# 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?

- What happens to hydrates during chemical reactions?

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### 1. Big Idea:

Hydrates are compounds that contain water molecules as part of their structure, and understanding their formulas and reactions helps in predicting chemical behavior.

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### 2. Phenomenon-Based Learning:

### # Unit Phenomenon:

How can chemical reactions help improve safety features?

### # Chapter Phenomenon:

How do quantities of matter relate in chemical equations, and how do they affect reactions?

### # Lesson Phenomenon:

Water often participates in chemical reactions in surprising ways. In hydrates, water is chemically bound within a compound. How does this water influence the compound’s behavior and reactions?

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### 3. Vocabulary:

- **Hydrates**: Compounds that have water molecules chemically bonded within their structure.

- **Anhydrous formula**: The formula of a compound after all water has been removed.

- **Greek prefix**: A prefix used to indicate the number of water molecules in a hydrate (e.g., mono-, di-, tri-).

- **Hydrate formula**: The chemical formula of a hydrate, showing the ratio of water molecules to the compound (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 hidden in a compound affect a chemical reaction?

### # Mini-task:

**Objective**: Observe how heating a hydrate releases water.

**Materials**:

- A small amount of copper(II) sulfate pentahydrate (CuSO₄·5H₂O)

- A heatproof dish

- A Bunsen burner or candle

- Tongs

- Safety goggles

**Procedure**:

1. Place a small amount of copper(II) sulfate pentahydrate in the heatproof dish.

2. Observe its blue color.

3. Gently heat the compound over a flame while wearing safety goggles.

4. Watch for changes in color and texture as the water is released.

5. Allow the dish to cool and observe the final product.

**Questions**:

1. What color was the copper(II) sulfate before heating?

\*Answer: Blue.\*

2. What happened to the color after heating?

\*Answer: It turned white or grayish, indicating the loss of water.\*

3. What does this tell you about the role of water in hydrates?

\*Answer: Water is part of the structure and affects the compound’s properties, like color.\*

**AI Tool Task**:

Use an online molecular visualization tool (such as MolView or ChemSketch) to model the structure of a hydrate and its anhydrous form. Compare the two and discuss how the water molecules are arranged in the hydrate.

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

### # Prior Knowledge:

- Chemical formulas represent the types and numbers of atoms in a compound.

- Water (H₂O) is a common molecule in the environment and can participate in various chemical reactions.

- Heating substances can cause physical and chemical changes.

### # Real-World Connection:

Have you ever noticed that some packets of food or electronics come with small silica gel packets labeled "Do not eat"? These packets absorb water to keep things dry. Similarly, hydrates naturally absorb water into their structure, but they release it when heated.

### # Background Information:

Hydrates are special compounds that include water molecules as part of their chemical structure. The water is not simply mixed in; it is bonded to the compound in a fixed ratio. For example, copper(II) sulfate pentahydrate (CuSO₄·5H₂O) contains five water molecules for every one unit of copper(II) sulfate. When hydrates are heated, they lose their water and become "anhydrous" (without water).

This concept is important in stoichiometry because the water in hydrates must be included when calculating the mass or reacting quantities of a compound. For example, the blue color of CuSO₄·5H₂O is due to the water in its structure. When the water is removed by heating, the compound turns white, indicating it has become anhydrous.

### # Interactive Notes:

- Discuss the Greek prefixes used in naming hydrates (e.g., mono- for one, di- for two, tri- for three).

- Ask: What does the formula CuSO₄·5H₂O tell us about this compound?

\*Answer: It contains one unit of copper(II) sulfate and five water molecules.\*

- Relate to the phenomenon: How might the water in hydrates affect a chemical reaction?

