Totally Epic AP Chem Review: Hydrates and Combustion Analysis!

Combustion Analysis: Technique used to obtain the empirical formula of a hydrocarbon

→ Remember a standard (unbalanced) combustion reaction? (This formula is unbalanced!)

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

How to Solve a Combustion Analysis Problem

1. Convert mass of CO_2 and mass of H_2O to moles of each compound.
2. Convert moles of CO_2 to moles of carbon, and moles of H_2O to moles of hydrogen.
3. If compound contains something which is not C or H, find its mass by subtraction, and convert the mass to moles.
4. Now you have mole numbers! Complete the empirical formula calculation (divide by small, multiply til whole).

Practice:

1. Combustion analysis determined that a compound containing only carbon and hydrogen produces 1.83 g CO_2 and 0.901 g H_2O. Find the empirical formula of the compound.

$$C: 1.83\,g\,CO_2 \times \frac{1\,mol\,CO_2}{44.01\,g\,CO_2} \times \frac{1\,mol\,C}{1\,mol\,CO_2} = 0.0416\,mol \quad \div 0.0416 = 1 \quad \times 2 = 2$$

$$H: 0.901\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.016\,g\,H_2O} \times \frac{2\,mol\,H}{1\,mol\,H_2O} = 0.100\,mol \quad \div 0.0416 = 2.4 \quad \times 2 \approx 5$$

$$\boxed{C_2H_5}$$

2. Upon combustion, a 0.8233 g sample of a compound containing only carbon, hydrogen, and oxygen produces 2.445 g CO_2 and 0.6003 g H_2O. What is the empirical formula of the compound?

$$C: 2.445\,g\,CO_2 \times \frac{1\,mol\,CO_2}{44.01\,g\,CO_2} \times \frac{1\,mol\,C}{1\,mol\,CO_2} = 0.05555\,mol \times 12.01\,g/mol = 0.6672\,g\,C$$

$$H: 0.6003\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.016\,g\,H_2O} \times \frac{2\,mol\,H}{1\,mol\,H_2O} = 0.06664\,mol \times 1.008\,g/mol = 0.06717\,g\,H$$

$$O: 0.8233\,g - 0.6672 - 0.06717 = 0.0889\,g\,O \times \frac{1\,mol\,O}{16.00\,g\,O} = 0.00556\,mol\,O$$

$$C: 0.05555\,mol \quad \div 0.00556 = 10$$

$$H: 0.06664\,mol \quad \div 0.00556 = 12$$

$$O: 0.00556\,mol \quad \div 0.00556 = 1$$

$$\boxed{C_{10}H_{12}O}$$

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
Waters of hydration	The embedded water molecules within the hydrate
Heating to "constant mass"	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.

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: A student attempts to experimentally determine the number of moles of water in one mole of $BeC_2O_4 \cdot 3\ H_2O$. The student collects the data shown in the table below.

Mass of empty crucible	36.48 g
Initial mass of sample and crucible	39.69 g
Mass of sample and crucible after first heating	38.82 g

a. Calculate the total number of moles of water lost when the sample was heated.

b. Determine the formula of the hydrated compound.

a.) $39.69 - 38.82 = 0.87\text{ g } H_2O \times \dfrac{1\text{ mol } H_2O}{18.016\text{ g } H_2O} = \boxed{0.049\text{ mol } H_2O}$

b.) $38.82 - 36.48 = 2.34\text{ g } BeC_2O_4 \times \dfrac{1\text{ mol } BeC_2O_4}{97.03\text{ g } BeC_2O_4} = 0.0241\text{ mol } BeC_2O_4$

$BeC_2O_4 : 0.0241$ ÷ 0.0241 = 1

$H_2O : 0.049$ ÷ 0.0241 = 2

$\boxed{BeC_2O_4 \cdot 2H_2O}$

Common Lab Errors when Determine the Formula of a Hydrate

Error	Effect on Calculated % H_2O
Crucible (container) not heated "to Constant mass" • 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 % H_2O will be larger than the actual % H_2O 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 % H_2O will be larger than the actual % H_2O 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 % 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 Combustion Analysis and Hydrates: Fiery yet Thirst-Quenching!

1. When the unbalanced reaction below occurs at STP, 1.5 L of CO_2 and 1.0 L of H_2O are created. What is the empirical formula of the hydrocarbon?

$$C_xH_y(g) + O_2 \rightarrow CO_2(g) + H_2O(g)$$

a. CH_2 b. C_2H_3 c. C_2H_5 d. C_3H_4

$$C: 1.5\,L \times \frac{1\,mol}{22.4\,L} \times \frac{1\,mol\,C}{1\,mol\,CO_2} = \frac{1.5}{22.4}\,mol$$

$$H: 1.0\,L \times \frac{1\,mol}{22.4\,L} \times \frac{2\,mol\,H}{1\,mol\,H_2O} = \frac{2}{22.4}\,mol$$

$$\div \frac{1.5}{22.4}: \quad = 1, \quad = \frac{2}{1.5} = \frac{4}{3}$$

$$\times 3: \quad = 3, \quad = 4$$

2. 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?

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

$$160\,g\,CuSO_4 \times \frac{1\,mol}{\underbrace{(64+32+16\cdot 4)}_{160}\,g} = 1\,mol\,CuSO_4$$

$$250-160 = 90\,g\,H_2O \times \frac{1\,mol}{18.016\,g} \approx \frac{90}{20} \approx 5$$