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Do the equilibrium constants have units? A discussion on how general chemistry textbooks calculate and report the equilibrium constants

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Abstract

An analysis of the literature concerning the dimensionality of equilibrium constants reveals that this topic manifests as a controversial issue. Based on this previous examination, this work studies if general chemistry textbooks accurately define and calculate equilibrium constants. In order to evaluate those textbooks, in the first part of this study the experimental equilibrium constants, K_p and K_c, and the thermodynamic equilibrium constant, K^o, are defined. Also, the equations that relate each constant to the other two are given. In the specific presentation of these quantities, an example is discussed both performing their accurate calculation and reporting them using the proper units. In the second part of this study, it is examined the way both first-year university chemistry textbooks and pre-university chemistry textbooks determine equilibrium constants, concentrating on how they handle the units of these quantities. Many textbooks treat K_p and K_c as dimensionless quantities. This misleading assumption is caused by a problem in the terminology used as in many cases K_p (or K_c) plays the role K^o. In order to avoid this misleading treatment of the equilibrium constants some suggestions are provided.

Keywords

Equilibrium constants, General chemistry text books, Definition, Calculation and Units

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State of the literature

- An important subject of science textbook research has been the detection of possible inconsistencies and erroneous treatments that may exist.
- Educational studies suggest that one of the sources of students' learning difficulties in physical chemistry lies in how textbooks and teachers deal with key chemistry concepts.
- The discussion concerning the dimensionality of equilibrium constants has revealed to be controversial.

Contribution of this paper to the literature

- This work strives for clarifying how to properly calculate and report each of the equilibrium constants.
- This study reports the main misrepresentations and problematic issues concerning the treatment of equilibrium constants performed by general chemistry textbooks.
- Several suggestions are provided in order to avoid current textbook misrepresentations.

Introduction

Purpose

The equilibrium constant is a basic and important thermodynamic concept that is usually introduced in pre-university chemistry courses. This concept has revealed to be controversial as some authors (Abrantes & Nieto de Castro, 1985; Boggs, 1958; Centellas, 2008; Cox, Dixon, Morris & Roberts, 1979; Harris, 1978, 1982; Gil & Paiva, 1999; Gordus, 1991; Quintero, 1987; Ronneau, 1993; Rosenberg & Kotz, 1999; Tykodi, 1986)state that equilibrium constants are unitless quantities while other studies (Antonik, 1993; Delorme, 1985 ; Depovere & Weiler, 1993a, 1993b; Henry, 1967;Laidler, 1990; Mills, 1989, 1995; Molyneux, 1991; Pethybridge & Mills, 1979; Treptow, 1999; Vickerman, 1979; Wright, 1979) discuss that the experimental equilibrium constants, viz. K_p and K_c, must be expressed with units, whereas the thermodynamic equilibrium constant, K^o, is a dimensionless quantity. Thus, the existence of these two opposed views is not of a minor importance as students are regularly requested by their teachers to be careful when reporting quantities as they must be expressed using their corresponding correct units. The clarification of this issue should be resolved in order to avoid confusion on students (Quílez, 2016; Quílez-Díaz & Quílez-Pardo-2015).

The aim of this study is to describe how chemistry textbooks report equilibrium constants. I hypothesized that the aforementioned controversy had permeated textbook presentation and calculation of equilibrium constants. Thus, I considered that many textbooks would be treating experimental equilibrium constants as unitless quantities.

The discussion that follows consists of two main parts. The first one is based on initial examinations dealing with the term standard in thermodynamics (Cox, 1982; Freeman, 1985; Treptow, 1999) and focuses on how to define and report the equilibrium constants. This analysis establishes that the aforementioned controversial subject matter is essentially a terminological problem. In order to avoid erroneous treatments, this work has tried to provide a clear presentation of each of the quantities involved and has exemplified how to report them. Taking into account this initial theoretical foundation, this educational research inspects in its second part how both pre-university chemistry textbooks and first-year university chemistry textbooks deal with equilibrium constants.

Equilibrium constants: definition, calculation and units

The aim of this section is to present a clear differentiation on two related quantities: the thermodynamic equilibrium constant, K^o, and the experimental equilibrium constants, K_p and K_c. In this regard, this article tries to provide a proper level of understanding of the implications of how different units lead to different numerical values of the equilibrium constants for a particular reaction.

Following the IUPAC recommendation (Ewing, Lilley, Oloffson, Rätzsch & Somsen, 1994; Mills, Cvitas, Homann, Kallay &Kuchitsu, 1993)and the equilibrium constant expressions provided in

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several educational papers (Borge, 2015; Quílez, 2012, 2016), in **Boxes 1** and **2** are summarized the equations that define each quantity, as well as their mathematical relationships. The equations reported in those boxes assume ideal equilibrium behaviour of the species involved.

Box 1. Glossary of equilibrium constant terms and their units for a given chemical equilibrium represented as:a A(g) + b B(g) $\rightleftharpoons$ c C(g) + d D(g).

Practical equilibrium constants, K_p and K_c:

$$K_p = \left\{ \frac{(p_C)^c (p_D)^d}{(p_A)^a (p_B)^b} \right\}_{eq}; \quad (1)$$

its units are (unit of pressure)$^{\Delta n(g)}$

$$K_c = \left(\frac{[C]^c [D]^d}{[A]^a [B]^b} \right)_{eq}; \quad (2)$$

its units are (unit of concentration)Δn(g)

Mathematical equation relating K_p to K_c:

$$K_p = K_c (RT)^{\Delta n(g)} \quad (3)$$

Thermodynamic equilibrium constant, K^o (unitless quantity):

$$K^o = \left\{ \frac{(p_C / p^o)^c (p_D / p^o)^d}{(p_A / p^o)^a (p_B / p^o)^b} \right\}_{eq}; \quad (4)$$

p^o = 1 bar

$$\Delta\mu^0 = \Delta_r G^0 = -RT \ln\left\{ \frac{(p_C / p^o)^c (p_D / p^o)^d}{(p_A / p^o)^a (p_B / p^o)^b} \right\}_{eq} = -RT \ln K^o \quad (5)$$

$$K^o = \exp[-\Delta_r G^o / RT] \quad (6)$$

Mathematical equations relating K_p/K_c to K^o:

$$K_p = K^o (p^o)^{\Delta n(g)} \quad (7a)$$

$$K_c = K^0 \left(\frac{p^0}{RT} \right)^{\Delta n(g)} \quad (7b)$$

p^o = 1 bar

Units of K_p:

$$K_p = K^o \,(\text{bar})^{\Delta n(g)} = K^o \left(\frac{1}{1.01325}\,\text{atm} \right)^{\Delta n(g)} = K^o \,(10^5\ \text{Pa})^{\Delta n(g)} \quad (8)$$

Box 2. Glossary of equilibrium constant terms and their units for a given chemical equilibrium represented as:a A(aq) + b B(aq) $\rightleftharpoons$ c C(aq) + d D(aq).

Practical equilibrium constant K_c:

$$K_c = \left(\frac{[C]^c [D]^d}{[A]^a [B]^b} \right)_{eq}; \quad (9)$$

its units are (unit of concentration)$^{\Delta n}$

Thermodynamic equilibrium constant, K^o (unitless quantity):

$$K^o = \left\{ \frac{(c_C / c^o)^c (c_D / c^o)^d}{(c_A / c^o)^a (c_B / c^o)^b} \right\}_{eq}; \quad (10)$$

c^o = 1 mol L^{-1}

$$\Delta\mu^0 = \Delta_r G^0 = -RTln \left\{ \frac{\left({}^{c_C}/_{c^0}\right)^c \left({}^{c_D}/_{c^0}\right)^d}{\left({}^{c_A}/_{c^0}\right)^a \left({}^{c_B}/_{c^0}\right)^b} \right\}_{eq} = -RTlnK^0 \quad (11)$$

$$K^o = \exp[-\Delta_r G^o / RT] \quad (12)$$

Mathematical equation relating K_c to K^o:

$$K_c = K^o (c^o)^{\Delta n};\ c^o = 1\ \text{mol L}^{-1} \quad (13)$$

Practical gas phase equilibrium examples discussing equations grouped in **Box 1** may help students in the differentiation between both the experimental equilibrium constants, K_p and K_c, and the thermodynamic constant, K^o, which also may allow them to perform a sound treatment concerning their units. This analysis can be exemplified calculating the values of K^o, K_p and K_c for chemical equilibrium systems (Quílez-Díaz & Quílez Pardo, 2015).In the discussion that follows, these calculations are performed for the following equilibrium (Quílez, 2016):

$$2\ NO_2(g) \rightleftharpoons N_2O_4(g)$$

In this particular case, $\Delta_r G^o$ = - 4,728.7 J mol^{-1}. Then, from equation (6) we obtain:

$K^o = \exp[-\Delta_r G^o/RT] = 6.7364$; T = 298.15 K.

