How Many Liters Of Water Can Be Made From 34 Grams Of Oxygen Gas And 6.0 Grams Of Hydrogen Gas At Stp?

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This question explores a fundamental chemical reaction—the synthesis of water from its elemental gases—and involves understanding stoichiometry, molar calculations, and gas laws under Standard Temperature and Pressure (STP). Determining how much water can be produced from specific quantities of oxygen and hydrogen requires a step-by-step approach, involving the balanced chemical equation, molar conversions, and volume calculations. In this article, we will delve into the details of these calculations, explain the underlying principles, and provide insights into the chemistry involved.

Understanding the Chemical Reaction: The Formation of Water

The Reaction Equation

The formation of water from hydrogen and oxygen gases is represented by a simple, well-known chemical reaction:

$$ 2H_2(g) + O_2(g) \rightarrow 2H_2O(g) $$

This balanced equation indicates that 2 molecules of hydrogen gas react with 1 molecule of oxygen gas to produce 2 molecules of water vapor. When considering molar ratios, it simplifies to:

- 2 moles of H_2 react with 1 mole of O_2 to produce 2 moles of H_2O.

Implications of the Reaction

From the molar ratio, we see that:

- 1 mole of O_2 is needed to react with 2 moles of H_2.
- The amount of water produced is directly proportional to the limiting reactant—either oxygen or hydrogen—depending on their initial quantities.

Understanding which reactant is limiting is critical in calculating the maximum amount of water that can be formed from given masses of oxygen and hydrogen.

Step 1: Calculating the Moles of Reactants

The Molar Masses of Oxygen and Hydrogen

To convert grams to moles, we use molar masses:

- Oxygen gas (O_2): approximately 32.00 g/mol
- Hydrogen gas (H_2): approximately 2.02 g/mol

Calculating Moles of Oxygen

Given 34 grams of oxygen gas:

$$ \text{Moles of O}_2 = \frac{\text{Mass of O}_2}{\text{Molar mass of O}_2} = \frac{34\, \text{g}}{32.00\, \text{g/mol}} \approx 1.0625\, \text{mol} $$

Calculating Moles of Hydrogen

Given 6.0 grams of hydrogen gas:

$$ \text{Moles of H}_2 = \frac{6.0\, \text{g}}{2.02\, \text{g/mol}} \approx 2.97\, \text{mol} $$

Step 2: Determining the Limiting Reactant

Using the molar ratios from the balanced equation:

- For every 1 mol of O_2, 2 mol of H_2 are required.
- For the available oxygen, the required hydrogen is:

$$ 1.0625\, \text{mol O}_2 \times 2 = 2.125\, \text{mol H}_2 $$

- Since we have approximately 2.97 mol of H_2, which exceeds the 2.125 mol needed to react with all the available oxygen, hydrogen is in excess.

Conclusion:
Oxygen is the limiting reactant because there isn't enough hydrogen to react with all the oxygen. The maximum amount of water produced will be based on the available oxygen.

Step 3: Calculating Moles of Water Produced

From the balanced reaction, 1 mol of O_2 produces 2 mol of H_2O:

$$\text{Moles of H}_2O = \text{Moles of O}_2 \times 2 = 1.0625 \times 2 = 2.125\, \text{mol}$$

This is the maximum number of moles of water vapor that can be formed from the given reactants.

Step 4: Converting Moles of Water to Volume at STP

Understanding Gas Volumes at STP

Standard Temperature and Pressure (STP) is defined as 0°C (273.15 K) and 1 atm pressure. Under these conditions, 1 mol of an ideal gas occupies approximately 22.4 liters.

Calculating the Volume of Water Vapor

Since water is produced as vapor, and assuming complete conversion and ideal gas behavior:

$$\text{Volume of H}_2O = \text{Moles of H}_2O \times 22.4\, \text{L/mol}$$

$$\text{Volume of H}_2O = 2.125\, \text{mol} \times 22.4\, \text{L/mol} \approx 47.6\, \text{L}$$

Note: If the water condenses into liquid form, the volume calculation would differ. However, since the question specifies volume at STP, we consider water vapor.

Additional Considerations: Liquid Water Yield

The calculation above provides the volume of water vapor produced, which is relevant in gaseous synthesis or high-temperature reactions. To find the amount of liquid water produced, we convert moles of water to grams and then to volume, considering the density of water.

Converting Moles of Water to Grams

Molar mass of water (H_2O):

$$2 \times 1.008\, \text{g} + 16.00\, \text{g} = 18.016\, \text{g}$$

Total mass:

$$ 2.125\, \text{mol} \times 18.016\, \text{g/mol} \approx 38.3\, \text{g} $$

Converting Grams to Volume of Liquid Water

Density of liquid water:

$$ \approx 1\, \text{g/mL} $$

Therefore, volume:

$$ \frac{38.3\, \text{g}}{1\, \text{g/mL}} = 38.3\, \text{mL} $$

This indicates that about 38.3 milliliters of liquid water can be produced from the reaction under ideal conditions.

Summary of Key Results

- Maximum moles of water vapor produced: approximately 2.125 mol
- Volume of water vapor at STP: approximately 47.6 liters
- Mass of liquid water produced: approximately 38.3 grams
- Volume of liquid water: approximately 38.3 milliliters

Understanding the Limitations and Real-World Factors

While the calculations above assume ideal conditions, real-world factors can influence the actual yields:

- Reaction efficiency: Not all reactants may react completely due to kinetic barriers.
- Impurities: Presence of impurities can reduce the purity of the products.
- Gas behavior deviations: Real gases deviate slightly from ideal behavior, especially at high pressures.
- Condensation: Water produced at high temperature remains as vapor unless cooled.

Conclusion

In conclusion, from 34 grams of oxygen gas and 6.0 grams of hydrogen gas at STP, the maximum amount of water vapor that can be produced is approximately 47.6 liters. This calculation is based on the limiting reactant—oxygen—and assumes ideal gas behavior and 100% reaction efficiency. Understanding these principles provides insight into stoichiometry, chemical reactions, and gas laws, which are foundational in chemistry and industrial applications involving water synthesis, fuel production, and atmospheric science.

Remember: Always consider the context of your reaction conditions and the nature of the products (gas or liquid) when interpreting such calculations.

Frequently Asked Questions

How do you determine the amount of water produced from given masses of oxygen and hydrogen gases at STP?

You use the balanced chemical equation for water formation ($2H_2 + O_2 \rightarrow 2H_2O$), convert the given masses to moles, determine the limiting reagent, and then calculate the moles of water produced to find the volume at STP.

What is the molar mass of oxygen and hydrogen gases used in the calculation?

The molar mass of O_2 is approximately 32 g/mol, and that of H_2 is about 2 g/mol.

How many moles of oxygen and hydrogen are present in 34 grams of O_2 and 6.0 grams of H_2?

Moles of O_2 = 34 g ÷ 32 g/mol = 1.0625 mol; Moles of H_2 = 6.0 g ÷ 2 g/mol = 3 mol.

Which reactant is limiting in this reaction, and how does it affect water production?

Oxygen is the limiting reagent because it has fewer moles relative to the stoichiometric ratio, so the amount of water produced is determined by the oxygen quantity.

What volume of water (in liters) can be produced from these reactants at STP?

From 1 mol of O_2, 2 mol of H_2O are formed, so 1.0625 mol of O_2 produces 2.125 mol of H_2O. At STP, 1 mol of water vapor occupies 22.4 liters, so the total volume is approximately 2.125 mol × 22.4 L/mol ≈ 47.6 liters.

Additional Resources

Water Production from Gases: An In-Depth Analysis of Reactant Quantities and Resulting Water Volume

When exploring the fascinating world of chemical reactions, especially those involving the synthesis of water from its constituent gases—oxygen and hydrogen—it's essential to understand the underlying principles that dictate how much water can be produced given specific quantities of reactants. In this detailed examination, we'll analyze the question: "How many liters of water can be made from 34 grams of oxygen gas and 6.0 grams of hydrogen gas at standard temperature and pressure (STP)?"

This query isn't just a theoretical exercise—it's fundamental to fields ranging from industrial chemistry to environmental science, where efficient resource utilization matters immensely. Let's delve into the chemistry step-by-step, breaking down the calculations and providing a comprehensive understanding of the process.

Understanding the Chemical Reaction: The Formation of Water

Before quantifying the amount of water produced, it's crucial to grasp the chemical basis of the reaction involved.

The Reaction Equation

The reaction between hydrogen and oxygen gases to produce water is represented by the balanced chemical equation:

$$ 2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l) $$

This equation indicates that:

- 2 molecules of hydrogen gas react with 1 molecule of oxygen gas
- To produce 2 molecules of liquid water

Key Point: The reaction occurs at the molecular level, and the stoichiometry (mole ratio) is central to calculating the yield based on given masses.

Step 1: Convert Given Masses to Moles

The first step is to convert the given masses of oxygen and hydrogen gases into moles, as the reaction's stoichiometry is expressed in moles.

Calculating Moles of Oxygen Gas (O_2)

Given:
- Mass of oxygen gas = 34 grams
- Molar mass of O_2 = 32.00 g/mol

Calculation:

$$
\text{Moles of O}_2 = \frac{\text{Mass of O}_2}{\text{Molar mass of O}_2} = \frac{34\,g}{32.00\,g/mol} = 1.0625\, \text{mol}
$$

Calculating Moles of Hydrogen Gas (H_2)

Given:
- Mass of hydrogen gas = 6.0 grams
- Molar mass of H_2 = 2.016 g/mol

Calculation:

$$
\text{Moles of H}_2 = \frac{6.0\,g}{2.016\,g/mol} \approx 2.976\, \text{mol}
$$

Step 2: Determine the Limiting Reactant

The limiting reactant is the reactant that will be entirely consumed first, thus limiting the amount of water produced.

From the balanced reaction:
- 2 mol H_2 reacts with 1 mol O_2

Compare the mole ratios:

Reactant	Moles Available	Moles Required per Reaction	Actual ratio
H_2	2.976 mol	2 mol per 1 reaction unit	2.976 / 2 ≈ 1.488

O_2	1.0625 mol	1 mol per 1 reaction unit	1.0625 / 1 ≈ 1.0625

Since:
- H_2 requires 1.488 reaction units to be fully consumed
- O_2 requires 1.0625 reaction units

Conclusion: O_2 is the limiting reactant because it will run out first (fewer reaction units are possible with the available amount).

Step 3: Calculate the Theoretical Water Yield

Knowing that oxygen is limiting, we proceed based on the amount of oxygen to determine the maximum amount of water produced.

From the balanced equation:
- 1 mol O_2 produces 2 mol H_2O

So:

$$
\text{Moles of H}_2O = 2 \times \text{Moles of O}_2 = 2 \times 1.0625 = 2.125\, \text{mol}
$$

Key Point: The maximum moles of water that can be formed, based on the limiting reactant, is 2.125 mol.

Step 4: Convert Moles of Water to Volume at STP

Standard Temperature and Pressure (STP) conditions imply:
- 1 mol of gas occupies approximately 22.4 liters

However, note that water in this reaction is formed as a liquid, not a gas. But since the question asks about liters of water, it's important to clarify whether we're referring to:

- The volume of liquid water produced (which is a mass-based measure)
- The volume that the water vapor would occupy if vaporized at STP

Assumption: Given the context, the question most likely refers to the volume of liquid water produced, which is typically measured in liters by mass conversion, knowing the density of water.

Step 5: Convert Moles of Water to Mass and Volume of Liquid Water

Given:
- Moles of water = 2.125 mol
- Molar mass of water = 18.015 g/mol

Calculations:
$$
\text{Mass of water} = 2.125\, \text{mol} \times 18.015\, g/mol \approx 38.33\, g
$$

Since the density of water is approximately 1 g/mL, the volume in milliliters is:
$$
\text{Volume} = \frac{\text{Mass}}{\text{Density}} = \frac{38.33\, g}{1\, g/mL} = 38.33\, mL
$$

Converted to liters:
$$
38.33\, mL = 0.03833\, L
$$

Result: Approximately 0.0383 liters of liquid water can be produced under ideal conditions.

Additional Considerations: Gaseous Water Vapor Volume

If we interpret the question as concerning the volume of water vapor (gas form) at STP, then:

- 1 mol of water vapor occupies 22.4 liters at STP
- Moles of water vapor = 2.125 mol

Thus:

$$
\text{Volume of water vapor} = 2.125\, \text{mol} \times 22.4\, \text{L/mol}
$$

\approx 47.6\, \text{L}
\]

This means that, if the produced water were in vapor form at STP, it would occupy approximately 47.6 liters.

Summary of Results and Practical Implications

Aspect	Quantity	Explanation
Moles of O_2 used	1.0625 mol	Based on initial mass of 34 g
Moles of H_2 used	2.976 mol	Based on initial mass of 6.0 g
Limiting reactant	Oxygen (O_2)	Reacts completely first
Moles of water produced	2.125 mol	From 1 mol O_2 reacting
Mass of water produced	~38.33 g	2.125 mol × 18.015 g/mol
Volume of liquid water	~0.0383 liters	Assuming 1 g/mL density
Volume of water vapor at STP	~47.6 liters	Vapor phase, ideal gas law

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Final Thoughts: Practical Applications and Limitations

While the theoretical calculations provide a clear picture of the maximum possible water yield, real-world scenarios involve additional factors:

- Efficiency of the reaction: In practice, reactions are not 100% efficient due to energy losses, side reactions, and incomplete mixing.
- Purity of reactants: Impurities can affect the amount of water produced.
- Phase considerations: The actual phase of water (liquid or vapor) depends on temperature and pressure conditions.

Nonetheless, this detailed calculation offers valuable insight into the fundamental chemistry governing water synthesis from hydrogen and oxygen gases. Whether for industrial water production, fuel cell technology, or scientific research, understanding these principles is essential for optimizing resource use and process efficiency.

In conclusion, from 34 grams of oxygen and 6.0 grams of hydrogen, at STP, the maximum amount of liquid water producible is approximately 0.0383 liters, with the potential to generate around 47.6 liters of water vapor if vaporized under identical conditions. This exemplifies the direct relationship between reactant quantities and product volume, reinforcing the importance of stoichiometry in chemical processes.

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