

Chemistry 11
Stoichiometry V

Name: Key
Date:
Block:

- 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

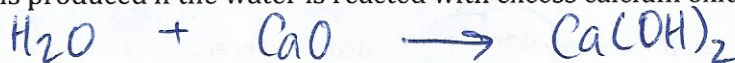
$$\text{Percent Purity} = \frac{\text{mass of pure substance}}{\text{mass of sample}} \times 100\%$$

Example 1.

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

$$95\% = \frac{\text{mass of pure substance}}{85.00\text{g H}_2\text{O}} \times 100\% \Rightarrow \frac{95}{100} \times 85.00 = 81\text{g H}_2\text{O (pure)}$$

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?



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

Example 2.

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

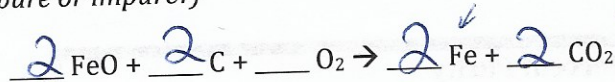
$$\% \text{ purity} = \frac{\text{pure}}{\text{Sample}} \times 100\%$$

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

$$x = \frac{65 \times 100}{35} = 190\text{g H}_2\text{O (impure)} \quad \leftarrow \text{Sample}$$

Example 3.

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



* pure produces pure!!

a. How much (mass) FeO was needed to produce Fe? (15.4 g FeO)

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

$$= 15.4 \text{ g FeO pure}$$

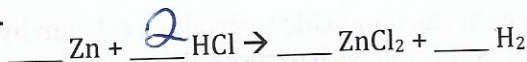
b. What is the percentage purity of FeO used? (15.4% FeO)

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

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

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? (3.18 g Zn)

↳ pure H₂

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

$$= 2.85 \text{ g Zn pure}$$

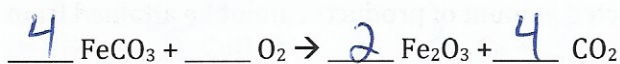
$$\% \text{ purity} = \frac{\text{pure}}{\text{sample}} \times 100\%$$

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

$$x = \frac{2.85 \times 100\%}{89.5\%} = \boxed{3.18 \text{ g Zn sample impure}}$$

Practice:

The roasting of siderite ore, FeCO_3 , produces iron (III) oxide:



a. What is the balanced equation?

b. A 15.0 g FeCO_3 sample is 42.0% pure. What mass of Fe_2O_3 can the sample produce? (4.34 g Fe_2O_3) pure

$$\% \text{ pure} = \frac{\text{pure}}{\text{Sample}} \times 100\% \quad \text{42.0\%} = \frac{x}{15.0g} \cdot 100\% \quad \frac{42.0}{100} \times 15.0 \downarrow = 6.30g \text{ FeCO}_3$$

$$6.30g \text{ FeCO}_3 \times \frac{1 \text{ mol FeCO}_3}{115.86g \text{ FeCO}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCO}_3} \times \frac{159.70g \text{ Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = \boxed{4.34g \text{ Fe}_2\text{O}_3 \text{ pure}}$$

c. A second sample of FeCO_3 , with a mass of 55.0 g is roasted so as to produce 37.0 g of Fe_2O_3 . What is the percentage purity of FeCO_3 ? (97.6% FeCO_3)

$$37.0g \text{ Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.70g \text{ Fe}_2\text{O}_3} \times \frac{4 \text{ mol FeCO}_3}{2 \text{ mol Fe}_2\text{O}_3} \times \frac{115.86g \text{ FeCO}_3}{1 \text{ mol FeCO}_3} = 53.7g \text{ FeCO}_3 \text{ (pure)}$$

$$\% \text{ pure} = \frac{\text{pure}}{\text{Sample}} \times 100\% = \frac{53.7g}{55.0g} \times 100\% = \boxed{97.6\% \text{ FeCO}_3}$$

d. What mass of siderite ore with a purity of 62.8% is needed to make 1.00 kg of Fe_2O_3 ? (2.31 kg FeCO_3)

