# Unit: Unit 3: Chemical Reactions and Stoichiometry

## Chapter: Chapter 10: Stoichiometry

### Lesson: Lesson 4: Hydrates: Their Formulas and Reactions

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### Essential Questions:

- How do hydrates form, and how can we determine their composition?

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### 1. Big Idea:

Hydrates are compounds that include water molecules in their structure, and understanding their formulas helps predict their behavior in chemical reactions.

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### 2. Phenomenon-Based Learning:

**Unit Phenomenon:**

How can chemical reactions help improve safety features?

**Chapter Phenomenon:**

Now you have several ways to measure matter, by quantity of particles, mass, or volume. But how do those quantities relate to each other in a chemical equation? What is the ratio in which they react?

**Lesson Phenomenon:**

Water is a tricky substance, and it is often involved in many chemical reactions. Sometimes water stands by itself in a reaction, but other times it is part of other substances, like it integrates their formula! These are called hydrates. How would hydrates affect a chemical reaction? How should they be counted?

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### 3. Vocabulary:

- **Hydrates**: Compounds that have water molecules chemically bonded to them.

- **Anhydrous formula**: The formula of a compound without water molecules attached.

- **Greek prefix**: Prefixes like mono-, di-, tri-, etc., used to indicate the number of water molecules in a hydrate.

- **Hydrate formula**: The chemical formula that shows both the compound and the number of water molecules attached to it (e.g., CuSO₄·5H₂O).

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### 4. SMART Objectives:

- Calculate the percent by mass of water in a hydrate.

- Predict the products of reactions involving hydrates.

- Analyze the factors that affect the percent yield of a reaction.

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### 5. Engage (Ignite):

### # Pathfinder: Phenomenon-related question:

How does water in hydrates change the way a chemical reaction happens?

### # Mini-Task:

**Objective:** Observe and analyze how water in hydrates can be removed and how this affects the compound.

**Materials:**

- Epsom salt (MgSO₄·7H₂O)

- Heat source (Bunsen burner or hot plate)

- Small beaker

- Tongs

- Scale

**Procedure:**

1. Weigh a small amount (around 5 g) of Epsom salt and record the mass.

2. Gently heat the salt in a beaker using a Bunsen burner or hot plate.

3. Observe any changes (e.g., color, texture, or steam release).

4. After heating, weigh the remaining substance and record the mass.

**Questions:**

1. What happened to the mass of the Epsom salt after heating?

\*Answer: The mass decreased because the water in the hydrate was removed.\*

2. What does this tell us about the role of water in hydrates?

\*Answer: Water is part of the hydrate’s structure, and heating removes it, leaving behind the anhydrous compound.\*

3. How could this process be useful in real-life applications?

\*Answer: Removing water from hydrates can change their properties, which might be important for storage, reactions, or industrial processes.\*

**AI Tool Integration:**

Use an online molecular modeling tool to visualize the structure of hydrates like CuSO₄·5H₂O before and after water is removed. This can help understand how water molecules are bonded to the compound.

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### 6. Pre-Explore (Direct Instruction):

### # Prior Knowledge:

- Understanding of chemical formulas and the use of Greek prefixes (e.g., mono-, di-, tri-).

- Familiarity with basic stoichiometry concepts, such as molar mass and percent composition.

- Knowledge of physical changes, such as evaporation and condensation, from earlier grades.

### # Real-World Connection:

Epsom salt (MgSO₄·7H₂O) is commonly used in baths to relieve muscle pain. When left open, it can lose water and become a dry, powdery substance. This is an example of a hydrate losing its water content.

### # Background Information:

Hydrates are compounds that include water molecules as part of their structure. These water molecules are not just "mixed in" but are chemically bonded to the compound. For example, copper(II) sulfate pentahydrate (CuSO₄·5H₂O) contains five water molecules for every one copper(II) sulfate unit. When heated, hydrates lose their water and become anhydrous compounds.

Relating this to the phenomenon: Hydrates can play a role in chemical reactions, such as those in airbags. If a hydrate is used, the water it contains might affect the reaction's speed or products. This is why knowing the exact formula of a hydrate is important in stoichiometry.

