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Stoichiometry of equations through the Inverse de Donder relation

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Abstract:

The stoichiometry of equations was revisited in light of the Law of Conservation of Matter at an atomic, elemental level. For a balanced chemical equation following the simplest general model $aA \rightarrow bB$, the fact that $\frac{a}{b} = \frac{n_A}{n_B}$, where n_A is the number of moles of A consumed and n_B is the number of moles of B produced in the reaction during the experiment, was used to address the de Donder relation, introduced by Theophile de Donder (1873–1957). While crediting the Belgian scientist for pointing out that “the reaction-ratio method” may be utilized for stoichiometry, the inverse de Donder relation was applied to problems in reaction stoichiometry. Several examples were used to show that the latter relation can be used to form proportions in order to rapidly solve such stoichiometry problems and to do so with fundamental chemical understanding. Educators in general chemistry were encouraged to teach the discussed method in their courses.

Keywords: de Donder, mass, mole, proportions, stoichiometry

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Introduction

The most fundamental chemical observation of the 18th century, stated by Antoine Lavoisier (1743–1794), is the law of mass conservation (Silberberg & Amateis, 2018), also called the Law of Conservation of Matter. In light of our current atomic view of matter, we may apply the Law of Conservation of Matter not only to the total mass of the reactants and products in a chemical reaction, but also to the mass of the atoms of each element in the reaction (not to mention the extremely small mass that is converted to energy or vice versa in a chemical reaction). In fact, if that were not true, we would have no reason to “balance” any chemical equation. The term “stoichiometry” was coined by Jeremias Benjamin Richter (1762–1807), the author of a book containing the rudiments of chemical calculations that appeared in Silesia in 1792 (Szabadváry & Oesper, 1962); however, what we call “stoichiometry of equations” today (that is, the study of the quantitative aspects of equations (Silberberg & Amateis, 2018)) heavily depends on the Law of Conservation of Matter at an atomic, elemental level.

Stoichiometry of equations is commonly covered in college-level introductory and general chemistry textbooks in the US today. According to DeToma (1994), in the article “Symbolic Algebra and Stoichiometry” published in Journal of Chemical Education, the 1994 trend in modern general chemistry textbooks was to treat basic problems in reaction stoichiometry as a part of a class “to be done exclusively by factor-label methods.” The current trend seems to be the same. In the factor-label procedure, in Tykodi’s words (1987), “the input information with a string of factors containing appropriate labels” is operated on “in such a way as to cause cancellation of all unwanted labels” so that “the sought-after quantity” is yielded in the end. Tykodi (1987) finds the factor-label procedure “excellent” for unit conversions, as this author does; however, his “final word of caution” (Tykodi, 1987) is not to “oversell” the idea of that method. This article is to increase chemical educators’ familiarity with an alternative method that this author finds more directly reflecting the Law of Conservation of Matter than the factor-label procedure in stoichiometry of equations and, therefore, more suitable for that purpose from a foundational perspective.

The de Donder relation

The simplest general model of a chemical equation is $A \rightarrow B$, and when the equation is balanced (that is, when it is written to reflect that the reaction is consistent with the Law of Conservation of Matter), it becomes $aA \rightarrow bB$.

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Since this balanced equation shows that for every a moles of A consumed, b moles of B are produced, it may be concluded that $\frac{a}{b} = \frac{n_A}{n_B}$, where n_A is the number of moles of A consumed and n_B is the number of moles of B produced in the reaction during the experiment. This proportion may be mathematically rearranged as the proportion $\frac{a}{n_A} = \frac{b}{n_B}$. The fractions on the two sides of the latter proportion may be inversed in order to yield the proportion $\frac{n_A}{a} = \frac{n_B}{b}$, which is a clear example of the de Donder relation (Tykodi, 1987), which is that for the general balanced equation $\alpha_1 A_1 + \alpha_2 A_2 + \dots \rightarrow \beta_1 B_1 + \beta_2 B_2 + \dots$, the numbers of moles of the reactants consumed and the products produced in the reaction during the experiment are related according to the following:

$$\frac{n_{A1}}{\alpha_1} = \frac{n_{A2}}{\alpha_2} = \dots = \frac{n_{B1}}{\beta_1} = \frac{n_{B2}}{\beta_2} = \dots$$

The Belgian scientist Theophile de Donder (1873–1957) is surely credited for pointing out that “the reaction-ratio method” may be utilized for stoichiometry (Tykodi, 1987). Still, the alternative stoichiometric method that this author would like to increase the readers’ familiarity with is based on an “educationally enhanced” version of the de Donder relation. It is important noting that this author’s basic introduction to this method was in Iran in 1990, through a chemistry teacher called Mr. Mohammadebrahim Adelimosayeb.

