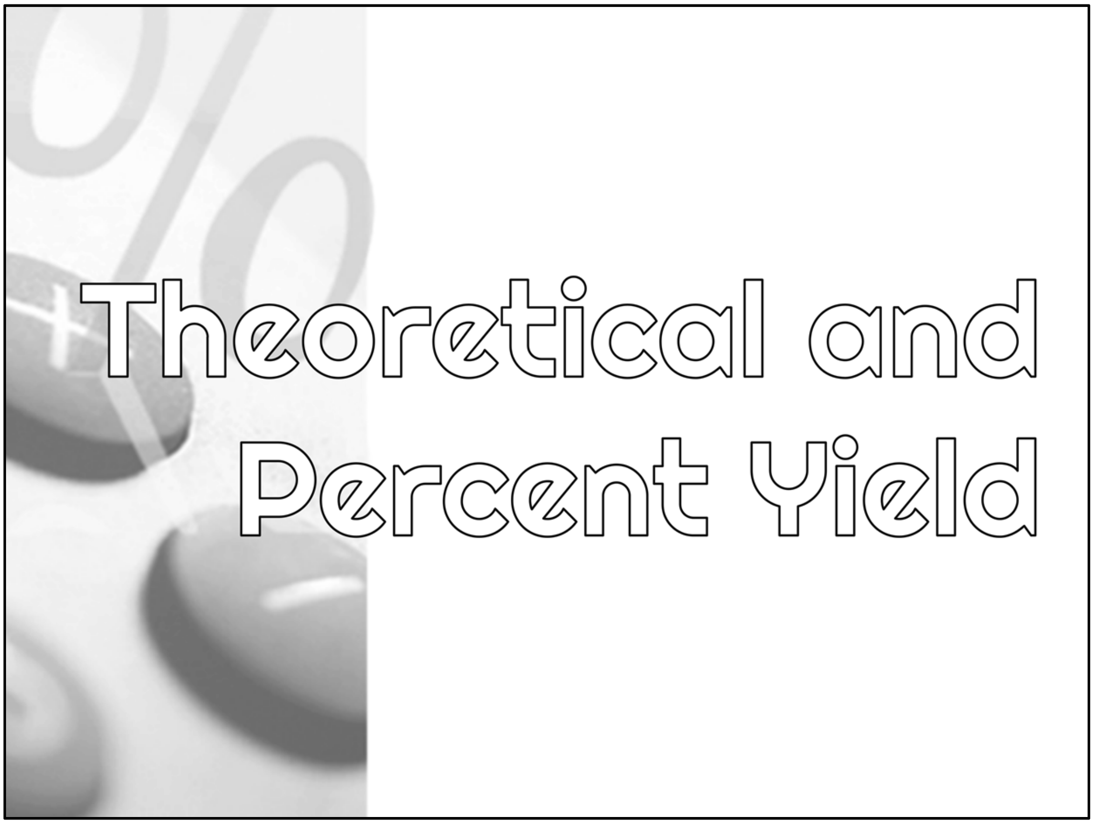
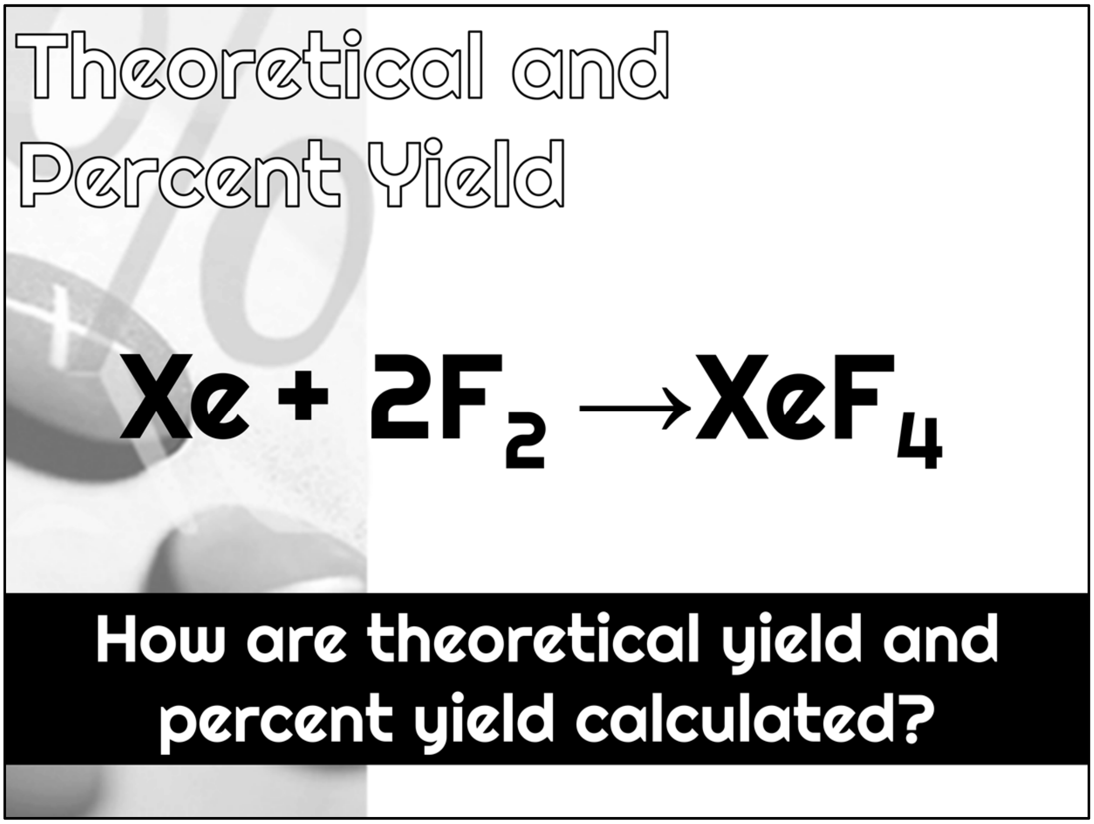
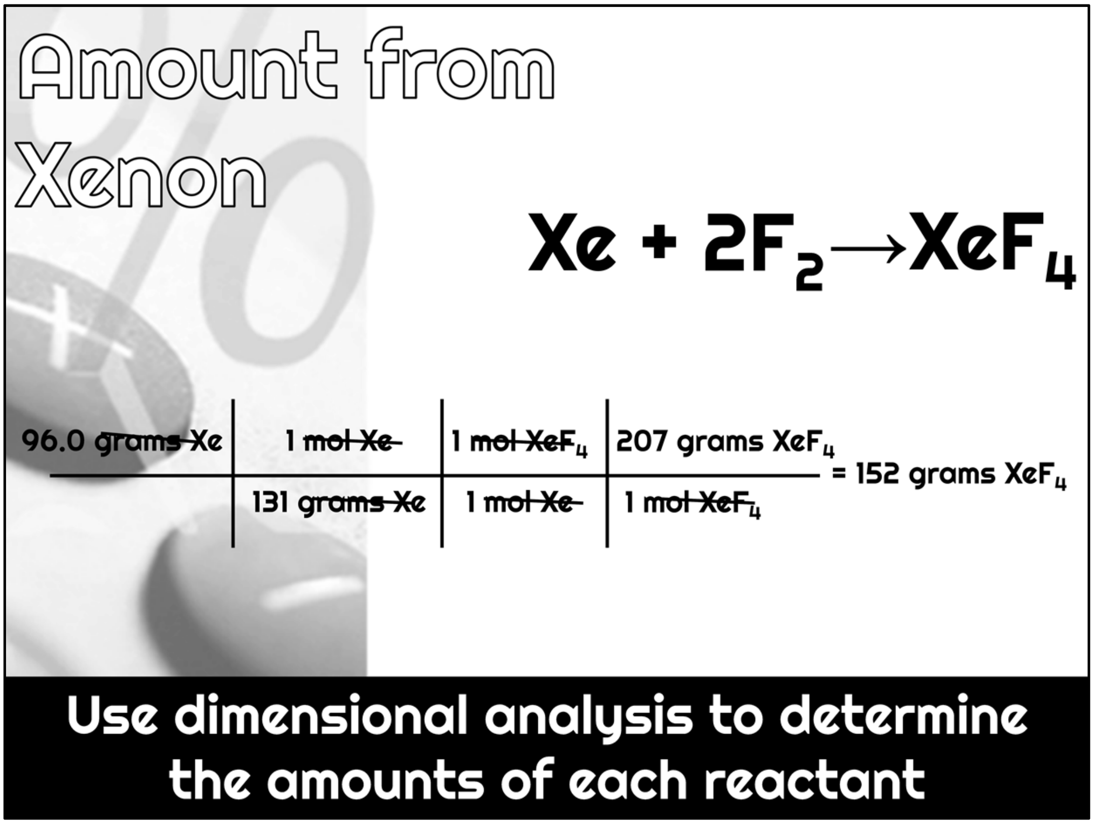
Module 7: Stoichiometry
Topic 5 Content: Percent and Theoretical Yield Presentation Notes



