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Measurement of Acid Dissociation Constants of Weak Acids by Cation Exchange and Conductometry

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A simple strategy is presented for the determination of acid dissociation constants based on the measurement of conductance of a known concentration of the acid and/or the conductance of a solution of its fully or partially neutralized alkali-metal salts. For an *n*-protic acid, 2*n* conductance measurements are minimally necessary. In the simplest case of a typical monoprotic acid, the conductance of its alkali salt solution is measured before and after passage through an H⁺-form exchanger. From these data both the pK_a of the acid and the equivalent conductance of the anion can be computed. The underlying equations are rigorously solved for monoprotic acids and some diprotic acid systems. For other diprotic and multiprotic acid systems, initial estimates are obtained by making approximations; the complete data set is then subjected to multiparametric fitting. The method does not require pH measurements; conductance can generally be measured accurately at low enough ionic strengths to obviate the need for major activity correction. Several experimental measurements are presented and excellent agreement with literature pK_a values is observed. The reliability of the equivalent conductance values computed in this fashion is limited, however.

Measurements of dissociation constants of weak acids/bases are routinely necessary, e.g., for new compounds, for well-established compounds at a temperature where previous data are unavailable, and in different solvent systems. Virtually all determination techniques involve pH measurement as an integral part; frequently a pH-dependent property of the solution of the substance is measured. Several monographs that review the existing techniques are available (1–4). Conductometry and differential vapor pressure osmometry

are the only two common techniques that allow measurement of K_a without a measurement of pH; a particularly thorough review of the several classical conductometric approaches is presented in ref 3. By far the best and best known of these methods is the so-called absolute method of Fuoss and Krauss (5) in which no a priori assumptions are needed and the dissociation constant of a monoprotic weak acid is calculated by conductance measurement of a series of successively diluted solutions of the acid and extrapolation.

We propose here a simple conductometric technique for the determination of K_a values of weak acids. The method requires no assumptions beyond that the equivalent ionic conductance of H⁺, OH⁻, and alkali-metal ions (Na⁺, K⁺, if Na or K salts of the acids are used) in the solvent medium and the autoionization constant of the solvent medium are known with certainty. The method evaluates the dissociation constants computed in terms of concentration rather than in terms of activities. However, since measurements are made typically at a concentration of 10^{-3} M, in most cases the difference of the computed value from the true activity constant is insignificant. If desired, activity corrections can be made, using, for example, Kielland's data (6). Other approaches of making such corrections have also been delineated (2).

The technique we describe is particularly attractive if instrumentation for conductometric ion chromatography (IC) is available. No pH measurement is necessary. Application to aqueous systems is demonstrated.

PRINCIPLES

Monoprotic Acids. *Moderately Weak Acid (pH $\lesssim$ 6).* Consider a C molar solution of MX (M = alkali metal: Na, K, etc.) where HX is a weak acid with a dissociation constant K_a. Neglecting the contribution of autoprotolysis of water, the conductivity G_{MX} of this solution is given by

$$G_{MX} = ([M^+]\lambda_{M^+} + [X^-]\lambda_{X^-} + [OH^-]\lambda_{OH^-})/1000K \quad (1)$$

where λ_{M^+} and λ_{OH^-} are known, K is the cell constant, and

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$$[M^+] = C \tag{2}$$

$$[X^-] = C - [OH^-] \tag{3}$$

and

$$[OH^-] = (-K_b + (K_b^2 + 4K_bC)^{1/2})/2 \tag{4}$$

where K_b is K_w/K_a and K_w is the autoionization constant for water. Substituting eqs 2-4 in eq 1 thus results in an equation where λ_{X^-} and K_a are the only unknowns, assuming that the cell constant has been previously measured.

Consider now that this NaX solution is ion exchanged for H^+, as is done routinely by a suppressor in suppressed IC (7).

The conductivity of this C molar HX solution thus generated, G_{HX}, is given by (neglecting autoprotolysis of water)

$$G_{HX} = ((\lambda_{H^+} + \lambda_{X^-})[H^+])/1000K \tag{5}$$

where

$$[H^+] = (-K_a + (K_a^2 + 4K_aC)^{1/2})/2 \tag{6}$$

Equations 5 and 6 yield

$$\lambda_{X^-} = (2000KG_{HX}/(-K_a + (K_a^2 + 4K_aC)^{1/2})) - \lambda_{H^+} \tag{7}$$

Substitution of eq 7 in eq 1 yields

$$1000KG_{MX} - C[\lambda_{M^+} - \lambda_{H^+} + 2000KG_{HX}/$$
$$(-K_a + (K_a^2 + 4K_aC)^{1/2})] - [(-(K_w/K_a) +$$
$$((K_w/K_a)^2 + 4K_wC/K_a)^{1/2})/2][\lambda_{OH^-} + \lambda_{H^+} -$$
$$2000KG_{HX}/(-K_a + (K_a^2 + 4K_aC)^{1/2})] = 0 \tag{8}$$

Equation 8 can be readily iteratively solved for K_a. Once K_a is known, λ_{X^-} may also be computed, most readily from eq 7. We designate this approach as method 1.

