

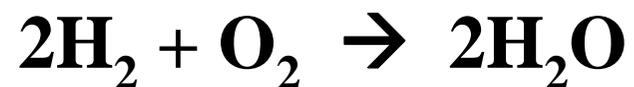


## Chemical equations

**Chemical equations are representations of reactions.**

**Instead of: 2moles of hydrogen react with one mole of oxygen to produce 2moles of water,**

**We write:**



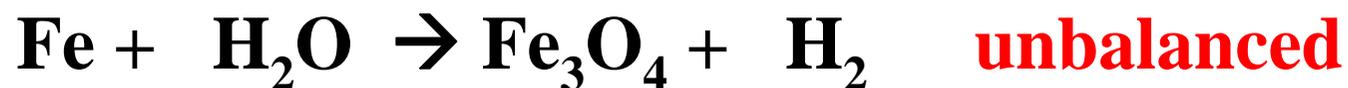
**Reactants      Products**



**In chemical equations:**

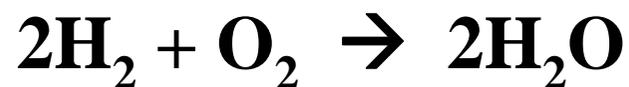
**Number of atoms of each element on the left =  
Number of the atoms on the right side  
(conservation of mass law)**

**Examples:**





**The equation:**



**can be read as follows:**

**2 molecules of H<sub>2</sub> react with 1 molecule of O<sub>2</sub> to produce 2 molecules of H<sub>2</sub>O**

**2 mol H<sub>2</sub> react with 1mol O<sub>2</sub> to produce 2mol H<sub>2</sub>O**

**2(2)g H<sub>2</sub> react with 1(32)g O<sub>2</sub> to produce 2(18)g H<sub>2</sub>O**



**The coefficients of the chemical equation give the ratios in which the substances react.**

**ratios can use to solve stoichiometric problems**

**we can answer questions of the type:**

**How much of a reactant we need to produce certain amount of a product?**

**How much product will be produced from certain amount of a reactant?**



## Example

Determine the number of  $O_2$  moles required to react with 4mol of  $C_2H_6$



the problem:  $X \text{ mol } O_2 = 4.0 \text{ mol } C_2H_6$

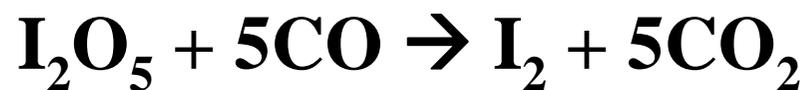
the stoichiometric ratio:  $7\text{mol } O_2 = 2\text{mol } C_2H_6$

$$X \text{ mol } O_2 = \frac{(7\text{mol } O_2)(4.0 \text{ mol } C_2H_6)}{(2\text{mol } C_2H_6)} = 14 \text{ mol } O_2$$



## Example:

The amount of CO in a sample of a gas can be determined by the reaction



If a gas sample liberates 0.192g of I<sub>2</sub>, how many grams of CO were present in the sample?

$$n_{\text{I}_2} = \frac{0.192 \text{ g}}{254 \text{ g / mol}} = 7.56 \times 10^{-4} \text{ mol}$$



**from the problem:  $X \text{ mol CO} = 7.56 \times 10^{-4} \text{ mol I}_2$**

**from the equation:  $5 \text{ mol CO} = 1 \text{ mol I}_2$**

$$X \text{ mol CO} = \frac{(5 \text{ mol CO})(7.56 \times 10^{-4} \text{ mol I}_2)}{(1 \text{ mol I}_2)} =$$

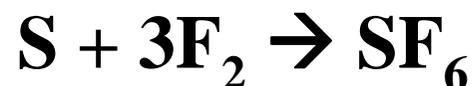
$$3.78 \times 10^{-3} \text{ mol CO}$$

$$3.78 \times 10^{-3} \text{ mol CO} \times 28.01 \text{ g/mol CO} = 0.106 \text{ g CO}$$



## Limiting reactant & yield of reactions

For the reaction



Suppose we have

20 moles of  $\text{F}_2$  and 4 moles of S

**which reactant will determine the quantity of the product?**

from the problem:  $4\text{mol S} = X\text{mol F}_2$

from the equation:  $1\text{mol S} = 3\text{mol F}_2$

#moles of  $\text{F}_2$  needed to react with 4 moles of S =



$$\frac{4\text{mol S} \times 3\text{mol F}_2}{1\text{mol S}} = 12 \text{ mol F}_2$$

**but we have 20 moles of F<sub>2</sub>**

**let's use the 20 moles of F<sub>2</sub>**

**from the problem: Xmol S = 20mol F<sub>2</sub>**

**from the equation: 1mol S = 3mol F<sub>2</sub>**

**#moles of S needed to react with 20moles of F<sub>2</sub> =**

$$\frac{1\text{mol S} \times 20\text{mol F}_2}{3\text{mol F}_2} = 6.7 \text{ mol S}$$



but we have only 4 moles of S

which reactant will be consumed first?

***Sulfur***

***Limiting reactant: the reactant that is consumed first.***

**An easy way to determine the limiting reactant:**

**For all reactants determine the ratio**

$$\frac{\text{amount of the reactant from the problem}}{\text{amount of the reactant from the equation}}$$



**The smallest number belongs to the limiting reactant.**

**In our example:**

$$\text{For S: } \frac{\text{amount of S from the problem}}{\text{amount of S from the equation}} = \frac{4 \text{ mol S}}{1 \text{ mol S}} = 4$$

$$\text{For F}_2: \frac{\text{amount of F}_2 \text{ from the problem}}{\text{amount of F}_2 \text{ from the equation}} = \frac{20 \text{ mol F}_2}{3 \text{ mol F}_2} = 6.7$$

**S has the smaller ratio,**

**∴ S is the limiting reactant**



## *Percent yield of a reaction*

**Theoretical yield** from a reaction is the yield calculated by assuming that the reaction goes to completion.

In practice: **Actual yield** is usually less than theoretically expected.

