Osmotic Coefficients and Mean Activity Coefficients of Uni-univalent Electrolytes in Water at 25 °C

Walter J. Hamer* **and Yung-Chi Wu**

Institute for Basic Standards
National Bureau of Standards, Washington, D.C. 20234

This paper gives values for the osmotic coefficients and mean activity coefficents of uni-univalent electrolytes in aqueous solutions at 25 °C. The values are expressed on the molality or weight basis. The data available in the literature have been corrected to the presently accepted scales of atomic weights (1969) and temperature (IPST 1968) and, where necessary, to the absolute electrical units of 1969 and the fundamental constants of 1963. The selected values of osmotic coefficients and mean activity coefficients for individual electrolytes have been made internally consistent thermodynamically. In some cases estimated values are given; in other cases, references only are given when the data are sparse or unsuited to critical evaluation. Values of the osmotic coefficients and mean activity coefficients of 79 compounds are given together with the standard deviation, variance, and normalized standard deviation of their fit to equations which express these quantities as functions of electrolyte concentration. Finally, literature references are given to data on 51 additional uni-univalent electrolytes.

Key words: Activity coefficients; osmotic coefficients; excess Gibbs energy for electrolytes; critically evaluated data.

Contents

---

*Present address: 3028 Dogwood Street, N.W., Washington, D.C. 20015.

Introduction

This paper gives values for the osmotic coefficients and mean activity coefficients of uni-univalent electrolytes in aqueous solutions at 25 °C. The values are expressed on the molality or weight basis. When the literature gives values for both the osmotic coefficients and the mean activity coefficients in the same investigation, the internal consistency of the two sets of data has been checked by iteration using equations which express the variation of osmotic coefficients and mean activity coefficients with molality. The osmotic coefficient, ϕ, and the mean activity coefficient, γ, are related through the excess Gibbs energy, ΔG^{ex} given by $\Delta G^{ex}= 2\ m\,RT(1-\phi+\ln\gamma)$, where m is molality, R the gas constant, and T the Kelvin temperature. When only osmotic coefficients or mean activity coefficients are reported in the literature, the data given were critically evaluated, and the other function was then calculated by the appropriate equation. The theoretical background of the equations used in analyzing the data is discussed. The strengths and weaknesses of the various experimental techniques used in determining activity coefficients are also discussed.

1. General Relations for Chemical Potentials, Activities, and Activity Coefficients

For closed isolated systems to which no matter is added or withdrawn, thermodynamics gives:

$$dU=TdS-pdV, \tag{1.1}$$

where U is energy, T the Kelvin temperature (K), S the entropy, p the pressure, and V the volume. For an open system to which matter may be added or withdrawn (for example, addition of a salt to water or removal of water from a salt solution by evaporation) one has

$$dU=TdS-pdV+\mu_A dn_A+\mu_B dn_B\ \ldots+\mu_Z dn_Z, \tag{1.2}$$

where the n's represent the number of moles of component A, B, etc., represented by subscripts and the μ's represent the Gibbs' chemical potential of the various components. Since from thermodynamic considerations,

$$G=U-TS+pV=\text{Gibbs energy}, \tag{1.3}$$

$$H=U+pV=\text{Enthalpy (or heat content)}, \tag{1.4}$$

$$A=U-TS=\text{Helmholtz energy}, \tag{1.5}$$

the chemical potential for component a, for example, may be defined in a multiplicity of ways as follows:

$$\mu_A=(\partial U/\partial n_a)_{S,V,n'}=(\partial H/\partial n_a)_{S,p,n'}=(\partial A/\partial n_a)_{T,V,n'}=(\partial G/\partial n_a)_{T,p,n'}. \tag{1.6}$$

Here the subscript n' indicates the constancy of all components except component a. Thus, the chemical potential can be expressed in various ways depending on what variables are held constant during an experiment. The last equality is the one generally used in defining the chemical potential, since experiments may be readily designed wherein the temperature and pressure are maintained constant.

As the energy, entropy, enthalpy (heat content), Helmholtz energy, and Gibbs energy are defined only by differential equations, we can determine only differences of these quantities between two states of a system containing the same quantity of matter. It is therefore customary to tabulate the differences from some standard state, which is equivalent to an arbitrary assumption that the various functions are each zero in the standard states. For general purposes it is customary to assume that the energy, or the enthalpy, of each element at some standard temperature, usually 25 °C, and some standard pressure, usually one atmosphere, and in its most stable form under these conditions is zero. This same standard state may be adopted for the entropy, Helmholtz energy, and Gibbs energy, although the entropy is sometimes taken as zero at a temperature of 0 K in accordance with the third law of thermodynamics. A knowledge of the individual heat capacities and volumes permits the calculation of these quantities under other conditions. See later for the choice of the standard state for electrolytic solutions.

The same considerations apply to the chemical potential, i.e., only differences from an arbitrarily selected standard state can be determined. Throughout this paper the chemical potential is defined in terms of *the partial molal Gibbs energy (fourth equality in equation* 1.6). The difference in chemical potential between two states (compositions) of an ideal solution is given by:

$$\mu_i-\mu_i'=RT\ln\frac{p_i}{p_i'}=RT\ln\frac{x_i}{x_i'}, \tag{1.7}$$

where i denotes component i, and p and x denote, respectively, the vapor pressure and mole fraction. If one of the compositions is the pure component with $x_i'=1$, $p_i'=p_i^\circ$, and $\mu_i'=\mu_i^\circ$, then

$$\mu_i=\mu_i^\circ+RT\ln\frac{p_i}{p_i^\circ}=\mu_i^\circ+RT\ln x_i, \tag{1.8}$$

where the relative vapor pressure of component i is equal to the mole fraction of component i. For real or non-ideal solutions this equation becomes

$$\mu_i=\mu_i^\circ+RT\ln\frac{f_i}{f_i^\circ}=\mu_i^\circ+RT\ln a_i, \tag{1.9}$$

where a_i is the activity of component i and defined as the relative fugacity, and μ_i° is the chemical potential when the activity is unity.

For solutions of electrolytes the standard state is chosen so that the ratio of the activity to the concentration is equal to unity in the limit of $c=0$ where the laws of ideal solutions are obeyed. G. N. Lewis [1][1] called this ratio the activity coefficient; thus $a=x\gamma_x$.

It is more usual, however, to express the composition of a solution either in terms of concentration (molarity), c, i.e., moles of solute (electrolyte) per liter of solution, or as molality, m, in moles of solute per 1000 grams of solvent. In each of these cases, the numerical values of the activity coefficient differ from those expressed on the basis of mole fraction; γ_c and γ_m are used here to denote the activity coefficient on the molarity and molality scales, respectively. The relations between γ_x, γ_c, and γ_m are given by:

$$\gamma_x=\gamma_c(d-0.001\ \mathrm{M}_2c+0.001\ \mathrm{M}_1c\nu)/d_\circ, \quad (1.10)$$

$$\gamma_x=\gamma_m(1+0.001\ \nu m\ \mathrm{M}_1), \quad (1.11)$$

$$c\gamma_c=d_\circ m\gamma_m, \quad (1.12)$$

where d = density of solution, $d_\circ$ = density of solvent, M_1 = molecular weight of solvent, M_2 = molecular weight of solute, and ν = the number of ions into which a molecule of the solute dissociates.

For electrolytic solutions, the activity of an electrolyte that ionizes into ν_+ cations and ν_- anions is defined by:

$$a_{\mathrm{solute}}=a_2=(a_+^{\nu_+})\,(a_-^{\nu_-})=a_\pm^{\nu}, \quad (1.13)$$

where a_+ and a_- are single ion activities (which cannot be measured individually) and $\nu=\nu_++\nu_-$. Therefore, for an electrolytic solution, the mean activity, $a_\pm$, of the solute would be given by:

$$\mu=\mu^\circ+\nu RT\ln a_\pm. \quad (1.14)$$

On the molality scale, $a_\pm=\gamma_\pm m_\pm$ and $a_2=(\gamma_\pm m_\pm)^\nu$ where $m_\pm$ and $\gamma_\pm$ denote, respectively, the mean molality and mean activity coefficient. The chemical potential equals μ° when the solution is at unit activity, on any concentration scale. The mean activity coefficient is defined by:

$$\gamma_\pm=(\gamma_+^{\nu_+}\gamma_-^{\nu_-})^{1/\nu}, \quad (1.15)$$

where γ_+ and γ_- are single ion activity coefficients (which cannot be measured individually), and the mean ionic concentration (in molality, for example), if the electrolyte is completely dissociated, by:

$$m_\pm=(\nu_+^{\nu_+}\nu_-^{\nu_-})^{1/\nu}m=Xm. \quad (1.16)$$

The properties of electrolytic solutions are, in general, directly related to the ionic strength of the solution, defined by:

$$I=1/2\sum_i z_i^2\, m=Ym, \quad (1.17)$$

where z_i is the ionic valence. Henceforth in this paper m and γ will be used to denote the mean quantities of molality and activity coefficient. In this paper which deals exclusively with uni-univalent electrolytes, ν_+, ν_-, z_i, X, and Y all have a value of unity.

The determination of the activity of the solute in terms of known concentrations involves reference to the limit of $m=0$, where the activity and concentration are equivalent.

At finite concentrations the ratio of the activity to the concentration gives a measure of the activity coefficient (it has unit value at $m=0$). Extrapolations to zero concentration are required, therefore, to obtain values free of interionic effects, and the Debye-Hückel [2] theory of interionic attraction, in simple or extended form, is generally used in the extrapolations to give an approximate measure of the interionic effects (or the activity coefficient) which becomes more exact as the concentration approaches zero. For this reason, the Debye-Hückel equations will be discussed in the next section before the experimental methods of obtaining activity coefficients are presented.

2. Theoretical Expressions for Activity Coefficients

Activity coefficients give a measure of the deviations of real solutions from ideality and include the magnitudes of all effects that lead to these deviations. In dilute solutions the main effect is that of interionic attraction, i.e., the attraction between electrical (ionic) charges of unlike sign, and for which Debye and Hückel [2], assuming that ions are point charges, gave:

$$\log\ \gamma_c=-\,|\,z_+z_-\,|\,A_c\sqrt{I_c}, \quad (2.1)$$

where γ is now used, as stated above, to represent $\gamma_\pm$. The subscript c refers to concentrations on the volume basis and the value of the constant A_c is given by:

$$A_c=\left(\frac{2\pi N}{1000}\right)^{1/2}\frac{e^3}{2.302585(4\pi\epsilon_\circ)^{3/2}\,k^{3/2}}\left(\frac{1}{T^{3/2}\epsilon^{3/2}}\right), \quad (2.2)$$

where ϵ denotes the dielectric constant of the solvent in SI units [3],

N = Avogadro constant $=6.02252\ (28)\times10^{26}$ kmol^{-1}
e = elementary charge = $1.60210\ (2)\times10^{-19}\ C$
$\epsilon_\circ$ = permittivity of free space $=8.85417\,(3)\times10^{-12}C^2\mathrm{J}^{-1}\mathrm{m}^{-1}$
k = Boltzmann constant $=1.38054\times10^{-23}$ J/K
$\pi=3.14159265$

and T = Kelvin temperature (defined on the thermodynamic scale by assigning 273.16 K to the triple point of water). The numbers in parentheses in each case represent established limits of error, namely, three standard errors based on the standard deviations of the

---

[1] Figures in brackets indicate literature references in section 8.

data and applied to the last digits in the listed values of the physical constant. These values of the physical constants are those recommended in 1963 by the committee on fundamental constants of the National Academy of Sciences-National Research Council [3]. In SI units, concentrations are expressed in the unit kmol m^{-3}, and are numerically equal to concentrations expressed in units of mol cm^{-3}. A_m on the molality (or weight) basis is given by

$$A_m = A_c (d_\circ)^{1/2} \tag{2.3}$$

and

$$A_m = 0.5108 \text{ at } 25\ {}^\circ\mathrm{C}\ [4].$$

When the size, s_i, of the ions [2] is taken into account, the Debye-Hückel equation becomes:

$$\log \gamma_c = \frac{-|z_+ z_-| A_c \sqrt{I_c}}{1 + B_c s_i \sqrt{I_c}}, \tag{2.4}$$

where

$$B_c = \left(\frac{8\pi N}{1000}\right)^{1/2} \frac{e}{(kT\epsilon)^{1/2}(4\pi\epsilon_\circ)^{1/2}}, \tag{2.5}$$

and where the symbols have the significance given above. B_m on the molality (or weight) basis is given by:

$$B_m = B_c (d_\circ)^{1/2}, \tag{2.6}$$

and $B_m = 0.3287$ at 25 °C [4]. Hückel [5] later modified equation 2.4 by adding an empirical term linear in the concentration or ionic strength. Using molal concentrations and the ionic strength, his equation can be written:

$$\log \gamma_m = \frac{-|z_+ z_-| A_m \sqrt{I_m}}{1 + B_m s_i \sqrt{I_m}} + \beta_m I_m. \tag{2.7}$$

The linear term was intended to take care of the variation of the dielectric constant of the solution with concentration. For convenience, the subscript m is dropped in what follows (however, the molality unit is retained).

In general equation 2.7 represents log γ for solutions more dilute than 1 molal. For solutions more concentrated than 1 molal, higher terms in I are generally required and in this paper the equation:

$$\log \gamma = \frac{-|z_+ z_-| A \sqrt{I}}{1 + B^* \sqrt{I}} + \beta I + CI^2 + DI^3 + \ldots, \tag{2.8}$$

is used to express γ at higher concentrations; here B^* is used to denote Bs_i. The constants C, D, etc., as well as B^* and β are taken as empirical and are not considered to have physical significance; in fact within experimental error values of γ as a function of m are frequently insensitive to the number of terms used in equation 2.8 or to the actual values of these constants. An illustration of this will be given later.

3. Determination of Activity Coefficients

A number of experimental methods have been developed or perfected for the determination of activity coefficients of solutes (electrolytes) in a selected solvent. These are based on:

(1) freezing-point depression,
(2) boiling-point elevation,
(3) vapor-pressure lowering,
(4) isopiestic or vapor-pressure equilibration,
(5) electromotive forces (emfs) of galvanic cells without liquid junction,
(6) emfs of galvanic cells with transference,
(7) solubility,
(8) diffusion.

The first four involve measurements of the escaping tendency of the solvent and the subsequent evaluation of the solute activity through the Gibbs-Duhem equation (see below). The last four give direct measures of the solute activity. The solubility method is generally applicable only to sparingly soluble salts; it is infrequently used and is not considered here. A ninth method involving the measurement of the osmotic pressure of a solution across a semi-permeable membrane is subject to many experimental difficulties. A tenth method involving bithermal equilibration in the vapor phase has not been tested sufficiently to be classed as a precision method. The freezing-point and boiling-point methods are not considered here as this paper is restricted to data at 25 °C.

3.1. Vapor-Pressure Lowering and Osmotic Coefficient

In this method the vapor pressure of the pure solvent and of the solution, or the difference, are measured by a static or dynamic method. In the static method the difference between the vapor pressure of a solution and that of the pure solvent is measured with a differential manometer. In the dynamic method the same air is first passed through water, then a desiccant to absorb the water, then through the solution, and finally through a second desiccant, all immersed in the same thermostat. The activity of the solvent, a_1, is then obtained from the relation:

$$a_1 = \frac{f_1}{f_1^\circ}, \tag{3.1}$$

where f_1 and f_1° are, respectively, the fugacity of the solution and pure solvent. In most cases, the vapor pressures are so low that they may be set equal to the fugacities, so that

$$a_1 = \frac{p_1}{p_1^\circ}, \tag{3.2}$$

---

[2] Frequently, in the literature the ion size is denoted by a_i; s_i is used here to avoid confusion with the ion activity, denoted earlier in this paper by a_i.

where p_1 and p_1° are the vapor pressures, respectively, of the solution and pure solvent. The activity of the solute is then determined from the Gibbs - Duhem equation:

$$d\ln a_2 = \frac{-x_1}{x_2}\, d\ln a_1\,, \tag{3.3}$$

where a_2 denotes the activity of the solute and x_1 and x_2 the mole fraction of the solvent and solute, respectively.

Equation 3.3 is difficult to integrate since x_1/x_2 approaches infinity for the infinitely dilute solution. This problem is overcome by using the special function, known as the practical osmotic coefficient, ϕ_p, defined by Bjerrum [6],

$$\bar{G}-\bar{G}^\circ = -\phi_p RT \sum^{s} m_i/1000 = RT\ln a_1, \tag{3.4}$$

or

$$\bar{G}_1-\bar{G}_1^\circ = -\phi_m RT\nu m M_1/1000 = RT\ln a_1, \tag{3.4$'$}$$

where $\bar{G}_1$ and $\bar{G}_1^\circ$ are the partial molal Gibbs energy of the solvent in solution and standard molal Gibbs energy of the solvent, respectively, R the gas constant, T the Kelvin temperature, m_i the molality of the ith ion, and M_1 the molecular weight of the solvent [1]. For a single electrolyte dissociating into ν ions per molecule $\sum^{s} m_i = \nu m$, and ϕ_p becomes ϕ_m. From equation (3.4)′ it follows that

$$\phi_m = \frac{-1000}{\nu m M_i}\ln a_1, \tag{3.5}$$

where ϕ_m is the molal osmotic coefficient. Hereafter, for convenience the subscript m is dropped. Substitution of equation 3.5 for $\ln a_1$ and $(\gamma m)^\nu$ for a_2 in equation 3.3 and rearrangement give:

$$d\ln\gamma = (\phi-1)\,\frac{dm}{m} + d\phi, \tag{3.6}$$

which upon conversion to common logarithms and integration from 0 to m gives:

$$\log\gamma = 0.4342945(\phi-1) + 0.4342945\int_0^m (\phi-1)\,\frac{dm}{m} \tag{3.7}$$

The integral may be evaluated by plotting values of $(\phi-1)/m$ against m. Equation (3.7) may be altered by substituting $2m^{1/2}dm^{1/2}$ for dm. Then $(\phi-1)/m^{1/2}$ is plotted against $m^{1/2}$ for the evaluation of the integral; this latter procedure is more convenient.

Alternately, the molal osmotic coefficient is given by:

$$\phi = 1 + \frac{2.302585}{m}\int_0^m m\, d\log\gamma, \tag{3.8}$$

and the integral evaluated by plotting m against $\log\gamma$. If $\log\gamma$ is given by equation 2.4, 2.7, or 2.8 (expressed on the m basis), ϕ is then given, respectively, by:

$$\phi = 1 - 2.302585\left\{\frac{|z_+z_-|A}{(B^*)^3 m}\left[(1+B^*\sqrt{m}) - 4.60517\log(1+B^*\sqrt{m}) - 1/(1+B^*\sqrt{m})\right]\right\}, \tag{3.9}$$

or

$$\phi = 1 - 2.302585\left\{\frac{|z_+z_-|A}{(B^*)^3 m}\left[(1+B^*\sqrt{m}) - 4.60517\log(1+B^*\sqrt{m}) - 1/(1+B^*\sqrt{m})\right] - \beta m/2\right\}, \tag{3.10}$$

or

$$\phi = 1 - 2.302585\left\{\frac{|z_+z_-|A}{(B^*)^3 m}\left[(1+B^*\sqrt{m}) - 4.60517\log(1+B^*\sqrt{m}) - 1/(1+B^*\sqrt{m})\right]\right.$$
$$\left. - \beta m/2 - \frac{2}{3}Cm^2 - \frac{3}{4}Dm^3\right\}. \tag{3.11}$$

A and B are Debye-Hückel constants, the values of which for 25 °C are given above in equations (2.3) and (2.6), respectively, (the subscript m is omitted here as a convenience). The other constants may be obtained by the method of least squares. The vapor-pressure lowering method suffers in that precise measurements in dilute solutions, needed for an accurate integration of equation 3.7, are difficult to obtain, and another method must be used for the dilute range.

3.2. Isopiestic or vapor-pressure equilibration

In this method the vapor pressure of a solution of unknown activity is compared or matched against that of a solution of known activity, whereby isopiestic solutions (by distillation or vapor-pressure equilibration) are obtained having the same vapor pressure and the same solvent activity. Obviously, this is a relative method, as the activity of the reference solution must be determined by another method. Aqueous solutions of potassium chloride, sodium chloride, calcium chloride, and sulfuric acid have been selected as standards with the first two best suited for the study of uni-univalent electrolytes except at high concentrations where sulfuric acid may be required as a standard.

Two solutions, A and B for example, are equilibrated until their vapor pressures become equal. The molalities of the two solutions are then determined, generally from changes in known weights of the original solutions. The isopiestic ratio is defined by:

$$\mathscr{R}=\frac{\nu_A m_A}{\nu_B m_B}, \tag{3.12}$$

where A is taken as the reference solution having known vapor pressures as a function of concentration. Usually, it is convenient to plot $\mathscr{R}$ as a function of m_B. When the vapor pressures of solutions A and B are equal

$$\nu_A m_A \phi_A = \nu_B m_B \phi_B, \tag{3.13}$$

or

$$\phi_B = \phi_A \mathscr{R}. \tag{3.14}$$

Accordingly, the osmotic coefficient of solution B, namely ϕ_B, is derived from the known value of ϕ_A and the isopiestic ratio. From ϕ_B the activity coefficients of solution B may be determined by equation 3.7.

If the activity coefficients are known for solution A (the standard) the activity coefficients of the unknown solution B may be obtained as follows:

$$\begin{aligned}\nu_A m_A\, d\ln\gamma_A + \nu_A m_A\, d\ln m_A \\ = \nu_B m_B\, d\ln\gamma_B + \nu_B m_B\, d\ln m_B,\end{aligned} \tag{3.15}$$

as the vapor pressures of the two solutions are equal at equilibrium. Equation 3.15 may be written in terms of $\mathscr{R}$ as:

$$d\ln\gamma_B + d\ln m_B = \mathscr{R}\, d\ln\gamma_A + \mathscr{R}\, d\ln m_A \tag{3.16}$$

$$= d\ln\gamma_A + d\ln m_A + (\mathscr{R}-1)\, d\ln\gamma_A m_A, \tag{3.16'}$$

and therefore,

$$\ln\gamma_B = \ln\gamma_A + \int_0^{m_A} d\ln\frac{m_A}{m_B} + \int_0^{m_A}(\mathscr{R}-1)\, d\ln\gamma_A m_A. \tag{3.17}$$

Now since $\nu_A m_A = \nu_B m_B$ as m_A approaches zero, equation 3.17 becomes

$$\ln\gamma_B = \ln\gamma_A + \ln\mathscr{R} + 2\int_0^{m_A}\frac{(\mathscr{R}-1)}{\sqrt{a_A}}\, d\sqrt{a_A}. \tag{3.18}$$

The integral is evaluated by plotting $(\mathscr{R}-1)/\sqrt{a_A}$ against $\sqrt{a_A}$ and multiplying by 2, or by plotting $(\mathscr{R}-1)/a_A$ against a_A.

Since this method is a relative one, values for the standard or reference solution must be determined by another method or methods, such as the emf methods.

3.3. Electromotive Forces (emfs) of Galvanic Cells without Liquid Junctions

a. Type 1: One Fluid

In this type of cell two electrodes of low solubility are immersed in the solution whose activity coefficient is to be determined; one electrode is reversible to the cation, the other to the anion. This cell may be represented for HCl using hydrogen and silver-silver chloride electrodes, given symbolically by:

$$\text{Pt, H}_2(g)|\text{HCl}(m)|\text{AgCl}(s),\ \text{Ag}(s),$$

where g = gas, m = molality, and s = solid.

The emf, E, of this cell is given by:

$$E = E^\circ - (RT/F)\ln m_\text{H} m_\text{Cl}\gamma_\text{H}\gamma_\text{Cl}, \tag{3.19}$$

where E° is the standard potential for the cell, R is the gas constant, F the faraday, and the other symbols have the significances given above. Equation 3.19 may be rearranged to:

$$E + (RT/F)\ln m_\text{H} m_\text{Cl} + (RT/F)\ln\gamma_\text{H}\gamma_\text{Cl} = E^\circ. \tag{3.20}$$

Values of $[E+(RT/F)\ln m_\text{H} m_\text{Cl}]$ for various values of m_HCl are then plotted against m_HCl and extrapolated to $m_\text{HCl}=0$ where the intercept gives the value of E° as $\gamma_\text{H}\gamma_\text{Cl}=1$ in the limit of $c=0$. Once E° is determined $\gamma_\text{H}\gamma_\text{Cl}$ can be determined by equation 3.19 at various m's for which E is measured.

b. Type 2: Two Fluids

This type of cell is used for those cases where the anode reacts relatively rapidly with the solution. In this case flowing amalgams are used and the amalgams are in contact with the solution for a very short time. This type of cell may be illustrated by:

$$\begin{aligned}&\text{Ag}(s),\ \text{AgCl}(s)|\text{KCl}(m)|\text{K–Hg}| \\ &\qquad\qquad \text{KCl}(m')|\text{AgCl}(s),\ \text{Ag}(s),\end{aligned}$$

which is used to determine activity coefficients of KCl. The emf of this cell is given by:

$$E = \frac{2RT}{F}\ln\frac{\gamma m}{\gamma' m'}. \tag{3.21}$$

As m/m' is known, the measurement of E permits an evaluation of γ/γ'. It is usually customary to keep m constant and at a low value, say 0.1, and vary m'. Then measurements of E at various m's give values of $\gamma_{0.1}/\gamma'$, and $\gamma_{0.1}$ is obtained using some form of the Debye-Hückel theory. Of course if $\gamma_{0.1}$ is known, γ' can be calculated directly from the ratio $\gamma_{0.1}/\gamma'$.

