Concept of limiting reactant

Suppose you have a part-time job in a sandwich shop. One very popular sandwich is always made as follows:

2 slices bread + 3 slices meat + 1 slice cheese → sandwich

Assume that you come to work one day and find the following quantities of ingredients:

8 slices bread
9 slices meat
5 slices cheese

How many sandwiches can you make? What will be left over?

How many sandwiches each component can make?

Bread

$$\boxed{8 \text{ slices bread} \times \frac{1 \text{ sand}}{2 \text{ slices bread}} = 4 \text{ sand}}$$

Meat

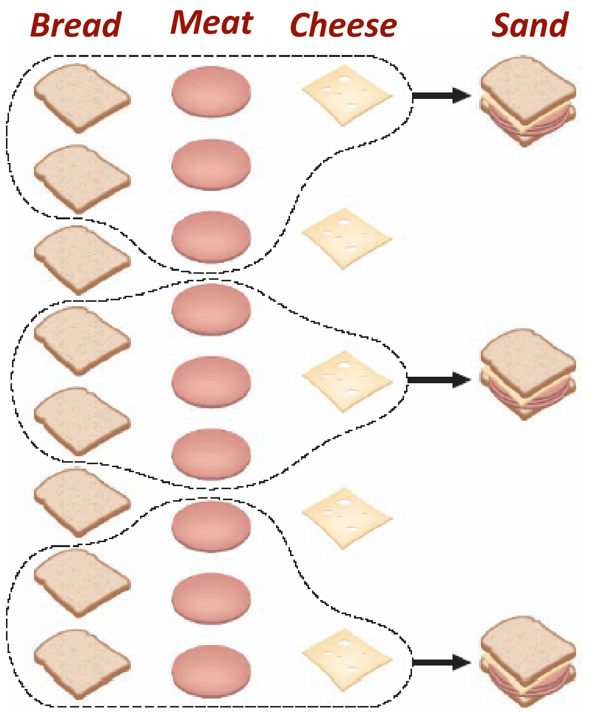
$$\boxed{9 \text{ slices meat} \times \frac{1 \text{ sand}}{3 \text{ slices meat}} = 3 \text{ sand}}$$

Cheese

$$\boxed{5 \text{ slices cheese} \times \frac{1 \text{ sand}}{1 \text{ slices cheese}} = 5 \text{ sand}}$$

- *Overall, how many sandwiches can you make? 3*
- *When you run out of meat, you must stop making sandwiches. The meat is the limiting ingredient.*
- *What do you have left over? 2 slices of bread + 2 pieces of cheese*

The Limiting reactant (meat) is the component that limits the number of products (sandwiches) you can make.



Exercise

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper (II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1 g of NH_3 is reacted with 90.4 g of CuO, which is the limiting reactant? How many grams of N_2 will be formed?

Solution: Find limiting reactant and mass of N_2 produced

we know:

- *The chemical reaction*

$$NH_3\ (g) + CuO\ (s) \rightarrow N_2\ (g) + Cu\ (s) + H_2O\ (g)$$

- *18.1 g NH_3*
- *90.4 g CuO*

Balance the equation

$$2NH_3\ (g) + 3CuO\ (s) \rightarrow N_2\ (g) + 3Cu\ (s) + 3H_2O\ (g)$$

What are the moles of NH_3 and CuO?

Molar masses NH_3 =17.03 g/mol CuO =79.55 g/mol

$$18.1\ \cancel{g\ NH_3} \times \frac{1\ mol\ NH_3}{17.03\ \cancel{g\ NH_3}} = 1.06\ mol\ NH_3$$

$$90.4\ \cancel{g\ CuO} \times \frac{1\ mol\ CuO}{79.55\ \cancel{g\ CuO}} = 1.14\ mol\ CuO$$

Moles of CuO required to react with 1.06 mol NH_3

$$1.06\ \cancel{mol\ NH_3} \times \frac{3\ mol\ CuO}{2\ \cancel{mol\ NH_3}} = 1.59\ mol\ CuO$$

However, we have only 1.14 mol CuO. Therefore CuO is limiting (CuO will run out before NH_3 does).