Thus, considering that p^o = 1 bar, K_p can be easily calculated by using equation (7a), as $\Delta n(g) = -1$, $K_p = K^o(p^o)^{-1}$= 6.7364 bar^{-1}. Moreover, as 1 atm = 1.01325 bar, other values of K_p can be reported considering equation (8): K_p= 6.8257 atm^{-1} and K_p= 6.7364 × 10^{-5} Pa^{-1}. Notice that the values of K^o and K_p are the same when K_p is expressed in bar units. Finally, using equation (3),

$K_c = K_p(RT)$ = 166.99 (mol L^{-1})$^{-1}$;T = 298.15 K.

But K_p can also be calculated employing equation (1), instead of equation (7a), which means to determine previously the corresponding equilibrium values of partial pressures. This case can be exemplified using the data (T = 298.15 K) reported on a previous article (Raff, 2014): $p(NO_2)_{eq}$ = 0.34987 bar; $p(N_2O_4)_{eq}$ = 0.82507 bar. In the equilibrium we are exemplifying, we can express K_p as follows:

$$K_p = \left[\frac{p(N_2O_4)}{p(NO_2)^2}\right]_{eq} \tag{14}$$

Then, K_p = [0.82507 bar/(0.34987 bar)^2] = 6.7403 bar^{-1} (it is stressed that units must be reported). Notice that this value is slightly different from the numerical quantity reported when using equation (7a).That is, the calculation of K_p from the thermodynamic equilibrium constant, K^o, produces a value that differs from the one obtained from equation (1), where experimental values of partial pressures are considered. Those unequal values for K_p calculated using distinct equations are due to deviations from the ideal behaviour of the gaseous equilibrium mixture.

Box3 summarises the expressions corresponding to each equilibrium constant and their corresponding values calculated for the equilibrium 2 $NO_2(g) \rightleftharpoons N_2O_4$ (g); T = 298.15 K.

Box 3. Equilibrium constant equations and their corresponding values for the equilibrium 2 $NO_2(g) \rightleftharpoons N_2O_4$ (g); T = 298.15 K.

2 $NO_2(g) \rightleftharpoons N_2O_4$ (g); T = 298.15 K				
$K^o = \exp[-\Delta_r G^o/RT]$ (p^o = 1 bar)	$K_p = K^o(p^o)^{-1}$ (method A) $K_p = \left[\frac{p(N_2O_4)}{p(NO_2)^2}\right]_{eq}$ (method B)			$K_c = K_p(RT)$ $K_c = K^0\left(\frac{p^0}{RT}\right)^{-1}$
6.7364	(A) 6.7364bar^{-1} (B) 6.7403 bar^{-1}	(A) 6.8257 atm^{-1}	(A) 6.7364× 10^{-5} Pa^{-1}	(A) 166.99 (mol L^{-1})$^{-1}$

We have calculated one value for K^o (p^o = 1 bar), which is dimensionless (K^o= 6.7364). Using equations (7a) and (8), this value has allowed to calculate three different values for K_p, depending on the units of pressure used (K_p= 6.7364 bar^{-1}, K_p= 6.8257 atm^{-1}and K_p= 6.7364 × 10^{-5} Pa^{-1}). From equation (14), we have calculated a different bar unit value for this quantity: K_p= 6.7403 bar^{-1}.

Analogous examples can be performed in cases involving aqueous solution equilibria. In these situations, we must employ the equations summarized in **Box 2**. Concerning equilibrium constants, an analogous discussion as the one performed in the case of a gas phase equilibrium can be carried out. That is, when applying equation (13), we must note that as we have taken c^o = 1 mol/L, only if the units of K_c are (mol/L)$^{\Delta n}$, does its value equal that of K^o. However, it must be stressed that although K_c and K^o may yield the same value, it does not mean that there is only one quantity. In this particular case, two identical calculated values correspond to two different quantities. K^o is a unitless quantity and in this circumstance, K_c has the dimensions of mol/L raised to the power of Δn. But, if other concentration units are employed, a different value for K_c is obtained. For example, at 298.15 K, for acetic acid (Pethybridge & Mills, 1979): K^o = 1.751·10^{-5} (c^o = 1 mol L^{-1}); K_c = 1.751·10^{-5} molL^{-1} and K_c = 1.756·10^{-5} molkg^{-1}.

Presentation and calculation of equilibrium constants in general chemistry textbooks

Sample

The sample of textbooks studied consists of two groups (Annex):
a) Grade-12 chemistry textbooks (N = 45);
b) first-year university chemistry textbooks (N = 39).

The sample of Grade-12 chemistry textbooks covers an ample collection of textbooks through a large span of time (more than thirty years, including the recent editions corresponding to the most important publishers of pre-university chemistry textbooks in the Spanish context). Moreover, two international textbooks are part of this sample as there is an increasing number of Spanish chemistry students that chose to study this subject in English. Also, several textbooks devoted to solving chemical equilibrium exercises have been taken into consideration.

On the other hand, the sample of university chemistry textbooks covers a wide range of books, which have been selected considering that they are well known by a great audience as most of them have been translated into different languages. Particularly, most of the books in the sample have been translated into Spanish and can be found in the libraries of Spanish universities. Also, in many cases, they have been subjected to several editions, which represents a signal of their approval by many chemistry teachers. The selected books correspond to editions made during the last thirty years. Eventually, they have been chosen for similar examinations in previous educational research papers. As in both samples the span of time covered for textbook editions is large, it may help to analyse if the terminological issue examined in this study is a long-standing problem.

Research issues

Considering that general chemistry textbooks may not define equilibrium constants properly and even may confuse these quantities, this study has focused on trying to find out if textbooks:

1) define each of the equilibrium constants (i.e. K_p, K_c and K^o);
2) report K_p and K_c using the appropriate units;
3) calculate K^o as a unitless quantity;
4) present and use the mathematical equation relating K_p to K^o;
5) assume that K_p plays the role of K^o.

That is, those items have been used to guide the examination on how chemistry textbooks define, calculate and report equilibrium constants. Particularly, this research aims to establish if there is confusion when naming and determining both K_p and K^o. In this regard, this work also studies if units are used properly in each case and if textbooks differentiate those quantities, establishing the corresponding mathematical relationship between them.

Results and discussion

Keeping in mind the aforementioned research issues, this analysis has concentrated on finding textbook inconsistencies and erroneous treatments concerning the definition and calculation of chemical equilibrium constants. The major results obtained in this examination are provided in **Table 1**.

As there are some differences in the results obtained for each of the two levels analysed, the discussion that follows firstly concentrates on the examination of Grade-12 chemistry textbooks and later on the particular cases for first-year chemistry textbooks are discussed.

In the analysis of each sample, the main features describing how textbooks deal with equilibrium constants will be presented. Several explicit references will be provided when some particular significant cases are worth to be mentioned.

Grade-12 textbooks

Spanish Grade-12 chemistry textbooks introduce, without exception, the experimental equilibrium constants K_p and K_c, as in equations (1) and (2).They also present equation (3). In 56%of the sample those experimental constants are reported as dimensionless quantities in two situations: a) when determining their values after introducing experimental data in equations (1) and (2) and b) when using equation (3)for the calculation of one of the two constants, knowing the value of the other one.

