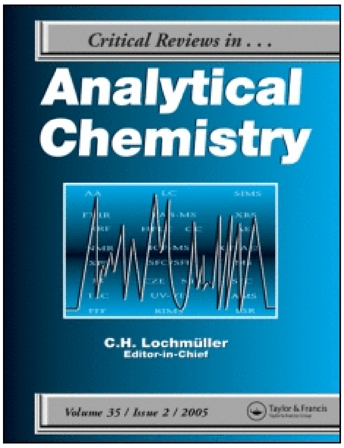

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THE EFFECTS OF SALTS ON TITRATIONS

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I. INTRODUCTION

As long ago as 1899, Arrhenius,[1] on the basis of catalytic experiments, concluded that addition of a neutral salt, such as sodium chloride, to a solution of an uncharged weak acid caused an increase in its dissociation. This is now understood to be the result of decreased activity coefficients of the ions.[2] In 1916, Kolthoff[3] determined that "indicators with an alkaline character are given an acid tint by neutral salts, while the indicators with acid character are given an alkaline tint." Concentrated salts are known to increase the acidity of mineral acids, and their effects were quantitatively studied in the early 1950's by several investigators, using measurements of various acidity functions. It was not until 1958 that Critchfield and Johnson[4] took analytical advantage of this increased acidity to titrate very weak bases in aqueous solution. Although relatively few studies have been made since then on the effects of salts in titrations, a number of specific studies have demonstrated enhanced sharpness of potentiometric end points for particular titrations. The mechanisms of such enhancements appear to vary markedly, depending on the solvent, the sample, and the titrant.

The aims of this review are to summarize the titrations that have been studied and to discuss the various theories and mechanisms that have been proposed for the salt effects observed. Titrations of weak bases and weak acids in both aqueous and nonaqueous solution are described, and some effects of salts on oxidation-reduction reactions are summarized, with possible implications for redox titrations.

II. ACID-BASE TITRATIONS

A. Measures of Acidity

This review will concentrate on the influence of salts on titration curves and on the interpretation of such influence. In order to interpret the results of acid-base titrations, it is necessary to make acidity measurements. Hence, we will briefly summarize some of the salient features of acidity measurements in both aqueous and nonaqueous media, and their applications to concentrated salt solutions. The reader is referred to more comprehensive reviews of acidity functions for further details.

There are a number of ways in which acidity of a solution can be estimated. Three measurements that sometimes can conveniently be made are pMH, pH_{GE}, and H_o.[5,6] pMH is equal to -log [strong acid], where the acid is completely dissociated and its concentration is equal to the molar concentration of the solvated proton. This is obviously limited to strong acids. pH_{GE} is the pH measured using a cell with a glass indicator electrode and a saturated calomel reference electrode. The determination and interpretation of pH_{GE} are frequently uncertain, due to the liquid-junction potential, and are of little use in comparing different solvents. Baumann[7] has described the measurement of pH in concentrated salt solutions when using glass electrodes. Salted buffers of known empirical pH are used to calibrate the glass electrodes. In contrast to techniques involving calibration with unsalted buffers, the pH values measured for a fixed buffer composition are constant when using different glass electrode-reference electrode combinations. The method was demonstrated by pH determinations in alkaline 4*M* sodium nitrate and in acidic 5.4*M* lithium nitrate solutions.

H_o is the acidity function originally described by Hammett and Deyrup in 1932.[8] Rosenthal and Dwyer[5] have determined that in acid solution $pMH\text{-}pH_{GE}$ and $pMH\text{-}H_o$ (and therefore $pH_{GE}\text{-}H_o$) are constants for a given salt solution and are independent of the value of pMH.

The acidity function is probably the most useful quantitative measure of acidity under a wide variety of conditions. It has been reviewed in detail by Paul and Long,[6] and more recently in the book by Rochester,[6a] and readers are referred to these references for a complete discussion of its applications and limitations. The acidity function is defined by

$$H_o = pk_{BH^+} - \log \frac{C_{BH^+}}{C_B} \qquad (1)$$

where B is one of a series of uncharged indicator bases, usually a substituted aniline; these are called the H_o indicators. K_{BH^+} is the thermodynamic dissociation constant for the conjugate acid of the base in terms of molar concentrations, referred to ideal dilute solution in water:

$$BH^+ = B + H^+$$

C_{BH^+} and C_B are directly observable from the absorbance of the two different colored forms of the indicator. H_0 is formally equivalent to

$$H_0 = -\log \frac{a_{H^+} f_B}{f_{BH^+}} = -\log[H^+] f_{H^+} f_B / f_{BH^+} \qquad (2)$$

where a_{H^+} is the hydrogen ion activity and the f's are the molar-concentration activity coefficients. H_0 is specifically a measure of the tendency for the solution in question to transfer a proton to an uncharged base molecule. Increasingly negative values correspond to higher acidities.

Analogous acidity functions can be defined for the ionization equilibria of positively or negatively charged bases:

$$BH^{++} = B^+ + H^+;\ H_+ \equiv pk_{BH^{++}} - \log \frac{C_{BH^{++}}}{C_{B^+}} \qquad (3)$$

$$BH = B^- + H^+;\ H_- \equiv pk_{BH} - \log \frac{C_{BH}}{C_{B^-}} \qquad (4)$$

and so forth.

At infinite dilution, the activity coefficients in Equation 2 and in the analogous equations for Equations 3 and 4 approach unity, and a_{H^+} becomes equal to C_{H^+}. Under these conditions, each of the above acidity functions becomes equal to pa_H ($-\log a_{H^+}$). However, with increasing acid concentration, in nonaqueous media, or in concentrated salt solutions, the various indicator acidity functions deviate from one another. No theory exists for estimating the extent of deviation from ideality of solutes at either high electrolyte concentrations or in nonaqueous media, but indicator concentration ratios can be measured directly and the acidity function applied on an empirical basis.

Differences in the indicator acidity functions arise due to differences in charge type of the indicators, not the least effect being the differences in their activity coefficients. Most investigators using indicator acidities in high concentrations of electrolytes measure H_0. This is because it has been empirically established that a series of different neutral base indicators gives reasonably uniform values of H_0 up to highly acid solutions, whereas such systematic investigations are lacking for most indicators of other charge types.

A few systematic studies have been made on the influence of concentrated neutral salts on acidity functions. Harbottle,[9] in a study of perchloric acid-sodium perchlorate solutions, noted a striking decrease in H_0 upon adding increasing amounts of sodium perchlorate; it decreased from −0.22 to −1.46 upon adding 5*M* salt to 1*M* perchloric acid. A decrease is expected from Equation 1 because the added salt should "salt out" the indicator base; i.e., f_B in Equation 2 should increase. Depending on the agent, there is a linear increase or decrease in log f_B for *p*-nitroaniline with increasing salt concentrations.[6] In the order $NaClO_4 < Me_4NBr < Et_4NBr <$ sodium *p*-toluenesulfonate, these salts cause it to decrease, whereas the alkali halide salts, in the order $KCl < NaBr < NaCl < LiCl$, cause it to increase; the sodium perchlorate effect is in the opposite direction to what one would predict from the H_0 measurements (see below).

In addition to the effect on f_B, the ratio f_{H^+}/f_{BH^+} in Equation 2 should increase with increased ionic strength. Lucas and Steigman[10] measured the activity coefficients of *p*-nitroanilinium chloride in dilute acid solutions containing chloride salts. It was highest in lithium chloride solutions and decreased in the order $NaCl > KCl > CsCl > Me_4NCl > Et_4NCl$ over the concentration range of about 1 to 5*m*. In bromide solutions, the order was $NaBr > KBr > Et_4NBr$. 2,4-Dichloranilinium chloride and diphenylammonium chloride exhibited different behavior.

Paul[11] and Long and McIntyre[12] studied the effects of salts on H_0 in hydrochloric acid solutions. Figure 1 shows the dependence of H_0 on the various salts they reported. The figure shows that the salt concentration effect is specific for the particular salt and is linear up to high concentrations of salt. The apparent acidity is increased by salts in the order $LiCl > NaCl, NaBr > NaClO_4 > KCl$, whereas it is decreased by sodium *p*-toluenesulfonate, Me_4NBr, , and Et_4NBr. The effects follow much the same order as the changes in log f_B, except for $NaClO_4$ and sodium *p*-toluenesulfonate (Figure 2). Both of these latter salts have relatively large anions and show a larger relative tendency to salt in the neutral base than would be suggested by their effects on H_0. (Perchlorate salts also have a greater effect in increasing the acidity of the proton, relative to the behavior of halide salts – see Section II.C.1 and Reference 29.)

The salt effects on H_0 are dependent on the indicator base.[11] The difference is reportedly due largely to an observed difference in the log f_B term. Such specific effects on different indicators are sometimes not observed with strong mineral

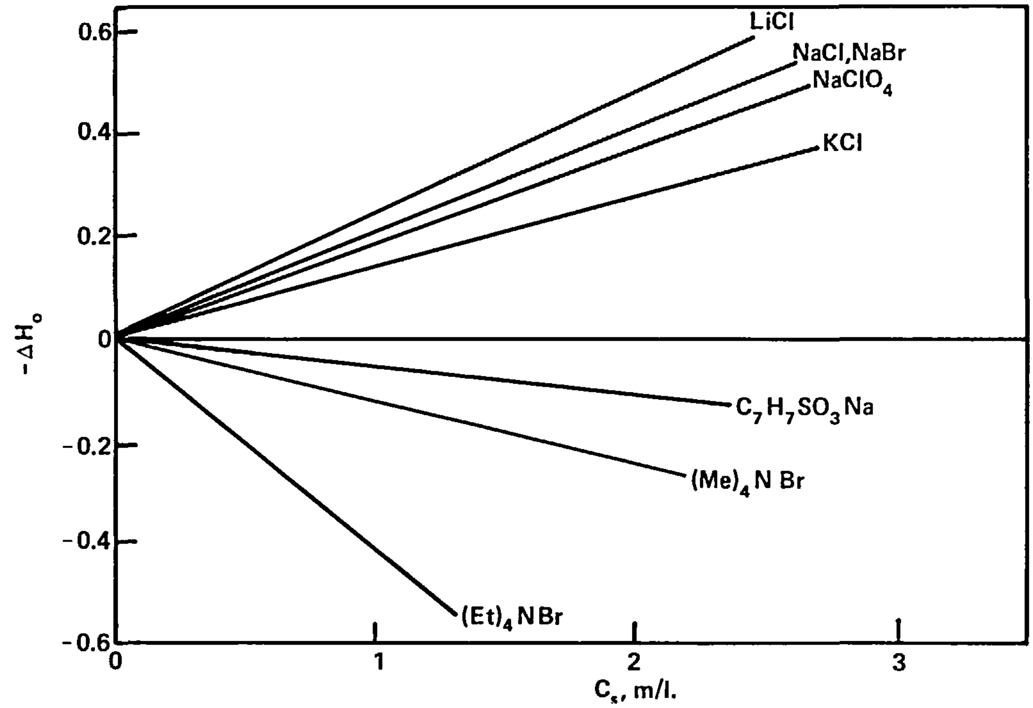

FIGURE 1. Salt effects on H_0 of 1.0 *M* hydrochloric acid. The indicator is *p*-nitroaniline. (From Paul, M. A. and Long, F. A., *Chem. Rev.*, 57, 1 (1957). With permission.)

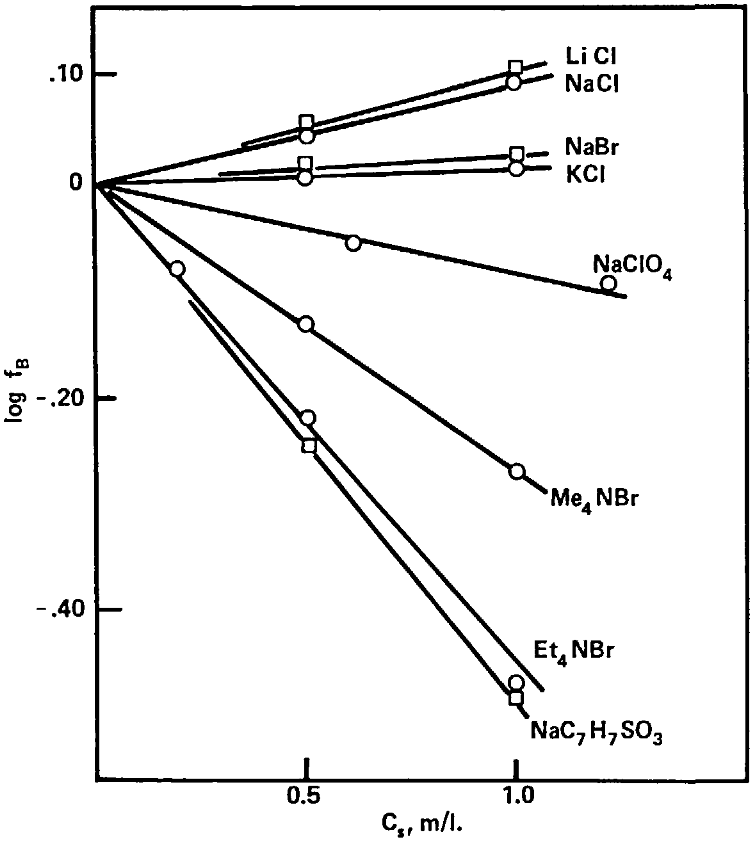

FIGURE 2. Salt effects on the activity coefficient of *p*-nitroaniline. (From Paul, M. A. and Long, F. A., *Chem. Rev.*, 57, 1 (1957). With permission.)

acids, possibly because their "salt" effects on log f_B are small.[12,13]

Ojeda and Wyatt,[14] in a more eloquent treatment of acidity functions in salt solutions, applied a correction of the log f_B term to H_0 by calculating the activity coefficient of the basic form of the indicators from solubility measurements in the salt solutions. In this way, the salting-out effect is eliminated, leaving a rather larger effect, which is interpreted in terms of the dehydration of the hydrogen ion as the water activity is reduced. Details of results are given below under the discussion of acidities in salt solution.

Long and McIntyre,[12] in a study of acid-catalyzed hydrolysis rates of methylal in the presence of salts, found poor correlation between measured H_0 values for the salt solutions and the hydrolysis rates, even though with strong acids a good correlation of rate and H_0 had previously been obtained. They did obtain direct correlation *if they added a log* f_M/f_B *term to the* $-\Delta H_0$ value, where f_M is the activity coefficient of the methylal. They interpreted this to mean that the rate does indeed depend on H_0 but that a correction must be made for the different effects of salts on methylal and the neutral base used to determine H_0.

Moiseev and Flid[15,16] have reported results that are in fair agreement with the studies by Paul.[11] They used the empirical equation*

$$y = pk_{BH^+} + \log C_{H^+} + L\mu \qquad (5)$$

where y is defined as $\log(C_{BH^+}/C_B)$, μ is the ionic strength due to the added salt, and L is a constant derived from the dependence of y on μ. From the definition of H_0, L should be the same as the slope (α) of $-H_0$ vs. [salt] for a 1:1 electrolyte:[6]

$$H_0 = -\log C_{H^+} - L\mu \qquad (6)$$

Moiseev and Flid[15,16] obtained a value for L of 0.10 for the addition of potassium chloride, calcium chloride, or magnesium chloride to hydrochloric acid solutions, using *p*-nitroaniline indicator. This compares with a value of α of 0.14 reported by Paul[11] for potassium chloride added to hydrochloric acid, using the same indicator. Note, however, that Paul[11] obtained quite different values of α for different salts containing widely differing cations and anions.

In similar studies with sulfuric acid, Moiseev and Flid[15,16] found that most sulfates decreased y, probably because of a lowered acidity due to the reaction $H^+ + SO_4^= = HSO_4^-$. Mercuric sulfate, on the other hand, caused a large increase due to formation of $Hg(SO_4)_2^=$.

Values for H_0 can be determined for solutions in nonaqueous and in mixed solvents, but it is important to establish the generality of the H_0 function for any particular solvent system. Braude and Stern[17] have made measurements of H_0 for hydrochloric acid in ethanol, acetone, and dioxane containing varying proportions of water. The constancy of $H_0 + \log C_H{}^+$ in dilute solutions suggests that hydrochloric acid behaves as a strong acid in ethanol and acetone. In dioxane it appears to behave as a weak acid, probably due to the low dielectric constant of the solvent. The effect of water on H_0 is similar for each solvent. Figure 3 illustrates the effect for 0.1*M* hydrochloric acid. A maximum is obtained in each case, indicating a minimum in the acidity. Addition of the organic solvent to an aqueous solution of the acid probably tends to decrease the f_B of the base, i.e., to "salt-in" the base and increase H_0[6] (Equations 1 and 2). Further addition increases $f_H{}^+$ as H_3O^+ gives way to the stronger acid SH^+, the conjugate of the organic solvent S.

The above assumptions, however, are complicated by the expectation that hydrochloric acid would not be completely dissociated in acetone or ethanol, except at very low concentrations, and that a significant fraction of the protonated indicator may be present as ion pairs.

Rochester[6a] has pointed out that the acidity function approach is generally as satisfactory for increasing concentrations of strong acids in a *particular* mixed solvent composition as for solutions of strong acids in water. However, in some cases this is not true, particularly for solvents of low dielectric constant, where ion association becomes significant. He also concludes that the acidity function concept is less useful when applied to fixed concentrations of strong acids in a series of solutions with *varying* solvent composition.

More detailed discussions and comparisons of

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*See footnote in Reference 6, page 20, concerning an error in the *Chemical Abstracts* statement of this equation.

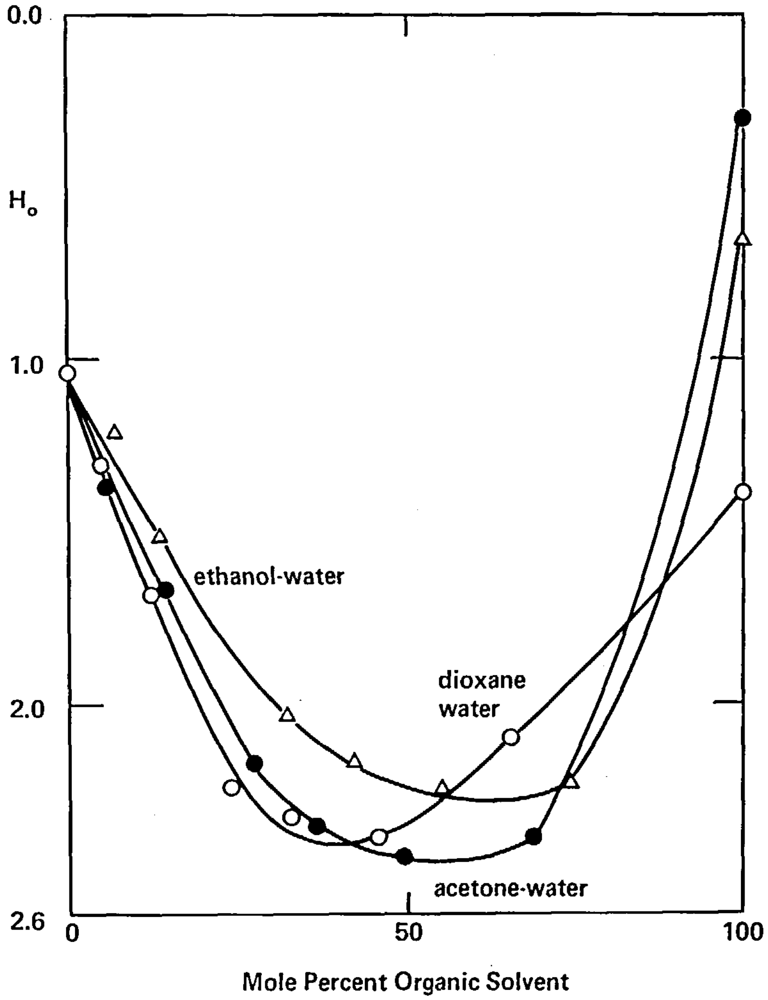

FIGURE 3. H_0 in mixed solvents containing 0.1 *M* hydrochloric acid. (From Paul, M. A. and Long, F. A., *Chem. Rev.*, 57, 1 (1957). With permission.)

different acidity functions are available elsewhere.[6,6a,18,19]

The papers by Bunnett and Olsen,[19a] Hinman and Lang,[19b] and Arnett and Mach[19c] also discuss the need for different acidity functions for different kinds of bases and indicate that the acidity function H_0 is not as generally applicable as was thought at one time.

