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?

$$95\% = \frac{\text{mass of pure substance}}{85.00\text{g } H_2O} \times 100\% \Rightarrow \frac{95}{100} \times 85.00 = 81\text{g } H_2O \text{ (pure)}$$

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?

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

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

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?

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

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

$$x = \frac{65 \times 100}{35} = 190\text{g } H_2O \text{ (impure)}$$

sample

Example 3.

100.0 g of FeO produces 12.0 g of pure Fe according to the following reaction.

pure

(Is the 100.0 g sample of FeO pure or impure?)

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

* pure produces pure!!

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

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

$= 15.4$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%.

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

What mass of this impure zinc is required to produce 975 mL of hydrogen gas at STP? *(3.18 g Zn)*

→ pure H_2

$$975\text{mL } H_2 \times \frac{1\text{L } H_2}{1000\text{mL } H_2} \times \frac{1\text{ mol } H_2}{22.4\text{L } H_2} \times \frac{1\text{ mol Zn}}{1\text{ mol } H_2} \times \frac{65.39\text{g Zn}}{1\text{ mol Zn}}$$

$= 2.85$ 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:

$$\underline{4}\ FeCO_3 + \underline{\quad}\ O_2 \rightarrow \underline{2}\ Fe_2O_3 + \underline{4}\ CO_2$$

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

$\% \text{ pure} = \frac{\text{pure}}{\text{sample}} \times 100\%$ $\quad 42.0\% = \frac{x}{15.0g} \cdot 100\%$ $\quad \frac{42.0}{100} \times 15.0 = 6.30g\ FeCO_3$

$6.30g\ FeCO_3 \times \frac{1mol\ FeCO_3}{115.86g\ FeCO_3} \times \frac{2mol\ Fe_2O_3}{4mol\ FeCO_3} \times \frac{159.70g\ Fe_2O_3}{1mol\ Fe_2O_3} = \boxed{4.34g\ Fe_2O_3}$ 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% $FeCO_3$)*

pure

$37.0g\ Fe_2O_3 \times \frac{1mol\ Fe_2O_3}{159.70g\ Fe_2O_3} \times \frac{4mol\ FeCO_3}{2mol\ Fe_2O_3} \times \frac{115.86g\ FeCO_3}{1mol\ FeCO_3} = 53.7g\ FeCO_3$ (pure)

$\% \text{ pure} = \frac{\text{pure}}{\text{sample}} \times 100\% = \frac{53.7g}{55.0g} \times 100\% = \boxed{97.6\%\ FeCO_3}$

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 $FeCO_3$)*

$FeCO_3$ pure

$1.00kg\ \cancel{FeCO_3}\ Fe_2O_3 \times \frac{1000g\ Fe_2O_3}{1kg\ Fe_2O_3} \times \frac{1mol\ Fe_2O_3}{159.70g\ Fe_2O_3} \times \frac{4mol\ FeCO_3}{2mol\ Fe_2O_3}$

$\times \frac{115.86g\ FeCO_3}{1mol\ FeCO_3} = 1450g\ FeCO_3$ (pure)

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

$62.8\% = \frac{1450g}{x} \times 100\%$

$x = \frac{1450}{62.8} \times 100\%$

$= 2310g\ FeCO_3 \times \frac{1kg}{1000g} = \boxed{2.31kg\ FeCO_3}$

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?

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

Example 1.
Given the following reaction:

limiting

____ CH_4 + ____ Cl_2 → ____ CH_3Cl + ____ HCl

When 15.0 g of CH_4 reacts with excess Cl_2, a total of 29.7 g of CH_3Cl is formed. What is the percentage yield of the reaction? *(62.9% CH_3Cl)*

↳ actual

if all of this reacted then how much CH_3Cl would have formed?

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

$= 47.2g\ CH_3Cl$ (theoretical)

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

$$= \frac{29.7}{47.2} \times 100\% \Rightarrow \boxed{62.9\%\ CH_3Cl}$$

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

____ K_2O + ____ CO_2 → ____ K_2CO_3

how much actually produced

if the reaction has a 76.0% yield? *(1.67 g K_2CO_3)*

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

$= 2.20g\ K_2CO_3$ (theoretical)

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

actual

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

Example 3.

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

actual

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

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

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

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

$58.0\% = \frac{12.5}{x} \times 100\%$ $\boxed{x = 21.6\,g\ CuO\ (Theoretical)}$

Practice:

The reaction:

actual

$\underline{\quad}\ SiO_{2\,(g)} + \underline{4}\ HF_{(g)} \rightarrow \underline{\quad}\ SiF_{4\,(g)} + \underline{2}\ H_2O_{(g)}$

produces 2.50 g of H_2O when 12.20 g of SiO_2 is treated with a small excess of HF.

↳ limiting

a. If we assume that SiO_2 is 100% pure, what is the percent yield of the reaction? *(34.2% H_2O)*

$12.20\,g\ SiO_2 \times \frac{1\,mol\ SiO_2}{60.09\,g\ SiO_2} \times \frac{2\,mol\ H_2O}{1\,mol\ SiO_2} \times \frac{18.02\,g\ H_2O}{1\,mol\ H_2O} = 7.317\,g\ H_2O$ (theoretical)

$\%\ yield = \frac{actual}{theoretical} \times 100\% = \frac{2.50}{7.317} \times 100\%$

$= \boxed{34.2\%\ H_2O}$

b. If we assume the reaction has 100% yield, what is the percent purity of SiO_2? *(34.2% SiO_2)*

(pure) (actual) $2.50\,g\ H_2O \times \frac{1\,mol\ H_2O}{18.02\,g\ H_2O} \times \frac{1\,mol\ SiO_2}{2\,mol\ H_2O} \times \frac{60.09\,g\ SiO_2}{1\,mol\ SiO_2}$

$= 4.17\,g\ SiO_2$ (actual) (pure)

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

$= \frac{4.17}{12.20} \times 100\% =$ ~~[illegible]~~ $\boxed{34.2\%\ SiO_2}$