

## Limiting Reactant Problems

limiting reactant - the reactant which determines the amount of product that will be produced

-the reactant that gets used up completely

-the reactant that makes the smallest amount of product

excess reactant-any reactant not completely used up during a reaction

-the reactant that remains unreacted at the end of the experiment

e.g. In a reaction, 5.00g of CH<sub>4</sub> reacts with 7.00g of O<sub>2</sub> according to the following equation:



Determine the limiting reactant and the mass of water produced.

First we will determine the limiting reactant. From the equation we can see that 1 mole of CH<sub>4</sub> requires 2 moles of O<sub>2</sub> to completely react.

Lets determine how many moles of each reactant is present and compare our results to the number of moles required for a complete reaction as indicated in the balanced chemical equation.

moles of CH<sub>4</sub> present in 5.00g of CH<sub>4</sub>...

$$5.00\text{g CH}_4 \times \frac{1 \text{ mole CH}_4}{16.05\text{g CH}_4} = 0.312 \text{ mol CH}_4$$

now lets determine the number of moles of O<sub>2</sub> required to completely react with the CH<sub>4</sub> present...

1 mol CH<sub>4</sub> = 2 mol O<sub>2</sub> c.f. (from balanced equation)

$$0.312 \text{ mol CH}_4 \times \frac{2 \text{ mol O}_2}{1 \text{ mol CH}_4} = 0.624 \text{ mol O}_2$$

so the balanced equation indicates that 0.312 mol of CH<sub>4</sub> require 0.624 mol of O<sub>2</sub> to completely react. Now we need only to find how many moles of O<sub>2</sub> are present in 7.00g of O<sub>2</sub> and compare it to the number of moles of O<sub>2</sub> required by the balanced equation as indicated above (0.624 mol O<sub>2</sub>)...

lets determine the number of moles of O<sub>2</sub> present in 7.00g of O<sub>2</sub>...

$$M_{\text{O}_2} = 32.00\text{g} \quad 7.00\text{g O}_2 \times \frac{1 \text{ mol O}_2}{32.00\text{g O}_2} = 0.219 \text{ mol O}_2$$

the balanced equation requires 0.624 mol of O<sub>2</sub> to react with 0.312 mol of CH<sub>4</sub> but only 0.219 mol of O<sub>2</sub> are present in 7.00g of O<sub>2</sub>...

Moles of reactant required by balanced equation > moles of reactant present

$$0.624 \text{ mol O}_2 > 0.219 \text{ mol O}_2$$

∴ O<sub>2</sub> is the limiting reactant

Now that we know that  $O_2$  is the limiting reactant we can determine how many moles of  $H_2O$  are produced by 0.219 moles of  $O_2$ . We use the number of moles of  $O_2$  present because  $H_2O$  will continue being produced only as long as  $O_2$  is present. We know this because as the limiting reactant,  $O_2$  will run out first. So the mass of water produced is as follows...

first lets determine how many moles of  $H_2O$  will be produced by 0.219 moles of  $O_2$ ...

2 mol  $O_2$  = 2 mol  $H_2O$  (c.f. from balanced chemical equation)

$$0.219 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } O_2} = 0.219 \text{ mol } H_2O$$

now lets determine the mass of 0.219 moles of  $H_2O$ ...

$$M_{H_2O} = 18.02g \quad 0.219 \text{ mol } H_2O \times \frac{18.02g \text{ } H_2O}{1 \text{ mol } H_2O} = 3.94g \text{ } H_2O$$

$\therefore$  as dictated by the balance chemical equation, the reaction of 5.00g of  $CH_4$  with 7.00g of  $O_2$  will have  $O_2$  as the limiting reactant and produce 3.94g of  $H_2O$

It is very important to determine the limiting reactant so as to be able to correctly predict how much product will actually form!!!