B. Acidities in Salt Solutions

The success of certain titrations in concentrated neutral salt solutions is the result, partially at least, of enhanced acidity or activity of the proton.[4] Hence, a summary of acidities determined for concentrated salt solutions along with the quantitative correlation of the enhancing effects with salt concentrations is of interest.

In the 1920s, Harned and co-workers[20,21] investigated the effects of concentrated salt solutions on the activity coefficients of acids and bases. Representative results are shown in Figure 4. In acid solution, chloride salts at concentrations greater than about 1*M* cause the activity coefficient of the acid to increase, and at higher concentrations it becomes greater than unity. The order of increase is $LiCl > NaCl > KCl > CaCl_2 > SrCl_2 > BaCl_2$, similar to the order of enhancement of the activity coefficients of indicator bases and their salts. In base solution, however, the activity coefficients do not increase appreciably at high concentrations. The same is true of sulfuric acid solutions in the presence of sulfates, again probably due to suppressed ionization of the sulfuric acid.

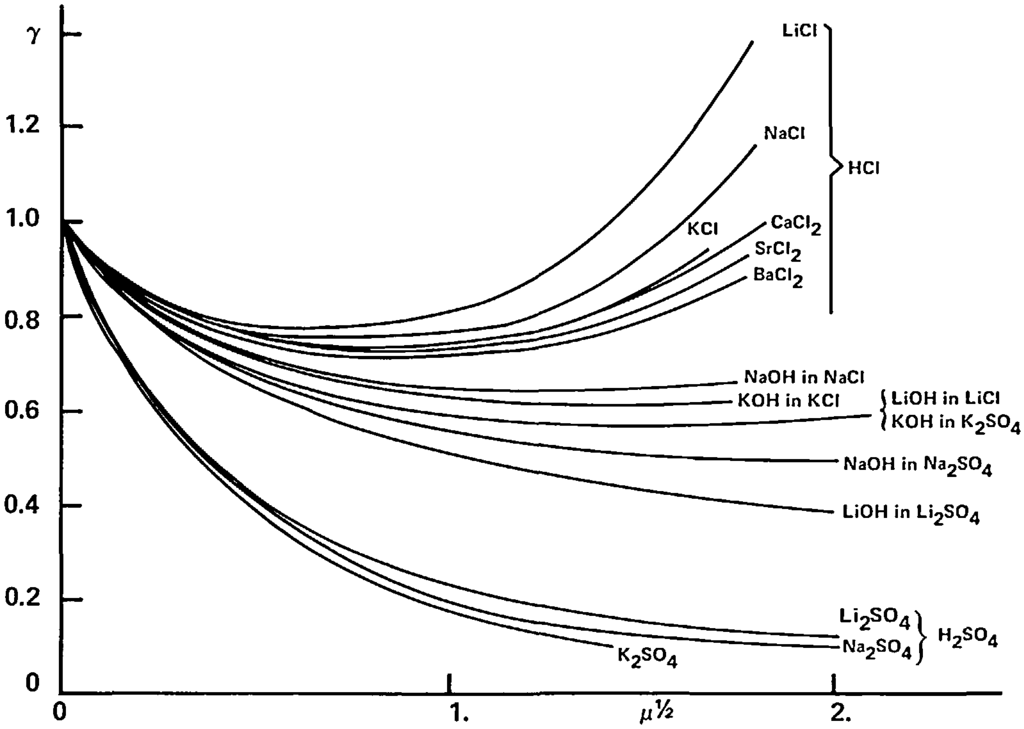

FIGURE 4. Activity coefficients of acids and hydroxides in salt solutions. (From Harned, H. S., *Trans. Faraday Soc.*, 23, 462 (1927). With permission.)

Hydration of the salt ions has been presented as a plausible explanation of the decreasing activity coefficients followed by the rapid increase. In a solution of increasing salt concentration, hydration has a double effect. The first is to increase the mole fraction of the salt by removing the water molecules. The second is to dehydrate the ions of the acid under consideration as the water is used up in hydrating the salt ions. But the effect is undoubtedly much more complex than simple hydration.[2] Other factors probably include electrostatic Debye-Huckel effects, ionic association, change in dielectric constant, and so forth, acting in combination.[2]

Harned[22] studied the dissociation of water in the presence of salts and verified the expected increase in dissociation.

In 1953, Kilpatrick and co-workers[23-26] measured the dissociation constants of several weak acids in aqueous and nonaqueous salt solutions. Most measurements were made with the cell

Au	HA 0.005*M* MB 0.005*M* MCl X*M* Quinhydrone	MCl 1.00*M*	HCl 0.005*M* MCl 0.005*M* Quinhydrone	Au

The measured emf gives the ratio of the hydrogen ion concentrations in the two half-cells, and, assuming that the hydrogen ion concentration in the right half-cell is equal to the stoichiometric concentration of the strong acid, the dissociation constant is calculated from the stoichiometry of the buffer solution

$$K_c = C_{H_3O^+} \cdot \frac{C_{A^-}}{C_{HA}}$$

From a knowledge of the thermodynamic dissociation constant, $K_a = K_c \frac{f_{H^+} f_{A^-}}{f_{HA}}$, and of the activity coefficient of the molecular weak acid, HA, in the salt solutions, the mean activity coefficient of the ions of weak acid, $\sqrt{f_{H^+} f_{A^-}}$, can be calculated. Table 1 lists representative dissociation constants for benzoic acid (K_a = 6.32 X 10^{-5}) in aqueous salt solutions along with the change in activity coefficients with changing salt

TABLE 1

The Dissociation Constant of Benzoic Acid in Aqueous Salt Solutions at 25°C, $K_0 \times 10^5$

	Electrolyte		
Mol/l	KCl	NaCl	LiCl
0.05	9.10	9.12	9.16
.10	10.08	10.05	11.08
.20	10.43	10.81	11.47
.30	10.83	11.38	11.86
.40	11.18	11.79	12.93
.50	11.32	11.90	12.95
.60	11.38	11.85	13.00
.70	11.12	11.80	12.95
.80	11.06	11.72	12.94
.90	11.10	11.55	12.90
1.00	10.79	11.36	12.88
1.50	9.94	10.48	12.60
2.00	8.84	9.20	11.25
2.50	7.58	7.58	9.46
3.00	6.52	6.49	8.99

From Kilpatrick, M., *J. Am. Chem. Soc.*, 75, 584 (1953). With permission.

concentrations. The dissociation constant increases in the order KCl < NaCl < LiCl, due to the difference in the activity coefficients of molecular benzoic acid as well as in the mean activity coefficient of the ions. The differences are more marked at the higher salt concentrations. Note that at high salt concentrations the increase in the dissociation constant is reversed and a maximum is reached at somewhat less than 1*M*. These studies have been extended for acetic acid and formic acid to 8*M* lithium chloride.[36,44]

The dependence of K_c on added salts is much more marked in nonaqueous solvents, and K_c changes markedly in going from one solvent to another. For example, in ethanol K_c changes from 0.87 to 13.8 to 39.0 × 10^{-10} in going from 0 to 0.005 to 0.10*M* LiCl. The K_c/K_a ratio for 0.10*M* LiCl increases from 1.76 to 9.00 to 44.8 for the solvents water, methanol, and ethanol.

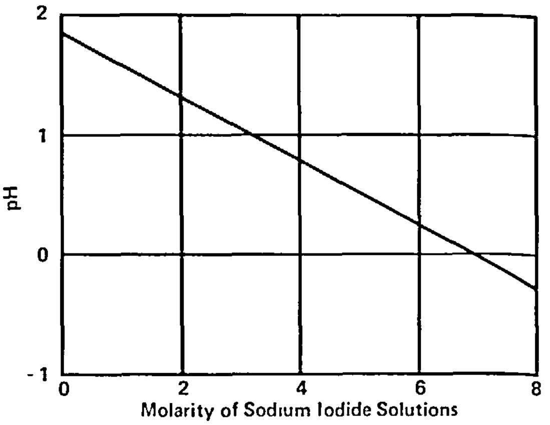

FIGURE 5. Effect of sodium iodide concentration on pH of 0.0192 *M* hydrochloric acid. (From Critchfield, F. E. and Johnson, J. B., *Anal. Chem.*, 30, 1247 (1958). With permission.)

Similar studies have been made for acetic acid,[24] glycolic acid,[25] and cyclohexanecarboxylic acid,[26] except that activity coefficients of the molecular acids were not available. Results were similar for the dependence of K_c on the salt concentration and the salt type.

The first application of concentrated salts in acid-base titrations was by Critchfield and Johnson.[4] They titrated weak bases in aqueous salt solutions with hydrochloric acid titrant. In this study, they demonstrated that the acidity of the titrant was enhanced at least two orders of magnitude in the presence of 8*M* sodium iodide. Figure 5 illustrates that a linear decrease in pH was found with increasing molar concentration of the salt.

Although Paul[11] had shown that acidity functions in concentrated salt solutions decrease linearly with increasing neutral salt, the magnitude of the effect was explained on the basis of the tendency to salt-out the indicator rather than a real decrease in the acidity of the solution (see above, Figure 2 and discussion of Equation 2). The equation developed by Moiseev and Flid[15] (Equation 6), on the other hand, indicates that the effect is independent of the nature of the salt because L is a constant. Critchfield and Johnson, therefore, undertook a study of the effects of neutral salts on the pH and acidity of aqueous solutions.[27] They observed a direct correlation between the heat of solution of the salt and the measured pH of an acid solution, shown in Figure 6. All the salts are at 4I concentration, where I is

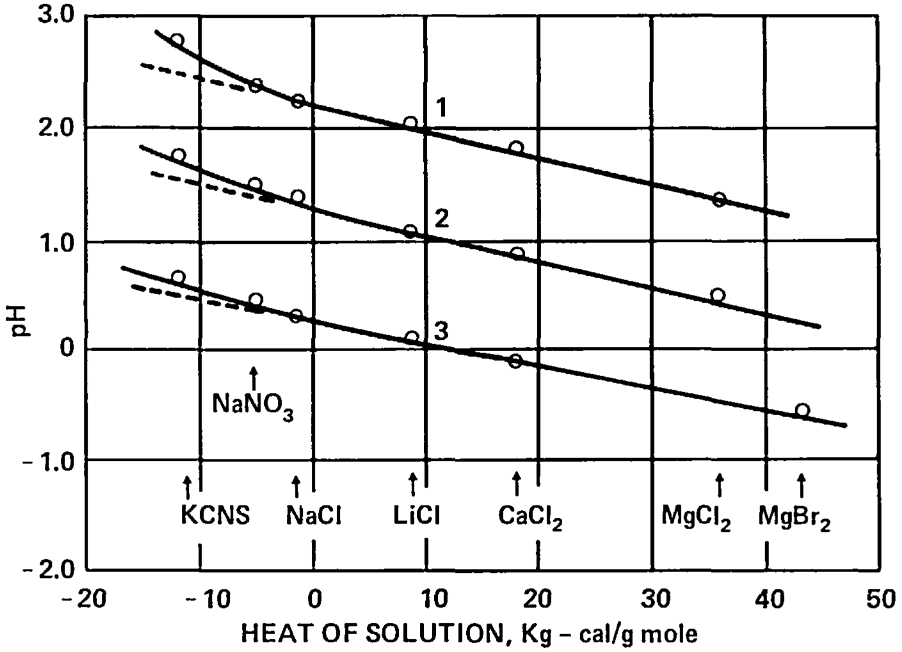

FIGURE 6. Correlation between pH and heat of solution of neutral salts. 1. 0.1 *M* acetic acid $K_a = 1.75 \times 10^{-5}$. 2. 0.1 *M* chloroacetic acid $K_a = 1.4 \times 10^{-3}$. 3. 0.1 *M* hydrochloric acid. Each solution is 4I in salt. (From Critchfield, F. E. and Johnson, J. B., *Anal. Chem.*, 31, 570 (1959). With permission.)

defined as $M(\frac{n}{2})$; this is referred to as the ionic *concentration. M* is the molar concentration of the salt, and n is the number of ions present in the salt (e.g., 3 for $CaCl_2$). I is similar to ionic strength, but it ignores the charge of the ions. A linear correlation was not found if solutions of equal ionic strength were used.

For hydrochloric acid, the following empirical equation was derived from the straight line portion of the curve in Figure 6.*

$$pH = -\log C_{H^+} - I(0.18 + 6 \times 10^{-3}\ \Delta H_s) \qquad (7)$$

where ΔH_s is the heat of solution of the salt in kilocalories per gram-mole. The curves for acetic acid and chloroacetic acid follow the same equation, except that $-\log C_{H^+}$ is replaced by $-\log(K_aM)^{1/2}$. Salts with negative heats of solution tended to deviate from these straight line correlations, particularly for the weak acids.

Critchfield and Johnson[27] compared measured pH values and acidity function measurements (using *p*-nitroaniline) for 0.1*M* hydrochloric acid solutions in 4I concentrations of neutral salts. Their results are summarized in Table 2. The pH_{GE} and H_0 values generally agree within about 0.15 unit. This is similar to the difference of 0.12 unit between the pH and H_0 of 0.1*M* hydrochloric acid in water.[6] As already mentioned, pH and H_0 are identical at infinite dilution. The largest differences occur in solutions of salts with negative heats of solution. The relationship between H_0 and the heat of solution is similar to that between pH and heat of solution. The equation defining the linear dependence is

$$H_0 = -\log C_{H^+} - I\,(0.20 + 5 \times 10^{-3}\ \Delta H_s) \qquad (8)$$

This is identical to Equation 7 for pH, except that the constants 0.18 and 6×10^{-3} are changed slightly (due to the slight difference between pH and H_0). Equation 8 can be compared with Equation 6, derived by Moiseev and Flid.[15] They are similar in form, but the latter predicts that the salt effect is independent of the nature of the salt because L is a constant; by contrast, ΔH_s is dependent on the salt. No explanation has been offered for the independence reported by Moiseev and Flid.

A plot of measured pH, calculated pH, and H_0

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*In the original paper the minus sign in front of the log term is omitted.

TABLE 2

pH and Acidity Function (H_0) of Hydrochloric Acid Solutions Containing Neutral Salts

Salt[a]	pH	H_0[b]	H_0[c]	Heat of solution Kcal/g-Mol
Sodium chloride	0.32	0.18	0.17	-1.28
Potassium chloride	0.58	0.42	0.40	-4.44
Lithium chloride	0.12	0.03	0.01	8.51
Potassium thiocyanate	0.65	0.49	–	-12.20
Sodium nitrate	0.45	0.27	–	-5.02
Calcium chloride	-0.08	-0.13	–	17.99
Magnesium chloride	-0.31	-0.31	–	36.00
Magnesium bromide	-0.65	-0.52	–	43.25

[a] Each solution 4I in salt and 0.1*M* in hydrochloric acid.
[b] H_0 values obtained with *p*-nitroaniline.
[c] Calculated from data of Paul.[11]

From Critchfield, F. E. and Johnson, J. B., *Anal. Chem.*, 31, 570 (1959). With permission.

for 0.1*M* hydrochloric acid as a function of the ionic concentration of calcium chloride gives a linear dependence up to 5I, similar to Figure 5 for sodium iodide, and the maximum deviation between any of the quantities is 0.1 unit. Similarly, a plot of measured pH, calculated pH, and H_0 as a function of hydrochloric acid concentration in solutions containing 4I lithium chloride results in a maximum deviation of 0.1 unit between the quantities up to 0.2*M* acid. Hence, pH and H_0 are comparable in magnitude in neutral salt solutions over a wide range of acid and salt concentrations and do not differ by much more than would be expected in water alone. Hence, both pH and H_0 are an approximate quantitative measure of the acidity of these solutions. This is supported by the fact that the liquid-junction potential at the calomel electrode in the salt solution is small. Critchfield and Johnson,[27] using the Planck equation,[28] calculated the liquid-junction potential for 3.75*N* sodium chloride to be 4.4 mV, corresponding to an error of only 0.07 pH unit.

In water-methanol mixtures, however, the pH and H_0 of 0.1*M* hydrochloric acid and 4I lithium chloride differ widely, as shown in Figure 7. Although the pH is unaffected by the methanol up to 70 or 80%, the acidity function increases to a maximum, both in the absence (curve 1) and in the presence (curve 3) of salt. This is similar to Figure 2, in which added organic solvent tends to "salt in" the indicator base and increase H_0, whereas at higher organic solvent concentrations SH^+ replaces H_3O^+ to increase the acidity. Critchfield and Johnson[27] suggested that the effect of the methanol should be greater in the presence of added salts, because the salts themselves have an initial salting-out effect, i.e., the methanol should decrease the salting-out tendency. They interpret the similarity of the two H_0 curves (1 and 3) to mean that the increase in H_0 by the addition of methanol is due to some other effect. It appears that the salting-in effect of the organic solvent proposed by Braude and Stern[17] may be that other effect, and the salting-out effect of the salts may tend to counterbalance this. (The net effect actually appears to be greater in the absence of the salt, but different indicators were used, so direct comparison cannot be made.) In the salt solution, it appears that pH and H_0 may merge at high methanol concentrations. Above 70 to 80% methanol, the medium can no longer be

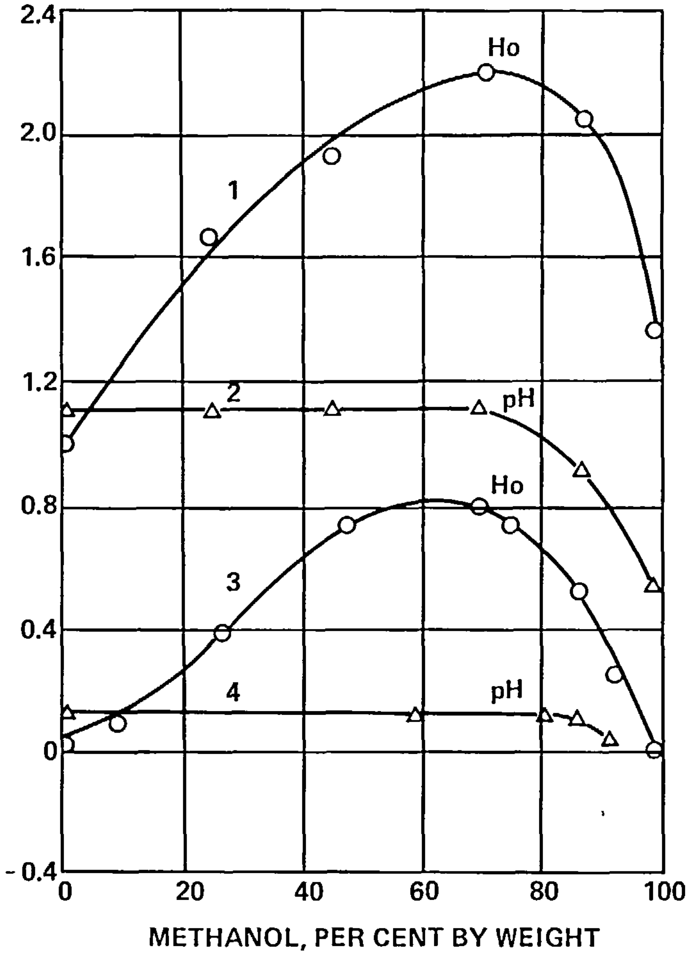


FIGURE 7. pH and acidity function (H_0) of methanol-water solutions containing 0.1 *M* hydrochloric acid. Curves 1 and 2 – Solutions contain no salt. H_0 measurements using *m*-nitroaniline. Curves 3 and 4 – Solutions contain 41 lithium chloride: H_0 measurements using *p*-nitroaniline. (From Critchfield, F. E. and Johnson, J. B., *Anal. Chem.*, 31, 570 (1959). With permission.)

considered aqueous and pH is not a quantitative measure of acidity.