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### 7. Explore:

**Hands-on Investigation**:

**Objective**: Determine the percent by mass of water in a hydrate.

**Materials**:

- Epsom salt (MgSO₄·7H₂O)

- Heatproof dish

- Bunsen burner or candle

- Scale

- Tongs

- Safety goggles

**Procedure**:

1. Weigh a small amount of Epsom salt and record its mass.

2. Heat the salt in a dish until it turns white (indicating the water has been removed).

3. Allow the dish to cool and weigh the anhydrous salt.

4. Calculate the mass of the water lost by subtracting the mass of the anhydrous salt from the original mass.

5. Use the formula below to calculate the percent by mass of water:

\[

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

\]

**Questions**:

1. How much water was lost during heating?

2. What is the percent by mass of water in the hydrate?

3. Why is it important to include the water when calculating the formula of a hydrate?

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### 8. Explain:

- Hydrates are compounds that include water molecules in their structure.

- Heating a hydrate removes the water, leaving behind an anhydrous compound.

- The percent by mass of water in a hydrate can be calculated using the mass of the water lost and the original mass of the hydrate. This helps chemists determine the formula of the hydrate.

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### 9. Elaborate:

**Real-World Application**:

Hydrates are used in many industries. For example, gypsum (CaSO₄·2H₂O) is a hydrate used to make drywall. When heated, it loses water and becomes plaster of Paris, which is used to make molds and casts.

**Discussion**:

- How might the water in hydrates affect their use in construction or medicine?

- Why is it important to know the exact formula of a hydrate?

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### 10. Evaluate:

**Task**:

Given the formula for a hydrate, such as BaCl₂·2H₂O, calculate the percent by mass of water.

**Challenge Question**:

If 10 grams of a hydrate lose 4 grams of water when heated, what is the percent by mass of water in the hydrate?

\*Answer: 40%\*

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This lesson plan provides a structured, engaging, and hands-on approach to learning about hydrates, ensuring that learners connect the content to real-world applications and develop a deeper understanding of chemical reactions.

# Lesson 4: Hydrates – Their Formulas and Reactions

## Introduction: What Are Hydrates?

Have you ever noticed how salt sometimes clumps together when left in a humid place? That’s because it absorbs water from the air. Some substances, like salt, can trap water molecules within their structure. When water becomes part of a compound's structure, the compound is called a **hydrate**. Hydrates are fascinating because they show how water can sneak into chemical formulas and become part of a solid substance.

For example, **copper(II) sulfate pentahydrate** (CuSO₄·5H₂O) is a hydrate. The "·5H₂O" means that for every one copper sulfate molecule, there are five water molecules attached to it. These water molecules are called **water of hydration**, and they play an important role in the properties of the compound.

But what happens when hydrates are heated? The water can leave, and the compound becomes **anhydrous**, meaning it no longer has water in its structure. This change is important in chemical reactions, as the presence or absence of water can affect how substances behave.

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## Section 1: Hydrate Formulas – Understanding the Basics

### What Do Hydrate Formulas Look Like?

A hydrate formula shows the main compound and the number of water molecules attached to it. For example:

- **Calcium chloride dihydrate (CaCl₂·2H₂O)**: This means there are two water molecules for every one calcium chloride molecule.

- **Magnesium sulfate heptahydrate (MgSO₄·7H₂O)**: This means there are seven water molecules for every one magnesium sulfate molecule.

The number of water molecules is often shown using **Greek prefixes**:

- Mono- (1)

- Di- (2)

- Tri- (3)

- Tetra- (4)

- Penta- (5)

- Hexa- (6)

- Hepta- (7)

- Octa- (8)

### Why Do Hydrates Matter?

Hydrates are important in many real-world applications. For example:

- **Concrete**: When cement reacts with water, it forms hydrates that make the concrete hard and strong.

- **Medicine**: Some drugs are stored as hydrates to keep them stable.

- **Airbags**: The chemical reactions that inflate airbags can involve hydrates.

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### Solved Example 1: Writing a Hydrate Formula

**Problem**: Write the formula for a hydrate that contains one molecule of sodium carbonate (Na₂CO₃) and ten water molecules.

