Activity-based Analysis of Potentiometric pH Titrations

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Abstract

The discrepancy between concentrations and activities is a predicament well known to the analytical chemist. Because of the difficulty of determining activity coefficients, the standard technique for quantitative equilibrium studies is to work under a particular ‘constant ionic strength’ by adding an excess of an inert salt. Under such conditions, activity coefficients are approximately constant and can be taken into the equilibrium constants which are defined for the chosen ionic strength (I). Here we propose a fundamentally different approach. Throughout the numerical analysis of the titration data, activity coefficients for all individual species are approximated by well-known equations based on the work of Debye-Hückel. The computational analysis of the measurements strictly obeys the law of mass conservation and obeys the law of mass action only approximately. The main novelty is that now the addition of inert salts is no longer required and measurements are done at minimal I. Consequently, the thermodynamic equilibrium constants are now determined much more robustly based on experiments taken at low I. The approach has been tested and validated with the two very well investigated 3-protic phosphoric and citric acids. In summary: the technique of artificially keeping ionic strength constant has been made replaced by improved computational analysis.

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**Introduction**

Analysis and modelling of chemical processes are often based on the law of mass action. This is the case for both equilibrium and kinetic investigations. However, whilst the law of mass action is based on activities, chemists are used to dealing with concentrations, a mismatch that is not easily overcome. The activity coefficients that quantify the difference between concentration and activity are parameters that are very difficult to determine experimentally.

There are two main reasons why concentrations (molarity) are almost universally used: (a) solutions of known concentration are straightforwardly prepared, using balances and volumetric flasks; (b) computational analysis software is almost universally based on concentrations, in order to satisfy point (a) but also because the computations are considerably simpler. (Note that very similar arguments apply to molality-based concentrations, they are less commonly used in analytical chemistry and we will not further elaborate on molalities).

The relationship between concentrations and activities is formally very simple, however, the activity coefficients which relate the two are very difficult to determine.

The standard way of dealing with the predicament is to work under ‘constant ionic strength’ by adding an excess of an inert salt [1, 2, 3, 4]. This approach is attractive as the activity coefficients are largely depending on the ionic strength (I) of the solution. If I is constant the activity coefficients are constant as well and can be taken into the equilibrium constants (or rate constants in the case of kinetic investigations). Importantly, standard analysis software can be employed. Finally, equilibrium (or rate) constants are reported for the specific ionic strength at which the investigation was conducted. There are several disadvantages with this approach.

(a) Almost invariably the ionic strength of a solution changes with progress of the reaction. Addition of an excess of an inert salt will not reduce the absolute change in ionic strength, it will only reduce the relative change. While the effect on activity coefficients is more pronounced low ionic strengths, they are only approximately constant at high ionic strength and thus, the assumption is that at high I the activity coefficients will not change much and can be considered to be ‘constant’.

(b) An additional and important problem is that ‘inert’ salts are never completely inert and tend to interfere with other ionic species either specifically or via ion pairs or outer-sphere complexes [5, 6].

(c) The thermodynamically relevant equilibrium (or rate) constants at infinite dilution or zero ionic strength are not directly experimentally accessible. The accepted method for their determination is to repeat the experiment under a range of sufficiently high ionic strengths with subsequent extrapolation to I=0 [7, 8, 9, 10]. The disadvantages of this approach are obvious: many experiments need to be taken and measurements at high I result in large extrapolation to zero ionic strength. Obviously, the larger the extrapolation the less robust it is.

In this contribution, we propose a better approach which is based on fundamental thermodynamics and advanced computational analysis.
Throughout all computations, activities are used, and they are computed from known concentrations and estimates for the activity coefficients. While activity coefficients still are only estimated, this approach has several advantages.

(a) As there is no requirement for constant $I$ there is no need to add inert salts and thus investigations are carried out at minimal ionic strength. Under such conditions approximations for the activity coefficients are most reliable.

(b) Thermodynamic equilibrium constants are resulting directly from the analysis. No ionic strength dependences have to be determined.

While the principle can be applied for the analysis of any chemical process, in this contribution we develop and experimentally evaluate the approach for the analysis of potentiometric pH titrations.