3.4. Emfs of Galvanic Cells with Transference

In this method, cells of the type:

$$\text{Ag}(s),\ \text{AgCl}(s)|\text{KCl}(m)|\text{KCl}(m')|\text{AgCl}(s),\ \text{Ag}(s),$$

are used. The emf of a cell of this type is given by:

$$dE=\frac{-2RT}{F}t_+(d\ln m\gamma), \tag{3.22}$$

where t_+ is the cationic transference number. To obtain γ by this method values of t_+ as a function of m must be known. Longsworth [7] used the following method. He first set

$$t_+=t_r+\Delta t_+, \tag{3.23}$$

with t_r being the transference number at some reference molality, say 0.1 m. Then substitution in equation 3.22 and rearrangement give:

$$\frac{-dE}{t_r}=(2RT/F)^*[d\log m+(\Delta t_+/t_r)d\log m+d\log\gamma + (\Delta t_+/t_r)d\log\gamma], \tag{3.24}$$

where $(2RT/F)^*$ includes the conversion factor in going to common logarithms. Integrating between the 2 molalities m and m_r and rearranging give:

$$-\Delta\log\gamma=\log\gamma-\log\gamma_r=-E/(2RTt_r/F)^* -(\log m-\log m_r)-(1/t_r)\int_{m_r}^{m}\Delta t_+d\log m -(1/t_r)\int_{\gamma_r}^{\gamma}\Delta t_+d(\Delta\log\gamma). \tag{3.25}$$

The first two terms on the right are directly computable from the experimental data (E, m, and m'). The third term is evaluated from a plot of Δt_+ values against values of log m. The fourth term is evaluated by plotting preliminary values of Δ log λ (adding the first three terms on right side of equation 3.25) against Δt_+. This last process is repeated until convergence is attained in the values of Δ log γ and the solution of equation 3.25. This process gives only log $\gamma-\log\gamma_r$ or $-\Delta$ log γ. Now

$$\log\gamma=G-\Delta\log\gamma, \tag{3.26}$$

where G is a constant and equal to log γ_r. If γ_r is known the γ can be obtained at a series of concentrations. If γ_r is not known equation 2.4 (expressed on the m basis) may be used for its evaluation. Substituting the value of log γ_r given by equation 2.4 for G in equation 3.26 and rearranging yield:

$$\Delta\log\gamma+A\sqrt{m}=G+Bs_i(G-\Delta\log\gamma)\sqrt{m}. \tag{3.27}$$

By plotting values of the left side of this equation against values of the second term on the right side, the constants G and Bs_i can be evaluated. The value of G then gives the value of log γ_r and hence γ_r; then γ can be determined at other concentrations. This method is especially suited for very dilute solutions, below 0.1 m, and serves to give values of the activity coefficients of standard solutions used in the isopiestic method.

3.5. Diffusion

As shown by Onsager and Fuoss [8] the diffusion of an electrolyte may be represented by the equation:

$$\mathscr{D}=\nu(1000RT)\left(\frac{\bar{\mathscr{M}}}{c}\right)\left(1+c\frac{d\ln\gamma}{\partial c}\right), \tag{3.28}$$

where $\mathscr{D}$ is the diffusion coefficient, and $\overline{\mathscr{M}}/c$ is given by:

$$\frac{\bar{\mathscr{M}}}{c}=\left[\frac{\lambda_{+\circ}\lambda_{-\circ}\times10^{-7}}{1000\nu_+F^2|z_+|\Lambda_\circ} -\frac{(|z_-|\lambda_{+\circ}-|z_+|\lambda_{-\circ})^2(1+\kappa s)^{-1}\sqrt{I}}{1000|z_+z_-|(\nu_++\nu_-)\Lambda_\circ^2(6\pi N\eta)(\epsilon T)^{1/2}} +\left(\frac{z_-^2\lambda_{+\circ}+z_+^2\lambda_{-\circ}}{(\nu_++\nu_-)^2\Lambda_\circ^2}\right)\frac{\kappa^4c\phi(\kappa s)}{48\pi^2N^2\eta(\epsilon T)^{1/2}}\right] \text{cm}^2\text{s}^{-2}\text{ volt}^{-1}\text{ amp}^{-1},$$

and where ϕ (κs) is given by:

$$\phi(\kappa s)=e^{2\kappa s}Ei(2\kappa s)/(1+\kappa s)^2, \tag{3.29}$$

where Ei denotes the exponential integral function, in which the distance of closest approach, s, is used. Here $\lambda_{+\circ}$ and $\lambda_{-\circ}$ are the limiting equivalent conductances, respectively, of the positive and negative ions, $\Lambda_\circ$ the limiting equivalent conductance of the solution, η is the viscosity of the solution, and κ is the Debye-Hückel quantity given by:

$$\kappa=\left(\frac{8\pi Ne^2}{1000\,\epsilon kT}\right)^{1/2}\sqrt{I}\text{ cm}^{-1}, \tag{3.30}$$

where N, ϵ, k, and T have the significance given above.

Rearrangement of equation 3.28 yields [9]:

$$\frac{\mathscr{D}}{\nu(1000RT)(\bar{\mathscr{M}}/c)}-1\equiv\mathscr{D}'=c\frac{d\ln\gamma}{\partial c}. \tag{3.31}$$

Accordingly,

$$2.302585\log\gamma=\int_0^c\frac{2\mathscr{D}'}{c^{1/2}}dc^{1/2}. \tag{3.32}$$

At the lower limit of concentration

$$\lim_{c\to0}\left[\frac{\mathscr{D}'}{c^{1/2}}\right]=\frac{\partial\ln\gamma}{2\partial c^{1/2}}=-\frac{2.302585}{2}A. \tag{3.33}$$

By plotting $\mathscr{D}'/c^{1/2}$ versus $c^{1/2}$ to the limiting value, equation 3.33, the integral of equation 3.32 may be evaluated and thus log γ and γ. This method is applicable to very dilute solutions and has been so used [9].

4. Literature Data on Osmotic and Mean Activity Coefficients

The data on activity coefficients of electroyltes in aqueous solutions in the literature and in this paper

TABLE 1. Values for the activity coefficients

m	A[a] (3.5)[b]	B (3.4)	C (3.4)	D (3.3.b)	E (3.3.b)	F (3.3.b)	G (3.3.b)	H (3.3.b)	I (3.3.b)	J (*)[c]	K (*)	L (*)	M (*)	N (*)	O (*)	P (3.2)	Q (3.3.b)
0.0005	0.975	[d]*0.9751*															
.001	.9651	*.9653*	0.9650	0.997	0.979	0.997	0.977				0.9663						
.002	.9517	.9520	.9516								.9532						
.005	.9272	.9274	.9270	[e](.927)	.923	.940	.946				.9289						
.01	.9016	.9021	.9015	.872	.890	.907	.922	0.922	0.893		.9033						
.02		.8701	.8694					.892	.864		.8711						
.05		.8171	.8164	.801	.790	.805	.841	.840	.814		.8171		0.842				0.813
.10		.7696	.7692	.751	.745	.759	.796	.794	(.769)	0.766	.7707	0.766	.797	0.779	0.766	0.766	(.767)
.20					.700	.712	.749	.749	.725	.716		.717	.749		.717	.717	.716
.30					.673	.685				.686				.687			.686
.40																	
.50					.638	.650	.682	.682	.661	.650	.6561	.652	.680	.660	.657	.657	.649
.60																	
.70					.618	.629				.627							
.80																	
.90																	
1.0					.593	(.604)	.634	.634	.614	.604	.6126	.609	.632	.598	.608	.608	.604
1.2																	
1.4																	
1.5											.5949	.591	.609	.575			.582
1.6																	
1.8													.605				
2.0														.566	.579	.579	.574
2.5														.568			.569
3.0														.575	.576	.576	.573
3.5																	.577
4.0																	
4.5																	
4.8																	

[a] Notes at bottom of table.

[b] Experimental method (see text).

[c] See Footnotes.

[d] Italic values are calculated ones.

[e] Values in parentheses taken as standard.

[f] Gave molality as 5.

A–Ref. [9].

B–Ref. [11]; concentrations are in moles/liter.

C–T. Shedlovsky [15].

D–D.A. MacInnes and K. Parker [16]; these authors did not give values for γ. Values listed here for γ were calculated from the measured values of a/a' where a and a' are the activities of KCl at two different concentrations; γ for 0.005 m was assumed to be the same as in column C.

E–A. A. Noyes and D. A. MacInnes [17]; these authors calculated the values listed from the measurements of MacInnes and Parker [16] and Beattie [18].

F–Values in column E were corrected by taking the value for γ at 1.0 m to be the same as on column J.

G–G. N. Lewis and M. Randall [19]; a recalculation of the values given by Noyes and MacInnes [17].

H–G. N. Lewis and M. Randall [20]; based on same data as in Ref. [19].

I–Values in column H were corrected by taking the value for γ at 0.10 m to be the same as in Column C.

J–G. Scatchard and S. S. Prentiss [21]; freezing-point data corrected to 25 °C using available heat data (see Ref. 23).

K–H. M. Spencer [22]; freezing-point data corrected to 25 °C using available heat data.

L–Values in column K corrected by R. A. Robinson [23].

M–E. R. Jones and C. R. Bury [24]; freezing-point data corrected to 25 °C using available heat data.

are, for the most part, composite, i.e., several methods have been used in arriving at "best" values. Furthermore, in the relative methods, notably the isopiestic method, adjustments have been made to bring values obtained by various methods into consistency with isopiestic activity ratios. For example, the activity coefficients of dilute solutions of KCl, outside of the concentration range suitable for isopiestic measurements have been determined with high precision in this dilute range from measurements of galvanic cells with transference [10, 11]. Similar emf measurements have been made for NaCl [12, 13] from which the activity coefficients of NaCl were determined. Now the ratio of the activity coefficients of KCl and NaCl obtained from the emf measurements must be consistent with the isopiestic ratio of these two salts.

Once this standardization is achieved the isopiestic method is used effectively to measure osmotic and activity coefficients to high concentrations, in some cases for the saturated solution [14].

How the various methods are used in conjunction to evaluate activity coefficients of a selected electrolyte over a range of concentrations is illustrated here (table 1) for KCl; at the same time the advances in the precision of potassium chloride in water at 25 °C

	R	S	T	U	V	W	X	Y	Z	A′	B′	C′	D′	E′	F′	G′	H′	I′
m	(*)	(3.3.b)	(3.2)	(*)	(3.1)	(3.4)	(3.3.b)	(*)	(3.2)	(3.2)	(3.2)	(*)	(3.2)	(*)	(*)	(*)	(*)	(*)
0.0005	....	....	....	....	....	....	....	....	....	....	....	....	....	....	....	0.9747	0.9749	....
.001	....	0.965	....	....	....	....	....	....	....	....	....	....	....	....	....	.9648	.9650	0.9649
.002	....	....	....	....	....	....	....	....	....	....	....	....	....	....	....	.951	.9517	.9515
.005	....	.926	....	....	....	0.9275	*0.927*	*0.927*	....	....	....	....	....	....	....	....	.9269	.9266
.01	....	.899	....	....	....	.9021	*.901*	*.901*	....	....	....	....	....	....	....	.901	.9009	.9011
.02	....	....	....	....	....	....	....	....	....	....	....	....	....	....	....	.868	.8693	.8689
.05	0.849	.815	....	....	....	.8172	*.815*	.818	....	....	....	....	....	....	....	.816	.8162	.8155
.10	.789	(.764)	0.766	0.766	....	.7701	.769	(.769)	0.770	0.770	0.770	0.769	0.770	0.769	0.770	.769	.7689	.7682
.20	.728	.712	.715	.714	....	.7191	.719	.717	.718	.718	.719	.717	.718	.717	.718	.718	.7177	.7170
.30	.692	....	.684	....	.668	.6875	.688	....	.687	.688	.688	.687	.687	.687	.688	.687	.6871	.6865
.40	.667	....	....	....	.659	....	....	....	....	.666	.667	....	.666	....	.666	.665	.6657	.6653
.50	.652	.644	.648	.646	.651	.6516	.651	.648	.650	.649	.651	.650	.649	.650	.649	.649	.6495	.6492
.60	.640	....	....	....	.642	....	....	....	....	.637	.638	....	.637	....	.637	.636	.6366	.6365
.70	....	....	.624	....	.634	....	.628	....	.627	.626	.628	.626	.626	.626	.626	.626	.6261	.6261
.80	....	....	....	....	.626	....	....	....	....	.618	.619	....	.618	....	.618	.617	.6173	.6174
.90	....	....	....	....	.619	....	....	....	....	.610	.612	....	.610	....	.610	.610	.6098	.6101
1.0	(.605)	.597	.602	.599	.613	.6069	.606	.601	.607	.604	.606	.605	.604	.605	.604	.603	.6035	.6038
1.2	....	....	....	....	.603	....	....	....	....	.593	.596	....	.593	....	.593	....	.5932	.5937
1.4	....	....	....	....	.595	....	....	....	....	.586	.587	....	.586	....	.586	....	.5854	.5861
1.5	....	.576	.581	.577	.592	....	.585	.581	.585	....	....	.585	....	.585	....	.582	.5822	....
1.6	....	....	....	....	.589	....	....	....	....	.580	.580	....	.580	....	.580	....	.5795	.5803
1.8	....	....	....	....	.586	....	....	....	....	.576	.575	....	.576	....	.576	....	.5750	.5759
2.0	.574	.569	.572	.570	.584	.5781	.576	.574	.577	.573	.571	.575	.573	.575	.573	.572	.5717	.5726
2.5	....	.568	.569	.569	.584	....	.572	.572	.574	.569	.567	.572	....	.572	.569	....	.5674	.5684
3.0	.569	.571	.570	.572	.589	.5741	.571	.575	.575	.569	.567	.573	.569	.573	.569	.568	.5672	.5682
3.5	....	.571	.573	.572	.597	.5764	.574	.575	.578	.572	.570	.576	....	.576	.572	....	.5700	.5712
4.0	.580	.581	.578	.583	.607	.5806	.579	.585	.583	.575	.575	.582	.577	.582	.577	.576	5750	.5765
4.5	....	....	....	....	....	....	....	....	.591	.583	.583	.590	....	.590	.585	....	.5814	.5838
4.8	....	....	....	....	....	....	....	....	....	....	.588	.595	.588	....	.588	.590[f]	.5860	.588

N–H. S. Harned [25]; calculations from freezing points, emfs, and vapor pressures.

O–Landolt-Bőrnstein's "Tabellen", 5th Ed, Zweiter Ergänzungsband, p 1112 (1931); calculated by O. Redlich and P. Rosenfeld from *freezing-point data and molal heats of dilution.*

P–R. A. Robinson [23]; isopiestic measurements.

Q–R. Caramazza [26]; amalgam concentration cells.

R–C. Scatchard [27]; calculated values of γ for KCl at 20 °C from the vapor-pressure data of B. F. Lovelace, J. C. W. Frazer, and V. B. Sease [28] and corrected to 25 °C by taking γ as 0.605 for 1.0 *m*.

S–H. S. Harned [29]; amalgam concentration cells.

T–R. A. Robinson and D. A. Sinclair [30]; got their standard for KCl from freezing-point, boiling-point, vapor-pressure, emf, and heat data.

U–*Recalculation by R. A. Robinson [23] of data of H. S. Harned* [29].

V–Z. Shibata and K. Niva [31].

W–T. Shedlovsky and D. A. MacInnes [10].

X–H. S. Harned and M. A. Cook [32].

Y–Recalculation here of data of H. S. Harned [29].

Z–R. A. Robinson [33].

A′–*R. A. Robinson and R. H. Stokes [34].*

B′–Ref. [14].

C′–R. A. Robinson and H. S. Harned [35]; review.

D′–R. A. Robinson [36].

E′–H. S. Harned and B. B. Owen [37]; compilation.

F′–R. A. Robinson and R. H. Stokes [38]; compilation.

G′–K. S. Pitzer and L. Brewer [39]; calculated values from data in Refs. [11] and [10] and emf and isopiestic data on KCl and NaCl; adjusted data of Ref. [38] by adding 0.0004 to all values of $-\log \gamma$ and 0.0020 to all ϕ values.

H′–M. H. Lietzke *and R. W. Stoughton [40]; method of evaluation* used herein.

I′–Present evaluation.

with which the values are obtained is shown. The data are not given in time (or date) sequence but rather in order of concentration range for which the method was used or was suitable. It will be noted that the final values are arrived at from a combination of the results of all five experimental methods discussed in section 3 of this paper. The method used in the final evaluation is given in the next section.

5. Evaluation of Data on Osmotic and Mean Activity Coefficients

In critically evaluating the literature data on activity coefficients and osmotic coefficients all data were made consistent with the ^{12}C scale of atomic weights; as this change corresponds to only 0.004 percent, adjustments from the oxygen scale to the ^{12}C scale were necessary only in rare cases. Furthermore, where necessary, the data were corrected to the presently accepted values for the ice point (273.15) and the Faraday (96487.0 ± 1.6 coulombs g-equivalent $^{-1}$), and all electrical data were converted to the "absolute" electrical units of 1969. It is estimated that 95 percent of the literature has been covered.

Earlier data have been critically reviewed and compiled in the International Critical Tables by Randall [41]. Later data have been assembled by Harned and Owen [37] and Robinson and Stokes [38] in their monographs; graphical methods were used for the most part in their evaluations. The data given herein are based mainly on these latter compilations and have been extended by more recent experimental data. In section 6.2 which follows, additional data to that given by Harned and Owen and Robinson and Stokes are given for those compounds underlined. Data for those compounds marked with an *n* are not given in these two reference sources.

The data based on isopiestic measurements were evaluated, using the method of least squares, by first applying equations (3.9), (3.10), or (3.11) to the values of ϕ obtained by the experimenters from smoothed data. The constants B^*, β, C, D, etc. so obtained were then used in equation (2.8) to obtain γ. In some cases, B^* was first set equal to 1.0, 1.5, etc., as has been suggested by various authors [4] and the other constants obtained by the method of least squares. In those cases where the experimenter reported values for ϕ and derived values of γ, the internal consistency of the two sets of values was checked by using equations (3.9), (3.10), (3.11) etc., until deviations from the input values of ϕ and γ were at a minimum. Values of ϕ and γ were related through the excess Gibbs energy given by:

$$\Delta G^{\rm ex} = \nu m RT(2.302585)\left\{\frac{A}{(B^*)^3}\left[(2-B^*_m\sqrt{m})B^*_m\sqrt{m}-2\log(1+B^*_m\sqrt{m})\right]+\beta^2 m/2+Cm^3/3+Dm^4/4+\ldots\right\}. \quad (4.1)$$

Where independent values of γ were available, they were used to check the process of obtaining γ from isopiestic measurements of ϕ.

In those cases where the activity coefficients were the primary data the reverse procedure was followed. In those cases where only γ or ϕ was reported, iteration was still used to bring the primary data and the derived quantities into good fits with equations (2.8) and (3.11).

Although it has been advanced that it would be better to select a common or standard value for B^* owing to the cross-differentiation relation for a mixture of electrolytes [38, 39, 42, 43], B^* has nevertheless been allowed to vary in order to obtain the best fit for experimental values of γ and ϕ through the excess Gibbs energy relation. The cross-differentiation relation requires a common B^* if equation 2.4 is used to represent activity coefficients; more leeway is permitted in the selection of values for B^* if equations 2.7 or 2.8 are used.

As stated in section 2, values of γ and ϕ as functions of m are frequently insensitive to the number of terms used in equation 2.8 or to the actual values of those constants. These facts may be illustrated for data on silver nitrate given in table 2. Although the values in columns 2 and 3 are "best" values owing to the low standard deviations of fit for γ and ϕ, the values given in the other columns are not greatly different, when rounded to the third decimal, from the observed or experimental values. In other words, the various sets of constants given in table 2 reproduce the activity coefficients and osmotic coefficients within the accuracy of the experimental data.

When values for only γ or ϕ are reported in the literature, checks on the internal consistency of γ and ϕ are not possible. In these cases γ or ϕ was evaluated by procedures discussed under each compound, and iteration was followed as described above to obtain the best fits to equations (2.8) and (3.11).

The standard deviations, $\sigma(\phi)$ and $\sigma(\gamma)$, and the variances, VAR(ϕ) and VAR(γ), of fit to equations 3.11 and 2.8, respectively, are given at the bottom of each table in section 6.2 which follows. Since data for some compounds cover a wide range of concentrations while for others the data are restricted to a smaller range of concentrations, "normalized" standard deviations, $\sigma^*(\phi)$ and $\sigma^*(\gamma)$, defined by

$$\sigma^* = \Sigma[(100\Delta/\text{predicted values})^2]/100,$$

are also given; Δ = predicted value − observed value.

Although the standard deviations and "normalized" standard deviations give a measure of the overall fit of equations (2.8) and (3.11) to the experimental data and derived quantities, deviation plots could also have been given to indicate the extent of the agreement of calculated quantities with original data throughout the concentration range. However, it was not possible to include such deviation plots here. Agreement of data herein presented with original data may be ascertained by reference to the literature which is cited here. In general the equations agree with the input data within the experimental error, except for the very highest concentrations, where the experimental uncertainty is greater.

In the tables, values of ϕ and γ are given from m = 0.001 to the highest molality for which data are available. In some cases, the experimental data extend over the entire concentration range; in other cases the values for the lower concentrations are those given by the equations which fit the experimental data at higher concentrations. In each case (table) a line is drawn if necessary, to show the concentration above which experimental data are available.

TABLE 2. Values of the osmotic and mean activity coefficients of $AgNO_3$

m	"best" values		ϕ	γ	ϕ	γ	ϕ	γ
	ϕ	γ						
0.001	0.9880	0.9643	0.9880	0.9643	0.9879	0.9642	0.9880	0.9642
.002	.9833	.9503	.9832	.9502	.9832	.9502	.9832	.9501
.005	.9742	.9238	.9741	.9237	.9740	.9235	.9740	.9235
.01	.9645	.8958	.9644	.8956	.9642	.8952	.9642	.8952
.02	.952	.859	.951	.859	.951	.858	.951	.858
.05	.928	.794	.928	.794	.927	.793	.927	.792
.1	.904	.732	.903	.731	.902	.729	.903	.729
.2	.872	.656	.872	.656	.871	.653	.870	.653
.3	.849	.606	.849	.605	.848	.603	.847	.603
.4	.829	.567	.829	.567	.829	.565	.829	.565
.5	.812	.536	.813	.536	.812	.534	.812	.534
.6	.797	.509	.797	.509	.797	.508	.797	.507
.7	.782	.486	.783	.486	.784	.485	.784	.485
.8	.769	.465	.770	.465	.771	.465	.771	.464
.9	.756	.447	.757	.447	.759	.446	.759	.446
1.0	.744	.430	.745	.430	.747	.430	.747	.430
1.2	.721	.400	.723	.401	.725	.401	.726	.401
1.4	.701	.375	.702	.376	.705	.376	.706	.376
1.6	.681	.353	.683	.354	.687	.355	.687	.355
1.8	.663	.334	.665	.334	.669	.335	.670	.335
2.0	.647	.316	.647	.317	.652	.318	.653	.318
2.5	.609	.280	.609	.281	.614	.282	.615	.282
3.0	.577	.252	.576	.252	.580	.253	.581	.254
3.5	.549	.229	.547	.229	.550	.230	.550	.230
4.0	.524	.210	.522	.210	.523	.210	.523	.211
4.5	.503	.194	.501	.194	.499	.194	.499	.194
5.0	.484	.181	.483	.180	.478	.180	.478	.180
5.5	.467	.169	.466	.169	.460	.168	.460	.168
6.0	.452	.159	.452	.159	.444	.158	.443	.158
7.0	.427	.142	.428	.142	.418	.141	.417	.141
8.0	.405	.129	.407	.129	.400	.128	.398	.127
9.0	.386	.117	.387	.118	.386	.117	.385	.117
10.0	.370	.108	.369	.108	.377	.109	.376	.109
11.0	.357	.101	.355	.100	.369	.102	.371	.102
12.0	.350	.0945	.348	.0942	.363	.0957	.367	.0961
13.0	.351	.0897	.357	.0903	.356	.0903	.364	.0911
B^*	0.950		0.900		0.835		0.825	
β	-1.1493×10^{-1}		-1.0407×10^{-1}		-9.3685×10^{-2}		-9.1936×10^{-2}	
C	1.0846×10^{-2}		6.638×10^{-2}		5.6009×10^{-3}		5.2927×10^{-3}	
D	-6.8460×10^{-4}		1.0609×10^{-4}		-1.4281×10^{-4}		-1.2585×10^{-4}	
E	1.8730×10^{-5}		-4.7637×10^{-5}					
F			2.0147×10^{-6}					
$\sigma(\phi)$	0.000698		0.000810		0.004312		0.000861	
$\sigma(\gamma)$	0.000129		0.000181		0.001662		0.000314	
VAR(ϕ)	4.872×10^{-7}		6.561×10^{-7}		1.859×10^{-5}		7.413×10^{-7}	
VAR(γ)	1.664×10^{-8}		3.276×10^{-8}		2.762×10^{-6}		9.860×10^{-8}	
$\sigma^*(\phi)$	0.00135		0.00157		0.008369		0.00167	
$\sigma^*(\gamma)$	0.000378		0.000531		0.004873		0.000921	

6. Compounds

6.1. Index to Compounds

In this section and the next one, compounds are considered in the following order:

Table	Compound	Table	Compound	Table	Compound
		30	KI	59	Na *p*-tol
		31	KOH	60	K *p*-tol
3	HF[n]	32	$KClO_3$	61	Na formate
4	HCl	33	$KBrO_3$	62	Li acetate
5	HBr[c]	34	KNO_3	63	Na acetate
6	HI	35	KH_2PO_4	64	K acetate
7	$HClO_4$	36	KH_2AsO_4	65	Rb acetate
8	HNO_3	37	KCNS	66	Cs acetate
9	LiCl	38	KPF_6[n]	67	Tl acetate
10	LiBr	39	RbF[n]	68	Na propionate
11	LiI	40	RbCl	69	Na butyrate
12	LiOH	41	RbBr	70	Na valerate
13	$LiClO_4$	42	RbI	71	Na caproate
14	$LiNO_3$	43	$RbNO_3$	72	Na heptylate
15	NaF	44	CsF[n]	73	Na caprylate
16	NaCl	45	CsCl	74	Na pelargonate
17	NaBr	46	CsBr	75	Na caprate
18	NaI	47	CsI	76	Na H malonate
19	NaOH	48	CsOH	77	Na H succinate
20	$NaClO_3$	49	$CsNO_3$	78	Na H adipate
21	$NaClO_4$	50	$AgNO_3$[b]	79	KH malonate
22	$NaBrO_3$	51	TlCl[n]	80	KH succinate
23	$NaNO_3$	52	$TlClO_4$	81	KH adipate
24	NaH_2PO_4	53	$TlNO_3$	*	$NaBO_2$[n]
25	NaH_2AsO_4	54	$TlNO_2$[n]	*	$NaBF_4$[n]
26	NaCNS	55	NH_4Cl	*	$NaHCO_3$[n]
27	KF	56	NH_4ClO_4[n]	*	AgF[n]
28	KCl[a]	57	NH_4NO_3	*	TlN_3[n]
29	KBr	58	Li *p*-tol[d]	*	Additional compounds referenced only.[n]

[a] Table 1 gives additional KCl data.
[b] Table 2 gives additional $AgNO_3$ data.
[c] Those underlined include extensions of data in Refs. 37 and 38.
[d] Para-toluenesulfonate.
[n] Not given in Refs. 37 and 38.
* Referenced only; see section 6.2 on compounds.

6.2. Compounds Evaluated

The recommended data are given in tables in the above order. References to experimental measurements follow each table; emphasis in the evaluation was given to references italicized.

