How Many Moles Of Water Are Produced When 6.0 Moles Of Hydrogen Gas React With 2.5 Moles Of Oxygen Gas?

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Understanding the chemical reaction between hydrogen and oxygen gases to produce water is fundamental in chemistry. When hydrogen gas reacts with oxygen gas, it forms water—a process that involves a well-known chemical equation. Specifically, the question asks: How many moles of water are produced when 6.0 moles of hydrogen gas react with 2.5 moles of oxygen gas? To accurately answer this, we need to analyze the balanced chemical equation, determine the limiting reactant, and then calculate the amount of water produced.

Balanced Chemical Equation for Hydrogen and Oxygen Reaction

The first step in solving any stoichiometry problem is to write and balance the chemical equation. The reaction between hydrogen gas (H_2) and oxygen gas (O_2) to form water (H_2O) is as follows:

Unbalanced Equation:

```plaintext
$H_2 + O_2 \rightarrow H_2O$
```

Balanced Equation:

```plaintext
$2\ H_2 + O_2 \rightarrow 2\ H_2O$
```

This balanced equation tells us that:
- 2 moles of hydrogen gas react with 1 mole of oxygen gas
- To produce 2 moles of water

Understanding the mole ratio from the balanced equation is critical for calculating the amount of water produced.

Determining the Limiting Reactant

Given:
- Hydrogen gas (H_2): 6.0 moles
- Oxygen gas (O_2): 2.5 moles

From the balanced equation:
- 2 moles of H_2 react with 1 mole of O_2
- Therefore, for 6.0 moles of H_2, the required O_2 is:

$$
\text{Required O}_2 = \frac{1 \text{ mole O}_2}{2 \text{ moles H}_2} \times 6.0 \text{ moles H}_2 = 3.0 \text{ moles O}_2
$$

But only 2.5 moles of O_2 are available, which is less than the 3.0 moles needed to fully react with 6.0 moles of H_2. Hence, Oxygen gas is the limiting reactant.

This means:
- All the available O_2 (2.5 moles) will be consumed
- Hydrogen gas remains in excess

Calculating the Moles of Water Produced

Since O_2 is limiting, the amount of water produced depends on the amount of O_2 available.

From the balanced equation:
- 1 mole of O_2 produces 2 moles of H_2O

Thus, for 2.5 moles of O_2:

$$
\text{Moles of H}_2O = 2 \times 2.5 \text{ moles} = 5.0 \text{ moles}
$$

Therefore, 5.0 moles of water are produced when 6.0 moles of hydrogen react with 2.5 moles of oxygen.

Summary of the Calculation Steps

1. Write the balanced chemical equation: $2\ H_2 + O_2 \rightarrow 2\ H_2O$
2. Determine the molar ratios from the balanced equation: 2 moles H_2 : 1 mole O_2 : 2 moles H_2O
3. Calculate the amount of O_2 needed for 6.0 moles of H_2: 3.0 moles
4. Compare with available O_2 (2.5 moles): O_2 is limiting
5. Calculate water produced based on limiting reactant: 2.5 moles O_2 produce 5.0 moles H_2O

Additional Considerations in Stoichiometry

Understanding the stoichiometric relationships helps in various chemical calculations beyond just water formation. Let's explore some related concepts:

Role of Limiting Reactants

- Limiting reactants determine the maximum amount of product formed.
- Excess reactants remain unreacted after the reaction reaches completion.

Calculating Theoretical Yield

- Theoretical yield refers to the maximum amount of product achievable based on limiting reactant.
- In this case, the theoretical yield of water is 5.0 moles.

Practical Implications

- In real-world scenarios, yields may be less due to side reactions or incomplete reactions.
- However, theoretical calculations provide essential benchmarks for chemical processes.

Practical Applications of Hydrogen and Oxygen Reactions

The reaction between hydrogen and oxygen to produce water is not just a theoretical exercise—it has practical significance in various fields:

Fuel Cells

- Hydrogen fuel cells generate electricity through the reaction $H_2 + \frac{1}{2} O_2 \rightarrow H_2O$.
- The process produces water as the only byproduct, making it environmentally friendly.

Industrial Water Production

- Large-scale synthesis of water in industries may involve catalytic processes involving hydrogen and oxygen.

Rocket Propulsion

- Hydrogen-oxygen rockets utilize the exothermic reaction to produce thrust, with water as a byproduct.

Conclusion: Summing Up the Moles of Water Produced

To summarize, when 6.0 moles of hydrogen gas react with 2.5 moles of oxygen gas, the limiting reactant is oxygen. Given the stoichiometry of the reaction, 5.0 moles of water are produced. This calculation exemplifies fundamental principles of stoichiometry, including balancing chemical equations, identifying limiting reactants, and calculating product yields.