In one textbook (Barrio, Sánchez, Bárcena & Caamaño, 2016) K_p is reported as a unitless quantity, but K_c does have units of (mol $L^{-1})^{(\Delta n)}$. Among these textbooks that do not report the value of $K_p$with units,there is not usually an explanation stating why they treat K_p as dimensionless quantity. In three textbooks (Cardona, Pozas, Martín &Ruíz, 2003; García & García-Serna, 2003; Morcillo, Fernández, Carrión, 1998)it is warned that in spite that pressure values must be reported in atm and concentration values in mol L^{-1}, both K_p and K_c are considered dimensionless quantities. A similar statement was found in a book (Climent, Domingo, Latre, Sanz, Silla, Soler, & Viché, 1989)aimed atinstructing the essential chemistry topics to Grade-12 Spanish students in their preparation for the final official chemistry exam to access university. Additionally, in two textbooks (Carriedo, Fernández & García, 2016; Illana, Araque, Liébana & Teijón, 2016)it is mentioned that both K_p and K_c should have units but in ideal conditions both constants are expressed in terms of activities. Then, it is stated that as these quantities are dimensionless, thus both constants can be reported as unitless quantities. Consequently, those textbooks that treat K_p and K_c as unitless quantities do not care about the units of the *RT* product when using equation (3).

Several additional cases have been found where both K_p and K_c are treated as dimensionless quantities; particularly, five books of problems for Grade-12 chemistry students (Domingo & Silla, 1991; Fidalgo & Fernández, 2005; Goldberg, 2001; Guardia, Menéndez & de Prada, 2009; López, 2000) and even a doctoral dissertation (Moncaleano, 2008) aimed at the pedagogical study of chemical equilibrium. Furthermore, the same case can be found in a series of textbooks (Fernández, 2016; Morcillo & Fernández, 1991; Zubiaurre & Arsuaga, 2007) that for more than twenty-five years have been collecting and solving the problems and questions corresponding to the final official exams for Grade-12 Chemistry made by the Spanish universities. Thus, as official exams are key references for Chemistry teachers and their students, this last finding may have contributed to not overcoming the reported confusion concerning the units of the equilibrium constants.

When units are reported, the common units for $K_c$are $(\mathrm{mol/L})^{\Delta n}$. The usual units for$K_p$ are $(\mathrm{atm})^{\Delta n}$; only in four textbooks(9%)the units of K_p are also $(\mathrm{Pa})^{\Delta n}$.There is no explicit discussion in any of the books in the sample on the units of K_c and K_p.

Concerning the thermodynamic equilibrium constant, K^o, this quantity is normally not defined (89 %). However, in one textbook (Brown & Ford, 2014) it is defined as K_c. Conversely, K^o is accurately introduced in only four textbooks (9 %)of the sample analysed. In three textbooks (Andrés, Antón, Barrio, de la Cruz & González, 2000; Caamaño & Obach, 2000; Caamaño, Obach & Servent, 1991) their authors state that the standard pressure is 1 atm and readers are explicitly warned that K^o is a unitless quantity. In their discussion, they state that as p^o = 1 atm, the values of K^o and K_p are identical when partial pressures are measured in atm. Only one textbook (Simón, García-Serna, & Romero, 2016) was found where both K^o is defined and the given standard state for gases is p^o = 1 bar.

Regarding the calculation of the thermodynamic equilibrium constant, K^o, in many cases (78 %) the following equation is introduced,

$$\Delta G^o = -RT \ln K_p \,, \qquad (15)$$

assuming, consequently, that K_p is the thermodynamic equilibrium constant. The values textbooks report for K_p in this situation usually have no units, but they do not explain why the obtained figure is presented without units. Thus, in some textbooks K_p is firstly defined as an experimental quantity and then reported having units of pressure, but later on it is calculated as a unitless quantity when it is assumed to play the role of the thermodynamic equilibrium constant, as presented in equation (15). In one textbook (Pozas, Martín, Rodríguez, Ruiz & Vasco, 2016) their authors use in equation (15)the previous calculated value of K_p, having units of pressure (i.e. $(\mathrm{atm})^{\Delta n}$); that is, the experimental value obtained using equation (1) is now reported without units and then introduced in equation (15) in order to calculate ΔG^o. However, in two textbooks (Barrio & Belmonte, 1992; Belmonte, 1999) the reported value for K_p, after calculating this quantity as K_p= exp[$-\Delta G^o/RT$], has atm units. This is also the case of two books of problems (García, Lozano & Teijón, 1991; Fernández & Pérez, 2004) aimed at preparing Grade-12 students for their official Chemistry exam for entering university.

Table1. Main misrepresentations and problematic issues concerning the treatment of equilibrium constants performed by general chemistry textbooks.

Research issue	Grade-12 textbooks	First-year university textbooks
Definition of the equilibrium constants (i.e. K_p, K_c and K^o).	1) K_p and K_c are defined (100 %) 2) The thermodynamic equilibrium constant (i.e.K^o)is not explicitly defined (89 %).	1) K_p and K_c are defined (100 %). 2) The thermodynamic equilibrium constant (i.e.K^o)is not explicitly defined (92 %).
Appropriate units for K_p and K_c and K^o as a unitless quantity.	1)K_p and K_c are reported as unitless quantities (56 %). - There is not a clarification explaining why the authors of textbooks have treated the experimental equilibrium constants (i.e. K_p and K_c) as dimensionless quantities (53 %). - When using equation $K_p = K_c(RT)^{\Delta n(g)}$, it is assumed unitless numbers for both K_p and K_c and, nothing is said about the units of the product RT (56 %). 2) When K_p is reported having units, authors usually employ atm for partial pressures, neglecting other pressure units such as bar or Pa (91 %). 3) K^o is defined as a unitless quantity (9%). 4) Only one textbook refers to the current standard pressure value: $p^o = 1$ bar.	1) The experimental equilibrium constants (i.e. K_p and K_c) are reported as unitless quantities (74 %): - There is not clarification explaining why the authors of textbooks have treated the experimental equilibrium constants (i.e. K_p and K_c) as dimensionless quantities (18 %). - Both K_p and K_c are unitless because there are more rigorous methods for expressing those quantities (33 %). - It is customary to report those quantities without units (18 %). - K_p is defined in terms of activities (5%). - When using equation $K_p = K_c(RT)^{\Delta n(g)}$, it is assumed unitless numbers for both K_p and K_c and, usually, nothing is said about the units of the product RT (74 %). 2) When K_p is reported having units, authors usually employ atm for partial pressures, neglecting other pressure units such as bar or Pa (92 %). 3) K^o is defined as a unitless quantity (8 %). 4) There is not reference to the current standard pressure value: $p^o = 1$ bar (92 %).
Presentation and use of the mathematical equation relating K_p(or K_c) to K^o.	They are not presented in any textbook.	Few textbooks present and use these equations (8 %).
Assumption that K_p plays the role of K^o.	1) It is assumed that K_p plays the role of the thermodynamic equilibrium constant (K^o) (78 %). 2) The following mathematical relationship is usually used $\Delta G^o = -RT \ln K_p$ (78 %). 3) There is not an exposition clarifying why when performing the calculation of K_p, using the above equation, it must be reported as a unitless quantity (78 %).	1) There is not explicit distinction between the thermodynamic equilibrium constant, K^o, and the practical or experimental equilibrium constants, K_p and K_c (92 %). 2) In many cases, it is explicitly stated that K^o means K_p/K_c(62 %). 3) It is used the following mathematical relationship $\Delta G^o = -RT \ln K_p$ or a similar one as it is assumed that the thermodynamic equilibrium constant is K_p(or K_c) (62 %).

Finally, it is worthy to note that equation (15) contains two errors: a) it uses ΔG^{o} instead of $\Delta_r G^{o}$ and b) it is assumed that $K_p = K^{o}$ (or $K_c = K^{o}$). In this work, we have concentrated on the second inaccuracy. The correct expression is equation (5) (Borge, 2015; Quílez, 2012). This expression was found in only three Grade-12 textbooks (8 %) of the sample analysed. Also, it should be noted that in this logarithmic relationship the units of $\Delta_r G^{o}$ are those of the product RT (i.e. J mol^{-1}) as the other part of equation (5) (i.e. lnK^{o}) represents the Napierian logarithm of a number (i. e. $K^{o}$is unitless), instead of the logarithm of a dimensioned quantity (e. g. K_p or K_c) (Boggs, 1958; Copley, 1958; Laidler, 1995) because dimensions are not carried at all in a logarithmic function (Matta, Massa, Gubskaya & Knoll, 2011).