The lower limit of pK_a values that can be determined in this fashion is dictated by the requirement that when G_{HX} is measured, a significant fraction ($\geq 10\%$) of HX should be in the undissociated form. If we assume that the maximum permissible concentration, to avoid the necessity of activity corrections, is 1 mM, then the minimum pK_a that can be measured is 2.1.

Weak Acids ($6 \lesssim pK_a \lesssim 9$). As the pK_a value approaches 7, the measurement of K_a by the approach above is compromised by the accuracy with which G_{HX} can be measured. Unless scrupulous care is taken, the dissolution of atmospheric CO_2 in the test solutions poses a problem. In the foregoing measurement method, the typical concentration C of NaX taken is 10^{-3} M. The applicability of the technique is extended up to a pK_a of ~ 9 if a higher concentration of HX (either directly or a higher concentration of NaX solution passed through a suppressor) is used for the measurement of G_{HX}. This can be a concentration of $\geq 10^{-2}$ M, which considerably ameliorates the effect of the intrusion of small amounts of CO_2. Although this concentration is high, due to the high pK_a of the acid, the ionic strength of the solution is too small to require significant activity corrections. The only modification necessary is to replace C in eqs 6 and 7 with C' where $C' \gg C$ and is the concentration of HX used for the measurement of G_{HX}. We designate this approach as method 2.

Weak and Very Weak Acids ($pK_a > 7$). For acids with pK_a above 7, an accurate measurement of G_{HX} can be seriously affected by strong acid impurities. A solution to this problem is to measure the conductance of two different concentrations of NaX (e.g., C and C' molar) and not rely on measurements of G_{HX}. For the C molar solution, eqs 1-4 remain valid; rearranging eqs 1-3 yields

$$\lambda_{X^-} = (1000KG_{MX} - [OH^-]\lambda_{OH^-} - C\lambda_{M^+})/(C - [OH^-]) \tag{9}$$

and for the C' molar solution a corresponding equation is

obtained. These two equations combine into one that contains only K_b as the unknown and can be readily solved iteratively. We designate this approach as method 3.

Diprotic and Multiprotic Acids. For an n-protic acid, there are $2n$ unknowns: dissociation constants K_{a1} through K_{an} and equivalent conductances $\lambda_{HX^-_{n-1}}$ through $\lambda_{X^{n-}}$. Minimally, $2n$ conductance measurements are therefore necessary. Further, if these data are to yield meaningful values of the desired parameters, the solutions used for such measurements should be such that all forms of the protolyzable anion are fully represented. Conductance of a C molar solution of $M_{n-m}H_mX$, derived from a n-protic acid ($m = 0 \rightarrow n$) can be generally expressed as

$$1000KG = (n - m)C\lambda_{M^+} + [H^+]\lambda_{H^+} + [OH^-]\lambda_{OH^-} +$$
$$CQ^{-1}[\sum_{i=1\rightarrow n} (K_{a1} \dots K_{ai}[H^+]^{n-i}i\lambda[H_{n-i}X^{i-}])] \tag{10}$$

and

$$Q = [H^+]^n + \sum_{i=1\rightarrow n} (K_{a1} \dots K_{ai}[H^+]^{n-i}) \tag{11}$$

A general analytical solution of $2n$ sets of eq 10 is difficult because of the necessity to eliminate $[H^+]$ from each; $[H^+]$ is different for each solution. Auxiliary equations involving mass and/or charge-balance must be brought in and the general situation is considerably complicated. While a strict analytical solution has thus far eluded us, a less rigorous but nevertheless generally applicable approach is to make approximations to obtain initial estimates of the parameters of interest. The full set of data, which may well contain more than $2n$ measurements, is then subjected to nonlinear least-squares fitting using a Marquardt-type random walk algorithm (8) to obtain the best fit of the experimental data, using the initial set of approximate values as the starting point. For an n-protic acid, $2n + 1$ conductance measurements are minimally required to allow multiparametric fitting; more measurements are preferred. The input data are a $m \times 3$ array where m is the number of measurements and the three columns represent the measured conductance, metal ion concentration, and concentration of the acid. This approach is designated as method 4 and is demonstrated in the results section.