### # Interactive Elements:

- **Discussion Prompt:** Why do you think water is included in the structure of some compounds?

- **Quick Question:** If a hydrate loses water when heated, what happens to its mass?

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### 7. Explore (Hands-On Activity):

### # Activity: Determining the Formula of a Hydrate

**Objective:** Use experimental data to calculate the formula of a hydrate.

**Materials:**

- Crucible

- Copper(II) sulfate pentahydrate (CuSO₄·5H₂O)

- Heat source (Bunsen burner or hot plate)

- Scale

- Tongs

**Procedure:**

1. Measure the mass of an empty crucible.

2. Add about 5 g of CuSO₄·5H₂O to the crucible and record the total mass.

3. Heat the crucible gently until the blue color of the hydrate changes to white (indicating the water has been removed).

4. Allow the crucible to cool, then measure the mass again.

5. Calculate the mass of water lost and the remaining anhydrous compound.

**Questions:**

1. What is the ratio of water to copper(II) sulfate in the compound?

\*Answer: 5:1 (based on the hydrate formula CuSO₄·5H₂O).\*

2. How can this ratio be used to write the formula of the hydrate?

\*Answer: The ratio tells us how many water molecules are bonded to each formula unit of the compound.\*

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### 8. Explain (Concept Development):

- Hydrates are named using Greek prefixes to indicate the number of water molecules. For example, MgSO₄·7H₂O is magnesium sulfate heptahydrate.

- When hydrates are heated, they lose water and become anhydrous compounds. This process is called dehydration.

- The percent by mass of water in a hydrate can be calculated using the formula:

\[

\text{Percent by mass of water} = \frac{\text{Mass of water}}{\text{Total mass of hydrate}} \times 100

\]

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### 9. Elaborate (Extend Understanding):

- Research how hydrates are used in industry, such as in cement production or as drying agents (e.g., anhydrous calcium chloride).

- Explore how the presence of water in hydrates might affect their storage or use in chemical reactions.

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### 10. Evaluate (Check Understanding):

**Questions:**

1. Define a hydrate and give an example.

\*Answer: A hydrate is a compound with water molecules chemically bonded to it, such as CuSO₄·5H₂O.\*

2. What happens when a hydrate is heated?

\*Answer: It loses water and becomes an anhydrous compound.\*

3. Calculate the percent by mass of water in MgSO₄·7H₂O.

\*Answer: Using molar masses, the percent by mass of water is approximately 51%.\*

**Extension Task:**

Predict how the removal of water from a hydrate might change its physical or chemical properties.

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This lesson plan is designed to build curiosity, connect to real-world applications, and develop a deep understanding of hydrates and their role in stoichiometry.

### Lesson 4: Hydrates – Their Formulas and Reactions

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### # Introduction: What Are Hydrates?

Have you ever noticed that some substances seem dry but actually contain water? These substances are called **hydrates**. Hydrates are compounds that have water molecules attached to them as part of their structure. The water in a hydrate is not just mixed in like when you add water to a drink. Instead, the water molecules are chemically bonded to the compound in a fixed ratio.

For example, a common hydrate is **copper(II) sulfate pentahydrate**, which has the formula **CuSO₄·5H₂O**. This means that for every one copper sulfate molecule, there are five water molecules attached to it. These water molecules are called **water of hydration**.

Hydrates are important in many areas of chemistry because they can affect how a chemical reaction happens. In this lesson, we’ll learn how to calculate the percent by mass of water in a hydrate, predict what happens when hydrates react, and analyze how hydrates can influence the outcome of a reaction.

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### # Section 1: The Formulas of Hydrates

Hydrates have a unique way of showing water in their chemical formula. The formula is written with a dot separating the main compound and the water. For example:

- **Na₂CO₃·10H₂O**: Sodium carbonate decahydrate (10 water molecules for every sodium carbonate unit)

- **MgSO₄·7H₂O**: Magnesium sulfate heptahydrate (7 water molecules for every magnesium sulfate unit)

The number of water molecules in a hydrate is always a whole number. This is because the water molecules are part of the crystal structure of the compound.