Theoretical and Percent Yield



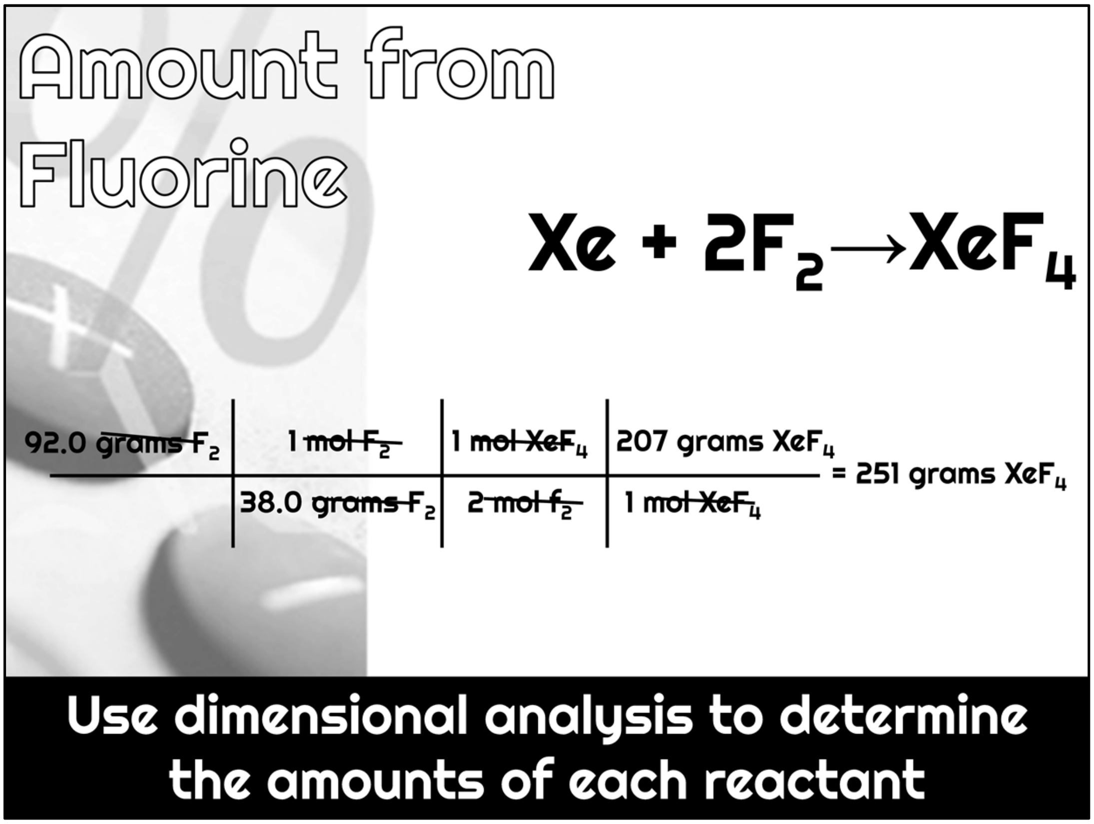
Most of the noble gases do not form stable compounds, but there are some known compounds of xenon. Given the right conditions, xenon combines with fluorine to produce xenon tetrafluoride, or XeF_4. You can see this reaction in the equation shown here. All reactions have a limiting agent and an excess agent, as well as a theoretical yield and a percent yield. How are the theoretical yield and percent yield calculated from analyzing a chemical equation?



