


Slide 1

**Limiting Reagents**

When you just can't get enough!



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Slide 2


**2 eggs + 1 cup flour → 3 cookies**

If I have a dozen eggs, how many cookies can I make?

NONE – I have no flour!

If I have 6 cups of flour, how many cookies can I make?

NONE – I have no eggs!



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Slide 3

**10 cups of flour + 12 eggs**

2 eggs + 1 cup flour = 3 "cookies"

A. 18  
B. 30  
C. 6  
D. 12  
E. 10



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Slide 4


**2 eggs + 1 cup flour → 3 cookies**

If I have a dozen eggs, and 10 cups of flour, how many cookies can I make?

18 – then I'm out of eggs.

If I have 6 eggs and 2 cups of flour, how many cookies can I make?

Six – then I'm out of flour.



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Slide 5

**CONGRATULATIONS!**

You now understand  
Limiting Reagents!



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Slide 6

**Emphasis on “limiting”**

The key to “limiting reagents” is the limiting.

In most reactions, there is neither an infinite supply of reactants nor a perfect balance between the finite supply.

In most reactions, you run out of 1 reactant before any of the others and the reaction stops.



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Slide 7

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

8.35 g of hydrogen and 15.32 g of oxygen are mixed and allowed to react. Which is the limiting reagent? What is the **theoretical yield** of water?

Grams is good, MOLES! MOLES! MOLES! IS BETTER!



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Slide 8


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

$8.35 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = 4.14 \text{ mol H}_2$

$15.32 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.479 \text{ mol O}_2$

$4.14 \text{ mol H}_2 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} = 4.14 \text{ mol H}_2\text{O}$

$0.479 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} = 0.958 \text{ mol H}_2\text{O}$



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Slide 9

**Apples to Apples  
Oranges to Oranges**

YOU CAN'T LOOK AT THE AMOUNT OF THE REACTANTS!!!!

IT'S NOT THE AMOUNT OF REACTANTS,  
IT'S THE AMOUNT OF PRODUCTS THEY CAN MAKE!!!!!!!!!!!!



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Slide 10

**4 cups flour + 1 can corn = 10 corn fritters**

I have 6 cups of flour and 2 cans of corn...which is the limiting reagent?

I have MORE FLOUR, but I also run out of flour first!

I can't compare flour to corn, I compare corn fritters to corn fritters.



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Slide 11

**O<sub>2</sub> is the Limiting Reagent**

It can make the least amount of product.

After you make 0.958 mol H<sub>2</sub>O, it doesn't matter how much hydrogen you have. You are DONE, because you are out of O<sub>2</sub>!



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Slide 12


**Theoretical Yield**

$$0.958 \text{ mol H}_2\text{O} \times \frac{18.01 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 17.2 \text{ g H}_2\text{O}$$

Notice, this is another case of The Dance!

g → mol → mol → g

Only we do The Dance twice and take the smallest quantity!



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Slide 13

**Clicker Question**

$N_2 + H_2 \rightarrow NH_3$

Ammonia can be made from the gas phase reaction of nitrogen and hydrogen. 16.8 g of nitrogen and 2.47 g of hydrogen were mixed and reacted for 72 hours at 850 °C. How many grams of ammonia would you expect to produce?

- A. 19.27 g
- B. 20.4 g
- C. 10.2 g
- D. 13.9 g
- E. 20.8 g



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Slide 14

**Clicker Question**

$N_2 + H_2 \rightarrow NH_3$

Ammonia can be made from the gas phase reaction of nitrogen and hydrogen. 16.8 g of nitrogen and 2.47 g of hydrogen were mixed and reacted for 72 hours at 850 °C. How many grams of ammonia would you expect to produce?

- A. 19.27 g (conservation of mass doesn't apply)
- B. 20.4 g (wrong limiting reagent)
- C. 10.2 g (unbalanced equation)
- D. 13.9 g (correct)
- E. 20.8 g (unbalanced equation and wrong limiting reagent)



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Slide 15

$N_2 + H_2 \rightarrow NH_3$

Ammonia can be made from the gas phase reaction of nitrogen and hydrogen. 36.06 g of nitrogen and 12.53 g of hydrogen were mixed and reacted for 72 hours at 850 °C. How many grams of ammonia would you expect to produce? If 22.86 g  $NH_3$  were collected, what is the % yield of the reaction?