### ## Real-World Connection: Why Hydrates Matter

Hydrates are used in many everyday products. For example:

- **Epsom salt** (used for relaxing baths) is magnesium sulfate heptahydrate (**MgSO₄·7H₂O**).

- **Gypsum**, a key ingredient in drywall, is calcium sulfate dihydrate (**CaSO₄·2H₂O**).

When hydrates are heated, they lose their water molecules and become **anhydrous** (without water). This process is called **dehydration**.

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### # Section 2: Calculating the Percent by Mass of Water in a Hydrate

One important property of hydrates is the **percent by mass of water**. This tells us how much of the hydrate's mass comes from water. To calculate this, we use the formula:

\[

\text{Percent by Mass of Water} = \left( \frac{\text{Mass of Water in the Hydrate}}{\text{Total Mass of the Hydrate}} \right) \times 100

\]

### ## Example Problem 1: Percent by Mass of Water

**Question:** What is the percent by mass of water in copper(II) sulfate pentahydrate (**CuSO₄·5H₂O**)?

**Step 1:** Find the molar masses of the components.

- Molar mass of CuSO₄ = 63.55 (Cu) + 32.07 (S) + 4 × 16.00 (O) = 159.62 g/mol

- Molar mass of 5H₂O = 5 × [2(1.01) + 16.00] = 90.10 g/mol

- Total molar mass of CuSO₄·5H₂O = 159.62 + 90.10 = 249.72 g/mol

**Step 2:** Calculate the percent by mass of water.

\[

\text{Percent by Mass of Water} = \left( \frac{90.10}{249.72} \right) \times 100 = 36.07\%

\]

**Answer:** The percent by mass of water in copper(II) sulfate pentahydrate is **36.07%**.

### ## Practice Problem 1:

What is the percent by mass of water in sodium carbonate decahydrate (**Na₂CO₃·10H₂O**)?

\*(Answer: 62.93%)\*

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### # Section 3: Reactions Involving Hydrates

When hydrates are heated, they lose their water molecules. This process is called **dehydration**. The reaction can be written as:

\[

\text{Hydrate} \xrightarrow{\text{heat}} \text{Anhydrous Compound} + \text{Water Vapor}

\]

For example:

\[

CuSO₄·5H₂O \xrightarrow{\text{heat}} CuSO₄ + 5H₂O

\]

The anhydrous compound often looks different from the hydrate. For example, copper(II) sulfate pentahydrate is blue, but anhydrous copper(II) sulfate is white.

### ## Example Problem 2: Predicting Products

**Question:** What happens when magnesium sulfate heptahydrate (**MgSO₄·7H₂O**) is heated?

**Answer:** When heated, magnesium sulfate heptahydrate loses its 7 water molecules and becomes anhydrous magnesium sulfate (**MgSO₄**). The reaction is:

\[

MgSO₄·7H₂O \xrightarrow{\text{heat}} MgSO₄ + 7H₂O

\]

### ## Practice Problem 2:

What happens when calcium sulfate dihydrate (**CaSO₄·2H₂O**) is heated?

\*(Answer: It becomes anhydrous calcium sulfate (**CaSO₄**) and releases 2 water molecules.)\*

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### # Section 4: Factors Affecting the Percent Yield of a Reaction

In real-life chemical reactions, we don’t always get the amount of product we expect. The **percent yield** tells us how much of the expected product we actually got. It is calculated using the formula:

\[

\text{Percent Yield} = \left( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100

\]

### ## Factors That Affect Percent Yield:

1. **Impurities in Reactants:** If the reactants aren’t pure, the reaction may not go as planned.

2. **Incomplete Reactions:** Sometimes, the reaction doesn’t go to completion.

3. **Loss of Product:** During handling or transfer, some product might be lost.

### ## Example Problem 3: Percent Yield

**Question:** A student heats 10.0 g of CuSO₄·5H₂O and expects to get 6.4 g of anhydrous CuSO₄. However, they only collect 5.8 g. What is the percent yield?

**Step 1:** Use the formula for percent yield.

\[

\text{Percent Yield} = \left( \frac{5.8}{6.4} \right) \times 100 = 90.63\%

\]

**Answer:** The percent yield is **90.63%**.

### ## Practice Problem 3:

A student heats 15.0 g of MgSO₄·7H₂O and expects to get 7.4 g of MgSO₄. If they collect 6.8 g, what is the percent yield?