The inverse de Donder relation for stoichiometry of equations

It was shown that for the balanced equation $aA \rightarrow bB$, $\frac{a}{n_A} = \frac{b}{n_B}$, where n_A is the number of moles of A consumed and n_B is the number of moles of B produced in the reaction during the experiment. Proportions of this order, where the coefficient representing the number of moles of the reactant or product in the balanced equation is in the *numerator* and the number of moles of the reactant consumed or product produced in the reaction during the experiment is in the *denominator*, apply to the general balanced equation $\alpha_1 A_1 + \alpha_2 A_2 + \dots \rightarrow \beta_1 B_1 + \beta_2 B_2 + \dots$, and the “inverse” de Donder relation may be written as follows:

$$\frac{\alpha_1}{n_{A1}} = \frac{\alpha_2}{n_{A2}} = \dots = \frac{\beta_1}{n_{B1}} = \frac{\beta_2}{n_{B2}} = \dots$$

How is the inverse de Donder relation useful in stoichiometry of equations and why does this author find it educationally enhanced?

Any problem in reaction stoichiometry involves using a balanced chemical equation. It also “gives” (that is, somehow makes available) the amount (in mol or g) of a substance (reactant or product) in a chemical reaction and “asks for” (that is, requires calculating) the amount (in mol or g) of another substance (reactant or product) in that reaction. Therefore, any problem as such may be solved using a simple proportion made of two ratios that follow the inverse de Donde relation, where the “unknown” to solve for (by multiplying the components of the known diagonal and dividing the algebraic product by the third known part of the proportion) is often one of the denominators (however, it can be in one of the numerators). Each of the two ratios will include the coefficient of a reactant or product in the balanced equation in its numerator and the amount (in mol or g) of that substance in its denominator. This author’s personal experience suggests that the two ratios are most easily formed when they are clearly labeled with the chemical formula of their corresponding substance and when the proportion is formed line by line, with the first line formed based on the balanced equation. It is important noting that the coefficient of the substance, which will be in the numerator, must be multiplied by the molar mass (M) of the substance if the denominator is in g (rather than mol), keeping in mind that the mass of a substance in g is the algebraic product of its number of moles and its molar mass.

The following examples represent several types of chemistry problems in stoichiometry of equations and their solutions using the inverse de Donder relation. In order to demonstrate an evaluation of the use of this article *versus* the traditional approach (that is, the factor-label procedure), the examples were also given to a chosen student (see the Acknowledgments section), whose solutions are included as a supplementary document.

Examples

Example 1

How many moles of chlorine gas are consumed when 0.10 mol of hydrogen chloride gas is formed from hydrogen gas and chlorine gas directly combining, according to the following equation?

$$n_{\text{HCl}}(\text{produced}) = 0.10 \text{ mol}$$

$$n_{\text{Cl}_2}(\text{consumed}) = ?$$

$$\begin{array}{ccc} \text{Cl}_2 & & \text{HCl} \\ 1 \text{ mol (consumed)} & & 2 \text{ mol (produced)} \\ \hline n_{\text{Cl}_2}(\text{consumed}) & = & 0.10 \text{ mol (produced)} \end{array}$$

$$\Rightarrow n_{\text{Cl}_2}(\text{consumed}) = \frac{1 \text{ mol} \times 0.10 \text{ mol}}{2 \text{ mol}} = 0.050 \text{ mol}$$

Example 2

How many grams of sodium hydroxide are formed when 0.25 mol of hydrogen gas is formed from metallic sodium reacting with water, according to the following equation?

$$n_{\text{H}_2}(\text{produced}) = 0.25 \text{ mol}$$

$$m_{\text{NaOH}}(\text{produced}) = ?$$

$$FW_{\text{NaOH}} = (1 \times 22.99 \text{ amu}) + (1 \times 16.00 \text{ amu}) + (1 \times 1.008 \text{ amu}) = 39.998 \text{ amu}$$

$$\Rightarrow M_{\text{NaOH}} = 39.998 \text{ g/mol}$$

$$\begin{array}{ccc} \text{NaOH} & & \text{H}_2 \\ 2 \times 39.998 \text{ g (produced)} & & 1 \text{ mol (produced)} \\ \hline m_{\text{NaOH}}(\text{produced}) & = & 0.25 \text{ mol (produced)} \end{array}$$

$$\Rightarrow m_{\text{NaOH}}(\text{produced}) = \frac{2 \times 39.998 \text{ g} \times 0.25 \text{ mol}}{1 \text{ mol}} = 20 \text{ g}$$

Example 3

How many grams of metallic mercury are formed from 1.0 g of mercury(II) sulfide reacting with oxygen gas, according to the following equation?