**Solution**:

1. Write the formula for sodium carbonate: Na₂CO₃.

2. Add the number of water molecules using a dot: Na₂CO₃·10H₂O.

3. The formula is **sodium carbonate decahydrate (Na₂CO₃·10H₂O)**.

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### Practice Question 1:

Write the formula for a hydrate that contains one molecule of barium chloride (BaCl₂) and two water molecules.

**Answer**: BaCl₂·2H₂O (barium chloride dihydrate)

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

### How Do We Calculate the Percent by Mass of Water?

The percent by mass of water in a hydrate tells us how much of the compound’s mass comes from water. Here’s the formula:

\[

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

\]

### Solved Example 2: Calculating Percent by Mass of Water

**Problem**: Calculate the percent by mass of water in copper(II) sulfate pentahydrate (CuSO₄·5H₂O). The molar masses are:

- CuSO₄ = 159.6 g/mol

- H₂O = 18.0 g/mol

**Solution**:

1. Find the mass of water: \( 5 \times 18.0 = 90.0 \, \text{g/mol} \).

2. Find the total mass of the hydrate: \( 159.6 + 90.0 = 249.6 \, \text{g/mol} \).

3. Use the formula:

\[

\text{Percent by mass of water} = \frac{90.0}{249.6} \times 100 = 36.1\%

\]

So, the percent by mass of water is **36.1%**.

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### Practice Question 2:

Calculate the percent by mass of water in magnesium sulfate heptahydrate (MgSO₄·7H₂O). The molar masses are:

- MgSO₄ = 120.4 g/mol

- H₂O = 18.0 g/mol

**Answer**: \( \text{Percent by mass of water} = 51.2\% \)

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## Section 3: Reactions Involving Hydrates

### What Happens When Hydrates Are Heated?

When hydrates are heated, the water of hydration is released, leaving behind the **anhydrous compound**. This process is called **dehydration**. For example:

\[

\text{CuSO₄·5H₂O (s)} \xrightarrow{\text{heat}} \text{CuSO₄ (s)} + 5\text{H₂O (g)}

\]

The blue crystals of copper(II) sulfate pentahydrate turn into a white powder of anhydrous copper(II) sulfate when heated.

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### Solved Example 3: Predicting Products of a Reaction

**Problem**: Predict the products when barium chloride dihydrate (BaCl₂·2H₂O) is heated.

**Solution**:

1. Identify the hydrate: BaCl₂·2H₂O.

2. Write the products: The water is released as steam, and the anhydrous compound remains.

\[

\text{BaCl₂·2H₂O (s)} \xrightarrow{\text{heat}} \text{BaCl₂ (s)} + 2\text{H₂O (g)}

\]

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### Practice Question 3:

Predict the products when magnesium sulfate heptahydrate (MgSO₄·7H₂O) is heated.

**Answer**: MgSO₄ (s) + 7H₂O (g)

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## Section 4: Factors That Affect Percent Yield of a Reaction

### What Is Percent Yield?

Percent yield measures how much product you actually get in a reaction compared to how much you expected to get. Here’s the formula:

\[

\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100

\]

### Why Is Percent Yield Important?

In real-world reactions, you rarely get 100% yield. Some reasons include:

- **Incomplete reactions**: Not all reactants turn into products.

- **Side reactions**: Other reactions might occur, forming unwanted products.

- **Loss of material**: Some product might be lost during handling.

For example, in the production of airbags, the percent yield of the gas-producing reaction must be high to ensure the airbag inflates properly.

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### Solved Example 4: Calculating Percent Yield

**Problem**: In a reaction, the theoretical yield of a product is 50.0 g, but the actual yield is 45.0 g. What is the percent yield?

**Solution**:

1. Use the formula:

\[

\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100

\]

2. Plug in the values:

\[

\text{Percent yield} = \frac{45.0}{50.0} \times 100 = 90.0\%

\]

So, the percent yield is **90.0%**.

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### Practice Question 4:

In a reaction, the theoretical yield is 75.0 g, but the actual yield is 60.0 g. What is the percent yield?

