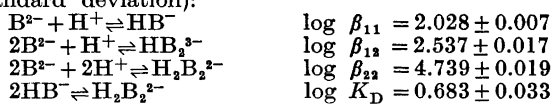


The Association of 3-Bromo-5-sulphosalicylic Acid in Aqueous Solution

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The pH equilibria of 3-bromo-5-sulphosalicylic acid have been studied, at 25°C, in acidic solution and 3 M sodium (perchlorate) medium, using a glass electrode. The total concentration, B , of disodium 3-bromo-5-sulphosalicylate, Na_2B , ranged from 0.0050 M to 0.2000 M. The data are best explained by the following equilibria and corresponding formation constants (given with 3σ , where σ is the standard deviation):



The preliminary constants were obtained by Sillén's "curve-fitting" method and refined using the generalized least squares program LETAGROP, both E_0 for the electrode and the analytical concentrations H and B being adjusted.

The equilibrium constants of dimerization, K_{12} and K_D , and the proton association constants, K_{11} and K_{22} , are critically compared with those of salicylic acid.

The association of salicylic acid in aqueous solution has been investigated previously.¹ The dimers H_2B_2 and HB_2^- were found to coexist with salicylic acid, HB , in aqueous solution, when the total concentration of salicylate ion, B , was greater than 0.01 M. Due to the slight solubility of salicylic acid, the dimers H_2B_2 and HB_2^- exist only in minor amounts in aqueous solution. It is, therefore, difficult to obtain further information about their structures in solution by X-ray or spectroscopic methods.

A derivative of salicylic acid, 3-bromo-5-sulphosalicylic acid, was chosen for the present investigation. The larger solubility (compared with salicylic acid) of its monosalt in water and the rather high atomic number of Br, ought to make further investigations on the structures of the dimers in aqueous solution possible. In the present work only their compositions and formation constants have been investigated. Experiments on the diffraction of X-rays by the equilibrium solution of the disodium and monosodium salts of 3-bromo-

5-sulphosalicylic acid are, however, in progress. Relatively definite information concerning the structures of the dimers in solution is expected and will be reported in the future.

The experiments and calculations in the present work are similar to those in the previous paper.¹ The most important symbols are listed below for reference:

B	total concentration of disodium 3-bromo-5-sulphosalicylate, Na_2B .
b	concentration of free B^{2-} ion.
$[\text{HB}^-]$	concentration of monosodium 3-bromo-5-sulphosalicylate
$[\text{HB}_2^{3-}]$	concentration of the dimer HB_2^{3-} .
$[\text{H}_2\text{B}_2^{2-}]$	concentration of the dimer $\text{H}_2\text{B}_2^{2-}$.
β_{pq}	equilibrium constant for the reaction $p\text{H}^+ + q\text{B}^{2-} \rightleftharpoons \text{H}_p\text{B}_q^{(p-2q)}$.
Z	average number of H^+ bound per B^{2-} .
E	emf in mV.

EXPERIMENTAL

Chemicals. Sodium perchlorate and perchloric acid solutions were prepared and analysed as described in a previous paper.¹

Disodium 3-bromo-5-sulphosalicylate, Na_2B , was prepared from the corresponding monosodium salt in the following way. The monosodium salt * was recrystallized three times from hot distilled water and neutralized to the disodium salt by adding sodium carbonate A. R. in hot solution. The sodium ion concentration was determined by an ion exchange technique.² The concentration of Na_2B was then calculated as one half of the value obtained for the sodium ion concentration.

Emf measurements. The experiments were carried out as a number of potentiometric titrations in which V_0 ml of an initial solution containing B M Na_2B and H_0 M HClO_4 in $(3 - 2B)$ M NaClO_4 was titrated with equal volumes, V_T ml, of an H_T M HClO_4 solution and a $2B$ M Na_2B solution, in $(3 - H_T)$ M and $(3 - 4B)$ M NaClO_4 , respectively. In each titration, the total hydrogen ion concentration, H , was varied under constant B by the addition of HClO_4 . As Na_2B becomes insoluble in 3 M $\text{Na}(\text{ClO}_4)$ when $2B > 0.5$ M, the values of B used were chosen to be 0.0050, 0.0100, 0.0250, 0.0500, 0.1000, 0.1500 and 0.2000 M. A summary of the titrations is given in Table 1.

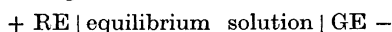
Table 1. Summary of the titrations.

Titn.	H_0 mM	B^a M	$H_{T(1)}$ M	$H_{T(2)}$ M
1	-0.074	0.0050	0.03985	0.09985
2	-0.074	0.0100	0.03985	0.09985
3	-0.074	0.0250	0.03985	0.09985
4	-0.074	0.0500	0.09985	0.3244
5	-0.070	0.0998	0.09992	0.3244
6	-0.068	0.1500	0.09992	0.3245
7	-0.065	0.2000	0.09992	0.3245

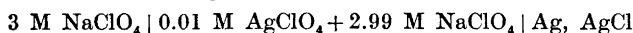
^a B was obtained from the analytical composition of the solution, as the phenolic group in B^{2-} is practically undissociated in acidic solution.

* The author is indebted to Mr. Jan Nilsson, B. Sc., for preparing the salt and determining the amount of dibrominated sulphosalicylic acid impurity, which was less than 0.5 mol %, as estimated by NMR methods.

During the course of a titration, h was measured by means of the following cell:



where GE denotes a glass electrode, and RE a reference half-cell of the composition



The titration vessel and the "Wilhelm bridge" used were kept in an oil bath thermostatically maintained at $25.00 \pm 0.05^\circ\text{C}$ in a room of temperature $25 \pm 2^\circ\text{C}$. The emf of the cell was determined with a Radiometer valve potentiometer (PHM 4C) to an accuracy of ± 0.2 mV.

From the measured emf, h was calculated using the relation:

$$E = E_0 + 59.15 \log h + E_j$$

where E_0 is a constant for the glass electrode and E_j is the liquid junction potential between the equilibrium solution and 3 M NaClO₄. E_j can be expressed as a function of h , namely $E_j = jh$, where j is a constant. Both E_0 and j were determined by performing a separate acid titration of 3 M NaClO₄ and calculated from the emf by the procedure described by Biedermann and Sillén.³

As the quantities, h , B , H_0 , and H_T were known, it was then possible to calculate Z , where Z is the average number of H⁺ bound per B⁻, by using the relation

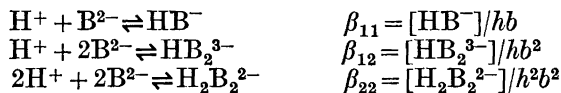
$$BZ = (H_0 V_0 + H_T V_T) / (V_0 + 2V_T) - h = H - h$$

In the preliminary calculations the analytical hydrogen ion concentration in the 3 M Na(ClO₄) medium was taken to be H_0 . As the stock solution of Na₂B was prepared by weighing an equivalent amount of Na₂CO₃ to neutralize the NaHB solution, and the concentration of Na₂B was determined as one half of the analytical value of the sodium ion concentration, any inaccuracy in the quantity of Na₂CO₃ will cause not only error in B but also in H_0 . Refined values of H_0 and B were later obtained by the LETAGROP treatment.

RESULTS AND CALCULATIONS

The experimental data obtained in the titrations are shown in Fig. 1, where Z has been plotted against $-\log h$ for various B values. The error bars indicated for the titration points correspond to an error in the measured emf of ± 0.1 mV. Within the experimental error of the measured emf these curves, $Z(-\log h)_B$, do not coincide for various B values. With decreasing B they approach a limiting curve. This indicates the formation of one or several polynuclear complexes when $B > ca. 0.005$ M.

To find the composition of these complexes and to determine the relevant formation constants, the "curve-fitting" and least squares methods were used. The most probable polynuclear complexes coexisting with HB⁻ were assumed to be HB₂³⁻ and H₂B₂²⁻. Because of the moderate concentrations of B used, the highest being only 0.2 M, the further association of dimers is less probable. The relevant reactions and equilibrium constants can thus be written:



The normalized function⁴

$$B = \frac{[u - (1 + u)Z][u(1 - 2\alpha) - 1]}{u[2Z(1 + \alpha u) - (1 + 2\alpha u)]^2}$$

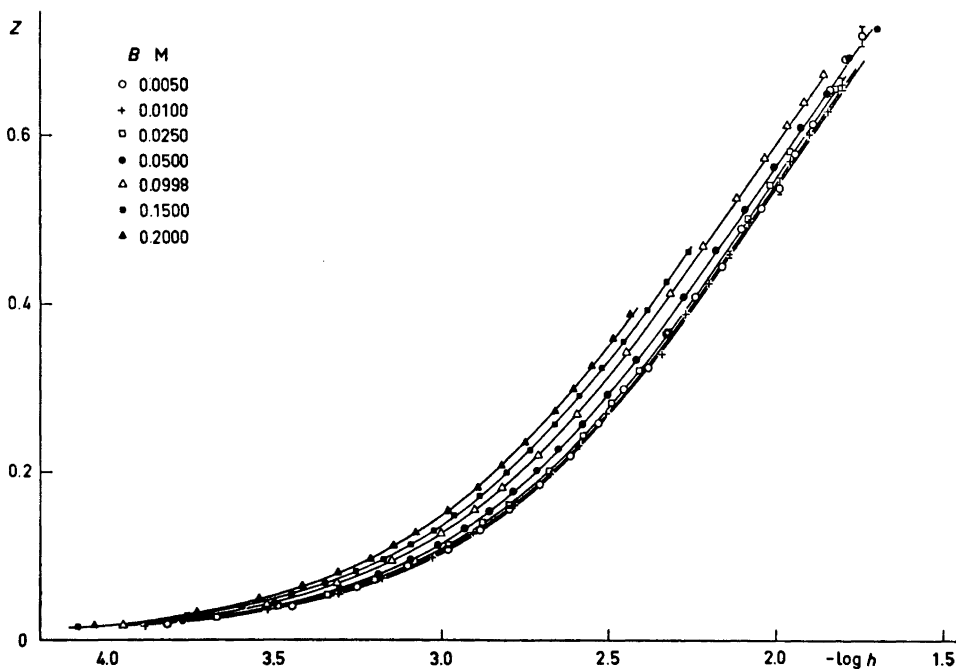


Fig. 1. Experimental data $Z(-\log h)_B$. Z is the average number of H^+ bound per B^{2-} . The error bars correspond to an error of ± 0.1 mV in the measured emf.

where $\log \alpha = \log \beta_{22} - \log \beta_{11} - \log \beta_{12}$, was then used in the "curve-fitting" method. The theoretical $\log \mathbf{B} = f(\log u)_{Z,\alpha}$ curves were calculated from the above equation and then compared with the experimental $\log B(\log h)_z$ curves

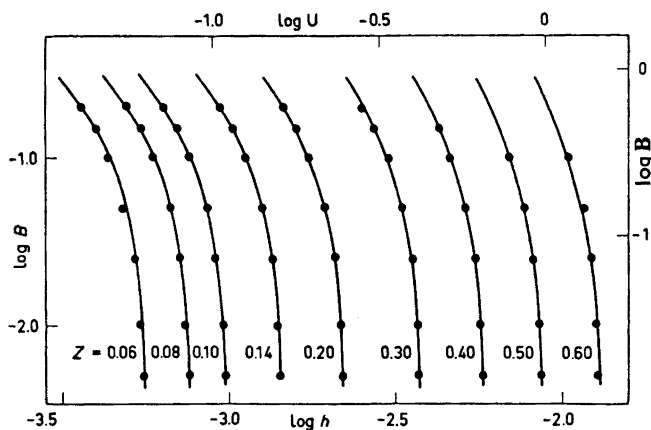


Fig. 2. Data for $\log B = f(\log h)_Z$ for the titrations of B^{2-} in 3 M $Na(ClO_4)$. The continuous curves are $\log \mathbf{B} = f(\log u)_{Z,\alpha}$ in the position of best fit for $\alpha = 1.4$.

obtained from Fig. 1. In Fig. 2, they are shown in the position of best fit, for which $\alpha = 1.4$. The equilibrium constants, β_{11} and β_{12} , were then calculated from the differences in intercept of the experimental and calculated graphs along the coordinate axes in Fig. 2, according to the following relations:

$$\log u = \log \beta_{11} + \log h, \quad \log \mathbf{B} - \log B = \log \beta_{12} - \log \beta_{11}.$$

The corresponding value of β_{22} was calculated from

$$\log \alpha = \log \beta_{22} - \log \beta_{11} - \log \beta_{12}.$$

The constants thus obtained were $\log \beta_{11} = 2.05$, $\log \beta_{12} = 2.52$ and $\log \beta_{22} = 4.72$.

These constants were then refined using the generalized least squares program LETAGROP.^{5,6} In this program, the computer searches for the "best" set of equilibrium constants, β_{pq} , which give the minimum value for the error squares sum $U = \sum (H_{\text{calc}} - H_{\text{tot}})^2$. In the computer calculations, the following equations were used:

$$E = E_0 + 59.156 \log h + jh$$

$$B_{\text{tot}} = Bf_B = b + \sum q\beta_{pq}h^pb^q$$

$$H_{\text{tot}} = V_0(H_0 + \delta H_0)/(V_0 + 2V_T) + V_T(H_T + \delta H_T)/(V_0 + 2V_T)$$

$$H_{\text{cal}} = h + \sum p\beta_{pq}h^pb^q$$

$$Z = (H_{\text{tot}} - h)/B_{\text{tot}}$$

where δH_0 and δH_T correspond to the analytical errors in H in the original titration vessel and the buret solution, respectively, and $f_B = B_{\text{tot}}/B$ is the analytical error factor for B . These parameters, δH_0 , δH_T , f_B , and E_0 , were thus treated as an unknown group of constants to be determined and adjusted alternatively with the equilibrium constants β_{pq} . The input information comprised the estimated values of β_{pq} and $j = -11$ mV/M (common for all data), E_0 , B , V_0 , and H_T (for each titration), V_T and E (for each point in the titration).

The calculations were performed under five different conditions. In each calculation, a number of group constants were adjusted alternatively with β_{pq} . The resulting "best" set of values for U_{min} , β_{pq} , δH_0 , δH_T , E_0 , and f_B , obtained in each calculations, is given in Tables 2 and 3.

Table 2. Equilibrium constants for the reactions $p\text{H}^+ + q\text{B}^{2-} \rightleftharpoons \text{H}_p\text{B}_q^{p-2q}$ calculated by LETAGROP. 25°C, 3 M Na(ClO₄), 144 points, $U = \sum (H_{\text{calc}} - H_{\text{tot}})^2$.

Calculation	$10^6 U_{\text{min}}$	$\log(\beta_{11} \pm 3\sigma)$	$\log(\beta_{12} \pm 3\sigma)$	$\log(\beta_{22} \pm 3\sigma)$	Group constants adjusted with β_{pq}
1	2.099	2.025 ± 0.009	2.566 ± 0.019	4.779 ± 0.022	δH_0
2	1.235	2.028 ± 0.016	2.548 ± 0.048	4.770 ± 0.046	$\delta H_0, \delta H_T$
3	1.096	2.021 ± 0.0001	2.536 ± 0.012	4.740 ± 0.001	$\delta H_0, E_0$
4	1.072	2.028 ± 0.007	2.537 ± 0.017	4.739 ± 0.019	$\delta H_0, f_B$
5	0.785	2.027 ± 0.004	2.541 ± 0.010	4.734 ± 0.011	$\delta H_0, E_0, f_B$

Table 3. E_0 , f_B , δH_0 , and δH_T calculated by LETAGROP for "best fit" for the calculation listed.

