

1. Percent Purity
2. Percent Yield

Percent Purity

Chemicals don't always exist in pure form.

- The purity of a chemical is indicated as the % purity
- The impure substance contains another substance to make the mass higher than a pure substance
- **ONLY THE PURE SUBSTANCE WILL REACT TO PRODUCE A PURE PRODUCT!**
- Affects reactants - what are you putting into the reaction to react

When doing stoch, pure makes pure

$$\text{Percent Purity} = \frac{\text{mass of pure substance}}{\text{mass of total sample (impure)}} \times 100\%$$

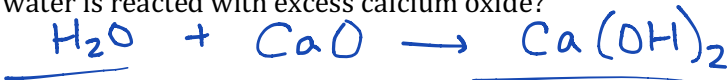
Example 1.

An 85.00 g sample of water is 95% pure. What is the mass of pure water that reacts?

$$\% = \frac{\text{mass pure water}}{85.00 \text{ g H}_2\text{O}} \times 100\% \rightarrow \frac{95\%}{100\%} \times 85.00 \text{ g H}_2\text{O} = \boxed{81 \text{ g pure H}_2\text{O}}$$

2sf

This impure water sample reacts with calcium oxide to produce calcium hydroxide. What mass of calcium hydroxide is produced if the water is reacted with excess calcium oxide?



$$81 \text{ g pure H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol Ca(OH)}_2}{1 \text{ mol H}_2\text{O}} \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 333 \text{ g} \rightarrow \boxed{330 \text{ g pure Ca(OH)}_2}$$

Example 2.

A sample of water is 35% pure. If the mass of pure water is 65g, what is the mass of the total sample?

$$\% \text{ Purity} = \frac{\text{pure}}{\text{sample}} \times 100\%$$

$$35\% = \frac{65 \text{ g H}_2\text{O}}{x} \times 100\%$$

$$x = \frac{65 \times 100}{35} = 185.7$$

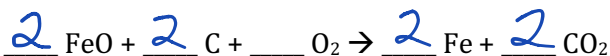
$$= \boxed{190 \text{ g impure/sample H}_2\text{O}}$$

pure produces pure.

← Sample

Example 3.

If 100.0g of FeO produces 12.0g of pure Fe according to the reaction. (Is the 100.0g sample of FeO pure or impure?)



a. How much (mass) FeO was needed to produce Fe?

$$12.0 \text{ g Pure Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{2 \text{ mol FeO}}{2 \text{ mol Fe}} \times \frac{71.85 \text{ g FeO}}{1 \text{ mol FeO}} = 15.4 \text{ g Pure FeO}$$

← pure

* pure produces pure!

b. What is the percentage purity of FeO used?

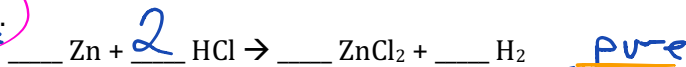
$$\% \text{ Purity} = \frac{\text{Pure}}{\text{Sample}} \times 100\%$$

$$= \frac{15.4 \text{ g Pure FeO}}{100.0 \text{ g Sample FeO}} \times 100\%$$

$$= 15.4 \% \text{ Purity FeO}$$

Example 4.

Zinc metal has a purity of 89.5%.



What mass of this impure zinc is required to produce 975 mL of hydrogen gas at STP?

$$975 \text{ mL H}_2 \times \frac{1 \text{ L H}_2}{1000 \text{ mL H}_2} \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 2.85 \text{ g Pure Zn}$$

$$\% \text{ Purity} = \frac{\text{Pure}}{\text{Sample}} \times 100\%$$

$$89.5 \% = \frac{2.85 \text{ g}}{x} \times 100\%$$

$$x = \frac{2.85 \times 100}{89.5}$$

$$= 3.18 \text{ g sample/impure Zn}$$

Percent Yield

Sometimes 100% of the expected amount of products cannot be attained from a reaction. This can occur because:

1. The reactants may not all react
2. Some of the products are lost due to the experiment procedures

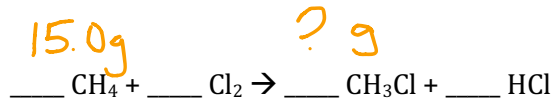
} actual < theoretical

- Affects products - how much product did you actually produce?

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

Example 1.

Given the following reaction:



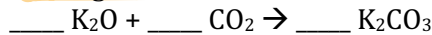
When 15.0g of CH₄ reacts with excess Cl₂, a total of 29.7g of CH₃Cl is formed. What is the percentage yield of the reaction?

$$15.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.05 \text{ g CH}_4} \times \frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol CH}_4} \times \frac{50.49 \text{ g CH}_3\text{Cl}}{1 \text{ mol CH}_3\text{Cl}} = 47.2 \text{ g CH}_3\text{Cl}$$

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% \rightarrow \frac{29.7 \text{ g CH}_3\text{Cl}}{47.2 \text{ g CH}_3\text{Cl}} \times 100\% = \boxed{62.9\% \text{ CH}_3\text{Cl}}$$

Example 2.

What mass of K₂CO₃ is produced when 1.50g of K₂O is reacted according to the reaction,



if the reaction has a 76.0% yield?

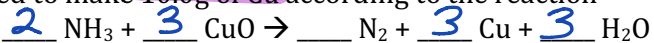
$$1.50 \text{ g K}_2\text{O} \times \frac{1 \text{ mol K}_2\text{O}}{94.20 \text{ g K}_2\text{O}} \times \frac{1 \text{ mol K}_2\text{CO}_3}{1 \text{ mol K}_2\text{O}} \times \frac{138.21 \text{ g K}_2\text{CO}_3}{1 \text{ mol K}_2\text{CO}_3} = 2.20 \text{ g K}_2\text{CO}_3$$

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

$$76.0\% = \frac{x}{2.20 \text{ g K}_2\text{CO}_3} \times 100\% \rightarrow x = \boxed{1.67 \text{ g K}_2\text{CO}_3}$$

Example 3.

What mass of CuO is required to make 10.0g of Cu according to the reaction



if the reaction has a 58.0% yield?

$$10.0 \text{g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{g Cu}} \times \frac{3 \text{ mol CuO}}{3 \text{ mol Cu}} \times \frac{79.55 \text{g CuO}}{1 \text{ mol CuO}} = 12.5 \text{g CuO}$$

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

$$58.0\% = \frac{12.5 \text{g}}{x} \times 100\%$$

$$x = 21.6 \text{g CuO}$$