VANCOUVER COMMUNITY COLLEGE

Limiting Reagent Problems

When we mix chemicals together and they react, they recombine their atoms according to the formula equation of the reaction, and they do so in fixed ratios. It's likely that we will add more of one reagent than the ratio calls for, and so we will run out of the other reagent before we've used everything up. The reagent that we run out of first is called the **limiting reagent**, because it limits how long the reaction will go on. We say that the other reagent(s) are **in excess**. There will be some of these reagents left over when the reaction has run its course.

To find the limiting reagent, we must compare the molar amounts of reagents to the mole-to-mole ratio given by the formula equation.

Example 1: A solution containing 17.7 g of $AgNO_3$ is added to a solution containing 7.53 g of $CaC\ell_2$. $AgC\ell$ precipitates according to the equation:

$$\underset{\text{silver nitrate}}{2\,AgNO_3} + \underset{\text{calcium chloride}}{CaCl_2} \rightarrow \underset{\text{silver chloride}}{2\,AgCl\downarrow} + \underset{\text{calcium nitrate}}{Ca(NO_3)_2}$$

Determine the mass of silver chloride that precipitates, and the mass and identity of the reagent in excess.

Solution: [1] **Convert both reacting quantities to moles**, since we're dealing with a reaction.

$$\text{silver nitrate: } 17.7\text{ g }AgNO_3 \times \frac{1\text{ mol }AgNO_3}{169.88\text{ g }AgNO_3} = 0.104\text{ mol }AgNO_3$$

$$\text{calcium chloride: } 7.53\text{ g }CaCl_2 \times \frac{1\text{ mol }CaCl_2}{110.98\text{ g }CaCl_2} = 0.0679\text{ mol }CaCl_2$$

[2] **Divide the molar amounts of each reagent by that reagent's coefficient in the equation. The lowest answer is the limiting reagent.**

silver nitrate: 0.104 mol ÷ 2 = 0.0520 mol

calcium chloride: 0.0679 mol ÷ 1 = 0.0679 mol

Silver nitrate is the limiting reagent.

Let's stop for a moment and look at what that means in this problem. Silver nitrate reacts with calcium chloride in a ratio of 2 : 1. We would expect 0.104 mol of AgNO3 to react completely with 0.0520 mol of $CaC\ell_2$, since that's also a 2 : 1 ratio. We would expect 0.0679 mol of $CaC\ell_2$ to react with twice that much $AgNO_3$, 0.136 mol, but we don't have that much. We will run out of silver nitrate before we've used up all our calcium chloride.

Authored by Gordon Wong

[3] **Calculate the amounts of any products and reagents involved in the reaction that the question asks for.** We'll use conversion fractions and the coefficients in the formula to do this, basing our calculations on the limiting reagent.

$$\text{calcium chloride:}\quad 0.104\,\text{mol AgNO}_3 \times \frac{1\,\text{mol CaCl}_2}{2\,\text{mol AgNO}_3} = 0.0520\,\text{mol CaCl}_2$$

$$\text{silver chloride:}\quad 0.104\,\text{mol AgNO}_3 \times \frac{2\,\text{mol AgCl}}{2\,\text{mol AgNO}_3} = 0.104\,\text{mol AgCl}$$

So 0.104 mol $AgC\ell$ is precipitated, and 0.0520 mol $CaC\ell_2$ is reacted.

[4] **Determine the mass of the remaining excess reagent.** Since 0.0520 mol of $CaC\ell_2$ is used up, and we started with 0.0679 mol of $CaC\ell_2$, there is 0.0159 mol of $CaC\ell_2$ remaining. We need to convert this to grams.

$$0.0159\,\text{mol CaCl}_2 \times \frac{110.98\,\text{g CaCl}_2}{1\,\text{mol CaCl}_2} = 1.76\,\text{g CaCl}_2$$

[5] **Determine the mass of the precipitate.**

$$0.104\,\text{mol AgCl} \times \frac{143.32\,\text{g AgCl}}{1\,\text{mol AgCl}} = 14.9\,\text{g AgCl}$$

EXERCISES

A. Consider the following balanced equation:

$$\underset{\text{chlorine dioxide}}{6\,\text{ClO}_2} + \underset{\text{water}}{3\,\text{H}_2\text{O}} \rightarrow \underset{\text{chloric acid}}{5\,\text{HClO}_3} + \underset{\text{hydrochlor ic acid}}{\text{HCl}}$$

4.25 g of chlorine dioxide and 0.853 g of water are reacted.

1) How many moles of $C\ell O_2$ are available for reaction?

2) How many moles of H_2O are available for reaction?

3) Which is the limiting reagent?

4) Which is the excess agent?

5) How many moles of the excess reagent were actually consumed in the reaction?

6) How many moles of the excess reagent remain after the reaction has finished?

7) How many grams of the excess reagent remain?

8) How many moles of $HC\ell O_3$ are produced?

9) How many grams of $HC\ell O_3$ are produced?

10) How many moles of $HC\ell$ are produced?

11) How many grams of $HC\ell$ are produced?

B. Consider the following balanced equation:

$$\underset{\text{arsenic}}{2\,As} + \underset{\text{bromine}}{3\,Br_2} \rightarrow \underset{\text{arsenic (III) bromide}}{2\,AsBr_3}$$

1) How many grams of arsenic (III) bromide can be produced when 55.3 g of arsenic is mixed with 125 g of bromine?

2) What is the mass and identity of the reagent in excess?

C. Silver tarnishes in the presence of hydrogen sulphide:

$$\underset{\text{silver}}{4\,Ag} + \underset{\text{hydrogen sulphide}}{2\,H_2S} + \underset{\text{oxygen}}{O_2} \rightarrow \underset{\text{silver sulphide}}{2\,Ag_2S} + \underset{\text{water}}{2\,H_2O}$$

Silver sulphide is responsible for the black colour. Hydrogen sulphide is released into the air during the decomposition of organic matter.

1) How many grams of Ag_2S could be obtained from a mixture of 0.750 g Ag, 0.155 g H_2S and 0.900 g O_2?

2) Which is the limiting reagent, and which are the excess reagents? How many grams of the excess reagents remain in excess?

3) How many grams of water are produced in the reaction?

SOLUTIONS

A. (1) 0.0630 mol (2) 0.0474 mol (3) chlorine dioxide (4) water (5) 0.0315 mol (6) 0.0159 mol (7) 0.286 g H_2O (8) 0.0525 mol (9) 4.43 g $HC\ell O_3$ (10) 0.0105 mol (11) 0.383 g $HC\ell$

B. (1) 164 g $AsBr_3$ (2) 16.2 g of arsenic

C. (1) 0.861 g (2) silver is the limiting reagent; 0.0366 g H_2S and 0.844 g O_2 (3) 0.0626 g