Understanding these principles is essential for chemists and engineers involved in chemical manufacturing, energy production, and environmental management. Accurate calculations enable the optimization of reactions, assessment of reactant requirements, and prediction of product quantities, all of which are vital for efficient and sustainable chemical processes.

Frequently Asked Questions (FAQs)

1. Why is oxygen the limiting reactant in this reaction?

Because the amount of oxygen available (2.5 moles) is less than the amount needed to react with all the hydrogen (which requires 3.0 moles), oxygen limits the amount of water produced.

2. Can all the hydrogen be used up if oxygen is limiting?

No, since oxygen is limiting, only a portion of hydrogen (specifically, 5.0 moles worth) reacts with oxygen, leaving some hydrogen unreacted.

3. How does this calculation change if more oxygen were available?

If more oxygen were available, hydrogen would become the limiting reactant, and the amount of water produced would be based on the 6.0 moles of hydrogen, resulting in 6.0 moles of water.

4. What is the significance of balancing chemical equations?

Balancing ensures the law of conservation of mass is obeyed, allowing accurate stoichiometric calculations for reactants and products.

5. How can these calculations be applied in real-world scenarios?

They help in designing chemical reactors, optimizing fuel cell operations, estimating reactant requirements, and understanding environmental impacts of chemical processes.

By mastering the concepts illustrated here, students and professionals can confidently approach various chemical reactions, ensuring precise and efficient chemical synthesis and energy conversion.

Frequently Asked Questions

What is the balanced chemical equation for the reaction between hydrogen gas and oxygen gas?

The balanced chemical equation is $2\ H_2 + O_2 \rightarrow 2\ H_2O$.

How many moles of water are produced when 6.0 moles of hydrogen gas react completely?

According to the balanced equation, 2 moles of H_2 produce 2 moles of H_2O, so 6.0 moles of H_2 produce 6.0 moles of H_2O.

Given 2.5 moles of oxygen gas, how many moles of water can be produced?

Since 1 mole of O_2 produces 2 moles of H_2O, 2.5 moles of O_2 can produce 5.0 moles of H_2O, provided hydrogen is in excess.

Is hydrogen or oxygen the limiting reactant in this reaction?

Hydrogen is in excess (6.0 mol) compared to the amount needed for 2.5 mol of oxygen. Therefore, oxygen is the limiting reactant.

How many moles of water are produced based on the limiting reactant?

Since oxygen is limiting and 2.5 mol of O_2 produce 5.0 mol of H_2O, the maximum water produced is 5.0 moles.

What is the total amount of water produced when 6.0 mol of hydrogen reacts with 2.5 mol of oxygen?

The limiting reactant is oxygen, so 5.0 moles of water are produced in this reaction.

Can all 6.0 moles of hydrogen react with 2.5 moles of oxygen?

No, only 5.0 moles of hydrogen can react with 2.5 moles of oxygen to produce water; the remaining hydrogen is in excess.

What is the molar ratio of hydrogen to water in the reaction?

The molar ratio of hydrogen to water is 1:1, as 2 H_2 produce 2 H_2O, so 1 H_2 produces 1 H_2O.

Additional Resources

How Many Moles Of Water Are Produced When 6.0 Moles Of Hydrogen Gas React With 2.5 Moles Of Oxygen Gas?

Understanding the quantitative aspects of chemical reactions is fundamental in chemistry, especially when predicting product formation based on reactant quantities. One common and classic example involves the reaction between hydrogen gas and oxygen gas to produce water. This reaction not only exemplifies fundamental stoichiometric principles but also has practical importance across various industries, including energy, manufacturing, and environmental science.

In this comprehensive review, we delve into the detailed process of determining how many moles of water are generated when specific amounts of hydrogen and oxygen gases react. We will explore the reaction's balanced chemical equation, stoichiometric ratios, the concept of limiting reactants, and the step-by-step calculations involved in deriving the final answer.

Understanding the Reaction: Hydrogen and Oxygen to Water

The chemical reaction between hydrogen gas (H_2) and oxygen gas (O_2) to produce water (H_2O) is a classic example of a combustion or synthesis process. The fundamental reaction can be written as:

$$\text{2 H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O} (g)$$

Key aspects of this reaction:

- It is a combination reaction where two reactants combine to form a single product.
- It is exothermic, releasing energy in the form of heat and light.
- The reaction is highly efficient and nearly complete under the right conditions.

