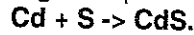


Stoichiometry WS # 5 c.
Stoichiometry of excess quantities and % Yield

Answer Key

A substance that is not present in sufficient quantities to react with all of another substance present is called a limiting reactant; the other substance is said to be in excess.

1. If 2.20 g of cadmium reacts with 0.570 g of sulphur to produce cadmium sulphide as shown by



- a) Which is the limiting reactant?

$$2.20 \text{ g Cd} \times \frac{1 \text{ mol Cd}}{112.4 \text{ g Cd}} = 0.01957$$

$$0.570 \text{ g S} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} = 0.017757 \text{ limiting}$$

(1:1 ratio)

- b) How much (mass) cadmium sulphide is produced?

$$0.570 \text{ g S} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} \times \frac{1 \text{ mol CdS}}{1 \text{ mol S}} \times \frac{144.5 \text{ g CdS}}{1 \text{ mol CdS}} = 2.57 \text{ g}$$

- b) Which reactant is the reactant with the excess quantity? Calculate the mass of the reactant in excess that is left over from the reaction above.

$$0.017757 \text{ mol S used} \times \frac{1 \text{ mol Cd}}{1 \text{ mol S}} = 0.017757 \text{ mol Cd used}$$

$$0.01957 - 0.017757 = 0.001813 \text{ mol Cd left} \times \frac{112.4 \text{ g Cd}}{1 \text{ mol Cd}} = 0.20 \text{ g}$$

2. In an experiment, 8.51 g of H_2 and 9.25 g O_2 are placed in a reaction vessel. The introduction of a spark causes a violent explosion that generates water.



- a) Calculate the mass of water generated from this reaction. (you need to find the limiting reactant first)

$$8.51 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.0 \text{ g H}_2} = 4.255 \text{ mol H}_2 \rightarrow 2.128 \text{ excess}$$

limiting $9.25 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} = 0.289 \text{ mol} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 10.4 \text{ g H}_2\text{O}$

- b) Calculate the mass of the reactant that was in excess amount.

$$0.289 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \times \frac{2.0 \text{ g H}_2}{1 \text{ mol H}_2} = 1.16 \text{ g H}_2 \text{ used}$$

$$8.51 - 1.16 = 7.35 \text{ g H}_2 \text{ left}$$

3. In a precipitation reaction, 50.0 mL of a 0.40 M CaCl_2 reacts with 60.0 mL of 0.30 M Na_2CO_3 solution. Calculate the mass of CaCO_3 produced? This is a double replacement reaction.

$$0.0500 \text{ L} \times 0.40 \text{ M} = 0.0200 \text{ mol CaCl}_2$$

$$0.0600 \text{ L} \times 0.30 \text{ M} = 0.0180 \text{ mol Na}_2\text{CO}_3 \text{ limiting}$$

$$\text{CaCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow 2\text{NaCl} + \text{CaCO}_3$$

$$0.0180 \text{ mol Na}_2\text{CO}_3 \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol Na}_2\text{CO}_3} \times \frac{100.1 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 1.80 \text{ g CaCO}_3$$

Percent Yield. The actual yield divided by the ideal, or theoretical yield, and multiplied by 100 equals the percent yield. Calculate the percent yield of the following:

1. The maximum amount of CaCO_3 that could be made during the experiment was 1.2 g. This is referred to as the theoretical yield. The actual amount that was obtained was 0.80 g. What is the percent yield?

$$\frac{0.80\text{g}}{1.20\text{g}} \times 100\% = 66.7\%$$

2. $\text{Ca} + \text{Cl}_2 \rightarrow \text{CaCl}_2$

If 40.0 g of Ca and 80.0 of Cl_2 were present as reactants. Only 106 g of CaCl_2 is produced.

What is the percent yield? You need to find the limiting reactant, and then calculate the theoretical yield first.

$$40.0\text{g Ca} \times \frac{1\text{mol Ca}}{40.1\text{g Ca}} = 0.9975\text{ mol } \star \text{ limiting}$$

$$80.0\text{g Cl}_2 \times \frac{1\text{mol Cl}_2}{71.0\text{g Cl}_2} = 1.127\text{ mol}$$

$$\frac{106\text{g}}{111\text{g}} \times 100\% = 95.5\%$$

$$0.9975\text{ mol Ca} \times \frac{1\text{mol CaCl}_2}{1\text{mol Ca}} \times \frac{111.1\text{g CaCl}_2}{1\text{mol CaCl}_2} = 111\text{g}$$

3. $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$

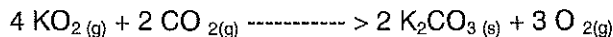
If 2.8 g of N_2 and 5.0 g of H_2 were present as reactants. Only 1.45 g of NH_3 are produced. What is the percent yield?

$$2.8\text{g N}_2 \times \frac{1\text{mol N}_2}{28.0\text{g N}_2} = 0.10\text{ mol} \quad \text{limiting} \quad 5.0\text{g H}_2 \times \frac{1\text{mol}}{2.0\text{g}} = \frac{2.5\text{g}}{3}$$

$$0.10\text{ mol N}_2 \times \frac{2\text{mol NH}_3}{1\text{mol N}_2} \times \frac{17.0\text{g NH}_3}{1\text{mol NH}_3} = 3.4\text{g NH}_3$$

$$\frac{1.45\text{g}}{3.4\text{g}} \times 100 = 42.6\%$$

4. What mass of K_2CO_3 is produced when 0.50 g of KO_2 is reacted with an excess of CO_2 according to the reaction:



if the reaction produces a 55.0% yield?

$$0.50\text{g KO}_2 \times \frac{1\text{mol KO}_2}{71.1\text{g KO}_2} \times \frac{2\text{mol K}_2\text{CO}_3}{4\text{mol KO}_2} \times \frac{138.2\text{g K}_2\text{CO}_3}{1\text{mol K}_2\text{CO}_3} = 0.4859\text{g}$$

$$0.4859\text{g} \times 0.55 = 0.27\text{g K}_2\text{CO}_3$$