Before you determine the percent yield and theoretical yield, you must first find the amount of xenon tetraflouride produced from a given mass of xenon. In this example, imagine that you have 96.0 grams of xenon. In order to calculate the amount, you will need to use the following conversion factors:

- Gram formula mass of xenon
- Mole equivalents from the balanced equation
- Gram formula mass of xenon tetraflouride

To solve this problem, you will need to use dimensional analysis. Multiply 96.0 grams of xenon by 1 mole of xenon. Next, divide by 131 grams of xenon. Then, multiply by 1 mole of xenon tetraflouride over 1 mole of xenon. Finally, multiply by 207 grams of xenon tetraflouride over 1 mole of xenon tetraflouride. After completing the dimensional analysis you will find that 152 grams of xenon tetraflouride are produced. You will need this amount to calculate the theoretical and percent yield.



Now you must find the amount of xenon tetraflouride produced from a given mass of fluorine. In this example, imagine that you have 92.0 grams of fluorine. In order to calculate the percent yield, you will need to use the following conversion factors:

- Gram formula mass of fluorine
- Mole equivalents from the balanced equation
- Gram formula mass of xenon tetraflouride

To solve this problem you will need to use dimensional analysis. Multiply 92.0 grams of fluorine by 1 mole of fluorine. Next, divide by 38.0 grams of fluorine. Then, multiply by 1 mole of xenon tetraflouride over 2 moles of fluorine. Finally, multiply by 207 grams of xenon tetraflouride over 1 mole of xenon tetraflouride. After completing the dimensional analysis you will find that 251 grams of xenon tetraflouride are produced. You will need this amount to calculate the theoretical and percent yield.