HF

The stoichiometric mean activity coefficients of HF have been determined at 25 °C and from 0.001 to 1.0 *m* by Broene and De Vries [*44*] from the emfs of the cell:

$$\text{Pb-Hg}|\text{PbF}_2(s)|\text{HF}(m)|\text{H}_2(g),\ \text{Pt}.$$

Anthony and Hudleston [45]; Kendall, Booge, and Andrews [46]; and Paterno and Peratoner [47] measured the freezing-point depression of HF from 0.025 to 4.14, 0.266 to 2.882, and 0.114 to 5.765 *N*, respectively, and Parker [48] summarized values for the molal heat capacity and heats of dilution from 0 to 27.753 *m*. The vapor pressures of aqueous solutions of HF have been measured by Fredenhagen and Wellman [49] from 0 to 20 *N*; Brosheer, Lenfesty, and Elmore [*50*] from 1 to 20 *m*; General Electric Company [51] from 30 to 100 percent; and Vdovenko, Lazarev, and Shirvinskaya [52] from 0.5 to 5.0 *m*. Elmore, Hatfield, Mason, and Jones [*53*] calculated values of the solvent activity from 0.001 to 1.0 *m* from conductivity data [54], the dissociation constants of Wooster [55], and Kielland's [56] values for the activity coefficients of the hydrogen and fluoride ions; their value for log *a* (solvent) at 1.0 *m* differs by 0.0001 from that calculated by their equation for the data of Brosheer, Lenfesty, and Elmore. Hamer and Wu [*57*] calculated values of the stoichiometric mean activity coefficient from 0 to 4 *m* from recalculated dissociation constants [58], from emf data, and from freezing-point depressions and apparent molal heat capacity data. They also gave values for the ionic activity coefficients. Weight is given to emf measurements [*44*, *57*] for dilute solutions and to vapor-pressure measurements [*50*, *53*] for the more concentrated solutions.

TABLE 3. Osmotic and mean activity coefficients of HF

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.792	.551	.800	.513	.0279	6.000	.511	.0102
.002	.715	.429	.900	.511	.0263	7.000	.524	.00955
.003	.685	.368	1.000	.509	.0249	8.000	.537	.00909
.005	.654	.302	1.200	.507	.0227	9.000	.550	.00873
.007	.634	.262	1.400	.504	.0210	10.000	.564	.00845
.010	.614	.225	1.600	.502	.0196	11.000	.578	.00822
.020	.580	.163	1.800	.500	.0184	12.000	.592	.00804
.050	.551	.106	2.000	.498	.0175	13.000	.606	.00789
.100	.539	.0766	2.500	.494	.0155	14.000	.620	.00778
.200	.530	.0550	3.000	.491	.0141	15.000	.635	.00769
.300	.526	.0452	3.500	.489	.0130	16.000	.649	.00763
.400	.522	.0393	4.000	.487	.0122	17.000	.663	.00758
.500	.520	.0352	4.500	.493	.0115	18.000	.678	.00755
.600	.517	.0322	5.000	.499	.0110	19.000	.693	.00753
.700	.515	.0296	5.500	.505	.0105	20.000	.707	.00752

$$\phi = a + bm^{-3/2} + cm^{1/2} + dm$$

$$\log \gamma = 0.4342945[a - 1 + 1/3\ bm^{-3/2} + 3cm^{1/2} + 2dm] + (a - 1) \log m + I$$

concentration range

	0.001 – .05m	0.05 – 0.5m	0.5 – 4.0m	4.0 – 20.0m
a =	0.76331	0.54732	0.54932	0.45868
b =	2.7429×10^{-6}	1.5151×10^{-4}	1.5151×10^{-4}	1.3351×10^{-2}
c =	−1.98157	−0.0458	−0.0488	−0.020858
d =	4.6108	0.008825	0.008825	0.017092
I =	−0.801	−1.3543	−1.3518	−1.3596
$\sigma(\phi)$ =	0.00250	0.00093	0.00086	0.00983
$\sigma(\gamma)$ =	0.00132	0.00121	0.00003	0.00013
VAR(ϕ) =	6.25×10^{-6}	8.65×10^{-7}	7.40×10^{-7}	9.66×10^{-4}
VAR(γ) =	1.74×10^{-6}	1.46×10^{-6}	$9. \times 10^{-10}$	1.69×10^{-8}
$\sigma^*(\phi)$ =	3.83×10^{-3}	1.75×10^{-3}	1.71×10^{-3}	1.68×10^{-2}
$\sigma^*(\gamma)$ =	4.39×10^{-3}	2.03×10^{-2}	1.36×10^{-3}	1.46×10^{-2}

HCl

As early as 1898 Dolezalek [59] used hydrogen and chlorine electrodes to study the properties of HCl at 30 °C. In 1900, Jahn [60] used silver-silver chloride electrodes to study HCl concentration cells at 18 °C and Gahl [61] made measurements of the vapor-pressure lowering on HCl solution with uncertain accuracy. Measurements of cells without liquid junction and with H_2 and Hg, HgCl electrodes have been made at 25 °C by Ellis (0.003–4.5 *m*) [62], Linhart (0.01–16 *m*) [63], Getman (0.001–1.0 *N*) [64], Duboux and Rochat (1–4.1 *m*) [65], Hills and Ives (0.0016–0.12 *N*) [66], Grzybowski (0.001–0.02 *m*) [67], and Schwabe and Ziegenbalg (0.0014–0.1 *m*) [68], and with transference by Lewis, Brighton, and Sebastian (0.01–0.10 *N*) [69], and Gupta, Hills, and Ives (0.004–1 *m*) [70]. Many more studies have been made with hydrogen and silver-silver chloride electrodes cells without transference by Noyes and Ellis (0.01–4.5 *m*) [71], Linhart (0.0001–0.05 *m*) [72], Scatchard (0.01–1.5 *m*) [73], Nonhebel (0.00015–0.04 *m*) [74], Güntelberg (0.01–1.0 *m*) [75], Roberts (0.0016–0.01 *m*) [76], Carmody (0.003–0.1 *m*) [77], Harned and Ehlers (0.005–0.1 *m*) [78], Harned and Ehlers (0.0042–4 *m*), [79], Anderson (0.00002–0.003 *m*) [80], Åkerlöf and Teare (3–16 *m*) [81], Bates and Bower (0.001–0.1 *m*) [82], and Zielen (0.001—4 *m*) [83], and with transference by Wilke and Schränkler (1–6 *N*) [84], Shedlovsky and MacInnes (0.0035–0.1 *N*) [85], Giber, Lengyel, Tamas, and Tahi (0.815–15.43 *m*) [86], and Faita and Mussini (0.0056–9.25 *m*) [87]. Covington and Prue [88] made measurements with glass and silver-silver chloride electrodes from 0.005–0.1 *m* and Lohonyai and Proszt [89] investigated cells with transference from 0.001–1 *m* using quinhydrone electrodes. Harpst, Holt, and Lyons [90] used the diffusion method for dilute solutions (0.0063–0.05 *N*). The vapor-pressure lowering of HCl was measured by Bates and Kirschman (3.24–10 *m*) [91], Dobson and Masson (0–10 *m*) [92], Dunn and Rideal (0.35–5.3 *m*) [93], Haase, Naas, and Thumm (0.1–16 *m*) [94], and Ionin and Kurina (16–19.35 *m*) [95]. The activity coefficient of HCl in HCl-salt mixtures has also been studied extensively [see Ref. 37]; these studies, however, are not given weight in determining the activity coefficients of aqueous solutions of pure HCl. Freezing-point depressions are not considered here. Weight is given to results obtained at 25 °C from galvanic cells without liquid junction.

TABLE 4. Osmotic and mean activity coefficients of HCl

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.800	1.011	.785	5.000	1.682	2.38
.002	.984	.952	.900	1.025	.797	5.500	1.766	2.77
.005	.977	.929	1.000	1.038	.811	6.000	1.849	3.23
.010	.970	.905	1.200	1.067	.842	7.000	2.011	4.38
.020	.961	.876	1.400	1.095	.877	8.000	2.165	5.90
.050	.950	.832	1.600	1.125	.917	9.000	2.309	7.88
.100	.944	.797	1.800	1.155	.961	10.000	2.441	10.40
.200	.946	.768	2.000	1.186	1.009	11.000	2.560	13.52
.300	.953	.758	2.500	1.265	1.148	12.000	2.667	17.32
.400	.963	.756	3.000	1.346	1.316	13.000	2.765	21.9
.500	.974	.759	3.500	1.429	1.517	14.000	2.854	27.4
.600	.986	.765	4.000	1.513	1.757	15.000	2.941	34.0
.700	.998	.774	4.500	1.597	2.040	16.000	3.030	42.3

$B^* = 1.525$ $\beta = 1.0494 \times 10^{-1}$ $C = 6.5360 \times 10^{-3}$ $D = -4.2058 \times 10^{-4}$

$E = -4.0700 \times 10^{-6}$ $F = 5.2580 \times 10^{-7}$

$\sigma(\phi) = 2.38 \times 10^{-3}$ $\sigma(\gamma) = 3.22 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 5.643 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 1.034 \times 10^{-5}$

$\sigma^*(\phi) = 1.40 \times 10^{-3}$ $\sigma^*(\gamma) = 2.33 \times 10^{-3}$

HBr

Measurements of cells without liquid junction and with hydrogen and silver-silver bromide electrodes have been made at 25 °C by Lewis and Storch (0.01–0.1 *m*) [96], Livingston (0.08–1.55 *m*) [97], Keston (0.0001262–0.003719 *m*) [98], Harned, Keston, and Donelson (0.001–1 *m*) [99], Bierman and Yamaski (0.5–4 *m*) [100], Bienert (2–11 *m*) [101], Hetzer, Robinson, and Bates (0.005–0.1 *m*) [102], and Faita, Mussini, and Ozzioni (0.055–5.5 *m*) [103], and with transference by Wilke and Schränkler (1–9 *m*) [84]. Gupta, Hills, and Ives measured cells without transference and with hydrogen and mercury-mercurous bromide electrodes for the concentration range 0.005–0.2 *m* [104] and Mussini and Pazzoli made measurements on cells with liquid junction for 0.1–1*m* using silver-silver bromide and calomel electrodes [*105*]. Bates and Kirschman [91] gave results of vapor-pressure measurements for the concentration range, 5.6–9.286 *m*.

TABLE 5. Osmotic and mean activitv coefficients of HBr

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.966	.700	1.022	.816	4.000	1.701	2.48
.002	.985	.953	.800	1.039	.833	4.500	1.822	3.06
.005	.977	.930	.900	1.055	.851	5.000	1.947	3.80
.010	.970	.907	1.000	1.072	.872	5.500	2.073	4.75
.020	.963	.879	1.200	1.107	.918	6.000	2.200	5.95
.050	.953	.837	1.400	1.143	.970	7.000	2.452	9.39
.100	.949	.806	1.600	1.181	1.029	8.000	2.692	14.7
.200	.954	.783	1.800	1.219	1.094	9.000	2.909	22.6
.300	.965	.778	2.000	1.258	1.167	10.000	3.088	33.4
.400	.978	.782	2.500	1.361	1.385	11.000	3.214	46.5
.500	.992	.790	3.000	1.470	1.665			
.600	1.007	.802	3.500	1.583	2.022			

$B^* = 1.6468$ $\beta = 1.2457 \times 10^{-1}$ $C = 8.8530 \times 10^{-3}$ $D = 2.4750 \times 10^{-5}$

$E = 3.7190 \times 10^{-5}$

$\sigma(\phi) = 8.29 \times 10^{-3}$ $\sigma(\gamma) = 2.83 \times 10^{-1}$ $\mathrm{VAR}(\phi) = 6.873 \times 10^{-5}$ $\mathrm{VAR}(\gamma) = 7.991 \times 10^{-2}$

$\sigma^*(\phi) = 3.45 \times 10^{-3}$ $\sigma^*(\gamma) = 9.87 \times 10^{-3}$

HI

Measurements of cells without liquid junction and with hydrogen and silver-silver iodide electrodes have been made at 25 °C by Noyes and Freed (0.03–0.1 *m*) [106], Pearce and Fortsch (0.005–0.25 *m*) [107], Küppers (3.5–10 *m*) [108], Hetzer, Robinson, and Bates (0.017–0.95 *m*) [*109*], and Kortum and Hausserman (0.00009–0.0587 *m*) [*110*] and with transference by Wilke and Schränkler (0.2–0.65 *N*) [84]. Bates and Kirschman measured the vapor-pressure lowering for the concentration range, 5.97 –9.77 *m* [*91*], and Harned and Robinson made isopiestic measurements from 0.1 to 3 *m* [*111*].

TABLE 6. Osmotic and mean activity coefficients of HI

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.966	.600	1.039	.865	3.000	1.531	2.002
.002	.985	.953	.700	1.057	.888	3.500	1.646	2.457
.005	.978	.931	.800	1.076	.913	4.000	1.765	3.04
.010	.972	.909	.900	1.095	.940	4.500	1.888	3.79
.020	.965	.884	1.000	1.115	.969	5.000	2.016	4.76
.050	.958	.847	1.200	1.154	1.033	5.500	2.146	6.01
.100	.958	.823	1.400	1.193	1.104	6.000	2.279	7.63
.200	.970	.811	1.600	1.234	1.182	7.000	2.548	12.4
.300	.985	.816	1.800	1.274	1.268	8.000	2.809	20.1
.400	1.002	.828	2.000	1.315	1.363	9.000	3.048	32.1
.500	1.020	.845	2.500	1.421	1.645	10.000	3.247	49.1

$B^* = 1.90$ $\beta = 1.6188 \times 10^{-1}$ $C = -2.8100 \times 10^{-4}$ $D = 1.1544 \times 10^{-3}$

$E = -8.2330 \times 10^{-5}$

$\sigma(\phi) = 5.23 \times 10^{-3}$ $\sigma(\gamma) = 1.56 \times 10^{-1}$ $\mathrm{VAR}(\phi) = 2.732 \times 10^{-5}$ $\mathrm{VAR}(\gamma) = 2.419 \times 10^{-2}$

$\sigma^*(\phi) = 2.63 \times 10^{-3}$ $\sigma^*(\gamma) = 8.87 \times 10^{-3}$

$HClO_4$

Covington and Prue [*112*] measured the emf of cells with transference and with hydrogen electrodes from 0.01 to 0.10 *m* at 25 °C. Pearce and Nelson [113] measured the vapor-pressure lowering from 0 to 12 *m*, and Robinson and Baker [*114*] and Haase, Dücker, and Küppers [*115*] made isopiestic measurements from 0.1 to 16 *m*.

TABLE 7. Osmotic and mean activity coefficients of $HClO_4$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.966	.800	1.016	.798	5.000	1.858	3.100
.002	.985	.953	.900	1.030	.811	5.500	1.98	3.83
.005	.977	.929	1.000	1.045	.826	6.000	2.11	4.76
.010	.970	.906	1.200	1.075	.861	7.000	2.37	7.46
.020	.962	.878	1.400	1.106	.901	8.000	2.63	11.87
.050	.952	.836	1.600	1.139	.946	9.000	2.90	19.08
.100	.947	.803	1.800	1.174	.998	10.000	3.17	30.8
.200	.949	.776	2.000	1.209	1.055	11.000	3.43	49.9
.300	.957	.767	2.500	1.303	1.226	12.000	3.68	80.4
.400	.967	.766	3.000	1.404	1.445	13.000	3.93	129.
.500	.978	.769	3.500	1.510	1.724	14.000	4.17	205.
.600	.990	.777	4.000	1.621	2.078	15.000	4.40	323.
.700	1.003	.786	4.500	1.738	2.528	16.000	4.61	501.

$B^* = 1.7125$ $\beta = 9.2552 \times 10^{-2}$ $C = 1.3495 \times 10^{-2}$ $D = -6.2113 \times 10^{-4}$

$E = 8.6809 \times 10^{-6}$

$\sigma(\phi) = 2.36 \times 10^{-3}$ $\sigma(\gamma) = 2.48 \times 10^{-1}$ $VAR(\phi) = 5.583 \times 10^{-6}$ $VAR(\gamma) = 6.145 \times 10^{-2}$

$\sigma^*(\phi) = 1.54 \times 10^{-3}$ $\sigma^*(\gamma) = 2.21 \times 10^{-3}$

HNO_3

Stonehill [116] measured the emf of cells of the type:

$Pt|Q(sat), HNO_3\ (m, \text{fixed})|HNO_3\ (m, \text{variable}), Q(sat)|Pt,$

where Q = quinhydrone. He varied the concentration of HNO_3 from 0.001021 to 0.204 *m*. Corrington and Prue [112] used glass electrodes in cells with transference and varied the concentration of HNO_3 from 0.01 to 0.10 *m*. Flatt and Bengnerel [*117*] measured the liquid-vapor equilibrium for the binary system HNO_3–H_2O for liquid-phase compositions of 0 to 68 percent HNO_3 while Davis and De Bruin [*118*] measured the partial pressures of HNO_3 for solution concentrations of 2 to 16 *N*. Haase, Dücker, and Küppers [*115*] made isopiestic measurements from 2 to 28 *m*.

TABLE 8. Osmotic and mean activity coefficients of HNO_3

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	1.400	1.009	.749	12.000	1.488	1.884
.002	.984	.952	1.600	1.023	.761	13.000	1.508	1.999
.005	.977	.929	1.800	1.036	.774	14.000	1.523	2.109
.010	.969	.905	2.000	1.050	.788	15.000	1.535	2.212
.020	.961	.875	2.500	1.083	.826	16.000	1.543	2.309
.050	.948	.829	3.000	1.114	.868	17.000	1.547	2.397
.100	.941	.792	3.500	1.145	.913	18.000	1.548	2.476
.200	.937	.756	4.000	1.175	.961	19.000	1.546	2.546
.300	.940	.739	4.500	1.203	1.011	20.000	1.542	2.607
.400	.944	.730	5.000	1.230	1.063	21.000	1.535	2.658
.500	.949	.725	5.500	1.256	1.116	22.000	1.526	2.700
.600	.955	.723	6.000	1.281	1.171	23.000	1.516	2.734
.700	.961	.723	7.000	1.327	1.285	24.000	1.504	2.761
.800	.968	.724	8.000	1.368	1.403	25.000	1.491	2.781

TABLE 8. Osmotic and mean activity coefficients of HNO_3—Continued

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.900	.975	.727	9.000	1.405	1.523	26.000	1.477	2.796
1.000	.981	.730	10.000	1.437	1.644	27.000	1.463	2.807
1.200	.995	.738	11.000	1.465	1.765	28.000	1.450	2.816

$B^* = 1.5824$ $\beta = 6.2432 \times 10^{-2}$ $C = -1.3137 \times 10^{-3}$ $D = -1.2866 \times 10^{-5}$
$E = 4.9168 \times 10^{-7}$

$\sigma(\phi) = 9.20 \times 10^{-3}$ $\sigma(\gamma) = 1.96 \times 10^{-2}$ $VAR(\phi) = 8.458 \times 10^{-5}$ $VAR(\gamma) = 3.834 \times 10^{-4}$
$\sigma^*(\phi) = 6.48 \times 10^{-3}$ $\sigma^*(\gamma) = 1.25 \times 10^{-2}$

LiCl

The emfs of cells without transference and with lithium-amalgam and silver-silver chloride electrodes have been measured by Pearce and Mortimer (0.005–1 *m*) [119], MacInnes and Beattie (0.01–3 *m*) [120], Harned (0.05–4 *m*) [29], Lengyel and Giber (1–14.7 *m*) [121], and Lengyel, Giber, and Tamas (0.695–17.73 *m*) [122]. Caramazza [123] and Wilke and Schränkler [84] employed concentration cells; the former used silver-silver chloride electrodes and concentrations from 0.05 to 5 *m* and the latter lithium-amalgam electrodes and concentrations from 1 to 13 *N*. Lebed and Aleksandro [124] used lithium-amalgam and silver-silver chloride electrodes in cells without liquid junction over the concentration range of 0.005 to 0.1 *m*. Harned [9] used diffusion measurements to study dilute solutions, 0.0005 to 0.005 *N*. The vapor pressures of LiCl solutions were measured by Pearce and Nelson (0.1 sat) [125] and Kangro and Groeneveld (0.1–18 *m*) [126] while isopiestic measurements were made by Robinson and Sinclair (0.1–3 *m*) [30] and Robinson (0.1–20 *m*) [127]. A nomogram on the vapor pressure of LiCl based on data of the Lithium Corporation of America, Inc. was prepared by Davis [128].

TABLE 9. Osmotic and mean activity coefficients of LiCl

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	1.000	1.020	.775	8.000	2.159	5.19
.002	.984	.952	1.200	1.044	.798	9.000	2.328	7.11
.005	.976	.928	1.400	1.068	.825	10.000	2.480	9.60
.010	.969	.904	1.600	1.093	.855	11.000	2.612	12.7
.020	.960	.874	1.800	1.119	.888	12.000	2.723	16.4
.050	.947	.827	2.000	1.145	.924	13.000	2.814	20.7
.100	.940	.789	2.500	1.213	1.029	14.000	2.887	25.5
.200	.940	.756	3.000	1.284	1.157	15.000	2.947	30.9
.300	.946	.743	3.500	1.361	1.312	16.000	2.995	36.9
.400	.954	.739	4.000	1.441	1.499	17.000	3.036	43.4
.500	.964	.739	4.500	1.525	1.726	18.000	3.066	50.3
.600	.974	.742	5.000	1.613	2.00	19.000	3.079	57.0
.700	.985	.748	5.500	1.703	2.33	19.219 [a]	3.079	58.4
.800	.996	.756	6.000	1.794	2.73			
.900	1.008	.765	7.000	1.979	3.76			

$B^* = 1.305$ $\beta = 1.1603 \times 10^{-1}$ $C = -7.7726 \times 10^{-3}$ $D = 2.9279 \times 10^{-3}$
$E = -3.1953 \times 10^{-4}$ $F = 1.4068 \times 10^{-5}$ $G = -2.2498 \times 10^{-7}$

$\sigma(\phi) = 1.07 \times 10^{-2}$ $\sigma(\gamma) = 2.83 \times 10^{-1}$ $VAR(\phi) = 1.316 \times 10^{-4}$ $VAR(\gamma) = 7.995 \times 10^{-2}$
$\sigma^*(\phi) = 4.31 \times 10^{-3}$ $\sigma^*(\gamma) = 1.04 \times 10^{-2}$

a – saturated solution.

LiBr

The emfs of cells without transference and with lithium-amalgam and silver-silver bromide electrodes have been measured by Harned (0.1–4 *m*) [29] while Wilke and Schränkler [84] measured cells with transference for 1 to 8 *N*. Isopiestic measurements have been made by Robinson (0.16–3.3 *m*) [*23*], Robinson and McCoach (4.8–20 *m*) [*129*], and Sarnowski and Baranowski (1.5–6 *N*) [*130*].

TABLE 10. Osmotic and mean activity coefficients of LiBr

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	1.000	1.035	.803	8.000	2.437	8.651
.002	.984	.952	1.200	1.064	.834	9.000	2.670	13.10
.005	.977	.929	1.400	1.095	.871	10.000	2.905	20.01
.010	.970	.905	1.600	1.127	.913	11.000	3.138	30.63
.020	.961	.877	1.800	1.160	.960	12.000	3.364	46.66
.050	.950	.832	2.000	1.195	1.012	13.000	3.574	70.16
.100	.944	.797	2.500	1.284	1.166	14.000	3.761	103.1
.200	.944	.767	3.000	1.377	1.359	15.000	3.92	147.
.300	.951	.755	3.500	1.473	1.598	16.000	4.04	201.
.400	.960	.752	4.000	1.572	1.891	17.000	4.13	264.
.500	.970	.754	4.500	1.674	2.252	18.000	4.18	332.
.600	.982	.759	5.000	1.777	2.696	19.000	4.20	404.
.700	.994	.767	5.500	1.883	3.242	20.000	4.22	486.
.800	1.007	.778	6.000	1.990	3.915			
.900	1.021	.789	7.000	2.210	5.778			

$B^* = 1.60$ $\beta = 8.5164 \times 10^{-2}$ $C = 1.8335 \times 10^{-2}$ $D = -2.5742 \times 10^{-3}$
$E = 2.4140 \times 10^{-4}$ $F = -1.1612 \times 10^{-5}$ $G = 2.0402 \times 10^{-7}$

$\sigma(\phi) = 7.49 \times 10^{-3}$ $\sigma(\gamma) = 1.43$ $\mathrm{VAR}(\phi) = 5.610 \times 10^{-5}$ $\mathrm{VAR}(\gamma) = 2.036$
$\sigma^*(\phi) = 3.31 \times 10^{-3}$ $\sigma^*(\gamma) = 8.21 \times 10^{-3}$

LiI

The emfs of cells with transference and lithium-amalgam electrodes were measured from 1 to 5 *N* by Wilke and Schränkler [84]. Isopiestic measurements have been made by Robinson and Sinclair (0.1–3 *m*) [*30*] and Sarnowski and Baranowski (1.5–6 *N*) [130].

TABLE 11. Osmotic and mean activity coefficients of LiI

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.966	.300	.978	.803	1.200	1.112	.957
.002	.985	.953	.400	.993	.811	1.400	1.143	1.007
.005	.978	.930	.500	1.008	.824	1.600	1.176	1.063
.010	.971	.908	.600	1.023	.838	1.800	1.211	1.126
.020	.964	.882	.700	1.037	.855	2.000	1.249	1.197
.050	.956	.843	.800	1.052	.872	2.500	1.353	1.420
.100	.956	.817	.900	1.067	.892	3.000	1.471	1.722
.200	.965	.802	1.000	1.082	.912			

$B^* = 1.695$ $\beta = 1.7810 \times 10^{-1}$ $C = -4.4794 \times 10^{-2}$ $D = 1.8282 \times 10^{-2}$
$E = -2.0254 \times 10^{-3}$

$\sigma(\phi) = 1.98 \times 10^{-3}$ $\sigma(\gamma) = 2.14 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 3.933 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 4.591 \times 10^{-6}$
$\sigma^*(\phi) = 1.83 \times 10^{-3}$ $\sigma^*(\gamma) = 1.97 \times 10^{-3}$

LiOH

Harned and Swindells [131] measured the emfs of cells without transference and with hydrogen and lithium-amalgam electrodes from 0.0505 to 3.926 *m*. Kangro and Groeneveld [126] measured the vapor pressure for concentrations from 0.5 to 5.0 *m*.