Because the increase in acidity of aqueous solutions in the presence of salts was obtained with acids that are completely ionized, it must be associated with an increase in the activity of the proton. Critchfield and Johnson[27] concluded that the hydrogen ion is probably made more active by losing water of hydration. This theory is supported by the correlation between pH and the heat of solution of the neutral salts, since the salts with the largest heat of solution would have a greater affinity for water and would more successfully compete with the proton for water of hydration. This implies also that the proton is not completely dehydrated, at least by the salts studied by Critchfield and Johnson.[27] Perchlorate salts cause a greater increase in acidity than do the salts shown in Figure 6.[29] Matkovich and Christian[30] have obtained indirect evidence that in about three-fourths-saturated aqueous calcium chloride solution (ca. 4.2*m*) all the water is tied up in solvating the calcium chloride; this serves as a basis for separation of acetone from water. At higher salt concentrations the solubility of acetone in water increases, as the former solvent is now used in solvating the extra calcium chloride. At 4.2*m* calcium chloride, thirteen molecules of water are associated with each molecule of the salt. This agrees fairly well with the hydration number of 12.0 for calcium chloride reported by Stokes and Robinson.[31] If all the water can indeed be removed by solvation, then either the proton becomes completely desolvated or else the

TABLE 3

Measurements of Dilute HCl Solutions in 4 *M* LiCl at 25°C

HCl concentration, mol/l	Indicator	H_o	pH_{GE} glass electrode	pH_H hydrogen electrode	$pMH-H_o = \log Q_{BH^+}$	$pMH-pH_{GE}$
4.98×10^{-3}	pNA[a]	1.237	1.36	1.35	1.066	0.94
7.97×10^{-3}	pNA	1.044	1.13	1.14	1.055	0.96
1.00×10^{-2}	pNA	0.970	1.06	1.07	1.030	0.94
2.99×10^{-2}	pNA	0.497	0.55		1.027	0.97
7.97×10^{-2}	oNA	0.007	0.15		1.092	0.94
					Average = 1.054	0.95
					SD_x = 0.012	0.006_3
					90% level = ±0.026	±0.013

[a] pNA is *p*-nitroaniline; oNA is *o*-nitroaniline.

From Rosenthal, D. and Dwyer, J. S., *J. Phys. Chem.*, 66, 2687 (1962). With permission.

competition equilibria are such that some water remains with the proton; the latter is more likely.

A simple experimental demonstration of the increased activity of the hydrogen ion in the presence of salt can be obtained by observing the reaction of metallic calcium in water and in 6*N* sodium chloride,[27] or of zinc granules in 0.1*N* hydrochloric acid in the absence and presence of saturated lithium iodide.[32] At the same hydrogen ion concentration, the metals liberate hydrogen much more vigorously in the salt solutions, in agreement with the 100-fold increase in acidity shown in Figure 5.

Rosenthal and Dwyer[33] measured pH (with glass electrode – pH_{GE} – and with hydrogen electrode – pH_H) and H_o values of dilute solutions of hydrochloric acid in 4*M* and 8*M* lithium chloride, and compared these with pMH (-log total strong acid concentration). Table 3 summarizes their data for 4*M* lithium chloride. Measurements were also made in 4*M* calcium chloride, 6*M* sodium nitrate, and 6*M* sodium perchlorate. In each case, the $pMH - pH_{GE}$ value was constant in a particular salt solution, independent of pMH. Likewise, $pMH - H_o$ was essentially constant in a given salt solution over a wide range of hydrochloric acid concentrations. Hence, $pH_{GE} - H_o$ is also essentially constant. The difference is small, 0.1–0.2 unit, as reported by Critchfield and Johnson.[27]

Measurements with a hydrogen electrode were in good agreement with glass electrode measurements, similar to results obtained by Critchfield for 0.025*M* hydrochloric acid in alkali and alkaline earth halide solutions and by Bruckenstein and Kolthoff for acetic acid and acetic acid-sodium acetate solutions in up to 5*M* calcium chloride.[34] These results demonstrate the meaningfulness of potentiometric measurements with the glass electrode.[35]

Rosenthal and Dwyer[33] described the quantitative aspects of the acidity indicator equilibrium in these solutions by the equation

$$K_{BH^+} = ([B]\,[\text{Total strong acid concentration}]/[BH^+])Q_{BH} \qquad (9)$$

where $Q_{BH^+} = f_B f_{H^+}/f_{BH^+}$ and depends on the nature and concentration of the salt and on the nature of the uncharged base, B. K_{BH^+} is the thermodynamic dissociation constant as defined before. The acidity function is, hence, given by

$$H_o = pK_{BH^+} - \log[BH^+]/[B] = pMH - \log Q_{BH^+} \qquad (10)$$

The log Q_{BH^+} values are given in Table 3. In 8*M* lithium chloride it is 2.492, using *o*-nitroaniline or 4-chloro-2-nitroaniline.

Dwyer and Rosenthal[36] performed similar acid-base equilibria studies for negatively charged carboxylate bases (acetate, formate) in dilute acid solutions containing concentrated lithium chloride. The results could be explained by the equation

$$K_{BH} = ([B^-][\text{Total strong acid}]/[BH])Q_{BH} \qquad (11)$$

which is identical to Equation 9 except that BH refers to an uncharged acid. Again, the various factors which are important in concentrated salt solutions can be satisfactorily incorporated into the Q_{BH} term. Experimental values of Q_{BH} for acetic acid in up to 8*M* lithium chloride solutions are reported in the paper, and they agree with values reported by others in more dilute lithium chloride solutions.[24,37,38] It was demonstrated that the Q_{BH} values could be empirically calculated using the equation

$$Q_{BH} = -[\sqrt{M}/(1 + A\sqrt{M})] + BM \qquad (12)$$

where M is the molarity of the salt and A and B are constants. Proudlock and Rosenthal[39] determined the activity coefficients of aniline derivatives and naphthalene (nonelectrolytes) in 0–10*M* lithium chloride. In several instances, small but significant deviations from the equation log f = BM were observed, where B is a constant for a particular base in a particular salt and M is the molarity of the salt. The equation holds up to certain salt concentrations and then deviates, presumably due to electrostatic or chemical interactions between the nonelectrolyte ions and water. One possible chemical interaction is variable hydration of the nonelectrolyte so that one or more equilibria of the type $B \cdot qH_2O \rightleftharpoons B + qH_2O$ needs to be considered.

Ojeda and Wyatt[14] corrected H_0 values for the log f_B term (see Equation 2) in order to eliminate the salting-out effect on the indicator base and calculated corrected H_0^* values (i.e., $H_0 + \log f_B$). Table 4 compares H_0 and H_0^* values for different acid concentrations in various concentrations of sodium bromide, calcium chloride, and magnesium chloride. H_0^* data are also listed in the original paper for lithium chloride (4 - 8*M*) and sodium chloride (1 - 4*M*) solutions.

Ojeda and Wyatt[14] compared the behavior of the corrected H_0^* values with that to be expected from a simple hydration model previously proposed[40] for concentrated acid solutions. In order to simplify their model they assumed that only the proton hydration need be considered in calculating the correct concentrations for the equilibrium $B + H_3O^+ \rightleftharpoons BH^+ + H_2O$. The experimental ratio C_{BH^+}/C_B will then be proportional to $X_{H_3O^+}\, f_B/a_{H_2O}$, where $X_{H_3O^+}$ represents the mole fraction of H_3O^+ in the solution, assuming the activity coefficients of the free species H_3O^+ and BH^+ cancel (this assumption is not justified, however, for charged bases). This leads to the following equation for H_0^*:

$$H_0^* = H_0 + \log f_B = -k - \log (X_{H_3O^+}/a_{H_2O}) \qquad (13)$$

The constant, k, can be regarded as the pK for the formation of the free species H_3O^+ from water, as suggested by Deno and Taft.[41] Deno and Taft argue for a value of 6.78 (corrected to the new Paul and Long scale[6]), but Wyatt gives reasons for believing it is actually nearer a value of 5.[14,40] Equation 13 predicts that in a plot of H_0^* for fixed acid concentration against the water activity of the salt solutions, the points for the different salts should be on a common curve. This was shown to be the case[14] and Ojeda and Wyatt therefore concluded that, apart from the log f_B effect which has already been allowed for, the principal reason why neutral salts increase the acidity in aqueous solutions is that they simply reduce the water activity and thus tend to dehydrate the H_3O^+ ion. This would tend to corroborate the postulation of Critchfield and Johnson.[27]

In developing a physical model for proton hydration, Ojeda and Wyatt[14] regarded the H_3O^+ ion as being able to take on four water molecules in stages. (See below for a discussion of the hydration number of the proton.) They then derived the mole fraction of H_3O^+ in terms of the total acid concentration, C_{H^+}, the water activity, the molar concentration of the salt, C, and the number of ions per molecule of the salt, ν:

$$X_{H_3O^+} = C_{H^+}(1-a)/(\nu C + 2C_{H^+}) \times (1 + 20a + 150a^2 + 500a^3 + 625a^4) \qquad (14)$$

The coefficients of a in the last parenthetical term represent K_1, K_1K_2, $K_1K_2K_3$, and $K_1K_2K_3K_4$, respectively, for the successive equilibrium constants of the hydration of H_3O^+. Substitution into Equation 13 gives

$$H_o^* + k = -\log\left\{\frac{C_{H^+}(1 - a)}{a(\nu C + 2C_{H^+})(1 + 20a + 150a^2 + 500a^3 + 625a^4)}\right\} \tag{15}$$

By inserting experimental values of H_0^* and a in this equation, a mean value of 5.02 was calculated for k from 40 measurements. This compares with 5.12 obtained for concentrated acid solutions,[40] indicating that the same model accommodates the effects of at least some of the salts. Calculated H_o^* values, adopting k = 5.02, are compared with measured values in Table 4. They generally agree to within about 0.1 unit.

Ojeda and Wyatt do note some anomalies in their model, particularly among those salts that "salt-in" the indicator. After allowing for salting-out and salting-in, the residual effect on H_o is much too large in the case of sodium *p*-toluene sulfonate (Table 5, Paul and Long[6]) to be explained in terms of their model. A more perplexing, though smaller, anomaly occurs with sodium perchlorate, which has a greater effect on H_o than predicted from Equation 15; this is puzzling since perchloric acid fits in well with the other acids.[40] Note also in Figure 2 that log f_B in sodium perchlorate solutions changes in an order opposite that of other salts that cause an increase in acidity.[11]

Ojeda, Pérez, and Wyatt[42] used their hydration model to derive an explicit expression for Q_{BH} for uncharged bases:

$$Q_{BH} = \left(\frac{1-a}{a}\right)\left(\frac{55.5}{\nu C}\right)\left(\frac{1296}{1 + 20a + 150a^2 + 500a^3 + 625a^4}\right) f_B \tag{16}$$

Log Q_{BH} values calculated directly with this equation using experimentally determined f_B values were in substantial agreement with those determined by Rosenthal and Dwyer.[5]

Saxton[43] has performed similar studies with a number of other weak bases and acids in 1–10*M* lithium bromide, and the H_o and log Q functions are given. On the basis of results, a possible interaction between nitro compounds and bromide solutions is suggested. Empirical equations for calculating log Q are fitted to the data. Rosenthal et al.[44] extended studies to −2, −1, 0, +1, and +2 charged bases in dilute acid solution in 1–8*M* lithium chloride. In all cases, the quantitative aspects of the acid-base equilibria were consistent with an equation similar to Equation 9. Values of Q_{BH} are reported.

Acidity functions in concentrated acid solutions increase markedly, much as in concentrated salt solutions, and it is of interest to compare them. Bascombe and Bell[45] used the acidity function to obtain information about the interaction in concentrated aqueous acid solutions. They point out that H_0 and h_0 deviate from pH and $[H^+]$ at high acid concentrations. For example, h_0 has a value of about 1,000 in a 10-molal solution of any mineral acid, corresponding to a value of about 100 for the activity coefficient ratio in Equation 2.

In attempting to account for this behavior, Bascombe and Bell[45] note that the salting-out effect[11] will certainly increase f_B in Equation 2, but that the observed changes in h_o are much too large to be accounted for in this way. They interpret this to mean that the increase in h_o with concentration is attributable primarily to an increase in f_{H^+}/f_{BH^+} and that this is due to the hydration of the hydrogen ion. This in turn is related to the decrease in the water activity. If the hydrogen ion is associated with *h* water molecules to form species H_h^+ then the equilibrium determining the acidity function should be written as

$$B + H_h^+ \rightleftharpoons BH^+ + hH_2O$$

and (neglecting any hydration of BH^+):

$$-H_o = \log [H_h^+] - h \log a_{H_2O} + \log (f_B f_{BH^+}/f_{BH^+}) \tag{17}$$

The activity of water is significantly less than unity in concentrated acid solutions and hence Equation 17 can account for a sharp rise in acidity with moderate values of h. Since H_h^+ is now a large

TABLE 4[a]

Salt	M	m[b]	a_{H_2O}	HCl, 0.0102 M			HCl, 0.0203 M			HCl, 0.0507 M			HCl, 0.1005 M		
				H_0 exptl[c]	H_0* exptl	H_0* calcd	H_0 exptl	H_0* exptl	H_0* calcd	H_0* exptl	H_0* exptl	H_0* calcd	H_0 exptl	H_0* exptl	H_0* calcd
				2.09			1.66			1.28			0.98		
NaBr[d]	2	2.11	0.9243	1.47	1.55	1.66	1.33	1.31	1.37	0.84	0.92	0.98	0.60	0.68	0.68
	4	4.49	0.8155	0.91	1.07	1.34	0.77	0.93	1.05	0.37	0.53	0.65	0.21	0.37	0.36
$CaCl_2$[e]	1	1.03	0.9431	1.61	1.72	1.70	1.29	1.40	1.40	0.88	1.00	1.01	0.57	0.69	0.73
	2	2.11	0.8507	0.98	1.23	1.39	0.81	1.06	1.09	0.43	0.68	0.70	0.11	0.36	0.41
	3	3.29	0.7241	0.61	0.97	1.00	0.28	0.63	0.71	-0.13	0.23	0.31	-0.43	-0.07	0.02
	4	4.56	0.5527	-0.09	0.39	0.43	-0.37	0.12	0.15	-0.65	-0.17	-0.25	-0.99	-0.51	-0.55
$MgCl_2$[e]	1	1.03	0.9400	1.45	1.59	1.63	1.23	1.37	1.34	0.84	0.98	0.95	0.53	0.68	0.65
	2	2.10	0.8368	0.87	1.15	1.32	0.77	1.05	1.02	0.43	0.71	0.63	0.01	0.29	0.34
	3	3.26	0.6597	0.35	0.77	0.75	0.24	0.66	0.45	-0.01	0,41	0.06	-0.55	-0.13	-0.06
	4	4.50	0.5082	-0.29	0.24	0.24	-0.30	0.23	0.06	-0.76	-0.23	-0.45	-1.05	-0.52	-0.75

[a] Value of k used: 5.02.

[b] Molalities were calculated from density data found in *International Critical Tables*, Vol. III, McGraw-Hill, New York, N.Y., 1928, p. 80 (NaBr), p. 72 ($CaCl_2$), p. 71 ($MgCl_2$).

[c] exptl = experimental, calcd = calculated.

[d] Corrected H_0 values for NaBr solutions were calculated by using a value of 0.040 for the k_s parameter.[12]

[e] Corrected H_0 values for $CaCl_2$ and $MgCl_2$ solutions were calculated by using our solubility data.

From Ojeda, M. and Wyatt, P. A. H., *J. Phys. Chem.*, 68, 1857 (1964). With permission.

cation, one can assume that $f_{H_h^+} \approx f_{BH^+}$, and Equation 17 becomes

$$-H_0 - \log c + h \log a_{H_2O} = \log f_B = Ac \quad (18)$$

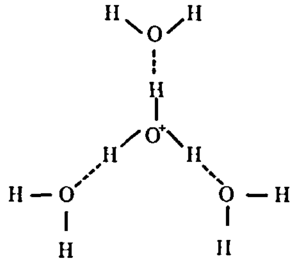
where c is the molar concentration of the acid and represents $[H_h^+]$, and A is the salting-out coefficient in the Setschenov equation. The water activities are available from vapor pressure measurements for hydrochloric, perchloric, and sulfuric acids. Making h = 4, Equation 18 gives reasonable values of A for these acids up to c = 8. The conclusion is that the assumption of a species $H(H_2O)_4^+$ is adequate to account for most of the increase in acidity in concentrated acid solutions.

Bascombe and Bell[45] derived a more general equation without recourse to measured water activities that, with certain assumptions, reduces to

$$-H_0 = \log c - h \log (c/m)(1 - 0.018hm) \quad (19)$$

where m is the molality of the acid. Making h = 4 as before, calculated values of H_0 for hydrochloric, hydrobromic, perchloric, and sulfuric acids agree within about 0.1 unit with measured values up to m ≈ 8. The equation begins to fail at concentrations higher than about 10*m* because less than h water molecules are available for each proton.

Others have derived a hydration number of 4 for the proton from activity coefficients of aqueous acids,[46] measurements of the uptake of water by acidic ion exchange resins,[47] and from the specific heats of aqueous acids.[48] Wicke, Eigen, and Ackermann[48] proposed the following symmetrical structure:



In addition, a more loosely bound (by hydrogen bonds) outer layer of water molecules is assumed,[49] with a total of up to 10 water molecules,[50] at least in dilute solution.

Wyatt[40] investigated the chemical hydration of concentrated sulfuric acid solutions. He concluded that the first stages of hydration of the H_3O^+ and HSO_4^- ions occur in 70 to 90% sulfuric acid and that the appearance of molecular H_2SO_4 Raman lines for these solutions is due to strong ionic hydrations:

$$2H_3O^+ + HSO_4^- \rightleftharpoons H_3O^+(H_2O) + H_2SO_4$$

rather than the weakness of the acid. At greater dilutions (65–20% sulfuric acid) it is shown that the Hammett acidity function is a function of the water activity for sulfuric, perchloric, hydrochloric, and nitric acids.

In a separate study, Rosenthal and Dwyer[51] showed that H_0 values in 1–9*M* lithium chloride, 1–9*M* perchloric acid, and 1–11*M* hydrochloric acid could be calculated using the equation

$$H_0 = pCH + n \log a_{H_2O} - BC \qquad (20)$$

where C is either the molar or molal concentration, CH is the corresponding molar or molal hydrogen ion concentration, and n and B are least-square constants; n is comparable to h in Equation 18. Using a value of $n = 4$, calculated H_0 values differ significantly from measured values at greater than 6*M*. These deviations were explained in terms of a competition for water of solvation between the hydrogen or lithium ion, chloride or perchlorate ion, indicator base, and conjugate acid, which results in a value of n less than 4. Bascombe and Bell[45] assumed that B and BH^+ were either unhydrated or hydrated to the same extent. If least-square values of n were used in Equation 20, Rosenthal and Dwyer[51] were able to calculate H_0 values within the experimental error of measured H_0 values. Even in dilute solution, they found no evidence for $n = 4$, and values of $n = 2$ better fit the data for hydrochloric acid and lithium chloride. Even then, deviation from experimental H_0 values occurred, and these were explained in terms of an effective change in n in going from dilute to more concentrated solutions. As implied before, in sufficiently concentrated salt solutions, there is insufficient water to satisfy the "normal hydration numbers" of all the species present.

Others have investigated the effects of concentrated salts on acidities.[51a-51f] Schwabe[51a] found that in acidic and basic solutions, addition of salts led to higher H^+ activity, proportional to the salt concentration. NR_4^+ salts caused decreasing H^+ activity. Schwabe concluded that there is a higher effective H^+ concentration because part of the water is bound as hydrated water.

For further reading concerning strong electrolyte solutions, the books by Robinson and Stokes[52] and Harned and Owen[53] are recommended.

C. Aqueous Solutions

1. Titration of Weak Bases

Ordinarily, bases with K_b values smaller than about 10^{-8} to 10^{-9} cannot be titrated with strong acids in water. Critchfield and Johnson[4] discovered that in the presence of concentrated neutral salts, bases with ionization constants as low as 10^{-12} could be titrated potentiometrically. Figure 8 illustrates the titration curve for aniline ($K_b = 3.8 \times 10^{-10}$) titrated with 0.5*N* hydrochloric acid in the absence and in the presence of 7*M* sodium iodide. The break is not analytically useful in the absence of the salt, but in the presence of the salt, a well-defined end point is obtained. The base end of the curve, prior to the end point, is not affected, suggesting that any change in the acid-base equilibria of the weak base in the presence of the salt is relatively small. (See, however, the studies below by Rosenthal and co-workers[5,39] for bases of different charge which indicate the contrary). The acid end of the titration curve, however, following the end point, is shifted markedly toward lower pH values; the apparent acidity is more than 100-fold greater than in the absence of salt.