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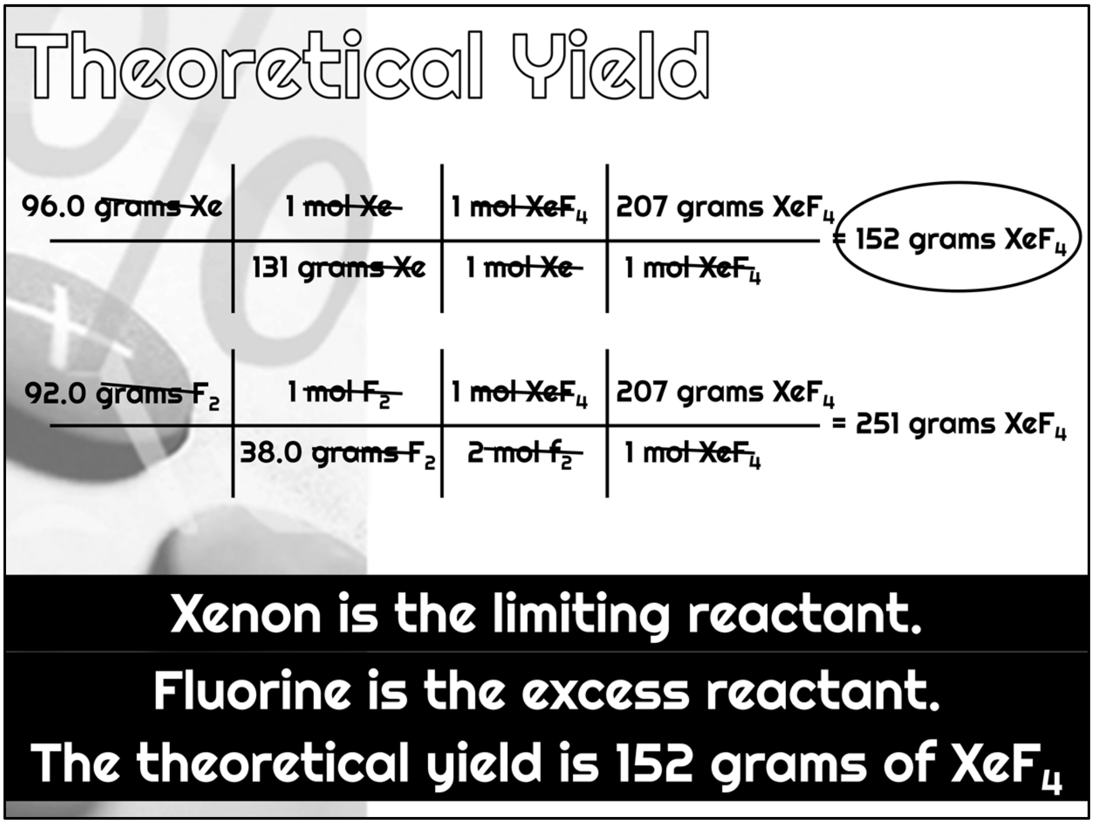
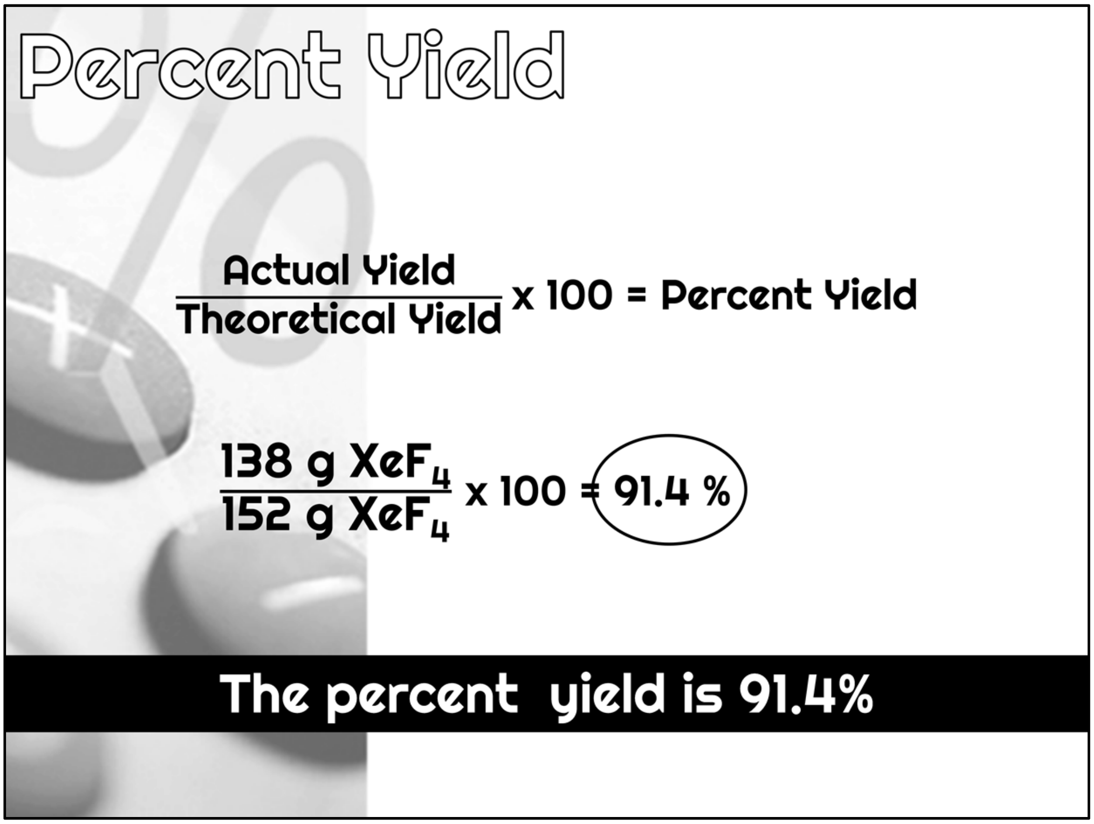


From the calculations, you know that 96.0 grams of xenon produces 152 grams of xenon tetraflouride and 92.0 grams of fluorine produces 251 grams of xenon tetraflouride. Therefore, only 152 grams of xenon tetraflouride can be formed and after that, the reaction stops. Therefore, xenon is the limiting reactant. There is more fluorine than needed, so it is the excess reactant. The theoretical yield is 152 grams of XeF_4.



Suppose that you actually carry out this reaction in the lab and from the given amounts of reactant only 138 grams of XeF_4 is produced. One reason could be some of the reactant is used in some kind of side reaction, or maybe some of it was lost in the process of obtaining and measuring the final product. In any instance, the lab procedure did not yield 100% of what it should have. To calculate the **percent yield**, the percent that was produced is divided by the amount produced by the theoretical yield. This result is multiplied by 100. The calculation shows that the percent yield when 138 grams of XeF_4 is produced is 91.4%.