We base the theory on a commonly used 3-component system of a $M$, $L$ and $H$ (metal, ligand and proton). From these components there are any number of species formed, e.g.

$$\begin{align*}
m_iM + l_iL + h_iH & \leftrightharpoons M_{m_i}L_{l_i}H_{h_i} \\
\beta_i & = \frac{(M_{m_i}L_{l_i}H_{h_i})}{(M)^{m_i} \times (L)^{l_i} \times (H)^{h_i}} = \frac{\gamma_i[M_{m_i}L_{l_i}H_{h_i}]}{(\gamma_M[M])^{m_i} \times (\gamma_L[L])^{l_i} \times (\gamma_H[H])^{h_i}}
\end{align*}$$

Equation 1

In the $i$-th equilibrium the $i$-th species is formed stoichiometrically from the components as given in equation 1. Each set of stoichiometric coefficients $m_i$, $l_i$, $h_i$ defines one of the different species formed. The formation constant $\beta_i$ is defined thermodynamically as a function of the activities of all species. The activity of a species $x$ is represented as $\{x\}$. While the quotient of all activity coefficients relates the thermodynamically correct $\beta$-value in Equation 1 with the apparent concentration-based value, it is composed of all individual activity coefficients and best determined and understood in this way.

The activity coefficients $\gamma$ of all species need to be approximated. There are many systems of approximations for activity coefficients, and we will expand on that later. In this work we concentrate on Debye-Hückel type equations where the activity coefficient of an ion is a relatively simple function of its charge and of overall ionic strength of the solution, Equation 2.

$$\gamma(M_{m_i}L_{l_i}H_{h_i}) = f(\text{charge}_i, \mu)$$

Equation 2

There are several adaptations and extensions of the original Debye-Hückel equation, but the principles are maintained. All these variations of the original Debye-Hückel equations deliver good approximations for low ionic strength and generally less accurate approximations for increasing ionic strengths.

In this contribution, in addition to the original Debye-Hückel equation (Equation 3), we use the Davies equation (4), and the Specific Interaction Theory (SIT), equation (5) [11, 12, 13, 14].
\[
\log \gamma_i = - \frac{A z_i^2 \sqrt{\mu}}{1+\sqrt{\mu}} \quad \text{Equation 3}
\]
\[
\log \gamma_i = - A z_i^2 \left( \frac{\sqrt{\mu}}{1+\sqrt{\mu}} - C \mu \right) \quad \text{Equation 4}
\]
\[
\log \gamma_i = - z_i^2 \frac{0.51 \sqrt{\mu}}{1+1.5 \sqrt{\mu}} + \sum_k b_k c_k \quad \text{Equation 5}
\]

The parameter \( A \) in equations (3) and (4) is independently defined by known equations and is a function of the dielectric constant, \( D \), of the solvent and the absolute temperature, \( T \): \( A = 1.8246 \times 10^6 / (DT)^{3/2} \) [15]. For measurements at 25°C in aqueous solution, the value of \( A \) is 0.5085 (dm³/mol)^{1/2}. The charge of the species \( i \) is given as \( z_i \) in all equations. The Davies equation (Equation 4) has the additional empirical parameter \( C \). A value of \( C = 0.3 \) is commonly used and has been employed in this work.

The SIT theory, Equation 5, is considerably more complex. For each species \( x \), the SIT equation requires a set of \( k \) interaction parameters, \( b_{xk} \), one for each independent ion in the solution. Many interaction parameters have been determined and published [e.g. 10], however, for a new system this can be demanding. It is not surprising that many researchers use simplified SIT equations, e.g. setting all or some interaction parameters as \( b_{xk} = 1 \). There are also adaptations of the SIT equation that include terms like \( D \mu \) and \( E \mu^2 \), instead of the interaction terms [155, 16, 17, 18]. The empirical parameters \( D \) and \( E \) need to be determined and activity coefficients will depend on the specific values.

There is a wide range of alternative and more advanced systems of approximations for activity coefficients. They are mainly developed for chemical engineering applications where high concentrations and ionic strengths are very common as are multiphase reactions and neutral molecules with strong polar interactions. A representative, but by no means complete list includes the Perturbed Chain Statistical Associating Fluid Theory (PC-SAFT) [19, 20], Conductor like Screening Model for Realistic Solvents (COSMO-RS) [21, 22], UNIQUAC [23] and UNIQUAC Functional-Group Activity Coefficients (UNIFAC) [24, 25] and the similar non-random two-liquid model (NRTL) [26]. Pitzer equations are equally powerful and can be used at high ionic strength, however there are even more adjustable parameters that need to be determined [27].