The previous findings show that among the current Spanish Grade-12 Chemistry textbooks there is only one case (Simón, García-Serna, & Romero, 2016)that accurately both introduces and calculates the three equilibrium constants (i.e. K_p, K_c and K^{o}). Moreover, this study has shown that during the last thirty years, few textbooks have properly defined and calculated those quantities. In more than half of the sample analysed, K_p and K_c are treated as dimensionless quantities and it is usually assumed that K_p (or K_c) plays the role of the thermodynamic equilibrium constant, K^{o}. Thus, the misleading way which Spanish Grade-12 Chemistry textbooks treat those quantities has revealed to be a long-standing problem when teaching chemical equilibrium to high school students.By the same token, complementary and referential material for the study of chemistry, such as books of problems and official chemistry exams to enter university, may have contributed to those inaccurate assumptions.

First-year textbooks

The examination carried out on first-year chemistry textbooks reveals that a small percentage (26%) express K_p and K_c having units. These cases mean that K_p usually has (atm)$^{\Delta n}$ units; Freemantle (1991)express $K_p$in (Pa)$^{\Delta n}$ units and Burrows (2017)report several examples in which K_p is measured in (Pa)$^{\Delta n}$,(atm)$^{\Delta n}$ and (bar)$^{\Delta n}$ units.

Among the textbooks that consider these quantities as dimensionless (74%), there is more diversity than in the case of Grade-12 textbooks. In a first group of seven textbooks(18 %)there is no mention to why they have omitted units. However, in one of them (Spencer, Bodner & Rickard, 2008) their authors warn that in spite that K_p is a dimensionless quantity, any calculation involving it requires that partial pressures must be reported in atm units. In a second group, textbooks (18 %) give explanations that may surprise their readers. For example, in one textbook it is stated the following "…the units of equilibrium constant can always be figured out from the equilibrium constant expression. Therefore, it is customary to omits the units, and we shall follow that custom here" (Moore, Stanitski, Wood, Kotz & Joesten, 1998, p. 574). Additionally, Zumdahl and Zumdahl (2007, p. 583) state:

> The reason for this is beyond the scope of this text, but it involves corrections for the nonideal behaviour of the substances taking part in the reaction. When these corrections are made, the units cancel out and the corrected K has no units. Thus, we will not use units for K in this text.

Finally, in a third group of textbooks (33 %), after presenting the general equations for both K_p and K_c, as in equations (1) and (2),and examples for the calculation of each of these two quantities, the reader is warned that the figures obtained are dimensionless as their authors make reference to a more rigorous treatment of the equilibrium constants in which these quantities are defined in terms of activities. They usually explain that each term in the equilibrium constant represents the ratio of the measured quantity of the substance (molar concentration or pressure) to the thermodynamic standard-state quantity of the substance, recalling that these standard states are mol L^{-1} for a substance in solution and 1 atm for gases. Thus, they conclude that although they consider experimental values of partial pressures in atm or concentrations in mol L^{-1}, as each term is understood to be a ratio of a concentration or partial pressure to a standard value, the units on each term cancel, which results in the value of K_p(or K_c)having no units. Moreover, when using equation (3) ($K_p = K_c (RT)^{\Delta n(g)}$), they assume unitless numbers for both K_p and K_c and, usually, nothing is said about the units of the product RT. However, in one textbook (Jespersen, Brady & Hyslop, 2012) the reader is advised that the following R = 0.0821 L atm mol^{-1}K^{-1} is the only value of R that can be used in equation (3). Additionally, a similar framework is embodied in two textbooks (Petrucci, Harwood, Herring & Madura, 2006; Atkins & Jones, 2010) as their authors define only one equilibrium constant in terms of activities although they name it as K_p.

In an advanced thermodynamic section of the topic chemical equilibrium all the textbooks analysed introduce an equation that usually has this customary presentation,

$$\Delta G^{o} = -RT \ln K \tag{16}$$

The main textbook inaccuracy is that their authors assume explicitly (62 %) that the calculation of K in equation (16) means to calculate K_p (and also K_c and K_{sp}). That is, in several textbooks (21 %) it is understood that the equilibrium constant that appears in the above equation is actually K_p. This assumption prompts that in some textbooks (10%), K_p is written in the above equation (16)

$$\Delta G^{\circ} = -RTln\left\{\frac{(p_C)^c(p_D)^d}{(p_A)^a(p_B)^b}\right\}_{eq} = -RTlnK_p \tag{17}$$

Similarly, in another group of textbooks (26 %) their authors state that $K = K_p$ in the case of gases (p^{o} = 1 atm) and $K = K_c$ in the case of aqueous diluted solutions (c^{o} = 1 mol/L). Additionally, in several textbooks (8 %),equation(16) is used to calculate K_{sp}.

Thus, the symbol K_p (and also K_c and K_{sp}) is misleadingly used to denote the thermodynamic equilibrium constant, K^{o}. Also, one may infer that the value of K_p introduced in equation (17) has units of pressure. Hence, equation (17) may lead to incorrect assumptions because it hides the fact that the argument of the logarithm is dimensionless. Similar to the above discussion on equation (15), equations (5) and (17) are not the same as K_p is a dimensioned quantity, but logarithmic quantities are unitless. Thus, one must remark that in order to explicitly make dimensionless the thermodynamic equilibrium constant, K^{o} (instead of K_p), equation (17) must accurately be rewritten introducing the standard pressure on it (Borge, 2015), as has been done in equation (5).

This misleading assumption in which the thermodynamic equilibrium constant, K^{o}, is represented by the symbol K_p may derive from a careless neglect of the units in equation (5) as p^{o} = 1 bar (1 atm for most textbooks) and then the p^{o} values may be considered unnecessary (because units cancel out) and thus the denominators are omitted (Zumdahl & DeCoste, 2017) which certainly may confuse our students as they may find (and use) mathematical expressions containing logarithms of quantities with units.

The previous examination shows that in few freshman textbooks (Burrows, Holman, Parsons, Pilling & Price, 2017; Oxtoby, Gillis & Campion, 2012; Petrucci, Herring, Madura & Bissonnette, 2017) the three equilibrium constants (i.e. K_p, K_c and K^{o}) are accurately defined. Concerning their units and the calculation of K_p (or K_c) using equations (7a) and (13), in only two textbooks (Burrows, Holman, Parsons, Pilling & Price, 2017; Petrucci, Herring, Madura & Bissonnette, 2017) there is a clear discussion on this issue. Thus, in three quarters of the sample analysed, K_p and K_c are treated as dimensionless quantities and it is usually assumed that K_p (or K_c) plays the role of the thermodynamic equilibrium constant, K^{o}. Hence, as in the case of pre-university textbooks, the misleading way which first-year textbooks treat those quantities has revealed to be a long-standing problem.

In the study that corresponds to the dimensionality of K^{o} in equation (16), all textbooks, except one case (Atkins, 1989),treat the thermodynamic equilibrium constant as a unitless quantity, although in a 15 % of the sample there is not any explanation stating why this quantity is dimensionless. In a textbook of chemistry problems of this university level (Adamson, 1975) there is another case in which K^{o} appears having units. Finally, only in three textbooks (8 %) their authors state that the standard value of pressure is p^{o} = 1 bar; thus, the majority of textbooks take for this value p^{o} = 1 atm.

As a final comment, an inspection of some advanced textbooks reveals that their treatment of the quantities analysed in this study is quite more accurate than in the case of first-year textbooks. However, confusing equation (17) was found in a higher-level textbook (Raff, 2001). Moreover, in some well-known physical chemistry textbooks (Chang, 2000; Engel & Reid, 2012; Levine, 2009; McQuarrie & Simon, 1997; Mortimer, 2000;Vemulapalli, 1993) the thermodynamic equilibrium constant is named using the K symbol keeping the subscript p on it, which may not help to overcome the confusion between these two different quantities, K^{o} and K_p. Conversely, this terminological association is not present in two textbooks of this type (Atkins & de Paula, 2010; Silbey, Alberty & Bawendi, 2005). Eventually, a clear presentation of each equilibrium constant can be found in two advanced thermodynamic textbooks (Brenon-Audat, Busquet & Mesnil, 1993; DeVoe, 2016).