TABLE 12. Osmotic and mean activity coefficients of LiOH

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.400	.864	.600	1.600	.866	.494
.002	.983	.950	.500	.859	.579	1.800	.870	.488
.005	.974	.923	.600	.857	.563	2.000	.875	.484
.010	.964	.895	.700	.855	.550	2.500	.884	.476
.020	.951	.858	.800	.855	.539	3.000	.891	.469
.050	.929	.794	.900	.855	.530	3.500	.895	.463
.100	.907	.735	1.000	.856	.522	4.000	.895	.457
.200	.884	.668	1.200	.858	.510	5.000	.933[a]	.493[a]
.300	.872	.628	1.400	.862	.501			

$B^* = 4.5000 \times 10^{-1}$ $\beta = 7.6900 \times 10^{-2}$ $C = -6.8870 \times 10^{-3}$

$\sigma(\phi) = 5.67 \times 10^{-3}$ $\sigma(\gamma) = 4.13 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 3.213 \times 10^{-5}$ $\mathrm{VAR}(\gamma) = 1.702 \times 10^{-5}$

$\sigma^*(\phi) = 6.47 \times 10^{-3}$ $\sigma^*(\gamma) = 6.26 \times 10^{-3}$

a – Values of Kangro and Groeneveld [126].

$LiClO_4$

Harned and Shropshire [132] used diffusion measurements to obtain activity coefficients for concentrations of 0.0005 to 0.020 *N*. Jones [133] made isopiestic measurements from 0.2 to 4.5 *m*.

TABLE 13. Osmotic and mean activity coefficients of $LiClO_4$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.966	.400	.983	.797	1.600	1.171	1.035
.002	.985	.953	.500	.997	.806	1.800	1.206	1.095
.005	.978	.931	.600	1.011	.818	2.000	1.240	1.161
.010	.971	.908	.700	1.026	.833	2.500	1.329	1.351
.020	.964	.882	.800	1.041	.849	3.000	1.418	1.581
.050	.956	.843	.900	1.057	.868	3.500	1.508	1.857
.100	.954	.815	1.000	1.072	.887	4.000	1.596	2.184
.200	.960	.795	1.200	1.105	.931	4.500	1.683	2.567
.300	.971	.792	1.400	1.138	.981			

$B^* = 1.925$ $\beta = 1.1420 \times 10^{-1}$ $C = 9.4314 \times 10^{-3}$ $D = -9.0320 \times 10^{-4}$

$\sigma(\phi) = 1.53 \times 10^{-3}$ $\sigma(\gamma) = 2.46 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 2.355 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 6.059 \times 10^{-6}$

$\sigma^*(\phi) = 1.33 \times 10^{-3}$ $\sigma^*(\gamma) = 1.88 \times 10^{-3}$

$LiNO_3$

Vapor pressures were measured by Pearce and Nelson (0.00–12.8693 *m*) [125] and Kangro and Groeneveld (0.5–20.0 *m*) [126]. Isopiestic measurements were made by Robinson from 0.1 to 3.5 *m* [134] and from 0.1 to 13.5 *m* [135]. Harned and Shropshire [132] used diffusion measurements for studies from 0.0005 to 0.020 *N*.

TABLE 14. Osmotic and mean activity coefficients of $LiNO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	1.000	.997	.743	8.000	1.55	1.96
.002	.984	.952	1.200	1.015	.758	9.000	1.61	2.22
.005	.976	.928	1.400	1.034	.775	10.000	1.66	2.50
.010	.969	.904	1.600	1.052	.794	11.000	1.70	2.79
.020	.960	.874	1.800	1.071	.815	12.000	1.74	3.08
.050	.947	.827	2.000	1.089	.837	13.000	1.77	3.38
.100	.939	.788	2.500	1.135	.898	14.000	1.80	3.68
.200	.936	.753	3.000	1.179	.966	15.000	1.82	3.96
.300	.940	.736	3.500	1.222	1.040	16.000	1.83	4.22
.400	.946	.729	4.000	1.264	1.120	17.000	1.83	4.46
.500	.953	.726	4.500	1.304	1.206	18.000	1.83	4.67
.600	.961	.726	5.000	1.344	1.298	19.000	1.82	4.84
.700	.970	.728	5.500	1.382	1.395	20.000	1.81	4.97
.800	.978	.732	6.000	1.418	1.49			
.900	.987	.737	7.000	1.49	1.72			

$B^* = 1.40$ $\beta = 8.5378 \times 10^{-2}$ $C = -1.3859 \times 10^{-3}$ $D = -2.1636 \times 10^{-5}$

$E = 1.9151 \times 10^{-7}$

$\sigma(\phi) = 1.72 \times 10^{-2}$ $\sigma(\gamma) = 5.96 \times 10^{-2}$ $VAR(\phi) = 1.011 \times 10^{-3}$ $VAR(\gamma) = 3.547 \times 10^{-3}$

$\sigma^*(\phi) = 1.01 \times 10^{-2}$ $\sigma^*(\gamma) = 1.75 \times 10^{-2}$

NaF

Ivett and De Vries [136] calculated activity coefficients from 0.05 to 0.9 *m* from emfs of the cell:

$$Na\text{-}Hg|NaF(m)|PbF_2|Pb\text{-}Hg.$$

Robinson [137] made isopiestic measurements over the concentration range 0.1–1 *m*. Ivett and De Vries gave 0.983 *m* for the saturated solution.

TABLE 15. Osmotic and mean activity coefficients of NaF

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.100	.924	.764	.700	.880	.604
.002	.984	.951	.200	.908	.710	.800	.877	.592
.005	.976	.926	.300	.899	.676	.900	.874	.582
.010	.967	.901	.400	.892	.652	.983 [a]	.871	.574
.020	.956	.868	.500	.887	.633	1.000	.871 [b]	.573 [b]
.050	.939	.813	.600	.883	.617			

$B^* = 1.28$ $\beta = -1.8000 \times 10^{-2}$

$\sigma(\phi) = 7.90 \times 10^{-4}$ $\sigma(\gamma) = 6.60 \times 10^{-4}$ $VAR(\phi) = 6.260 \times 10^{-7}$ $VAR(\gamma) = 4.300 \times 10^{-7}$

$\sigma^*(\phi) = 8.90 \times 10^{-4}$ $\sigma^*(\gamma) = 1.01 \times 10^{-3}$

a – saturated solution.
b – extrapolated values for supersaturated solution.

NaCl

Measurements of cells without transference and with sodium-amalgam and calomel electrodes were made by Allmand and Polack (1 *m*-sat) [138] and with sodium-amalgam and silver-silver chloride electrodes by Harned (0.05–4 *m*) [29], Harned and Nims (0.05–4 *m*) [139], Brown and MacInnes (0.005–0.1 *m*) [140], Sakong and Huang [141], and Lebed and Aleksandrov (0.0001–0.1 *m*) [124]. Measurements of cells with transference made with silver-silver chlorides were made by Wilke and Schränkler (1–5 *N*) [84] and Janz and Gordon (0.001–0.1 *m*) *[13]*. Caramazza *[142]* made measurements on concentration cells using silver-silver chloride electrodes for 0.001 to 5 *m*. Calomel electrodes were used in cells with liquid junction by Getman (0.001–3 *m*) [143] and Haas and Jellinek (0.005–4 *m*) [144] (Haas and Jellinek also used silver-silver chloride electrodes). Stokes [145] used a bithermal method over the concentration range of 4 to 6 *m*. Harned *[9]* used the diffusion method for dilute solutions, 0.0005 to 0.001 *N*. Vapor-pressure lowerings were measured by Negus (1–5 *m*) [146], Pearce and Nelson (0.1 *m*-sat) [125], Olynyk and Gordon (2–6 *m*) [147], and Petit (0.5–6 *m*) *[148]*. Hepburn [149] used the dew-point method to obtain a value for the saturated solution. Isopiestic measurements were made by Robinson and Sinclair (0.1–4 *m*) *[30]*, Scatchard, Hamer, and Wood (0.1 *m*-sat) *[14]*, Robinson (0.1–4 *m*) *[150]*, Janis and Ferguson (0.1–4 *m*) *[151]*, and Robinson (0.1–6 *m*) *[36]*.

TABLE 16. Osmotic and mean activity coefficients of NaCl

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.500	.921	.681	2.000	.984	.668
.002	.984	.952	.600	.923	.673	2.500	1.013	.688
.005	.976	.928	.700	.926	.667	3.000	1.045	.714
.010	.968	.903	.800	.929	.662	3.500	1.080	.746
.020	.959	.872	.900	.932	.659	4.000	1.116	.783
.050	.944	.822	1.000	.936	.657	4.500	1.153	.826
.100	.933	.779	1.200	.944	.655	5.000	1.191	.874
.200	.924	.734	1.400	.953	.656	5.500	1.231	.928
.300	.921	.709	1.600	.962	.658	6.000	1.270	.986
.400	.920	.693	1.800	.973	.662	6.144^{a}	1.281	1.004

$B^{*} = 1.4495$ $\beta = 2.0442 \times 10^{-2}$ $C = 5.7927 \times 10^{-3}$ $D = -2.8860 \times 10^{-4}$

$\sigma(\phi) = 6.40 \times 10^{-4}$ $\sigma(\gamma) = 4.90 \times 10^{-4}$ $\mathrm{VAR}(\phi) = 4.120 \times 10^{-7}$ $\mathrm{VAR}(\gamma) = 2.430 \times 10^{-7}$

$\sigma^{*}(\phi) = 6.60 \times 10^{-4}$ $\sigma^{*}(\gamma) = 6.80 \times 10^{-4}$

a – saturated solution.

NaBr

Measurements of cells without transference and with sodium-amalgam and silver-silver bromide electrodes have been made by Harned and Douglas (0.03–3 *m*) [152], Harned (0.05–4 *m*) [29], Murata (0.01–8.66 *m*) [153], Harned and Crawford (0.2–4 *m*) *[154]*, and Lebed and Aleksandrov (0.005–2 *m*) *[124]*. Measurements of cells with transference and with hydrogen electrodes were made by Wilke and Schränkler (1–6 *N*) [84]. Haas and Jellinek [144] used calomel and silver-silver bromide electrodes in cells with liquid junctions over the range of 0.01 to 4 *m*. Vapor-pressure measurements were made by Pearce, Taylor, and Bartlett (0.1–9 *m*) [155] and Ionin and Chezzanova (4–9.1 *m*) [156]. Isopiestic measurements have been made by Robinson (0.1–4 *m*) *[23, 150]*, Pencier and Marcus (2.8–8.9 *m*) *[157]*, and Makarov, Vlasov, and Azarko (3 9.1 *m*) *[158]*.

TABLE 17. Osmotic and mean activity coefficients of NaBr

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.600	.936	.692	3.000	1.112	.816
.002	.984	.952	.700	.941	.688	3.500	1.157	.871
.005	.976	.928	.800	.946	.687	4.000	1.203	.934
.010	.969	.903	.900	.951	.686	4.500	1.249	1.005
.020	.959	.873	1.000	.957	.687	5.000	1.296	1.083
.050	.945	.824	1.200	.969	.691	5.500	1.341	1.169
.100	.935	.783	1.400	.983	.697	6.000	1.386	1.261

TABLE 17. Osmotic and mean activity coefficients of NaBr—Continued

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.200	.929	.742	1.600	.997	.706	7.000	1.470	1.465
.300	.928	.720	1.800	1.012	.717	8.000	1.547	1.694
.400	.929	.706	2.000	1.027	.730	9.000	1.619	1.949
.500	.932	.697	2.500	1.068	.768			

$B^* = 1.49$ $\beta = 3.5230 \times 10^{-2}$ $C = 7.1023 \times 10^{-3}$ $D = -3.2348 \times 10^{-4}$

$E = -3.6943 \times 10^{-5}$ $F = 2.6380 \times 10^{-6}$

$\sigma(\phi) = 4.17 \times 10^{-3}$ $\sigma(\gamma) = 6.19 \times 10^{-3}$ $\text{VAR}(\phi) = 1.741 \times 10^{-5}$ $\text{VAR}(\gamma) = 3.826 \times 10^{-5}$

$\sigma^*(\phi) = 3.03 \times 10^{-3}$ $\sigma^*(\gamma) = 4.10 \times 10^{-3}$

NaI

Measurements of cells without transference and with sodium-amalgam and silver-silver iodide electrodes were made by Harned and Douglas (0.03–3 *m*) [152] and Lebed and Aleksandrov (0.005–0.2 *m*) [124]. Wilke and Schränkler [84] measured the emf of cells with transference and with sodium-amalgam electrodes from 1 to 6 *N*. Haas and Jellinek [144] used calomel and silver-silver iodide electrodes in cells with liquid junction over the concentration range, 0.01–4 *m*. Isopiestic measurements were made by Robinson (0.1–3.7 *m*) [*23*], Miller and Sheridan (4–12 *m*) [*159*], and Makarov, Vlasov, and Azarko (3 *m*–sat) [158].

TABLE 18. Osmotic and mean activity coefficients of NaI

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.700	.964	.722	4.000	1.298	1.153
.002	.984	.952	.800	.972	.725	4.500	1.356	1.269
.005	.977	.928	.900	.981	.729	5.000	1.415	1.402
.010	.969	.904	1.000	.989	.734	5.500	1.474	1.552
.020	.960	.874	1.200	1.007	.747	6.000	1.535	1.723
.050	.947	.827	1.400	1.025	.763	7.000	1.656	2.131
.100	.939	.789	1.600	1.044	.781	8.000	1.775	2.642
.200	.936	.753	1.800	1.063	.801	9.000	1.890	3.267
.300	.938	.735	2.000	1.083	.823	10.000	1.996	4.011
.400	.943	.726	2.500	1.134	.887	11.000	2.088	4.861
.500	.950	.722	3.000	1.187	.963	12.000	2.164	5.782
.600	.957	.721	3.500	1.242	1.051			

$B^* = 1.486$ $\beta = 6.8170 \times 10^{-2}$ $C = 3.0500 \times 10^{-3}$ $D = 0.0000$

$E = -1.0000 \times 10^{-5}$

$\sigma(\phi) = 7.20 \times 10^{-3}$ $\sigma(\gamma) = 3.03 \times 10^{-2}$ $\text{VAR}(\phi) = 5.150 \times 10^{-5}$ $\text{VAR}(\gamma) = 9.170 \times 10^{-4}$

$\sigma^*(\phi) = 3.90 \times 10^{-3}$ $\sigma^*(\gamma) = 1.56 \times 10^{-2}$

NaOH

Measurements of cells without transference and made with sodium-amalgam and hydrogen electrodes were made by Harned (0.0202–3.10 *m*) [160, 161], Ferguson and Schlucter (0.01004–2.825 *m*) [162], and Harned and Hecker (0.05–4.0) [*163*]. Kobayski and Wang [*164*] made measurements on the cell:

$$\text{Hg}|\text{HgO}|\text{NaOH}(m)|\text{H}_2,\ \text{Pt},$$

from 0.1 to 0.9 *m*. Vapor pressures were measured by Stokes (5.085–13.834 *m*) [*145*] and Kangro and Groeneveld (1.0–27.0) [126] and isopiestic measurements by Stokes (2–29 *m*) [*165*].

TABLE 19. Osmotic and mean activity coefficients of NaOH

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	1.600	.994	.692	14.000	2.489	8.083
.002	.984	.952	1.800	1.008	.702	15.000	2.575	9.796
.005	.976	.927	2.000	1.023	.714	16.000	2.645	11.65
.010	.968	.902	2.500	1.062	.749	17.000	2.697	13.60
.020	.958	.870	3.000	1.105	.794	18.000	2.735	15.57
.050	.943	.819	3.500	1.151	.847	19.000	2.758	17.52
.100	.931	.775	4.000	1.200	.911	20.000	2.771	19.41
.200	.924	.731	4.500	1.253	.987	21.000	2.774	21.25
.300	.923	.708	5.000	1.310	1.076	22.000	2.772	23.03
.400	.925	.694	5.500	1.370	1.181	23.000	2.767	24.78
.500	.928	.685	6.000	1.434	1.302	24.000	2.760	26.52
.600	.932	.679	7.000	1.568	1.609	25.000	2.752	28.26
.700	.937	.676	8.000	1.710	2.018	26.000	2.741	29.96
.800	.943	.674	9.000	1.855	2.559	27.000	2.725	31.48
.900	.948	.673	10.000	1.999	3.258	28.000	2.697	32.57
1.000	.954	.674	11.000	2.138	4.145	29.000	2.647	32.86
1.200	.967	.678	12.000	2.268	5.241			
1.400	.980	.684	13.000	2.386	6.556			

B* = 1.21 $\beta = 6.1509 \times 10^{-2}$ $C = -2.9784 \times 10^{-3}$ $D = 1.3293 \times 10^{-3}$

$E = -1.1134 \times 10^{-4}$ $F = 3.5452 \times 10^{-6}$ $G = -3.9985 \times 10^{-8}$

$\sigma(\phi) = 5.53 \times 10^{-3}$ $\sigma(\gamma) = 7.34 \times 10^{-2}$ $\mathrm{VAR}(\phi) = 3.060 \times 10^{-5}$ $\mathrm{VAR}(\gamma) = 5.380 \times 10^{-3}$

$\sigma^*(\phi) = 4.66 \times 10^{-3}$ $\sigma^*(\gamma) = 5.06 \times 10^{-3}$

$NaClO_3$

Isopiestic measurements were made by Jones [*166*] from 0.2 to 3.0 *m*.

TABLE 20. Osmotic and mean activity coefficients of $NaClO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.300	.905	.687	1.200	.878	.575
.002	.984	.952	.400	.899	.664	1.400	.876	.563
.005	.976	.927	.500	.895	.646	1.600	.875	.553
.010	.968	.902	.600	.891	.631	1.800	.874	.545
.020	.958	.870	.700	.888	.619	2.000	.873	.537
.050	.941	.818	.800	.885	.608	2.500	.875	.523
.100	.928	.771	.900	.883	.598	3.000	.880	.514
.200	.913	.719	1.000	.881	.590			

B* = 1.45 $\beta = -2.4867 \times 10^{-2}$ $C = 4.1573 \times 10^{-3}$

$\sigma(\phi) = 1.81 \times 10^{-3}$ $\sigma(\gamma) = 1.05 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 3.240 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 1.000 \times 10^{-6}$

$\sigma^*(\phi) = 2.00 \times 10^{-3}$ $\sigma^*(\gamma) = 1.71 \times 10^{-3}$

$NaClO_4$

Isopiestic measurements were made by Jones (0.2–6.5) [*133*], by Miller and Sheridan (4–16 *m*) [159], and Rush and Johnson (m=6–18) [167].

TABLE 21. Osmotic and mean activity coefficients of $NaClO_4$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.500	.911	.668	2.000	.933	.608
.002	.984	.952	.600	.910	.657	2.500	.947	.608
.005	.976	.928	.700	.910	.648	3.000	.961	.612
.010	.968	.903	.800	.911	.641	3.500	.976	.618
.020	.959	.872	.900	.912	.635	4.000	.991	.626
.050	.943	.821	1.000	.913	.630	4.500	1.007	.636
.100	.931	.777	1.200	.916	.622	5.000	1.023	.648
.200	.920	.729	1.400	.920	.616	5.500	1.042	.662
.300	.915	.702	1.600	.924	.612	6.000[a]	1.062	.678
.400	.912	.683	1.800	.929	.610			

$B^* = 1.50$ $\beta = -3.0000 \times 10^{-3}$ $C = 7.4800 \times 10^{-3}$ $D = -1.2000 \times 10^{-3}$
$E = 8.2600 \times 10^{-5}$

$\sigma(\phi) = 8.40 \times 10^{-4}$ $\sigma(\gamma) = 7.50 \times 10^{-4}$ $VAR(\phi) = 7.050 \times 10^{-7}$ $VAR(\gamma) = 5.550 \times 10^{-7}$
$\sigma^*(\phi) = 8.60 \times 10^{-4}$ $\sigma^*(\gamma) = 1.12 \times 10^{-3}$

a – for molalities above 6.0, the data of references [159] and [167] differ fairly widely; average values are as follows:

m	ϕ	γ	m	ϕ	γ
6.0	1.0570±.0009	0.680±.004	12.0	1.2007±.0013	0.876±.014
7.0	1.0859±.0029	0.710±.003	14.0	1.2339±.0121	0.940±.030
8.0	1.1122±.0052	0.740±.004	16.0	1.2626±.0234	1.008±.053
9.0	1.1385±.0045	0.773±.002	18.0	1.2626*	1.007*
10.0	1.1607±.0046	0.808±.003			

* Ref [167] only.

$NaBrO_3$

Jones and Froning [*168*] made isopiestic measurements from 0.2 to 2.617 *m* (saturated).

TABLE 22. Osmotic and mean activity coefficients of $NaBrO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.300	.884	.658	1.200	.825	.507
.002	.984	.951	.400	.873	.628	1.400	.817	.490
.005	.976	.926	.500	.864	.605	1.600	.810	.474
.010	.967	.900	.600	.856	.585	1.800	.803	.461
.020	.956	.867	.700	.849	.568	2.000	.798	.449
.050	.937	.811	.800	.844	.553	2.500	.793	.426
.100	.920	.759	.900	.838	.540	2.617[a]	.792[b]	.421[b]
.200	.899	.698	1.000	.833	.528			

$B^* = 1.475$ $\beta = -1.0264 \times 10^{-1}$ $C = 4.6481 \times 10^{-2}$ $D = -1.7654 \times 10^{-2}$
$E = 2.9209 \times 10^{-3}$

$\sigma(\phi) = 8.34 \times 10^{-3}$ $\sigma(\gamma) = 7.67 \times 10^{-4}$ $VAR(\phi) = 6.954 \times 10^{-5}$ $VAR(\gamma) = 5.887 \times 10^{-7}$
$\sigma^*(\phi) = 1.01 \times 10^{-2}$ $\sigma^*(\gamma) = 1.34 \times 10^{-3}$

a – saturated solution.
b – extrapolated values.

$NaNO_3$

Vapor pressures were measured by Pearce and Hopson (0.1–10.830 *m*) [169] and Kangro and Groeneveld (0.1–10.0 *m*) [*126*]; 10.830 *m* = saturation. Isopiestic measurements were made by Robinson [*134*] from 0.1 to 6.0 *m*. Harned and Shropshire used diffusion measurements to obtain activity coefficients for the concentration ranges, 0.005–0.020 *N* [*170*] and 0.003–0.015 *N* [*132*].

TABLE 23. Osmotic and mean activity coefficients of $NaNO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.700	.864	.585	4.000	.798	.408
.002	.984	.951	.800	.860	.571	4.500	.793	.396
.005	.975	.926	.900	.856	.559	5.000	.789	.386
.010	.967	.900	1.000	.852	.549	5.500	.787	.378
.020	.956	.867	1.200	.845	.530	6.000	.789	.372
.050	.938	.811	1.400	.839	.515	7.000	.808	.367
.100	.921	.760	1.600	.834	.501	8.000	.860	.378
.200	.903	.702	1.800	.830	.489	9.000	.965	.415
.300	.891	.666	2.000	.826	.478	10.000	1.146	.500
.400	.883	.639	2.500	.817	.456	10.830[a]	1.374	.641
.500	.876	.618	3.000	.810	.437			
.600	.870	.600	3.500	.803	.422			

$B^* = 1.30 \quad \beta = -4.6500 \times 10^{-2} \quad C = 9.4000 \times 10^{-3} \quad D = -1.5100 \times 10^{-3}$

$E = 1.0500 \times 10^{-4}$

$\sigma(\phi) = 1.56 \times 10^{-3} \quad \sigma(\gamma) = 9.40 \times 10^{-4} \quad VAR(\phi) = 2.435 \times 10^{-6} \quad VAR(\gamma) = 8.817 \times 10^{-7}$

$\sigma^*(\phi) = 1.62 \times 10^{-3} \quad \sigma^*(\gamma) = 1.86 \times 10^{-3}$

a – saturated solution.

NaH_2PO_4

Isopiestic measurements were made by Stokes (0.1–6.5 *m*) [*171*] and Scatchard and Breckenridge (0.1–1.3 *m*) [*172*].

TABLE 24. Osmotic and mean activity coefficients of NaH_2PO_4

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.500	.832	.564	2.000	.721	.371
.002	.984	.951	.600	.819	.539	2.500	.707	.343
.005	.975	.925	.700	.808	.518	3.000	.699	.322
.010	.966	.898	.800	.797	.499	3.500	.695	.306
.020	.954	.863	.900	.787	.482	4.000	.694	.294
.050	.933	.803	1.000	.778	.467	4.500	.695	.284
.100	.911	.745	1.200	.763	.441	5.000	.698	.275
.200	.884	.675	1.400	.750	.419	5.500	.703	.269
.300	.863	.628	1.600	.738	.401	6.000	.712	.265
.400	.847	.593	1.800	.729	.385	6.500	.726	.262

$B^* = 1.275 \quad \beta = -1.3156 \times 10^{-1} \quad C = 2.8874 \times 10^{-2} \quad D = -3.5926 \times 10^{-3}$

$E = 1.9455 \times 10^{-4}$

$\sigma(\phi) = 1.79 \times 10^{-3} \quad \sigma(\gamma) = 7.98 \times 10^{-4} \quad VAR(\phi) = 1.183 \times 10^{-7} \quad VAR(\gamma) = 2.372 \times 10^{-8}$

$\sigma^*(\phi) = 2.08 \times 10^{-3} \quad \sigma^*(\gamma) = 2.27 \times 10^{-3}$

NaH_2AsO_4

Isopiestic measurements were made by Scatchard and Breckenridge (0.1–1.3 *m*) [*172*].

TABLE 25. Osmotic and mean activity coefficients of NaH_2AsO_4

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.200	.903	.707	.900	.824	.536
.002	.984	.952	.300	.887	.668	1.000	.816	.522
.005	.976	.927	.400	.874	.636	1.100	.808	.509
.010	.968	.902	.500	.862	.611	1.200	.802	.497
.020	.958	.871	.600	.851	.589	1.300	.796	.486
.050	.941	.817	.700	.841	.569			
.100	.924	.767	.800	.832	.552			

$B^* = 1.75 \quad \beta = -1.1424 \times 10^{-1} \quad C = 1.7543 \times 10^{-2}$

$\sigma(\phi) = 6.60 \times 10^{-4} \quad \sigma(\gamma) = 5.20 \times 10^{-4} \quad VAR(\phi) = 4.413 \times 10^{-7} \quad VAR(\gamma) = 2.683 \times 10^{-7}$

$\sigma^*(\phi) = 7.90 \times 10^{-4} \quad \sigma^*(\gamma) = 8.30 \times 10^{-4}$

NaCNS

Isopiestic measurements were made by Robinson (0.1–4.0 *m*) [*222*] and Miller and Sheridan (1.0–18.0) [*159*].