The titration curves in Figure 9 illustrate that the enhancement of the potentiometric break is definitely a function of the concentration of the salt. Inspection of curve 2 in Figure 8 shows that the addition of 4.0 ml of excess hydrochloric acid (corresponding to 0.0192*M*) decreases the pH of the solution to about 0.0. This compares with the value of zero predicted from Figure 5 for 0.0192*M* hydrochloric acid in 7*M* sodium iodide.

It appears then, that after the base is neutralized, the pH of the solution depends on the excess hydrochloric acid and is independent of the amount BH^+. According to Critchfield and Johnson,[4] the mechanism of enhancement of the potentiometric break must be associated with the decrease in the pH of mineral acids by neutral salts. If the theory of the increased activity of the

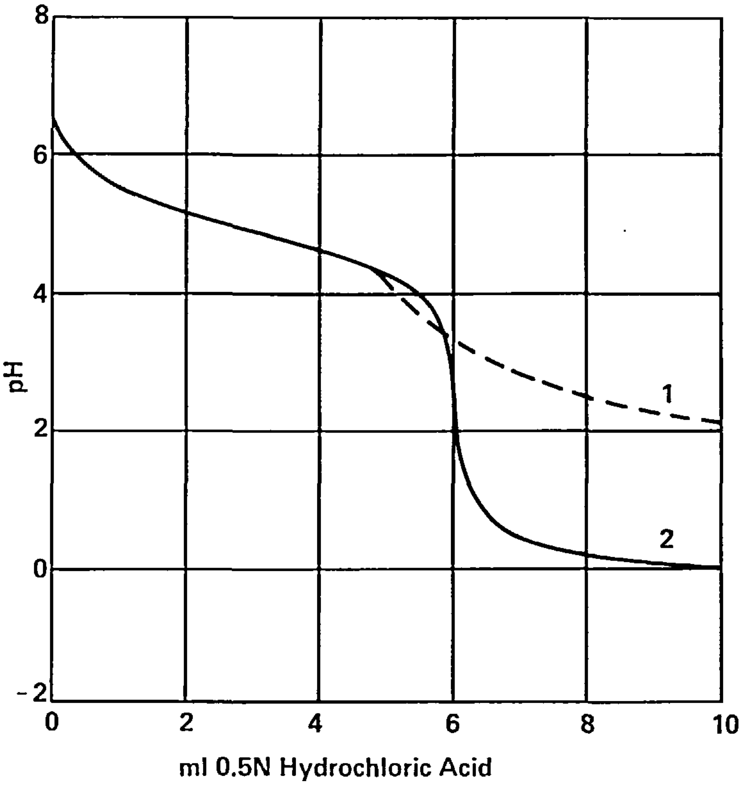


FIGURE 8. Potentiometric titration of aniline. 1. In water. 2. In 7 *M* aqueous sodium iodide. (From Critchfield, F. E. and Johnson, J. B., *Anal. Chem.*, 30, 1247 (1958). With permission.)

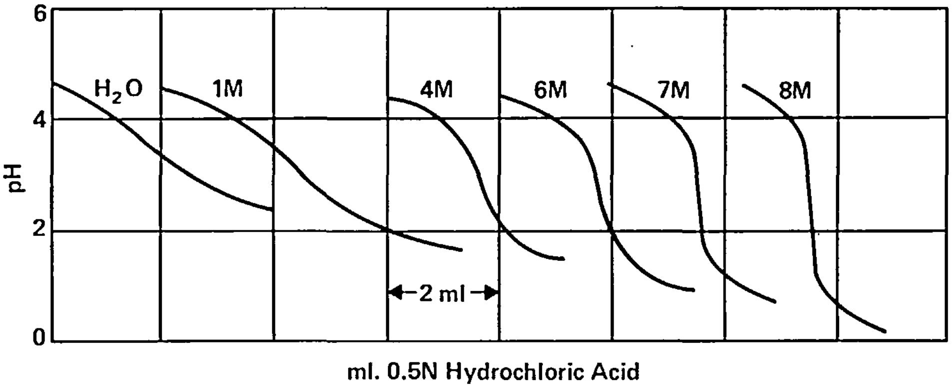


FIGURE 9. Potentiometric titration of aniline in aqueous sodium iodide solutions. (From Critchfield, F. E. and Johnson, J. B., *Anal. Chem.*, 30, 1247 (1958). With permission.)

hydrogen·ion is correct this would explain the effect.[27] (Again, see below the results of Rosenthal and co-workers[5,39] which suggest the improvement in titrations can be explained by factors affecting the equilibria involved).

Generally, any salt of a fairly strong base and a strong acid will enchance the potentiometric break for the titration of the weak uncharged bases studied by Critchfield and Johnson.[4] Examples are sodium chloride, lithium chloride, and calcium chloride. Sodium sulfate, on the other hand, inhibits the break for aniline, and so the salt must be derived from an acid stronger than sulfuric acid ($K_a > 10^{-2}$). Strong base-weak acid salts, such as sodium acetate, inhibit the break, which is not surprising since they are appreciably basic. Note that these conclusions may not be valid for charged bases.

Titrations of weak bases in concentrated salt solutions offer some unique advantages. Different amino nitrogens of poly-functional amines can sometimes be differentiated while they cannot in water or other solvents. For example, in the titration of diethylenetriamine in water, the first two amino nitrogens are titrated together to give a single end point. But the third amino nitrogen is too weakly basic ($K_3 = 4.7 \times 10^{-11}$) to be titrated. In *6M* iodide, however, a second end point is obtained, corresponding to the titration of the third nitrogen. Similarly, the fourth nitrogen in triethylenetetramine ($K_4 = 1 \times 10^{-12}$) cannot be titrated in water, but can be titrated in *8M* sodium iodide. Although nonionizing solvents such as acetonitrile or methylisobutylketone are usually good differentiating solvents, when the above compounds are titrated in these solvents, insoluble salts form prior to complete neutralization. Conversely, titrations in a leveling solvent, such as glacial acetic acid, allow the determination of the total neutralization values of the compounds, but the individual amino nitrogen cannot be differentiated.

Strong amines can be titrated satisfactorily in the presence of weak amines in aqueous solution, but weak amines cannot be titrated in the presence of strong amines. However, in concentrated salt solutions, the potentiometric break of weak amines is enhanced sufficiently that these can be determined in the presence of strong amines. Hence, triethylamine-pyridine mixtures can be titrated in *6M* sodium chloride. Similar titrations can be performed in nonionizing solvents, but sometimes the amines form insoluble salts.

Critchfield and Johnson[4] studied the effect of concentrated salt solutions on acid-base indicators. Table 5 summarizes the pH range in which a variety of indicators change colors, from pH 1.5 to 7.6. In general, the pH range is not appreciably affected by the presence of the salts. Some indicators, however, are subject to salt effects that shift the pH of their color transition. These include bromophenol blue, bromophenol green, and phenolphthalein. The indicators in Table 5 are recommended for titrations in salt solutions.

Table 6 extends the list of indicators, as studied by Rosenthal and Dwyer.[5] Note that they include bromoresol green and that there is an observable effect by the salt on the color transition. Kolthoff[3] made a series of experiments on the effect of several alkali and alkaline earth salts on the color of dimethyl yellow, methyl orange, congo red, methyl red, *p*-nitrophenol, sodium alizarinsulfonate, and azolitman in dilute carbon dioxide solutions. His general conclusion was that indicators with an alkaline character are given an acid tint by neutral salts, while the indicators with an acid character are given an alkaline tint.

Table 7 illustrates the successful titration of several bases ($K_b = 1.3 \times 10^{-3}$ to 4.7×10^{-11}) in concentrated salt solutions using indicators to detect the end point.[4] Results agreed closely with those obtained by titration in glacial acetic acid. The indicator method was limited, though, by the lack of indicators that could be used below pH 2. This is not usually a problem in water, but in the salt solution the pH frequently is less than zero. Note from Table 6, however, that it should be possible to find indicators which can be used below pH 2, e.g., erythrosin or possibly H_o indicators. Thymol blue cannot be used for amines with ionization constants less than 10^{-11}. Hence, the fourth amino nitrogen of triethylenetetramine ($K_4 = 1 \times 10^{-12}$) cannot be titrated using an indicator.

Rosenthal and Dwyer[5] extended the study of acid-base titrations in concentrated salt solutions to consider differences observed with different bases, the accuracy of titrations, and differentiating and leveling in the titration of mixtures. They compared experimental titration curves with calculated titration curves (using experimentally measured K_{BH}/Q_{BH} ratios[33,36]) and the two agreed very closely. Figure 10 shows the curves for

TABLE 5

Behavior of Indicators in Strong Salt Solutions

Indicator	Color change, acid to base	Medium[a]	Approximate pH range
Thymol blue	Red to orange	Lithium chloride	1.5 to 2.7
		Calcium chloride	1.5 to 3.0
		Water	1.9 to 2.6
M-Alka Ver	Red to green	Lithium chloride	3.0 to 4.0
		Calcium chloride	2.9 to 4.2
		Water	3.3 to 4.2
Methyl orange	Red to orange	Lithium chloride	4.1 to 5.5
		Calcium chloride	4.0 to 5.0
		Water	4.1 to 5.0
Methyl red	Red to yellow	Lithium chloride	5.6 to 6.4
		Calcium chloride	5.4 to 6.4
		Water	5.4 to 6.4
Bromothymol blue	Yellow to blue	Lithium chloride	6.2 to 7.5
		Calcium chloride	6.1 to 7.4
		Water	6.0 to 7.6

[a] All salt solutions are aqueous. Lithium chloride, 8 *M*; calcium chloride, 4.5 *M*.

From Critchfield, J. E. and Johnson, J. B., *Anal. Chem.*, 30, 1247 (1958). With permission.

TABLE 6

Some Indicator End Points in Water and LiCl Solutions

		Charge type		pH_{GE} Interval			
Indicator	pK_{BH}	Base	Acid	H_2O,b,e p.,a	4*M* LiCl,b,e.p.,a	8*M* LiCl,b,e.p.,a	Color change,[b] b,e.p.,a
Erythrosin		–	0	3.8,3.5,2.9	2.2,1.0,0	2.1,0.6,<0	Pk,e.p.,Or
Thymol blue	1.75	–	±	2.8,2.5,1.4,0.6	2.6,2.1,1.3,1.0	4.0,3.0,2.2,1.5	Y,Or,R(P)
Malachite green		+	+2	2.0,1.4,0.70,0.07	3.0,2.2,1.3,0.7	4,2.7,1.2	B,G,e.p.,Y
Bromcresol green	4.68	–2	–	5.4,4.4,3.6	4.7,3.9,3.1	4.5,3.5,2.9	B,G,Y
p-Aminoazobenzene	2.82	0	+	2.8,1.1,<0	3.8,2.2,∿1.6	4.4,3.1,∿2.4	Y,Pk Br, Br Or
H_0 Indicator[a]	(3.06)	0	+	(3.13)	(3.17)	(3.25)	
Ac^--like	(3.06)	–	0	(2.99)	(2.12)	(1.35)	

[a] Hypothetical indicators. The values in parentheses are calculated end points corresponding to the pH_{GE} at which the concentration of the acid form of the indicator equals the concentration of the basic form. Values in water are the calculated pa_H for solutions having an ionic strength of 0.03*M*.

[b] Pk = pink, Or = orange, Y = yellow, R = red, P = purple, B = blue, G = green, Br = brown, b = basic form, e.p. = end point, a = acid form.

From Rosenthal, D. and Dwyer, J. S., *Anal. Chem.*, 35, 161 (1963). With permission.

TABLE 7

Titration of Weak Bases in Strong Salt Solutions

Compound	Ionization constant	Medium	Indicator	Purity, % by weight Salt solution	Other[a]
Diethylamine	1.3×10^{-3}	Calcium chloride, 4.5*M*	Methyl red	98.0	98.2
Triethanolamine	4.5×10^{-7}	Sodium chloride, 6*M*	Methyl orange	102.1	102.6
γ-Picoline	1.1×10^{-8}	Lithium chloride, 8*M*	Thymol blue	95.6	95.4
Pyridine	1.7×10^{-9}	Lithium chloride, 8*M*	Thymol blue	99.1	99.6
Aniline	3.8×10^{-10}	Lithium chloride, 8*M*	Thymol blue	99.8	99.4
Diethylene triamine	4.7×10^{-11}[b]	Lithium chloride, 8*M*	Thymol blue	99.3	99.4
Sodium acetate	–	Sodium iodide, 8*M*	Thymol blue	100.0	99.4

[a] By titration in acetic medium with perchloric acid using crystal violet indicator.
[b] Corresponds to ionization of third amino nitrogen (K_3).

From Critchfield and Johnson, J. B., *Anal. Chem.*, 30, 1247 (1958). With permission.

the titration of aniline in lithium chloride. Similar titration curves are obtained in plotting, H_0 vs. percent titrated. Note that the base end of the curve is affected. This indicates that the base itself is indeed appreciably affected by the salt, presumably because the salt affects the concentration equilibrium constant of the base.[5]

For negatively charged bases, the enhancing effect at the end point is less than for uncharged bases. This would not be true though for positively charged bases.[44] Figure 11 shows titration curves for sodium acetate in different salt solutions. The base end of the curve is markedly shifted toward more acid values by the halide salts. In some instances, the change in pH at the end point for charged bases was smaller in salt solution than in water. This was never the case for uncharged bases. Table 8 shows the changes observed by Rosenthal and Dwyer.

Rosenthal and co-workers[5,39,53a] interpret the enhancing effects of salts on titrations of weak bases as follows: Neutral salts can affect activity coefficients and equilibria. For any given salt solution (neglecting the charge on the base),

$$K_{BH} = \frac{[B][H]}{[BH]} \; Q_{BH}$$

where Q_{BH} and f_{H^+} are constants for any given concentrated salt solution. If Q_{BH} is unity for a given base in a given concentrated salt solution, the hydrogen ion *concentration* would be the same at each point in the titration in the salt solution as it would be in dilute aqueous solutions. Generally, log f_{H^+} would be different from zero in a concentrated salt solution. If the proton is "more acid", log f_{H^+} would be positive. The *shape* of the pH_{GE} vs. % titrated titration curve would be exactly the same in the two solutions, but the pH_{GE} values would *all* be shifted by an amount equal to log f_{H^+}. Thus, despite the fact that the "acidity" of the proton is enhanced as determined by pH_{GE} measurements, the titration would not be improved in the salt solution. Therefore, Rosenthal suggests that an improvement in the

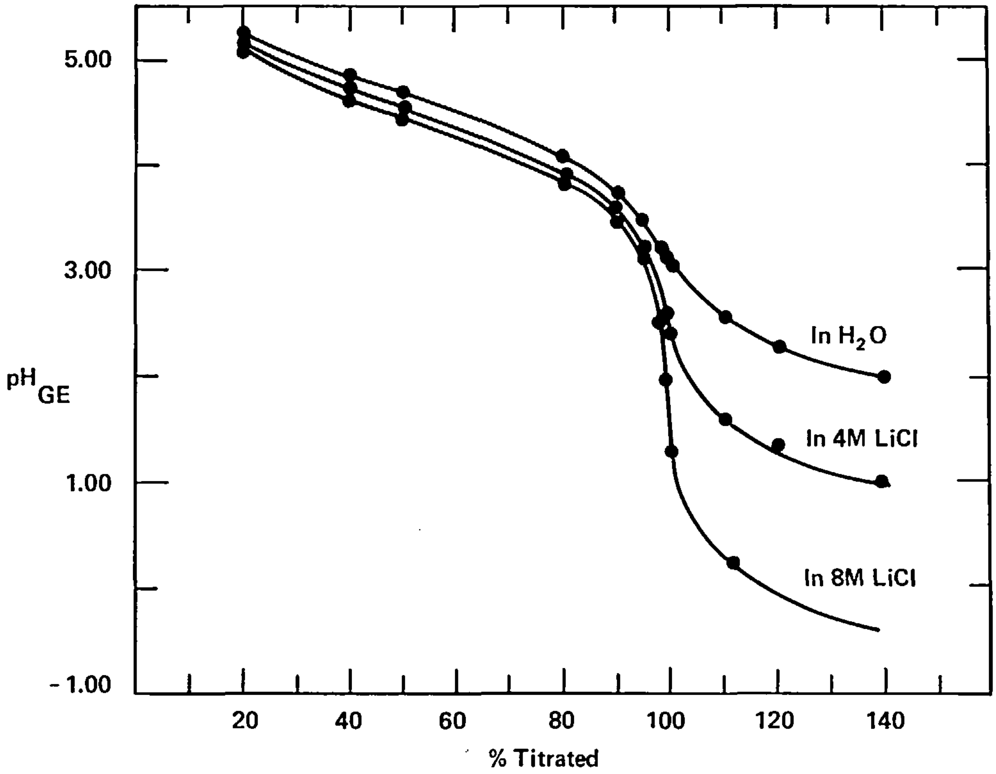


FIGURE 10. Calculated and experimental pH_{GE} values obtained for the titration of 0.03 *M* aniline. Points represent experimental values. Curves are calculated. (From Rosenthal, D. and Dwyer, J. S., *Anal. Chem.*, 35, 161 (1963). With permission.)

titration is due to a shift in the equilibrium from left to right for the reaction

$$B + H^+ \rightleftharpoons BH$$

and the related decrease in the concentration acid dissociation constant K_{BH}/Q_{BH}. This is because Q_{BH} is increased by the addition of salt ($Q_{BH} = \frac{f_B f_{H^+}}{f_{BH} a_w{}^n}$, where a_w represents the activity of water). This increase could be explained in terms of the effect of neutral salts upon activity coefficients and the activity of water.

From these arguments, the small effect of sodium iodide upon the titration curve for aniline in the buffer region (Figure 8) is due to the fact that $Q_{BH} \sim f_{H^+}$ for these solutions. Note in Figures 10 and 11 that pH_{GE} values are different in the buffer region in different salt solutions.

If the improvement in the titration is due solely to "increased acidity of the proton", one would expect the same improvement for all bases in a given salt solution in concentrated salt solutions. Considering the results given in Table 9,[5,39,44] it is clear there are substantial variations in the effect of 4*M* and 8*M* lithium chloride, even for bases of the same charge type.

Rosenthal and Dwyer[5] calculated that to obtain an accuracy of 2 parts/1,000 in the titration of 0.03*M* base, a salt solution must be found such that $pK_{BH} + \log Q_{BH} \geqslant 7$; more generally, $pK_{BH} + \log Q_{BH} + \log C_B > 5.5$.[54] Hence, from log Q_{BH} values,[33,36,51] pK_{BH} for an H_0 indicator-like base must be 4.5 in order to titrate a 0.03*M* solution of the base in 8*M* lithium chloride with an accuracy of 2 parts/1,000 pK_{BH} would have to be 4.4 for an aniline-like base and 6.4 for an acetate-like base. Ojeda, Pérez, and Wyatt[42] used a calculation of the titration error as a direct indication of the salt effect in sharpening the end point, and agreed substantially with the results of Rosenthal and Dwyer.[5]

The differences in Q_{BH} values for individual bases[5] suggests the possibility of achieving either differentiation (as already demonstrated by Critchfield and Johnson[4]) or leveling for the titration of two bases in a particular salt solution.

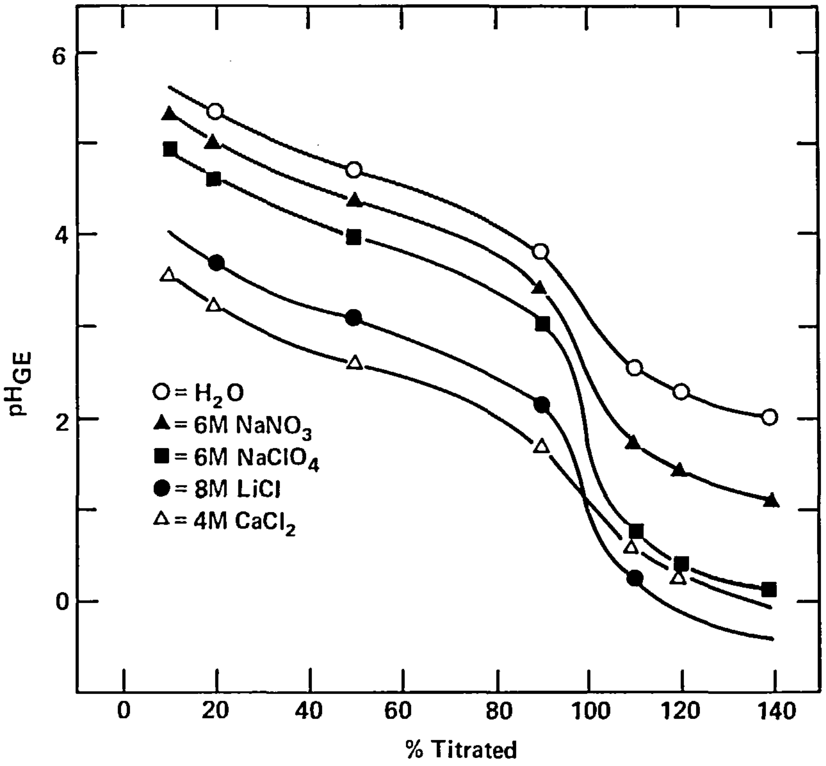


FIGURE 11. Calculated and experimental pH_{GE} values obtained for the titration of 0.03 *M* sodium acetate. (From Rosenthal, D. and Dwyer, J. S., *Anal. Chem.*, 35, 161 (1963). With permission.)