Relative simple and straightforward explicit solutions can be achieved for some diprotic acid cases, however. As with monoprotic acids, the best experimental approaches are often dictated by the precise range of pK_a values involved. These are enumerated below.

(a) Monoprotic Acid Approximation. Whenever $K_{a1} \gg K_{a2}$, K_{a1} and λ_{HX^-} can be determined by making the monoprotic acid assumption and applying in effect method 1 or 2 as delineated for monoprotic acids. Thus K_{a1} and $\lambda_{HCO_3^-}$ can be determined by measuring G_{NaHCO_3} and measuring $G_{H_2CO_3}$ after proton exchange. (Similarly, K_{a1} and $\lambda_{H_2PO_4^-}$ may be determined from $G_{NaH_2PO_4}$ and $G_{H_3PO_4}$, etc.) We designate these approaches as method 1' or 2', as appropriate. However, in general the monoprotic acid approximation is most aptly applied when both test solutions are acidic and the effect of the second dissociation step is minimum. Therefore, it is best to use the conductance data for two concentrations of H_2X when applying this approximation. Thus for C molar H_2X

$$G_{H_2X} = ((\lambda_{H^+} + \lambda_{HX^-})[H^+])/1000K \tag{12}$$

where

$$[H^+] = (-K_{a1} + (K_{a1}^2 + 4K_{a1}C)^{1/2})/2 \tag{13}$$

combining eq 12 with its counterpart from C' molar H_2X

$$G_{H_2X}[H^+]' = G_{H_2X}'[H^+] \tag{14a}$$

Equation 14a is solved iteratively for K_{a1} and this value is put

back in eq 12 to compute λ_{HX^-}. This approach is designated as method 5.

Once λ_{HX^-} is known, the determination of K_{a2} and $\lambda_{X^{2-}}$ may be carried out by measuring the conductance of two concentrations of M_2X and considering only the $X^{2-} + H_2O \rightleftharpoons HX^- + OH^-$ equilibrium, as in method 3 for monoprotic acids. The sole difference is that the hydrolysis of the divalent X^{2-} generates conducting HX^- rather than nonconducting HX formed in the hydrolysis of X^-. Recognizing that $[HX^-] = [OH^-]$ and $[X^{2-}] = C - [OH^-]$

$$G_{M_2X} = (2C\lambda_{M^+} + [OH^-](\lambda_{OH^-} + \lambda_{HX^-}) + 2(C - [OH^-])\lambda_{X^{2-}})/1000K \quad (14)$$

where

$$[OH^-] = (-K_{b1} + (K_{b1}{}^2 + 4K_{b1}C))^{1/2}/2 \quad (15)$$

and

$$K_{b1} = K_w/K_{a2} \quad (16)$$

Similar to method 3, this approach, designated as method 3', involves transposition of eq 14

$$\lambda_{X^{2-}} = (1000KG_{M_2X} - 2C\lambda_{M^+} - [OH^-] \times (\lambda_{OH^-} + \lambda_{HX^-}))/(2(C - [OH^-])) \quad (17)$$

A parallel equation results from C' molar M_2X. Combining these two and putting in the value of λ_{HX^-} determined in the first part during the determination of K_{a1} results in an equation that is readily solved iteratively for K_{b1}. Instead of M_2X solutions, MHX solutions can be used. The computational approach is slightly more involved. For C and C' molar MHX

$$G_{MHX} = (C\lambda_{M^+} + [H^+]\lambda_{H^+} + [HX^-]\lambda_{HX^-} + 2[X^{2-}]\lambda_{X^{2-}} + [OH^-]\lambda_{OH^-})/1000K \quad (18)$$

From eq 18 and the corresponding equation for C' molar MHX, $\lambda_{X^{2-}}$ is eliminated by multiplication with $[X^{2-}]'$ and $[X^{2-}]$, respectively, and subtraction from each other

$$1000K([X^{2-}]'G_{MHX} - [X^{2-}]G_{MHX'}) = \lambda_{M^+}(C[X^{2-}]' - C'[X^{2-}]) + \lambda_{H^+}([H^+][X^{2-}]' - [H^+]'[X^2]) + \lambda_{HX^-}([HX^-][X^{2-}]' - [HX^-]'[X^{2-}]) + \lambda_{OH^-}([OH^-][X^{2-}]' - [OH^-]'[X^{2-}]) \quad (19)$$

Equation 19 is iteratively solved for K_{a2} by a double iteration step: For every trial value for K_{a2}, the quartic equation for the MHX case (9) is solved iteratively for $[H^+]$. From this value of $[H^+]$, we compute $[X^{2-}]$ and $[HX^-]$. The same step is carried out for $[H^+]'$ and then $[X^{2-}]'$ and $[HX^-]'$ are computed. Thus eq 19 is iteratively solved. The general strategy of this approach, designated method 6, can be applied to combinations of M_2X and MHX solutions as well.

