**Hydrates: Salty salts with a hidden surprise!**

A **hydrate** is a **pure** substance (often ionic) that contains a fixed composition of water molecules (known as "waters of hydration") embedded in its crystal structure.

- Heating a hydrate "drives off" the water molecules, and the solid that remains behind is called **anhydrous**, meaning "**without water**."
- By measuring the mass of water removed when dehydrating a hydrate, we can determine the ratio of water molecules to anhydrous salt for a given hydrate, which allows us to find the formula of the hydrate.

**Notes about Language:** Talking about hydrates can be tricky! Here's a quick guide to the terminology used.

<table>
<thead>
<tr>
<th>Word/ Phrase</th>
<th>Meaning/ Context</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrate</td>
<td>Pure substance, typically crystalline, containing a fixed ratio of water molecules within its structure → this term is only used <strong>before</strong> water is removed by heating!</td>
</tr>
<tr>
<td>Waters of hydration</td>
<td>The embedded water molecules within the hydrate</td>
</tr>
<tr>
<td>Water molecules &quot;driven off&quot;</td>
<td>The process of forcing out the embedded water molecules in a hydrate through heating (do NOT call this evaporation – different context!)</td>
</tr>
<tr>
<td>Anhydrous salt</td>
<td>Compound remaining after all water molecules have been ‘driven off’ (removed)</td>
</tr>
<tr>
<td>Heating to “constant mass”</td>
<td>The process where the hydrate is heated multiple times and the mass measured each time until the mass stops changing, to ensure all of the water molecules have been driven off.</td>
</tr>
</tbody>
</table>

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**Hydrate Math**

The percent of water in a hydrate can be determined in a manner similar to determining the percent composition of a compound.

\[
\text{% water} = \frac{\text{mass of water lost}}{\text{mass of hydrate}} \times 100\%
\]

From the mass of water lost, moles of water can be calculated. From the mass of the anhydrous salt, moles of anhydrous salt can be calculated. The ratio of these quantities yields the number of water molecules per formula unit of hydrate.

\[
\frac{\text{moles of water}}{\text{moles of anhydrous salt}} = \text{molecules of water per formula unit of hydrate}
\]

**Steps to gravimetrically (by mass) determine the formula of a hydrate:**

1. Determine the **mass** of the water that has left the compound.
2. Convert the mass of water to **moles**.
3. Convert the mass of anhydrate that is left over to moles.
4. Find the water-to-anhydrate mole **ratio** (just like finding an **empirical** formula!)
5. Use the mole ratio to write the formula.
Example: The following data were obtained when a sample of barium chloride hydrate was analyzed:

<table>
<thead>
<tr>
<th>Mass</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of empty test tube</td>
<td>18.42 g</td>
</tr>
<tr>
<td>Mass of test tube and hydrate (before heating)</td>
<td>20.75 g</td>
</tr>
<tr>
<td>Mass of test tube and anhydrous salt (after heating)</td>
<td>20.41 g</td>
</tr>
</tbody>
</table>

a. Calculate the mass of water lost from the hydrate.

\[ 20.75 - 20.41 = \boxed{0.34 \text{ g } H_2O} \]

b. Calculate the percentage of water in the original sample.

Hydrate mass: \(20.75 - 18.42 = 2.33 \text{ g}\)

\[ \% H_2O = \frac{0.34}{2.33} \times 100 = \boxed{15 \% H_2O} \]

c. Calculate the moles of water lost from the sample.

\[ 0.34 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.016 \text{ g } H_2O} = 0.019 \text{ mol } H_2O \]

d. Calculate the moles of anhydrous salt remaining after the sample was heated.

\[ (\text{anhydrous}) \quad \text{BaCl}_2 : 20.41 - 18.42 = 1.99 \text{ g } \text{BaCl}_2 \times \frac{1 \text{ mol}}{208.23 \text{ g }} = 0.00956 \text{ mol } \text{BaCl}_2 \]

e. Determine the formula for the hydrate.

\[ \text{BaCl}_2 : 0.00956 \text{ mol} \]
\[ H_2O : 0.019 \text{ mol} \]
\[ \Rightarrow \left\{ \begin{array}{c} \text{BaCl}_2 \cdot 2H_2O \\ \end{array} \right\} \]

\[ \text{BaCl}_2 : 0.00956 \text{ mol} \]

\[ H_2O : 0.019 \text{ mol} \approx 2 \]
## Common Lab Errors when Determining the Formula of a Hydrate

