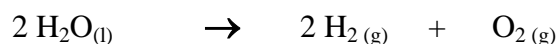


Chemistry 20

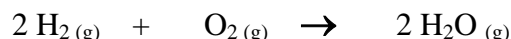
Lesson 6 – Stoichiometry I

I. Chemical reactions and the mole ratio

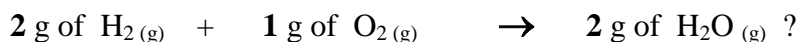
Chemistry is the study of how matter rearranges itself into different forms and combinations. For example, when water undergoes electrolysis the water molecules are broken apart resulting in the formation of hydrogen and oxygen gases.



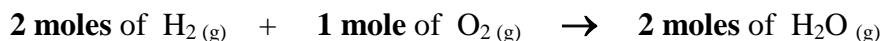
If the gases are then mixed together and a spark or flame ignites the gases, the atoms rearrange themselves again to form water molecules.



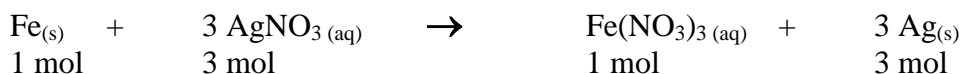
These are two simple examples of how matter rearranges itself into different forms. The next logical question is: **What do the coefficients in front of each molecule represent?** Do they represent a ratio of masses or a ratio of moles? If they represented a ratio of masses, then the following equation would hold:



If one were to test this hypothesis one would find that only 1.13 g of $\text{H}_2\text{O}_{(g)}$ would actually be produced. Therefore, the coefficients do not represent a mass ratio. The chemical reaction actually means:



If one were to test this hypothesis, it turns out to be correct. The coefficients represent a **mole ratio**. When chemicals combine with one another, they do so according to the mole amounts represented by the coefficients of the balanced chemical equation. For example, in the following reaction



the ratio of 1:3:1:3 remains **constant**. Further, if one started with 0.50 mol of Fe, the number of moles of Ag and $\text{Fe}(\text{NO}_3)_3$ produced could be calculated using the mole ratio:

To calculate moles of Ag we set up the mole ratio:

$$\begin{aligned} \frac{n_{\text{Ag}}}{3} &= \frac{n_{\text{Fe}}}{1} \\ \frac{n_{\text{Ag}}}{3} &= \frac{0.50 \text{ mol}}{1} \\ n_{\text{Ag}} &= \frac{3(0.50 \text{ mol})}{1} = 1.50 \text{ mol of Ag} \end{aligned}$$

Coefficients from the balanced equation.

To calculate moles of $\text{Fe}(\text{NO}_3)_3$ we set up the mole ratio:

$$\frac{n_{\text{Fe}(\text{NO}_3)_3}}{1} = \frac{n_{\text{Fe}}}{1}$$

Coefficients from the balanced equation.

$$\frac{n_{\text{Fe}(\text{NO}_3)_3}}{1} = \frac{0.50 \text{ mol}}{1}$$

$$n_{\text{Fe}(\text{NO}_3)_3} = \mathbf{0.50 \text{ mol of Fe}(\text{NO}_3)_3}$$

The mole ratio is a powerful principle. Once you know the number of moles of one of the compounds, you can calculate **any** of the others from the balanced equation.

