% Yield & % Purity

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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 100% reactants → 100% products.

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 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, less than 100% of the reactant are converted into products.

actually get (in the experiment)

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

(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 Yeild.

Calculations of Yield

In any chemical reaction, the amount of product you get is called the chemical yield. The yield always differs from the amount you expect to form:

Theoretical yield = expected amount (calculate based on stoichiometry)

Actual yield = amount actually produced

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

Why does the actual yield differ from the theoretical yield?

There are many possible reasons. Some are acceptable and some aren't. Here are a few examples. human error X

- Carelessness or sloppiness. Reactants improperly measured out or lost due to spills, etc.
- Product left behind in the reaction container when collected or transferred.
- Products may be contaminated. The mass of the impurity results in greater mass than expected.
  - Note: Water is frequently a source of error. Important to dry samples thoroughly before weighing.

} experimental errors ✓

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.

3 Types of % Yield Calculations:

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)

Limiting reactant

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.

experimental mass.... this is the "actual yield"

What is the percent yield of the reaction?

$$\% \text{ yield} = \frac{\text{actual mass}}{\text{theoretical mass}} = \frac{29.7\text{g } CH_3Cl}{?\ \boxed{47.3\text{g}}} \times 100 = \boxed{62.8\%}$$

Find the theoretical mass from the limiting reactant:

$$\frac{15.0\text{g } CH_4}{} \left| \frac{1 \text{ mol } CH_4}{16.0\text{g } CH_4} \right| \frac{1 \text{ mol } CH_3Cl}{1 \text{ mol } CH_4} \left| \frac{50.5\text{g } CH_3Cl}{1 \text{ mol } CH_3Cl} \right. = \underline{47.3}\text{ g } CH_3Cl$$

Example 2 A student carried out an experiment in which she completely reacted 45.8 g of potassium carbonate with an excess of hydrochloric acid. As a result, 46.3 g of potassium chloride were produced. What was the percentage yield of potassium chloride?

______________ + ______________ → ______________ + _____ H_2O + _____ CO_2

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

Example 3 When a second student carried out the same reaction as above, 52.7 g of potassium chloride was obtained. Calculate the percentage yield. What most likely explains the result?

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

actual mass

Example 4 If the reaction has a 76.0% yield, what mass of $K_2CO_{3(s)}$ is produced when 1.50 g of $KO_{2(s)}$ is reacted with an excess of $CO_{2(g)}$ according to the reaction:

starting amount (calculate theoretical yield)

L.R.

$$4\ KO_{2(s)} + 2\ CO_{2(g)} \dashrightarrow 2\ K_2CO_{3(s)} + 3\ O_{2(g)}$$

excess

① Calc. Theoretical Yield:

1.50g KO_2	1 mol KO_2	2 mol K_2CO_3	138.2g K_2CO_3
	71.1g KO_2	4 mol KO_2	1 mol K_2CO_3

$= \boxed{1.458\text{g } K_2CO_3}$

Theoretical yield

② if % yield is 76.0 = $\frac{\text{actual}}{\text{theoretical}}$

actual yield = (% yield)(theoretical)

(0.760)(1.458g K_2CO_3)

$= 1.11\text{g } K_2CO_3$

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

Example 5:

= [illegible] g K_2CO_3

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

Example 5:

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:

$$2\ NH_3(aq) + 3\ CuO(s) \rightarrow N_2(g) + 3\ Cu(s) + 3H_2O(l)$$

Sample Purity

When a sample is "impure" it means that only part of its mass is due to the pure substance.

e.g. If sample of NaCl is 75% pure, then **<u>100.0 g of impure sample contains 75.0 g of pure NaCl.</u>**

x % purity *means* x grams of ______ ____________ = 100 grams of ________ _________

From this equivalence statement, we get the conversion factors: $\boxed{\dfrac{x\ g\ \text{pure}}{100.0\ g\ \text{impure}}\ \text{or}\ \dfrac{100.0\ g\ \text{impure}}{x\ g\ \text{pure}}}$

Example 1 A sample of potassium carbonate is 58.5% pure. What mass of sample contains 87.3 g of potassium carbonate? What mass of potassium carbonate is in 295.3 g of the sample?