<table>
<thead>
<tr>
<th>Error</th>
<th>Effect on Calculated % H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Crucible (container) not heated “to constant mass”</td>
<td></td>
</tr>
</tbody>
</table>
| • You didn’t heat the crucible to remove extra water molecules before getting the mass of the container and hydrate sample | • Water from the crucible will be lost as well as the expected water loss from heating the hydrate.  
• The calculated % H₂O will be larger than the actual % H₂O in the hydrate. |
| Excess heating caused the dehydrated sample to decompose. |
| • Often times, a gas will be released during the decomposition | • Gas from the decomposition will be lost as well as the expected water loss from heating the hydrate.  
• The calculated % H₂O will be larger than the actual % H₂O in the hydrate. |
| Strong initial heating caused some of the hydrate sample to spatter out of the crucible. |  
• Hydrated salt will be lost as well as the expected water loss from heating the hydrate.  
• The calculated % H₂O will be larger than the actual % H₂O in the hydrate. |
| The dehydrated sample absorbed moisture from the air after heating (but before the mass is measured). |  
• Not all of the waters of hydration will be removed.  
• The calculated % H₂O will be smaller than the actual % H₂O in the hydrate. |
| The hydrate is not heated to “constant mass” |  
• The hydrate should be heated multiple times and the mass measured each time, to ensure all of the water molecules have been driven off.  
• Not all of the water molecules will have been driven off, so the remaining salt is not completely anhydrous.  
• The calculated % H₂O will be smaller than the actual % H₂O in the hydrate. |

### Practice with Combustion Analysis and Hydrates: Fiery yet Thirst-Quenching!

**Multiple Choice Practice:**

1. The anhydrous salt X₂CO₃ has a molar mass of 106 g/mol. A hydrated form of this salt is heated until all of the water is removed and it loses 54% of its mass. The formula of the hydrate is:

   - (a) X₂CO₃·7H₂O  
   - (b) X₂CO₃·5H₂O  
   - (c) X₂CO₃·3H₂O  
   - (d) X₂CO₃·H₂O

   ⇒ H₂O ≈ 110 g H₂O  
   \( \frac{110 g H_2O}{18.016 g} \times \frac{1 mol}{18.016 g} \approx \frac{110}{20} < 5 mol H_2O \)

     
     \( (a \text{ little more than } 50\%) \)

   2. When the unbalanced reaction below occurs at STP, 1.5 L of CO₂ and 1.0 L of H₂O are created. What is the empirical formula of the hydrocarbon?

   \[ C_nH_m(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g) \]

   - b. CH₂  
   - b. C₂H₃  
   - c. C₂H₇  
   - d. C₃H₄

   \[ C: 1.5 L \times \frac{1 mol CO_2}{22.4 L} \times \frac{1 mol C}{1 mol CO_2} = \frac{1.5}{22.4} \text{ mol} \]

   \[ H: 1.0 L \times \frac{1 mol H_2O}{22.4 L} \times \frac{2 mol H}{1 mol H_2O} = \frac{2}{22.4} \text{ mol} \]

   \[ = 1 \]

   \[ \div \frac{1.5}{22.4} = \frac{2}{1.5} = \frac{4}{3} \]

   \[ = 3 \]

   \[ x^3 = 4 \]
3. A sample of a hydrate of CuSO₄ with a mass of 250 grams was heated until all the water was removed. The sample was then weighed and found to have a mass of 160 grams. What is the formula for the hydrate?

c. CuSO₄ · 10 H₂O  
b. CuSO₄ · 7 H₂O  
c. CuSO₄ · 5 H₂O  
d. CuSO₄ · 2 H₂O

\[
160 \text{ g CuSO}_4 \times \frac{1 \text{ mol CuSO}_4}{(16 + 32 + 16 \times 4) \text{ g}} = 1 \text{ mol CuSO}_4
\]

\[
250 - 160 = 90 \text{ g H₂O} \times \frac{1 \text{ mol}}{18.016 \text{ g}} \approx 9.0 \approx 5
\]