Industrial processes will always be conducted at highest possible concentrations and thus high ionic strengths. Simple Debye-Hückel approximations are only of very limited value. The main difficulty with these advanced approximations is that they require typically several additional and specific parameters which are often listed for common ions such as \( \text{Na}^+ \), \( \text{Cl}^- \), etc., but the determination of those parameters for new species is very difficult and/or time consuming. Even so, thermodynamically correct, activity-based analyses of chemical processes are wide-spread in the chemical engineering literature [28].
There are several publications on the determination of thermodynamic protonation constants from potentiometric titrations, a representative collection of references is given in [29,30,31,32]. While similar in the aim, there are substantial differences. Here, we estimate an individual activity coefficient for each ionic species, and these are further used in the law of mass action. In the above previous publications, group activities are estimated, an overall factor which corrects for the non-ideal behaviour. In some instances, only activities for the proton are introduced, this of course reflect the activity response of the pH electrode. More importantly, the activity correction is assumed to be constant throughout the titration, rather than continuously changing during the titration due to changing ionic strength.

The present version of the software package for the analysis of potentiometric titrations does not include such advanced approximations. The critical advantage is that it can straightforwardly be applied to any new investigations as, except for the SIT equation, no new parameters have to be determined. Potentiometric titrations are commonly performed at low concentrations in the range $10^{-4}$ – $10^{-2}$ M and as no inert salts are added measurements are taken at minimal ionic strength.

The numerical analysis of solution equilibria is based on the law of mass conservation and the law of mass action. The law of mass conservation dictates that the known total concentrations of the components must match the sum over all stoichiometrically corrected species concentrations. In the example of the components M, L and H there are 3 mass balance equations, as shown in equation 6.

\[
[M]_{tot} = \sum_{i=1}^{\text{species}} m_i \beta_i [M]^{m_i} [L]^{l_i} [H]^{h_i}
\]

\[
[L]_{tot} = \sum_{i=1}^{\text{species}} l_i \beta_i [M]^{m_i} [L]^{l_i} [H]^{h_i}
\]

\[
[H]_{tot} = \sum_{i=1}^{\text{species}} h_i \beta_i [M]^{m_i} [L]^{l_i} [H]^{h_i}
\]

Equation 6

In addition to the mass balance equations, for each species formed, the law of mass action, as exemplified in equation 1, is also obeyed. If \textit{nspecies} species are formed, there are \textit{nspecies-ncomponents} mass action equations. Note, the components themselves are also contained in the list of species, but their mass action equation is trivial, e.g. \{M\}=1×\{M\}. These trivial equations for the components are replaced by the same number of mass balance equations. Thus, altogether there are \textit{nspecies} equations and the same number of unknowns, i.e. all species concentrations. These systems of equations always have one solution, the one established in the actual mixture.

Newton-Raphson type algorithms are most appropriate for the calculation of the composition of an equilibrated solution of known equilibrium constants and total concentrations of the components \([M]_{tot}, [L]_{tot}\) and \([H]_{tot}\). Here we briefly summarise the working of the Newton-Raphson algorithm and the adaptations required for activity-based analysis. The species concentrations as calculated by the adapted Newton-Raphson algorithm still comply with the law
of mass conservation, equation 1. In addition, the activities as approximated by any of the equations 5 – 7 must obey the law of mass action, equation 6. The Newton-Raphson algorithm requires initial estimates for the free species concentrations, in the example [M], [L], and [H]. From these all species concentrations are calculated using equation 1. Next the component total concentrations are computed, as in equations 6, and compared with the known total concentrations. If the differences are below a given limit the calculation is finished and the present species concentration are returned. If the calculated total concentrations differ significantly an iterative refinement is started which is usually based on a Taylor series expansion of the total component differences as a function of the free component concentration. See [33] for extensive development of the procedure as well as the Matlab code.

Adaptation for activity-based calculations requires an inner loop starting with an initial guess for the ionic strength. This allows the approximation of the component activities for the present I which in turn allows the computation of all species activities using equation 3. Using the activity coefficients, all species concentrations are calculated. Rather than checking the total concentration match, now the ionic strength is re-calculated, and the inner loop is repeated till convergence. Only then the total concentration check is performed in the outer, standard Newton-Raphson iteration. It is beyond the scope of this contribution and the scope of this journal to include extended elaborations on the novel algorithm. Free software (ReactLab EQSIM) is available that calculates or simulates complete titrations based on the above ideas [34].