TABLE 26. Osmotic and mean activity coefficients of NaCNS

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.900	.962	.707	6.000	1.338	1.213
.002	.984	.952	1.000	.968	.709	7.000	1.416	1.390
.005	.977	.928	1.200	.980	.714	8.000	1.491	1.591
.010	.969	.904	1.400	.993	.721	9.000	1.561	1.815
.020	.960	.875	1.600	1.006	.731	10.000	1.624	2.06
.050	.948	.828	1.800	1.019	.742	11.000	1.678	2.31
.100	.939	.789	2.000	1.033	.754	12.000	1.722	2.57
.200	.934	.751	2.500	1.068	.790	13.000	1.752	2.81
.300	.934	.732	3.000	1.105	.832	14.000	1.768	3.02
.400	.937	.720	3.500	1.142	.881	15.000	1.767	3.18
.500	.941	.713	4.000	1.181	.935	16.000	1.747	3.27
.600	.946	.709	4.500	1.220	.996	17.000	1.705	3.28
.700	.951	.707	5.000	1.259	1.062	18.000	1.639	3.19
.800	.956	.707	5.500	1.299	1.135			

$B^* = 1.60 \quad \beta = 4.4018 \times 10^{-2} \quad C = 2.9953 \times 10^{-3} \quad D = -1.4925 \times 10^{-4}$

$E = -1.0518 \times 10^{-6}$

$\sigma(\phi) = 5.57 \times 10^{-2} \quad \sigma(\gamma) = 1.41 \times 10^{-1} \quad VAR(\phi) = 3.105 \times 10^{-3} \quad VAR(\gamma) = 2.001 \times 10^{-2}$

$\sigma^*(\phi) = 3.36 \times 10^{-2} \quad \sigma^*(\gamma) = 5.30 \times 10^{-2}$

KF

Isopiestic measurements were made by Robinson (0.1–4.0 *m*) [*137*] and Tamas and Kosza (2–17.5 *m*) [*174*].

TABLE 27. Osmotic and mean activity coefficients of KF

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.900	.928	.647	6.000	1.283	.989
.002	.984	.952	1.000	.932	.645	7.000	1.367	1.131
.005	.976	.927	1.200	.941	.643	8.000	1.451	1.298
.010	.968	.902	1.400	.950	.644	9.000	1.533	1.493
.020	.958	.870	1.600	.961	.647	10.000	1.611	1.715
.050	.942	.818	1.800	.972	.651	11.000	1.684	1.962
.100	.930	.773	2.000	.984	.658	12.000	1.751	2.23
.200	.920	.726	2.500	1.015	.678	13.000	1.810	2.52
.300	.916	.700	3.000	1.048	.705	14.000	1.860	2.82
.400	.915	.682	3.500	1.084	.738	15.000	1.900	3.12
.500	.916	.670	4.000	1.122	.777	16.000	1.930	3.41
.600	.918	.662	4.500	1.161	.821	17.000	1.948	3.68
.700	.921	.655	5.000	1.201	.871	17.500	1.953	3.80
.800	.924	.650	5.500	1.242	.927			

$B^* = 1.29$ $\beta = 2.7845 \times 10^{-2}$ $C = 5.0000 \times 10^{-3}$ $D = -2.5309 \times 10^{-4}$

$E = 2.6790 \times 10^{-6}$

$\sigma(\phi) = 2.60 \times 10^{-3}$ $\sigma(\gamma) = 3.90 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 6.631 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 1.494 \times 10^{-5}$

$\sigma^*(\phi) = 1.99 \times 10^{-3}$ $\sigma^*(\gamma) = 3.44 \times 10^{-3}$

KCl

In addition to the results listed in table 1, the following measurements have been made. Harned [175] used calomel electrodes in cells with transference over the concentration range, 0.1–3 *N*, Getman [143] from 0.001 to 3 *N*, and Bird and Nixon [176] from 0.001 to 1 *m*. Cells without transference and with potassium-amalgam and silver-silver chloride electrodes were measured by Parton (0.3–1.82 *m*) [177], and Lebed and Aleksandrov (0.005–0.2 *m*) [124]. Calomel and silver-silver chloride electrodes were used in cells with liquid junction by Haas and Jellinek [144] for the concentration range, 0.005–4 *m*. Cells with transference and potassium-amalgam electrodes were used by Wilke and Schränkler [84] over the concentration range, 1–6 *N*. Vapor pressures were measured by Hill (0.067–2 *m*) [178], Pearce and Nelson (0.1–sat) [125], Weir (0.269–1.408 *m*) [179], Brown and Delaney (0.025–2.38 *m*) [180], and Petit (0.25–*m*) [148]. Isopiestic measurements have been made by Janis and Ferguson (0.1–4 *m*) [*151*], and Robinson (0.1–4.5) [*150*].

TABLE 28. Osmotic and mean activity coefficients of KCl [#]

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.500	.900	.649	2.000	.912	.573
.002	.984	.951	.600	.899	.636	2.500	.923	.568
.005	.976	.927	.700	.898	.626	3.000	.936	.568
.010	.967	.901	.800	.898	.617	3.500	.950	.571
.020	.957	.869	.900	.898	.610	4.000	.965	.576
.050	.940	.816	1.000	.898	.604	4.500	.981	.584
.100	.927	.768	1.200	.900	.594	4.803 [a]	.990	.589
.200	.913	.717	1.400	.902	.586	5.000 [b]	.997 [b]	.593 [b]
.300	.906	.687	1.600	.905	.580			
.400	.902	.665	1.800	.908	.576			

$B^* = 1.295$ $\beta = 7.0000 \times 10^{-5}$ $C = 3.5990 \times 10^{-3}$ $D = -1.9540 \times 10^{-4}$

$\sigma(\phi) = 7.10 \times 10^{-4}$ $\sigma(\gamma) = 7.10 \times 10^{-4}$ $\mathrm{VAR}(\phi) = 4.999 \times 10^{-7}$ $\mathrm{VAR}(\gamma) = 4.980 \times 10^{-7}$

$\sigma^*(\phi) = 7.69 \times 10^{-4}$ $\sigma^*(\gamma) = 1.04 \times 10^{-3}$

a – saturated solution.
b – calculated values for supersaturated solution.
\# – see also Table 1; values here are rounded values of those given in Column I′ of Table 1.

KBr

Measurements have been made of potassium-amalgam cells without transference and with silver-silver bromide electrodes by Pearce and Hart (0.001–2.8 *N*) [181], Harned and Douglas (0.03–3 *m*) [152], and Harned (0.05–4 *m*) [182], and with mercury-mercurous bromide electrodes by Murata (0.01–5.66 *m*) [*183*]. Cells with transference and silver-silver bromide electrodes were used by Wilke and Schränkler (1–4 *N*) [84] and MacWilliam and Gordon (0.01–0.1 *m*) [*184*]. Concentration cells with silver-silver bromide electrodes were used by Scatchard and Orthing [*185*] over the concentration range of 0.01 to 0.0625 *m*. Haas and Jellinek [144] used calomel and silver-silver bromide electrodes in cells with liquid junction over the concentration range, 0.01–4 *m*. Robinson made isopiestic measurements from 0.1 to 4 *m* [*23, 150*].

TABLE 29. Osmotic and mean activity coefficients of KBr

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.500	.906	.658	2.000	.927	.593
.002	.984	.952	.600	.905	.646	2.500	.940	.592
.005	.976	.927	.700	.905	.637	3.000	.954	.595
.010	.968	.902	.800	.905	.629	3.500	.969	.600
.020	.958	.870	.900	.906	.622	4.000	.985	.608
.050	.942	.817	1.000	.907	.617	4.500	1.000	.617
.100	.928	.771	1.200	.910	.608	5.000	1.015	.626
.200	.916	.722	1.400	.913	.602	5.500	1.029	.637
.300	.910	.693	1.600	.917	.598			
.400	.907	.673	1.800	.922	.595			

$B^* = 1.35$ $\beta = 3.9749 \times 10^{-3}$ $C = 3.9450 \times 10^{-3}$ $D = -2.9951 \times 10^{-4}$

$\sigma(\phi) = 6.63 \times 10^{-4}$ $\sigma(\gamma) = 4.14 \times 10^{-4}$ $\mathrm{VAR}(\phi) = 4.397 \times 10^{-7}$ $\mathrm{VAR}(\gamma) = 1.714 \times 10^{-7}$

$\sigma^*(\phi) = 7.11 \times 10^{-4}$ $\sigma^*(\gamma) = 6.51 \times 10^{-4}$

KI

Measurements have been made with potassium-amalgam cells without transference and with silver-silver iodide electrodes by Harned and Douglas (0.03–3 *m*) [152]. Wilke and Schränkler [84] used potassium-amalgam cells with transference over the concentration range, 1–6 *N*. Gelbach [*186*] used cells with and without transference for a concentration range of 0.002–2 *m*; he used potassium-amalgam and silver-silver iodide electrodes. Torto [*187*] measured concentration cells for the range, 0.1–0.5 *N*; he used silver-silver iodide electrodes. Haas and Jellinek [144] used calomel and silver-silver iodide electrodes in cells with liquid junction over the concentration range, 0.01–4 *m*. Vapor pressures were measured by Pearce, Taylor, and Bartlett (0.1 *m*-sat) [155], while isopiestic measurements were made by Robinson (0.1–4.58 *m*) [23] and Robinson and Wilson (0.1–4.5 *m*) [*188*].

TABLE 30. Osmotic and mean activity coefficients of KI

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.918	.688	1.600	.945	.638
.002	.984	.952	.500	.918	.676	1.800	.951	.637
.005	.976	.927	.600	.919	.666	2.000	.957	.638
.010	.968	.902	.700	.921	.659	2.500	.972	.643
.020	.958	.871	.800	.923	.654	3.000	.987	.650
.050	.943	.820	.900	.925	.650	3.500	1.003	.660
.100	.932	.776	1.000	.928	.646	4.000	1.018	.671
.200	.922	.731	1.200	.933	.641	4.500	1.033	.683
.300	.919	.705	1.400	.939	.639			

$B^* = 1.381$ $\beta = 2.5012 \times 10^{-2}$ $C = -1.1700 \times 10^{-4}$

$\sigma(\phi) = 1.49 \times 10^{-3}$ $\sigma(\gamma) = 1.26 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 2.213 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 1.588 \times 10^{-6}$

$\sigma^*(\phi) = 1.53 \times 10^{-3}$ $\sigma^*(\gamma) = 1.87 \times 10^{-3}$

KOH

Chow [189] calculated values for activity coefficients from 0.003 to 1.0 *N* from measurements of the emf of the cell:

$$Hg(l)|HgO(s)|KOH(c_1)|K\text{-}Hg|KOH(c_2)|HgO(s)|Hg(l).$$

Knobel [190] made revisions in Chow's measurements, and measured the emf of the cell:

$$Pt,\ H_2(g)|KOH(c_1)|K\text{-}Hg|KOH(c_2)|H_2(g),\ Pt,$$

from 0.001 to 3.0 *N*; Harned and Cook [*191*] and Åkerlöf and Bender [*192*] also measured the emf of this cell, the former for the range, 0.05–4.0 *m* and the latter for 0.2240–17.544 *m*. Harned [161] measured the emf of the cell:

$$Pt,\ H_2(g)|KOH(c_2),\ KCl(c)|K\text{-}Hg|KOH(c_1)|H_2,$$

for c_2 and $c_1 \cong 0.01$ and 0.1 *m* and $c \cong 0.02$ to 3.0 *m*. Kangro and Groeneveld [*126*] measured the vapor pressure for concentrations from 1.0 to 20.0 *m*.

TABLE 31. Osmotic and mean activity coefficients of KOH

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	1.000	1.000	.733	8.000	1.957	3.65
.002	.984	.952	1.200	1.023	.752	9.000	2.095	4.73
.005	.976	.927	1.400	1.047	.774	10.000	2.229	6.11
.010	.968	.902	1.600	1.072	.800	11.000	2.357	7.86
.020	.958	.871	1.800	1.097	.828	12.000	2.480	10.0
.050	.944	.821	2.000	1.123	.860	13.000	2.594	12.7
.100	.934	.779	2.500	1.189	.950	14.000	2.700	16.0
.200	.930	.740	3.000	1.256	1.058	15.000	2.796	19.9
.300	.934	.722	3.500	1.324	1.184	16.000	2.880	24.3
.400	.940	.713	4.000	1.393	1.331	17.000	2.952	29.4
.500	.948	.710	4.500	1.463	1.501	18.000	3.009	34.9
.600	.957	.711	5.000	1.533	1.697	19.000	3.052	40.6
.700	.967	.714	5.500	1.604	1.923	20.000	3.079	46.4
.800	.978	.718	6.000	1.675	2.18			
.900	.989	.725	7.000	1.817	2.82			

$B^* = 1.15$ $\beta = 1.0000 \times 10^{-1}$ $C = 2.6270 \times 10^{-3}$ $D = -1.3000 \times 10^{-4}$
$E = 1.7000 \times 10^{-7}$

$\sigma(\phi) = 1.19 \times 10^{-2}$ $\sigma(\gamma) = 2.92 \times 10^{-1}$ $VAR(\phi) = 1.426 \times 10^{-4}$ $VAR(\gamma) = 8.556 \times 10^{-2}$
$\sigma^*(\phi) = 4.98 \times 10^{-3}$ $\sigma^*(\gamma) = 1.27 \times 10^{-2}$

$KClO_3$

Jones and Froning [*168*] made isopiestic measurements from 0.2 to 0.7 *m*.

TABLE 32. Osmotic and mean activity coefficients of $KClO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.050	.934	.805	.500	.833	.569
.002	.984	.951	.100	.914	.749	.600	.818	.543
.005	.975	.926	.200	.887	.681	.700	.803	.519
.010	.966	.899	.300	.867	.635			
.020	.955	.865	.400	.849	.599			

$B^* = 1.325$ $\beta = -1.1815 \times 10^{-1}$ $C = 1.5718 \times 10^{-3}$

$\sigma(\phi) = 4.40 \times 10^{-4}$ $\sigma(\gamma) = 4.60 \times 10^{-4}$ $VAR(\phi) = 1.650 \times 10^{-7}$ $VAR(\gamma) = 2.290 \times 10^{-7}$
$\sigma^*(\phi) = 2.20 \times 10^{-5}$ $\sigma^*(\gamma) = 5.80 \times 10^{-4}$

$KBrO_3$

Jones [*193*] made isopiestic measurements from 0.15 to 0.50 *m*.

TABLE 33. Osmotic and mean activity coefficients of $KBrO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.020	.954	.863	.300	.858	.624
.002	.984	.951	.050	.933	.803	.400	.838	.585
.005	.975	.925	.100	.911	.744	.500	.819	.552
.010	.966	.898	.200	.881	.673			

$B^* = 1.30$ $\beta = -1.3900 \times 10^{-1}$

$\sigma(\phi) = 1.14 \times 10^{-3}$ $\sigma(\gamma) = 7.90 \times 10^{-4}$ $VAR(\phi) = 1.292 \times 10^{-6}$ $VAR(\gamma) = 6.211 \times 10^{-7}$

$\sigma^*(\phi) = 1.34 \times 10^{-3}$ $\sigma^*(\gamma) = 1.16 \times 10^{-3}$

KNO_3

Kangro and Groeneveld [126] measured the vapor pressures for concentrations for 1.0 to 3.0 *m*. Robinson [*134*] made isopiestic measurements from 0.1 to 3.5 *m*. Harned [*9*] used diffusion measurement to determine activity coefficients for dilute solutions, 0.0005 to .001 *N*; Harned and Shropshire [*132*] extended the measurements to 0.02 *N*.

TABLE 34. Osmotic and mean activity coefficients of KNO_3

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.300	.855	.614	1.200	.737	.415
.002	.983	.950	.400	.837	.577	1.400	.717	.390
.005	.974	.924	.500	.821	.546	1.600	.699	.369
.010	.965	.896	.600	.807	.521	1.800	.682	.350
.020	.952	.860	.700	.793	.498	2.000	.667	.332
.050	.929	.797	.800	.781	.478	2.500	.631	.297
.100	.906	.735	.900	.769	.460	3.000	.602	.269
.200	.876	.662	1.000	.758	.444	3.500	.579	.246

$B^* = 9.7500 \times 10^{-1}$ $\beta = -1.0090 \times 10^{-1}$ $C = 6.8017 \times 10^{-3}$

$\sigma(\phi) = 2.22 \times 10^{-3}$ $\sigma(\gamma) = 1.26 \times 10^{-3}$ $VAR(\phi) = 4.946 \times 10^{-6}$ $VAR(\gamma) = 1.575 \times 10^{-6}$

$\sigma^*(\phi) = 2.80 \times 10^{-3}$ $\sigma^*(\gamma) = 2.24 \times 10^{-3}$

KH_2PO_4

Isopiestic measurements have been made by Stokes (0.1–1.8 *m*) [*171*] and Scatchard and Breckenridge (0.1–1.3 *m*) [*172*].

TABLE 35. Osmotic and mean activity coefficients of KH_2PO_4

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.200	.869	.652	.900	.749	.439
.002	.983	.950	.300	.845	.601	1.000	.737	.422
.005	.974	.924	.400	.824	.561	1.200	.716	.393
.010	.964	.896	.500	.806	.529	1.400	.698	.369
.020	.951	.859	.600	.790	.502	1.600	.682	.348
.050	.927	.793	.700	.775	.478	1.800	.668	.331
.100	.902	.730	.800	.762	.457			

$B^* = 9.5000 \times 10^{-1}$ $\beta = -1.2876 \times 10^{-1}$ $C = 1.6212 \times 10^{-2}$

$\sigma(\phi) = 1.50 \times 10^{-3}$ $\sigma(\gamma) = 1.66 \times 10^{-3}$ $VAR(\phi) = 2.242 \times 10^{-6}$ $VAR(\gamma) = 2.742 \times 10^{-6}$

$\sigma^*(\phi) = 1.84 \times 10^{-3}$ $\sigma^*(\gamma) = 2.48 \times 10^{-3}$

KH_2AsO_4

Isopiestic measurements have been made by Scatchard and Breckenridge [*172*] for concentrations from 0.1 to 1.3 *m*.

TABLE 36. Osmotic and mean activity coefficients of KH_2AsO_4

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.100	.913	.748	.700	.802	.516
.002	.984	.951	.200	.884	.678	.800	.791	.496
.005	.975	.926	.300	.863	.630	.900	.781	.479
.010	.966	.899	.400	.845	.594	1.000	.771	.463
.020	.955	.865	.500	.829	.564	1.200	.756	.437
.050	.934	.805	.600	.815	.538	1.300	.750	.426

$B^* = 1.425$ $\beta = -1.5314 \times 10^{-1}$ $C = 2.9695 \times 10^{-2}$

$\sigma(\phi) = 1.32 \times 10^{-3}$ $\sigma(\gamma) = 1.01 \times 10^{-3}$ $VAR(\phi) = 1.750 \times 10^{-6}$ $VAR(\gamma) = 1.026 \times 10^{-6}$
$\sigma^*(\phi) = 1.59 \times 10^{-3}$ $\sigma^*(\gamma) = 1.74 \times 10^{-3}$

KCNS

Measurements of vapor pressures over the concentration range, 0–10 *m*, were made by Pearce and Hopson [169]. Isopiestic measurements from 0.1 to 5.0 *m* were made by Robinson [*173*].

TABLE 37. Osmotic and mean activity coefficients of KCNS

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.901	.664	1.600	.892	.569
.002	.984	.951	.500	.898	.647	1.800	.893	.562
.005	.976	.927	.600	.896	.634	2.000	.894	.556
.010	.967	.901	.700	.895	.623	2.500	.897	.545
.020	.957	.869	.800	.894	.614	3.000	.901	.537
.050	.940	.815	.900	.893	.605	3.500	.906	.532
.100	.926	.768	1.000	.892	.598	4.000	.911	.528
.200	.913	.716	1.200	.892	.586	4.500	.917	.526
.300	.906	.686	1.400	.892	.577	5.000	.923	.525

$B^* = 1.30$ $\beta = -1.8501 \times 10^{-3}$ $C = 8.5729 \times 10^{-4}$

$\sigma(\phi) = 1.09 \times 10^{-3}$ $\sigma(\gamma) = 7.00 \times 10^{-4}$ $VAR(\phi) = 1.182 \times 10^{-6}$ $VAR(\gamma) = 4.928 \times 10^{-7}$
$\sigma^*(\phi) = 1.20 \times 10^{-3}$ $\sigma^*(\gamma) = 1.17 \times 10^{-3}$

KPF_6

Isopiestic measurements from 0.1 to 0.5 *m* were made by Robinson, Stokes, and Stokes [*194*].

TABLE 38. Osmotic and mean activity coefficients of KPF_6

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.020	.947	.851	.300	.792	.537
.002	.983	.949	.050	.916	.776	.400	.761	.488
.005	.973	.922	.100	.881	.699	.500	.737	.450
.010	.962	.892	.200	.831	.602			

$B^* = 1.00$ $\beta = -3.3970 \times 10^{-1}$ $C = 1.4050 \times 10^{-1}$

$\sigma(\phi) = 1.75 \times 10^{-3}$ $\sigma(\gamma) = 2.15 \times 10^{-3}$ $VAR(\phi) = 3.056 \times 10^{-6}$ $VAR(\gamma) = 4.614 \times 10^{-6}$
$\sigma^*(\phi) = 2.14 \times 10^{-3}$ $\sigma^*(\gamma) = 4.00 \times 10^{-3}$

RbF

Isopiestic measurements from 0.1 to 3.5 *m* were made by Tien [195].

TABLE 39. Osmotic and mean activity coefficients of RbF

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.300	.930	.719	1.200	.977	.701
.002	.984	.952	.400	.934	.708	1.400	.987	.706
.005	.976	.927	.500	.939	.701	1.600	.996	.711
.010	.968	.902	.600	.944	.698	1.800	1.005	.718
.020	.958	.871	.700	.950	.696	2.000	1.013	.724
.050	.944	.821	.800	.955	.695	2.500	1.032	.742
.100	.934	.780	.900	.961	.696	3.000	1.051	.762
.200	.929	.739	1.000	.967	.697	3.500	1.073	.786

$B^*=1.198$ $\beta = 9.1520 \times 10^{-2}$ $C = -1.7980 \times 10^{-2}$ $D = 2.1090 \times 10^{-3}$

$\sigma(\phi) = 5.81 \times 10^{-4}$ $\sigma(\gamma) = 1.25 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 3.533 \times 10^{-7}$ $\mathrm{VAR}(\gamma) = 1.570 \times 10^{-4}$

$\sigma^*(\phi) = 7.60 \times 10^{-4}$ $\sigma^*(\gamma) = 2.70 \times 10^{-2}$

RbCl

Harned [9] used diffusion measurements to determine activity coefficients for dilute solutions, 0.0005–0.001 *N*. Lebed and Aleksandrov [124] measured the emf of cells without transference from 0.005 to 0.2 *m*; they used rubidium-amalgam and silver-silver chloride electrodes. Isopiestic measurements were made by Robinson and Sinclair (0.1–4 *m*) [30], Robinson (0.1–5 *m*) [196], and Makarov, Evstropev, and Vlasov (5–7.78 *m*) [197].

TABLE 40. Osmotic and mean activity coefficients of RbCl

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.600	.888	.619	3.000	.916	.536
.002	.984	.951	.700	.887	.608	3.500	.928	.535
.005	.975	.926	.800	.886	.598	4.000	.940	.537
.010	.967	.900	.900	.886	.590	4.500	.952	.540
.020	.956	.867	1.000	.886	.583	5.000	.964	.544
.050	.938	.811	1.200	.887	.571	5.500	.976	.549
.100	.923	.761	1.400	.888	.562	6.000	.987	.554
.200	.907	.707	1.600	.891	.555	7.000	1.005	.564
.300	.899	.674	1.800	.893	.550	7.500	1.012	.568
.400	.894	.651	2.000	.896	.546	7.800	1.015	.570
.500	.891	.633	2.500	.906	.539			

$B^*=1.1439$ $\beta = 8.1000 \times 10^{-4}$ $C = 3.2460 \times 10^{-3}$ $D = -2.2672 \times 10^{-4}$

$\sigma(\phi) = 2.24 \times 10^{-3}$ $\sigma(\gamma) = 2.39 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 5.015 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 5.710 \times 10^{-6}$

$\sigma^*(\phi) = 2.38 \times 10^{-3}$ $\sigma^*(\gamma) = 3.28 \times 10^{-3}$

RbBr

Isopiestic measurements were made by Robinson (0.1–5 *m*) [23, 196]

TABLE 41. Osmotic and mean activity coefficients of RbBr

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.892	.648	1.600	.883	.547
.002	.984	.951	.500	.889	.630	1.800	.885	.541
.005	.975	.926	.600	.886	.616	2.000	.886	.535
.010	.967	.900	.700	.884	.604	2.500	.892	.525
.020	.956	.866	.800	.883	.594	3.000	.899	.519
.050	.938	.811	.900	.882	.585	3.500	.907	.515
.100	.922	.760	1.000	.882	.578	4.000	.916	.514
.200	.906	.705	1.200	.881	.565	4.500	.925	.514
.300	.898	.672	1.400	.882	.555	5.000	.934	.514

$B^* = 1.1343$ $\beta = -8.9000 \times 10^{-4}$ $C = 2.2110 \times 10^{-3}$ $D = -1.3210 \times 10^{-4}$

$\sigma(\phi) = 6.82 \times 10^{-4}$ $\sigma(\gamma) = 8.56 \times 10^{-4}$ $VAR(\phi) = 4.656 \times 10^{-7}$ $VAR(\gamma) = 7.320 \times 10^{-7}$

$\sigma^*(\phi) = 7.64 \times 10^{-4}$ $\sigma^*(\gamma) = 1.33 \times 10^{-3}$

RbI

Isopiestic measurements were made by Robinson (0.2–5.1 *m*) [*23*] and (0.1–5 *m*) [*196*].

TABLE 42. Osmotic and mean activity coefficients of RbI

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.891	.646	1.600	.882	.544
.002	.984	.951	.500	.887	.627	1.800	.883	.537
.005	.975	.926	.600	.884	.613	2.000	.886	.532
.010	.967	.900	.700	.882	.601	2.500	.893	.523
.020	.955	.866	.800	.881	.591	3.000	.901	.517
.050	.937	.810	.900	.880	.582	3.500	.910	.515
.100	.921	.759	1.000	.880	.574	4.000	.921	.514
.200	.905	.703	1.200	.880	.562	4.500	.931	.515
.300	.896	.670	1.400	.880	.552	5.000	.941	.517

$B^* = 1.109$ $\beta = -1.2600 \times 10^{-3}$ $C = 2.8000 \times 10^{-3}$ $D = -1.7740 \times 10^{-4}$

$\sigma(\phi) = 1.02 \times 10^{-3}$ $\sigma(\gamma) = 1.07 \times 10^{-3}$ $VAR(\phi) = 1.307 \times 10^{-6}$ $VAR(\gamma) = 1.152 \times 10^{-6}$

$\sigma^*(\phi) = 1.72 \times 10^{-3}$ $\sigma^*(\gamma) = 1.16 \times 10^{-3}$

$RbNO_3$

Isopiestic measurements were made by Robinson [*196*] for concentrations of 0.1 to 4.5 *m*.

