles of water driven off per mole of the anhydrous compound is a simple whole number. If the formula of the anhydrous compound is known, you can use your data to determine the formula of the hydrate.

**OBJECTIVES**

**Determine** that all the water has been driven from a hydrate by heating a sample to constant mass.

**Use** experimental data to calculate the number of moles of water released by a hydrate.

**Infer** the empirical formula of the hydrate from the formula of the anhydrous compound and experimental data.

**MATERIALS**

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Procedure

1. Throughout the experiment, handle the crucible and cover with clean crucible tongs only. Place the crucible and cover on the triangle as shown in Figure A. Position the cover slightly tipped, leaving only a small opening for any gases to escape. Preheat the crucible and its cover until the bottom of the crucible turns red.

   CAUTION The crucible and cover are very hot after each heating. Remember to handle them only with tongs.

2. Using tongs, transfer the crucible and cover to a desiccator. Allow them to cool 5 min in the desiccator. Never place a hot crucible on a balance. When the crucible and cover are cool, determine their mass to the nearest 0.01 g. Record this mass in the Data Table.

3. Using a spatula, add approximately 5 g of magnesium sulfate hydrate crystals to the crucible. Determine the mass of the covered crucible and crystals to the nearest 0.01 g. Record this mass in the Data Table.

4. Place the crucible with the magnesium sulfate hydrate on the triangle, and again position the cover so that there is a small opening. Too large an opening may allow the hydrate to spatter out of the crucible. Heat the crucible very gently with a low flame to avoid spattering any of the hydrate. Increase the temperature gradually for 2 or 3 min. Then, heat strongly, but not red-hot, for at least 5 min.

5. Using tongs, transfer the crucible, cover, and contents to the desiccator, and allow them to cool for 5 min. Then, using the same balance you used in Step 2, determine their mass. Be sure the crucible is sufficiently cool, because heat can affect your measurement. Record the mass in the Data Table.
Water of Hydration continued

6. Again heat the covered crucible and contents strongly for 5 min. Allow the crucible, cover, and contents to cool in the desiccator, and then use the same balance as before to determine their mass. If the last two mass measurements differ by no more than 0.01 g, you may assume that all the water has been driven off. Otherwise, repeat the heating process until the mass no longer changes. Record this constant mass in your Data Table.

DISPOSAL

7. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of the MgSO₄ in your crucible as your teacher directs. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

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<th>16.825</th>
</tr>
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<tbody>
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Analysis

1. **Organizing Data** Calculate the mass of anhydrous magnesium sulfate (the residue that remained after driving off the water). Record the mass in the Calculations Table.

2. **Organizing Data** Calculate the number of moles of anhydrous magnesium sulfate. Record the number of moles in the Calculations Table.

   \[
   \text{MgSO}_4, \quad 2.294 \text{ g} \times \frac{1 \text{ mol}}{120.37 \text{ g}} = \frac{2.294}{120.37} = 0.019057 \text{ mol}
   \]

3. **Organizing Data** Calculate the mass of water driven off from the hydrate. Record the mass in the Calculations Table.

   \[
   19.119 - 16.738 = 2.381
   \]
4. **Organizing Data** Calculate the number of moles of water driven off from the hydrate. Record the number of moles in the **Calculations Table**.

\[
\text{H}_2\text{O} \quad 2.381g \times \frac{1\text{ mol}}{18.02\text{ g}} = 0.1321\text{ moles}
\]

5. **Organizing Ideas** Write the equation for the reaction that occurred when you heated hydrated MgSO₄ in this experiment. Use the letter \( n \) to represent the number of moles of water driven off per mole of anhydrous magnesium sulfate.

\[
\text{MgSO}_4 \cdot \text{H}_2\text{O} \xrightarrow{\text{heat}} \text{MgSO}_4 + \text{H}_2\text{O} (g)
\]

6. **Organizing Conclusions** Using your answers to Calculations Items 2, 4, and 5, determine the mole ratio of MgSO₄ to H₂O to the nearest whole number. Record the ratio in the **Calculations Table**.

\[
\frac{0.19057}{0.1321} = 1.47
\]

7. **Organizing Conclusions** Use your answer to Calculations Item 6 to write the formula for the magnesium sulfate hydrate. Record the formula in the **Calculations Table**.

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Conclusions

1. Applying Conclusions The following results were obtained when a solid was heated by three different lab groups. In each case, the students observed that when they began to heat the solid, drops of a liquid formed on the sides of the test tube.

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a. Could the solid be a hydrate? What evidence supports your answer?

Yes, the solid could be a hydrate. After heating, the solid’s mass decreased significantly. This could be attributed to the evaporation of water, which could only be present in a hydrate.

b. If, after heating, the solid has a molar mass of 208 g/mol and a formula of XY, what is the formula of the hydrate? Use Lab group 1’s results.

2. Applying Conclusions Some cracker tins include a glass vial of drying material in the lid to keep the crackers crisp. In many cases, the material is a mixture of magnesium sulfate and cobalt chloride indicator. As the magnesium sulfate absorbs moisture (MgSO₄ · H₂O + 6H₂O → MgSO₄ · 7H₂O), the indicator changes color from blue to pink (CoCl₂ · 4H₂O + 2H₂O → CoCl₂ · 6H₂O). When this drying mixture becomes totally pink, it can be restored if it is heated in an oven. What two changes are caused by the heating?

The heat changes the magnesium sulfate into cobalt chloride in exchange for one water molecule from each side.

It's a mixture of both.
3. **Applying Conclusions**  How does the experiment you performed exemplify the law of definite composition?

Our lab proves the fact that compounds always exist in definite proportions which in turn make it a patent example of the law of definite composition. The H₂O₂ in our final equation matches hydrogen sulfide in our experiment.

---

4. **Analyzing Methods**  Why did you use the same balance each time you measured the mass of the crucible and its contents?

To keep the error in the scales at a minimum.

---

5. **Predicting Outcomes**  Use handbooks to investigate the properties of the compounds listed in the following table. Write answers to these questions in the table: Does the compound form a hydrate? How would you describe the appearance of the anhydrous compound? If your teacher approves, repeat the experimental procedure, using one of the listed compounds, and verify its hydrate formula. Explain any large deviation from the correct hydrate formula.

<table>
<thead>
<tr>
<th>Name of Compound</th>
<th>Water of crystallization?</th>
<th>Appearance of anhydrous compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium carbonate</td>
<td>Yes</td>
<td>White powder</td>
</tr>
<tr>
<td>Sodium sulfate</td>
<td>Yes</td>
<td>White powder</td>
</tr>
<tr>
<td>Sodium aluminum sulfate</td>
<td>Yes</td>
<td>White powder</td>
</tr>
<tr>
<td>Potassium chloride</td>
<td>Yes</td>
<td>White crystal powder</td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>Yes</td>
<td>White powder</td>
</tr>
<tr>
<td>Copper sulfate</td>
<td>Yes</td>
<td>White powder</td>
</tr>
</tbody>
</table>