TABLE 8

Change in pH_{GE} in the Vicinity of the Equivalence Point ΔpH in Going from 90 to 110% Neutralized in the Titration of 0.03*N* Base with HCl

	Uncharged bases				Charged bases				
Solution	2AP[a] pK_{BH} 3.637	Aniline, 4.597	THAM[a] 8.077	Pip[a] 11.124	Fo^{-a} 3.739	Ac^{-a} 4.757	Ph^{-a} 9.998	OH^-	$p\text{-}PS^{-2}$,[a] 9.054
Water	0.50	1.22	(4.4)[b]	(6.9)[b]	0.53	1.28	(6.0)[b]	8.82	(5.3)[b]
4*M* LiCl	1.36	2.00	(5.9)	(7.9)	0.51	1.31	(6.5)	8.43	(5.8)
8*M* LiCl	2.81	3.30	(6.4)	(8.5)	0.81	1.90	(7.2)	8.75	(6.4)
6*M* $NaClO_4$	2.16					2.31		9.15	
6*M* $NaNO_3$	1.45					1.71		8.88	
4*M* $CaCl_2$	2.45					1.21		7.94	

[a] 2AP is 2-aminopyrimidine; THAM is tris(hydroxymethyl)aminomethane; Pip is piperidine; Fo^- is formate; Ac^- is acetate; Ph^- is phenolate; OH^- is hydroxide; $p\text{-}PS^{-2}$, *p*-phenolate-sulfonate.

[b] Values in parentheses were determined with an automatic titrator and may be in error by 0.2 or more.

From Rosenthal, D. and Dwyer, J. S., *Anal. Chem.*, 35, 161 (1963). With permission.

TABLE 9

Log Q_{BH} Values for Bases of Different Charge Type in 4*M* and 8*M* Lithium Chloride

Base	pK_{BH}	Base charge type	Log Q_{BH} values	
			4*M* LiCl	8*M* LiCl
Chloromethylphosphonate	6.59	-2	-0.95	–
Formate	3.746	-1	-0.14	0.31
Acetate	4.757	-1	-0.02	0.62_5
Aniline	4.596	0	0.87	2.13
p-Nitroaniline	0.99	0	0.98	2.39
N-methylaniline	0.525	0	1.31	2.96
N,N-Dimethylaniline	0.65	0	1.56	3.37
2-aminopyrimidine	3.663	0	1.14	2.57_5
N,N-dimethylethylenediamine	6.59	+1	1.31	2.52
p-Tetramethyl phenylenediamine		+1	2.13	4.07
$(C_2H_5(CH_3)_2N^+CH_2CH_2)_2NCH_3$		+2	2.66	4.47
$pMH\text{-}pH_{GE}$			0.93	2.31

*Compiled from data in References 5, 39, and 44.

The most striking example would be the differentiation of two bases with identical thermodynamic pK_{BH} values. In order for the first base (B_1) to be titrated with an accuracy of 2 parts/1,000, a salt solution would have to be used in which $pK'_{B_1H} - pK'_{B_2H} = (\log Q_{B_1H} - \log Q_{B_2H})$ is greater than 5.5, where K′ is the concentration equilibrium constant. For the bases considered by Rosenthal and Dwyer,[5] no salt solution was found to achieve this degree of differentiation.

The converse case, leveling, is possible, where the extreme case would be for two bases with different pK'_{BH} values in water but similar values in concentrated salt solutions. For example, 2-aminopyrimidine and acetate ion have pK'_{BH} values differing by about one unit in water, but the values are identical in appropriate concentrations of lithium chloride or calcium chloride.

Rosenthal and Dwyer[5] suggest that salts such as lithium bromide will probably give larger Q_{BH} values for uncharged bases than most other univalent salts because of the large effect of the small cation and large anion. Saxton[43] has confirmed this prediction. Lithium iodide might be even better except for its lower solubility, and the ease of air oxidation of iodide.

Ojeda, Pérez, and Wyatt[42] suggest in their hydration explanation of the salt effect via dehydration of the hydrogen (whereby protonation equilibria are displaced to the right: $H_3O^+ \cdot nH_2O + B = BH^+ + (n+1)\ H_2O$) that sufficiently soluble nonelectrolytes should have a similar effect, provided they are not comparable with water in basic strength or solvating power. In confirmation of this, they were able to produce a sharpening of the end point in the aniline-hydrochloric acid titration by the addition of sugars. This is noteworthy in view of the fact that Matkovich and Christian[30,55] were able to achieve separation of acetone from water by high concentrations of sucrose, similar to their results with high concentrations of certain salts. It is interesting that two relatively simple theories have been proposed for the effectiveness of the "salting-out" abilities of electrolytes and nonelectrolytes for miscible solvents. These are the medium effect theory and the salt effect theory, respectively.[56] The former assumes that the salting-out agent is preferentially solvated by one of the solvents, making this solvent unavailable for dissolving the second. Sucrose probably lowers the solubility of alcohol[57] or acetone[30] in water in this manner. Hence, sucrose presumably can satisfy the above suggested mechanism[42] via hydration. The salt effect theory applies only to ionic compounds and assumes that the more polar

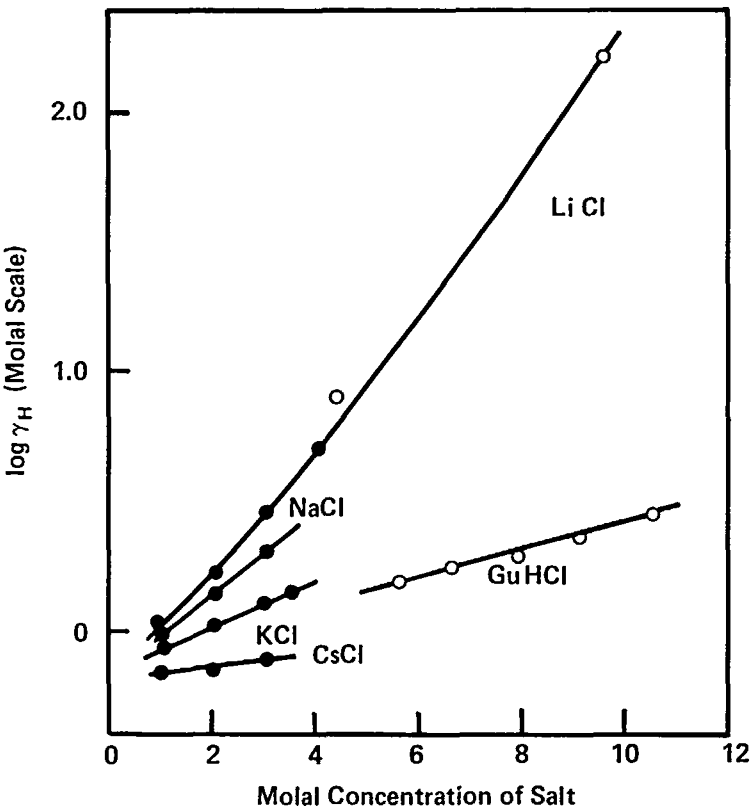


FIGURE 12. Apparent activity coefficients of hydrogen ions in solutions of GuHCl and other chlorides at 25°. Points shown as filled circles are based on $\gamma\pm$ values determined in cells without liquid junction, with assumptions given in the text. Open circles are based on pH measurements with the glass electrode, assuming pH = $-\log C_H\gamma_H$. (From Nozaki, Y. and Tanford, C., *J. Am. Chem. Soc.*, 89, 736 (1967). With permission.)

of the two solvents (e.g., water) preferentially congregates around the salt because of electrostatic attraction, leading to nonideal behavior and positive deviation from Raoult's law.

Nozaki and Tanford[58] reported that acid-base titrations could be performed in concentrated guanidine hydrochloride solutions, not significantly different from concentrated solutions of alkali metal chlorides, except at high pH where the dissociation of the guanidinium ion becomes appreciable. The pK for this dissociation in 6*M* solution is 13.74 at 25°C. Although titration curves are not presented, titration of basic amino acids in 6*M* guanidinium hydrochloride were apparently successful. Nozaki and Tanford measured the pH of 0.001 to 0.06*M* hydrochloric acid in 4 to 6*M* guanidinium hydrochloride and from these measurements calculated activity coefficients for the hydrogen ion in the different salt solutions. The logarithms of these are plotted as a function of salt concentration in Figure 12 and are compared with the corresponding values calculated from the results reported by Rosenthal and Dwyer[33] for alkali metal halide salts. It is seen that the behavior of hydrochloric acid in guanidinium hydrochloride is comparable to its behavior in the other halide salts, being intermediate between potassium chloride and cesium chloride. Note that the activity coefficients in the concentrated guanidinium hydrochloride solutions are greater than unity. These results are interesting in that guanidinium hydrochloride is not a neutral salt, but is slightly acidic (a 6*M* solution should have a pH of approximately 6.5, based only on its dissociation constant in water and neglecting a "self-salt" effect).

The pK values of a number of amino acids were determined in 6*M* guanidium hydrochloride and were found to be close to the corresponding values in dilute salt solutions.[58] Hence, intramolecular electrostatic interactions between charged groups on amino acid molecules are apparently not greatly affected by the presence of high concentrations of this salt.

Christian[59] took advantage of the salt-effect to titrate microequivalent quantities of weak bases with coulometrically generated[60] acid. Halide salts could not be employed because the halide ions are oxidized at the generating anode in preference to the oxidation of water. Attempts were made to react the product (e.g., I_3^-) with an electrolyte to produce an equivalent amount of hydrogen ion, e.g., by reaction with sulfite:

$$H_2O + I_3^- + SO_3^{-2} = 3I^- + SO_4^{-2} + 2H^+$$

Unfortunately, the buffering capacity of the sulfite/bisulfite was sufficiently great to preclude good end point breaks. Several other systems behaved similarly. Although sulfate ions would not be oxidized at the anode, it has already been noted that concentrated sulfate salts inhibit the end point breaks.[27]

Because of the above difficulties, perchlorate salts were selected as probable concentrated salts for the coulometric titrations. The pH of equal concentrations of hydrochloric acid (ca. 0.01*N*) in various solutions was measured as follows: H_2O, 2.16; 3*M* $LiClO_4$, 0.80; 8*M* $NaClO_4$, -0.5. Volumetric titration of 0.1 mg of aniline in 40 ml of solution with 0.02*N* hydrochloric acid showed a pH change at the end point of 0.5, 1.4, and 1.5 in 3*M* lithium perchlorate, 8*M* sodium iodide, and 8*M* sodium perchlorate, respectively. Hence, sodium perchlorate compares favorably with sodium iodide for the enhancement of the acidity (lithium perchlorate has lower solubility than sodium perchlorate in water). In fact, in a separate study, Schertz and Christian[29] observed that the perchlorates of lithium and magnesium (4I concentration[4]) cause the pH of 0.1*M* hydrochloric acid as a function of heat of solution to be over one unit more acid than predicted from Figure 5 for the halides and pseudohalides. The two points appeared to parallel the halide line and perhaps a different family of curves results for different classes of salts. At any rate, the perchlorates appear to exhibit superior acidity enhancement in aqueous solution. This agrees with the anomalous behavior of sodium perchlorate observed by Ojeda and Wyatt[14] (see above).

The enhancing effect of sodium perchlorate is indeed fortunate for coulometric titrations, and Christian was able to obtain excellent accuracy in the titration of small quantities of bases.[59] Table 10 summarizes the results. Only one end point break was observed for ethylenediamine, corresponding to titration of both amine groups.

To summarize some of the advantages of titration of weak bases in aqueous solution containing concentrated salts, these systems are beneficial for titrating basic salts and nonelectrolytes that are not readily soluble in the usual nonaqueous solvents but are soluble in water.[4,5] Sometimes the acid salt product is insoluble in nonaqueous solvents but soluble in water. It should be remembered, however, that most electrolytes[61] and many nonelectrolytes[4,5] are less soluble in concentrated salt solutions than in water. Aniline, for example, has a solubility of 0.39*M* in water but only 0.049*M* in 11.5*M* lithium chloride. In some cases, these amines can be titrated into solution because their salts are normally more soluble. Small amounts of methanol can be used to solubilize amines without precipitating excessive quantities of salts.[4] When this is necessary, lithium chloride is usually the salt of choice because of its organic solubility.

Usually a 10 to 15% dilution of the salt solution by addition of aqueous titrant can be tolerated without appreciably affecting the end point if the salt solution is originally 7 to 8*M*.[4] Coulometric titrations offer the advantage in that the salt solution is not diluted by the titrant.[59]

It should be pointed out that some concentrated salt solutions may have basic impurities that become titratable in these solutions.[30]

It should be mentioned that under certain conditions, some weak bases may be titrated in molten salts. Courgnaud and Tremillon,[61a] for example, carried out potentiometric acid-base titrations with glass electrodes in molten $Ca(NO_3)_2 \cdot 4H_2O$ at 75°C. This solvent is more acidic than water at the same temperature. As a result, bases which are too weak to titrate in water are easily titrated. Because calcium ion can complex numerous bases, its presence strengthens some acids and diminishes the strength of bases. A mixture of pyridine and acetate, which cannot be titrated in water, is titratable in this medium. Janz[61b] has reviewed acid-base relations and domains of acidity in molten carbonates and Masuko[61c] has described acid-base equilibria and

TABLE 10

Coulometric Titrations of Bases in Aqueous Salt Solutions

Electrolyte	Base	μequiv. taken	μequiv. found
6*M* $NaClO_4$	NaOH	16.23	16.15
	NH_3	68 2	69.0
		68.2	69.2
		12.35	12.50
8*M* $NaClO_4$	Aniline	62.4	61.4
	Ethylenediamine	38.2	38.4
		38.2	38.4

From Christian, G. D., *Anal. Chim. Acta,* 46, 77 (1969). With permission.

basicity in fused salt solutions and the effect of the basicity on redox reactions.

2. *Titrations of Weak Acids*

In spite of the success of titrations of weak bases in aqueous concentrated salt solutions, practically no titrations of weak acids in concentrated salt solutions have been successfully attempted. This is apparently primarily because the activity of the hydroxide ion is not enhanced, as in the case of the hydrogen ion (although there is some conflicting evidence – see below), coupled with the fact that any change in the acidity of the weak acid is comparatively small.

Leithe[62] was, however, able to demonstrate a significant sharpening of the end points in the titration of weak acids in saturated sodium chloride solutions with sodium hydroxide titrant. This allowed him to titrate carbonic acid to carbonate and phosphoric acid to phosphate, and to titrate boric acid, hydrogen cyanide, and phenol using either azo violet or thymolphthalein indicator. Figure 13 shows the potentiometric titration curves of potassium bicarbonate and sodium monohydrogen phosphate in the presence and absence of salt, along with the transition range for azo violet. The effect is slight but significant. The pH of the excess sodium hydroxide is actually decreased in the presence of the salt but the initial pH is also more acidic. Note also that the transition range of the indicator is affected by the salt. Using this system, Leithe was able to titrate bicarbonate-cabonate mixtures, and hydroxide in carbonates and vice versa. Sodium nitrate or potassium chloride could be used in place of sodium chloride, but ammonium salt had to be absent.

Christian[63] attempted to titrate weak acids in the presence of alkaline salts with the possibility of "leveling" the acids. Figure 14 shows the titration curves of boric acid ($pK_a \approx 9.2$) and sodium monohydrogen phosphate ($pK_a \approx 12.3$) in the absence and presence of saturated sodium acetate. Although the initial pH was decreased in the case of the weaker acid, no potentiometric break could be obtained. For boric acid, the initial portion of the curve was flatter, suggesting a stronger acid, but the net effect, as in the former case, is to shift the entire curve toward the pH of the saturated sodium acetate. The titration of acetic acid in the presence of either saturated sodium acetate or ammonium chloride is shown in Figure 15. Not surprisingly, the ionization of the acid is greatly suppressed by the common ion acetate, whereas in the presence of ammonium chloride, the final pH is acidic from the untitrated ammonium ion.

An interesting special case of weak acid titrations in 10*M* lithium chloride was reported by Kubota and Costanzo.[64] They were able to titrate free acid in the presence of a hydrolyzable metal ion without the necessity of adding a separate complexing agent (to prevent hydrolysis of the metal), followed by titration of the hydrolyzable cation. The titrant was either lithium or sodium hydroxide dissolved in 10*M* lithium chloride. This system is particularly useful for free acid determination in the presence of ferric iron because there is always some tendency of base to react partially with the complexed iron. The titration of free acid in the presence of either iron (III) or aluminum exhibits a sharp free-acid break followed by a second break corresponding to the total equivalents of hydrolyzable ion. The single break for the cations indicates there is no clear demarcation between the mono-, di-, and trihy-

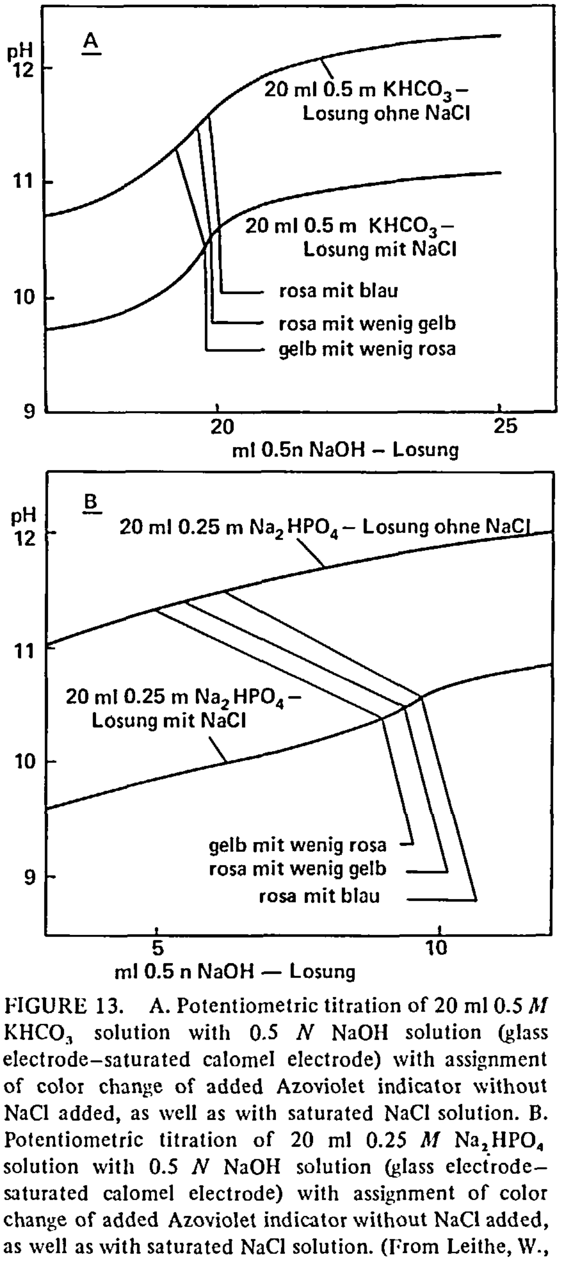


FIGURE 13. A. Potentiometric titration of 20 ml 0.5 *M* $KHCO_3$ solution with 0.5 *N* NaOH solution (glass electrode–saturated calomel electrode) with assignment of color change of added Azoviolet indicator without NaCl added, as well as with saturated NaCl solution. B. Potentiometric titration of 20 ml 0.25 *M* Na_2HPO_4 solution with 0.5 *N* NaOH solution (glass electrode–saturated calomel electrode) with assignment of color change of added Azoviolet indicator without NaCl added, as well as with saturated NaCl solution. (From Leithe, W., *Z. Anal. Chem.*, 189, 396 (1962). With permission.)

droxy species. The titration of free acid in the presence of beryllium, however, shows two breaks for the beryllium which correspond to the mono- and di-hydroxy beryllium compounds. This suggests that the two species have widely separated formation constants, probably greater than 10^7 (see below). This is the only cation studied that exhibited this behavior.

