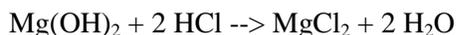


LIMITING REAGENT PROBLEMS

The first step in solving a limiting reagent problem is being able to recognize that you have a limiting reagent problem.

Suppose you were given the following problem:

A 50.6g sample of $\text{Mg}(\text{OH})_2$ is reacted with 45.0g of HCl according to the reaction:



What is the theoretical yield of MgCl_2 ?

Is this a limiting reagent problem? One way to find out is to write down what is known about any component of the reaction below that component:



Notice how quantities of both reactants are known. Which one will be used up first? You can't tell, nor should you jump to any conclusions. Just because it looks like there is less $\text{Mg}(\text{OH})_2$ present does not automatically mean it will be used up before all of the HCl is consumed. This is a limiting reagent problem.

There are numerous ways to solve a limiting reagent problem. I will present 3 different methods. You choose the method that you feel most comfortable with.

Method 1

In order to find out which reactant is the limiting reagent, you have to compare them to each other. This comparison must be done in moles; therefore, the next step will be to convert each of the grams of reactants to moles using their molar mass:

$$50.6 \text{ g Mg}(\text{OH})_2 \times \frac{1 \text{ mol Mg}(\text{OH})_2}{58.3 \text{ g Mg}(\text{OH})_2} = 0.868 \text{ mol Mg}(\text{OH})_2$$

$$45.0 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} = 1.23 \text{ mol HCl}$$

Again, you should not jump any conclusions about which reactant is the limiting reagent. Just because there are fewer moles of magnesium hydroxide does not mean it is the limiting reagent. Arbitrarily pick one of these reactants and calculate how many moles of the other reactant is needed to completely use up the reactant picked. In this case, magnesium hydroxide is arbitrarily chosen:

$$0.868 \text{ mol Mg}(\text{OH})_2 \times \frac{2 \text{ mol HCl needed}}{1 \text{ mol Mg}(\text{OH})_2} = 1.74 \text{ mol HCl needed}$$

Compare the moles HCl needed to the actual moles HCl available. In this case, 1.74 mole of HCl is needed and 1.23 mole HCl is available--that's not enough. So, even though it appears that there are more moles of HCl than $\text{Mg}(\text{OH})_2$, the HCl is the limiting reagent. The HCl will be run out before the magnesium hydroxide and thereby limit the amount of product formed. For this reason, use the moles of HCl to calculate the theoretical yield of magnesium chloride:

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$$1.23 \text{ mol HCl} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol HCl}} \times \frac{95.3 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 58.6 \text{ g MgCl}_2$$

The theoretical yield is the maximum amount of product that can be produced (in an ideal world). In the "real" world it is difficult to produce the amount obtained for the theoretical yield. A percent yield is often used to show how close to ideality one has obtained in a chemical synthesis. Suppose in the reaction discussed a chemist actually obtained 55.4 g of MgCl₂. This is called the actual yield and would be given to you in the problem. To calculate the percent yield:

$$\% \text{ yield} = \frac{55.4 \text{ actual g MgCl}_2}{58.6 \text{ theoretical g MgCl}_2} \times 100 = 94.5 \%$$

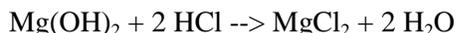
Method 2

An alternative method involves comparing the theoretical ratio of the reactants to the actual ratio of reactants available. First, calculate the moles of each reactant using their molar mass:

$$50.6 \text{ g Mg(OH)}_2 \times \frac{1 \text{ mol Mg(OH)}_2}{58.3 \text{ g Mg(OH)}_2} = 0.868 \text{ mol Mg(OH)}_2 \text{ available}$$

$$45.0 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} = 1.23 \text{ mol HCl available}$$

Consider the balanced reaction:



From the balanced equation, the theoretical mole ratio is:

$$\frac{2 \text{ moles of HCl needed}}{1 \text{ mol Mg(OH)}_2}$$

Let's see what the actual mole ratio is based on actual the amounts of reactants present:

$$\frac{1.23 \text{ mol HCl}}{0.868 \text{ mol Mg(OH)}_2} = \frac{1.42 \text{ mol HCl present}}{1 \text{ mol Mg(OH)}_2}$$

Comparing these two ratios, one can easily see that there is not enough HCl, so HCl must be the limiting reagent:

$$\frac{2 \text{ mol HCl needed}}{1 \text{ mol Mg(OH)}_2} \text{ compared to } \frac{1.42 \text{ mol HCl present}}{1 \text{ mol Mg(OH)}_2}$$

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Since HCl is the limiting reagent, use the moles of HCl available to calculate the theoretical yield of MgCl₂:

$$1.23 \text{ mol HCl} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol HCl}} \times \frac{95.3 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 58.6 \text{ g MgCl}_2$$

Method 3

A third method is to calculate the theoretical yield of product produced by each reactant and chose the lesser amount:

$$50.6 \text{ g Mg(OH)}_2 \times \frac{1 \text{ mol Mg(OH)}_2}{58.3 \text{ g Mg(OH)}_2} \times \frac{1 \text{ mol MgCl}_2}{1 \text{ mol Mg(OH)}_2} \times \frac{95.3 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 82.7 \text{ g MgCl}_2$$

$$45.0 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol HCl}} \times \frac{95.3 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 58.6 \text{ g MgCl}_2$$

Since HCl produced less product, HCl is the limiting reagent and 58.6 g MgCl₂ is the theoretical yield.

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