Concluding remarks, final comments and educational suggestions

The first part of this article has provided a brief summary concerning the definition and calculation of the different equilibrium constants (i.e. K_p, K_c and K^{o}).This analysis has also presented an example in order to aid in the discussion involving the calculation and proper units of those quantities. Particularly, the thermodynamic equilibrium constant, K^{o}, is a unitless quantity because it is calculated in reference to specific standard states (e.g. 1 bar or 1mol L^{-1}) (i.e. equations (4) and (10)). These reference states should be specified as they have calculational effects depending on which is used. Conversely, K_p has units of pressure (e.g. bar, atm, Pa, etc.) and K_c has units of concentration (e.g. mol L^{-1}, mol kg^{-1}). The terminological examination of these three different quantities has served to base the study on how chemistry textbooks deal with equilibrium constants.

Concerning how chemistry textbooks define and report equilibrium constants, the main findings of this study are the following:

1) Many textbooks present K_c and K_p as dimensionless quantities.
    a) Particularly, in more than half of the Grade-12 textbooks' sample and in three quarters of first-year textbooks their authors report experimental equilibrium constants without units.
    b) In several cases, first-year textbooks refer to a more rigorous method (i.e. activities) when defining K_c and K_p, assuming that units cancel out. In other instances, an explanation stating why K_c and K_p are reported as unitless quantities is not given or simply it is asserted that it is customary to omit units.
    c) Authors usually do not care about the units of the product RT when performing calculations using the equation $K_p = K_c(RT)^{\Delta n(g)}$.
2) When K_p is reported having units, authors usually employ atm as the unit for partial pressures, neglecting other pressure units such as bar or Pa.
3) The thermodynamic equilibrium constant, K^{o}, is normally not defined.
4) In most of the chemistry textbooks analysed there is not explicit distinction between both K^{o} and K_p; thus, this quantity usually plays the role of K^{o}.
5) The mathematical relationship between K^{o} and K_p is normally not given; a similar case was found concerning K^{o} and K_c.
6) Most textbooks still refer to the value 1 atm as the standard state pressure, thus few of them use the current value, p^{o} = 1 bar.
7) In the equation $\Delta_r G^{o} = -RT \ln K^{o}$, it is usually assumed that K^{o} is K_p; in several cases, that equation is written as $\Delta G^{o} = -RT \ln K_p$, reinforcing the assumption that K_p plays the role of K^{o}.
8) In the introduction and discussion of the meaning of equation $\Delta_r G^{o} = -RT \ln K^{o}$, it is not normally explained why K^{o} is a dimensionless quantity.

The cases discussed in this article demonstrate how deep the terminological misrepresentations on the equilibrium constants are rooted both in pre-university chemistry textbooks and in those corresponding to first-year level. This confusion in terminology leads to report K_p and K_c as dimensionless quantities. As most educators usually follow what is stated in textbooks, this misleading treatment of equilibrium constants may have also permeated the way how chemistry teachers deal with these quantities, which warns about the importance of possible re-education of teachers in order to guarantee that students learn these topics properly.

In order to avoid current misrepresentations, the thermodynamic equilibrium constant, K^o, should be defined as well as the practical equilibrium constants, K_p and K_c, stating the equations that relate both K^o to K_p and K^o to K_c. These equations may help in understanding why K^o is a unitless quantity whereas the experimental equilibrium constants, viz. K_p and K_c, must be expressed having units. In addition, this study has showed how different units lead to different numerical values for the practical equilibrium constants corresponding to a particular equilibrium equation. Also, it has discussed that although K^o and K_p (or K^o and K_c) may yield the same value (i.e. gas pressure values measured in bar; concentrations in mol L^{-1}), it does not mean that there is only one quantity, which first-year chemistry textbooks usually name K_p (or K_c).

Selected examples, as the one performed in this study, may allow students to accurately deal with equilibrium constants' calculation and units. Still, it is suggested a progression in this presentation as the correct treatment of the thermodynamic equilibrium constant, K^o, would be beyond the scope of a pre-university level as students of these introductory chemistry courses may be unable to understand the formalities, concepts and equations that allow the calculation of K^o when connecting it with $\Delta_r G^{\circ}$ using equation (6). Thus, Grade-12chemistry textbooks should solely define both K_p and K_c, reporting them with their proper units. That is, when solving chemical equilibrium problems, the performance of ICE tables should allow students the introduction of the corresponding values (and their units) in the corresponding equilibrium constant expression (i.e. equations (1) and (2)), which eventually could lead them to report the calculated constant as a dimensioned quantity. This case also applies when using equation (3) in the study of gaseous equilibrium reactions. Hence, the complete treatment discussed in this study, which involves the definition and calculation of the thermodynamic constant, K^o, as well as its relationship to K_p and K_c, would be left to first-year university chemistry textbooks. That is, in this college level, after reminding students the dimensioned experimental equilibrium constants (i.e. K_p and K_c), firstly introduced in their pre-university general chemistry course, the equilibrium constant, K^o, should be defined as a dimensionless quantity that can be calculated using equation (6). Additionally, K_p can be calculated from K^o employing equation (7a); similarly, K_c can be obtained from equation (7b) or equation (13), taking in mind that these expressions assume ideal behaviour of the species involved in the equilibrium mixture.

References

Abrantes, L.M. & Nieto de Castro, C. (1985). O conceito de constante de equilíbrio – suaintroduçao. *Boletim Sociedade Portuguesa de Química, 19*, 21-23.

Adamson, A. W. (1975). *Understanding Physical Chemistry*. New York: Benjamin.

Andrés, D. M., Antón, J. L., Barrio, J. I., de la Cruz, M. C. & González, F. (2000). *Química*. Madrid: Editex.

Antonik, S. (1993). Aspects dimensionnels de la constante d'équilbre en phase gazeuse. Usage des tables de thermodynamique. *Bulletin de l'Union des Physiciens, 87*, 909-917.

Atkins, P. W. (1989). *General Chemistry*. New York: Scientific American Books.

Atkins, P. & de Paula, J. (2010). *Atkins' physical chemistry*. Oxford: Oxford University Press.

Atkins, P. W. & Jones, L. L. (2010). *Chemical Principles. The Quest for Insight*. New York: Freeman and Company.

Barrio, J. I. & Belmonte, M. (1992). *Química COU*. Reacción. Madrid: SM.

Barrio, J. I., Sánchez, A., Bárcena, A. I. & Caamaño, A. (2016). *Química 2*. Madrid: SM.

Belmonte, M. (1999). *Química 2*. Barcelona: Ediciones del Serbal.

Bergqvist, A., Drechsler, M., De Jong, O. & Rundgren, S. N. C. (2013). Representations of chemical bonding models in school textbooks – help or hindrance for understanding, *Chemistry Education Research and Practice, 14*, 589 – 606. DOI: 10.1039/C3RP20159G

Boggs, J. E. (1958). The logarithm of "ten apples". *Journal of Chemical Education, 35*(1), 30-31. DOI: 10.1021/ed035p30

Borge, J. (2015). Reviewing some crucial concepts of Gibbs energy in chemical equilibrium using a computer-assisted, guided-problem-solving approach, *Journal of Chemical Education, 92*(2), 296–304. DOI: 10.1021/ed5005992

Brenon-Audat, F., Busquet, C. & Mesnil, C. (1993). *Thermodynamique Chimique*. Paris: Hachette.

Brown, C. & Ford, M. (2014). *Higher Level Chemistry*, 2nd ed.; Harlow: Pearson.

Burrows, A., Holman, J., Parsons, A., Pilling, G. & Price, G. (2017). *Chemistry. Introducing inorganic, organic and physical chemistry*. Oxford: Oxford University Press.

Caamaño, A. & Obach, D. (2000). *Química 2*. Barcelona: Teide.

Caamaño, A., Obach, D. & Servent, A. (1991). *Química COU*. Barcelona: Teide.

