7.5 Percentage Yield

Every summer, hundreds of young Canadians plant trees as part of the reforestation effort in BC. Imagine you and your group planted 5015 Douglas fir saplings on a mountainside on Vancouver Island. Five years later, a group of forestry surveyors checking for survival find that 4655 of these saplings have taken root and are growing. the percentage yield of your group's effort would be 93%.

So far, we have been making the assumption that all reactions always go entirely to completion.

Meaning that **all** of the limiting reactant has been converted into product, leaving only the excess reactant with none of the limiting reactant remaining at all. "in a perfect world"

In real life, this is not always the case. Like the tree planting example, many reactions complete themselves only **partially.**

Such reactions give only a partial percentage yield, <100% less than 100% of the reactant are converted into products.

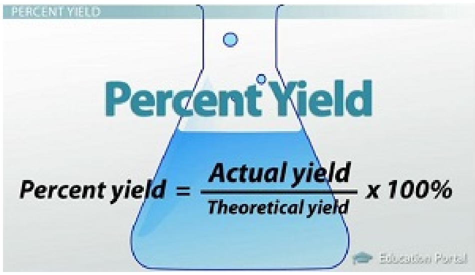
$$\text{Percentage Yield} = \frac{\text{mass of product obtained}}{\text{mass of product expected}} \times 100\%$$

what we actually got

based on stoich. calc. "theoretically"

(the terms %yield is used to express how much of the product is actually obtained)

The amount of product expected (calculated using stoichiometry) is commonly referred to as the theoretical yield.



Actual yield: mass of product collected / obtain

Theoretical yield: how much we "should" get, based on stoich. calculations.

There are 2 major reasons for this reduced yield of products:

1. The reactants may not all react.
2. Some of the products are lost during procedures such as solvent extraction, filtration and crystallization, which are needed to physically separate and purify the products.

NOTE: we can never make more product than expected *(unless we didn't make what we thought we did!)*

The actual yield of the PRODUCTS should be Lower than the *theoretical yield*.

<u>3 Types of % Yield Calculations:</u>

1. *Find the percent yield*, given the mass of reactant used and mass of product formed.
2. *Find the mass of product* formed, given the mass of reactant used and the percentage yield.
3. *Find the mass of the reactant* used, given the mass of product formed and the percentage yield.

Type 1 Problems: *Find the percent yield*, given the mass of reactant used and mass of product formed.

↳ actual yield ↳ use to solve for theoretical yield.

When 15.0 g of $CH_{4(g)}$ is reacted with an excess of $Cl_{2(g)}$ according to the reaction:

$CH_{4(g)} + Cl_{2(g)} \rightarrow CH_3Cl_{(g)} + HCl_{(g)}$ A total of 29.7 g of $CH_3Cl_{(g)}$ is formed.

15.0g 29.7g Actual yield.

What is the percent yield of the reaction?

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

mass → mols CH_4 → mols CH_3Cl → mass CH_3Cl

$$\frac{15.0\,\text{g}\,\cancel{CH_4}}{} \left|\frac{1\,\cancel{\text{mol}\,CH_4}}{16.0\,\text{g}\,\cancel{CH_4}}\right| \frac{1\,\cancel{\text{mol}\,CH_3Cl}}{1\,\cancel{\text{mol}\,CH_4}} \left|\frac{50.5\,\text{g}\,CH_3Cl}{1\,\cancel{\text{mol}\,CH_3Cl}}\right. = \underbrace{47.34\,\text{g}\,CH_3Cl}_{\text{Theoretical yield.}}$$

molar mass molar ratio

$$\%\ \text{yield} = \frac{29.7\,\text{g}}{47.34\,\text{g}} \times 100 = \boxed{62.7\%}$$

Type 2 Problems: *Find the mass of product* formed, given the mass of reactant used and the percentage yield.

find the actual yield

$$\%\ \text{yield} = \frac{\text{Act.}}{\text{Theor.}} \leftarrow ①$$

If the reaction has a 76.0% yield, **what mass of K2CO3(s) is produced** when 1.50 g of KO2(s) is reacted with an excess of CO2(g) according to the reaction:

4 KO2(s) + 2 CO2(g) → 2 K2CO3(s) + 3 O2(g)

1.50g ?g

① Find theoretical yield:

$$\frac{1.50\,\text{g}\,KO_2}{} \left|\frac{1\,\text{mol}\,KO_2}{71.1\,\text{g}\,KO_2}\right| \frac{2\,\text{mol}\,K_2CO_3}{4\,\text{mol}\,KO_2} \left|\frac{138.2\,\text{g}\,K_2CO_3}{1\,\text{mol}\,K_2CO_3}\right. = \underbrace{1.458\,\text{g}\,K_2CO_3}_{\text{Theoretical yield.}}$$