\*(Answer: 91.89%)\*

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### Progress Check: Questions and Answers

1. **What is a hydrate? Give one example.**

\*(Answer: A hydrate is a compound that contains water molecules chemically bonded to it. Example: Copper(II) sulfate pentahydrate (**CuSO₄·5H₂O**).)\*

2. **What happens to a hydrate when it is heated? Write the reaction for heating sodium carbonate decahydrate.**

\*(Answer: When a hydrate is heated, it loses its water molecules and becomes anhydrous. Reaction: **Na₂CO₃·10H₂O → Na₂CO₃ + 10H₂O**.)\*

3. **A student calculates the percent by mass of water in a hydrate and gets 40%. If the total mass of the hydrate is 250 g, how much of the mass comes from water?**

\*(Answer: 40% of 250 g is 0.40 × 250 = 100 g. So, 100 g of the mass comes from water.)\*

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By understanding hydrates, their formulas, and their reactions, you can see how water plays a key role in chemistry. Whether it’s in the safety features of cars or the products we use every day, chemical reactions involving hydrates are all around us!

# Lesson 4: Hydrates – Their Formulas and Reactions

## Introduction: What Are Hydrates?

Have you ever noticed how some substances seem to "trap" water? These substances are called **hydrates**. A hydrate is a compound that has water molecules chemically bonded to it. The water isn’t just sitting there—it’s part of the compound’s structure. For example, a common hydrate is gypsum, which is used to make drywall. Its formula is **CaSO₄·2H₂O**, meaning it contains calcium sulfate and two water molecules for every formula unit.

Hydrates are fascinating because they show us how water can integrate into other substances. But why does this matter? Well, hydrates can affect chemical reactions, and understanding them helps us predict how much of a product we’ll get from a reaction. This is important in industries like medicine, construction, and even car safety systems!

In this lesson, we’ll explore how to calculate the percent by mass of water in a hydrate, predict what happens when hydrates react, and analyze factors that affect the percent yield of a reaction. Let’s dive in!

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## Section 1: The Formula of a Hydrate

### What Does the Formula Tell Us?

The formula of a hydrate tells us how many water molecules are attached to each formula unit of the compound. For example:

- **CuSO₄·5H₂O** means that for every one unit of copper(II) sulfate, there are five water molecules.

- **MgSO₄·7H₂O** means that for every one unit of magnesium sulfate, there are seven water molecules.

The dot (·) in the formula separates the main compound from the water molecules.

### Why Are Hydrates Important?

Hydrates are important because they behave differently from anhydrous compounds (compounds without water). For example, **anhydrous CuSO₄** is a white powder, while **CuSO₄·5H₂O** is blue. The water molecules in hydrates can affect their color, texture, and even how they react in chemical processes.

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### Sample Problem 1: Determining the Percent by Mass of Water in a Hydrate

Let’s calculate the percent by mass of water in **CuSO₄·5H₂O**.

### # Step 1: Find the molar masses of the compound and water.

- Molar mass of CuSO₄ = 63.55 (Cu) + 32.07 (S) + 4 × 16.00 (O) = **159.62 g/mol**

- Molar mass of 5H₂O = 5 × [2 × 1.01 (H) + 16.00 (O)] = **90.10 g/mol**

### # Step 2: Add them together to get the total molar mass.

- Total molar mass = 159.62 + 90.10 = **249.72 g/mol**

### # Step 3: Calculate the percent by mass of water.

\[

\text{Percent by mass of water} = \left( \frac{\text{Mass of water}}{\text{Total mass}} \right) \times 100

\]

\[

\text{Percent by mass of water} = \left( \frac{90.10}{249.72} \right) \times 100 = 36.08\%

\]

So, **36.08%** of the mass of CuSO₄·5H₂O is water.

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### Practice Question 1:

Calculate the percent by mass of water in **MgSO₄·7H₂O**.