Finally, it is possible to determine K_{a2}, $\lambda_{X^{2-}}$, and λ_{HX^-} from M_2X solutions alone if the determination of K_{a1} is not important. Consider that $\lambda_{X^{2-}}$ an be obtained by multiplying eq 14 with $[OH^-]'$ and the corresponding equation for C' molar M_2X with $[OH^-]$ and combining them

$$\lambda_{X^{2-}} = (500K(G_{M_2X}[OH^-]' - G_{M_2X}'[OH^-]) + \lambda_{M^+}(C'[OH^-] - C[OH^-]'))/(C[OH^-]' - C'[OH^-]) \quad (20)$$

λ_{HX^-} can be similiarly obtained by multiplying eq 14 with $(C' - [OH^-]')$ and the corresponding equation for C' molar M_2X with $(C - [OH^-])$ and eliminating $\lambda_{X^{2-}}$

$$\lambda_{HX^-} = 1000K(G_{M_2X}(C' - [OH^-]') - G_{M_2X}'(C - [OH^-])) + 2\lambda_{M^+}(C[OH^-]' - C'[OH^-])/(C'[OH^-] - C[OH^-]') - \lambda_{OH^-} \quad (21)$$

In iterative trials of varying K_{a2}, computing $[OH^-]$ via eq 15 and thence $\lambda_{X^{2-}}$ and λ_{HX^-} via eqs 20 and 21, we have gen-

erally observed that $\lambda_{X^{2-}}$ and λ_{HX^-} have meaningful positive values (e.g., between 10 and 100) over a narrow range of K_{a2}. If the range of this uncertainty is adequate, no further computation is necessary. Otherwise, the computed values can be substituted into an expression of conductance for a third M_2X solution, an equation similar to eq 14. This approach is designated method 7.

(b) *Strong Acid Case.* A special diprotic acid case involves the complete dissociation of the first proton. Thus, effectively there are three unknowns, K_{a2}, λ_{HX^-}, and $\lambda_{X^{2-}}$. These parameters can be determined by measuring two concentrations each (C and C' molar) of a pair of compounds chosen from HX, MHX, and M_2X. When K_{a2} represents a moderately weak acid, the HX/MHX pair is preferred. When K_{a2} represents a very weak acid, the MHX/M_2X pair is preferred. The general computational approach is outlined below. For C and C' molar H_2X

$$G_{H_2X} = (\lambda_{H^+}[H^+] + \lambda_{HX^-}[HX^-] + 2\lambda_{X^{2-}}[X^{2-}])/1000K \quad (22)$$

Recognizing that $[H^+] = C + [X^{2-}]$ and $[HX^-] = C - [X^{2-}]$ and rearranging

$$1000KG_{H_2X} = C(\lambda_{H^+} + \lambda_{HX^-}) + [X^{2-}](\lambda_{H^+} + 2\lambda_{X^{2-}} - \lambda_{HX^-}) \quad (23)$$

The second term on the righ-hand side (rhs) of eq 23 is eliminated by multiplying eq 23 with $[X^{2-}]'$ and the corresponding equation for the C' molar solution with $[X^{2-}]$, respectively

$$\lambda_{HX^-} = 1000K([X^{2-}]'G_{H_2X} - [X^{2-}]G_{H_2X}')/([X^{2-}]'C - [X^{2-}]C') - \lambda_{H^+} \quad (24)$$

where

$$[X^{2-}] = (-(C + K_{a2}) + ((C + K_{a2})^2 + 4K_{a2}C)^{1/2})/2 \quad (25)$$

and $[X^{2-}]'$ is expressed correspondingly. For C molar MHX

$$1000KG_{MHX} = C\lambda_{M^+} + [H^+](\lambda_{H^+} - \lambda_{HX^-} + 2\lambda_{X^{2-}}) + C\lambda_{HX^-} \quad (26)$$

The second term on the rhs of eq 26 is eliminated by multiplying with $[H^+]'$ and the corresponding eq for C' molar MHX with $[H^+]$, respectively

$$\lambda_{HX^-} = (1000K([H^+]'G_{MHX} - [H^+]G_{MHX}') - \lambda_M([H^+]'C - [H^+]C'))/([H^+]'C - [H^+]C') \quad (27)$$

where

$$[H^+] = (-K_{a2} + (K_{a2}{}^2 + 4K_{a2}C)^{1/2})/2 \quad (28)$$

and $[H^+]'$ is expressed correspondingly. For C and C' molar M_2X, λ_{HX^-} is given by eq 21.

