Limiting and Excess Reagents

Limiting reactant: will be completely used up during the chemical reaction → determines all other amount

Excess reactant: will NOT be completely used up during the chemical reaction → has some left over at the end

Percent Yield: a method to calculate the effectiveness of a chemical reaction.

- **Actual yield**: what you produce from actually doing the reaction in a lab setting.
- **Theoretical yield:** what "should have been" produced from chemical reaction. This is calculated with Stoich !

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

How to Find the Limiting Reactant: Two Methods

1. Calculate the number of moles of each reactant and then divide each one by its coefficient from the balanced reaction.

   → The reactant which produces the smaller number will be the limiting reactant.

   Advantage: very fast (great for multiple choice!)

   Disadvantage: you still have to actually calculate whatever info is asked for by the problem.

2. Use a BCA table to calculate the ending number of moles for both reactants and products.

   → The reactant which is entirely used up will be the limiting reactant.

   Advantage: you will calculate the ending amount of your excess reactant AND products, so any calculation you need is already complete.

   Disadvantage: you have to be in MOLES, so you might have extra math to do before and after.

Let's try BOTH methods! Which do you like best?

1. Methanol is combusted according to the balanced equation below. If you have 50.0 g of methanol and 50.0 g of molecular oxygen, how many grams of excess reactant remain unreacted?

method 1:

$$2\,CH_3OH(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

$\frac{50.0\text{ g }CH_3OH}{32.042\text{ g/mol}} = \frac{1.56\text{ mol }CH_3OH}{2} = 0.780$

$\frac{50.0\text{ g }O_2}{32\text{ g/mol}} = \frac{1.56\text{ mol }O_2}{3} = 0.520$

limiting! smaller!

$1.56\text{ mol }O_2 \times \frac{2\text{ mol }CH_3OH}{3\text{ mol }O_2} \times \frac{32.042\text{ g}}{1\text{ mol}} = \underbrace{33.3\text{ g}}_{\text{used up}}$ CH_3OH

$50.0 - 33.3 = \boxed{16.7\text{ g left over}}$

method 2:

	$2CH_3OH$	$+ 3O_2 \rightarrow$	$2CO_2$	$+ 4H_2O$
B	1.56	1.56	Ø	Ø
C	−1.04	−1.56	+1.04	+2.08
A	0.52	0	1.04	2.08

← not needed for this problem!

$1.56\text{ mol }O_2 \times \frac{2\,CH_3OH}{3\,O_2}$

$\underbrace{0.52}_{2\text{ s.f.}}\text{ mol }CH_3OH \times \frac{32.042\text{ g}}{1\text{ mol}} = 16.7$

$= \boxed{17\text{ g left over}}$

2. Aqueous sodium hydroxide reacts with phosphoric acid to give sodium phosphate and water in the following balanced chemical equation: $3\ NaOH\ (aq) + H_3PO_4\ (aq) \rightarrow Na_3PO_4\ (aq) + 3\ H_2O\ (l)$

If 20.80 g of NaOH is mixed with 29.40 g of H_3PO_4 (NaOH = 40.00 g/mol, H_3PO_4 = 98.00 g/mol):

a. How many grams of **Na_3PO_4** can be formed?

limiting!

$$\frac{20.80\ g\ NaOH}{40.00\ g/mol} = \frac{0.5200\ mol\ NaOH}{3} = 0.1733$$

Smaller!

$$\frac{29.40\ g\ H_3PO_4}{98.00\ g/mol} = \frac{0.3000\ mol}{1} = 0.3000$$

	$3\ NaOH$	$+\ H_3PO_4$	$\rightarrow Na_3PO_4$	$+\ 3H_2O$
B	0.5200	0.3000	Ø	Ø
C	-0.5200	-0.1733	0.1733	0.5200
A	Ø	$\boxed{0.1267}$	0.1733	0.5200

$$0.5200\ mol\ NaOH \times \frac{1\ H_3PO_4}{3\ NaOH} = 0.1733$$

$$0.1733\ mol\ Na_3PO_4 \times \frac{163.94\ g}{1\ mol} = \boxed{28.41\ g\ Na_3PO_4}$$

theoretical yield!

b. How many grams of the excess reactant remain unreacted?

$$\boxed{0.1267\ mol\ H_3PO_4} \times \frac{98.00\ g}{1\ mol} = \boxed{12.42\ g\ H_3PO_4 \text{ left over}}$$

c. If the actual yield of Na_3PO_4 was 15.00 g, what is the percent yield of Na_3PO_4?

$$\%\ yield = \frac{actual}{theor.} \times 100 = \frac{15.00\ g}{28.41\ g} \times 100 = \boxed{52.80\%}$$