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Slide 16


**$N_2 + 3 H_2 \rightarrow 2 NH_3$**

$36.06 \text{ g } N_2 \cdot \frac{1 \text{ mol } N_2}{28.014 \text{ g } N_2} \cdot \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 2.57 \text{ mol } NH_3$

$12.53 \text{ g } H_2 \cdot \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \cdot \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 4.14 \text{ mol } NH_3$

$N_2$  is the limiting reagent

$2.57 \text{ mol } NH_3 \cdot \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 43.77 \text{ g } NH_3$



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Slide 17

**Yield!**

Theoretical Yield = amount you expect to get

Actual Yield = amount you actually get

% yield =  $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$



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Slide 18

**Yield!**

$\frac{22.86 \text{ g } NH_3 \text{ actual}}{43.77 \text{ g } NH_3 \text{ expected}} \times 100 = 52.23 \% \text{ yield}$

UNITS! UNITS! UNITS!

$52.23 \% \text{ yield} = \frac{52.23 \text{ g actual}}{100 \text{ g expected}}$



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Slide 19

**$N_2 + 3 H_2 \rightarrow 2 NH_3$**

Ammonia can be made from the gas phase reaction of nitrogen and hydrogen. 25.00 g of nitrogen and 8.00 g of hydrogen were mixed and reacted for 72 hours at 850 °C. How many grams of ammonia would you expect to produce if the reaction proceeds with a 76% yield?

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Slide 20

**$N_2 + 3 H_2 \rightarrow 2 NH_3$**

$25.00 \text{ g } N_2 \cdot \frac{1 \text{ mol } N_2}{28.014 \text{ g } N_2} \cdot \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 1.785 \text{ mol } NH_3$

$8.00 \text{ g } H_2 \cdot \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \cdot \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 2.65 \text{ mol } NH_3$

$N_2$  is the limiting reagent

$1.785 \text{ mol } NH_3 \cdot \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 30.40 \text{ g } NH_3$  theoretical

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Slide 21

**Actual Yield?**

30.40 g  $NH_3$  expected  
76% yield

Just another UNITS! UNITS! UNITS! problem:

$30.40 \text{ g } NH_3 \text{ expected} \cdot \frac{76 \text{ g } NH_3 \text{ actual}}{100 \text{ g } NH_3 \text{ expected}} = 23.10 \text{ g } NH_3 \text{ actual}$

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Slide 22

$$\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3$$

Ammonia can be made from the gas phase reaction of nitrogen and hydrogen. If this reaction is known to proceed with 85% yield, how much nitrogen and hydrogen would you need to start with to get 10.00 g of ammonia?



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Slide 23


10.00 g NH<sub>3</sub> actual \*  $\frac{100 \text{ g theoretical}}{85 \text{ g actual}}$  = 11.765 g NH<sub>3</sub>

$11.765 \text{ g NH}_3 * \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} * \frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} * \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 2.089 \text{ g H}_2$

$2.089 \text{ g H}_2 * \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} * \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} * \frac{28.014 \text{ g N}_2}{1 \text{ mol N}_2} = 9.6766 \text{ g N}_2$

Notice  $2.089 \text{ g H}_2 + 9.6766 \text{ g N}_2 = 11.766 \text{ g NH}_3$

Of course, since I only get 10.00 g NH<sub>3</sub>, what happened to the other 1.766 g of stuff?



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Slide 24

Of course, since I only get 10.00 g NH<sub>3</sub>, what happened to the other 1.766 g of stuff?

- A. It is still N<sub>2</sub> and H<sub>2</sub>, it didn't react.
- B. It is hydrazine, N<sub>2</sub>H<sub>2</sub>
- C. Oxygen from the air created H<sub>2</sub>O and N<sub>2</sub>O.
- D. None of the above.



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