The cations in whose presence free acid could be determined also include chromium (III), copper (II), nickel, niobium, rare earths, thorium, uranium, and zirconium. In the case of niobium, the lithium chloride medium should be as close to saturation as possible (13.9*M*). All the above cations can be determined as the hydroxides. The magnitude of the free acid break in the presence of ions that form chloro complexes (copper, iron, niobium, uranium, and zirconium) is independent of the amount of cation. But in the presence of a nonchloro complexing ion, e.g., aluminum, the break is suppressed by high concentrations of the cation.

The titration of a mixture of acid, iron, chromium, and nickel ions is shown in Figure 16, illustrating the differentiating capability of this system. The order of titration is from the least soluble hydroxide to the most soluble. Comparison of pK_{sp} values for various metal hydroxides indicates that, empirically, these values must differ by at least seven in order for differentiation to be possible. Some pK_{sp} values are: $Fe(OH)_3$, 37.2; $Cr(OH)_3$, 30.1; $Ni(OH)_2$, 14.7; $Cu(OH)_2$, 19.7; $Al(OH)_3$, 31.7. Hence, copper can be differentiated from iron, chromium, or aluminum, but not from nickel. Likewise, aluminum can be distinguished from copper or nickel, but not from chromium or iron.

Aluminum exhibits an unusual property in that when relatively large amounts are titrated, two potentiometric breaks are obtained, the first corresponding to aluminum hydroxide, and the second close to the aluminate break. No explanation was found for this differing behavior compared to more dilute solutions.

The efficacy of titrations of hydrolyzable cations in lithium chloride solutions probably has to do with the dehydrating properties of the salt.[64] Hogfeldt and Leifer[65] have calculated that the average hydration of lithium in lithium chloride solutions decreases from six in dilute solution to 3.1 in concentrated solution, while the chloride hydration decreases from 2 to 0. Above 6*M* lithium chloride, the lithium ions do not have the maximum hydration. Kubota and Costanzo, therefore, postulate that the salt would be expected to take up any free water or waters of hydration from cations which do not retain waters of hydration as readily as lithium ion. Cations, such as copper or iron, which form chloro complexes would give up their waters of hydration easily.

While on the subject of titrations with bases, it is relevant to review the findings of a few studies

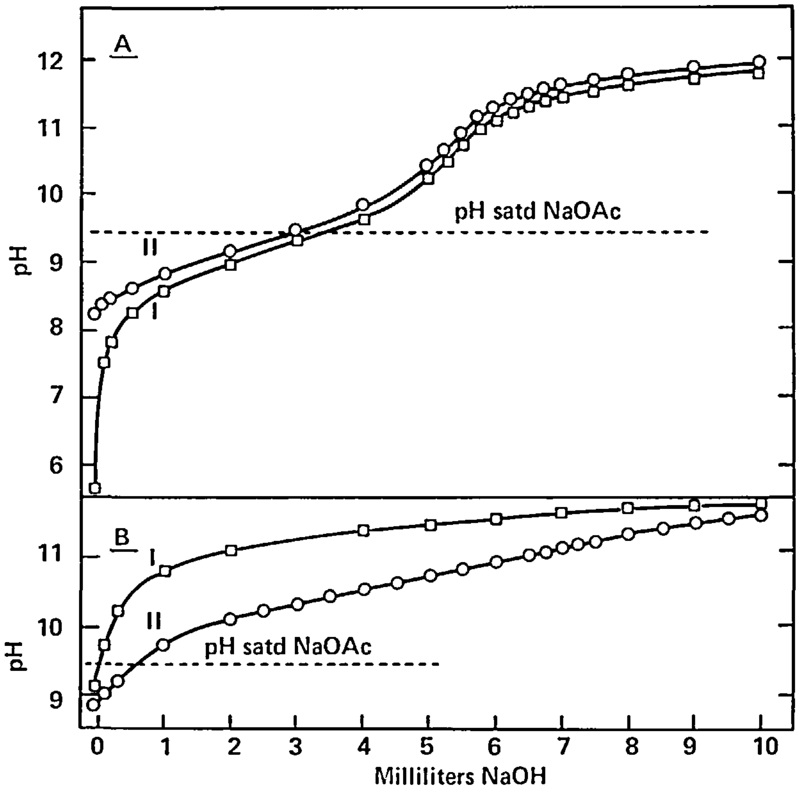


FIGURE 14. Titration of weak acids in the presence of sodium acetate. A. Boric acid. Curve I – 1 ml 1 *M* H_3PO_3 in 50 ml H_2O vs. 0.2 *N* NaOH. Curve II – 1 ml 1 *M* H_3PO_3 in 50 ml saturated NaOAc vs. 0.2 *N* NaOH. B. Sodium monohydrogen phosphate. Curve I – 2 ml 0.5 *M* $Na_2HPO_4 \cdot 7H_2O$ in 50 ml H_2O vs. 0.2 *N* NaOH. Curve II – 2 ml 0.5 *M* $Na_2HPO_4 \cdot 7H_2O$ in 50 ml saturated NaOAc vs. 2 *N* NaOH. (From Christian, G. D., unpublished results.)

concerning the basicity of the hydroxide ion in concentrated electrolyte solutions. Rosenthal and Dwyer[5] measured the pH of 0.03*M*, 0.024*M*, 0.15*M*, and 0.003*M* sodium hydroxide solutions in concentrated salt solutions using a glass electrode-S.C.E. pair. Their results are summarized in Table 11. As in acid solution, the pH values were more acidic in the salt solutions than in water. However, $pMOH + pH_{GE}$ was not a constant, i.e., $\Delta pMOH \neq -\Delta pH_{GE}$. The authors concluded that because of this and because of the known response of the glass electrode to alkali metal ions in basic solution, the meaningfulness and utility of pH_{GE} as a measure of basicity is limited. Nozaki and Tanford[58] point out also that the liquid junction potential cannot be assumed to be negligible (<0.02 pH units) in alkaline solutions, whereas it can be assumed so in acid solutions.

These latter authors estimated the apparent activity coefficient of the hydroxide ion in concentrated guanidinium hydrochloride solutions from pH measurements with liquid junction. Their results are shown in Table 12. Whereas the activity coefficient of the hydrogen ion was considerably greater than unity, that of the hydroxide ion was less than 0.1. It was concluded that the reason for this is the removal of hydroxide ions from solution by combination with guanidinium ions to form the free base guanidine; i.e., the free concentration of hydroxide ions cannot be considered to be the same as the added concentration of potassium hydroxide. The activity coefficient of alkali hydroxides remains less than unity in up to 4*M* solutions of alkali salts (see Figure 4).[20]

Bowden[19] has reviewed measurements of acidity functions H_- [defined as $pK_a - \log([HA]/[A^-])$ when using a weak acid indicator and equivalent to $-\log(a_{H^+} f_A/f_{HA})$] in strongly basic solutions. In this case, the function measures the ability of the solution to remove a proton from the acid. In dilute aqueous solutions up to 0.1*M*, H_- is identical with pH. In more concentrated

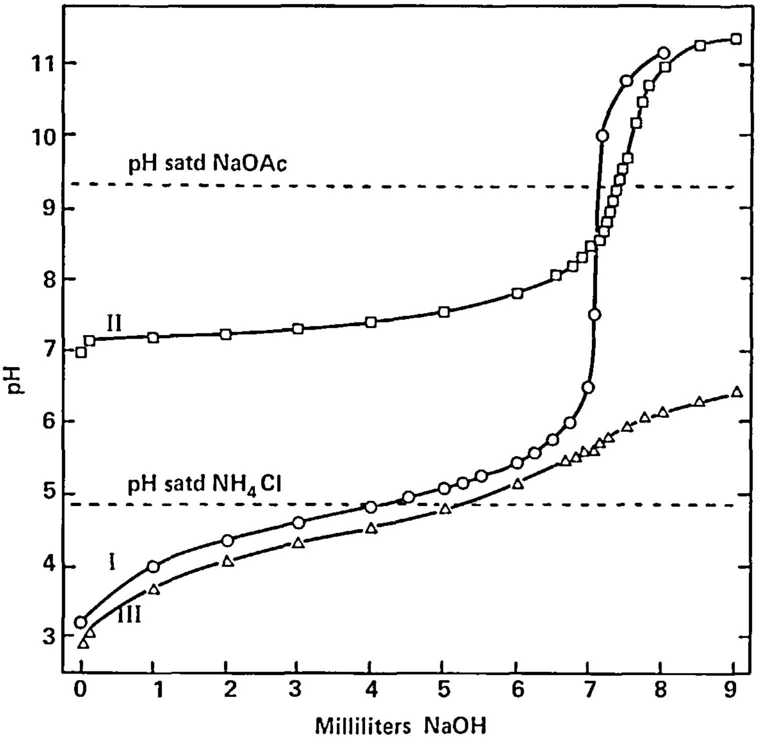


FIGURE 15. Titration of acetic acid in the presence of salts. Curve I – 1 ml 1 *N* HOAc in 50 ml H_2O vs. 0.2 *N* NaOH. Curve II – 1 ml 1 *N* HOAc in 50 ml saturated NaOAc vs. 0.2 *N* NaOH. Curve III – 1 ml 1 *N* HOAc in 50 ml saturated NH_4Cl vs. 0.2 *N* NaOH. (From Christian, G. D., unpublished results.)

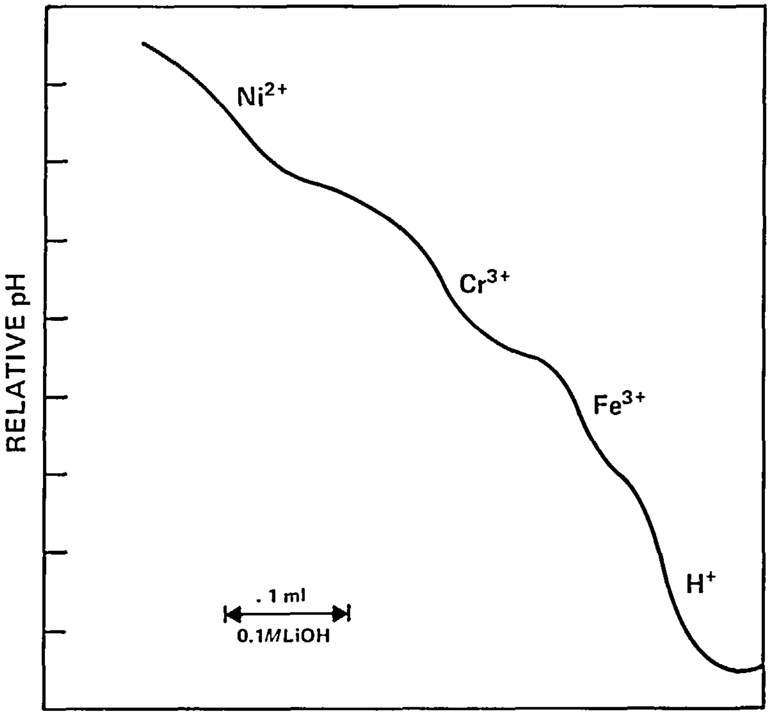


FIGURE 16. Titration of free acid and stainless steel corrosion products. Titration from left to right, pH increasing from bottom to top. (From Kubota, H. and Costanzo, D. A., *Anal. Chem.*, 36, 2454 (1964). With permission.)

TABLE 11

pH of Sodium Hydroxide Solutions in Concentrated Salts*

Solution	Concentration NaOH 0.003 M	0.015 M	0.024 M	0.03 M
H_2O	11.40	11.98	12.09	12.16
4 M LiCl	10.03	10.41	10.50	10.54
8 M LiCl	9.01	9.14	9.19	9.22
6 M $NaClO_4$	9.84	10.20	10.28	10.34
6 M $NaNO_3$	10.58	10.98	11.02	11.09
4 M $CaCl_2$	8.55	9.24	9.42	9.50

*Compiled from data in Reference 5.

TABLE 12

Apparent Activity Coefficients of OH^- Ions in Aqueous Guanidine Hydrochloride at 25°C[a]

GuHCl concentration		Log γ'OH	
M	m	M scale	m scale
5.0	7.8	-0.975	-1.165
5.5	9.1	-1.015	-1.230
6.0	10.5	-1.045	-1.290

[a]These data are independent of pH.

From Nozaki, Y. and Tanford, C., *J. Am. Chem. Soc.*, 89, 736 (1967). With permission.

solutions, H_- increases more rapidly than pH. The order of basicity for the same concentration in aqueous solution is LiOH < NaOH < KOH < $BzMe_3NOH$. This is in the opposite order to that of the water content of the very concentrated solutions. In addition, the basicity rises rapidly when the water content of the solutions is sharply reduced. The reason for the rise is suggested by Yagil and Anbar[66] to be similar to that given by Bascombe and Bell[45] for the change in H_0 in concentrated acid solutions. The equilibria may be represented as

$$HA + OH^- (H_2O)_n = A^- + (n + 1) H_2O. \qquad (21)$$

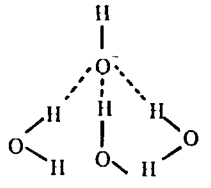
where n is the effective hydration number of the hydroxide ion. Assuming the ratio of activity coefficients in this equilibria to be unity, the following expression results:

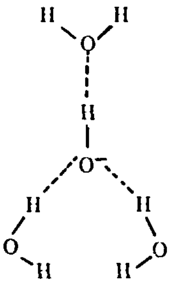
$$H_- = pK_w + \log([OH^-]/[H_2O]^{n+1}) \qquad (22)$$

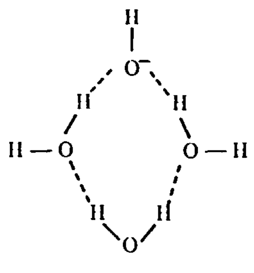
The sharp rise in basicity is therefore attributed to the fall in the free water concentration.

The primary hydrated hydroxide species in solution has been proposed to normally be $OH^-(H_2O)_3$[66] and a hydration number, n, of 3 gives the best agreement of Equation 22 with the data of Schwarzenbach and Sulzberger.[67] Bowden[19] suggests that it is likely that in the absence of full hydration the activity of the hydroxide ion is very high.[66]

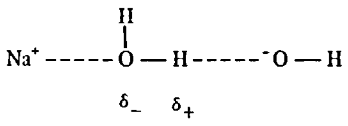
Proposed structures for the hydrated hydroxide ion are:


I


II


III

Structure II is similar to that proposed by Ackermann.[49] Robinson and Harned[68] suggest that in greater than $0.1M$ alkali metal hydroxide solutions, "solvent-bridged" ion pairs exist:



$$\mathrm{Na^+}\text{-----}\underset{\delta_-}{\overset{\overset{\mathrm{H}}{|}}{\mathrm{O}}}\text{—}\underset{\delta_+}{\mathrm{H}}\text{-----}{}^{-}\mathrm{O}\text{—}\mathrm{H}$$

Such ion pairs would explain the order of activity coefficients ($\mathrm{Li} < \mathrm{Na} < \mathrm{K} < \mathrm{Cs}$), since the smaller lithium ion would more strongly polarize the water molecules that are involved in hydration and increase the bridging tendency. Formation of such solvent-bridged ion pairs as solutions become more concentrated would prevent the activity of the hydroxyl ion from rising as rapidly as it would otherwise.

Stewart and co-workers[69,70] have measured other acidity functions in strongly basic solutions and Saxton[43] has measured log Q values for various acidity indicators in dilute base-concentrated lithium chloride solutions.

Finally, it should be noted that anions such as chloride have very little water of solvation[30] and hence would be expected to be poor competitors with hydroxide ions for water removal.

D. Nonaqueous Solutions

1. Titration of Weak Bases

Little work has been done on the effect of neutral salts on titrations in nonaqueous solvents. Christian[59] demonstrated that the potentiometric end point in the titration of weak bases in nonaqueous solvents could be enhanced in the presence of concentrated neutral salts. Microequivalent quantities of bases with K_b values as small as 4×10^{-12} were titrated with coulometrically generated acid in acetone containing $3M$ lithium perchlorate. One percent water was added as a source of protons. Small amounts of water were demonstrated later to cause a significant decrease in the potentiometric break.[29] The titration of 0.1 mg of aniline in 40 ml exhibited an apparent potentiometric break of 2.7 pH units at the end point, compared to 1.4 in $8M$ aqueous sodium perchlorate. About 20 μequiv. (microequivalents) of sodium hydroxide, aniline, ethylenediamine, and *m*-nitroaniline were titrated. The results for the last base were 17% high, probably due to some direct electrolysis at the anode. No end point break for this weak base ($K_b = 4 \times 10^{-12}$) could be obtained in acetone containing only $0.1M$ lithium perchlorate. Both amine groups of ethylenediamine titrated together as in aqueous sodium perchlorate solution[59].

Schertz and Christian[29] undertook a detailed investigation of this titration system to determine how general the salt effect might be, using volumetric titration. They were able to titrate bases as weak as caffeine ($K_b = 4.1 \times 10^{-14}$) in aprotic solvents, and even obtained a small potentiometric break for urea ($K_b = 1.6 \times 10^{-14}$). A number of solvent and salt systems were investigated for possible combinations. Primarily because of solubility limitations, only lithium perchlorate, lithium iodide, or magnesium perchlorate dissolved in either acetone, methylisobutylketone, or 2-butanone resulted in enhanced end point breaks. Several salts dissolved in glacial acetic acid, but none resulted in an end point enhancement (see below).

The titration of aniline in acetone containing increasing concentrations of lithium perchlorate is shown in Figure 17. Note that the potential readings are decreased (toward more alkaline values) by the addition of the salt, but the decrease is larger before the end point than beyond it, causing an overall enhancement of the break. Actually, the break reaches a maximum at about $1M$ salt, because at higher concentrations, the potential is shifted downward to a greater extent after the end point than it is before the end point. The shift of the apparent pH toward more basic values before the end point is in opposition to the results found in aqueous salt solutions for the titration of neutral bases.[4] Also, the portion of the curve beyond the end point is shifted in the acid direction in water, whereas here it is shifted in the basic direction, only slightly at first and then more greatly.

Similar results were found for the titration of aniline in acetone containing magnesium perchlorate or lithium iodide, except that the maximum break obtained was somewhat less than with lithium perchlorate. The same was true for titrations of aniline in 2-butanone or methylisobutylketone containing lithium perchlorate.

The perchloric acid titrant in acetone used for these systems was fairly unstable, but if the perchloric acid was dissolved in dioxane, it was stable. The latter titrant gave similar results to the

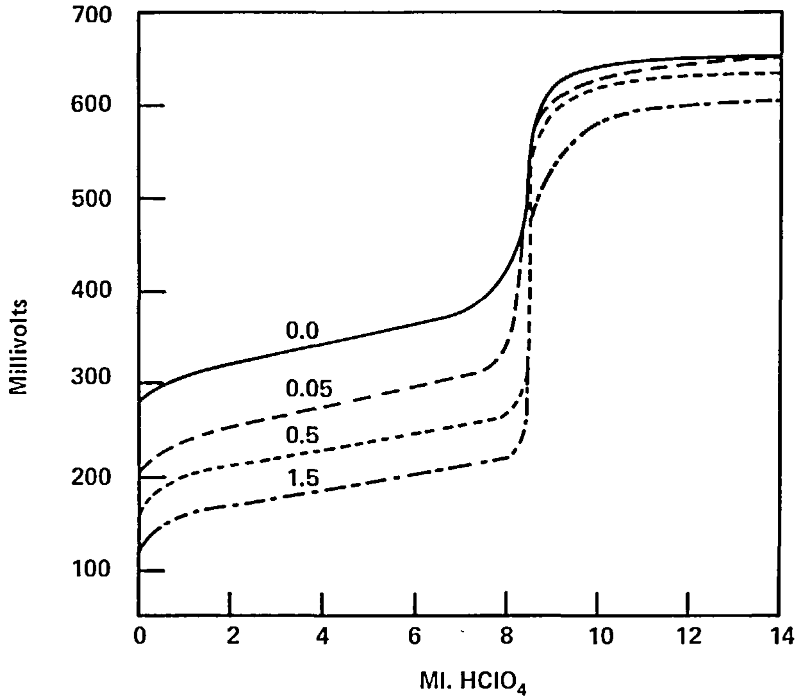


FIGURE 17. Effect of lithium perchlorate concentration on potentiometric titration of aniline in 50 ml of acetone solvent. Numbers on the curves represent molarity of salt. Titrant is 0.1 *M* $HClO_4$ in acetone. (From Schertz, W. L. and Christian, G. D., *Anal. Chem.*, 44, 755 (1972). With permission.)

former, except that at very high salt concentrations, the potentiometric break was diminished more greatly as the acid end of the curve was shifted toward more basic potentials. This may have been due to a higher water content in the dioxane than in the acetone.[29]

The water content had a marked effect on the potentiometric end point breaks in titrations of the weak bases. (The magnitude of the break was determined by extending a line from the linear portion of the titration curve before the end point. A line parallel to this was drawn tangent to the upper portion of the curve, and the distance between these two parallel lines was taken as the magnitude of the potentiometric end point break.) In the titration of aniline in acetone with added water, the end point breaks were: 0% water, 235 mV; 1% water, 130 mV; 2% water, 80 mV; 3% water, 60 mV. Only the portion of the curve beyond the end point was affected by the water. Hence, it is possible that the addition of concentrated salts effectively removes traces of water by hydration, thereby contributing to the potential break enhancement.