We illustrate the differences between concentration and activity analysis using the example of a titration of 10 ml of 0.03M Na$_2$CO$_3$ in sea water with 0.8 ml of 1M HCl. Sea water is assumed to be 3.5% NaCl with translates into 0.61 M NaCl. As the final concentration of free CO$_2$ is close but below its solubility in sea water, no corrections for potential evaporation of CO$_2$ after the second protonation of the carbonate ion was included. The example was chosen as there is a natural high ionic strength and there are ions of different charge involved in the equilibria.

Figure 1 displays the results in 5 panels. Panel (A) shows the concentration profiles of the relevant species vs. pH as calculated with the different γ-approximations. They are very similar; the difference reflects the different protonation constants that result for the different activity approximations and additionally different activity coefficients for the proton, the pH is defined as -log[H$^+$]. The pH scale is decreasing to reflect the direction of the titration from left to right. The titration Panel (B) shows the concentrations as a function of added acid, it is reassuring that the quantitative analytical chemistry is not affected by the different calculations, the concentrations are the same for all approximations. Panel (C) shows the activities of all relevant species as a function of the pH, it also shows the total ionic strength, its y-axis on the right. The ionic strength does not change much (from 0.7 to 0.63) as there is large excess of the inert NaCl in sea water. Due to the 2- charge, the activities of the carbonate ion are much lower than the concentrations; the effect is less pronounced for the singly charged bicarbonate ion and the proton. For the neutral carbon dioxide activity and concentration are identical; within the range of Debye-Hückel approximations, neutral molecules have an activity coefficient of 1 irrespective of ionic strength. The activities of the proton appear
to be identical for all approximations, this is because the pH is calculated as \(-\log([H^+])\), the difference is in the x-axis. Panel (D) shows the activities as a function of addition of acid. Here the proton activities show up differently.

**Figure 1**: Concentration and activity profiles vs pH and vs addition for the titration of a carbonate solution in sea water with concentrated acid. Refer to the text for details.

The above ideas are tested based on the analyses of potentiometric titrations of the two tri-protic acids, phosphoric acid and citric acid. These two systems are reasonably complex, they cover the range of ionic charges from +1 to -3 and a wide range of protonation constants, three very different ones for phosphoric acid and three moderately similar ones for citric acid. Further, these two acids are well investigated with many published estimates for their thermodynamic protonation constants which can be compared with the results of this analysis. The thermodynamic equilibrium constants, extrapolated from ionic strength dependences, were taken from extensive data bases [35,36,37]. Note, in the Supplementary Materials we have compiled a list of published values for the two acids, including the original references. The averages of these values are used to compute the ‘published’ value for the extrapolated thermodynamic constants used in the Figure 2. Note, for such acids it is common to use stepwise protonation constants, logK, rather than cumulative formation constants, \(\beta\)-values, as defined in Equation 1. We report logK values for the protonation constants which are numerically identical to the dissociation constants reported as pKₐ values.
Experimental

Materials
Boiled, CO$_2$-free Millipore Milli-Q water was used for all solutions. Sodium hydroxide, and phosphoric acid were purchased from Sigma-Aldrich. Potassium hydrogen phthalate (KHP), hydrochloric acid, citric acid from Ajax Finechem and high purity N$_2$ from Coregas. All chemicals were analytical grade and used without further purification. All solutions were standardised in the normal ways.

Potentiometric titrations
Potentiometric titrations were performed using a Metrohm 665 Dosimat automated burette system, or a Metrohm 905 Titrando. Both systems were equipped with a jacketed titration vessel that was connected to a Julabo FL601 circulating water bath system for temperature control. The titrations were performed at T = 298 K ($\pm$0.1 K). A Metrohm combined micro-pH glass electrode (Model 6.0234.100) was used for the titrations. The electrode voltage in mV was read and recorded by a National Instruments NI-DAQ 7 board or directly with the Metrohm Tiamo 2.5 software. The mV signal was used directly in the analysis of the titration data. Nitrogen was bubbled through the acidified solutions for 10 minutes prior to beginning the titrations and subsequently passed above the solutions during the actual titrations to ensure exclusion of CO$_2$.