Cakmakci, G. (2009). Emerging issues from textbook analysis in the area of chemical kinetics. *Australian Journal of Education in Chemistry, 70*, 31-38.

Cardona, A. R., Pozas, A., Martín, R. & Ruíz, A. (2003). *Química 2*. Madrid:McGrawHill.

Carriedo, G. A., Fernández, J. M. & García, M. J. (2016). *Química 2*. Madrid: Paraninfo.

Castillo, J., Ogaz, R., Merino, C. & Quiroz, W. (2016). An ontological and epistemological analysis of the presentation of the first law of thermodynamics in school and university textbooks, *Chemistry Education Research and Practice, 17*, 1041-1053. DOI: 10.1039/C6RP00105J

Centellas, F. A. (2008). Algunesreflexions al voltant del tractament de l'equilibriquímic en el batxillerat. *Educació Química, 1*(1), 10-16.

Chang, R. (2000). *Physical chemistry for the chemical and biological sciences*. Susalito: University Science Books.

Climent, M. D., Domingo, R.,Latre, F., Sanz, V., Silla, E., Soler, V.& Viché, J. (1989). *Química de COU*. Una perspectiva práctica. Valencia: Cespusa.

Copley, G. N. (1958). Letter to the Editor. *Journal of Chemical Education, 35*(7), 366-367. DOI: 10.1021/ed035p366

Cox, J. D. (1982). Notation for states and processes, significance of the word standard in chemical thermodynamics, and remarks on commonly tabulated forms of thermodynamic functions. *Pure & Applied Chemistry, 54*, 1239–1250.

Cox, G. A., Dixon, D., Morris, R. & Roberts, J. P. (1979): Does K have units? *Education in Chemistry, 16*, 101.

Delorme, P. (1985). A propos des constants d'équilibre. *Bulletin de l'Union des Physiciens, 79*, 1087-1092.

Depovere, P.& Weiler, J. (1993a). Mise au point a propos du caractère dimensionné ou indimensionné des constants d'équilibre. *Association Belge des Professeurs de Physique et Chimie, 117*, 107-112.

Depovere, P. & Weiler, J. (1993b). Réflexions concernant "le dimensionnement ou le non dimensionnement" des constantes d'équilibre. *Réponse des auteurs, Association Belge des Professeurs de Physique et Chimie, 119*, 234.

DeVoe, H. (2016). *Thermodynamics and Chemistry*. www.chem.umd.edu/thermobook.

Domingo, R. & Silla, E. (1991). *Exercicisnumèrics en química*. València: Generalitat Valenciana.

Drechler, M. & Schmidt, H. J. (2005). Textbooks' and teachers' understanding of acid-base models in chemistry teaching. *Chemistry Education Research and Practice, 6*(1), 19-35. DOI: 10.1039/B4RP90002B

Eltinge, E. M. & Roberts, C. W. (1993). Linguistic content analysis: A method to measure science as inquiry in textbooks. *Journal of Research in Science Teaching, 30*(1), 65-83. https://doi.org/10.1002/tea.3660300106

Engel, T. & Reid, P. (2012). *Physical Chemistry*. Englewood Cliffs: Prentice Hall.

Ewing, M. B., Lilley, T. H., Oloffson, G. M., Rätzsch, M. T. &Somsen, G. (1994). Standard quantities in chemical thermodynamics. *Pure & Applied Chemistry, 66*(3), 533-552.

Fernández, E. (2016). Selectividad 2015. *Pruebas de acceso a la Universidad: Química*. Madrid: Anaya.

Fernández, A. & Pérez, C. (2004). *Química Schaum bachillerato*. Madrid: McGrawHill.

Fidalgo, J. A. & Fernández, M. R. (2005). *Problemas de Química. 2º de Bachillerato*. León: Everest.

Freeman, R. D. (1985). Conversion of standard thermodynamic data to the new standard-state pressure. *Journal of Chemical Education, 62*(8), 681-686. DOI: 10.1021/ed062p681

Freemantle, M. (1991). *Chemistry in Action*. London: MacMillan.

García, J. A., Lozano, R. &Teijón, J. M. (1991): *Problemas de Química de Selectividad*. Madrid: Tebar Flores.

García, T. & García-Serna, J. R. (2003). *Química*. Barcelona: Edebé.

Gil, V. M. S. & Paiva, J. C. (1999). Equilibrium constant units revisited. *Chemical Educator, 4*(4), 128-130.

Goedhart, M. J. & Kaper, W. (2002). From chemical energetics to chemical thermodynamics. In Gilbert, J., De Jong O., Justi, R., Treagust, D. & Van Driel J. (eds.), *Chemical Education: Towards research-based practice* (pp. 339-362). Dordrecht: Kluwer.

Goldberg, D. E. (2001). *How to solve word problems in chemistry*. New York: McGraw Hill.

Gordus, A. A. (1991). Chemical equilibrium. The thermodynamic constant. *Journal of Chemical Education, 68*(2), 138-140. DOI: 10.1021/ed068p138

Guardia, C., Menéndez, A. I. & de Prada, F. (2009). *Química 2 Solucionario*. Madrid: Santillana.

Harris, W. F. (1978). The Plethora of Equilibrium Constants. *Chemsa, 4*, 170-172.

Harris, W. F. (1982). Clarifying the concept of equilibrium in chemically reacting systems. *Journal of Chemical Education, 59*(12), 1034-1036. DOI: 10.1021/ed059p1034

Henry, A. J. (1967). The dimensions of physical quantities. *Education in Chemistry, 4*, 81-86.

Illana, J., Araque, J. A., Liébana, A. & Teijón, J. M. (2016). *Química 2*. Madrid: Anaya.

Jespersen, N. D.; Brady, J. E. &Hyslop, A. (2012). *Chemistry. The Molecular Nature of Matter*. Hoboken: Wiley.

Justi, R. & Gilbert J. K. (2002). Models and modelling in chemical education. In Gilbert, J., De Jong, O., Justi, R., Treagust, D. & Van Driel, J. (eds.). *Chemical education: towards research-based practice* (pp. 213–234). Dordrecht: Kluwer.

Khine, M. S. (2013). *Critical Analysis of Science Textbooks. Evaluating Instructional Effectiveness*. Dordrecht: Springer.

López, J. A. (2000). *Problemas de Química*. Madrid: Prentice Hall.

Laidler, K. J. (1990). Units of an equilibrium constant. *Journal of Chemical Education, 67*(1), 88. DOI: 10.1021/ed067p88.1

Levine, I. (2009). *Physical chemistry*. New York: McGraw-Hill.

Matta, G. F., Massa, L., Gubskaya, A. W. & Knoll, E. (2011). Can one take the logarithm or the sine of a dimensioned quantity or a unit? Dimensional analysis involving transcendental functions. *Journal of Chemical Education, 88*(1), 67-70. DOI: 10.1021/ed1000476

McQuarrie, D. A. & Simon, J. D. (1997). *Physical chemistry: a molecular approach*. Susalito: University Science Books.

Mills, I. M. (1989). The choice of names and symbols for quantities in chemistry. *Journal of Chemical Education, 66*(11), 887-889. DOI: 10.1021/ed066p887

Mills, I. M. (1995). Dimensions of logarithmic quantities. *Journal of Chemical Education, 72*(10), 954-955. DOI: 10.1021/ed072p954

Mills, I., Cvitas, T., Homann, K., Kallay, N. & Kuchitsu, K. (1983). *Quantities, units, and symbols in physical chemistry*. Oxford: Blackwell.

Molyneux, P. (1991). The dimensions of logarithmic quantities: Implications for the hidden concentration and pressure units in pH values, acidity constants, standard thermodynamic functions, and standard electrode potentials. *Journal of Chemical Education, 68*(6), 467-469. DOI: 10.1021/ed068p467

Moncaleano, H. (2008).La enseñanza del concepto de equilibrio químico. Análisis de las dificultades y estrategias didácticas para superarlas. Doctoral Dissertation. Departament de Didàctica de les CiènciesExperimentals. Universitat de València.

Moore, J. W., Stanitski, C. L., Wood, J. L.,Kotz, J.C.&Joesten, M. D. (1998). *The Chemical World. Concepts and Applications*. Orlando: Saunders.