TABLE 43. Osmotic and mean activity coefficients of $RbNO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.400	.829	.568	1.600	.685	.355
.002	.983	.950	.500	.812	.536	1.800	.669	.336
.005	.974	.924	.600	.796	.509	2.000	.654	.320
.010	.965	.896	.700	.782	.486	2.500	.622	.285
.020	.952	.859	.800	.769	.465	3.000	.594	.258
.050	.928	.795	.900	.756	.447	3.500	.569	.236
.100	.904	.733	1.000	.744	.430	4.000	.545	.218
.200	.872	.657	1.200	.723	.401	4.500	.521	.201
.300	.849	.607	1.400	.703	.376			

$B^* = 1.00$ $\beta = -1.2621 \times 10^{-1}$ $C = 1.6706 \times 10^{-2}$ $D = -1.3194 \times 10^{-3}$

$\sigma(\phi) = 3.16 \times 10^{-3}$ $\sigma(\gamma) = 2.23 \times 10^{-3}$ $VAR(\phi) = 9.986 \times 10^{-6}$ $VAR(\gamma) = 4.973 \times 10^{-6}$

$\sigma^*(\phi) = 5.30 \times 10^{-3}$ $\sigma^*(\gamma) = 3.61 \times 10^{-3}$

CsF

Isopiestic measurements were made by Tien [*195*] for concentrations of 0.1 to 3.5 *m*.

TABLE 44. Osmotic and mean activity coefficients of CsF

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.937	.736	1.200	.997	.736
.002	.984	.952	.400	.941	.726	1.400	1.014	.750
.005	.977	.929	.500	.946	.721	1.600	1.032	.765
.010	.969	.905	.600	.952	.718	1.800	1.050	.783
.020	.961	.876	.700	.958	.718	2.000	1.069	.803
.050	.948	.830	.800	.965	.719	2.500	1.113	.857
.100	.940	.792	.900	.973	.722	3.000	1.155	.915
.200	.936	.755	1.000	.980	.726	3.500	1.189	.972

$B^* = 1.674$ $\beta = 3.9269 \times 10^{-2}$ $C = 1.4799 \times 10^{-2}$ $D = -2.3270 \times 10^{-3}$

$\sigma(\phi) = 1.30 \times 10^{-3}$ $\sigma(\gamma) = 1.10 \times 10^{-3}$ $VAR(\phi) = 1.680 \times 10^{-6}$ $VAR(\gamma) = 1.210 \times 10^{-4}$

$\sigma^*(\phi) = 2.76 \times 10^{-3}$ $\sigma^*(\gamma) = 3.23 \times 10^{-3}$

CsCl

Measurements of cells without transference and constructed with cesium-amalgam and silver-silver chloride electrodes were made by Harned and Schupp (0.001–3 *m*) [*198*], Caramazza (0.1–6 *m*) [*199*], and Lebed and Aleksandrov (0.005–0.2 *m*) [124]. Harned [*9*] used diffusion measurements to determine activity coefficients for dilute solutions, 0.0005–0.001 *N*. Isopiestic measurements were made by Robinson and Sinclair (0.1–5 *m*) [*30*] and Makarov, Evstropev, and Vlasov (5–11 *m*) [*197*].

TABLE 45. Osmotic and mean activity coefficients of CsCl

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.700	.866	.577	4.000	.899	.473
.002	.984	.951	.800	.864	.565	4.500	.911	.473
.005	.975	.925	.900	.862	.555	5.000	.923	.474
.010	.966	.898	1.000	.861	.546	5.500	.935	.477
.020	.954	.864	1.200	.859	.532	6.000	.947	.480
.050	.935	.805	1.400	.859	.520	7.000	.969	.488
.100	.917	.751	1.600	.859	.510	8.000	.987	.495
.200	.897	.691	1.800	.860	.503	9.000	1.002	.503
.300	.886	.654	2.000	.862	.496	10.000	1.013	.508
.400	.879	.627	2.500	.869	.485	11.000	1.019	.512
.500	.873	.607	3.000	.877	.478			
.600	.869	.590	3.500	.888	.474			

$B^* = 9.7500 \times 10^{-1}$ $\beta = -8.5074 \times 10^{-3}$ $C = 4.8702 \times 10^{-3}$ $D = -4.0260 \times 10^{-4}$

$E = 1.0233 \times 10^{-5}$

$\sigma(\phi) = 2.90 \times 10^{-3}$ $\sigma(\gamma) = 1.63 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 8.434 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 2.663 \times 10^{-6}$

$\sigma^*(\phi) = 3.24 \times 10^{-3}$ $\sigma^*(\gamma) = 2.69 \times 10^{-3}$

CsBr

Isopiestic measurements were made by Robinson [*23*, *196*] for the concentrations from 0.1 to 5 *m*.

TABLE 46. Osmotic and mean activity coefficients of CsBr

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.877	.626	1.600	.849	.501
.002	.984	.951	.500	.870	.605	1.800	.849	.493
.005	.975	.925	.600	.865	.588	2.000	.850	.485
.010	.966	.898	.700	.861	.573	2.500	.856	.472
.020	.954	.864	.800	.858	.561	3.000	.864	.464
.050	.935	.806	.900	.856	.550	3.500	.874	.460
.100	.917	.752	1.000	.853	.540	4.000	.884	.457
.200	.897	.691	1.200	.851	.524	4.500	.894	.455
.300	.885	.654	1.400	.849	.512	5.000	.902	.454

$B^* = 1.05$ $\beta = -2.6208 \times 10^{-2}$ $C = 8.7500 \times 10^{-3}$ $D = -7.1511 \times 10^{-4}$

$\sigma(\phi) = 2.48 \times 10^{-3}$ $\sigma(\gamma) = 1.56 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 6.149 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 2.437 \times 10^{-8}$

$\sigma^*(\phi) = 2.86 \times 10^{-3}$ $\sigma^*(\gamma) = 2.55 \times 10^{-3}$

CsI

Isopiestic measurements were made by Robinson for the concentration ranges, 0.2–3 *m*[*23*] and 0.1–3 *m* [*196*].

TABLE 47. Osmotic and mean activity coefficients of CsI

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.300	.883	.649	1.200	.845	.517
.002	.983	.951	.400	.875	.622	1.400	.841	.503
.005	.975	.925	.500	.868	.601	1.600	.838	.491
.010	.966	.898	.600	.863	.583	1.800	.835	.480
.020	.954	.863	.700	.859	.568	2.000	.832	.470
.050	.934	.804	.800	.855	.555	2.500	.827	.450
.100	.915	.749	.900	.852	.544	3.000	.822	.434
.200	.895	.688	1.000	.850	.534			

$B^* = 9.3650 \times 10^{-1}$ $\beta = -8.3798 \times 10^{-3}$

$\sigma(\phi) = 2.66 \times 10^{-3}$ $\sigma(\gamma) = 1.80 \times 10^{-3}$ $VAR(\phi) = 7.049 \times 10^{-6}$ $VAR(\gamma) = 3.233 \times 10^{-6}$

$\sigma^*(\phi) = 3.08 \times 10^{-3}$ $\sigma^*(\gamma) = 2.82 \times 10^{-3}$

CsOH

Harned and Schupp [200] determined the activity coefficients from the emfs of the cell:

$$Pt, H_2(g)|CsOH(m)|Cs{-}Hg|CsOH(0.05)|H_2(g), Pt,$$

with *m* extending from 0.1 to 1.2.

TABLE 48. Osmotic and mean activity coefficients of CsOH

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.966	.100	.946	.802	.700	.984	.762
.002	.985	.953	.200	.946	.772	.800	.993	.768
.005	.977	.930	.300	.952	.760	.900	1.002	.774
.010	.970	.906	.400	.959	.756	1.000	1.011	.782
.020	.962	.878	.500	.966	.755	1.200	1.029	.799
.050	.952	.836	.600	.975	.758			

$B^* = 1.763$ $\beta = 7.8000 \times 10^{-2}$

$\sigma(\phi) = 5.49 \times 10^{-3}$ $\sigma(\gamma) = 2.09 \times 10^{-3}$ $VAR(\phi) = 3.014 \times 10^{-5}$ $VAR(\gamma) = 4.376 \times 10^{-6}$

$\sigma^*(\phi) = 5.66 \times 10^{-3}$ $\sigma^*(\gamma) = 2.71 \times 10^{-3}$

$CsNO_3$

Isopiestic measurements were made by Robinson [196] for the concentration range, 0.1–1.5 *m*.

TABLE 49. Osmotic and mean activity coefficients of $CsNO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.200	.870	.655	.900	.746	.438
.002	.983	.951	.300	.844	.603	1.000	.735	.421
.005	.974	.924	.400	.822	.562	1.200	.715	.393
.010	.965	.897	.500	.804	.529	1.400	.698	.369
.020	.952	.860	.600	.787	.501	1.500	.690	.358
.050	.929	.796	.700	.772	.477			
.100	.904	.733	.800	.759	.457			

$B^* = 1.225$ $\beta = -1.9036 \times 10^{-1}$ $C = 5.3954 \times 10^{-2}$ $D = -9.2416 \times 10^{-3}$

$\sigma(\phi) = 1.34 \times 10^{-3}$ $\sigma(\gamma) = 7.23 \times 10^{-4}$ $VAR(\phi) = 1.800 \times 10^{-6}$ $VAR(\gamma) = 5.230 \times 10^{-7}$

$\sigma^*(\phi) = 1.87 \times 10^{-3}$ $\sigma^*(\gamma) = 1.50 \times 10^{-3}$

$AgNO_3$

Harned [*9*] used the diffusion method to determine activity coefficients for dilute solutions, 0.0005–0.001. MacInnes and Brown [*201*] measured the emfs of the cell:

$$Ag(s)|AgNO_3(c_1)|AgNO_3(c_2)|Ag(s),$$

for concentrations from 0.002 to 0.10 *N*. Kangro and Groeneveld [126] measured vapor pressures for concentrations from 1.0 to 15.0. Isopiestic measurements were made for concentrations from 0.1 to 13.5 *m* by Robinson and Tait [*202*].

TABLE 50. Osmotic and mean activity coefficients of $AgNO_3$#

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.800	.769	.465	5.000	.484	.181
.002	.983	.950	.900	.756	.447	5.500	.467	.169
.005	.974	.924	1.000	.744	.430	6.000	.452	.159
.010	.964	.896	1.200	.721	.400	7.000	.426	.142
.020	.952	.859	1.400	.701	.375	8.000	.404	.128
.050	.928	.794	1.600	.681	.353	9.000	.386	.117
.100	.904	.732	1.800	.663	.334	10.000	.370	.108
.200	.872	.656	2.000	.647	.316	11.000	.357	.101
.300	.849	.606	2.500	.609	.280	12.000	.350	.0945
.400	.829	.567	3.000	.577	.252	13.000	.351	.0897
.500	.812	.536	3.500	.549	.229	14.000	.351 [a]	.0856[a]
.600	.797	.509	4.000	.524	.210	15.000	.347 [a]	.0851[a]
.700	.782	.486	4.500	.503	.194			

$B^* = 9.5000 \times 10^{-1}$ $\beta = -1.1493 \times 10^{-1}$ $C = 1.0846 \times 10^{-2}$ $D = -6.8493 \times 10^{-4}$

$E = 1.8730 \times 10^{-5}$

$\sigma(\phi) = 6.98 \times 10^{-4}$ $\sigma(\gamma) = 1.29 \times 10^{-4}$ $VAR(\phi) = 4.872 \times 10^{-7}$ $VAR(\gamma) = 1.664 \times 10^{-8}$

$\sigma^*(\phi) = 1.35 \times 10^{-3}$ $\sigma^*(\gamma) = 3.78 \times 10^{-4}$

a – these data are those reported by Kangro and Groeneveld [126].

\# – see also Table 2.

TlCl

Cowptherwaite, La Mer, and Barksdale [203] measured the emf of the cell:

$$Tl(s)|TlCl(m)|AgCl(s)|Ag(s),$$

from 0.0005 to 0.01 *m*.

TABLE 51. Osmotic and mean activity coefficients of TlCl

m	ϕ	γ	m	ϕ	γ
.001	.987	.962	.005	.969	.913
.002	.981	.946	.010	.953	.876

$B^* = 1.00$ $\beta = -1.12$

$\sigma(\phi) = 1.10 \times 10^{-2}$ $\sigma(\gamma) = 8.30 \times 10^{-4}$ $VAR(\phi) = 1.212 \times 10^{-4}$ $VAR(\gamma) = 6.942 \times 10^{-7}$

$\sigma^*(\phi) = 1.15 \times 10^{-2}$ $\sigma^*(\gamma) = 9.10 \times 10^{-4}$

$TlClO_4$

Isopiestic measurements were made for concentrations from 0.025 to 0.5 *m* by Robinson [*196*].

TABLE 52. Osmotic and mean activity coefficients of $TlClO_4$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.020	.951	.859	.300	.844	.601
.002	.983	.950	.050	.927	.794	.400	.823	.561
.005	.974	.924	.100	.903	.730	.500	.803	.527
.010	.964	.896	.200	.869	.653			

$B^* = 9.4560 \times 10^{-1}$ $\beta = -1.2273 \times 10^{-1}$

$\sigma(\phi) = 1.58 \times 10^{-3}$ $\sigma(\gamma) = 1.44 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 2.486 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 2.067 \times 10^{-6}$

$\sigma^*(\phi) = 1.80 \times 10^{-3}$ $\sigma^*(\gamma) = 2.16 \times 10^{-3}$

$TlNO_3$

Isopiestic measurements were made for concentrations from 0.025 to 0.4 *m* by Robinson [*196*].

TABLE 53. Osmotic and mean activity coefficients of $TlNO_3$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.020	.946	.849	.300	.804	.546
.002	.983	.949	.050	.916	.774	.400	.776	.500
.005	.973	.921	.100	.883	.699			
.010	.961	.890	.200	.837	.607			

$B^* = 1.7000 \times 10^{-1}$ $\beta = -2.4000 \times 10^{-2}$

$\sigma(\phi) = 3.62 \times 10^{-3}$ $\sigma(\gamma) = 1.95 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 1.309 \times 10^{-5}$ $\mathrm{VAR}(\gamma) = 3.819 \times 10^{-6}$

$\sigma^*(\phi) = 4.39 \times 10^{-3}$ $\sigma^*(\gamma) = 2.85 \times 10^{-3}$

$TlNO_2$

Nardelli, Braibanti, and Chierici [204] measured the emfs of the cell:

$$Tl(s)|TlNO_2(m)|TlNO_2(m')|TlNO_2(0.00295)|Tl(s),$$

for values of *m* from 0.01512 to 1.3514; *m′* was varied with successive values intermediate to *m* and 0.00295.

TABLE 54. Osmotic and mean activity coefficients of $TlNO_2$

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.987	.963	.100	.843	.645	.700	.552	.275
.002	.982	.948	.200	.762	.520	.800	.538	.255
.005	.971	.917	.300	.698	.438	.900	.532	.240
.010	.958	.883	.400	.647	.378	1.000	.534	.229
.020	.938	.836	.500	.606	.334	1.200	.563	.217
.050	.896	.744	.600	.574	.300	1.400	.624	.216

$B^* = 6.0000 \times 10^{-1}$ $\beta = -5.6846 \times 10^{-1}$ $C = 2.4715 \times 10^{-1}$

$\sigma(\phi) = 2.91 \times 10^{-1}$ $\sigma(\gamma) = 4.19 \times 10^{-2}$ $VAR(\phi) = 8.480 \times 10^{-2}$ $VAR(\gamma) = 1.756 \times 10^{-3}$

$\sigma^*(\phi) = 4.38 \times 10^{-1}$ $\sigma^*(\gamma) = 6.80 \times 10^{-2}$

NH_4Cl

Vapor pressures for concentrations from 0.1–7.38 *m*(sat) were measured by Pearce and Pumplin [205]. Isopiestic measurements were made by Wishaw and Stokes (0.1–7.39 *m*, sat) [*206*] and Shul'ts, Makarov, and Su Yu-jêng (5.0–7.42 *m*) [207].

TABLE 55. Osmotic and mean activity coefficients of NH_4Cl

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.600	.898	.636	3.000	.926	.560
.002	.984	.952	.700	.897	.626	3.500	.935	.560
.005	.976	.927	.800	.897	.617	4.000	.944	.560
.010	.967	.901	.900	.897	.609	4.500	.953	.561
.020	.957	.869	1.000	.897	.603	5.000	.960	.563
.050	.941	.816	1.200	.898	.592	5.500	.966	.564
.100	.927	.769	1.400	.900	.584	6.000	.970	.565
.200	.913	.718	1.600	.902	.578	7.000	.975	.566
.300	.906	.687	1.800	.905	.573	7.405 [a]	.976	.565
.400	.902	.666	2.000	.908	.569			
.500	.900	.649	2.500	.917	.563			

$B^* = 1.325$ $\beta = -4.5787 \times 10^{-3}$ $C = 5.2712 \times 10^{-3}$ $D = -7.0557 \times 10^{-4}$

$E = 2.8434 \times 10^{-5}$

$\sigma(\phi) = 8.67 \times 10^{-4}$ $\sigma(\gamma) = 6.27 \times 10^{-4}$ $VAR(\phi) = 7.520 \times 10^{-7}$ $VAR(\gamma) = 3.930 \times 10^{-7}$

$\sigma^*(\phi) = 9.27 \times 10^{-4}$ $\sigma^*(\gamma) = 1.09 \times 10^{-3}$

a – saturated solution.

NH_4ClO_4

Isopiestic measurements were made by Esval and Tyree [*208*] for concentrations from 0.1 to 2.1 *m*.

TABLE 56. Osmotic and mean activity coefficients of NH_4ClO_4

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	.300	.862	.619	1.200	.797	.458
.002	.983	.950	.400	.849	.586	1.400	.789	.440
.005	.974	.924	.500	.839	.560	1.600	.782	.425
.010	.964	.895	.600	.830	.539	1.800	.776	.411
.020	.951	.859	.700	.823	.521	2.000	.770	.399
.050	.928	.794	.800	.817	.505	2.100	.768	.394
.100	.906	.734	.900	.811	.491			
.200	.879	.663	1.000	.806	.479			

$B^* = 6.7500 \times 10^{-1}$ $\beta = -1.4567 \times 10^{-2}$

$\sigma(\phi) = 2.60 \times 10^{-3}$ $\sigma(\gamma) = 1.85 \times 10^{-3}$ $VAR(\phi) = 6.776 \times 10^{-6}$ $VAR(\gamma) = 3.407 \times 10^{-6}$

$\sigma^*(\phi) = 3.17 \times 10^{-3}$ $\sigma^*(\gamma) = 3.39 \times 10^{-3}$

NH_4NO_3

Isopiestic measurements were made by Wishaw and Stokes [*209*] for concentrations from 0.1 to 25.954 (sat.).

TABLE 57. Osmotic and mean activity coefficients of NH_4NO_3

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.964	1.400	.801	.462	12.000	.592	.200
.002	.983	.951	1.600	.792	.446	13.000	.584	.192
.005	.975	.925	1.800	.784	.432	14.000	.578	.185
.010	.965	.897	2.000	.777	.419	15.000	.572	.179
.020	.953	.862	2.500	.760	.391	16.000	.566	.173
.050	.932	.801	3.000	.744	.368	17.000	.562	.168
.100	.912	.744	3.500	.730	.349	18.000	.557	.163
.200	.889	.678	4.000	.717	.332	19.000	.552	.158
.300	.873	.637	4.500	.705	.316	20.000	.548	.154
.400	.862	.606	5.000	.693	.303	21.000	.543	.150
.500	.853	.582	5.500	.682	.291	22.000	.537	.146
.600	.845	.561	6.000	.672	.280	23.000	.531	.142
.700	.837	.544	7.000	.654	.261	24.000	.524	.138
.800	.831	.528	8.000	.638	.245	25.000	.516	.134
.900	.825	.515	9.000	.624	.231	25.954[a]	.507	.131
1.000	.820	.502	10.000	.612	.220			
1.200	.810	.481	11.000	.601	.209			

$B^* = 9.2500 \times 10^{-1}$ $\beta = -3.4748 \times 10^{-2}$ $C = 1.1978 \times 10^{-3}$ $D = -1.9075 \times 10^{-5}$

$\sigma(\phi) = 2.99 \times 10^{-3}$ $\sigma(\gamma) = 1.07 \times 10^{-3}$ $VAR(\phi) = 8.940 \times 10^{-6}$ $VAR(\gamma) = 1.145 \times 10^{-6}$

$\sigma^*(\phi) = 4.76 \times 10^{-3}$ $\sigma^*(\gamma) = 3.06 \times 10^{-3}$

a – saturated solution.

p-Toluene Sulfonates

Isopiestic measurements on lithium, sodium, and potassium *p*-toluene sulfonates were made by Robinson [*134*].

TABLE 58. Osmotic and mean activity coefficients of Li p–toluene sulfonate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.911	.679	1.600	.893	.583
.002	.984	.952	.500	.909	.664	1.800	.891	.575
.005	.976	.927	.600	.908	.652	2.000	.891	.568
.010	.968	.902	.700	.906	.642	2.500	.897	.558
.020	.958	.870	.800	.904	.633	3.000	.911	.556
.050	.942	.818	.900	.903	.624	3.500	.932	.561
.100	.930	.773	1.000	.901	.617	4.000	.955	.569
.200	.919	.726	1.200	.898	.604	4.500	.969	.575
.300	.914	.698	1.400	.895	.592			

$B^* = 1.25$ $\beta = 4.4115 \times 10^{-2}$ $C = -3.7795 \times 10^{-2}$ $D = 1.2146 \times 10^{-2}$

$E = -1.1792 \times 10^{-3}$

$\sigma(\phi) = 3.00 \times 10^{-3}$ $\sigma(\gamma) = 2.47 \times 10^{-3}$ $VAR(\phi) = 9.029 \times 10^{-6}$ $VAR(\gamma) = 6.110 \times 10^{-6}$

$\sigma^*(\phi) = 3.22 \times 10^{-3}$ $\sigma^*(\gamma) = 3.84 \times 10^{-3}$

TABLE 59. Osmotic and mean activity coefficients of Na p–toluene sulfonate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.889	.649	1.600	.810	.492
.002	.984	.951	.500	.882	.628	1.800	.798	.475
.005	.976	.926	.600	.875	.610	2.000	.787	.460
.010	.967	.901	.700	.868	.594	2.500	.764	.427
.020	.956	.868	.800	.862	.580	3.000	.747	.402
.050	.939	.813	.900	.855	.566	3.500	.738	.383
.100	.924	.764	1.000	.849	.554	4.000	.734	.368
.200	.908	.709	1.200	.836	.531			
.300	.897	.675	1.400	.823	.511			

$B^* = 1.25$ $\beta = -8.5622 \times 10^{-3}$ $C = -2.8439 \times 10^{-2}$ $D = 8.2367 \times 10^{-3}$

$E = -7.0371 \times 10^{-4}$

$\sigma(\phi) = 8.62 \times 10^{-4}$ $\sigma(\gamma) = 5.42 \times 10^{-4}$ $VAR(\phi) = 7.435 \times 10^{-7}$ $VAR(\gamma) = 2.938 \times 10^{-7}$

$\sigma^*(\phi) = 9.88 \times 10^{-4}$ $\sigma^*(\gamma) = 1.11 \times 10^{-3}$

TABLE 60. Osmotic and mean activity coefficients of K p–toluene sulfonate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.300	.888	.665	1.200	.773	.476
.002	.984	.951	.400	.874	.634	1.400	.751	.449
.005	.976	.926	.500	.861	.607	1.600	.732	.425
.010	.967	.900	.600	.848	.583	1.800	.715	.405
.020	.956	.868	.700	.835	.562	2.000	.700	.387
.050	.939	.813	.800	.821	.542	2.500	.667	.348
.100	.923	.762	.900	.809	.524	3.000	.635	.317
.200	.903	.704	1.000	.796	.506	3.500	.609	.291

$B^* = 1.25$ $\beta = -1.1728 \times 10^{-2}$ $C = -1.0250 \times 10^{-1}$ $D = 5.9065 \times 10^{-2}$

$E = -1.4631 \times 10^{-2}$ $F = 1.3520 \times 10^{-3}$

$\sigma(\phi) = 1.76 \times 10^{-3}$ $\sigma(\gamma) = 1.19 \times 10^{-3}$ $VAR(\phi) = 3.098 \times 10^{-6}$ $VAR(\gamma) = 1.427 \times 10^{-6}$

$\sigma^*(\phi) = 2.63 \times 10^{-3}$ $\sigma^*(\gamma) = 2.22 \times 10^{-3}$

Salts of Monocarboxylic Acids

Isopiestic measurements on the sodium salts of formic, acetic, propionic, butyric, valeric, caproic, heptylic, caprylic, pelargonic, and capric acids were made by Smith and Robinson [*210*]. Isopiestic measurements on lithium, sodium, and potassium acetates were made by Robinson [*134*] and on thallium acetate by Robinson [*196*].