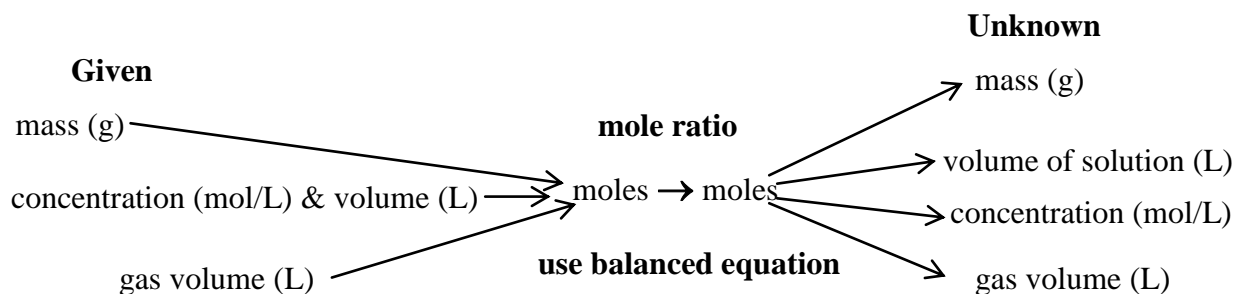
II. The stoichiometry method

From the previous section we can see that atoms combine with each other as a mole ratio and yet we measure compounds and molecules in terms of mass, gas volume, concentration, volume of solution, etc. **The process of stoichiometry combines the ideas of real world measurement (mass, volumes, etc.) and the chemical reaction represented by a balanced chemical equation.** In past lessons you have learned how to convert mass to moles and moles to mass. (In future lessons you will learn about concentrations & volumes of solutions, and gas volumes.) In addition, you have learned to predict, complete, and balance chemical equations. Now we will learn to apply these tools together to predict the results of any chemical reaction.

Steps for stoichiometry:

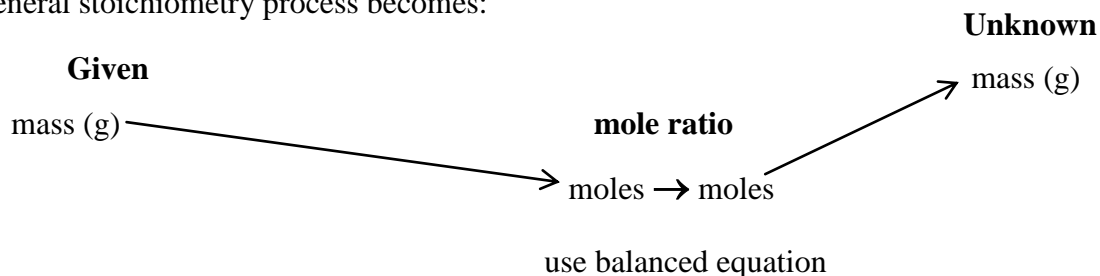
1. Read the problem.
2. Write down the complete, balanced chemical equation including all states of matter.
3. List what you are given and what is unknown under the appropriate compounds.
4. Convert the given amount into moles.
5. Apply the mole ratio.
6. Calculate the desired unknown.

Another way to represent the **general** stoichiometry process is:



III. Gravimetric stoichiometry

In our work to date, we have learned about masses and moles. **Gravimetric** stoichiometry is the term used to describe stoichiometry involving masses only. For gravimetric stoichiometry, the general stoichiometry process becomes:

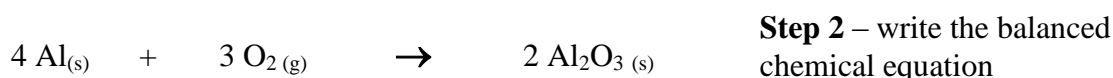


There are literally dozens of variations of methods for solving stoichiometric problems. Here are two basic ways that you can use to solve stoichiometry problems.

Mole ratio method

Step 1 – read the question

1. In the formation of aluminum oxide from its elements, what mass of aluminum is needed to form 152.9 g of aluminum oxide?



$$m_{\text{Al}} = ?$$

$$m_{\text{Al}_2\text{O}_3} = 152.9 \text{ g}$$

$$n_{\text{Al}_2\text{O}_3} = \frac{m_{\text{Al}_2\text{O}_3}}{M_{\text{Al}_2\text{O}_3}} = \frac{152.9 \text{ g}}{101.96 \text{ g/mol}}$$

$$n_{\text{Al}_2\text{O}_3} = 1.50 \text{ mol}$$

Step 5 - set up mole ratio

$$\frac{n_{\text{Al}}}{4} = \frac{n_{\text{Al}_2\text{O}_3}}{2}$$

$$\frac{n_{\text{Al}}}{4} = \frac{1.50 \text{ mol}}{2}$$

$$n_{\text{Al}} = 3.00 \text{ mol}$$

Step 6 - calculate mass of unknown

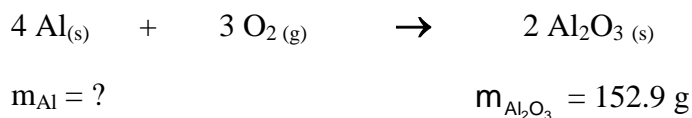
$$m_{\text{Al}} = n_{\text{Al}} \times M_{\text{Al}} = 3.00 \text{ mol} \times 26.98 \text{ g/mol}$$

$$m_{\text{Al}} = \mathbf{80.92 \text{ g}}$$

Unit analysis method

Step 1 – read the question

1. In the formation of aluminum oxide from its elements, what mass of aluminum is needed to form 152.9 g of aluminum oxide?