$$m_{\text{HgS}}(\text{consumed}) = 1.0 \text{ g}$$

$$FW_{\text{HgS}} = (1 \times 200.6 \text{ amu}) + (1 \times 32.07 \text{ amu}) = 232.67 \text{ amu}$$

$$\Rightarrow M_{\text{HgS}} = 232.67 \text{ g/mol}$$

$$m_{\text{Hg}}(\text{produced}) = ?$$

$$AW_{\text{Hg}} = 200.6 \text{ amu}$$

$$\Rightarrow M_{\text{Hg}} = 200.6 \text{ g/mol}$$

$$\begin{array}{ccc} \text{HgS} & & \text{Hg} \\ 1 \times 232.67 \text{ g (consumed)} & & 1 \times 200.6 \text{ g (produced)} \\ \hline 1.0 \text{ g (consumed)} & = & m_{\text{Hg}}(\text{produced}) \end{array}$$

$$\Rightarrow m_{\text{Hg}}(\text{produced}) = \frac{1.0 \text{ g} \times 1 \times 200.6 \text{ g}}{1 \times 232.67 \text{ g}} = 0.86 \text{ g}$$

Example 4

A 6.5-g sample of a metal (X) in its elemental form is dissolved in sulfuric acid, according to the following equation, and 0.10 mol of hydrogen gas is formed:

What is the atomic weight of the metal?

$$\begin{aligned} m_{\text{X}}(\text{consumed}) &= 6.5 \text{ g} \\ n_{\text{H}_2}(\text{produced}) &= 0.10 \text{ mol} \\ AW_{\text{X}} &= ? \end{aligned}$$

$$\begin{array}{ccc} \text{X} & & \text{H}_2 \\ 1 \times M \text{ g (consumed)} & & 1 \text{ mol (produced)} \\ \hline 6.5 \text{ g (consumed)} & = & 0.10 \text{ mol (produced)} \end{array}$$

$$\begin{aligned} \Rightarrow 1 \times M \text{ g (consumed)} &= \frac{6.5 \text{ g} \times 1 \text{ mol}}{0.10 \text{ mol}} \\ \Rightarrow M &= \frac{6.5 \text{ g} \times 1 \text{ mol}}{0.10 \text{ mol} \times 1 \text{ g}} = 65 \\ \Rightarrow AW_{\text{X}} &= 65 \text{ amu} \end{aligned}$$

Example 5

What volume of a 0.20 M solution of barium hydroxide is needed to precipitate all the sulfate that is present in 50.0 mL of a 1.0 M solution of sodium sulfate?

The balanced equation is as follows:

$$C_{\text{Ba(OH)}_2} = 0.20 \text{ M}$$

$$V_{\text{Na}_2\text{SO}_4 \text{ solution}} = 50.0 \text{ mL} = 0.0500 \text{ L}$$

$$C_{\text{Na}_2\text{SO}_4} = 1.0 \text{ M}$$

$$V_{\text{Ba(OH)}_2 \text{ solution}} = ?$$

$$\begin{array}{ccc}
 \text{Ba(OH)}_2 & & \text{Na}_2\text{SO}_4 \\
 1 \text{ mol (consumed)} & & 1 \text{ mol (consumed)} \\
 \hline
 0.20 \text{ M} \times V_{\text{Ba(OH)}_2 \text{ solution (consumed)}} & = & 1.0 \text{ M} \times 0.0500 \text{ L (consumed)}
 \end{array}$$

$$\begin{aligned}
 \Rightarrow 0.20 \text{ M} \times V_{\text{Ba(OH)}_2 \text{ solution (consumed)}} &= \frac{1 \text{ mol} \times 1.0 \text{ M} \times 0.0500 \text{ L}}{1 \text{ mol}} \\
 \Rightarrow V_{\text{Ba(OH)}_2 \text{ solution}} &= \frac{1 \text{ mol} \times 1.0 \text{ M} \times 0.0500 \text{ L}}{1 \text{ mol} \times 0.20 \text{ M}} = 0.25 \text{ L}
 \end{aligned}$$

A pedagogical comparison

Comparisons between using the factor-label procedure and using the inverse de Donder relation are of pedagogical importance. In this article, the above examples (Examples 1–5) were given to a chosen student in General Chemistry I (see the Acknowledgments section) who solved them using the factor-label procedure (except Example 5, which the student solved using a formula). The student's solutions are available as the Supplementary Material accompanying this article. When the author compares the solutions given in the article to the student's solutions, the following points are observed:

1. Due to the use of proportions, the author finds using the inverse de Donder relation more reflective of scientific understanding than using the factor-label procedure.
2. In some cases, such as those of mass-mass calculations, using the inverse de Donder relation requires fewer steps than using the factor-label procedure (see Example 3).
3. The student's solution to Example 5 involves a formula that is only valid if the number of moles of one reactant consumed is equal to the number of moles of another reactant consumed. That approach not only requires memorization, but also is not useful if the numbers of moles of the two reactants consumed are unequal.

Conclusion

The inverse de Donder relation is a very useful method for solving problems in reaction stoichiometry. The examples presented in this article show the systematic nature of the method and how fast it can lead to the final answer when it is properly understood. This author hopes that this article opens a new avenue to general chemistry educators in teaching the stoichiometry of equations.

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