For each acid several different titrations were undertaken, they mainly differ in the total concentrations of the acid, ranging from 0.001 to 0.18 M for phosphoric acid and from 0.03 M to 0.009 M for citric acid. All acid solutions were additionally acidified with HCl (1 to 3 moles of HCl per mole of acid) to ensure close to complete protonation and improved definition of the third protonation constant. 40 to 80 additions of standardised NaOH solution (0.5 to 2.8 M) were used, the total volume of base ensured there was at least a 20% excess of hydroxide ions with respect to the total proton concentration. Some titrations were executed in duplicates some in triplicates. The relevant details of the titrations including all initial concentrations, volumes etc. are listed in the Supplementary Materials.

In an adaptation of the well-known Gran method, the pH-electrode was calibrated for proton activity by titrating hydrochloric acid of known concentration with standardised NaOH solution. This enables the determination of the slope and intercept that define the linear relationship between the potential delivered by the electrode and the logarithm of the proton activity [38].

Data Analysis
The titrations were analysed by the ReactLab pH software [39]. ReactLab pH is activity based analytical program which fits the protonation constants and, if required, the reagent concentrations and the electrode parameters. Activity approximations can be chosen, using any of the previously discussed equations 5 to 7. ReactLab pH also allows the classical concentration-based analysis by fixing all activity coefficients to a value of 1.
**Results**

The claim of this contribution is that a single and simple titration at minimal ionic strength delivers good estimates for the thermodynamic equilibrium constants. Figure 2 serves to support this assertion. The dashed lines are the averages of the published values for the thermodynamic protonation constants, determined by extrapolation; the dotted lines represent the standard deviations. A list of published values and their references are given in Table 1 of the Supplementary Materials. The markers represent the average over all individual titrations, as described in the experimental section; the vertical bars represent their standard deviations. The results are for the analyses based on different activity coefficient approximations as indicated on the x-axis. Note, the scales on the y-axes are maintained for all constants for both acids at 0.7 log units.

The results are clearer and more discernible for citric acid, seen at the bottom of Figure 2. First, for the activity-based analyses all protonation constants are consistently close to their published thermodynamic values. Equally, the standard deviations for activity-based analyses are considerably smaller than those for the classical concentration-based analysis. The rationale is straightforward: there is a wide range of experimental conditions, total concentrations and number of points taken during the titrations, resulting in relatively large differences for the classical analyses as the ionic strength changes are substantial and thus the results more variable than for the activity-based analyses. In contrast, the activity-based analyses result in increased consistencies due to ionic strength effects being incorporated into the analysis. Their standard deviations are comparable to those of the published values.

The above observations are similar but slightly less consistent for the phosphoric acid titrations. The range of concentrations for these titrations were substantially higher than for the citric acid titrations, thus the ionic strengths were higher and consequently their changes during the titration more prominent as well. As with citric acid, this is reflected in the wider spread for the concentration-based analysis results. As with citric acid, the second and third protonation constant are closer to the literature values for the activity-based analyses. The results for the first protonation constant are different, the standard deviations are similar for all analyses and the concentration-based average is closest to the literature values. The most likely explanation is that high protonation constants are notoriously difficult to determine as the deprotonation of water which exists in much higher concentration is interfering. This interpretation is supported by the larger standard deviation for the literature values for this protonation constant. However, a significant number of additional and carefully designed experiments may be required for the ultimate understanding of the fact, such an undertaking would be beyond the scope of this contribution.
Figure 2: Protonation constants for phosphoric acid (top) and citric acid (bottom), analysed in concentration mode and with different approximations for the activity coefficients. The markers represent the averages of the analysis results and their standard deviations. The dashed lines represent the averages of the published values for zero ionic strength and their standard deviation.

During any titration the ionic strength will always change and consequently each concentration based, classical analysis is expected to be inferior and this should be reflected in a larger sum of squares. This expectation is partially met, while the fits are consistently slightly better for the activity-based analyses, the differences are statistically not significant, and this result is not further analysed. The distortion of a titration curve due to changing activity coefficients are very subtle, and it appears these changes are absorbed into the calculations of the individual equilibrium constants and thus the activity-based fits are only marginally better.

