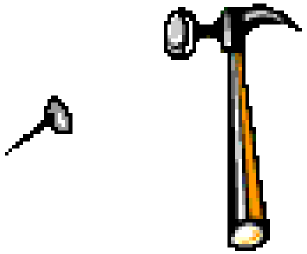
NOTES: 12.3 – Limiting Reagent & Percent Yield

Calculations need to be based on the limiting reactant.

- **<u>Example 1</u>**<u>:</u> Suppose a box contains 87 bolts, 110 washers and 99 nails. How many sets of 1 bolt, 2 washers and 1 nail can you use to create? What is the limiting factor?



55 sets; <u>washers</u> limit the amount

Calculations need to be based on the limiting reactant.

- **Example 2**: What is the maximum mass of sulfur dioxide that can be produced by the reaction of 95.6 g carbon disulfide with 100.0 g oxygen?

  $$CS_2 + O_2 \rightarrow CO_2 + SO_2$$

  Start by balancing the equation…

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- **<u>Example 2</u>**: What is the maximum mass of sulfur dioxide that can be produced by the reaction of 95.6 g carbon disulfide with 100.0 g oxygen?

  $$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$$

  Now solve the problem…

- **<u>Example 2</u>:** What is the maximum mass of sulfur dioxide that can be produced by the reaction of 95.6 g carbon disulfide with 100. g oxygen?

  $CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$

$$\frac{95.6g\ CS_2}{} \times \frac{1mol\ CS_2}{76.2g\ CS_2} \times \frac{2mol\ SO_2}{1mol\ CS_2} \times \frac{64.1g\ SO_2}{1mol\ SO_2} = 160.8g\ SO_2$$

$$\frac{100.0g\ O_2}{} \times \frac{1mol\ O_2}{32.0g\ O_2} \times \frac{2mol\ SO_2}{3mol\ O_2} \times \frac{64.1g\ SO_2}{1mol\ SO_2} = 133.5g\ SO_2$$

Which reactant is LIMITING?

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O_2 limits the amount of SO_2 that can be produced. CS_2 is in excess.

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133.5 g of SO_2 can be produced in this reaction.

Limiting Reagent Example 3:

- **Example 3**: What mass of CO_2 could be formed by the reaction of 8.0 g CH_4 with 48.0 g O_2? What mass of the excess reagent will be left over?

  $CH_4 + O_2 \rightarrow CO_2 + H_2O$

  Start by balancing the equation…

- **<u>Example 3</u>**: What mass of CO_2 could be formed by the reaction of 8.0 g CH_4 with 48.0 g O_2? What mass of the excess reagent will be left over?

  $$\mathbf{CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O}$$

  Now determine which reactant is the LIMITING REAGENT…

- **<u>Example 3</u>**: What mass of CO_2 could be formed by the reaction of 8.0 g CH_4 with 48.0 g O_2? What mass of the excess reagent will be left over?

$$\mathbf{CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O}$$

$$\frac{8.0g\ CH_4}{} \times \frac{1mol\ CH_4}{16.0g\ CH_4} \times \frac{1mol\ CO_2}{1mol\ CH_4} \times \frac{44.0g\ CO_2}{1mol\ CO_2} = 22.0g\ CO_2$$

$$\frac{48.0g\ O_2}{} \times \frac{1mol\ O_2}{32.0g\ O_2} \times \frac{1mol\ CO_2}{2mol\ O_2} \times \frac{44.0g\ CO_2}{1mol\ CO_2} = 33.0g\ CO_2$$

Which reactant is LIMITING?

- **<u>Example 3</u>: What mass of CO_2 could be formed by the reaction of 8.0 g CH_4 with 48.0 g O_2?** What mass of the excess reagent will be left over?

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22.0 g of CO_2 will be produced!

Which reactant is LIMITING?

8.0 g CH_4 is L.R.…& oxygen is in EXCESS

- **<u>Example 3</u>**: What mass of the excess reagent will be left over?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

***calculate how much oxygen will react!*

(use the 8.0 g of CH_4)

- **<u>Example 3</u>**: What mass of the excess reagent will be left over?