TABLE 61. Osmotic and mean activity coefficients of Na formate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.300	.923	.714	1.200	.941	.657
.002	.984	.952	.400	.923	.698	1.400	.947	.656
.005	.976	.928	.500	.924	.687	1.600	.953	.655
.010	.969	.903	.600	.926	.679	1.800	.960	.656
.020	.959	.873	.700	.928	.672	2.000	.966	.658
.050	.945	.824	.800	.930	.668	2.500	.984	.666
.100	.934	.781	.900	.932	.664	3.000	1.003	.678
.200	.926	.738	1.000	.935	.661	3.500	1.024	.694

$B^* = 1.50$ $\beta = 2.4781 \times 10^{-2}$ $C = -5.4350 \times 10^{-4}$ $D = 2.8402 \times 10^{-4}$

$\sigma(\phi) = 2.77 \times 10^{-4}$ $\sigma(\gamma) = 1.63 \times 10^{-3}$ $VAR(\phi) = 7.660 \times 10^{-8}$ $VAR(\gamma) = 2.665 \times 10^{-6}$

$\sigma^*(\phi) = 2.92 \times 10^{-4}$ $\sigma^*(\gamma) = 2.69 \times 10^{-3}$

TABLE 62. Osmotic and mean activity coefficients of Li acetate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.930	.707	1.600	.997	.708
.002	.984	.952	.500	.933	.699	1.800	1.011	.718
.005	.976	.928	.600	.937	.694	2.000	1.025	.730
.010	.969	.904	.700	.942	.690	2.500	1.060	.763
.020	.959	.873	.800	.947	.689	3.000	1.093	.800
.050	.946	.825	.900	.953	.688	3.500	1.124	.839
.100	.936	.784	1.000	.958	.689	4.000	1.151	.877
.200	.929	.743	1.200	.971	.693			
.300	.928	.721	1.400	.984	.700			

$B^* = 1.50$ $\quad \beta = 3.5673 \times 10^{-2}$ $\quad C = 8.1453 \times 10^{-3}$ $\quad D = -1.1625 \times 10^{-3}$

$\sigma(\phi) = 2.45 \times 10^{-3}$ $\quad \sigma(\gamma) = 9.66 \times 10^{-4}$ $\quad VAR(\phi) = 5.998 \times 10^{-6}$ $\quad VAR(\gamma) = 9.330 \times 10^{-8}$

$\sigma^*(\phi) = 2.51 \times 10^{-3}$ $\quad \sigma^*(\gamma) = 1.32 \times 10^{-3}$

TABLE 63. Osmotic and mean activity coefficients of Na acetate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.944	.745	1.200	1.019	.772
.002	.984	.952	.400	.951	.739	1.400	1.037	.789
.005	.977	.928	.500	.958	.737	1.600	1.054	.808
.010	.969	.905	.600	.967	.738	1.800	1.072	.828
.020	.961	.875	.700	.975	.740	2.000	1.090	.850
.050	.949	.830	.800	.984	.745	2.500	1.134	.911
.100	.942	.793	.900	.992	.750	3.000	1.179	.981
.200	.940	.760	1.000	1.001	.757	3.500	1.226	1.060

$B^* = 1.50$ $\quad \beta = 8.7320 \times 10^{-2}$ $\quad C = -4.8128 \times 10^{-3}$ $\quad D = 6.9295 \times 10^{-4}$

$\sigma(\phi) = 1.60 \times 10^{-3}$ $\quad \sigma(\gamma) = 1.47 \times 10^{-3}$ $\quad VAR(\phi) = 2.560 \times 10^{-6}$ $\quad VAR(\gamma) = 2.129 \times 10^{-6}$

$\sigma^*(\phi) = 1.51 \times 10^{-3}$ $\quad \sigma^*(\gamma) = 1.75 \times 10^{-3}$

TABLE 64. Osmotic and mean activity coefficients of K acetate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.950	.753	1.200	1.039	.804
.002	.984	.952	.400	.958	.749	1.400	1.060	.827
.005	.977	.929	.500	.967	.749	1.600	1.081	.853
.010	.969	.905	.600	.977	.753	1.800	1.103	.881
.020	.961	.876	.700	.987	.758	2.000	1.124	.910
.050	.949	.831	.800	.997	.766	2.500	1.176	.992
.100	.943	.795	.900	1.008	.774	3.000	1.230	1.086
.200	.944	.765	1.000	1.018	.783	3.500	1.286	1.195

$B^* = 1.50$ $\quad \beta = 1.0244 \times 10^{-1}$ $\quad C = -5.1136 \times 10^{-3}$ $\quad D = 7.5650 \times 10^{-4}$

$\sigma(\phi) = 2.64 \times 10^{-3}$ $\quad \sigma(\gamma) = 3.01 \times 10^{-3}$ $\quad VAR(\phi) = 6.959 \times 10^{-6}$ $\quad VAR(\gamma) = 9.072 \times 10^{-6}$

$\sigma^*(\phi) = 2.07 \times 10^{-3}$ $\quad \sigma^*(\gamma) = 2.63 \times 10^{-3}$

TABLE 65. Osmotic and mean activity coefficients of Rb acetate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.951	.755	1.200	1.047	.815
.002	.984	.952	.400	.960	.752	1.400	1.069	.841
.005	.977	.929	.500	.970	.753	1.600	1.092	.870
.010	.969	.905	.600	.980	.758	1.800	1.114	.900
.020	.961	.876	.700	.991	.764	2.000	1.137	.933
.050	.950	.831	.800	1.002	.772	2.500	1.192	1.023
.100	.944	.796	.900	1.013	.782	3.000	1.248	1.126
.200	.945	.766	1.000	1.024	.792	3.500	1.304	1.243

$B^* = 1.50$ $\beta = 1.0630 \times 10^{-1}$ $C = -3.6957 \times 10^{-3}$ $D = 4.3613 \times 10^{-4}$

$\sigma(\phi) = 7.24 \times 10^{-4}$ $\sigma(\gamma) = 1.04 \times 10^{-3}$ $VAR(\phi) = 5.240 \times 10^{-8}$ $VAR(\gamma) = 1.080 \times 10^{-6}$

$\sigma^*(\phi) = 8.00 \times 10^{-4}$ $\sigma^*(\gamma) = 1.11 \times 10^{-3}$

TABLE 66. Osmotic and mean activity coefficients of Cs acetate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.953	.759	1.200	1.053	.828
.002	.984	.952	.400	.963	.757	1.400	1.076	.855
.005	.977	.929	.500	.973	.759	1.600	1.099	.885
.010	.970	.905	.600	.984	.764	1.800	1.122	.917
.020	.961	.876	.700	.996	.772	2.000	1.144	.951
.050	.950	.832	.800	1.007	.781	2.500	1.199	1.044
.100	.945	.798	.900	1.019	.791	3.000	1.252	1.147
.200	.946	.769	1.000	1.030	.802	3.500	1.305	1.262

$B^* = 1.50$ $\beta = 1.1357 \times 10^{-1}$ $C = -5.2680 \times 10^{-3}$ $D = 4.5087 \times 10^{-4}$

$\sigma(\phi) = 2.20 \times 10^{-3}$ $\sigma(\gamma) = 2.02 \times 10^{-3}$ $VAR(\phi) = 4.858 \times 10^{-6}$ $VAR(\gamma) = 4.071 \times 10^{-6}$

$\sigma^*(\phi) = 2.06 \times 10^{-3}$ $\sigma^*(\gamma) = 2.44 \times 10^{-3}$

TABLE 67. Osmotic and mean activity coefficients of Tl acetate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.500	.857	.591	2.000	.807	.444
.002	.984	.951	.600	.849	.571	2.500	.802	.423
.005	.975	.926	.700	.842	.554	3.000	.797	.406
.010	.966	.899	.800	.837	.539	3.500	.792	.391
.020	.955	.865	.900	.832	.526	4.000	.784	.377
.050	.935	.806	1.000	.828	.515	4.500	.777	.365
.100	.916	.751	1.200	.821	.495	5.000	.769	.354
.200	.893	.687	1.400	.816	.479	5.500	.763	.344
.300	.878	.646	1.600	.812	.466	6.000	.760	.336
.400	.866	.616	1.800	.809	.454			

$B^* = 1.25$ $\beta = -8.7283 \times 10^{-2}$ $C = 3.2612 \times 10^{-2}$ $D = -7.3907 \times 10^{-3}$

$E = 7.7672 \times 10^{-4}$ $F = -2.9141 \times 10^{-5}$

$\sigma(\phi) = 1.54 \times 10^{-3}$ $\sigma(\gamma) = 1.09 \times 10^{-3}$ $VAR(\phi) = 2.373 \times 10^{-6}$ $VAR(\gamma) = 1.190 \times 10^{-6}$

$\sigma^*(\phi) = 1.80 \times 10^{-3}$ $\sigma^*(\gamma) = 2.22 \times 10^{-3}$

TABLE 68. Osmotic and mean activity coefficients of Na propionate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.954	.759	1.200	1.059	.835
.002	.984	.952	.400	.964	.758	1.400	1.083	.865
.005	.977	.929	.500	.975	.761	1.600	1.107	.897
.010	.970	.905	.600	.986	.767	1.800	1.130	.931
.020	.961	.876	.700	.998	.775	2.000	1.152	.966
.050	.950	.832	.800	1.010	.785	2.500	1.206	1.060
.100	.945	.798	.900	1.022	.796	3.000	1.254	1.160
.200	.947	.769	1.000	1.034	.808			

$B^* = 1.50$ $\beta = 1.1403 \times 10^{-1}$ $C = -1.6798 \times 10^{-3}$ $D = -6.1530 \times 10^{-4}$

$\sigma(\phi) = 1.08 \times 10^{-3}$ $\sigma(\gamma) = 1.91 \times 10^{-3}$ $VAR(\phi) = 1.164 \times 10^{-6}$ $VAR(\gamma) = 3.660 \times 10^{-6}$

$\sigma^*(\phi) = 1.05 \times 10^{-3}$ $\sigma^*(\gamma) = 2.84 \times 10^{-3}$

TABLE 69. Osmotic and mean activity coefficients of Na butyrate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.962	.770	1.200	1.104	.907
.002	.984	.952	.400	.976	.774	1.400	1.133	.951
.005	.977	.929	.500	.991	.783	1.600	1.160	.996
.010	.970	.905	.600	1.007	.795	1.800	1.183	1.040
.020	.961	.877	.700	1.023	.810	2.000	1.203	1.083
.050	.951	.833	.800	1.040	.827	2.500	1.240	1.182
.100	.946	.800	.900	1.056	.846	3.000	1.267	1.271
.200	.951	.775	1.000	1.073	.865	3.500	1.297	1.368

$B^* = 1.50$ $\beta = 1.2547 \times 10^{-1}$ $C = 3.3688 \times 10^{-2}$ $D = -2.0442 \times 10^{-2}$

$E = 2.7440 \times 10^{-3}$

$\sigma(\phi) = 2.16 \times 10^{-3}$ $\sigma(\gamma) = 2.26 \times 10^{-3}$ $VAR(\phi) = 4.686 \times 10^{-6}$ $VAR(\gamma) = 5.129 \times 10^{-6}$

$\sigma^*(\phi) = 1.96 \times 10^{-3}$ $\sigma^*(\gamma) = 2.19 \times 10^{-3}$

TABLE 70. Osmotic and mean activity coefficients of Na valerate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.300	.963	.771	1.200	1.103	.908
.002	.984	.952	.400	.977	.775	1.400	1.128	.948
.005	.977	.929	.500	.992	.785	1.600	1.147	.984
.010	.970	.905	.600	1.008	.797	1.800	1.157	1.012
.020	.962	.877	.700	1.025	.813	2.000	1.158	1.030
.050	.951	.834	.800	1.041	.830	2.500	(1.124)[a]	(1.027)
.100	.947	.801	.900	1.058	.848	3.000	(1.065)	(0.982)
.200	.952	.776	1.000	1.074	.868	3.500	(0.980)	(0.901)

$B^* = 1.50$ $\beta = 1.2902 \times 10^{-1}$ $C = 2.8803 \times 10^{-2}$ $D = -1.3844 \times 10^{-2}$

$E = -1.1394 \times 10^{-3}$

$\sigma(\phi) = 2.69 \times 10^{-3}$ $\sigma(\gamma) = 3.02 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 7.230 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 9.115 \times 10^{-6}$

$\sigma^*(\phi) = 2.52 \times 10^{-3}$ $\sigma^*(\gamma) = 3.73 \times 10^{-3}$

a – values in parenthesis are for the range of complex ionic and molecular equilibria; listed values were obtained from a smooth curve through experiment data for ϕ and subsequent integration for γ.

TABLE 71. Osmotic and mean activity coefficients of Na caproate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.400	.982	.783	1.600	(1.000)	(.830)
.002	.984	.952	.500	.999	.794	1.800	(0.960)	(.799)
.005	.977	.929	.600	1.014	.808	2.000	(0.925)	(.763)
.010	.970	.906	.700	1.029	.822	2.500	(0.827)	(.673)
.020	.962	.877	.800	1.041	.836	3.000	(0.770)	(.612)
.050	.951	.834	.900	1.050	.848	3.500	(0.746)	(.576)
.100	.948	.802	1.000	1.056	.858	4.000	(0.745)	(.556)
.200	.954	.779	1.200	1.054	.865	4.500	(0.761)[a]	(.542)
.300	.967	.777	1.400	(1.033)[a]	(.855)			

$B^* = 1.50$ $\beta = 1.3039 \times 10^{-1}$ $C = 7.1752 \times 10^{-2}$ $D = -6.4283 \times 10^{-2}$

$\sigma(\phi) = 2.49 \times 10^{-3}$ $\sigma(\gamma) = 1.91 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 6.190 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 3.659 \times 10^{-6}$

$\sigma^*(\phi) = 2.43 \times 10^{-3}$ $\sigma^*(\gamma) = 2.31 \times 10^{-3}$

a – values in parenthesis are for the range of complex ionic and molecular equilibria; listed values were obtained from a smooth curve through experiment data for ϕ and subsequent integration for γ.

TABLE 72. Osmotic and mean activity coefficients of Na heptylate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.989	.965	.400	.977	.779	1.600	(.635)	(.468)
.002	.984	.952	.500	.986	.783	1.800	(.596)	(.430)
.005	.977	.929	.600	(.986)[a]	(.781)	2.000	(.562)	(.398)
.010	.970	.906	.700	(.982)	(.775)	2.500	(.505)	(.340)
.020	.962	.877	.800	(.958)	(.754)	3.000	(.492)	(.306)
.050	.952	.835	.900	(.892)	(.700)	3.500	(.495)	(.284)
.100	.948	.803	1.000	(.833)	(.650)	4.000	(.502)	(.267)
.200	.955	.781	1.200	(.740)	(.562)	4.500	(.511)	(.255)
.300	.966	.777	1.400	(.683)	(.512)	5.000	(.523)	(.245)

$B^* = 1.50$ $\beta = 1.3607 \times 10^{-1}$ $C = 8.2551 \times 10^{-2}$ $D = -1.5707 \times 10^{-1}$

$\sigma(\phi) = 9.68 \times 10^{-4}$ $\sigma(\gamma) = 6.16 \times 10^{-4}$ $VAR(\phi) = 9.378 \times 10^{-7}$ $VAR(\gamma) = 3.799 \times 10^{-7}$

$\sigma^*(\phi) = 1.01 \times 10^{-3}$ $\sigma^*(\gamma) = 7.63 \times 10^{-4}$

a – values in parenthesis are for the range of complex ionic and molecular equilibria; listed values were obtained from a smooth curve through experiment data for ϕ and subsequent integration for γ.

TABLE 73. Osmotic and mean activity coefficients of Na caprylate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.400	(.960)[a]	(.758)	.900	(.568)	(.434)	1.800	(.394)	(.253)
.500	(.882)	(.693)	1.000	(.535)	(.401)	2.000	(.386)	(.236)
.600	(.802)	(.621)	1.200	(.486)	(.349)	2.500	(.387)	(.206)
.700	(.722)	(.553)	1.400	(.448)	(.309)	3.000	(.394)	(.185)
.800	(.644)	(.491)	1.600	(.420)	(.279)			

a – values in parenthesis are for the range of complex ionic and molecular equilibria; listed values were obtained from a smooth curve through experiment data for ϕ and subsequent integration for γ.

TABLE 74. Osmotic and mean activity coefficients of Na pelargonate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.300	(.812)[a]	(.550)	.800	(.385)	(.264)	1.600	(.300)	(.152)
.400	(.607)	(.460)	.900	(.362)	(.239)	1.800	(.298)	(.140)
.500	(.521)	(.390)	1.000	(.343)	(.219)	2.000	(.296)	(.130)
.600	(.458)	(.335)	1.200	(.316)	(.189)	2.500	(.289)	(.126)
.700	(.416)	(.295)	1.400	(.304)	(.168)			

a – values in parenthesis are for the range of complex ionic and molecular equilibria; listed values were obtained from a smooth curve through experiment data for ϕ and subsequent integration for γ.

TABLE 75. Osmotic and mean activity coefficients of Na caprate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.400	(.448)[a]	(.344)	.800	(.270)	(.184)	1.400	(.231)	(.107)
.500	(.370)	(.285)	.900	(.251)	(.169)	1.600	(.234)	(.097)
.600	(.326)	(.244)	1.000	(.235)	(.147)	1.800	(.236)	(.089)
.700	(.293)	(.212)	1.200	(.230)	(.120)			

a – values in parenthesis are for the range of complex ionic and molecular equilibria; listed values were obtained from a smooth curve through experiment data for ϕ and subsequent integration for γ.

Acid Salts of Dicarboxylic Acids

Isopiestic measurements were made by Stokes [211] on the sodium and potassium acid salts of malonic, succinic, and adipic acids.

TABLE 76. Osmotic and mean activity coefficients of Na acid malonate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.888	.646	1.600	.858	.525
.002	.984	.951	.500	.883	.626	1.800	.857	.515
.005	.975	.926	.600	.879	.610	2.000	.855	.507
.010	.967	.900	.700	.875	.596	2.500	.854	.490
.020	.956	.867	.800	.872	.584	3.000	.854	.477
.050	.938	.812	.900	.869	.574	3.500	.855	.467
.100	.922	.762	1.000	.867	.565	4.000	.856	.459
.200	.906	.705	1.200	.863	.549	4.500	.857	.451
.300	.896	.671	1.400	.860	.536	5.000	.858	.445

$B^* = 1.25$ $\beta = -2.5969 \times 10^{-2}$ $C = 5.4455 \times 10^{-3}$ $D = -5.1150 \times 10^{-4}$
$E = 1.1165 \times 10^{-5}$

$\sigma(\phi) = 1.82 \times 10^{-3}$ $\sigma(\gamma) = 1.63 \times 10^{-3}$ $VAR(\phi) = 3.314 \times 10^{-6}$ $VAR(\gamma) = 2.659 \times 10^{-6}$
$\sigma^*(\phi) = 2.09 \times 10^{-3}$ $\sigma^*(\gamma) = 2.70 \times 10^{-3}$

TABLE 77. Osmotic and mean activity coefficients of Na acid succinate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.893	.652	1.600	.878	.546
.002	.984	.951	.500	.889	.634	1.800	.880	.539
.005	.976	.926	.600	.885	.619	2.000	.883	.534
.010	.967	.900	.700	.883	.606	2.500	.894	.527
.020	.956	.868	.800	.881	.595	3.000	.907	.524
.050	.939	.813	.900	.879	.586	3.500	.921	.524
.100	.924	.764	1.000	.878	.578	4.000	.933	.525
.200	.908	.709	1.200	.877	.565	4.500	.945	.528
.300	.899	.676	1.400	.877	.554	5.000	.958	.532

$B^* = 1.25$ $\beta = -1.3050 \times 10^{-2}$ $C = -1.2759 \times 10^{-3}$ $D = 4.2800 \times 10^{-3}$
$E = -1.0581 \times 10^{-3}$ $F = 8.0100 \times 10^{-5}$

$\sigma(\phi) = 1.70 \times 10^{-3}$ $\sigma(\gamma) = 1.63 \times 10^{-3}$ $VAR(\phi) = 2.890 \times 10^{-6}$ $VAR(\gamma) = 2.643 \times 10^{-6}$
$\sigma^*(\phi) = 1.87 \times 10^{-3}$ $\sigma^*(\gamma) = 2.91 \times 10^{-3}$

TABLE 78. Osmotic and mean activity coefficients of Na acid adipate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.050	.943	.820	.500	.912	.670
.002	.984	.952	.100	.932	.776	.600	.911	.659
.005	.976	.927	.200	.922	.730	.700	.911	.650
.010	.968	.902	.300	.917	.703			
.020	.958	.871	.400	.914	.684			

$B^* = 1.25$ $\beta = 6.3902 \times 10^{-2}$ $C = -8.0250 \times 10^{-2}$ $D = 4.6890 \times 10^{-2}$

$\sigma(\phi) = 1.60 \times 10^{-4}$ $\sigma(\gamma) = 2.73 \times 10^{-4}$ $\mathrm{VAR}(\phi) = 2.544 \times 10^{-8}$ $\mathrm{VAR}(\gamma) = 7.445 \times 10^{-8}$
$\sigma^*(\phi) = 1.66 \times 10^{-4}$ $\sigma^*(\gamma) = 3.93 \times 10^{-4}$

TABLE 79. Osmotic and mean activity coefficients of K acid malonate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.874	.631	1.600	.808	.474
.002	.984	.951	.500	.864	.607	1.800	.804	.461
.005	.976	.927	.600	.856	.587	2.000	.799	.450
.010	.967	.901	.700	.848	.569	2.500	.792	.427
.020	.956	.868	.800	.842	.554	3.000	.785	.408
.050	.938	.812	.900	.836	.540	3.500	.779	.392
.100	.921	.761	1.000	.831	.528	4.000	.772	.378
.200	.900	.700	1.200	.822	.507	4.500	.764	.365
.300	.885	.661	1.400	.814	.489	5.000	.757	.353

$B^* = 1.50$ $\beta = -9.5652 \times 10^{-2}$ $C = 2.6791 \times 10^{-2}$ $D = -4.4686 \times 10^{-3}$

$E = 2.8346 \times 10^{-4}$

$\sigma(\phi) = 1.87 \times 10^{-3}$ $\sigma(\gamma) = 1.46 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 3.490 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 2.131 \times 10^{-6}$
$\sigma^*(\phi) = 2.19 \times 10^{-3}$ $\sigma^*(\gamma) = 2.52 \times 10^{-3}$

TABLE 80. Osmotic and mean activity coefficients of K acid succinate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.400	.883	.642	1.600	.846	.510
.002	.984	.951	.500	.876	.620	1.800	.845	.501
.005	.976	.926	.600	.871	.603	2.000	.845	.493
.010	.967	.901	.700	.866	.588	2.500	.849	.478
.020	.956	.868	.800	.862	.575	3.000	.856	.469
.050	.939	.813	.900	.858	.563	3.500	.864	.462
.100	.922	.762	1.000	.855	.553	4.000	.870	.457
.200	.904	.705	1.200	.850	.536	4.500	.875	.452
.300	.892	.668	1.400	.847	.522			

$B^* = 1.39$ $\beta = -5.8664 \times 10^{-2}$ $C = 1.6865 \times 10^{-2}$ $D = -1.7901 \times 10^{-3}$
$E = 3.8068 \times 10^{-5}$

$\sigma(\phi) = 9.23 \times 10^{-4}$ $\sigma(\gamma) = 5.64 \times 10^{-4}$ $\mathrm{VAR}(\phi) = 8.516 \times 10^{-7}$ $\mathrm{VAR}(\gamma) = 3.185 \times 10^{-7}$
$\sigma^*(\phi) = 1.06 \times 10^{-3}$ $\sigma^*(\gamma) = 9.61 \times 10^{-4}$

TABLE 81. Osmotic and mean activity coefficients of K acid adipate

m	ϕ	γ	m	ϕ	γ	m	ϕ	γ
.001	.988	.965	.100	.929	.772	.700	.895	.628
.002	.984	.952	.200	.916	.722	.800	.895	.619
.005	.976	.927	.300	.909	.692	.900	.896	.612
.010	.968	.902	.400	.904	.670	1.000	.900	.608
.020	.957	.870	.500	.900	.653			
.050	.942	.817	.600	.897	.640			

$B^* = 1.25$ $\beta = 3.8476 \times 10^{-2}$ $C = -7.1126 \times 10^{-2}$ $D = 4.3579 \times 10^{-2}$

$\sigma(\phi) = 1.25 \times 10^{-3}$ $\sigma(\gamma) = 1.54 \times 10^{-3}$ $\mathrm{VAR}(\phi) = 1.558 \times 10^{-6}$ $\mathrm{VAR}(\gamma) = 2.369 \times 10^{-6}$

$\sigma^*(\phi) = 1.39 \times 10^{-3}$ $\sigma^*(\gamma) = 2.41 \times 10^{-3}$

6.3. Compounds Referenced Only

$NaBO_2$

Isopiestic measurements were made on sodium metaborate by Platford [212] for molalities from 0.1 to 4.0. Values of ϕ and γ show complex behavior as the concentration is varied. Platford estimated that the uncertainties in ϕ are about 0.002 and in γ about 0.02.

$NaBF_4$

Isopiestic measurements on sodium fluoroborate were made by Platford [212] from 0.1 to 9.0 *m*. Values of ϕ and γ show complex behavior as the concentration is varied. Since sodium fluoroborate hydrolyzes slightly to give H_3BO_3 and HF, Platford considered his values of ϕ and γ to be uncertain.

$NaHCO_3$

Han and Bernardin [213] obtained estimates of the *ionic* activity coefficients of aqueous solutions of sodium bicarbonate as a function of ionic strength from ionization constants of carbonic acid, vapor pressure of carbon dioxide in water and in carbonate-bicarbonate solutions and activity coefficients of sodium carbonate. A graph was given showing the values for ionic strengths from 0.2 to 24 at 25 °C; no tabulation of data was given. Values above an ionic strength of about 1.1 are for supersaturated solutions.

AgF

Jahn-Held and Jellinek [214] determined the freezing-point depression for a few concentrations of AgF. Using heat data which they had obtained, they converted their activity coefficients to 25 °C.

TlN_3

Using silver azide and thallium amalgam electrodes, Brouty [215] determined the activity coefficients of thallium azide from 0.001 to 0.0145 molar; concentrations were expressed in moles per liter.

Symmetrical Tetraalkyl Ammonium Halides

Isopiestic measurements on tetramethyl–, tetraethyl–, tetrapropyl- and tetrabutylammonium chlorides, bromides, and iodides were made by Lindenbaum and Boyd [216] from 0.1 to 19.0, 9, 15, 18, 5.5, 12.0, 9.0, 27.0, 1.0, 0.23, 1.9, and 0.50 *m*, respectively. Bower and Robinson [217] also made isopiestic measurements on tetraethylammonium iodide; their measurements extended from 0.1 to 0.7 *m*, whereas those of Lindenbaum and Boyd extended to 1.9 *m*. Levien [218] also made isopiestic measurements on tetramethylammonium bromide and iodide over the range of 0.10 to 1.0 *m* for the former and 0.10 to 0.20 *m* for the latter. Wen, Saito, and Lee [219] made isopiestic measurements on tetramethyl–, tetraethyl–, tetrapropyl–, and tetrabutylammonium fluorides from 0.1 to 7.0, 5.5, 5.0, and 1.6 *m*, respectively.

Quaternary Ammonium Halides

Boyd, Schwarz, and Lindenbaum [220] made isopiestic measurements on triethylammonium (β-hydroxethyl) chloride (or choline chloride) from 0.1 to 6 *m*, on triethylammonium (β-hydroxethyl) bromide (or choline bromide) from 0.1 to 7 *m*, on trimethylbenzylammonium chloride from 0.1 to 8 *m*, triethylbenzylammonium bromide from 0.1 to 12 *m*, and dimethyl-β-hydroxyethylbenzylammonium chloride and bromide from 0.1 to 9.0 *m* and from 0.1 to 13.0 *m*, respectively. Fleming [221] made isopiestic measurements on choline chloride from 0.1 to 4 *m*. Wen and Saito [222] made isopiestic measurements from 0.1 to 6.5 *m* on tetraethanolammonium fluoride and bromide.