Verification: Compare the required and actual mole ratios of CuO and NH_3 in the balanced equation:

$$Required = \frac{mol\ CuO}{mol\ NH_3} = \frac{3}{2} = 1.5$$

$$Actual = \frac{mol\ CuO}{mol\ NH_3} = \frac{1.14}{1.06} = 1.08$$

Since the actual ratio is smaller than required (1.5), CuO is the limiting reactant.

A second verification: the limiting reactant should have the lowest ratio of moles available/coefficient in the balanced equation.

$$For\ NH_3 = \frac{Moles\ available}{Stoichiometry} = \frac{1.06}{2} = 0.535$$

$$For\ CuO = \frac{Moles\ available}{Stoichiometry} = \frac{1.14}{3} = 0.38$$

Since the ratio of CuO is smaller, CuO is the limiting reactant.

Find the mass of N_2 produced:

We must use the amount of limiting reactant (CuO) to calculate the amount of N_2 formed.

First calculate the moles of N_2

$$1.14\ \cancel{mol}\ CuO \times \frac{1\ mol\ N_2}{3\ \cancel{mol}\ CuO} = 0.38\ mol\ N_2$$

Next calculate the mass of N_2

- ✓ *Using the molar mass of N_2 (28.02 g/mol)*

$$0.38\ \cancel{mol}\ N_2 \times \frac{28.02\ g\ N_2}{1\ \cancel{mol}\ N_2} = 10.6\ g\ N_2$$

Percent yield

- **Theoretical yield:** the amount of a product that is expected by calculations to be obtained assuming the limiting reactant is completely consumed. Side reactions usually reduce the yield than calculations.

- **Actual yield:** the amount of a product that is obtained actually.

$$\boxed{Percent\ Yield,\% = \frac{Actual\ Yield}{Theoritical\ Yield} \times 100}$$

Exercise

- Methanol (CH_3OH) is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combining gaseous carbon monoxide and hydrogen. Suppose 68.5 kg CO(g) is reacted with 8.60 kg H_2(g).
  - Calculate the theoretical yield of methanol?
  - If 3.57×10^4 g CH_3OH is actually produced, what is the percent yield of methanol?

Solution   Write a balanced equation

$$\boxed{2H_2\ (g) + CO\ (g) \rightarrow CH_3OH\ (l)}$$

Convert mass to moles

$$68.5\,\cancel{kg}\,CO \times \frac{1000\,\cancel{g}\,CO}{1\,\cancel{kg}\,CO} \times \frac{1\,mol\,CO}{28.02\,\cancel{g}\,CO} = 2.44\times10^{3}\ mol\,CO$$

$$8.60\,\cancel{kg}\,H_2 \times \frac{1000\,\cancel{g}\,H_2}{1\,\cancel{kg}\,H_2} \times \frac{1\,mol\,H_2}{2.016\,\cancel{g}\,H_2} = 4.27\times10^{3}\ mol\,H_2$$

Which is limiting

$$For\ H_2 = \frac{Moles\ available}{Stoichiometry} = \frac{4.27\times10^{3}}{2} = 2.135\times10^{3}$$

$$For\ CO = \frac{Moles\ available}{Stoichiometry} = \frac{2.44\times10^{3}}{1} = 2.44\times10^{3}$$

H_2 is limiting

The theoretical Yield

$H_2 : CH_3OH = 1:1$

$$2.135\times10^{3}\ \cancel{mol}\,CH_3OH \times \frac{32.04\,g\,CH_3OH}{1\,\cancel{mol}\,CH_3OH} = 6.86\times10^{4}\ g\,CH_3OH$$

$$Percent\ Yield,\% = \frac{Actual\ Yield}{Theoritical\ Yield}\times100$$

$$= \frac{3.57\times10^{4}\ g}{6.86\times10^{4}\ g}\times100 = 52\%$$

Exercise

Consider the equation

$$2A + B \rightarrow A_2B$$

If you mix 1.0 mole of A with 2.0 mole of B, the number of moles of A_2B produced if the reaction is 100% complete is --------

A is the limiting reactant and consequently 0.5 mole A_2B produced

Exercise

The limiting reactant in a chemical reaction

- ❑ *has the lowest coefficient in a balanced equation.* ×
- ❑ *has the lowest ratio of moles available/coefficient in the balanced equation.* √
- ❑ *has the lowest ratio of coefficient in the balanced equation/moles available.* ×