Equating the rhs of any two of the trio of eq 21, 24, and 27 leads to an equation that contains only K_{a2} as the unknown and which can therefore be iteratively solved. This approach is designated as method 8.

Other diprotic acid cases are difficult to solve rigorously and are approached by method 4.

EXPERIMENTAL SECTION

All experiments were performed at 24 ± 1 °C. Distilled deionized water with specific resistivity ≥16 MΩ cm was used for all solutions, which were prepared immediately before use. All chemicals used were of reagent or analytical reagent grade.

The experimental system utilized essentially a single-channel flow-injection configuration. Pure water was pumped at 0.5 mL/min by a model B (Eldex Laboratories, Menlo Park, CA) reciprocating pump through a 4.6 × 250 mm column packed with 10-μm macroreticular poly(styrene divinylbenzene) functioning as a pulse dampener. All subsequent

Table I. Summary of Calculation Methods

calculation method	equations	description
1	1–8	computes K_a and λ_{X^-} for moderately weak monoprotic acids using MX and HX solutions of same concentration
2	1–8	as above, except for weak acids, HX concentration used is much higher than MX concentration
3	1–4, 9	as above, except for weak and very weak acids, using two solutions of MX
4		multiparametric fitting of K_a and λ values of conductance data, $2n + 1$ measurements minimally necessary, initial estimates obtained by other calculation methods and using approximation
1′	1–8	computes K_{a1} and λ_{HX^-} for a diprotic acid where $K_{a1} \gg K_{a2}$, ignoring K_{a2}, using method 1
2′	1–8	as above, except using method 2
5	12–14a	as above, except two solutions of H_2X are used
3′	14–16	computes K_{a2} and $\lambda_{X^{2-}}$ for a diprotic acid where $K_{a1} \gg K_{a2}$, using two M_2X solutions, K_{a1} is ignored, requires previous determination of λ_{HX^-}
6	18, 19	as above, except two MHX solutions (or one MHX and one M_2X solution) are used, K_{a1} is not ignored, requires previous determination of K_{a1} and λ_{HX^-}, uses nested iteration steps
7	14–16, 20, 21	computes K_{a2}, λ_{HX^-}, $\lambda_{X^{2-}}$ for a diprotic acid from three solutions of M_2X, ignores K_{a1}
8	22–28	computes K_{a2}, λ_{HX^-}, and $\lambda_{X^{2-}}$ for a diprotic acid where the first proton is fully dissociated using two solutions each of a pair of compounds chosen from H_2X, MHX, and M_2X

components, as enumerated below, were located inside a thermostated enclosure to minimize temperature fluctuations. The water carrier passed through a 0.5 × 5000 mm stainless steel thermal equilibration coil into a pneumatically actuated slider valve equipped with a 2-ml sample loop. The large loop volume is to ensure a plateau response and provides results equivalent to pumping the sample solution through the system. The effluent from the valve is divided at a tee; one stream proceeds through a 0.3 × 2000 mm PTFE restriction coil to a conductivity detection cell. The second stream also flows to a conductivity detection cell, via a 3 × 300 mm glass column packed with H^+-form cation exchanger (early version of a packed column anion suppressor available from Dionex Corp., Sunnyvale, CA). Identical Model 213A (Wescan Instruments, Santa Clara, CA) conductivity detectors were used; the cell constants were determined with aqueous KCl standards to be 33.1 ± 0.1 cm^{-1}.

The multiparametric nonlinear least-squares fitting program used (method 4) is a commercially available software package (NLLSQ, CET research group, Norman, OK). A listing of the other iterative algorithms (all methods except 4) in BASIC are available from the authors on request.