Some water (about 0.04 g[29]) is added with the titrant. It is possible that if titrations could be performed under strictly anhydrous conditions, a somewhat larger end point break could be obtained. Preparation and storage of anhydrous perchloric acid by conventional means[71] is somewhat dangerous, and addition to an organic solvent is not advised. Coetzee and Bertozzi,[72] however, were able to prepare dilute solutions of anhydrous perchloric acid in sulfalone by bubbling hydrochloric acid gas through a sulfuric acid scrubber, then into sulfalone. This solution was then titrated potentiometrically under nitrogen with anhydrous silver perchlorate in sulfalone to precipitate silver chloride and leave anhydrous perchloric acid. This titrant solution was stable for a year. Also, Pocker and Buchholz[72a] prepared anhydrous perchloric acid in diethyl ether by titrating dried silver perchlorate with a slight excess of hydrochloric acid in ether, followed by filtration of the silver chloride precipitate.

The titration of other bases with different strengths in acetone containing lithium perchlorate was qualitatively similar to the titration of aniline in that the end point break was enhanced to a maximum value and then diminished at higher salt concentrations.[29] In general, the weaker the base, the greater the salt enhancement, but the more salt

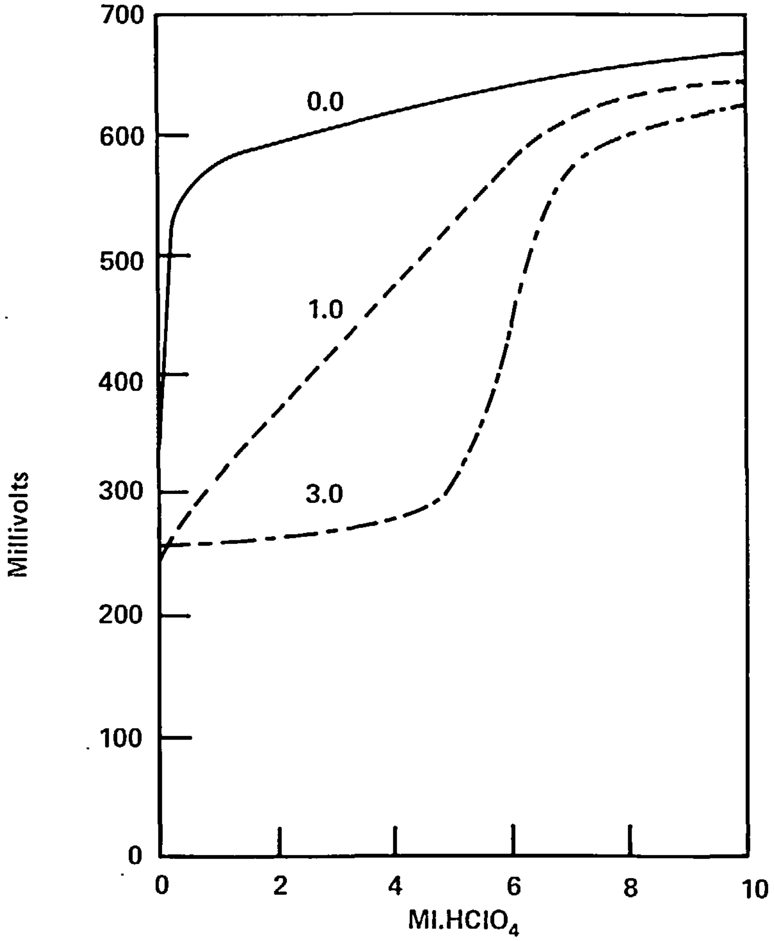


FIGURE 18. Effect of lithium perchlorate concentration on potentiometric titration of caffeine in 50 ml of acetone solvent. Numbers on the curves represent molarity of salt. Titrant is 0.1 *M* $HClO_4$ in acetone. (From Schertz, W. L. and Christian, G. D., *Anal. Chem.*, 44, 755 (1972). With permission.)

required to achieve the maximum break. Figure 18 illustrates the dramatic effect lithium perchlorate has on the titration of caffeine in acetone. Table 13 summarizes the maximum potentiometric breaks obtained (defined as above) and the lithium perchlorate concentrations for the titration of bases of different strengths in acetone, and compares them with results in the absence of salt, and with titrations in glacial acetic acid. Some interesting trends can be noted from the data. First, except for the weakest bases, all the bases give approximately the same magnitude of maximum end point break. Also, the base end of their titration curves are shifted to approximately the same potential region for the maximum break. Caffeine does not quite give the same break as the other bases and bases which are weaker than caffeine are too weak to be titrated in salt solutions of acetone. The acetone-lithium perchlorate system for titrating weak bases provides a definite advantage over the usual titrations in the leveling solvent glacial acetic acid. There appears to be no advantage, though, for the titration of bases as strong as pyridine with a pK_b of 8.81 in water. It would be interesting, in view of the widely differing behavior of charged bases in aqueous salt solutions,[5] to study their behavior in nonaqueous solvents.

Attempts were made to apply the salt enhancing system to titrations in glacial acetic acid because of its advantageous properties compared to acetone in the absence of salt.[29] Figure 19 shows the titration of aniline in acetic acid both in the presence and absence of added lithium perchlorate. Unfortunately, the end point break is actually suppressed rather than enhanced. The

TABLE 13

Maximum Potentiometric Breaks Obtained for Weak Bases Titrated in Solutions of Acetone Containing Lithium Perchlorate

Base	pK_b	mV break in acetic acid	mV break in acetone	mV break in acetone-$LiClO_4$	Break enhancement, mV	*M* of $LiClO_4$
Urea	13.80	–	none	20	(20)[a]	4.00
Caffeine	13.39	50	none	315	315	3.00
m-Nitroaniline	11.40	155	70	350	280	1.00
Aniline	9.42	265	235	360	125	1.00
Pyridine	8.81	370	250	340	90	1.50
N-Dimethylaniline	8.79	–	185	365	(180)[a]	0.50

[a] Titrant in dioxane.

From Schertz, W. L. and Christian, G. D., *Anal. Chem.*, 44, 755, (1972). With permission.

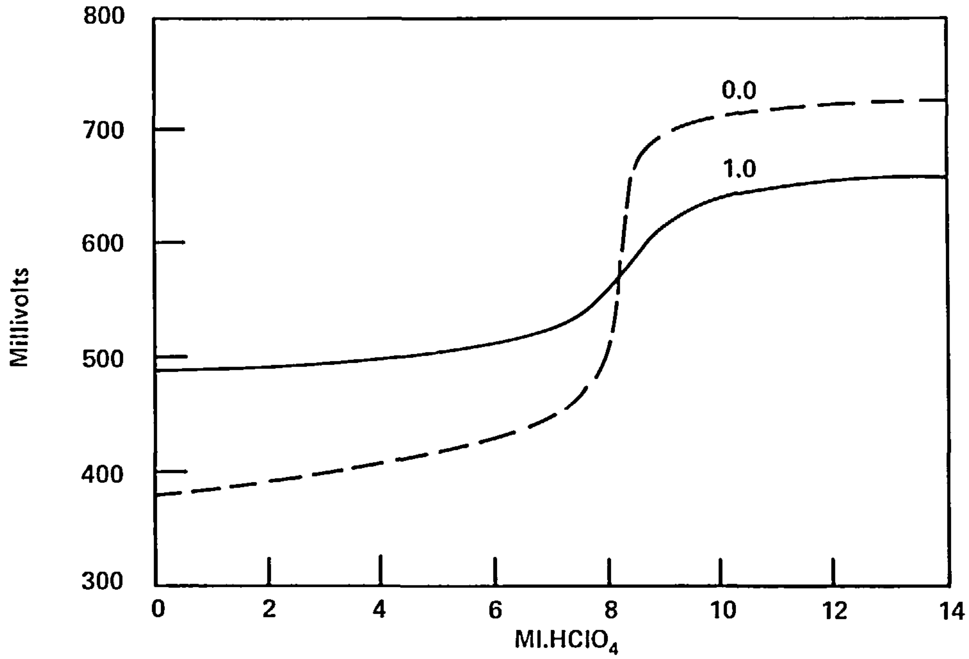


FIGURE 19. Effect of lithium perchlorate on potentiometric titration of aniline in 40 ml of glacial acetic acid solvent. Numbers on the curves represent molarity of salt. Titrant is 0.1 *M* $HClO_4$ in glacial acetic acid. (From Schertz, W. L. and Christian, G. D., *Anal. Chem.*, 44, 755 (1972). With permission.)

reasons for this are unclear, but they agree with results for the titration of other bases in glacial acetic acid in the presence of 0–0.6 *M* sodium perchlorate, 0–0.6 *M* N,N-dimethyl-anilinium perchlorate, or 0–0.01 *M* pyridinium perchlorate using perchloric acid in dioxane as the titrant.[73] Kolling[74,75] investigated the effects of sodium and potassium salts on acid-base indicators in glacial acetic acid and found that the results could be represented by the following equilibrium:

$$B^+A^- + HOAc + I_B \rightleftharpoons B^+OAc^- + I_BHA$$

where I_B and I_BHA are the base and acid forms of the indicator. He concluded that the solvolytic reaction on the salt ion pair to form detectable concentrations of the conjugate acid and base is the principal influence of added salts on the color ratio of the indicators.

Visual indicators can be used in the acetone-lithium perchlorate system. Methyl violet and orange IV gave suitable color changes at the equivalence point in the titration of aniline with 0.1 *N* perchloric acid in dioxane.[29] Thymol blue, methyl red, and malachite green were unsatisfactory.

An unambiguous mechanism for the enhancement of the titrations of weak bases in the ketone solvents is not available. The fact that several bases exhibited similar initial potential readings and maximum end point breaks suggests that they are somehow leveled to similar strengths. It is possible the salt acts as a Lewis acid to level the bases.[29] In this respect, it is noteworthy that a number of metal ions can be titrated as acids in nonaqueous solvents,[76-78] including lithium and sodium.[79] In a leveling mechanism, the weaker the base, the more salt would be required to achieve maximum leveling. This leveling seems to occur for the four strongest bases studied but appears to be incomplete for the very weak bases caffeine and urea. The shift of the curve beyond the end point could be due to water taken up by the salt during the titration since the salts are hygroscopic; higher salt concentrations would be expected to absorb water more rapidly. The high concentrations of salts would, of course, also have an effect on the activity coefficients of the various species in solution, which would be expected to generally increase the degree of ionization of the bases. In water, this effect is relatively small compared to other effects. Finally, it seems probable that ion-pair formation with the salt could have an appreciable effect on the acid-base equilibria in these solvents.

Ion pairs are known to occur in glacial acetic acid[80] and other solvents. It is probable that a protonated base could form ion pairs with the anion of the added salt in acetone, for example, perchlorate ion.[81] If such were to occur, then the base would be made to appear stronger based upon the equilibria involved:

$$BH^+ + ClO_4^- \rightleftharpoons BH^+ClO_4^- \qquad (23)$$

$$B + H^+ \rightleftharpoons BH^+ \qquad (24)$$

If the formation constant for Reaction 23 and/or the perchlorate concentration were sufficiently large that the reaction went at least partially to completion, then Reaction 24 would be shifted to the right. For a charged base, such as acetate ion, ion pairing (e.g., with lithium ion) would be expected to have an opposite effect, i.e., shift the equilibrium, $OAc^- + H^+ \rightleftharpoons HOAc$, to the left and cause the acetate ion to appear to be a weaker base. In view of this, as well as the results obtained for the titration of charged bases in water[5] (see Section II.C.1.), it would be revealing to investigate the effect of salts on the titration of charged bases in nonaqueous solvents.

Finally, it should be noted that Pocker and Buchholz,[72a] in a study of the effect of lithium perchlorate on the ionization of triphenylmethyl chloride and of hydrochloric acid in diethyl ether, found a large increase in the dissociation constants in the presence of the salt. For hydrochloric acid, the dissociation constant increased from 5.35×10^{-5} in 0.478 *M* lithium perchlorate to 2.41 in 5.05 *M* lithium perchlorate. The ionization rate constants also increased markedly. These results imply that in diethyl ether, hydrochloric acid becomes a stronger acid in the presence of lithium perchlorate.

2. *Titration of Weak Acids*

Marple and Fritz[82] obtained an enhancement in end point breaks in the titration of 2,6-di-*tert*-butylphenol in butyl alcohol by adding tetrabutylammonium bromide. Typical titration curves are shown in Figure 20. The entire curve is generally shifted toward more acid potentials, but the portion before the end point is shifted more than that past the end point. This is similar to the results of Schertz and Christian[29] for the titration of weak bases except that the shift is in the opposite direction. Note that a relatively small amount of salt has considerable effect on the shape of the titration curve (curve D vs. curve E). Marple and Fritz[82] attribute this initial change (curve D) primarily to the increased ionic strength of the solution and the effect on the dissociation of the 2,6-di-*tert*-butylphenol.

The end point enhancement was explained on the basis of an increase in the acidity of the solution of acid, HX, and its salt, Bu_4NX, relative to a dilute solution of tetrabutylammonium hydroxide, owing to the incomplete dissociation of the salt of the acid. Consider the equilibria involved in the titration:

$$Bu_4NOH \rightleftharpoons Bu_4N^+ + OH^- \qquad K_1 \qquad (25)$$

$$Bu_4NX \rightleftharpoons Bu_4N^+ + X^- \qquad K_2 \qquad (26)$$

$$HX \rightleftharpoons H^+ + X^- \qquad K_3 \qquad (27)$$

X^- is the anion of the acid being titrated in both Reactions 26 and 27. An increase in Bu_4N^+ by adding Bu_4NBr has a relatively small effect on the hydroxide concentration because K_1 is fairly large. But if $K_1 >> K_2$ and K_1 is at least $10^{-4}-10^{-3}$, the added salt will increase the acidity of HX relative to that of a dilute base solution because Reaction 26 will be driven to the left to a greater extent than Reaction 25. In aqueous solutions, the salt of the titrated acid is normally completely ionized (Reaction 26). But the dissociation of most tetrabutylammonium salts in *tert*-butyl alcohol is not extensive because of the low dielectric constant of the solvent. Most have dissociation constants in the range of 10^{-4} to 10^{-6}.[83] Hence, the ratio of $[X^-]/[HX]$ (in $pH \approx pK_a + \log[X^-]/[HX]$) is decreased in the presence of added Bu_4N^+ as Reaction 26 is shifted to the left and Reaction 25 is shifted to the right in the buffer region of the titration curve (in the presence of decreased X^- from Reaction 26). The portion of the titration curve beyond the end point would still be slightly less alkaline than in the absence of the added salt because Reaction 25 would be suppressed somewhat, although to a less extent than Reaction 26.

Only slight enhancement of end point breaks was obtained for the titration of other weak acids (succinimide, ethyl acetoacetate) and none for acetoacetanilide. Hence, this technique is useful only in cases where the acid is just barely titratable in pure *tert*-butyl alcohol (whose conjugate bases are essentially undissociated).

Smagowski and Jasinski[73] titrated acetic acid

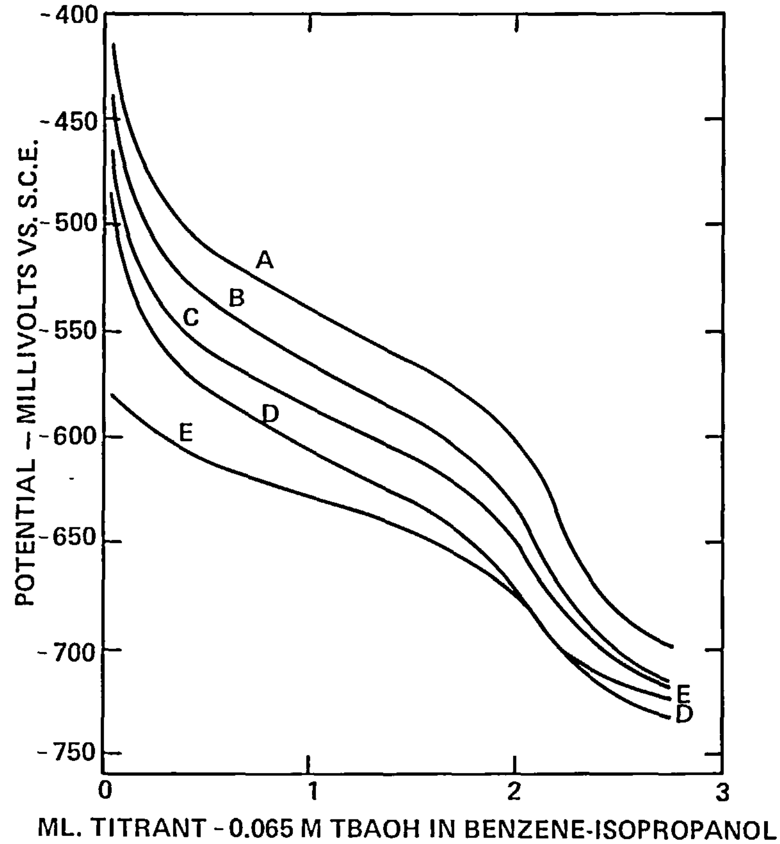


FIGURE 20. Effects of tetrabutylammonium bromide on the titration of 2,6-di-*tert*-butylphenol. A. 0.33 *M* Bu_4NBr at start of titration. B. 0.166 *M*. C. 0.066 *M*. D. 0.033 *M*. E. No added bromide. (From Marple, L. W. and Fritz, J. S., *Anal. Chem.*, 35, 1431 (1963). With permission.)

and benzoic acid in pyridine in the presence of 0–0.1 *M* tetraethylammonium perchlorate or 0–0.1 *M* tetraethylammonium chloride using tetrapropylammonium hydroxide titrant, and in the presence of 0–0.3 *M* sodium perchlorate using sodium methoxide titrant. As in the case of the acidic solvent, glacial acetic acid, titration curves in this basic solvent are generally suppressed by the added salts. It appears that advantageous salt effects occur primarily in neutral solvents only.

Wharton[84] described a unique and dramatic salt effect in the titration of weak acids, particularly fatty acids, in the presence of lithium chloride in acetonitrile. He used as the titrant, tri-*n*-propylamine. This is a weak base titrant and allows the titration of the free carboxylic acid content without titration of the anhydrides. Figure 21 shows the titration curves of different acids in the presence and absence of saturated (0.034–0.036*M*) lithium chloride. The end point breaks obtained with acids of varying strengths are summarized in Table 14. In the absence of lithium chloride, only acids with pK_a values less than 3.13 could be titrated, but in the presence of the salt, monocarboxylic acids with pK_a values as large as 5.5 (picolinic acid) could be titrated. Also, all other acidic groups with pK_a's up to 11.9 in a molecule could be titrated, provided one acidic group of the molecule had a pK_a less than 5.5. The acid groups were not differentiated, though, but titrated together. For example, the third proton of phosphoric acid could be titrated along with the second proton. An exception was maleic acid (pK_1 = 1.92, pK_2 = 6.23) in which the two acid groups were titrated separately.