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

*****calculate how much oxygen will react! (use the 8.0 g of CH_4)***

$$\frac{8.0g\ CH_4}{} \times \frac{1mol\ CH_4}{16.0g\ CH_4} \times \frac{2mol\ O_2}{1mol\ CH_4} \times \frac{32.0g\ O_2}{1mol\ O_2} = 32.0g\ O_2$$

32.0 g of O_2 will react with 8.0 g of CH_4

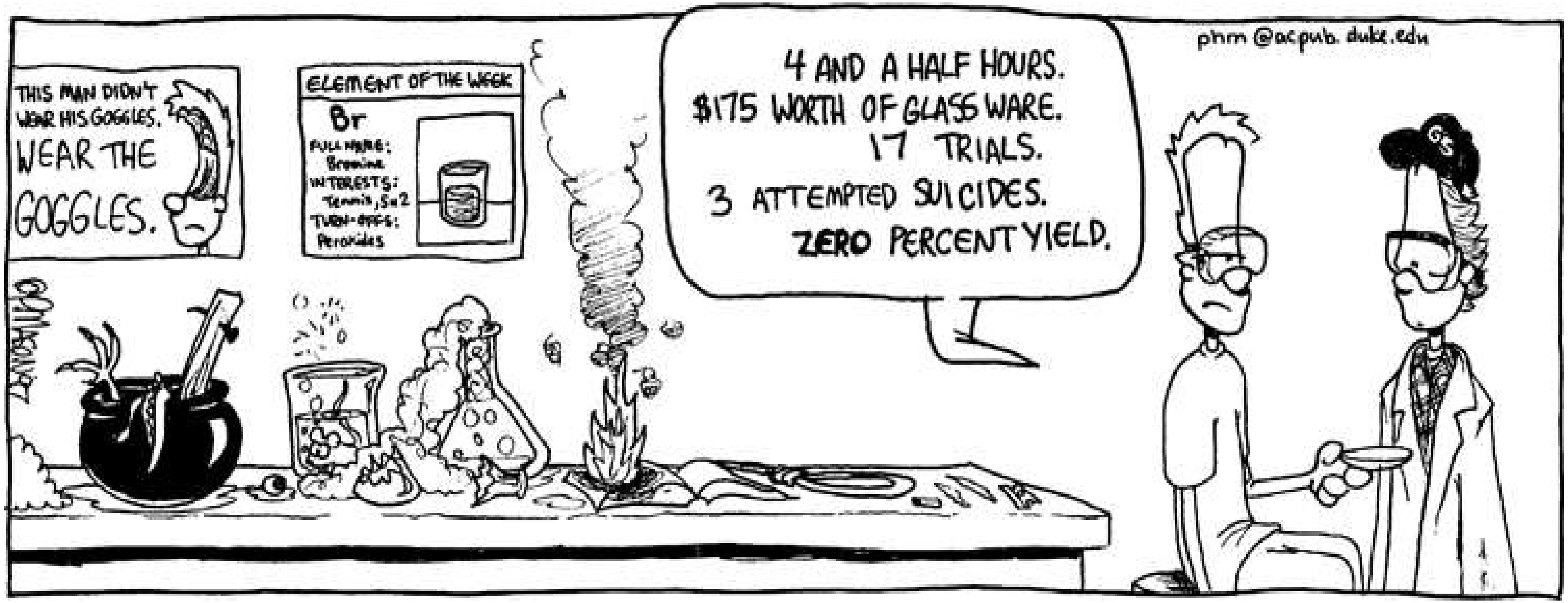
- **<u>Example 3</u>:** **What mass of the excess reagent will be left over?**

We started with 48.0 g of O_2…we calculated that 32.0 g of O_2 would react…SO, the amount left over is:

48.0 g – 32.0 g = $\boxed{\text{16.0 g } O_2 \text{ left over}}$

Percent Yield!

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$



Many chemical reactions do not go to completion (reactants are not completely converted to products).

<u>Percent Yield:</u> indicates what percentage of a desired product is obtained.

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

- So far, the masses we have calculated from chemical equations were based on the assumption that each reaction occurred 100%.

- The **THEORETICAL YIELD** is the amount calculated from the balance equation.

- The **ACTUAL YIELD** is the amount "actually" obtained in an experiment.

- Look back at Example 2. We found that 133.5 g of SO_2 could be formed from the reactants.

- Let's say that in an experiment, you collected 130.0 g of SO_2. **What is your percent yield?**

$$\% \, Yield = \frac{130.0g}{133.5g} \times 100 = 97.4\%$$

Percent Yield Example #1:

Calcium carbonate is decomposed by heating.

$$CaCO_{3\,(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$$

A) What is the **theoretical yield** of CaO when 24.8 g of $CaCO_3$ decompose?