B M	$\delta H_0 \pm 3\sigma$ mM	$\delta H_T \pm 3\sigma$ mM	$E_0 \pm 3\sigma$ mV	$f_B \pm 3\sigma$
Calculation 1				
0.005	-0.097 ± 0.072		47.59	1.000
0.010	-0.136 ± 0.066		48.33	1.000
0.025	-0.168 ± 0.089		48.14	1.000
0.050	-0.226 ± 0.108		47.91	1.000
0.100	-0.192 ± 0.121		47.12	1.000
0.150	-0.107 ± 0.178		47.90	1.000
0.200	-0.246 ± 0.251		48.04	1.000
Calculation 2				
0.005	$+0.001 \pm 0.052$	-0.299 ± 0.118	47.59	1.000
0.010	-0.043 ± 0.037	-0.248 ± 0.077	48.33	1.000
0.025	-0.076 ± 0.080	-0.197 ± 0.133	48.14	1.000
0.050	-0.185 ± 0.151	-0.085 ± 0.234	47.91	1.000
0.100	-0.312 ± 0.187	$+0.165 \pm 0.261$	47.12	1.000
0.150	-0.371 ± 0.186	$+0.257 \pm 0.257$	47.90	1.000
0.200	-0.133 ± 0.282	-0.549 ± 0.306	48.04	1.000
Calculation 3				
0.005	-0.024 ± 0.040		47.30 ± 0.08	1.000
0.010	-0.074 ± 0.039		48.04 ± 0.08	1.000
0.025	-0.114 ± 0.073		47.80 ± 0.11	1.000
0.050	-0.246 ± 0.127		47.66 ± 0.12	1.000
0.100	-0.334 ± 0.192		46.78 ± 0.12	1.000
0.150	-0.334 ± 0.172		47.48 ± 0.11	1.000
0.200	-0.144 ± 0.241		47.25 ± 0.11	1.000
Calculation 4				
0.005	-0.001 ± 0.053		47.59	0.9508 ± 0.0201
0.010	-0.050 ± 0.036		48.33	0.9775 ± 0.0071
0.025	-0.086 ± 0.076		48.14	0.9861 ± 0.0046
0.050	-0.199 ± 0.133		47.91	0.9956 ± 0.0042
0.100	-0.318 ± 0.161		47.12	0.9965 ± 0.0025
0.150	-0.333 ± 0.157		47.90	0.9951 ± 0.0024
0.200	-0.178 ± 0.210		48.04	0.9862 ± 0.0024
Calculation 5				
0.005	-0.036 ± 0.052		47.10 ± 0.39	1.046 ± 0.081
0.010	-0.052 ± 0.042		48.31 ± 0.48	0.978 ± 0.048
0.025	-0.095 ± 0.075		47.49 ± 1.38	1.018 ± 0.063
0.050	-0.164 ± 0.081		46.28 ± 0.66	1.045 ± 0.022
0.100	-0.201 ± 0.169		46.51 ± 0.75	1.007 ± 0.017
0.150	-0.291 ± 0.198		47.86 ± 0.96	0.994 ± 0.022
0.200	-0.223 ± 0.246		48.90 ± 1.14	0.996 ± 0.029

DISCUSSION

It can be seen from Tables 2 and 3 that a marked improvement in the fit was obtained in Calculations 3 and 4, where the values of U are much lower and the values for β_{pg} are more certain than those in 1 and 2. The adjustment

of δH_0 and f_B in Calculation 4 indicated in Table 3 seems to be reasonable, considering the experimental conditions. The values of f_B indicate that the error in the determination of B is less than 1.4 % for all titrations except for those where the concentrations of B are low, namely 0.0100 and 0.0050 M (where the error is about 2 % and 5 %, respectively). The value of δH_0 increases with increasing B and the ratio of δH_0 to B is always less than 0.5 %.

The standard deviations found for the E_0 in Calculation 3 are reasonable, considering the reading uncertainty in measuring the emf of the electrodes.

It is obvious from Table 3, that the results of Calculation 5 had not been improved by adjusting the parameters δH_0 , E_0 , and f_B together with β_{pq} . The rather high standard deviations found for E_0 and f_B (up to 1.6 mV and 0.08, respectively) are not feasible in view of the uncertainty in the measurement of the emf and the analytical errors in B . Thus the preferred "best" set of equilibrium constants, β_{pq} , is that of Calculation 4 (given together with 3σ , where σ is the standard deviation), *i.e.*

$$\log \beta_{11} = 2.028 \pm 0.007$$

$$\log \beta_{12} = 2.537 \pm 0.017$$

$$\log \beta_{22} = 4.739 \pm 0.019$$

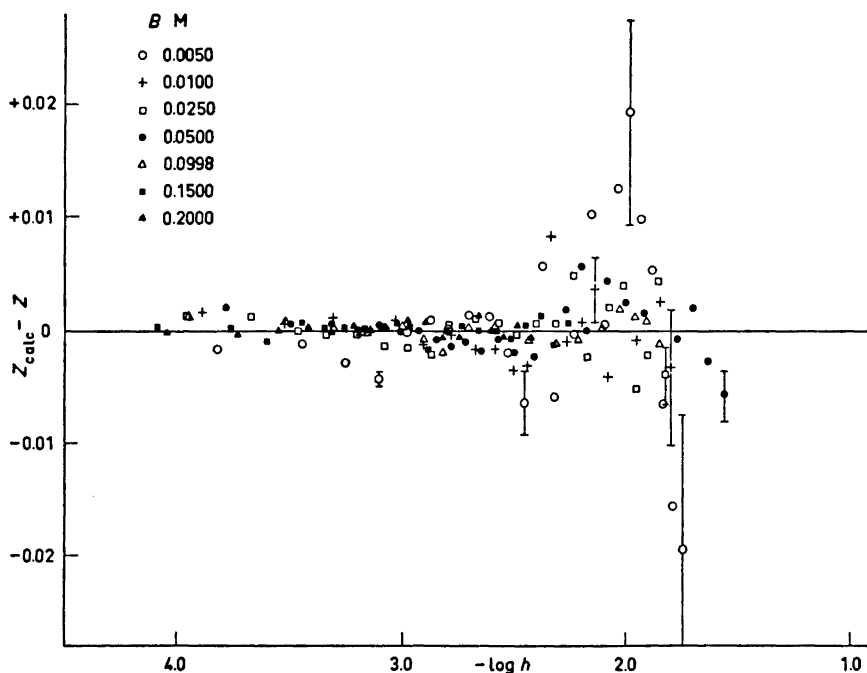
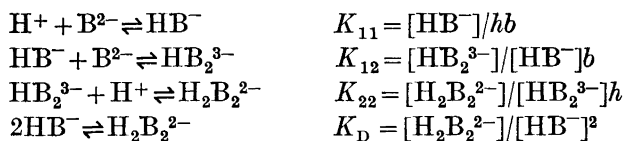