RESULTS AND DISCUSSION

Configurational Aspects. The described measurements can be accomplished by a number of experimental arrangements depending on the availability of equipment and inclinations of the experimenter. Some of these have unique merits worthy of discussion. The greatest sample conservation is achieved in a serial arrangement of the detectors with the H^+ exchanger located between the two detectors while a single detector can be used to perform two sequential conductance measurements from one injection by using a nested loop injection system (10) with the H^+-exchanger column comprising the inner loop. Potentially, the most information-rich system is a controlled dispersion flow-injection system in which a small volume of a known concentration of an alkali-metal hydroxide is injected into a flowing stream of a weak acid solution of a known concentration (or vice versa). It is necessary that the dispersion behavior of the sample bolus be accurately modeled by an independent dispersion calibration experiment involving the injection of a nonprotolyzable comparably sized solute into a water carrier. Thence, the exact composition corresponding to any instantaneous conductance value is known and an enormous amount of data can be obtained from a single injection. We have previously exploited such information-rich systems for computing pK_{In} for acid–base indicators and metal–ligand binding constants using simultaneous multiwavelength detection with a diode-array detector (11). However, it should be emphasized that these

experiments were internally calibrated for sample dispersion based on the known isosbestic wavelength of the indicator.

Choice of an Ion Exchanger Device. Membrane-based "suppressor" devices are almost exclusively used in the present practice of ion chromatography (12). The preferred use of a packed column ion exchanger in the present experiments may seem anachronic but is based on sound reasons. Chemically regenerated membrane suppressors inevitably lead to some penetration of the regenerant acid through the membranes (13) and this will seriously compromise the reliability of the data obtained in the present case. Although we have not specifically studied its utility, should a free acid be more easily available in pure form than an alkali-metal salt, it should be possible to use an M^+-form (preferred M^+ is likely to be Li^+) cation exchanger. Existing literature indicates that this may be easier to accomplish with a packed-column exchanger (14) rather than a membrane-based device (15).

Experimental Determinations. The pK_a values determined experimentally are listed in Table II along with the nature and approximate concentrations of the solutions used for the measurements. The calculation method, computed pK_a, and λ values are also listed along with literature values. The agreement for the pK_a values determined for HF, HCOOH, and CH_3COOH with the corresponding literature values is excellent, as is the attainable precision of the measurement. In comparison, the agreement of the computed ionic mobilities with literature data is not necessarily good and relative precision is poor. Note that the lack of relative precision in comparing λ vs pK_a measurements goes beyond the fact that the pK_a value is a logarithmically transformed version of the K_a value actually computed. The underlying reason for this is that in one or more of the test solutions, the overall conductance is dominated by H^+ or OH^-. For example, in a solution of HX, the conductance is linearly related to the sum of the ionic mobilities of H^+ and X^-. However, the former is so much larger than the latter that the computation is insensitive to minor changes in the value of λ_{X^-}. Without careful choice of test solutions such that the determination of the mobility of the anion is not dependent on computing a difference between two large numbers, the ionic mobility/equivalent conductance values determined by the present method are likely to be unreliable.

The determination of the pK_a value of a monoprotic acid is particularly straightforward by the present method when an alkali-metal salt is available; only small quantities of the sample are required. When new compounds are synthesized, the ability to characterize with a small amount of sample is particularly desirable. At this institution, Bartsch et al. (17) have a long-standing program of synthesizing crown ether

