Percent yield and percentage purity

- Sometimes 100% of the expected amount of product cannot be obtained from a reaction.
- Percentage Yield – amount of product actually produced as a percentage of the expected amount.

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

- Reasons for reduced yield:

  1) reactants may not react
  2) reactants may be less than 100% pure
  3) products are lost

- Another important term is percent purity

$$\% \text{ purity} = \frac{\text{Amount of chemical in sample (mol or mass)}}{\text{Sample (mol or mass)}} \times 100\ \%$$

Examples:

1) When 15.0 g of CH_4 is reacted with an excess of Cl_2 according to the reaction

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ a total mass of 29.7 g of CH_3Cl is formed. What is the percent yield of the reaction?

First find the mass of CH_3Cl expected (assume 100% yield)

% yield =

2) What mass of K_2CO_3 is produced when 1.50 g of KO_2 is reacted with an excess of CO_2 according to the reaction

$4KO_2 + 2CO_2 \rightarrow 2\ K_2CO_3 + 3\ O_2$ If the reaction has 76.0 % yield?

First calculate the mass of K_2CO_3 produced (assuming 100% yield)

Now introduce the fact that the percentage yield is 76% which reduces the mass of K_2CO_3 expected.

3) If 100.0 g of FeO produce 12.9 g of pure Fe according to the reaction

$2\ FeO + 2C + O_2 \dashrightarrow 2\ Fe + 2\ CO_2$

What is the percentage purity of the FeO used?

First calculate the mass of pure FeO which forms 12.9 g of Fe

Percentage purity =

4. The roasting of Siderite ore, $FeCO_3$, produces iron III oxide:

$4\ FeCO_3 + O_2 \dashrightarrow 2\ Fe_2O_3 + 4\ CO_2$

a) A 15.0 g $FeCO_3$ sample is 42% pure. What mass of Fe_2O_3 can the sample produce?

b) A second sample of FeCO3, 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 $FeCO_3$?

c) A 35 g sample of pure $FeCO_3$ produces 22.5 g of Fe_2O_3. What is the percentage yield of the reaction?

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

Answers: a) 4.34 g (b) 97.6% (c) 93.3% (d)2.31×10^{3} g

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$$\% \text{ purity} = \frac{\text{Amount of chemical in sample (mol or mass)}}{\text{Sample (mol or mass)}} \times 100\ \%$$

Examples:

molar mass = (12.01) + 4(1.01) = 16.05 g/mol

1) When 15.0 g of CH_4 is reacted with an excess of Cl_2 according to the reaction

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ a total mass of 29.7 g of CH_3Cl is formed. What is the percent yield of the reaction?

First find the mass of CH_3Cl expected (assume 100% yield)

molar mass (12.01) + 3(1.01) + 35.45 = 50.49 g/mol

$$15.0\text{ g } CH_4 \times \frac{1\text{ mol } CH_4}{16.05\text{ g } CH_4} \times \frac{1\text{ mol } CH_3Cl}{1\text{ mol } CH_4} \times \frac{50.49\text{ g } CH_3Cl}{1\text{ mol } CH_3Cl} = 47.2\text{ g } CH_3Cl$$

$$\% \text{ yield} = \frac{29.7\text{ g}}{47.2\text{ g}} \times 100 = 62.9\ \%$$

(39.10) + 2(16.00) = 71.10 g/mol

2) What mass of K_2CO_3 is produced when 1.50 g of KO_2 is reacted with an excess of CO_2 according to the reaction

$4KO_2 + 2CO_2 \rightarrow 2K_2CO_3 + 3\ O_2$ If the reaction has 76.0 % yield?

2(39.10) + 12.01 + 3(16.00) = 138.21 g/mol

First calculate the mass of K_2CO_3 produced (assuming 100% yield)

$$1.50\text{ g } KO_2 \times \frac{1\text{ mol } KO_2}{71.10\text{ g } KO_2} \times \frac{2\text{ mol } K_2CO_3}{4\text{ mol } KO_2} \times \frac{138.21\text{ g } K_2CO_3}{1\text{ mol } K_2CO_3} = 1.46\text{ g}$$

Now introduce the fact that the percentage yield is 76% which reduces the mass of K_2CO_3 expected.

$$(1.46)(0.76) = 1.11\text{ g}$$

3) If 100.0 g of FeO produce 12.9 g of pure Fe according to the reaction

$2\ FeO + 2C + O_2 \rightarrow 2\ Fe + 2\ CO_2$

What is the percentage purity of the FeO used?

First calculate the mass of pure FeO which forms 12.9 g of Fe

$$12.9\text{g Fe} \times \frac{1\text{ mol Fe}}{55.85\text{g Fe}} \times \frac{2\text{ mol FeO}}{2\text{ mol Fe}} \times \frac{71.85\text{g FeO}}{1\text{ mol FeO}} = 16.60\text{ g}$$

Percentage purity = $\frac{16.60\text{ g}}{100.0\text{ g}} \times 100 = 16.6\ \%$

4. The roasting of Siderite ore, $FeCO_3$, produces iron III oxide:

$4\ FeCO_3 + O_2 \rightarrow 2\ Fe_2O_3 + 4\ CO_2$

a) A 15.0 g $FeCO_3$ sample is 42% pure. What mass of Fe_2O_3 can the sample produce?

$$15.0\text{g} \times 0.42 = 6.3\text{g}$$

$$6.3\text{g } FeCO_3 \times \frac{1\text{ mol } FeCO_3}{115.86\text{g } FeCO_3} \times \frac{2\text{ mol } Fe_2O_3}{4\text{ mol } FeCO_3} \times \frac{159.7\text{ g } Fe_2O_3}{1\text{ mol } Fe_2O_3} = 4.34\text{g}$$

b) A second sample of FeCO3, 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 $FeCO_3$?

$$55.0\text{g } FeCO_3 \times \frac{1\text{ mol } FeCO_3}{115.86\text{g } FeCO_3} \times \frac{2\text{ mol } Fe_2O_3}{4\text{ mol } FeCO_3} \times \frac{159.7\text{g } Fe_2O_3}{1\text{ mol } Fe_2O_3} = 37.91\text{g}$$

$$\frac{37.0\text{ g}}{37.9\text{ g}} \times 100 = 97.6\ \%$$

c) A 35 g sample of pure $FeCO_3$ produces 22.5 g of Fe_2O_3. What is the percentage yield of the reaction?

$$35\text{g } FeCO_3 \times \frac{1\text{ mol } FeCO_3}{115.86\text{g } FeCO_3} \times \frac{2\text{ mol } Fe_2O_3}{4\text{ mol } FeCO_3} \times \frac{159.7\text{g}}{1\text{ mol } Fe_2O_3} = 24.1\text{g}$$

$$22.5\text{g} / 24.1\text{g} = 93.3\%$$

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

$$1.00\text{kg} \times \frac{1000\text{g}}{1\text{ kg}} \times \frac{1\text{ mol } Fe_2O_3}{159.7\text{g } Fe_2O_3} \times \frac{4\text{ mol } FeCO_3}{2\text{ mol } Fe_2O_3} \times \frac{115.86\text{g } FeCO_3}{1\text{ mol } FeCO_3} = 1450.97$$

Answers: a) 4.34 g (b) 97.6% (c) 93.3% (d) 2.31×10^3 g

$$\frac{1450.97\text{g}}{0.628} = 2.31 \times 10^3\text{ g}$$