Fig. 3. $(Z_{\text{calc}} - Z)$ as a function of $\log h$, plotted for the same titrations as in Fig. 1. The error bars correspond to an error of ± 0.1 mV in the measured emf.

Using the values of β_{11} , β_{12} , β_{22} , δH_0 , and B found in Calculation 4, the differences ($Z_{\text{calc}} - Z$), have been plotted in Fig. 3, as a function of $\log h$, for the same titrations as shown in Fig. 1. It can be seen that the differences are randomly distributed and small, namely < 0.005 (with the exception of the 0.0050 M titration). The error bars represent the possible error in Z due to an error in the measurement of the emf E of ± 0.1 mV. It can also be seen that the possible errors become very large as Z and $\log h$ increase, especially for low concentrations of B .

The proton association constants, K_{11} and K_{22} , and the equilibrium constants of dimerization, K_{12} and K_D , corresponding to the following reactions, were calculated:



They are related to the values of β_{11} , β_{12} , and β_{22} by the following equations:

$$\beta_{11} = K_{11}, \quad \beta_{12} = K_{11}K_{12}, \quad \beta_{22} = K_{11}K_{12}K_{22}, \quad \beta_{22}/\beta_{11}^2 = K_D$$

so that $\log K_{11} = 2.028 \pm 0.007$, $\log K_{12} = 0.509 \pm 0.024$, $\log K_{22} = 2.202 \pm 0.036$, and $\log K_D = 0.683 \pm 0.033$. By comparing the values of these constants for 3-bromo-5-sulphosalicylic acid with those for salicylic acid,¹ namely $\log K_{11} = 3.173 \pm 0.001$, $\log K_{12} = 0.173 \pm 0.020$, $\log K_{22} = 3.359 \pm 0.048$, and $\log K_D = 0.359 \pm 0.031$, we see that the values of K_{11} and K_{22} of 3-bromo-5-sulphosalicylic acid are about 14 times smaller than those of salicylic acid, while the values of K_{12} and K_D are only slightly larger.

The higher acid strength of HB^- and $\text{H}_2\text{B}_2^{2-}$ for 3-bromo-5-sulphosalicylic acid, compared with that of salicylic acid, can be attributed to the $-I$ inductive effect of the substituents $-\text{Br}$ and $-\text{SO}_3^-$. The electron-attracting $-I$ effect of $-\text{Br}$ and $-\text{SO}_3^-$ causes a displacement of the bonding electron pair between the C and S, and C and Br atoms, in the direction of S and Br, respectively. This produces a succession of electron displacements along the chain towards the carboxyl group, thus increasing the tendency to dissociation of the acidic group by stabilizing the carboxylate anion.

The rather small variation in K_{12} and K_D between salicylic acid and 3-bromo-5-sulphosalicylic acid is quite understandable. It has been observed that in certain solvents, such as benzene, CCl_4 , etc., the equilibrium constants of dimerization are not particularly sensitive to a change in the enthalpy change, ΔH , of hydrogen bond formation,^{7,8} as both ΔH and the entropy change, ΔS , affect K_D according to the equation $-RT \ln K_D = \Delta H - T\Delta S$. An increase in both ΔH and ΔS will thus to a certain extent balance each other. In aqueous solution the acids associate with the water molecules rather than with their own species. Thus not only the degree of dimerization for carboxylic acids is extremely decreased but also the variation in K_{12} and K_D is much less than that in the benzene or CCl_4 solvents.

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