Figures 3 reveals the differences between concentration (full lines) and activity (dashed lines) for one particular titration of citric acid, CA. The activity coefficients for this graph were estimated using the Davies Equation. This Figure displays the concentration and activity profiles for all species during the titration. The secondary y-axis shows the ionic strength (bold black line) of the solution during the titration, it covers a range of approximately 0.005 to 0.03M. The Figure clearly indicates the larger difference between activity and concentration for the 3- charged citrate ion, compared to the 2- charged hydrogen-citrate and all the 1- charged species, dihydrogen-citrate, hydroxide and the 1+ charged proton.
Figure 3: Concentration (full lines) and activity profiles (dashed lines) for all species in the titration of 10 ml 0.003 M acidified citric acid with 0.38 ml 0.53 M NaOH. The ionic strength of the solution is plotted as a bold black line with the secondary y-axis on the right.

Discussion

A novel method needs to be examined, the results must be tested and compared with those of established methods. In the present case we deal with a new algorithm and its outcomes need to be tested. In addition to the adherence to the law of mass conservation (Equation 6), all species activities are required to satisfy the law of mass action (Equation 1). This, while all activity coefficients are calculated from the species concentrations and one of the approximations in equations 3 to 5. This
cannot efficiently be demonstrated in a published text. Therefore, a fully worked ReactLab pH excel workbook is included in the Supplementary Materials. The workbook contains the complete analysis results (concentration and activity profiles for all species, all activity coefficients etc.) of an example titration (CitricAcid_Activity.xlsx) as established by ReactLab pH. In addition, it shows the tests for adherence to the law of mass action and of mass conservation, i.e. all species concentrations add up to the known total component concentrations, equation (6), and all activities introduced into the appropriate versions of equation (1) result in the given equilibrium constants. A summary of these results is given in Figure 1 in the Supporting Materials.

We stress again that the corrections are not perfect, but generally, an imperfect approximation is expected to be better than the assumption of constant ionic strength.

The analysis of potentiometric pH titrations based on activities rather than concentrations serves several purposes. (a) The law of mass action is based on activities and thus the suggested analyses are close to theoretically correct. Classical concentration-based analyses rely on constant ionic strength during the titration which is never completely achieved. (b) The results of the analysis include a set of thermodynamically correct equilibrium constants. These are otherwise only accessible via analysis of extended ionic strength dependences. (c) In classical analyses, inert salts need to be added to the titration solutions; inert salts are never perfectly inert, being of ionic nature they will always interfere with other ions in the solutions. Avoiding inert salts can only be advantageous. Additionally, as no salts need to be added, the ionic strength is minimal and consequently the activity approximations are optimal. The higher the excess the less the relative ionic strength is changing. However, a high ionic strength is far from the ideal of low ionic strength and thus increasing I will result in an increasingly less robust extrapolation to zero ionic strength. Software for concentration-based, classical analyses is well established, [e.g. 40, 41].

We propose a software solution to the above predicament: the analysis of the process is based on activities rather than concentrations. Activity coefficients are approximated by relatively simple and easily accessible Debye-Hückel based equations. The software thus enables the analysis of processes under non-constant ionic strengths, thus no added inert salts are required to maintain the ionic strength. Most importantly this allows the analysis of processes at minimal ionic strength, conditions where the approximations for the activity coefficients are optimal. The overall experimental advantages include also the avoidance of repeated experiments at different ionic strengths.

A valuable comment by a reviewer prompted us to state that the ionic strength effects on the junction potential of the pH electrode are not covered by our developments. However, the activity coefficient of the proton is naturally estimated in the same way as those of the other ionic species.

The software approach has been successfully tested for potentiometric pH titration of two tri-protic acids, phosphoric and citric acid. Implementation of the above principles for the analysis of spectrophotometric titrations as well as kinetics are in process.
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Highlights

- Analysis of potentiometric pH titrations based on the law of mass action, calculations are based on activities rather than concentrations.
- This avoids the necessity of maintaining constant ionic strength.
- This, in turn, avoids the addition of ‘inert’ salts, which usually are not completely inert.
- Thermodynamically correct equilibrium constants are determined without the investigation of ionic strength dependences.
- The overall result is a substantial saving in experimental effort, time and chemicals.
$$\beta_i = \frac{\{M_{m_i} L_{l_i} H_{h_i}\}}{\{M\}^{m_i} \times \{L\}^{l_i} \times \{H\}^{h_i}}$$
Declaration of interests

☒ The authors declare that they have no known competing financial interests or personal relationships that could have appeared to influence the work reported in this paper.

☐ The authors declare the following financial interests/personal relationships which may be considered as potential competing interests: