

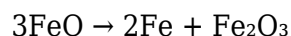
How Many Grams Of Fe Are Produced From 76.3 G Of FeO Given The Reaction $3\text{FeO} \rightarrow 2\text{Fe} + \text{Fe}_2\text{O}_3$

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When working with chemical reactions, one of the most common questions students and professionals alike ask is: "How much of a product can be obtained from a given amount of reactant?" In this case, we are exploring the reaction involving iron(II) oxide (FeO) and determining how many grams of iron (Fe) can be produced from 76.3 grams of FeO, based on the reaction $3\text{FeO} \rightarrow 2\text{Fe} + \text{Fe}_2\text{O}_3$ (or similar). This article provides a detailed, step-by-step explanation of how to calculate the amount of iron produced, including the relevant chemical concepts, molar calculations, and practical tips to approach such problems efficiently.

Understanding the Chemical Reaction

Before delving into calculations, it's essential to understand the chemical reaction involved. The reaction given is:



This reaction describes the reduction of iron(II) oxide (FeO) to elemental iron (Fe), with the formation of iron(III) oxide (Fe₂O₃) as a byproduct. The reaction is a simplified representation often used in metallurgy and chemistry to understand how metallic iron is extracted from its oxide.

Key Points:

- FeO is the reactant.
- Fe (iron metal) is the desired product.
- The stoichiometry of the reaction indicates the molar ratios: 3 mol of FeO produce 2 mol of Fe.

Note: For the purpose of this calculation, we focus on the molar ratio between FeO and Fe, which is 3:2.

Step-by-Step Calculation Process

Calculating the amount of Fe produced involves several steps:

1. Identify the known quantities.
2. Convert the given mass of FeO to moles.
3. Use the molar ratio from the balanced equation to find moles of Fe produced.
4. Convert moles of Fe to grams.

Let's explore each step in detail.

1. Identifying the Known Quantities

- Mass of FeO given: 76.3 grams
- Molar masses:
- FeO: Approximately 71.85 g/mol
- Fe: Approximately 55.85 g/mol

Sources for molar masses:

Element	Atomic Mass (g/mol)
Fe	55.85
O	16.00

Calculating molar mass of FeO:

$$\text{Molar mass of FeO} = 55.85 + 16.00 = 71.85 \text{ g/mol}$$

2. Converting Mass of FeO to Moles

Using the formula:

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Calculate moles of FeO:

$$\text{Moles of FeO} = \frac{76.3 \text{ g}}{71.85 \text{ g/mol}} \approx 1.061 \text{ mol}$$

3. Using the Molar Ratio to Find Moles of Fe

From the balanced reaction:



The molar ratio of FeO to Fe is 3:2, which means:

$$\text{Moles of Fe} = \text{Moles of FeO} \times \frac{2}{3}$$

Calculate moles of Fe:

$$\text{Moles of Fe} = 1.061 \times \frac{2}{3} \approx 0.707 \text{ mol}$$

4. Converting Moles of Fe to Grams

Using the molar mass of Fe:

$$\text{Mass of Fe} = \text{Moles of Fe} \times \text{Molar mass of Fe}$$

Calculate:

$$\text{Mass of Fe} = 0.707 \times 55.85 \text{ g/mol} \approx 39.50 \text{ g}$$

Result:

From 76.3 grams of FeO, approximately 39.50 grams of elemental Fe can be produced.

Additional Considerations and Practical Insights

While the above calculations provide an ideal theoretical yield, real-world scenarios often involve factors such as reaction efficiency, purity of reactants, and experimental conditions that can affect the actual yield.

Factors Affecting Actual Yield:

- Reaction Efficiency: Not all FeO may react fully; impurities or incomplete reactions can reduce yield.
- Purity of Reactants: Impurities in FeO can alter the amount of Fe produced.

- Process Conditions: Temperature, pressure, and catalysts can influence the reaction's completeness.

Calculating Percent Yield

If an actual experiment yields a different amount of Fe, the percent yield can be calculated:

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

Common Mistakes to Avoid

- Using incorrect molar masses: Always verify the atomic weights.
- Mixing units: Keep units consistent; convert grams to moles before calculations.
- Ignoring reaction stoichiometry: Remember the molar ratios from the balanced chemical equation.
- Forgetting significant figures: Report your final answer with appropriate precision based on the given data.

Summary of the Calculation

Step	Calculation	Result
1	Moles of FeO	1.061 mol
2	Moles of Fe (using ratio)	0.707 mol
3	Grams of Fe	39.50 g

Final Answer: Approximately 39.50 grams of iron (Fe) are produced from 76.3 grams of FeO via the reaction $3\text{FeO} \rightarrow 2\text{Fe} + \text{Fe}_2\text{O}_3$.

Conclusion

Understanding how to convert between masses, moles, and product yields is fundamental in chemistry. This step-by-step approach demonstrates how stoichiometry, balanced chemical equations, and molar mass calculations work together to determine the amount of product formed from a given reactant. By mastering these techniques, students and professionals can accurately predict yields, optimize reactions, and understand the quantitative aspects of chemical processes.

Whether you're preparing for exams or working in a laboratory, these skills are invaluable for analyzing chemical reactions and ensuring successful outcomes in your experiments and industrial processes.

Frequently Asked Questions

How do you determine the amount of Fe produced from 76.3 g of FeO in the reaction $3\text{FeO} \rightarrow \text{Fe} + \text{Fe}_2\text{O}_3$?

First, calculate moles of FeO: $76.3 \text{ g} \div 71.85 \text{ g/mol} =$ approximately 1.062 mol. Using the reaction ratio (3 mol FeO : 1 mol Fe), determine moles of Fe: $1.062 \text{ mol} \div 3 \approx 0.354 \text{ mol}$. Convert to grams: $0.354 \text{ mol} \times 55.85 \text{ g/mol} \approx 19.78 \text{ g}$ of Fe.

What is the molar mass of FeO used in calculations?

The molar mass of FeO is approximately 71.85 g/mol, calculated as 55.85 g/mol for Fe plus 16.00 g/mol for O.

In the reaction $3\text{FeO} \rightarrow \text{Fe} + \text{Fe}_2\text{O}_3$, what is the stoichiometric ratio between FeO and Fe?

The ratio is 3 mol FeO to 1 mol Fe, meaning three parts FeO produce one part Fe.

How many grams of Fe are produced from 76.3 grams of FeO according to the reaction?

Approximately 19.78 grams of Fe are produced from 76.3 grams of FeO.

What assumptions are made when calculating the mass of Fe produced from FeO?

Assumptions include complete reaction with no side reactions or losses, and pure reactants without impurities.

Can this calculation be applied to other quantities of FeO? How?

Yes, by adjusting the initial mass of FeO, calculating moles, using the stoichiometric ratio, and converting back to grams.

What is the significance of the molar ratio in this type of stoichiometric calculation?

It allows you to relate the amount of reactant to product, enabling conversion between different substances based on the balanced chemical equation.

How do you ensure accuracy when performing these

stoichiometric conversions?

Use precise molar masses, carefully perform each step, and double-check calculations to minimize errors.

What practical applications does this type of stoichiometry calculation have?

It is essential in industries for calculating yields, designing chemical processes, and ensuring proper quantities of reactants are used.

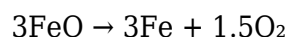
Additional Resources

How Many Grams Of Fe Are Produced From 76.3 G Of FeO Given The Reaction 3FeO

Understanding chemical reactions and their stoichiometric relationships is fundamental in chemistry, especially when it comes to calculating the amount of products formed from given reactants. Today, we explore a specific scenario: determining how many grams of iron (Fe) can be produced from 76.3 grams of iron(II) oxide (FeO) based on the reaction involving three moles of FeO. This inquiry not only highlights the importance of balanced equations but also demonstrates the practical application of molar conversions in real-world chemistry.

The Chemical Reaction Involving FeO and Iron

Before delving into calculations, it is essential to understand the chemical reaction in question. The reaction involving iron(II) oxide (FeO) typically pertains to its reduction to metallic iron (Fe). A common reduction process involves carbon or other reducing agents, but for simplicity, we consider a generalized reaction:



This equation indicates that three moles of FeO decompose or reduce to produce three moles of elemental iron and oxygen gas. The reaction is balanced, meaning the number of atoms for each element is the same on both sides, ensuring the conservation of mass. From this, we can see that the molar ratio of FeO to Fe is 3:3, or simply 1:1, implying a straightforward conversion between moles of FeO and Fe.

Step 1: Understanding the Molar Masses

To perform any stoichiometric calculation, the first step involves knowing the molar masses of the reactants and products:

- Iron(II) oxide (FeO):
- Iron (Fe): approximately 55.85 g/mol
- Oxygen (O): approximately 16.00 g/mol

- Molar mass of FeO = 55.85 + 16.00 = 71.85 g/mol

- Iron (Fe):

- Molar mass = 55.85 g/mol

These values are critical for converting between grams and moles, which are the fundamental units in stoichiometry.

Step 2: Convert the Given Mass of FeO to Moles

The first concrete step involves converting 76.3 grams of FeO into moles:

Number of moles of FeO =

$$\frac{\text{Mass of FeO}}{\text{Molar mass of FeO}} = \frac{76.3 \text{ g}}{71.85 \text{ g/mol}}$$

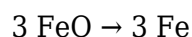
Calculating:

$$\frac{76.3}{71.85} \approx 1.061 \text{ mol}$$

This means that 76.3 grams of FeO corresponds to approximately 1.061 moles.

Step 3: Apply the Stoichiometric Relationship

Given the balanced reaction:



The molar ratio of FeO to Fe is 1:1, which simplifies the calculation. Since 3 moles of FeO produce 3 moles of Fe, then 1 mole of FeO produces 1 mole of Fe.

Therefore, the moles of Fe produced from 1.061 moles of FeO:

$$\text{Moles of Fe} = 1.061 \text{ mol}$$

Step 4: Convert Moles of Fe to Grams

To find the mass of iron produced:

$$\text{Mass of Fe} = \text{Moles of Fe} \times \text{Molar mass of Fe}$$

$$= 1.061 \text{ mol} \times 55.85 \text{ g/mol} \approx 59.26 \text{ g}$$

Thus, approximately 59.26 grams of iron can be produced from 76.3 grams of FeO under ideal conditions.

Practical Considerations and Real-World Implications

While the stoichiometric calculation provides a theoretical maximum, real-world reactions often encounter inefficiencies. Factors such as incomplete reactions, side reactions, impurities, and practical limitations of reaction conditions can reduce the actual yield. Nonetheless, these calculations are vital for industrial processes, laboratory experiments, and resource planning.

Applications:

- Metallurgical Industry: Iron extraction from iron ores relies heavily on stoichiometric calculations to optimize yield and cost.
- Chemical Manufacturing: Understanding the precise amount of reactants needed minimizes waste and environmental impact.
- Educational Purposes: Demonstrates the importance of molar relationships and reaction balancing in chemistry.

Additional Insights: Calculating from Different Starting Masses

Suppose a scenario where a different amount of FeO is available, or a different product yield is desired, similar calculations can be performed:

- Convert the amount of FeO to moles.
- Use the molar ratio from the balanced reaction to find moles of Fe.
- Convert the moles of Fe to grams for the final answer.

This process underscores the universality of stoichiometric principles in chemical computations.

Conclusion

In summary, starting with 76.3 grams of FeO, it is possible to theoretically produce approximately 59.26 grams of Fe based on the balanced chemical reaction $3\text{FeO} \rightarrow 3\text{Fe} + 1.5\text{O}_2$. This calculation illustrates the power of stoichiometry in bridging the gap between measured quantities and the quantities of substances involved in chemical reactions. Whether in research laboratories or industrial settings, such calculations are indispensable tools for chemists and engineers alike,

enabling efficient and informed decision-making.

Understanding these principles not only enhances our grasp of chemical processes but also underscores the significance of precise measurements, balanced equations, and molar relationships in the science of transformation.

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