Morcillo, J. & Fernández, M. (1991). *Selectividad. Pruebas de 1990: Química*. Anaya: Madrid.

Morcillo, J., Fernández, M. & Carrión, V. E. (1998). *Química 2*, Madrid: Anaya.

Mortimer, R. G. (2000). *Physical chemistry*. San Diego: Hartcourt.

Oxtoby, D. W., Gillis, H. P. & Campion, N. H. (2012). *Principles of Modern Chemistry*. Belmont: Brooks/Cole.

Pedrosa, M. A. & Dias, M. H. (2000). Chemistry textbook approaches to chemical equilibrium and student alternative conceptions. *Chemistry Education Research and Practice, 1*(2), 227-236. DOI: 10.1039/A9RP90024A

Pekdağ, B. & Azizoğlu, N. (2013). Semantic mistakes and didactic difficulties in teaching the "amount of substance" concept: a useful model. *Chemistry Education Research and Practice, 14*, 117 – 129. DOI: 10.1039/C2RP20132A

Pethybridge, A. D. & Mills, I. M. (1979). Answer to: does K have units? *Education in Chemistry, 79*, 191.

Petrucci, R. H., Harwood, W. S., Herring, G. E. & Madura, J. (2006). *General Chemistry: Principles and Modern Applications*. New Jersey: Prentice-Hall.

Petrucci, R. H., Herring, G. E., Madura, J. & Bissonnette, C. (2017). *General Chemistry: Principles and Modern Applications*. New Jersey: Prentice-Hall.

Quílez, J. (2012). First-year university chemistry textbooks' misrepresentations of Gibbs energy. *Journal of Chemical Education, 89*(1), 87-93. DOI: 10.1021/ed100477x

Quílez, J. (2016). K_p does not play the role of the thermodynamic equilibrium constant, K°: a discussed example on general chemistry textbooks' misrepresentations. *Chemistry: Bulgarian Journal of Science Education, 25*(6), 815-825.

Quílez, J. & Solaz, J. J. (1995). Students' and teachers' misapplication of Le Châtelier's principle. Implications for teaching chemical equilibrium, *Journal of Research in Science Teaching, 32*(9), 939-958. https://doi.org/10.1002/tea.3660320906

Quílez-Díaz, A. & Quílez-Pardo, J. (2015). Avoiding general chemistry textbooks' misrepresentations of chemical equilibrium constants. *African Journal of Chemical Education, 5*(2), 72-86.

Quintero, G. (1987). Le Châtelier – right or wrong? *Journal of Chemical Education, 64*(12), 1069. DOI: 10.1021/ed064p1069.2

Raff, L. M. (2001). *Principles of physical chemistry*. Englewood Cliffs: Prentice-Hall.

Raff, L. M. (2014).Spontaneity and Equilibrium: Why "$\Delta G < 0$ Denotes a Spontaneous Process" and "$\Delta G = 0$ Means the System Is at Equilibrium" Are Incorrect? *Journal of Chemical Education, 91*(3), 386-395. DOI: 10.1021/ed400453s

Ronneau, C. (1993). Reflexions concernant "le dimensionnement ou le non dimensionnement" des constantes d'équilibre. *Association Belge des Professeurs de Physique et Chimie, 119*, 231-233.

Rosenberg, R. M.&Kotz, I. M. (1999). Spontaneity and the equilibrium constant: advantages of the Planck function. *Journal of Chemical Education, 76*(10), 1448-1451. DOI: 10.1021/ed076p1448

Sanger, M. J. & Greenbowe, T. J. (1999). An analysis of college chemistry textbooks as sources of misconceptions and errors in electrochemistry. *Journal of Chemical Education, 76*(6), 853-860. DOI: 10.1021/ed076p853

Silbey, R. J., Alberty, R. A. & Bawendi, M. G. (2005). *Physical chemistry*. Hoboken: Wiley.

Simón, B., García-Serna, J. & Romero, J. J. (2016). *Química 2*. Barcelona: Edebé.

Sözbilir, M. (2004). What makes physical chemistry difficult? Perceptions of Turkish chemistry undergraduates and lecturers. *Journal of Chemical Education, 81*(4), 573-578. DOI: 10.1021/ed081p573

Spencer, J. N., Bodner, G. M. & Rickard, L. H. (2008). *Chemistry. Structure and Dynamics*. New York: Wiley.

Treptow, R. S. (1999). How thermodynamic data and equilibrium constants changed when the standard-state pressure became 1 bar. *Journal of Chemical Education, 76*(2), 212-215. DOI: 10.1021/ed076p212

Tulip, D. & Cook, A. (1993). Teacher and student usage of science textbooks, *Research in Science Education, 23*(1), 302–307. https://doi.org/10.1007/BF02357074

Tykodi, R. J. (1986). A better way of dealing with chemical equilibrium. *Journal of Chemical Education, 63*(7), 582-585. DOI: 10.1021/ed063p582

Vemulapalli, G. K. (1993). *Physical chemistry*. Englewood Cliffs: Prentice-Hall.

Vickerman, C. (1979). Answer to: does K have units? *Education in Chemistry, 79*, 191-192.

Wright, P. G. (1979). Answer to: does K have units? *Education in Chemistry, 79*, 192-193.

Zubiaurre, S. & Arsuaga, J.M. (2007). *Selectividad. Pruebas de 2006: Química*. Anaya: Madrid.

Zumdahl, S. S. & Zumdahl, S. S. (2007). *Chemistry*. Boston: Houghton Mifflin.

Zumdahl, S. S. & DeCoste, D. J. (2018). *Chemical Principles*. Boston: Cengage Learning.

Appendix: List of General Chemistry textbooks examined

A) Grade-12 Spanish Chemistry textbooks

Guillén, J., Julia, S., Masana, J. & Pascual, A. (1978). *Química COU*. Madrid: Magisterio Español.

Salinas, F.&de Manuel, E. (1978). *Química COU*. Zaragoza: Edelvives.

Guillem, C. (1979). *Química COU*. Alcoy: Marfil.

Miralles, L.; Nagore, E. & Pascual, R. M. (1982). *Química COU*. Valencia: ECIR.

Royo, P.; Serrano, R. & Otero, A. (1982). *Química COU*. Madrid: Santillana.

Martínez, A. (1989). *Química COU*. Madrid: Bruño.

Oroz, J. (1989). *Química COU*. Madrid: SM.

Alonso, P., Cebeira, R., García, M. J. & Ortega, E. (1990). *Química COU*. Madrid: McGraw Hill.

Morcillo, J. & Fernández, M. (1990). *Química COU*. Madrid: Anaya.

Caamaño, A., Obach, D. & Servent, A. (1991). *Química COU*. Barcelona: Teide.

Latre, F. & Usó, J. (1991). *Química*. Vila-real: Miralles.

Sauret, M. (1991). *Química COU*. Madrid: Bruño.

Masjuan, M. D., Dou, J. M. & Pelegrín, J. (1991). *Química COU*. Barcelona: Casals.

Arróspide, M. C. & Piñar, M. I. (1992). *Química COU*. Zaragoza: Luis Vives.

Barrio, J. I. & Belmonte, M. (1992). *Química COU*. Reacción. Madrid: SM.

Pozas, A., Illana, J., Garrido, A., Romero, A. & Teijón, J. M. (1993). *Química COU*. Madrid: McGraw Hill.

Ruíz, A., Rodríguez, A., Martín, R. & Pozas, A. (1997). *Química*. Madrid: McGraw Hill.

Morcillo, J., Fernández, M. & Carrión, V. E. (1998). *Química 2*. Madrid: Anaya.

Belmonte, M. (1999). *Química 2*. Barcelona: Ediciones del Serbal.

Fornells, M., García, R. M., López, V., Miró, A. & Pérez, E. (1999). *Química 2*. Barcelona: Búixoloa.

Oro, L. A., Andreu, J. L., Fernández, M. C. & Pérez-Torrente, J. J. (1999). *Química 2*. Madrid: Santillana.

Andrés, D. M., Antón, J. L., Barrio, J. I., de la Cruz, M. C. & González, F. (2000). *Química 2*. Madrid: Editex.

