Slide 1

Limiting Reagents

When you just can't get enough!

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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10 cups of flour + 12 eggs

2 eggs + 1 cup flour = 3 "cookies"

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

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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CONGRATULATIONS!

You now understand Limiting Reagents!

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

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

$8.35\ g\ H_2 * \frac{1\ mol\ H_2}{2.016\ g\ H_2} = 4.14\ mol\ H_2$

$15.32\ g\ O_2 * \frac{1\ mol\ O_2}{32.00\ g\ O_2} = 0.479\ mol\ O_2$

$4.14\ mol\ H_2 * \frac{2\ mol\ H_2O}{2\ mol\ H_2} = 4.14\ mol\ H_2O$

$0.479\ mol\ O_2 * \frac{2\ mol\ H_2O}{1\ mol\ O_2} = 0.958\ mol\ H_2O$

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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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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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O_2 is the Limiting Reagent

It can make the least amount of product.

After you make 0.958 mol H_2O, it doesn't matter how much hydrogen you have. You are DONE, because you are out of O_2!

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Theoretical Yield

$$0.958 \text{ mol } H_2O \times \frac{18.01 \text{ g } H_2O}{1 \text{ mol } H_2O} = 17.2 \text{ g } H_2O$$

Notice, this is another case of The Dance!

g → mol → mol → g

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

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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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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$N_2 + H_2 \rightarrow NH_3$

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

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

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

N_2 is the limiting reagent

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

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Yield!

Theoretical Yield = amount you expect to get

Actual Yield = amount you actually get

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

Slide 18

Yield!

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

UNITS! UNITS! UNITS!

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

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

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

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

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

N_2 is the limiting reagent

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

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Actual Yield?

30.40 g NH_3 expected
76% yield

Just another UNITS! UNITS! UNITS! problem:

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

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

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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$$10.00\ g\ NH_3\ actual * \frac{100\ g\ theoretical}{85\ g\ actual} = 11.765\ g\ NH_3$$

$$11.765\ g\ NH_3 * \frac{1\ mol\ NH_3}{17.03\ g\ NH_3} * \frac{3\ mol\ H_2}{2\ mol\ NH_3} * \frac{2.016\ g\ H_2}{1\ mol\ H_2} = 2.089\ g\ H_2$$

$$2.089\ g\ H_2 * \frac{1\ mol\ H_2}{2.016\ g\ H_2} * \frac{1\ mol\ N_2}{3\ mol\ H_2} * \frac{28.014\ g\ N_2}{1\ mol\ N_2} = 9.6766\ g\ N_2$$

Notice $2.089\ g\ H_2 + 9.6766\ g\ N_2 = 11.766\ g\ NH_3$

Of course, since I only get 10.00 g NH_3, what happened to the other 1.766 g of stuff?

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Of course, since I only get 10.00 g NH_3, what happened to the other 1.766 g of stuff?

A. It is still N_2 and H_2, it didn't react.
B. It is hydrazine, N_2H_2.
C. Oxygen from the air created H_2O and N_2O.
D. None of the above.