B) When this is done in the lab, 13.1 g of CaO is produced. **What is the percent yield of this reaction?**

Percent Yield Example #1:

$CaCO_{3\,(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$

A) What is the **theoretical yield** of CaO when 24.8 g of $CaCO_3$ decompose?

$$\frac{24.8\text{g } CaCO_3}{} \times \frac{1\text{mol } CaCO_3}{100.1\text{g } CaCO_3} \times \frac{1\text{mol } CaO}{1\text{mol } CaCO_3} \times \frac{56.1\text{g } CaO}{1\text{mol } CaO}$$

Percent Yield Example #1:

$CaCO_{3\,(s)} \rightarrow CaO_{\,(s)} + CO_{2\,(g)}$

A) What is the **theoretical yield** of CaO when 24.8 g of $CaCO_3$ decompose?

$$\frac{24.8\text{g } CaCO_3}{} \times \frac{1\text{mol } CaCO_3}{100.1\text{g } CaCO_3} \times \frac{1\text{mol } CaO}{1\text{mol } CaCO_3} \times \frac{56.1\text{g } CaO}{1\text{mol } CaO}$$

= 13.9 g CaO

Percent Yield Example #1:

B) When this is done in the lab, 13.1 g of CaO is produced. **What is the percent yield of this reaction?**

$$\%\,Yield = \frac{13.1g}{13.9g} \times 100$$

Percent Yield Example #1:

B) When this is done in the lab, 13.1 g of CaO is produced. **What is the percent yield of this reaction?**

$$\% \, Yield = \frac{13.1g}{13.9g} \times 100$$ **= 94.2%**

Percent Yield Example #2:

What is the **percent yield** if 4.65 g of copper is produced when 1.87 g of aluminum reacts with an excess of copper(II) sulfate?

$$\mathbf{2Al_{(s)} + 3CuSO_{4\,(aq)} \rightarrow Al_2(SO_4)_{3\,aq)} + 3Cu_{(s)}}$$

Percent Yield Example #2:

1st find the "theoretical yield":

$2Al_{(s)} + 3CuSO_{4\,(aq)} \rightarrow Al_2(SO_4)_{3\,aq)} + 3Cu_{(s)}$

$$\frac{1.87\text{g Al}}{} \; x \; \frac{1\,\text{mol Al}}{27.0\text{g Al}} \; x \; \frac{3\,\text{mol Cu}}{2\,\text{mol Al}} \; x \; \frac{63.5\,\text{g Cu}}{1\,\text{mol Cu}}$$

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= 6.60 g Cu

Percent Yield Example #2:

Now find % yield:

$$\%\,Yield = \frac{4.65g}{6.60g} \times 100$$

Percent Yield Example #2:

Now find % yield:

$$\% \, Yield = \frac{4.65g}{6.60g} \times 100$$

= 70.5%

<u>Example #3a</u>: Aluminum chloride, $AlCl_3$, can be made by the reaction of aluminum metal with chlorine gas, Cl_2, according to the following equation: $\mathbf{2Al + 3Cl_2 \rightarrow 2AlCl_3}$

What is the limiting reagent if 20.0 g of Al and 30.0 g of Cl_2 are used, and how much $AlCl_3$ can theoretically form?

Example #3a:

2Al + 3Cl$_2$ → 2AlCl$_3$

What is the limiting reagent if 20.0 g of Al and 30.0 g of Cl_2 are used, and how much $AlCl_3$ can theoretically form?

$$\frac{20.0g\ Al}{} \times \frac{1mol\ Al}{27.0g\ Al} \times \frac{2mol\ AlCl_3}{2mol\ Al} \times \frac{133.5g\ AlCl_3}{1mol\ AlCl_3} = 98.9g\ AlCl_3$$

$$\frac{30.0g\ Cl_2}{} \times \frac{1mol\ Cl_2}{71.0g\ Cl_2} \times \frac{2mol\ AlCl_3}{3mol\ Cl_2} \times \frac{133.5g\ AlCl_3}{1mol\ AlCl_3} = 37.6g\ AlCl_3$$

Example #3a:

2Al + 3Cl$_2$ ➔ 2AlCl$_3$

What is the limiting reagent if 20.0 g of Al and 30.0 g of Cl_2 are used, and how much $AlCl_3$ can theoretically form?

$$\frac{20.0g\ Al}{} \times \frac{1mol\ Al}{27.0g\ Al} \times \frac{2mol\ AlCl_3}{2mol\ Al} \times \frac{133.5g\ AlCl_3}{1mol\ AlCl_3} = 98.9g\ AlCl_3$$

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37.6 g of AlCl$_3$ will be produced (theoretical yield)!

Which reactant is LIMITING?

30.0 g Cl$_2$ is L.R.

Example #3b: In the previous reaction, when done in the lab, you produce 32.8 grams of $AlCl_3$. **What is your percent yield?**

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

Example #3b: In the previous reaction, when done in the lab, you produce 32.8 grams of $AlCl_3$. **What is your percent yield?**

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

$$\%\ \text{yield} = \frac{32.8\ \text{g}}{37.6\ \text{g}} \times 100$$

$$\%\ \text{yield} = 87.2\%$$