**Answer**: \( \text{Percent yield} = 80.0\% \)

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## Real-World Connection: Hydrates in Safety Features

Hydrates play a role in the chemical reactions that inflate airbags. For example, sodium azide (NaN₃) decomposes to produce nitrogen gas, which inflates the airbag. Some reactions involve hydrates to control the reaction rate or ensure safety. The water in hydrates can also help absorb heat, making the reaction safer for passengers.

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## Progress Check

### Question 1:

What is a hydrate? Provide an example.

**Answer**: A hydrate is a compound that contains water molecules within its structure. Example: Copper(II) sulfate pentahydrate (CuSO₄·5H₂O).

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### Question 2:

Calculate the percent by mass of water in calcium chloride dihydrate (CaCl₂·2H₂O). The molar masses are:

- CaCl₂ = 111.0 g/mol

- H₂O = 18.0 g/mol

**Answer**: \( \text{Percent by mass of water} = 24.3\% \)

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### Question 3:

Why is it important to consider percent yield in chemical reactions?

**Answer**: Percent yield is important because it shows how efficient a reaction is. In real-world applications, like airbags, a high percent yield ensures the reaction produces enough gas to inflate the airbag properly.

# Lesson 4: Hydrates – Their Formulas and Reactions

Have you ever noticed how water seems to be everywhere in chemistry? It’s not just a liquid we drink or swim in—it’s also a key player in many chemical reactions. Sometimes, water is a separate part of a reaction. Other times, it’s locked into the structure of a compound. These compounds are called **hydrates**, and they play a big role in chemistry. But what exactly are hydrates? How do they affect chemical reactions? And how can we calculate their properties? Let’s dive in and explore these questions step by step.

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## What Are Hydrates?

Hydrates are special compounds that contain water molecules as part of their structure. The water isn’t just mixed in—it’s actually bonded to the compound in a specific ratio. For example, the compound copper(II) sulfate pentahydrate has the formula **CuSO₄·5H₂O**. This means that for every one copper(II) sulfate molecule (CuSO₄), there are five water molecules (H₂O) attached to it.

The water in a hydrate is called **water of hydration**, and it can be removed by heating the compound. When the water is removed, the compound becomes **anhydrous**, which means “without water.” For example, if you heat copper(II) sulfate pentahydrate, the water evaporates, leaving behind anhydrous copper(II) sulfate (**CuSO₄**).

### Real-World Example: Hydrates in Construction

Have you ever seen a bag of cement? Cement contains a hydrate called gypsum (**CaSO₄·2H₂O**). When water is added to cement, the gypsum reacts to form a hard, solid material. This reaction is what makes cement such a strong building material. Without the water in the gypsum, the cement wouldn’t work the same way!

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## The Formula of a Hydrate

The formula of a hydrate shows how many water molecules are attached to each formula unit of the compound. This is usually written with a dot separating the compound and the water. For example:

- **Na₂CO₃·10H₂O**: Sodium carbonate decahydrate (10 water molecules)

- **MgSO₄·7H₂O**: Magnesium sulfate heptahydrate (7 water molecules)

The number of water molecules is often indicated using **Greek prefixes**:

- Mono = 1

- Di = 2

- Tri = 3

- Tetra = 4

- Penta = 5

- Hexa = 6

- Hepta = 7

- Octa = 8

- Nona = 9

- Deca = 10

For example, **pentahydrate** means there are 5 water molecules in the hydrate.

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## 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 what percentage of the hydrate’s mass comes from water. To calculate it, we use the following formula:

\[

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

\]

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

**Problem:**

A sample of magnesium sulfate heptahydrate (**MgSO₄·7H₂O**) has a molar mass of 246.47 g/mol. The water in the hydrate has a total molar mass of 126.07 g/mol. What is the percent by mass of water in the hydrate?

**Solution:**

1. Use the formula:

\[

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

\]

2. Plug in the values:

\[

\text{Percent by mass of water} = \left( \frac{126.07}{246.47} \right) \times 100

\]

3. Simplify:

\[

\text{Percent by mass of water} = 51.16\%

\]

So, 51.16% of the mass of magnesium sulfate heptahydrate comes from water.

