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 tricksy! Here's a quick guide to the terminology used.

Word/ Phrase	Meaning/ Context
Hydrate (w/ H_2O)	Pure substance, typically crystalline, containing a fixed ratio of water molecules within its structure → this term is only used *before* water is removed by heating!
Water molecules "driven off"	The process of forcing out the embedded water molecules in a hydrate through heating (<u>do NOT call this evaporation</u> – different context!)
Anhydrous salt (no H_2O)	Compound remaining after all water molecules have been 'driven off' (removed)

→ or "anhydrate"

Hydrate Math

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

$$\% \ water = \frac{mass \ of \ water \ lost}{mass \ of \ hydrate} \times 100\%$$

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, but be careful: you can't multiply til whole! The mole ratio of the anhydrous salt must always be 1; only the number of waters can be a whole number greater than 1)
5. Use the mole ratio to write the formula.

Example: The following data were obtained when a sample of barium chloride hydrate was analyzed:

Mass of empty test tube	18.42 g
Mass of test tube and hydrate (before heating)	20.75 g
Mass of test tube and anhydrous salt (after heating)	20.41 g

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

$$20.75 - 20.41 = \boxed{0.34\ g\ H_2O}$$

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

hydrate mass $= 20.75 - 18.42 = 2.33\ g$

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

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

$$0.34\ g\ H_2O \times \frac{1\ mol}{18.016\ g} = \boxed{0.019\ mol\ H_2O}$$

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

anhydrous $BaCl_2 = 20.41 - 18.42$
$= 1.99\ g\ BaCl_2$

$$1.99\ g\ BaCl_2 \times \frac{1\ mol}{208.23\ g} = \boxed{0.00956\ mol\ BaCl_2}$$

e. Determine the formula /and name! for the hydrate.

$BaCl_2 : 0.00956\ mol$
$H_2O : 0.019\ mol$
$\div 0.00956$
$= 1$
≈ 2

$$\boxed{BaCl_2 \cdot 2H_2O}$$

barium chloride dihydrate

Common Lab Errors when Determine the Formula of a Hydrate

Error	Effect on Calculated % H_2O
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 % H_2O will be larger than the actual % H_2O 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 % H_2O will be Smaller than the actual % H_2O 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 % H_2O will be Smaller than the actual % H_2O in the hydrate.

Practice with Hydrates: Thirst-Quenching!

Error Analysis Practice:

1. A student wants to experimentally determine the number of moles of water in one mole of $BeC_2O_4 \cdot 3\,H_2O$. After heating the sample and measuring the new mass, the student calculated the amount of water driven off. Using that value and her experimental data, she derives the formula of the hydrate as $BeC_2O_4 \cdot 2\,H_2O$. Provide a reasonable explanation for an error that might have caused this outcome and explain how this error affected the student's results.

→ It is unlikely that all H_2O was driven off after 1 heating, so the measured mass of anhydrate will be too large (since it still contains H_2O), and the calculated mass of water will be too low. Thus, the student's experimentally determined hydrate formula shows less waters of hydration than actually exist.

2. Another student repeats the attempt to experimentally derive the empirical formula of of $BeC_2O_4 \cdot 3\,H_2O$. He decides to heat the sample multiple times over a higher heat to be certain to drive off all water from the hydrate sample. Using his data, this student calculated the formula of the hydrate to be $BeC_2O_4 \cdot 4\,H_2O$. Explain what error might have resulted in this outcome.

Because this student specifically heated his sample multiple times at high heat, it's likely some of the anhydrate decomposed, releasing a gas. The mass lost into the air caused the calculated mass of H_2O to be too high (since the mass of gas is included when calculating change in mass).

Multiple Choice Practice:

1. A sample of a hydrate of $CuSO_4$ 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?

   a. $CuSO_4 \cdot 10\ H_2O$ b. $CuSO_4 \cdot 7\ H_2O$ c. $CuSO_4 \cdot 5\ H_2O$ d. $CuSO_4 \cdot 2\ H_2O$

$$MM(CuSO_4) = 64 + 32 + 16 \cdot 4 = 160\ g/mol$$

$$\frac{160\ g}{160\ g/mol} = 1\ mol\ CuSO_4$$

mass H_2O $= 250 - 160 = \cancel{90}^{10}\ g\ H_2O \times \frac{1\ mol}{\cancel{18}_{2}\ g} = 5\ mol\ H_2O$

2. The anhydrous salt X_2CO_3 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:

   b. $X_2CO_3 \cdot 7\ H_2O$ b. $X_2CO_3 \cdot 5\ H_2O$ c. $X_2CO_3 \cdot 3\ H_2O$ d. $X_2CO_3 \cdot H_2O$

We don't know sample size ⇒ assume 1 mol for simplicity

hydrate mass ↓

$$\frac{106}{x} = 0.46 \Rightarrow x = \frac{106}{0.46} \approx 106 \times 2 = 212$$

if H_2O is ~50% of hydrate mass, ⇒ hydrate mass is about 2X MM (anhydrate)

mass H_2O $= 212 - 106 = 106\ g\ H_2O \times \frac{1\ mol}{18\ g}$

$$= \frac{106}{18} > \frac{100}{20} > 5$$