Caamaño, A. & Obach, D. (2000). *Química 2*. Barcelona: Teide.

Barrrio, J. I. &Montejo, C. (2003). *Química 2*. Madrid: SM.

Cardona, A. R., Pozas, A., Martín, R. & Ruíz, A. (2003). *Química 2*. Madrid: McGraw Hill.

Fidalgo, J. A. & Fernández, M. R. (2003). *Química 2*. León: Everest.

García, T. & García-Serna, J. R. (2003). *Química*. Barcelona: Edebé.

Pérez, M. A., Pastor, E.&Ortiz, M. J. (2003). *Química 2*. Madrid: Santillana.

Rodríguez, A., Pozas, A., Martín, R. & Ruíz, A. (2007). *Química*. Madrid: McGraw Hill.

Barrio, J. I., Bárcena, A. I., Sánchez, A. & Caamaño, A. (2009). *Química 2*. Madrid: SM.

Fontanet, À. (2009). *Química*. Barcelona: Vicens Vives.

Zubiaurre, S., Arsuaga, J. M.& Garzón, B. (2009). *Química 2*. Madrid: Anaya.

Guardia, C., Menéndez, A. I. & de Prada, F. (2009). *Química*. Madrid: Santillana.

Brown, C. & Ford, M. (2014). *Higher Level Chemistry*, (2nd ed). Harlow: Pearson.

Ryan, L. & Norris, R. (2014). *A Level Chemistry*, (2nd ed). Cambridge: Cambridge University Press.

Barrio, J. I.; Sánchez, A. Bárcena, A. I. & Caamaño, A. (2016). *Química 2*. Madrid: SM.

Carriedo, G. A., Fernández, J. M. & García, M. J. (2016). *Química 2*. Madrid: Paraninfo.

Guardia, C. & Menéndez, A. I. (2016). *Química 2*. Madrid: Santillana.

Illana, J., Araque, J. A., Liébana, A. & Teijón, J. M. (2016). *Química 2*. Madrid: Anaya.

Manuel, M. M. & Pizarro, F. (2016). *Química 2*. Zaragoza: Edelvives.

Masjuan, M. D. & Pelegrín, J. (2016). *Química 2*, Barcelona: Casals.

Pozas, A.; Martín, R.; Rodríguez, A.; Ruiz, A.&Vasco, A. J. (2016). *Química 2*. Madrid: McGraw Hill.

Sauret, M. (2016). *Química 2*. Madrid: Bruño.

Simón, B.; García-Serna, J. & Romero, J. J. (2016). *Química 2*. Barcelona: Edebé.

Vidal, M. C.; Peña, J. (2016). *Química 2*. Madrid: Oxford.

B) Fist-year university Chemistry textbooks

Kotz, J. C. & Purcell, K. F. (1987). *Chemistry & Chemical Reactivity*. Philadelphia: Saunders.

McQuarrie, D. A. & Rock, P. A. (1987). *General Chemistry*. New York: Freeman and Company.

Atkins, P. W. (1989). *General Chemistry*. New York: Scientific American Books.

Bailar, J. C., Moeller, T., Kleinberg, J., Guss, C. O., Castellion, M. E. & Metz, C. (1989). *Chemistry*. San Diego: Hartcourt.

Bodner, G. M. & Pardue, H. L. (1989). *Chemistry: an Experimental Science*. New York: Wiley.

Segal, B. G. (1989). *Chemistry. Experiment and Theory*. New York: Wiley.

Freemantle, M. (1991). *Chemistry in Action*. London: MacMillan.

Gillespie, R. J., Humphreys, D. A., Baird, N. C. & Robinson, E. A. (1991). *Chemistry Massachusetts*: Allyn and Bacon.

Gillespie, R. J., Eaton, D. R., Humphreys, D. A. & Robinson, E. A. (1994). *Atoms, Molecules, and Reactions. An Introduction to Chemistry*. New Jersey: Prentice-Hall.

Brown, T. L., LeMay, H. E. & Bursten, B. E. (1997). *Chemistry: the Central Science*. New Jersey: Prentice Hall.

Whitten, K. W., Gailey, K. D. & Davis, R. E. (1997). *General Chemistry*. Orlando: Saunders.

Moore, J. W., Stanitski, C. L., Wood, J. L., Kotz, J. C. & Joesten, M. D. (1998). *The Chemical World. Concepts and Applications*. Orlando: Saunders.

Umland, J. B. & Bellama, J. M. (1999). *General Chemistry*. Pacific Grove: Brooks/Cole.

Brady, J. E., Russell, J. W. & Holum, J. R. (2000). *Chemistry. Matter and its Changes*. New York: Wiley.

Clugston, M. & Flemming, R. (2000). *Advanced Chemistry*. Oxford: Oxford University Press.

Masterton, W. L. & Hurley C. N. (2001). *Chemistry: Principles and Reactions*. New York: Harcourt.

Hill, J. W. & Petrucci, R. H. (2002). *General Chemistry. An integrated approach*. New Jersey: Prentice Hall.

Olmsted, J. & Williams, G. M. (2002). *Chemistry*. New York: Wiley.

Kotz, J. C. & Treichel, P. M. (2003*). Chemistry and Chemical Reactivity*. Pacific Grove: Brooks-Cole.

Tro, N. J. (2003). *Introductory Chemistry*. New Jersey: Prentice Hall.

Petrucci, R. H., Harwood, W. S., Herring, G. E. & Madura, J. (2006). *General Chemistry: Principles and Modern Applications*. New Jersey: Prentice-Hall.

Zumdahl, S. S. & Zumdahl, S. S. (2007). *Chemistry*. Boston: Houghton Mifflin.

Buthelezi, T., Dingrando, L., Hainen, N., Winstrom, C. & Zike, D. (2008). *Chemistry. Matter and change*. Columbus: McGraw-Hill.

McMurry, J., & Fay. C. (2008). *Chemistry*. New Jersey: Prentice Hall.

Spencer, J. N., Bodner, G. M. & Rickard, L. H. (2008). *Chemistry. Structure and Dynamics*. New York: Wiley.

Silberberg, M. S. (2009). Chemistry. *The Molecular Nature of Matter and Change*. New York: McGraw-Hill.

Atkins, P. W. & Jones, L. L. (2010). *Chemical Principles. The Quest for Insight*. New York: Freeman and Company.

Ebbing, D. D. & Common, S. D. (2011). *General Chemistry*. Belmont: Brooks/Cole.

Moore, J. W., Stanitski, C. L. & Jurs, P. C. (2011). Chemistry. *The molecular science*. Belmont: Brooks/Cole.

Jespersen, N. D., Brady, J. E. & Hyslop, A. (2012). Chemistry. *The Molecular Nature of Matter*. Hoboken: Wiley.

Brown, T.L., LeMay, H. E., Bursten, B. E., Murphy, C-J. & Woodward, P. M. (2012). *Chemistry: the Central Science*. New Jersey: Prentice Hall.

Oxtoby, D. W., Gillis, H. P. & Campion, N. H. (2012). *Principles of Modern Chemistry*. Belmont: Brooks/Cole.

Tro, N. J. (2013). *Chemistry: a molecular approach*. New Jersey: Prentice Hall.

Kotz, J. C., Treichel, P. M.; Townsend, J. R. & Treichel, D. A. (2015). *Chemistry & Chemical Reactivity*. Stamford: Cengage Learning.

Chang, R. & Goldsby, K. A. (2016). *Chemistry*. New York: McGraw-Hill.

Burrows, A., Holman, J., Parsons, A., Pilling, G. & Price, G. (2017). *Chemistry. Introducing inorganic, organic and physical chemistry*. Oxford: Oxford University Press.

Zumdahl, S. S. & DeCoste, D. J. (2018). Chemical Principles. Boston: Cengage Learning.

Petrucci, R. H., Herring, G. E., Madura, J. & Bissonnette, C. (2017). *General Chemistry: Principles and Modern Applications*. New Jersey: Prentice-Hall.

Gilbert, T. R., Kirss, R. V., Foster, N., Bretz, S. L., Davies, G. (2018). *Chemistry. The science in context*. New York: Norton & Co.

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