Separate acids with pK_a's differing by at least four could be differentiated if the stabilities of their lithium salts were comparable or if the lithium salt of the weaker acid were less dissociated (see mechanism below). In addition, mixtures of monocarboxylic acids could be

TABLE 14

Typical End Point Potential Breaks ($\Delta E_{e.p.}$) for Titration with Tri-*n*-propylamine in Acetonitrile Containing LiCl

Acid	pK_a (H_2O)	$\Delta E_{e.p.}$, mv	Number of equivalents titrated
Oxalic	1.27, 4.27	437[a]	2[b]
Malonic	2.86, 5.70	255[a]	2
Adipic	4.41, 5.28	280[a]	2
Sebacic	4.55, 5.52	415[a]	2
Citric	3.13, 4.76, 6.40	310[a]	3[b]
Lactic	3.86	270[a]	1
Phenylphosphonic	1.83, 7.07	475[a]	2[b]
Nicotinic	4.87	60[c]	1
Picolinic	5.50	185[c]	1
Benzoic	4.17	95[a]	1
H_3PO_4	2.1, 7.2, 11.9	385[c]	3[b]
Acetic	4.76	75[a] (55[c])	1
Hexanoic (caproic)	4.85	120[a] (110[c])	1
Dodecanoic (lauric)		230[a] (190[c])	1
Octadecanoic (stearic)		290[a] (250[c])	1

[a] Glass-calomel (saturated NaCl in CH_3OH) electrode pair.
[b] 1 equivalent titrated in absence of LiCl.
[c] Glass-calomel (0.1 *M* LiCl in CH_3OH) electrode pair.

From Wharton, H. W., *Anal. Chem.*, 37, 730 (1965). With permission.

differentiated if their chain lengths differed by at least six carbon atoms.

Note in Figure 21 that the portion of the acidity of the sample (before the end point) is apparently increased in the presence of lithium chloride and that the increase is larger the longer the carbon chain. Wharton[84] explained the mechanism of enhancement by a simple metathesis reaction:

$$HA + LiCl \rightleftharpoons LiA + HCl$$

The hydrochloric acid becomes the titratable species. The equilibrium is shifted to the right as the stability of the lithium salt of the acid increases (or its solubility decreases). The longer chain fatty acids give more stable salts. Short chain acids required considerable excess of lithium chloride before all the acid present could be titrated, indicating a relatively high value for the dissociation constant of LiA. But long chain (C_{18}) acids required only a slight excess of salt over that needed for stoichiometry.

Similar titrations could be performed by adding a methanol solution of calcium chloride (since anhydrous calcium chloride is insoluble in acetonitrile). In this case, less salt was required and larger end point breaks were obtained for the

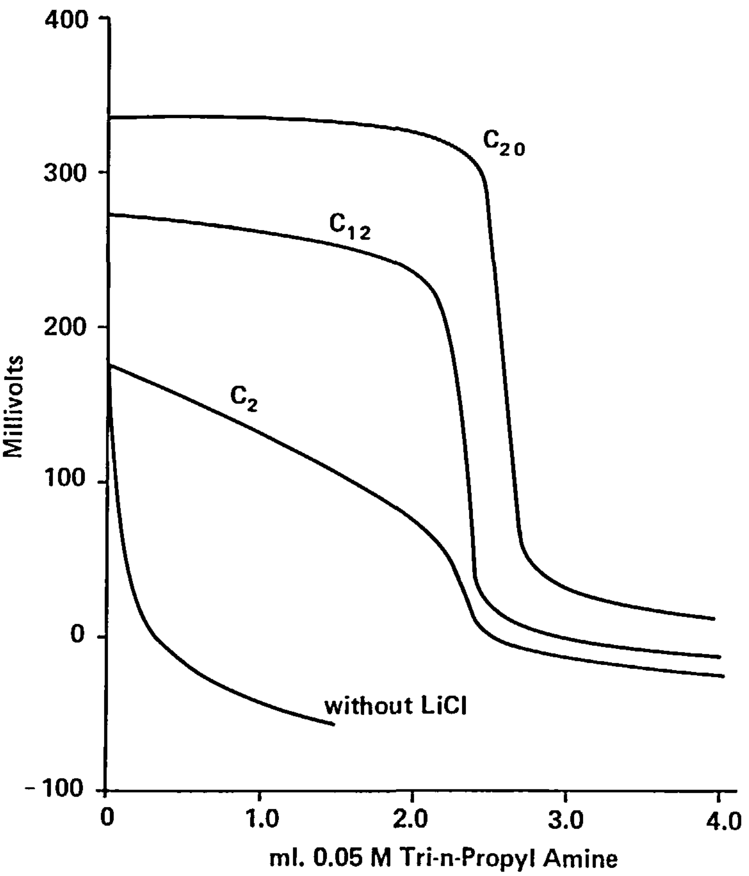


FIGURE 21. Titration of typical monocarboxylic acids in acetonitrile with and without LiCl. Tri-*n*-propylamine is the titrant. C_2, C_{12}, and C_{20} indicate the number of carbons. (From Wharton, H. W., *Anal. Chem.*, 37, 730 (1965). With permission.)

shorter chain acids, indicating the dissociation constants for the calcium carboxylates in acetonitrile are smaller than those of the corresponding lithium carboxylates. The methanol added with the salt, however, precluded titrations in the presence of anhydrides. A variety of other solvent and salt systems were investigated,[84] but the lithium chloride-acetonitrile system was generally the most satisfactory. Other lithium salts apparently have smaller dissociation than lithium chloride in the nonaqueous solvents.

Juillard[85] has reported the existence of hydrogen bonds between anions and acid molecules of acids dissolved in dimethylformamide. He used the Debye-Hückel limiting law to obtain a correct description of the salt effect in this solvent by taking into account the "imperfect nature" of the dissociation of the salt. A classification of the strengths of different carboxylic acids and phenols is proposed.

The basicity of strong bases (e.g., alkali metal alkoxides) in alcoholic solutions has been reviewed by Bowden.[19] For dilute (0.1*M*) solutions, the basicity increases in the order: methanol ($H_- = 12.66$) $<$ ethanol ($H_- = 14.57$) $<$ isopropyl alcohol ($H_- = 16.95$) $<$ *t*-pentyl alcohol ($H_- = 18.09$) $<$ *t*-butyl alcohol ($H_- = 19.14$). The basicity of a methanolic solution is comparable to that of an aqueous solution of the same base concentration. H_- is related to the base concentration up to 0.1*M* by

$$H_- = B + \log[OR^-] \qquad (28)$$

where B is a constant dependent on the nature of alcohol.[86] The basicity of an alcoholic solution appears to be a function of the autoprolysis constant and the proton affinity of the solvent.

At high base concentrations, as in water, the basicity increases more rapidly than predicted by

Equation 28. The stable alkoxide ion would be expected to be solvated, similar to the hydroxide ion, to give a structure like that in I, II, or III. The first and third would result in $OR^-(ROH)_3$ and the second $OR^-(ROH)_2$. At high alkoxide concentrations, insufficient solvent would be available to fully solvate the ions, and a sharp rise in the activity of the base would be expected. This has been observed.[87-89] The deviation from Equation 28 is much greater than in the case of water because of the smaller number of moles of alcohol per liter; i.e., the solvating species are more quickly exhausted. It is interesting that the stable complex, *t*-BuOK · *t*-BuOH, has been isolated,[90] suggesting the existence of solvent-bridged ion pairs similar to those suggested for water above. The kinetic basicity of solutions containing alkali metal *t*-butoxide in dimethysulfoxide is in the order Li < Na < K,[19] similar to the order of activity coefficients of hydroxide ion in aqueous alkali metal solutions (Li < Na < K < Cs [see Section II.C.2.]). It seems likely, then, that potassium or cesium *t*-butoxide might be a good titrant for weak acids in nonaqueous solvents.

Mixed solvents containing water, a dipolar aprotic solvent, and added base such as a tetraalkylammonium hydroxide behave similarly to the aqueous and alcohol systems, but the effects are greatly enhanced.[19] As the dipolar aprotic solvent content is increased, H_- rises considerably, with the rates of increase being in the order: pyridine < tetramethylene sulfone < dimethyl sulfoxide. Apparently, the water forms hydrogen bonds with the aprotic solvent, and the solvent effectively competes with hydroxide ion in this way.[91] A solution of sodium methoxide in methanolic dimethylsulfoxide is less basic than a similar sodium hydroxide solution in aqueous dimethylsulfoxide because of the more stable structure formed between water and dimethylsulfoxide and the intrinsically greater basicity of the hydroxide ion.

III. OXIDATION-REDUCTION REACTIONS

No direct studies have been made on possible effects of concentrated salts on redox titrations. Salts do have an effect on the rates of oxidation-reduction reactions, though, and it is pertinent to summarize some of the studies that have been made. Although equilbrium effects are of more interest in titrations, kinetic effects may be of interest in titrations that are kinetically slow.

Schmid and Bähr[92] noted an increase in the velocity constant for the reaction

$$HNO_3 + 2NO + H_2O \rightarrow 3HNO_2$$

upon adding sodium perchlorate, similar to the results obtained when the nitric acid concentration was increased above 1 *m*. They concluded that the increase lies with "the usual neutral salt effect in concentrated electrolyte solutions and follows the same empirical laws." Jaksic et al.[93] studied the effects of neutral salts on the conversion of hypochlorous acid and hypochlorite to chlorate:

$$2HOCl + OCl^- + 2H_2O \rightarrow ClO_3^- + 2H_3O^+ + 2Cl^-$$

or

$$HOCl + 2OCl^- + H_2O \rightarrow ClO_3^- + H_3O^+ + 2Cl^-$$

The expected pH for maximum conversion was 7.0–7.5, while the experimental value in the presence of 5*M* sodium chloride was 5.8–6.2. The difference from the expected values was explained in terms of the influence of the salt on the activity coefficients of the hydronium ion, hypochlorous acid, and hypochlorite ion, and, in turn, on the dissociation constant of hypochlorous acid since it was found that 5*M* sodium chloride shifted the pH of buffer solutions about one unit towards lower values. Jaksic et al.[93] also noted that at the same time, salts with positive heat of hydration caused dehydration and an increase in the hydronium ion acitvity.

Wrónska and Banas[94] obverved a strong negative salt effect on the rate of iron (II) oxidation by chlorate in acidic media which was very dependent on the anion of the salt. In general, the rate initially decreased with increasing salt concentration and then, with an exception of potassium and sodium sulfates, the rate increased. The initial inhibition decreased in the order: $SO_4^{-2} > Cl^- > ClO_4^-$ and $Na^+ > K^+ \approx Mg^{+2} > Ca^{+2} > Ba^{+2}$. A close correlation was observed between the activity coefficients of these salts and their inhibition of the oxidation. Hence, it appears that the effect may be largely one of affecting the activities of the reacting species. Valiente[95] found a negative salt effect on the velocity constant of the chlorate-sulfite reaction for several salts, especially at high sodium chloride concentrations.

Aluminum chloride, however, had the opposite effect.

The reaction of bromide with bromate is increased in high concentrations of chloride salts in the order: $Mg^{+2} > Ca^{+2} > Ba^{+2} > Na^{+} > K^{+}$.[96] The reaction accelerating action is due to the positive catalysis of chloride ions and not the salt effect. Nitrates and expecially sulfates cause a decrease in the reaction rate, the latter probably due to bisulfate ion formation.

Alkali and alkaline earth metal halides and thiocyanates inhibit the oxidation of ascorbic acid by oxygen.[97] The reaction of 0.01*M* ascorbic acid in 0.1*M* acetate buffer at pH 5.5 is suppressed to 5% by 0.01*M* salts. The anions of the salts (iodide, chloride, thiocyanate) were considered to have electron-donor properties, which suggested to the authors that free radical intermediates were stabilized by the salts by adding an electron to the radicals. On the other hand, the oxygen uptake increases more than 150% by transition metal cations (manganese (II), iron (II), cobalt (II), copper (II)).

Oxidation-reduction equilibria have been characterized by the concept of "electron activity".[98-101] However, this concept does not represent activity in the usual sense but, instead, is just an expression of the electrical potential, E. It is defined as $-\log e^- = pE = E/(RTF^{-1} \ln 10) \approx E/0.060$, where e^- is the electron activity. pE is large and positive in strongly oxidizing solutions (low electron activity), just as pH is high in strongly alkaline solutions (low proton activity). That is, an oxidizing solution prefers to accept electrons rather than release them in reaction and so has a low electron activity.

It would be interesting to consider the salt effect if electron "activity" in the usual sense were defined. In particular, a solvated electron might enter into competition with salt ions for solvent molecules. Kenney and Walker[101a] have summarized evidence of the existence of the hydrated electron along with proposals for its structure. The salt molecules probably would have little success in such a competition, however, in view of the high charge density on the electron. Taube[102] has summarized concepts of the oxidation-reduction reaction mechanisms. Since many inorganic ions are hydrated, and the waters of hydration, as well as other ligands, appear to be involved in bridging mechanisms of electron transfer, concentrated salts would undoubtedly affect rates of redox reactions involving such ions.

Redox titrations have been performed satisfactorily in nonaqueous solvents[103] and electron-exchange reactions have been studied in nonaqueous media[104] where the electron is in a different solvent medium compared to in water. Reactions in concentrated aqueous salt solutions (dehydrated) might be compared to such media. Masuko[61c] has described the effect of basicity in fused salt solutions on redox reactions and Landresse[104a] has reviewed redox reactions in molten salts.

Most oxidation-reduction reactions involved in titrations are rapid and so there would be little salt effect expected. The enhanced activity of the proton in aqueous concentrated salt solutions would undoubtedly alter the rates of reactions that involve protons if they could be measured (by fact reaction kinetics), and certainly would noticeably alter reaction rates of kinetically slow reactions. It would be interesting to study the effects of concentrated salts on reactions of oxidizing and reducing agents containing oxygen that are reduced or oxidized irreversibly and require a catalyst to assure successful titrations.[105] An example is the oxidation of arsenic (III) by cerium (IV), which is catalyzed by osmium tetroxide.

IV. PRECIPITATION TITRATIONS

Little work has been reported on the effects of salts on precipitation titrations, but such titrations represent the earliest application of salts for improvement of titrations. Mixtures of chloride and bromide or of bromide and iodide can not generally be titrated accurately because of solid solution formation and adsorption phenomena.[106] In 1922, Foerster[107] stated that in the presence of barium nitrate, iodide in a ratio of iodide: bromide > 1:3,600 can be estimated accurately, but bromide in chloride only when bromide:chloride – 1:10. Müller[108] described the addition of ions which are strongly adsorbed by silver halides to limit this error. He writes:

> "The accuracy of estimation of iodide in the presence of bromide and of iodide in the presence of chloride is good in all cases in the presence of 5% barium nitrate. Even the estimation of small amounts of one halide in presence of excess of the other is not affected ... In the estimation of bromide and chloride in the presence of one another, we

must, unfortunately, reckon on an error of 1% if too much bromide and correspondingly too little chloride, if equal molar concentrations of both are taken. This error increases with increasing excess of chloride, and falls off with increasing bromide excess... until it can be neglected when the molar concentration of bromide exceeds that of chloride by five times."

Müller also mentioned potassium alum in the same connection as barium nitrate, but gave no data concerning its action. Others confirmed Müller's observations that correct end points are obtained in the presence of 5% barium nitrate solution. Lathermoser[109] used ammonium nitrate to remove adsorbed silver from silver halides, and Fajans and co-workers[110] studied the adsorption of dyes and adsorption replacement by dyes on silver halides.

Clark[106] extended the studies of Fajan and Müller in improving the potentiometric end points for the titration of halide mixtures. A mixture of 5.0 ml of a 0.1 *N* solution of each of the halides in the presence of 50 ml of the solution of the substance tested was titrated with 0.1 *N* silver nitrate. Erythrosin (1 in 2,000), fluorescein (1 in 2,000), and rose Bengal 3 (1 in 2,000) gave correct end points with the bromide-iodide mixtures. In the case of the chloride-bromide mixture, fluorescein (1 in 2,000) gave a partial correction. Correct end points for both chloride-bromide and bromide-iodide mixtures were obtained with 5% solutions of either barium nitrate, ammonium nitrate, or sodium acetate. These salts gave correct end points for a wide variation in the proportions of the halides. They also shortened the titration by increasing the sensitivity of the titration and decreasing the time required for equilibrium potential in the region of the end points. In this respect, the presence of the salts is beneficial for the titration of even single halides. The results with the sodium acetate demonstrate the effect is not due to the nitrate ion, and this was confirmed by the fact that 1 *N* sulfuric acid gave practically complete correction. One tenth normal nitric acid also gave correct end points which were very sensitive. In the absence of added salts, the silver halides formed remain in colloidal suspension until the final end point is reached, but in the presence of salt, they begin to precipitate shortly after the first addition of silver nitrate.

Clark was able to titrate a mixture of these halides in the presence of added salt. Twenty ml of a mixture containing 3% iodide and 1% chloride in 0.1 *N* bromide (which represents possible proportions in a photographic preparation) plus 200 ml of 5% barium nitrate solution were titrated accurately.

De la Mare et al.[111] used Clark's method (5% barium nitrate) to titrate mixtures of bromide and chloride.

V. SUMMARY

The addition of salts to solutions can cause marked changes in titration curves, particularly in acid-base titrations. The effects can be summarized as follows:

1. Titration of weak bases in aqueous solutions. The activity of the proton is enhanced at least 100-fold in the presence of concentrated neutral salts, probably due to partial dehydration, which allows the titration of uncharged weak bases ($K_b \approx 10^{-12}$) with mineral acids.[4,5,27] Perchlorate salts appear to give abnormally large effects when their properties are compared with halide salts. The enhancing effect is less for negatively charged bases, primarily because the base is made to appear less alkaline.

2. Titration of weak acids in aqueous solutions. The enhancing effect by salts is less general than in the above case. This is apparently because the activity of the hydroxide ion is not increased as in the case of the hydrogen ion, although there is some evidence for increased activity under certain conditions. Leithe[62] was able to titrate several weak acids in the presence of halide or nitrate solutions. Free acid can be titrated in the presence of hydrolyzable metal ion in 10 *M* lithium chloride.[64] The hydrolyzable cations can be titrated separately or in mixture. The salt appears to play a salt dehydrating role, decreasing the degree of hydrolysis of the metal ion.

3. Titrations of weak bases in nonaqueous solutions. Very weak uncharged bases ($K_b < 10^{-13}$) can be titrated in ketones (acetone) in the presence of lithium perchlorate.[29,59] The acidity of the titrant does not appear to be increased but rather the bases appear to be leveled to stronger bases. Small amounts of water decrease the sharpness of the end point by causing the acid end of the curve to shift toward more alkaline values. Charged bases have not been titrated. End points

of titrations in glacial acetic acid are made less sharp in the presence of salts.

4. Titration of weak acids in nonaqueous solutions. Two special cases of favorable salt effects have been reported. The end point for the titration of 2,6-di-*tert*-butylphenol in butyl alcohol is enhanced by relatively small concentrations of tetrabutylammonium bromide.[82] The entire titration curve is shifted toward more acid potentials, but the portion beyond the end point is shifted more than that past the end point. The enhancement is explained on the basis of an increase in the acidity of the solution of acid, HX, and its salt, Bu_4NX, relative to a dilute solution of tetrabutylammonium hydroxide, owing to incomplete dissociation of the salt of the acid. The effect only occurs in cases where the acid is just barely titratable in pure *tert*-butyl alcohol (whose conjugate bases are essentially undissociated).

The second case is the titration of fatty acids with tri-*n*-propylamine (a weak base titrant that does not react with the anhydrides) in the presence of lithium chloride in acetonitrile.[84] Acids that are untitratable in acetonitrile become titratable in the presence of the salt by a simple metathesis reaction: $HA + LiCl = LiA + HCl$. The equilibrium is shifted to the right as the stability of the lithium salt increases and the hydrochloric acid becomes the titratable species. Longer chain acids form more stable salts and require less excess salt than the shorter chain acids for titration.

5. Oxidation-reduction reactions. Salts affect the rates of redox reactions, but no advantage is reported for redox titrations, probably because of the rapidity of reactions. It is possible that some reactions consuming protons may be aided in the presence of salts by virtue of the increased activity of the proton.

6. Precipitation titrations. Mixtures of chloride and bromide, bromide and iodide, or chloride, bromide, and iodide can be accurately titrated in the presence of 5% barium nitrate and other salts.

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