Balancing the Chemical Equation

Before embarking on any stoichiometric calculations, it is crucial to have a correctly balanced chemical equation, ensuring the conservation of mass. The provided unbalanced equation is:

$$\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$$

Balancing it, we get:

$$2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O}$$

This balanced equation indicates:

- 2 moles of hydrogen gas react with 1 mole of oxygen gas to produce 2 moles of water.

Stoichiometric Ratios and Implications

From the balanced equation, the molar ratios are:

- Hydrogen to water: 2 mol H_2 → 2 mol H_2O (a 1:1 ratio)
- Oxygen to water: 1 mol O_2 → 2 mol H_2O (a 1:2 ratio)

This means:

- For every 2 moles of hydrogen, 1 mole of oxygen is needed to produce 2 moles of water.
- The amount of water produced depends on the limiting reactant—either hydrogen or oxygen.

Given Quantities and the Problem Setup

The problem specifies:

- 6.0 moles of hydrogen gas (H_2)
- 2.5 moles of oxygen gas (O_2)

Our goal:

- To determine how many moles of water (H_2O) are produced when these reactants react completely under ideal conditions.

Determining the Limiting Reactant

The concept of limiting reactant is central in stoichiometry. It is the reactant that gets completely consumed first, thus limiting the amount of product formed.

Step 1: Calculate the theoretical amount of water produced from each reactant

- From hydrogen (H_2):

Using the molar ratio from the balanced equation:

\[
\text{2 mol H}_2 \rightarrow \text{2 mol H}_2\text{O}
\]

Therefore,

\[
6.0 \text{ mol H}_2 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} = 6.0 \text{ mol H}_2\text{O}
\]

- From oxygen (O_2):

Using the molar ratio:

\[
\text{1 mol O}_2 \rightarrow \text{2 mol H}_2\text{O}
\]

So,

\[
2.5 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} = 5.0 \text{ mol H}_2\text{O}
\]

Step 2: Identify the limiting reactant

- Hydrogen can produce 6.0 mol of water if it is in excess.
- Oxygen can produce 5.0 mol of water.

Since oxygen produces fewer moles of water, oxygen is the limiting reactant in this scenario.

Conclusion:

- The maximum amount of water that can be produced with the given quantities is limited by oxygen, which yields 5.0 mol of water.

Calculating the Actual Water Yield

Because oxygen is the limiting reactant, the total amount of water formed will correspond to the amount of oxygen available:

Final answer:

\[
\boxed{
\textbf{5.0 moles of water}
}
\]

This is the theoretical maximum amount of water produced when 6.0 moles of hydrogen react with 2.5 moles of oxygen under ideal, complete reaction conditions.

Additional Considerations and Real-World Factors

While the above calculation provides a clear theoretical framework, real-world reactions often involve additional factors:

- Reaction completeness: In ideal conditions, the reaction goes to completion. In practice, reaction conditions such as temperature, pressure, and catalysts influence the extent of reaction.
- Reaction efficiency: Not all reactants might convert fully into products; some may remain unreacted due to kinetic or thermodynamic constraints.
- Side reactions: In complex systems, other reactions might occur, affecting the overall yield.

However, in the context of this problem, assuming ideal conditions simplifies the calculation and provides a maximum theoretical yield.

Summary of Key Steps

To succinctly recap the process:

1. Write and balance the chemical equation: \(2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \)
2. Identify molar ratios from the balanced equation: 2 mol H_2 : 1 mol O_2 : 2 mol H_2O
3. Calculate the amount of water each reactant can produce:
- Hydrogen: 6.0 mol → 6.0 mol water
- Oxygen: 2.5 mol → 5.0 mol water
4. Determine the limiting reactant: Oxygen limits water formation to 5.0 mol.
5. Conclude the maximum product formed: 5.0 mol of water.

Final Thoughts and Practical Implications

Understanding how to determine the amount of water produced from given quantities of hydrogen and oxygen is fundamental in chemical manufacturing, energy production (such as fuel cells), and environmental management. This exercise demonstrates the importance of stoichiometry in predicting yields, optimizing reactions, and designing processes.

In real applications, additional factors like reaction kinetics, purity of reactants, and process conditions influence actual yields. Nevertheless, theoretical calculations like this serve as vital benchmarks for planning and efficiency assessments.

In conclusion, when 6.0 moles of hydrogen gas react with 2.5 moles of oxygen gas, a maximum of 5.0 moles of water can be produced under ideal conditions, with oxygen being the limiting reactant.

Note: Always double-check your balanced equations and calculations, and consider real-world variables when applying theoretical results to practical scenarios.

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