Free Response Practice (1991 Form B)

1. Answer the following questions about BeC₂O₄ and its hydrate.

   a. Calculate the mass percent of carbon in the hydrated solid with the formula BeC₂O₄·3 H₂O. (2 points)

   \[
   \% \text{ C} = \frac{2 \times 12.01}{151.08} \times 100 = 15.90 \%
   \]

   b. When heated to 220.°C, BeC₂O₄·3 H₂O dehydrates completely as represented below:

   \[
   \text{BeC}_2\text{O}_4 \cdot 3 \text{H}_2\text{O(s)} \rightarrow \text{BeC}_2\text{O}_4(s) + 3 \text{H}_2\text{O(g)}
   \]

   If 3.21 g of BeC₂O₄·3 H₂O is heated to 220.°C, calculate each of the following:

   i. The mass of BeC₂O₄ formed. (1 point)

   \[
   \frac{3.21 \text{ g BeC}_2\text{O}_4 \cdot 3 \text{H}_2\text{O}}{151.08 \text{ g}} \times \frac{1 \text{ mol BeC}_2\text{O}_4 \cdot 3 \text{H}_2\text{O}}{1 \text{ mol}} = 0.0212 \text{ mol}
   \]

   ii. The volume of H₂O(g) released, measured at 220.°C and 735 mmHg. (2 points)

   \[
   \frac{0.0212 \text{ mol BeC}_2\text{O}_4 \cdot 3 \text{H}_2\text{O}}{\text{k-fold}} \times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol BeC}_2\text{O}_4 \cdot 3 \text{H}_2\text{O}} = \frac{6.37 \times 10^{-2} \text{ mol H}_2\text{O}}{\text{k-fold}}
   \]

   \[
   \frac{\text{mmHg}}{\text{mol H}_2\text{O}} = \frac{6.37 \times 10^{-2} \text{ mol H}_2\text{O}}{735 \text{ mmHg}} = 2.67 \text{ L H}_2\text{O}
   \]
A student repeats the dehydration from part (b) in an attempt to experimentally determine the number of moles of water in one mole of BeC₂O₄·3H₂O. The student collects the data shown in the table below.

<table>
<thead>
<tr>
<th>Mass of empty crucible</th>
<th>36.48 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial mass of sample and crucible</td>
<td>39.69 g</td>
</tr>
<tr>
<td>Mass of sample and crucible after first heating</td>
<td>38.82 g</td>
</tr>
</tbody>
</table>

c. Use the data above to:
   
   i. Calculate the total number of moles of water lost when the sample was heated. (1 point)

   \[
   39.69 - 38.82 = 0.87 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 4.9 \times 10^{-2} \text{ mol H}_2\text{O}
   \]

   ii. Determine the formula of the hydrated compound. (2 points)

   \[
   38.82 - 36.48 = 2.34 \text{ g BeC}_2\text{O}_4 \times \frac{1 \text{ mol BeC}_2\text{O}_4}{97.03 \text{ g BeC}_2\text{O}_4} = 2.41 \times 10^{-2} \text{ mol BeC}_2\text{O}_4
   \]

   \[
   \text{BeC}_2\text{O}_4 : 2.41 \times 10^{-2} \text{ mol}
   \]

   \[
   \text{H}_2\text{O} : 4.9 \times 10^{-2} \text{ mol}
   \]

   \[
   \frac{2.41 \times 10^{-2}}{4.9 \times 10^{-2}} = 2\Rightarrow \text{BeC}_2\text{O}_4 \cdot 2\text{H}_2\text{O}
   \]

d. Is the student’s experimentally determined waters of hydration greater than, less than, or equal to the waters of hydration in the accepted formula? Provide a reasonable explanation for error and how this error affected the student’s results. (2 points)

   Less than! (experimental = BeC₂O₄·2H₂O, accepted = BeC₂O₄·3H₂O)

   It is unlikely that all H₂O was driven off after 1 heating, so the measured mass of anhydrous salt will be too high (i.e. it still contains H₂O) and the calculated mass of H₂O will be too low. Thus, the mole ratio as shown in the hydrated formula will show a lower ratio of H₂O : BeC₂O₄ than there should be (i.e., less waters of hydration than in the accepted formula).