**Answer:**

- Molar mass of MgSO₄ = 24.31 (Mg) + 32.07 (S) + 4 × 16.00 (O) = **120.38 g/mol**

- Molar mass of 7H₂O = 7 × [2 × 1.01 (H) + 16.00 (O)] = **126.14 g/mol**

- Total molar mass = 120.38 + 126.14 = **246.52 g/mol**

- Percent by mass of water = \(\left( \frac{126.14}{246.52} \right) \times 100 = 51.18\%\)

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## Section 2: Reactions Involving Hydrates

### What Happens When Hydrates Are Heated?

When hydrates are heated, they lose their water molecules. This process is called **dehydration**. For example:

\[

CuSO₄·5H₂O \xrightarrow{\text{heat}} CuSO₄ + 5H₂O

\]

In this reaction, blue copper(II) sulfate pentahydrate turns into white anhydrous copper(II) sulfate.

### Why Is This Useful?

Dehydration reactions are useful because they help us determine the number of water molecules in a hydrate. By measuring how much mass is lost when a hydrate is heated, we can figure out how much water was in the compound.

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### Sample Problem 2: Predicting the Products of a Reaction

**Problem:** What are the products when **MgSO₄·7H₂O** is heated?

**Solution:**

When heated, MgSO₄·7H₂O loses its water molecules:

\[

MgSO₄·7H₂O \xrightarrow{\text{heat}} MgSO₄ + 7H₂O

\]

The products are anhydrous magnesium sulfate (**MgSO₄**) and water (**H₂O**).

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### Practice Question 2:

What are the products when **BaCl₂·2H₂O** is heated?

**Answer:**

\[

BaCl₂·2H₂O \xrightarrow{\text{heat}} BaCl₂ + 2H₂O

\]

The products are anhydrous barium chloride (**BaCl₂**) and water (**H₂O**).

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## Section 3: Percent Yield of a Reaction

### What Is Percent Yield?

In a chemical reaction, the **percent yield** tells us how much product we actually got compared to how much we expected to get. It’s calculated using this formula:

\[

\text{Percent yield} = \left( \frac{\text{Actual yield}}{\text{Theoretical yield}} \right) \times 100

\]

### Why Does Percent Yield Matter?

Percent yield is important because it helps us understand how efficient a reaction is. In real-world applications, reactions rarely produce 100% of the expected product due to factors like impurities or incomplete reactions.

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### Sample Problem 3: Calculating Percent Yield

**Problem:** A reaction is expected to produce 10.0 g of anhydrous CuSO₄, but only 8.5 g is obtained. What is the percent yield?

**Solution:**

\[

\text{Percent yield} = \left( \frac{8.5}{10.0} \right) \times 100 = 85.0\%

\]

The percent yield is **85.0%**.

---

### Practice Question 3:

A reaction is expected to produce 15.0 g of MgSO₄, but only 12.0 g is obtained. What is the percent yield?

**Answer:**

\[

\text{Percent yield} = \left( \frac{12.0}{15.0} \right) \times 100 = 80.0\%

\]

The percent yield is **80.0%**.

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## Real-World Connection: Hydrates in Car Safety

Hydrates play a role in car safety systems like airbags. The gas that inflates an airbag often comes from a chemical reaction involving a compound that contains water. For example, sodium azide (**NaN₃**) reacts to produce nitrogen gas, which fills the airbag. The reaction must happen quickly and safely, so engineers carefully design the system to control the amount of gas produced.

Understanding hydrates and their reactions helps engineers predict how much gas will be released and ensure the airbag inflates properly to protect passengers.

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## Progress Check

### Question 1: What is a hydrate?

**Answer:** A hydrate is a compound that has water molecules chemically bonded to it.

### Question 2: What happens to a hydrate when it is heated?

**Answer:** When a hydrate is heated, it loses its water molecules in a process called dehydration.

### Question 3: Why is percent yield important in chemical reactions?

**Answer:** Percent yield is important because it shows how efficient a reaction is and helps us understand how much product we actually got compared to how much we expected.

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By understanding hydrates and their reactions, we can predict how substances will behave in real-world situations, from industrial processes to life-saving technologies like airbags. Keep practicing, and you’ll see just how useful chemistry can be!