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100$$

**Example:**

How many moles of  $\text{H}_2$  can theoretically be produced from 4.00mol of Fe and 5.00mol of  $\text{H}_2\text{O}$ ?





a. determine the limiting reagent

for Fe:  $\frac{4.00}{3} = 1.33$

for H<sub>2</sub>O:  $\frac{5.00}{4} = 1.25$  (smaller ratio)

the limiting reagent **H<sub>2</sub>O**

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b. theoretical yield

4 mol H<sub>2</sub>O = 4mol H<sub>2</sub>

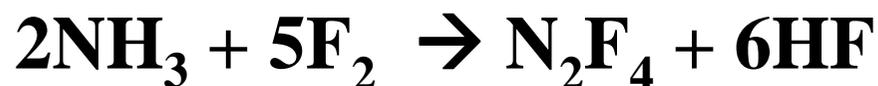
5.00mol H<sub>2</sub>O = Xmol H<sub>2</sub>

theoretical yield of H<sub>2</sub> = 5.00 mol H<sub>2</sub>



## Example

**4.80g of  $N_2F_4$  were obtained from the reaction of 4.00g of  $NH_3$  and 14.0g of  $F_2$ . What is the %yield of  $N_2F_4$ ?**



$$4.00g \text{ } NH_3 = \frac{4.00g}{17.0g / mol} = 0.235mol$$

$$14.0g \text{ } F_2 = \frac{14.0g}{38.00g / mol} = 0.368mol$$

**determine the limiting reagent:**



$$\text{for NH}_3 : \frac{0.235}{2} = 0.118$$

$$\text{for F}_2 : \frac{0.368}{5} = 0.0736 \quad \text{(smaller ratio)}$$

**F<sub>2</sub> is the limiting reagent**

**from the problem: 0.368mol F<sub>2</sub> = X mol N<sub>2</sub>F<sub>4</sub>**

**from the equation: 5 mol F<sub>2</sub> = 1 mol N<sub>2</sub>F<sub>4</sub>**

$$\mathbf{X} = \frac{0.368 \text{ mol F}_2 \times 1 \text{ mol N}_2\text{F}_4}{5 \text{ mol F}_2} = 0.0736 \text{ mol N}_2\text{F}_4$$



$$\text{M.wt N}_2\text{F}_4 = 104\text{g/mol}$$

$$0.0736\text{mol N}_2\text{F}_4 = 0.0736\text{mol} \times 104\text{g/mol} = 7.65\text{g}$$

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100$$

$$\% \text{ yield} = \frac{4.80\text{g}}{7.65\text{g}} \times 100 = 62.7\%$$



## Example

47.7g of CuO was left to react with excess amount of H<sub>2</sub> according to the equation



How many grams of Cu were produced if the percent yield of Cu was 55%?

**Excess H<sub>2</sub> means that CuO is the limiting reagent**

$$\# \text{ moles of CuO} = \frac{47.7 \text{ g}}{79.5 \text{ g / mol}} = 0.600 \text{ mol}$$

from the equation: 1 mol CuO = 1mol Cu

from the problem: 0.600mol CuO = Xmol Cu



**Theoretical yield=0.600 moles of Cu**

**From the relation: % yield =  $\frac{\text{Actual yield}}{\text{theoretical yield}} \times 100$**

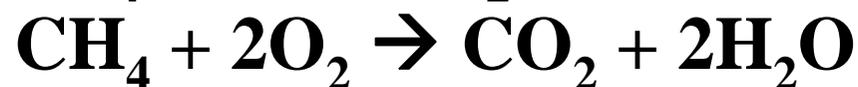
$$\frac{55}{100} = \frac{\text{Actual yield}}{0.600\text{mol}} \Rightarrow \text{Actual yield} = 0.33\text{mol of Cu}$$

$$\mathbf{0.33\text{mol Cu} = 0.33\text{mol} \times 63.5\text{g/mol} = 21 \text{ g}}$$



## Example

How many grams of  $\text{CO}_2$  can be prepared from the reaction of 8.0g  $\text{CH}_4$  and 48g  $\text{O}_2$ ?



$$8.0\text{g CH}_4 = \frac{8.0\text{g}}{16.04\text{g/mol}} = 0.50\text{mol}$$

$$48\text{g O}_2 = \frac{48\text{g}}{32\text{g/mol}} = 1.5\text{mol}$$



**Determine the limiting reagent**

$$\text{For CH}_4 : \frac{0.5 \text{ mol}}{1 \text{ mol}} = 0.5 \quad \text{(smaller ratio)}$$

$$\text{For O}_2 : \frac{1.5 \text{ mol}}{2 \text{ mol}} = 0.75$$

**CH<sub>4</sub> is the limiting reagent**

from the equation: 1 mol CH<sub>4</sub> = 1 mol CO<sub>2</sub>

from the problem: 0.5 mol CH<sub>4</sub> = X mol CO<sub>2</sub>

$$\text{X} = 0.5 \text{ mol CO}_2 = 0.5 \text{ mol} \times 44 \text{ g/mol} = 22 \text{ g}$$