**Practice Question:**

A sample of copper(II) sulfate pentahydrate (**CuSO₄·5H₂O**) has a molar mass of 249.68 g/mol. The water in the hydrate has a molar mass of 90.08 g/mol. What is the percent by mass of water in the hydrate?

**Answer:**

\[

\text{Percent by mass of water} = \left( \frac{90.08}{249.68} \right) \times 100 = 36.08\%

\]

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## Predicting Reactions Involving Hydrates

When hydrates are heated, they lose their water of hydration. This process is called **dehydration**, and it leaves behind the anhydrous compound.

### Example Reaction:

\[

\text{CuSO₄·5H₂O (s)} \xrightarrow{\text{heat}} \text{CuSO₄ (s)} + 5\text{H₂O (g)}

\]

In this reaction, copper(II) sulfate pentahydrate is heated to produce anhydrous copper(II) sulfate and water vapor.

**Practice Question:**

What are the products when calcium chloride dihydrate (**CaCl₂·2H₂O**) is heated?

**Answer:**

\[

\text{CaCl₂·2H₂O (s)} \xrightarrow{\text{heat}} \text{CaCl₂ (s)} + 2\text{H₂O (g)}

\]

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## Factors That Affect 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** of a reaction tells us how much product we actually got compared to how much we expected.

\[

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

\]

For reactions involving hydrates, the percent yield can be affected by:

1. **Incomplete reactions:** Not all of the hydrate may dehydrate.

2. **Loss of product:** Some of the anhydrous compound may be lost during the reaction.

3. **Impurities:** The hydrate sample may contain impurities that affect the reaction.

### Example Problem: Calculating Percent Yield

**Problem:**

A student heats 10.0 g of copper(II) sulfate pentahydrate (**CuSO₄·5H₂O**) and expects to produce 6.4 g of anhydrous copper(II) sulfate (**CuSO₄**). However, the student only collects 5.8 g of CuSO₄. What is the percent yield?

**Solution:**

1. Use the formula:

\[

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

\]

2. Plug in the values:

\[

\text{Percent yield} = \left( \frac{5.8}{6.4} \right) \times 100

\]

3. Simplify:

\[

\text{Percent yield} = 90.63\%

\]

**Practice Question:**

A student heats 15.0 g of magnesium sulfate heptahydrate (**MgSO₄·7H₂O**) and expects to produce 7.3 g of anhydrous magnesium sulfate (**MgSO₄**). The student collects 6.9 g of MgSO₄. What is the percent yield?

**Answer:**

\[

\text{Percent yield} = \left( \frac{6.9}{7.3} \right) \times 100 = 94.52\%

\]

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## Wrap-Up

Hydrates are fascinating compounds that combine water molecules with other substances in a fixed ratio. By understanding their formulas, calculating the percent by mass of water, and predicting their reactions, we can better understand how they behave in the lab and in real-world applications. Whether it’s in construction materials like cement or in chemical reactions in airbags, hydrates play an important role in our everyday lives.

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## Progress Check

1. **What is the difference between a hydrate and an anhydrous compound?**

**Answer:** A hydrate contains water molecules as part of its structure, while an anhydrous compound does not contain water.

2. **Calculate the percent by mass of water in sodium carbonate decahydrate (**Na₂CO₃·10H₂O**), given that the molar mass of the hydrate is 286.14 g/mol and the water is 180.18 g/mol.**

**Answer:**

\[

\text{Percent by mass of water} = \left( \frac{180.18}{286.14} \right) \times 100 = 62.96\%

\]

3. **A student heats 12.0 g of calcium chloride dihydrate (**CaCl₂·2H₂O**) and collects 8.5 g of anhydrous calcium chloride (**CaCl₂**). What is the percent yield?**

**Answer:**

\[

\text{Percent yield} = \left( \frac{8.5}{9.0} \right) \times 100 = 94.44\%

\]