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:

$$\underbrace{2H_2 + O_2}_{\text{Reactants}} \rightarrow \underbrace{2H_2O}_{\text{Products}}$$

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:

$Fe + H_2O \rightarrow Fe_3O_4 + H_2$ **unbalanced**

$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$ **balanced**

$KClO_3 \rightarrow KCl + O_2$ **unbalanced**

$2KClO_3 \rightarrow 2KCl + 3O_2$ **balanced**

The equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

can be read as follows:

2 molecules of H_2 react with 1 molecule of O_2 to produce 2 molecules of H_2O

2 mol H_2 react with 1mol O_2 to produce 2mol H_2O

2(2)g H_2 react with 1(32)g O_2 to produce 2(18)g H_2O

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

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

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

$$I_2O_5 + 5CO \rightarrow I_2 + 5CO_2$$

If a gas sample liberates 0.192g of I_2, how many grams of CO were present in the sample?

$$n_{I_2} = \frac{0.192\,g}{254\,g/mol} = 7.56 \times 10^{-4}\,mol$$

from the problem: Xmol CO = 7.56×10^{-4} mol I_2

from the equation: 5 mol CO = 1 mol I_2

$$X\ mol\ CO = \frac{(5mol\ CO)(7.56\times10^{-4}\,mol\ I_2)}{(1mol\ I_2)} =$$

$$3.78\times10^{-3}\ mol\ CO$$

3.78×10^{-3} mol CO × 28.01g/mol CO = 0.106g CO

Limiting reactant &yield of reactions

For the reaction

$$S + 3F_2 \rightarrow SF_6$$

Suppose we have

20moles of F_2 and 4 moles of S

which reactant will determine the quantity of the product?

from the problem: 4mol S = Xmol F_2

from the equation: 1mol S = 3mol F_2

#moles of F_2 needed to react with 4moles 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_2

let's use the 20 moles of F_2

from the problem: Xmol S = 20mol F_2

from the equation: 1mol S = 3mol F_2

#moles of S needed to react with 20moles of F_2 =

$$\frac{1\text{mol S} \times 20\text{mol } F_2}{3\text{mol } F_2} = 6.7 \text{ mol S}$$

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

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

The smallest number belongs to the limiting reactant.

In our example:

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

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,

$\therefore$ **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 H_2 can theoretically be produced from 4.00mol of Fe and 5.00mol of H_2O?

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

a. determine the limiting reagent

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

for H_2O: $$\frac{5.00}{4} = 1.25$$ (smaller ratio)

the limiting reagent H_2O

---

b. theoretical yield

4 mol H_2O = 4mol H_2

5.00mol H_2O = Xmol H_2

theoretical yield of H_2 = 5.00 mol H_2

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?

$$2NH_3 + 5F_2 \rightarrow N_2F_4 + 6HF$$

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

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

determine the limiting reagent:

$$\textbf{for } NH_3: \quad \frac{0.235}{2} = 0.118$$

$$\textbf{for } F_2: \quad \frac{0.368}{5} = 0.0736$$ **(smaller ratio)**

F_2 is the limiting reagent

from the problem: 0.368mol F2 = X mol N_2F_4

from the equation: 5 mol F_2 = 1 mol N_2F_4

$$\mathbf{X} = \frac{0.368\,mol\ F_2 \times 1\,mol\ N_2F_4}{5\,mol\ F_2} = 0.0736\,mol\ N_2F_4$$

M.wt N_2F_4 = 104g/mol

0.0736mol N_2F_4 = 0.736mol×104g/mol = 7.65g

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

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

Example

47.7g of CuO was left to react with excess amount of H_2 according to the equation

$$CuO\,(s) + H_2\,(g) \rightarrow Cu\,(s) + H_2O\,(l)$$

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

Excess H_2 means that CuO is the limiting reagent

 # moles of CuO = $\dfrac{47.7\,g}{79.5\,g/mol} = 0.600\,mol$

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

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

0.33mol Cu = 0.33mol × 63.5g/mol = 21 g

Example

How many grams of CO_2 can be prepared from the reaction of 8.0g CH_4 and 48g O_2?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

8.0g CH_4 = $$\frac{8.0g}{16.04g/mol} = 0.50mol$$

48g O_2 = $$\frac{48g}{32g/mol} = 1.5mol$$

Determine the limiting reagent

For CH_4 : $\frac{0.5\,mol}{1\,mol} = 0.5$ $\boxed{\textbf{(smaller ratio)}}$

For O_2 : $\frac{1.5\,mol}{2\,mol} = 0.75$

CH_4 is the limiting reagent

from the equation: 1 mol CH_4 = 1 mol CO_2

from the problem: 0.5mol CH_4 = X mol CO_2

X = 0.5mol CO_2 = 0.5mol ×44g/mol= 22g