Salts of Methane and Ethane Sulfonic Acid

Gregor, Rothenberg, and Fine, [223] made isopiestic measuresments on the lithium, sodium, potassium, ammonium, tetramethylammonium, tetraethylammonium, and tetrabutylammonium salts of methane and ethane sulfonic acids. They reported values only for activity coefficients and a direct check of their isopiestic data cannot be made.

Salts of Benzene Sulfonates

From isopiestic measurements Bonner and Rogers [224] obtained the osmotic and activity coefficients of lithium and sodium benzenesulfonates, lithium and sodium 2, 5-dimethylbenzenesulfonates, lithium and sodium 1, 3, 5-trimethylbenzenesulfonates, and of lithium and sodium *p*-ethylbenzenesulfonate.

7. Acknowledgement

The authors are greatly indebted to Dr. R. L. Nuttall for checking the internal consistency of tables 3 to 72, inclusive, and tables 76 to 81, inclusive, i.e., the relation of the constants given at the bottom of the tables and values for the osmotic and mean activity coefficients listed in the tables; and for the recalculations of data given in the literature [195] for RbF and CsF.

8. References

[1] Lewis, G. N., Proc. Am. Acad. Sci. **13**, 359 (1907).

[2] Debye, P., and Hückel, E., Physik. Z. **24**, 185 (1923).

[3] New Values for the Physical Constants-Recommended by NAS-NRC, NBS Tech. News Bull. **47**, No. 10, 175 (1963); Consistent Set of Physical Constants Proposed, Chem. and Eng. News **41**, No. 46, 43 (Nov. 18, 1963).

[4] Hamer, W. J., Theoretical Mean Activity Coefficients of Strong Electrolytes in Aqueous Solutions from 0 to 100 °C, Nat. Stand. Ref. Data Ser.-Nat. Bur. Stand. (U.S.), 24 table 3, p. 9 (Dec. 1968).

[5] Hückel, E., Physik. Z. **26**, 93 (1925).

[6] Bjerrum, N. Z., Elektrochem. **24**, 259 (1907); Proc. Int. Congr. App. Chem., Sect. X, London, 1909.

[7] Longsworth, L. G., The Principles of Electrochemistry by D. A. MacInnes, p. 160, Reinhold Publishing Corp., New York, N.Y., 1939.

[8] Onsager, L., and Fuoss, R. M., J. Phys. Chem. **36**, 2689 (1932).

[9] Harned, H. S., The Structure of Electrolytic Solutions, edited by W. J. Hamer, Chap. 10, John Wiley & Sons, New York, N.Y., 1959.

[10] Shedlovsky, T., and MacInnes, D. A., J. Am. Chem. Soc., **59**, 503 (1937).

[11] Hornibrook, W. J., Janz, G. J., and Gordon, A. R., J. Am. Chem. Soc. **64**, 513 (1942).

[12] Brown, A. S., and MacInnes, D. A., J. Am. Chem. Soc., **57**, 1356 (1935).

[13] Janz, G. J., and Gordon, A. R., J. Am. Chem. Soc., **65**, 218 (1943).

[14] Scatchard, G., Hamer, W. J., and Wood, S. E., J. Am. Chem. Soc., **60**, 3061 (1938).

[15] Shedlovsky, T., J. Am. Chem. Soc. **72**, 3680 (1950).

[16] MacInnes, D. A., and Parker, K., J. Am. Chem. Soc. **37**, 1445 (1915).

[17] Noyes, A. A., and MacInnes, D. A., J. Am. Chem. Soc. **42**, 239 (1920).

[18] Beattie, J. A., J. Am. Chem. Soc. **42**, 1128 (1920).

[19] Lewis, G. N., and Randall, M., J. Am. Chem. Soc. **43**, 1112 (1921).

[20] Lewis, G. N., and Randall, M., Thermodynamics and the Free Energy of Chemical Substances, Ch. 27, Table 12, p. 362, McGraw-Hill Book, Co., Inc., New York, N.Y., 1923.

[21] Scatchard, G., and Prentiss, S. S., J. Am. Chem. Soc. **55**, 4355 (1933).

[22] Spencer, H. M., J. Am. Chem. Soc. **54**, 4490 (1932).

[23] Robinson, R. A., J. Am. Chem. Soc. **57**, 1161 (1935).

[24] Jones, E. R., and Bury, C. R., Phil. Mag. (7) **3**, 1032 (1927).

[25] Harned, H. S., J. Am. Chem. Soc. **44**, 252 (1922).

[26] Caramazza, R., Gazz. chim. Ital., **90**, 1721 (1960).

[27] Scatchard, G., J. Am. Chem. Soc. **47**, 648 (1925).

[28] Lovelace, B. F., Frazer, J. C. W., and Sease, V. B., **43**, 102 (1921).

[29] Harned, H. S., J. Am. Chem. Soc. **51**, 416 (1929).

[30] Robinson, R. A., and Sinclair, D. A., J. Am. Chem. Soc. **56**, 1830 (1934).

[31] Shibata, Z., and Nira, K., Z. Physik. Chem. (Leipzig), **A173**, 415 (1935).

[32] Harned, H. S., and Cook, M. A., J. Am. Chem. Soc. **59**, 1290 (1937).

[33] Robinson, R. A., Trans. Far. Soc. **37**, 302 (1941).

[34] Robinson, R. A., and Stokes, R. H., Trans. Faraday Soc. **45**, 612 (1949).

[35] Robinson, R. A., and Harned, H. S., Chem. Rev. **28**, 419 (1941).

[36] Robinson, R. A., Proc. Roy. Soc. New Zealand **75**, 203 (1945).

[37] Harned, H. S., and Owen, B. B., The Physical Chemistry of Electrolytic Solutions, 3rd ed., Appendix A, Table (12–3–1A), p. 731, Reinhold Book Corp., New York, N.Y., 1958.

[38] Robinson, R. A., and Stokes, R. H., Electrolyte Solutions, 2nd ed., Appendix 8.3, Table 1, p. 476 and Appendix 8.10, Table 11, p. 494, Butterworths Scientific Publications, London, 1959.

[39] Pitzer, K. S., and Brewer, L., Thermodynamics-Gilbert Newton Lewis and Merle Randall, 2nd ed., Appendix 4, Table A4-2a, p. 643, McGraw-Hill Book Co., Inc., New York, N.Y., 1961.

[40] Lietzke, M. H., and Stoughton, R. W., J. Phys. Chem. **66**, 508 (1962).

[41] Randall, M., Int. Crit. Tables **7**, 224 (1930).

[42] Guggenheim, E. A., Thermodynamics, An Advanced Treatment for Chemists and Physicists, 3rd ed., p. 358, North-Holland Publishing Co., Amsterdam (1957).

[43] Harned, H. S., and Robinson, R. A., The International Encyclopedia of Physical Chemistry and Chemical Physics, Topic 15, Vol. 2, Multicomponent Electrolyte Solutions, p. 5, Pergamon Press, 1968.

[44] Broene, H. H., and De Vries, T., J. Am. Chem. Soc. **69**, 1644 (1947).

[45] Anthony, J. D. C., and Hudleston, L. J., J. Chem. Soc. **127**, 1122 (1925).

[46] Kendall, J., Booge, J. E., and Andrews, J. C., J. Am. Chem. Soc. **39**, 2303 (1917).

[47] Paterno, E., and Peratoner, A., Atti, Reale Acca. Livcei, CCLXXXVII, 306 (1890).

[48] Parker, V. B., Thermal Properties of Aqueous Uni-Univalent Electrolytes, Nat. Stand. Ref. Data Ser.-Nat. Bur. Stand. (U.S.), 2, p. 3 and fig. VIII (April 1, 1965).

[49] Fredenhagen, K., and Wellman, M., Z. Physik. Chem. **A162**, 454 (1932).

[50] Brosheer, J. C., Lenfesty, F. A., and Elmore, K. L., Ind. Engr. Chem. **39**, 423 (1947).

[51] General Electric Company, Tech. Service Bull., 30A, 13 (1945).

[52] Vdovenko, V. M., Lazarev, L. N., and Shirvinskaya, Radiokhim. **7**, 46 (1965).

[53] Elmore, K. L., Hatfield, J. D., Mason, C. M., and Jones, A. D., J. Am. Chem. Soc. **71**, 2710 (1949).

[54] Deussen, E., Z. Anorg. Chem. **44**, 303 (1905).

[55] Wooster, C. B., J. Am. Chem. Soc. **60**, 1609 (1938).

[56] Kielland, J., J. Am. Chem. Soc. **59**, 1675 (1937).

[57] Hamer, W. J., and Wu, Yung-Chi, J. Res. Nat. Bur. Stand. A. Phys. and Chem. **74A**, 761 (1970).

[58] Hamer, W. J., and De Wane, H. J., Electrolytic Conductance and the Conductances of the Halogen Acids in Water, Nat. Stand. Ref. Data Ser.-Nat. Bur. Standards (U.S.), 33, p. 23 (May 1970).
[59] Dolezalek, F., Z. Physik. Chem. **26**, 321 (1898).
[60] Jahn, H., Z. Physik. Chem. **33**, 545 (1900).
[61] Gahl, R., Z. Physik. Chem. **33**, 178 (1900).
[62] Ellis, J. H., J. Am. Chem. Soc. **38**, 737 (1916).
[63] Linhart, G. A., J. Am. Chem. Soc. **39**, 2601 (1917).
[64] Getman, F. H., J. Am. Chem. Soc. **42**, 1556 (1920).
[65] Duboux, M., and Rochat, J., Helv. Chim. Acta **22**, 161 (1939).
[66] Hills, G. J., and Ives, D. J. G., J. Chem. Soc. (1951), 305, 311, 318.
[67] Grzybowski, A. K., J. Phys. Chem. **52**, 550 (1958).
[68] Schwabe, K., and Ziegenbalg, B., Z. Elektrochem. **62**, 172 (1958).
[69] Lewis, G. N., Brighton, T. B., and Sebastian, R. L., J. Am. Chem. Soc. **39**, 2245 (1917).
[70] Gupta, S. R., Hills, G. J., and Ives, D. J. G., Trans. Faraday Soc. **59**, 1874 (1963).
[71] Noyes, A. A., and Ellis, J. H., J. Am. Chem. Soc. **39**, 2532 (1917).
[72] Linhart, G. A., J. Am. Chem. Soc. **41**, 1175 (1919).
[73] Scatchard, G., J. Am. Chem. Soc. **45**, 641, 696, 2098 (1925).
[74] Nonhebel, G., Phil. Mag. (7) **2**, 1085 (1926)
[75] Güntelberg, E., Z. Physik. Chem. **123**, 199 (1926).
[76] Roberts, E. J., J. Am. Chem. Soc. **52**, 3877 (1930).
[77] Carmody, W. R., J. Am. Chem. Soc. **54**, 188 (1932).
[78] Harned, H. S., and Ehlers, R. W., J. Am. Chem. Soc. **54**, 1350 (1932).
[79] Harned, H. S., and Ehlers, R. W., J. Am. Chem. Soc. **55**, 2179 (1933).
[80] Anderson, N. J., Dissertation, University of Chicago, 1934.
[81] Åkerlöf, G., and Teare, J. W., J. Am. Chem. Soc. **59**, 1855 (1937).
[82] Bates, R. G., and Bower, V. E., J. Res. Nat. Bur. Standards **53**, 283 (1954).
[83] Zielen, A. J., J. Phys. Chem. **67**, 1474 (1963).
[84] Wilke, E., and Schränkler, W., Z. Physik. Chem. **A162**, 361 (1932).
[85] Shedlovsky, T., and MacInnes, D. A., J. Am. Chem. Soc. **58**, 1970 (1936).
[86] Giber, J., Lengyel, S., Tamas, J., and Tahi, P., Magy. Kem. Folyoirat **66**, 170 (1960).
[87] Faita, G., and Mussini, T., J. Chem. & Eng. Data **9**, 332 (1964).
[88] Covington, A. K., and Prue, J. E., J. Chem. Soc. (1955), 3696.
[89] Lohonyai, N., and Proszt, J., Magy. Kem. Folyoirat **66**, 423 (1960).
[90] Harpst, J. A., Holt, E., and Lyons, P. A., J. Phys. Chem. **69**, 2333 (1965).
[91] Bates, S. J., and Kirschman, H. D., J. Am. Chem. Soc. **41**, 1991 (1919).
[92] Dobson, H. J. E., and Masson, I., J. Chem. Soc. **125**, 668 (1924).
[93] Dunn, J. S., and Rideal, E. K., J. Chem. Soc. **125**, 676 (1924).
[94] Haase, R., Naas, H., and Thumm, H., Z. Physik. Chem. Nene Folge (Frankfurt) **37**, 210 (1963).
[95] Ionin, M. V., and Kurina, N. V., Tr. Po. Khim & Khim Tech. No. 1, 40 (1964).
[96] Lewis, G. N., and Storch, H. H., J. Am. Chem. Soc. **39**, 2544 (1917).
[97] Livingston, R. S., J. Am. Chem. Soc. **48**, 45 (1926).
[98] Keston, A. S., J. Am. Chem. Soc. **57**, 1671 (1935).
[99] Harned, H. S., Keston, A. S., and Donelson, J. G., J. Am. Chem. Soc. **58**, 989 (1936).
[100] Bierman, W. J., and Yamaski, R. S., J. Am. Chem. Soc. **77**, 241 (1955).
[101] Bienert, K., Diplomarbeit, Aachen, 1961.
[102] Hetzer, H. B., Robinson, R. A., and Bates, R. G., J. Phys. Chem. **66**, 1423 (1962).
[103] Faita, G., Mussini, T., and Ozzioni, R., J. Chem. & Eng. Data **11**, 162 (1966).
[104] Gupta, S. R., Hills, G. J., and Ives, D. J. G., Trans. Faraday Soc. **59**, 1886 (1963).
[105] Mussini, T., and Pazzoli, S., Ann. Chim. (Rome) **53**, 569 (1963).
[106] Noyes, A. A., and Freed, E. S., J. Am. Chem. Soc. **42**, 476 (1920).
[107] Pearce, J. N., and Fortsch, A. R., J. Am. Chem. Soc. **45**, 2852 (1923).
[108] Küppers, H. A., Diplomarbeit, Aachen, 1960.
[109] Hetzer, H. B., Robinson, R. A., and Bates, R. G., J. Phys. Chem. **68**, 1929 (1964).
[110] Kortum, G., and Hausserman, W., Ber. des Bunsenois **69**, 594 (1965).
[111] Harned, H. S., and Robinson, R. A., Trans. Faraday Soc. **37**, 302 (1941).
[112] Covington, A. K., and Prue, J. E., J. Chem. Soc. (1957), 1567.
[113] Pearce, J. N., and Nelson, A. F., J. Am. Chem. Soc. **55**, 3080 (1933).
[114] Robinson, R. A., and Baker, O. J., Trans. & Proc. Royal Soc. N.Z., **76**, 250 (1946).
[115] Haase, R., Dücker, K. H., and Küppers, H. A., Ber. Bun. Physik. Chem. **69**, 97 (1965).
[116] Stonehill, H. J., J. Chem. Soc. (1943), 647.
[117] Flatt, R., and Bengnerel, F., Helv. Chim. Acta. **45**, 1765 (1962).
[118] Davis, W., Jr., and De Bruin, H. J., J. Inorg, & Nuc. Chem. **26**, 1069 (1964).
[119] Pearce, J. N., and Mortimer, F. S., J. Am. Chem. Soc. **40**, 509 (1918).
[120] MacInnes, D. A., and Beattie, J. A., J. Am. Chem. Soc. **42**, 1117 (1920).
[121] Lengyel, S., and Giber, J., Acta Chim. Acad. Sci. Hung. **32**, 235 (1962).
[122] Lengyel, S., Giber, J., and Tamas, J., Magy Kem Folyoirat, **66**, 161 (1960).
[123] Caramazza, R., Ann. Chim. (Rome) **53**, 472 (1963).
[124] Lebed, V. I., and Aleksandrov, V. V., Russian J. Phys. Chem. **38**, 1414 (1964).
[125] Pearce, J. N., and Nelson, A. F., J. Am. Chem. Soc. **54**, 3544 (1932).
[126] Kangro, W., and Groeneveld, A., Z. Physik. Chem. Nene Folge (Frankfurt) **32**, 110 (1962).
[127] Robinson, R. A., Trans. Faraday Soc. **41**, 756 (1945).
[128] Davis, D. S., Chem. & Process Eng. **39**, 175 (1958).
[129] Robinson, R. A., and McCoach, H. J., J. Am. Chem. Soc. **69**, 2244 (1947).
[130] Sarnowski, M., and Baranowski, B., Z. Physik. Chem. **228**, 310 (1965).
[131] Harned, H. S., and Swindells, F. E., J. Am. Chem. Soc. **48**, 128 (1926).
[132] Harned, H. S., and Shropshire, J. A., J. Am. Chem. Soc. **80**, 2968 (1958).
[133] Jones, J. H., J. Phys. Chem. **51**, 516 (1947).
[134] Robinson, R. A., J. Am. Chem. Soc. **57**, 1167 (1935).
[135] Robinson, R. A., J. Am. Chem. Soc. **68**, 2403 (1946).
[136] Ivett, R. W., and De Vries, T., J. Am. Chem. Soc. **63**, 2821 (1941).
[137] Robinson, R. A., J. Am. Chem. Soc. **63**, 628 (1941).
[138] Allmand, A. J., and Polack, W. G., J. Chem. Soc. **115**, 1020 (1919).
[139] Harned, H. S., and Nims, L. F., J. Am. Chem. Soc. **54**, 423 (1932).
[140] Brown, A. S., and MacInnes, D. A., J. Am. Chem. Soc. **57**, 1356 (1935).
[141] Sakong, G., and Huang, J. E., Ta Hau Hua Hsuch Hui-chik **6**, 113 (1962).
[142] Caramazza, R., Gazz. Chim. Ital. **90**, 1839 (1960).

[143] Getman, F. H., J. Am. Chem. Soc. **42**, 1556 (1920).
[144] Haas, K., and Jellinek, K., Z. Physik. Chem. **162**, 153 (1932).
[145] Stokes, R. H., J. Am. Chem. Soc. **69**, 1291 (1947).
[146] Negus, S. S., Thesis, Johns Hopkins University, 1922.
[147] Olynyk, P., and Gordon, A. R., J. Am. Chem. Soc. **65**, 224 (1943).
[148] Petit, M. C., J. Chim. Phys. **62**, 1119 (1965).
[149] Hepburn, J. R. I., Proc. Phys. Soc. **40**, 256 (1928).
[150] Robinson, R. A., Trans. Faraday Soc. **35**, 1217 (1939).
[151] Janis, A. A., and Ferguson, J. B., Canadian J. Res. **17B**, 215 (1939).
[152] Harned, H. S., and Douglas, S. M., J. Am. Chem. Soc. **48**, 3095 (1926).
[153] Murata, F., J. Chem. Soc. Japan **56**, 588 (1935).
[154] Harned, H. S., and Crawford, C. C., J. Am. Chem. Soc. **59**, 1903 (1937).
[155] Pearce, J. N., Taylor, M. D., and Bartlett, R. M., J. Am. Chem. Soc. **50**, 2051 (1928).
[156] Ionin, M. V., and Chezzanova, D. N., Tr. Po. Khim & Khim. Tech No. 2, 186 (1964).
[157] Pencier, J., and Marcus, Y., J. Chem. & Eng. Data **10**, 105 (1965).
[158] Makarov, C. L., Vlasov, Y. G., and Azarko, V. A., Russian J. Phys. Chem. **40**, 609 (1966).
[159] Miller, M. L., and Sheridan, L. L., J. Phys. Chem. **60**, 184 (1956).
[160] Harned, H. S., J. Am. Chem. Soc. **47**, 676 (1925).
[161] Harned, H. S., Z. Physik. Chem. **117**, 1 (1925).
[162] Ferguson, A. L., and Schlucter, A. W., Trans. Am. Electrochem. Soc. **52**, 369 (1927).
[163] Harned, H. S., and Hecker, J. C., J. Am. Chem. Soc. **55**, 4841 (1933).
[164] Kobayski, Y., and Wang, Hsiu-ying, J. Sci. Hiroshima Univ. **5A**, 71 (1934).
[165] Stokes, R. H., J. Am. Chem. Soc. **67**, 1690 (1945).
[166] Jones, J. H., J. Am. Chem. Soc. **65**, 1353 (1942).
[167] Rush, R. M., and Johnson, J. S., J. Phys. Chem. **72**, 767 (1968).
[168] Jones, J. H., and Froning, H. R., J. Am. Chem. Soc. **66**, 1673 (1944).
[169] Pearce, J. N., and Hopson, H., J. Phys. Chem. **41**, 536 (1937).
[170] Harned, H. S., and Shropshire, J. A., J. Am. Chem. Soc. **80**, 2618 (1958).
[171] Stokes, J. M., Trans. Faraday Soc. **41**, 686 (1945).
[172] Scatchard, G., and Breckenridge, R. C., J. Phys. Chem. **58**, 596 (1954).
[173] Robinson, R. A., J. Am. Chem. Soc. **62**, 3131 (1940).
[174] Tamas, J., and Kosza, G., Magy. Kem. Folyoirat **70**, 148 (1964).
[175] Harned, H. S., J. Am. Chem. Soc. **38**, 1986 (1916).
[176] Bird, E. W., and Hixon, R. M., J. Phys. Chem. **34**, 1412 (1930).
[177] Parton, H. N., Trans. Faraday Soc. **33**, 617 (1937).
[178] Hill, A. V., Proc. Roy. Soc. **A127**, 9 (1930).
[179] Weir, A. R., Coll. Czechoslov. Chem. Comm. **8**, 149 (1936).
[180] Brown, O. L. I., and Delaney, C. M., J. Phys. Chem. **58**, 255 (1954).
[181] Pearce, J. N., and Hart, H. B., J. Am. Chem. Soc. **43**, 2483 (1921).
[182] Harned, H. S., J. Am. Chem. Soc. **48**, 3095 (1926).
[183] Murata, F., J. Chem. Soc. Japan **54**, 379 (1933).
[184] MacWilliam, E. A., and Gordon, A. R., J. Am. Chem. Soc. **65**, 984 (1943).
[185] Scatchard, G., and Orthing, W. H., J. Colloid and Interface Sci. **22**, 12 (1966).
[186] Gelbach, R. W., J. Am. Chem. Soc. **55**, 4857 (1933).
[187] Torto, F. G., J. Chem. Soc. (1957), 3594.
[188] Robinson, R. A., and Wilson, J. M., Trans. Faraday Soc. **36**, 738 (1940).
[189] Chow, M., J. Am. Chem. Soc. **43**, 488 (1920).
[190] Knobel, M., J. Am. Chem. Soc. **45**, 70 (1923).
[191] Harned, H. S., and Cook, M. A., J. Am. Chem. Soc. **59**, 496 (1937).
[192] Åkerlöf, G., and Bender, P., J. Am. Chem. Soc. **70**, 2366 (1948).
[193] Jones, J. H., J. Am. Chem. Soc. **69**, 2066 (1947).
[194] Robinson, R. A., Stokes, J. M., and Stokes, R. H., J. Phys. Chem. **65**, 542 (1961).
[195] Tien, H. Ti, J. Phys. Chem. **67**, 532 (1963).
[196] Robinson, R. A., J. Am. Chem. Soc. **59**, 84 (1937).
[197] Makarov, L. L., Evstropev, K. K., and Vlasov, Y. G., Zh. Fiz. Khim. **32**, 1618 (1958).
[198] Harned, H. S., and Schupp, O. E., Jr., J. Am. Chem. Soc. **52**, 3886 (1930).
[199] Caramazza, R., Ann. Chim. (Rome) **53**, 481 (1963).
[200] Harned, H. S., and Schupp, O. E., Jr., J. Am. Chem. Soc. **52**, 3890 (1930).
[201] MacInnes, D. A., and Brown, A. S., Chem Rev. **18**, 335 (1936).
[202] Robinson, R. A., and Tait, D. A., Trans. Faraday Soc. **37**, 570 (1941).
[203] Cowptherwaite, I. A., La Mer, V. K., and Barksdale, J., J. Am. Chem. Soc. **56**, 544 (1934).
[204] Nardelli, M., Braibanti, A., and Chierici, I., Gazz. Chim. Ital. **87**, 510 (1957).
[205] Pearce, J. N., and Pumplin, G. G., J. Am. Chem. Soc. **59**, 1219 (1937).
[206] Wishaw, B. F., and Stokes, R. H., Trans. Faraday Soc. **49**, 27 (1953).
[207] Shul'ts, M. M., Makarov, L. L., and Su Yu-ĵeng, Russ. J. Phys. Chem. **36**, 1181 (1962).
[208] Esval, O. E., and Tyree, S. Y., Jr., J. Phys. Chem. **66**, 940 (1962).
[209] Wishaw, B. F., and Stokes, R. H., Trans. Faraday Soc. **49**, 30 (1953).
[210] Smith, E. R. B., and Robinson, R. A., Trans. Faraday Soc. **38**, 70 (1942).
[211] Stokes, J. M., J. Am. Chem. Soc. **70**, 1944 (1948).
[212] Platford, R. F., Canadian J. Chem. **47**, 2271 (1968).
[213] Han, S. T., and Bernardin, L. J., Tappi **41**, 540 (1958).
[214] Jahn-Held, W., and Jellinek, K., Z. Elektrochem. **42**, 608 (1936).
[215] Brouty, M^{me} Marie-Louise, Compt. Rend. **214**, 258 (1942).
[216] Lindenbaum, S., and Boyd, G. E., J. Phys. Chem **68**, 911 (1964).
[217] Bower, V. E., and Robinson, R. A., Trans. Faraday Soc. **59**, 1717 (1963).
[218] Levien, B. J., Australian J. Chem. **18**, 1161 (1965).
[219] Wen, W., Saito, S., and Lee, C., J. Phys. Chem. **70**, 1244 (1966).
[220] Boyd, G. E., Schwarz, A., and Lindenbaum, S., J. Phys. Chem. **70**, 821 (1966).
[221] Fleming, R., J. Chem. Soc. (1961), 3100.
[222] Wen, W., and Saito, S., J. Phys. Chem. **69**, 3569 (1965).
[223] Gregor, H. P., Rothenberg, M., and Fine, N., J. Phys. Chem. **67**, 1110 (1963).
[224] Bonner, O. D., and Rogers, O. C., J. Phys. Chem. **64**, 1499 (1960).