

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
- $$\text{CH}_4 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{HCl} \quad \text{a total mass of 29.7 g of}$$
- CH_3Cl is formed. What is the percent yield of the reaction?

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

% yield =

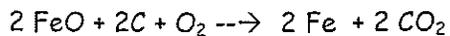
- 2) What mass of K_2CO_3 is produced when 1.50 g of KO_2 is reacted with an excess of CO_2 according to the reaction



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

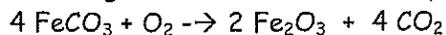


What is the percentage purity of the FeO used?

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:



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 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 ?

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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Examples:

- 1) When 15.0 g of CH_4 is reacted with an excess of Cl_2 according to the reaction $\text{CH}_4 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{HCl}$ 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)

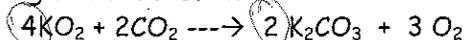
molar mass
 $(12.01) + 3(1.01) = 35.45$
 $= 50.49 \text{ g/mol}$

molar mass = $(12.01) + 4(1.01) = 16.05 \text{ g/mol}$

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

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

- 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



If the reaction has 76.0% yield?

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

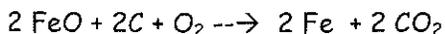
molar mass = $2(39.10) + 12.01 + 3(16.00) = 138.21 \text{ g/mol}$

$$1.50 \text{ g KO}_2 \times \frac{1 \text{ mol KO}_2}{71.10 \text{ g KO}_2} \times \frac{2 \text{ mol K}_2\text{CO}_3}{4 \text{ mol KO}_2} \times \frac{138.21 \text{ g K}_2\text{CO}_3}{1 \text{ mol K}_2\text{CO}_3} = 1.46 \text{ g}$$

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



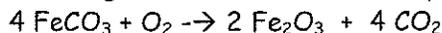
What is the percentage purity of the FeO used?

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

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

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

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



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

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

$$6.3 \text{g FeCO}_3 \times \frac{1 \text{ mol FeCO}_3}{115.86 \text{g FeCO}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCO}_3} \times \frac{159.7 \text{g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 4.34 \text{g}$$

b) 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 ?

$$55.0 \text{g FeCO}_3 \times \frac{1 \text{ mol FeCO}_3}{115.86 \text{g FeCO}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCO}_3} \times \frac{159.7 \text{g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 37.9 \text{g}$$

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

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?

$$35 \text{g FeCO}_3 \times \frac{1 \text{ mol FeCO}_3}{115.86 \text{g FeCO}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCO}_3} \times \frac{159.7 \text{g}}{1 \text{ mol Fe}_2\text{O}_3} = 24.1 \text{g}$$

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

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

$$1.00 \text{kg} \times \frac{1000 \text{g}}{1 \text{kg}} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{g Fe}_2\text{O}_3} \times \frac{4 \text{ mol FeCO}_3}{2 \text{ mol Fe}_2\text{O}_3} \times \frac{115.86 \text{g FeCO}_3}{1 \text{ mol FeCO}_3} = 1450.97 \text{g}$$

$$\text{Answers: a) } 4.34 \text{ g} \quad \text{(b) } 97.6\% \quad \text{(c) } 93.3\% \quad \text{(d) } \frac{1450.97 \text{g}}{0.628} = 2.31 \times 10^3 \text{ g}$$