Table II. Experimentally Determined Values

acid	solutions Used (mM)	calculation method	determined values[a]	literature values[a,b]
HF	NaF (1), HF (1)	1	pK_a 3.225 ± 0.004 λ_{F^-} 55.12 ± 0.06	pK_a 3.18 (3.16–3.32) λ_{F^-} 54.4
HCOOH	HCOONa (1), HCOOH (1)	1	pK_a 3.764 ± 0.003 λ_{HCOO^-} 47.52 ± 0.43	pK_a 3.75 (3.73–3.75) λ_{HCOO^-} 54.6
CH₃COOH	CH₃COONa (1), CH₃COOH (1)	1	pK_a 4.761 ± 0.002 $\lambda_{CH_3COO^-}$ 35.27 ± 0.48	pK_a 4.76 (4.749–4.762) $\lambda_{CH_3COO^-}$ 40.9
DCOA[c]	Na-DCOA (1), H-DCOA (1)	1	pK_a 4.602 ± 0.024 λ_{DCOA^-} 22.40 ± 0.15	pK_a 4.59 ± 0.22[d] λ_{DCOA^-} na[e]
MOPSO[f]	Na-MOPSO (1), H-MOPSO (1)	1	pK_a 6.625 ± 0.157 λ_{MOPSO^-} 20.86 ± 0.93	pK_a 6.9[g] λ_{MOPSO^-} na
	Na-MOPSO (2), H-MOPSO (2)	1	pK_a 6.701 ± 0.103 λ_{MOPSO^-} 19.23 ± 0.85	
	Na-MOPSO (10), H-MOPSO (10)	1	pK_a 6.823 ± 0.061 λ_{MOPSO^-} 13.96 ± 2.31	
	Na-MOPSO (1), H-MOPSO (10)	2	pK_a 6.839 ± 0.012 λ_{MOPSO^-} 20.72 ± 0.44	
H₂SO₄	KHSO₄ (1, 1.5) H₂SO₄ (1, 1.5)	8	pK_{a2} 2.012 ± 0.053 $\lambda_{HSO_4^-}$ 55.2 ± 1.0 $\lambda_{SO_4^{2-}}$ 76.8 ± 4.8	pK_a 1.96 (1.88–2.04) $\lambda_{HSO_4^-}$ 50 $\lambda_{SO_4^{2-}}$ 80.0
	K₂SO₄ (1, 1.5) K₂SO₄ (1, 1.5) KHSO₄ (1, 1.5) H₂SO₄ (1, 1.5)	4 4	pK_{a2} 2.03 ± 0.117 pK_{a2} 2.03 ± 0.117 $\lambda_{HSO_4^-}$ 60.1 ± 8.9 $\lambda_{SO_4^-}$ 79.9 ± 0.8	
C₆H₄(COOH)₂[h]	C₆H₄(COOH)₂ (1, 1.2)	5	pK_{a1} 2.935 ± 0.006 λ_{HX^-} 22.5 ± 1.0	pK_{a1} 2.950
	C₆H₄(COOK)(COOH)	6	pK_{a2} 5.381 ± 0.003 $\lambda_{X^{2-}}$ 35.6 ± 2.3	pK_{a2} 5.408
	C₆H₄(COOH)₂ (0.8, 1.0, 1.2)	4	pK_{a1} 2.940 ± 0.081 pK_{a2} 5.390 ± 0.056	λ values na
	C₆H₄(COOK)(COOH) (0.8, 1.0, 1.2)		λ_{HX^-} 41.5 ± 26.8 $\lambda_{X^{2-}}$ 58.9 ± 41.6	
H₂SeO₃	H₂SeO₃ (1, 1.2, 1.5) Na₂SeO₃ (1, 1.2, 1.5)	4 preliminary estimates by 5 and 3′	pK_{a1} 2.635 ± 0.078 pK_{a2} 8.329 ± 0.121 $\lambda_{HSeO_3^-}$ 19.8 ± 7.5 $\lambda_{SeO_3^{2-}}$ 48.6 ± 11.6	pK_{a1} 2.64 pK_{a2} 8.27 $\lambda_{HSeO_3^-}$ na $\lambda_{SeO_3^{2-}}$ na
H₂CO₃	NaHCO₃ (1), H₂CO₃ (1)	1′	pK_{a1} 6.352 ± 0.010 λ_{HCO^-} 40.70 ± 0.42	pK_{a1} 6.35 (6.351–6.366) pK_{a2} 10.33 (10.25–10.33)
	Na₂CO₃ (1.0), Na₂CO₃ (2.0)	3′	pK_{a2} 10.288 ± 0.28 $\lambda_{CO_3^{2-}}$ 64.53 ± 0.757	$\lambda_{HCO_3^-}$ 44.5 $\lambda_{CO_3^{2-}}$ 72
H₂B₄O₇ (?)	Na₂B₄O₇ (0.5, 1.0, 1.5)	3	pK_a 9.278 ± 0.075 λ_{X^-} 11.7 ± 2.9	
H₃PO₄	H₃PO₄ (0.5, 1.0, 1.2, 1.5)	4 estimates by 5′	pK_a 2.278 ± 0.052 $\lambda_{H_2PO_4^-}$ 32.8 ± 9.7	pK_{a1} 2.15 (1.90–2.15) pK_{a2} 7.20 (6.70–7.21) pK_{a3} 12.38 (11.921–12.375)
	NaH₂PO₄ (0.5, 1.5)	6	pK_{a2} 7.217 ± 0.068 $\lambda_{HPO_4^{2-}}$ 54.9 ± 1.3	$\lambda_{H_2PO_4^-}$ 33 $\lambda_{HPO_4^{2-}}$ 57[g]
	Na₂HPO₄ (1.0, 1.2)	6′	pK_a 12.213 ± 0.169 $\lambda_{PO_4^{3-}}$ 75.4 ± 4.1	$\lambda_{PO_4^{3-}}$ 69.0

