Chemistry 11

Stoichiometry V

Name: Notes
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

When doing stoich, pure makes pure

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

pure produces is pure.

Example 1.
An 85.00 g sample of water is 95% pure. What is the mass of pure water that reacts?

$$\% = \frac{\text{mass pure water}}{85.00\ g\ H_2O} \times 100\% \rightarrow \frac{95\%}{100\%} \times 85.00\ g\ H_2O = \boxed{81\ g\ \text{pure}\ H_2O}$$

2 sf

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?

$$H_2O + CaO \rightarrow Ca(OH)_2$$

$$81\ \cancel{g\ \text{pure}\ H_2O} \times \frac{1\ \cancel{mol\ H_2O}}{18.02\ \cancel{g\ H_2O}} \times \frac{1\ \cancel{mol\ Ca(OH)_2}}{1\ \cancel{mol\ H_2O}} \times \frac{74.10\ g\ Ca(OH)_2}{1\ \cancel{mol\ Ca(OH)_2}} = 333\ g \rightarrow \boxed{330\ g\ \text{pure}\ Ca(OH)_2}$$

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?

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

$$35\% = \frac{65\ g\ H_2O}{x} \times 100\%$$

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

$$= \boxed{190\ g\ \text{impure/sample}\ H_2O}$$

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?)

$$\underline{2}\ FeO + \underline{2}\ C + \underline{\quad}\ O_2 \rightarrow \underline{2}\ Fe + \underline{2}\ CO_2$$

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

*pure produces pure!

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

pure

b. What is the percentage purity of FeO used?

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

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

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

Example 4.
Zinc metal has a purity of 89.5%.

$$\underline{\quad}\ Zn + \underline{2}\ HCl \rightarrow \underline{\quad}\ ZnCl_2 + \underline{\quad}\ H_2$$

pure

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

$$975\ \cancel{mL\ H_2} \times \frac{1\ \cancel{L\ H_2}}{1000\ \cancel{mL\ H_2}} \times \frac{1\ \cancel{mol\ H_2}}{22.4\ \cancel{L\ H_2}} \times \frac{1\ \cancel{mol\ Zn}}{1\ \cancel{mol\ H_2}} \times \frac{65.39\ g\ Zn}{1\ \cancel{mol\ Zn}} = 2.85\ g\ pure\ Zn$$

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

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

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

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

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:

15.0g ? g

____ CH_4 + ____ Cl_2 → ____ CH_3Cl + ____ HCl

When 15.0g of CH_4 reacts with excess Cl_2, a total of 29.7g of CH_3Cl is formed. What is the percentage yield of the reaction?

$$15.0\ \cancel{g\,CH_4} \times \frac{1\ \cancel{mol}\ CH_4}{16.05\ \cancel{g\,CH_4}} \times \frac{1\ \cancel{mol}\ CH_3Cl}{1\ \cancel{mol}\ CH_4} \times \frac{50.49\ g\ CH_3Cl}{1\ \cancel{mol\ CH_3Cl}}$$

$$= 47.2\ g\ CH_3Cl$$

$$\%\ \text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% \rightarrow \frac{29.7\ g\,CH_3Cl}{47.2\ g\,CH_3Cl} \times 100\%$$

$$= \boxed{62.9\%\ CH_3Cl}$$

Example 2.
What mass of K_2CO_3 is produced when 1.50g of K_2O is reacted according to the reaction,

____ K_2O + ____ CO_2 → ____ K_2CO_3

if the reaction has a 76.0% yield?

1.50g ? g

$$1.50\ \cancel{g\,K_2O} \times \frac{1\ \cancel{mol\ K_2O}}{94.20\ \cancel{g\,K_2O}} \times \frac{1\ \cancel{mol}\ K_2CO_3}{1\ \cancel{mol}\ K_2O} \times \frac{138.21\ g\,K_2CO_3}{1\ \cancel{mol}\ K_2CO_3}$$

$$= 2.20\ g\,K_2CO_3$$

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

$$76.0\% = \frac{x}{2.20\ g\,K_2CO_3} \times 100\% \longrightarrow x = \boxed{1.67\ g\ K_2CO_3}$$

Example 3.

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

$\underline{2}\ NH_3 + \underline{3}\ CuO \rightarrow \underline{\quad}\ N_2 + \underline{3}\ Cu + \underline{3}\ H_2O$

if the reaction has a 58.0% yield?

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

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

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

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