Step 2 – write the balanced chemical equation

Step 3 – list given and unknown quantities

Step 4 - beginning with what is given, eliminate units until you have the unknown only

$$m_{\text{Al}} = 152.9 \text{ g}(\text{Al}_2\text{O}_3) \times \frac{1 \text{ mol}(\text{Al}_2\text{O}_3)}{101.96 \text{ g}(\text{Al}_2\text{O}_3)} \times \frac{4 \text{ mol}(\text{Al})}{2 \text{ mol}(\text{Al}_2\text{O}_3)} \times \frac{26.98 \text{ g}(\text{Al})}{1 \text{ mol}(\text{Al})} = \mathbf{80.92 \text{ g}(\text{Al})}$$

(Note how the units cancel out.)

$$m_{\text{Al}} = 152.9 \text{ g}(\cancel{\text{Al}_2\text{O}_3}) \times \frac{1 \text{ mol}(\cancel{\text{Al}_2\text{O}_3})}{101.96 \text{ g}(\cancel{\text{Al}_2\text{O}_3})} \times \frac{4 \text{ mol}(\cancel{\text{Al}})}{2 \text{ mol}(\cancel{\text{Al}_2\text{O}_3})} \times \frac{26.98 \text{ g}(\text{Al})}{1 \text{ mol}(\text{Al})} = \mathbf{80.92 \text{ g}(\text{Al})}$$

IV. Practice problems

1. What mass of ammonia is produced by the reaction of 5.40 g of hydrogen with nitrogen?

2. In the production of sulfur dioxide, how many moles of oxygen are required to react with 160 g of sulfur?

3. In the electrolytic decomposition of water, 0.500 kmol of hydrogen is formed. What mass of oxygen was formed at the same time?

4. How many moles of ammonia are produced when 0.60 mol of nitrogen reacts with hydrogen?

V. Assignment

1. Powdered zinc metal reacts violently with sulphur (S_8) when heated. Predict the mass of sulphur required to react with 25 g of zinc.
2. Bauxite ore contains aluminium oxide, which is decomposed using electricity to produce aluminium metal. What mass of aluminium metal can be produced from 100 g of aluminium oxide?
3. Determine the mass of oxygen required to completely burn 10.0 g of propane.
4. Calculate the mass of lead (II) chloride precipitate produced when 2.57 g of sodium chloride in solution reacts in a double replacement reaction with excess aqueous lead (II) nitrate.
5. Predict the mass of hydrogen gas produced when 2.73 g of aluminium reacts in a single replacement reaction with excess sulphuric acid.
6. What mass of copper(II) hydroxide precipitate is produced by the reaction in solution of 2.67 g of potassium hydroxide with excess aqueous copper(II) nitrate?
7. 34.0 g of lithium chloride and some water are products of a neutralisation reaction. What masses of reactants were required?
8. Potassium reacts with the air to form its oxide. What mass of oxide is formed when 4.57 g of potassium undergoes this reaction?
9. Calculate the mass of ammonia produced by the reaction of 5.40 g of nitrogen with hydrogen.
10. Ethyne gas (C_2H_2) is produced by adding water to calcium carbide (CaC_2).
$$CaC_{2(s)} + 2 H_2O_{(l)} \rightarrow C_2H_{2(g)} + Ca(OH)_{2(s)}$$
 - A. What mass of ethyne is produced by adding water to 5.00 g of calcium carbide?
 - B. What mass of calcium carbide is required to react completely with 98.0 g of water?
11. A student obtained the following data from a reaction between a zinc strip and a copper (II) sulphate solution. What mass of copper should theoretically be produced?

initial mass of zinc before reaction	15.42 g
final mass of zinc after reaction	14.15 g
12. Living things use the combustion of carbohydrates like $C_6H_{12}O_6$ in the process of cellular respiration to produce energy. If 0.450 mol of oxygen is used, what mass of the carbohydrate will be burned?