$$1.00 \text{ kg } \text{Fe}_2\text{O}_3 \times \frac{1000g \text{ Fe}_2\text{O}_3}{1 \text{ kg Fe}_2\text{O}_3} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.70g \text{ Fe}_2\text{O}_3} \times \frac{4 \text{ mol FeCO}_3}{2 \text{ mol Fe}_2\text{O}_3} \times \frac{115.86g \text{ FeCO}_3}{1 \text{ mol FeCO}_3} = 1480g \text{ FeCO}_3 \text{ (pure)}$$

$$\% \text{ pure} = \frac{\text{pure}}{\text{Sample}} \times 100\%$$

$$62.8\% = \frac{1480g}{x} \times 100\%$$

$$x = \frac{1480}{62.8} \times 100\%$$

$$= 2310g \text{ FeCO}_3 \times \frac{1 \text{ kg}}{1000g} = \boxed{2.31 \text{ kg FeCO}_3}$$

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
- 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 15.0 g of CH₄ reacts with excess Cl₂, a total of 29.7 g of CH₃Cl is formed. What is the percentage yield of the reaction? (62.9% CH₃Cl)

if all of this reacted then how much CH₃Cl would have formed?

limiting

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

↳ actual

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

$$= \frac{29.7}{47.2} \times 100\%$$

$$\Rightarrow \boxed{62.9\% \text{ CH}_3\text{Cl}}$$

Example 2.

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



if the reaction has a 76.0% yield? (1.67 g K₂CO₃)

how much actually produced

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

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

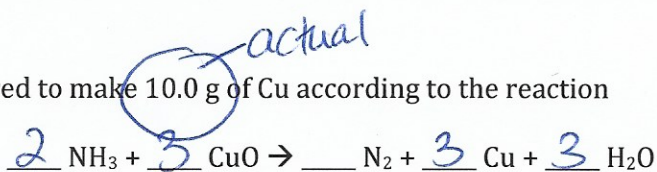
$$76.0\% = \frac{x}{2.20} \times 100\%$$

$$x = \boxed{1.67 \text{ g K}_2\text{CO}_3}$$

actual

Example 3.

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



if the reaction has a 58.0% yield? (21.5 g CuO)

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

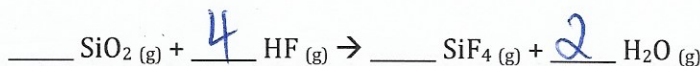
$$\% \text{ yield} = \frac{\text{actual}}{\text{Theoretical}} \times 100\%$$

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

$$x = 21.6 \text{g CuO (Theoretical)}$$

Practice:

The reaction:



produces 2.50 g of H₂O when 12.20 g of SiO₂ is treated with a small excess of HF.

a. If we assume that SiO₂ is 100% pure, what is the percent yield of the reaction? (34.2% H₂O)

$$12.20 \text{g SiO}_2 \times \frac{1 \text{mol SiO}_2}{60.09 \text{g SiO}_2} \times \frac{2 \text{mol H}_2\text{O}}{1 \text{mol SiO}_2} \times \frac{18.02 \text{g H}_2\text{O}}{1 \text{mol H}_2\text{O}} = 7.317 \text{g H}_2\text{O (theoretical)}$$

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% = \frac{2.50}{7.317} \times 100\% = 34.2\% \text{ H}_2\text{O}$$

b. If we assume the reaction has 100% yield, what is the percent purity of SiO₂? (34.2% SiO₂)

$$\begin{aligned} & \text{(pure)} \quad \text{(actual)} \quad 2.50 \text{g H}_2\text{O} \times \frac{1 \text{mol H}_2\text{O}}{18.02 \text{g H}_2\text{O}} \times \frac{1 \text{mol SiO}_2}{2 \text{mol H}_2\text{O}} \times \frac{60.09 \text{g SiO}_2}{1 \text{mol SiO}_2} \\ & = 4.17 \text{g SiO}_2 \text{ (actual)} \end{aligned}$$

(pure)

$$\% \text{ purity} = \frac{\text{pure}}{\text{Sample}} \times 100\%$$

$$= \frac{4.17}{12.20} \times 100\% = 34.2\% \text{ SiO}_2$$