[a] Uncertainties are based on three to five repeat measurements of the same solution, uncertainties arising from the presence of impurities or gravimetric errors in preparing solutions are not included. [b] The single cited values are taken from a compendium (ref 16). The pK_a ranges given in parentheses are from ref 2 or 3, as appropriate and represent values obtained by different investigators; λ values are all from ref 16. [c] sym-Dibenzo-16-crown-5-oxyacetic acid, see ref 17 for structure. [d] From ref 17. [e] Not available. [f] 3-[N-Morpholino]-2-hydroxy-propanesulfonic acid. [g] From ref 18. [h] o-Phthalic acid.

based ligands for the complexation of alkali metals. A few milliliters of a 1 mM solution of a crown ether carboxylic acid sample, DCOA, was sufficient to determine the pK_a of this compound. The result was in excellent agreement with the reported value of the pK_a of this compound previously obtained much more painstakingly by pH measurements. Determinations based on pH measurements required much greater amounts of sample and exhibited relatively poor precision at the same time.

MOPSO is a member of the zwitterionic Good buffer family (18) used widely as a biological buffer when noncomplexing and/or optically transparent buffers are required (11). The effects of different methods of computation and different measurement solutions are illustrated for this compound. The precision of the determined values as well as the agreement with the literature value improves as more appropriate measurement conditions and computation strategies are selected.

With diprotic acids, considering that the pK_{a2} of H₂SO₄ is near the lower limit of the pK_a values that can be determined by the present approach, the agreement with the literature value is good. The multiparametric fitting approach produces essentially the same pK_a, but the uncertainty limit is higher. With o-phthalic acid, both multiparametric fitting and rigorous computation produce essentially the same set of pK_a values. The λ values are considerably different, however. Both pK_a and λ values exhibit much greater uncertainty with multiparametric fitting. It should also be noted that considerably larger than usual sample volumes were needed to obtain reproducible data on H⁺-exchanged conductance with phthalate, presumably because of strong binding of the un-ionized acid to the poly(styrene–divinylbenzene)-based resin. Should small sample requirements be desired for use with aromatic acids, it will be necessary to use a different ion exchange resin, e.g., the perfluorocarbon cation exchange resin, Nafion. The pK_a values obtained for selenious acid are in excellent agreement

with literature data while the values obtained by multiparametric fitting were not actually markedly different from the preliminary estimates supplied to it. Both pK_a values are determined satisfactorily for H_2CO_3. The experimental data for tetraboric acid are not in agreement with data listed in typical tables. Most compendia continue to list two separate pK values for this acid (16, 19) even though it appears to be well established for some time (20, 21) that tetraboric acid does not exist in solution as such (albeit crystalline tetraborates such as borax are well characterized). The conductivity of a solution of H^+-exchanged $Na_2B_4O_7$, even at a concentration of 0.1 M, is so low that it is impossible for the corresponding acid to have a pK_1 of anywhere close to 4, regardless of the mobility of the resulting monoanion. Our results are consistent with a mole of $Na_2B_4O_7$ behaving as two equivalents of a monoprotic base NaX, where HX has the reported pK_a value. Indeed, attempts to model the situation in terms of a diprotic acid system and applying multiparametric fitting result in best fit values of K_{a1} and K_{a2} that significantly overlap each other, given the uncertainty limits on each. The behavior of borax as a monoprotic base may be explicable with a hydrolytic reaction such as $B_4O_7{}^{2-} + 7H_2O \rightarrow 2[B(OH)_3 \cdot B(OH)_4]^-$. The formation of aggregates between boric acid and borates in solution is well-known (20); note also that our measured pK_a value is essentially indistinguishable from that of boric acid, 9.234–9.237 (2). Regarding the pK_a values determined for H_3PO_4, pK_{a2} is in excellent agreement with the literature value. The agreement of the pK_{a1} and pK_{a3} values is not as good but is probably within an acceptable range. Large uncertainties are observed for the λ values although the mean values are close to literature data. Methods 5′ and 6′ connote the triprotic acid version of these calculation methods.

CONCLUSIONS

We have demonstrated a novel way of combining conductance measurements with complete consideration of the prevalent chemical equilibria to perform measurements of dissociation constants of weak acids. The same principles should be applicable to weak base systems using solutions of their salts and exchanging through OH^--form anion exchangers as necessary. Provided that a pure sample is available, we believe that the present approach offers sigificant advantages over existing methods due to small sample requirement, ease and speed of measurement, and the ease and reliability of calibrating a conductivity detector.

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