② Find actual yield:

$$\%\ \text{yield} = \frac{\text{actual}}{\text{theor.}} \qquad \therefore \text{actual yield} = \% \cdot \text{theoretical}$$

$$\text{actual yield} = (0.76)(1.458\,\text{g})$$

$$\boxed{= 1.11\,\text{g}\ K_2CO_3}$$

Type 3 Problems: *Find the mass of the reactant* used, given the mass of product formed and the percentage yield

If the reaction has a 58.0% yield, **what mass of CuO(s) is required** to make 10.0 g of Cu(s)
according to the reaction:

mass ← mass

2 NH3(aq) + 3 CuO(s) → N2(g) + 3 Cu(s) + 3H2O(l)

? 10.0g

① Calculate the mass of CuO required (assuming 100% yield)

$$\frac{10.0\,g\,Cu}{} \times \frac{1\,mol\,Cu}{63.5\,g\,Cu} \times \frac{3\,mol\,CuO}{3\,mol\,Cu} \times \frac{79.5\,g\,CuO}{1\,mol\,CuO} = 12.52\,g\,CuO$$

is required to make 10.0g Cu

② We made an assumption, BUT this mass of reactant is really for a 58% yield, not 100% like we assumed. So we need to correct that. To increase the amount of reactant & compensate:

theoretical

$$\rightarrow \frac{\text{actual mass}}{\text{CuO required}} = \frac{12.52\,g\,CuO}{0.580} = \boxed{21.6\,g\,CuO}$$

Summary:

$12.52\,g\,CuO \xrightarrow{100\%} 10.0\,g\,Cu$ "theoretical"

$\boxed{21.6\,g\,CuO} \xrightarrow{58\%} 10.0\,g\,Cu$ "Actual"

Sample Problem — Percentage Yield

One of the primary reasons for reforestation in British Columbia is the great loss of forests to fire each summer. The pithy cente of Douglas fir trees consists of cellulose, a starchy material made up of cross-connected molecules of glucose, $C_6H_{12}O_6$. During a small forest fire, the combustion of 8944 kg of $C_6H_{12}O_6$ molecules results in the formation of 5.75 x 10^6 L of carbon dioxide gas (measured at STP). What is the percentage yield?

What to Think about

1. Once again, in a stoichiometry problem, begin with a balanced equation.
   This is clearly a combustion reaction with a carbohydrate being combusted. The products are carbon dioxide and water vapour. Oxygen is a diatomic molecule.
2. Using standard stoichiometry, convert the given mass of reactant to product.
   This is the volume of product formed if the reaction were to proceed 100% to completion (and the gas were measured under STP conditions). In other words, it is the *theoretical yield*.
   There is a metric conversion required in this problem. "Kilo" means 1000.
3. Calculate the percentage yield as the quotient of the actual yield to the theoretical yield multiplied by 100%.
   The *entire* theoretical yield should be kept in the calculator throughout the calculation to avoid rounding errors.

How to Do It

$$C_6H_{12}O_6(s) + 6\,O_2(g) \rightarrow 6\,CO_2(g) + 6\,H_2O(g)$$

$$8944\,\cancel{kg} \times \frac{1000\,\cancel{g\,C_6H_{12}O_6}}{1\,\cancel{kg\,C_6H_{12}O_6}} \times \frac{1\,\cancel{mol\,C_6H_{12}O_6}}{180.0\,\cancel{g\,C_6H_{12}O_6}} \times \frac{6\,\cancel{mol\,CO_2}}{1\,\cancel{mol\,C_6H_{12}O_6}} \times \frac{22.4\,L\,CO_2}{1\,\cancel{mol\,CO_2}}$$

$$= 6.68 \times 10^6\,L\,CO_2$$

The *theoretical yield* is 6.68 x 10^6 L CO_2.

Thus the percentage yield is:

$$\frac{5.75 \times 10^6\,L}{6.68 \times 10^6\,L} \times 100\% = 86.1\%\ \text{yield}$$

Homework — Assignment 7.5 Percent Yield

1. A chunk of zinc metal reacts with an excess of hydrochloric acid solution. What is the percentage yield if a 7.23 g piece of zinc produces 2.16 L of hydrogen gas at STP? *Begin with a balanced equation.*

   = 87.2%

2. A solution containing 15.2 g of barium bromide is reacted with a solution containing excess sodium phosphate to form 9.5 g of precipitate. What is the percentage yield of the reaction? *Begin with a balanced equation.*

   $BaBr_2 + Na_3PO_4 \rightarrow NaBr + Ba_3(PO_4)_2$ = 93.0%

   precipitate

3. Copper(II) oxide reacts with hydrogen gas to form water and copper metal. From this reaction, 3.6g of copper metal was obtained with a yield of 32.5%. What mass of copper(II) oxide was reacted with the excess hydrogen gas? *Begin with a balanced equation.* = 14.0g CuO

***Hebden page 137 #35-37** ← after quiz

*HW: Review for Quiz! *
- stoich Review Assignment
- % yield Q's #1-3