[CONTRIBUTION FROM THE JOHN HARRISON LABORATORY OF CHEMISTRY OF THE UNIVERSITY OF PENNSYLVANIA]

THE ACTIVITY COEFFICIENT OF LITHIUM HYDROXIDE IN WATER AND IN AQUEOUS LITHIUM CHLORIDE SOLUTIONS, AND THE DISSOCIATION OF WATER IN LITHIUM CHLORIDE SOLUTIONS

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RECEIVED AUGUST 31, 1925 PUBLISHED JANUARY 8, 1926

The activity coefficients of the alkaline hydroxides as determined from the measurements of suitable cells are unusually interesting because they exhibit a behavior which is just the opposite to that of the corresponding chlorides. Thus, it has been found[1] that the activity coefficient of sodium hydroxide in aqueous solution is less than that of potassium hydroxide at a given concentration while it is well known that the reverse is true for the chlorides. Further, it is known that the activity coefficient of hydrochloric acid is greater in sodium than in potassium chloride solution at a given strength whereas the activity coefficient of sodium hydroxide in sodium chloride solution is less than that of potassium hydroxide in potassium chloride solution.[2] Similar measurements of cells containing lithium hydroxide solutions are here presented.

It has been shown[3] to be possible to evaluate the activity coefficient product of water in these uni-univalent chloride solutions by combining measurements of cells of three types by a method which does not involve uncertainties resulting from the presence of liquid-junction potentials. This determination is interesting because there are no other data available on the activity coefficient of a weak electrolyte in the presence of a strong one, and because this knowledge permits the calculation of the dissociation of water in the salt solutions. All the data necessary for such a calculation of the dissociation in lithium chloride solutions and the calculations are therefore presented.

Experimental Part

Lithium Amalgam was prepared electrolytically from mercury which was distilled four times and employed as a cathode in the electrolysis of lithium chloride. The anode was a graphite rod, separated from the rest of the cell by a porous clay cup. At the conclusion of the electrolysis, the amalgam was drawn through a capillary into a highly evacuated glass bulb, to which was sealed a capillary tube and stopcock. The bulb was then sealed off from the vacuum line and inverted. The further manipulation of the amalgam was similar to that described by other investigators.[4] It was finally diluted to form an amalgam containing 0.004% of lithium.

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[1] Harned, THIS JOURNAL, **47**, 676 (1925).

[2] Ref. 1, pp. 684, 689.

[3] Harned, THIS JOURNAL, **47**, 930 (1925).

[4] (a) MacInnes and Beattie, *ibid.*, **42**, 1117 (1920). (b) Knobel, *ibid.*, **45**, 70 (1923).

Lithium Hydroxide was prepared by mixing nearly equivalent amounts of hot, nearly saturated solutions of barium hydroxide and lithium sulfate, so that a trace of the former remained in excess. After several days, the clear supernatant liquid was decanted from the precipitate. In this way a solution free from carbonate was prepared. The minute traces of dissolved barium hydroxide and sulfate it contained were not detrimental to the accuracy of the results. A stronger solution was prepared by dissolving the lithium hydroxide (commercial analyzed product) in water, adding barium hydroxide in slight excess and decanting. These solutions, which were about 0.5 and 4 *M*, respectively, were preserved in bottles fitted with siphons and protected from the carbon dioxide of the air by soda-lime tubes. They were analyzed by titrating weighed portions with hydrochloric acid that had been standardized gravimetrically. The solutions used in the cells were prepared by diluting with redistilled water weighed amounts of those described above. Since dissolved air must be avoided when alkaline amalgam electrodes are employed, the solutions were boiled in a vacuum and introduced into the cells without coming in contact with the air. Correction was made for the water boiled off.

Lithium Chloride.—A concentrated solution of the pure salt was accurately analyzed gravimetrically.

The mixed solutions were prepared by adding together weighed amounts of the analyzed solutions of the hydroxide and chloride with a little more than enough water to bring them to the desired strength. They were boiled in a vacuum in weighed flasks, and correction was made for the water boiled off. They were then transferred to the cells without coming into contact with air.

The solutions containing hydrochloric acid and lithium chloride were made in an analogous manner.

The Activity Coefficient of Lithium Hydroxide in Aqueous Solution

From the electromotive forces of the cell, $H_2 \mid LiOH\ (m_2) \mid Li_xHg \mid LiOH\ (m_1) \mid H_2$, the ratios of the activity coefficients of the hydroxides in solutions of concentrations m_2 and m_1 may be computed by the equation

$$\log \frac{\gamma_2}{\gamma_1} = \frac{E}{0.1183} - \log \frac{m_2}{m_1} + \frac{1}{2} \log \frac{a_{H_2O(m_2)}}{a_{H_2O(m_1)}} \qquad (1)$$

where γ_2 and γ_1 are the activity coefficients of the hydroxide, $a_{H_2O(m_2)}$ and $a_{H_2O(m_1)}$ are the activities of water, in solutions of molal concentrations m_2 and m_1, respectively.[5]

The cell apparatus employed was a somewhat modified form of that used by Harned. It was essentially the same as Knobel's cell except that the electrolyte did not flow during the measurement. All the necessary precautions to exclude air from the apparatus and remove air from the solutions were scrupulously observed.

Since at times erratic behaviors unlike the behaviors of sodium and potassium amalgam cells were observed which seemed positively to depend upon the nature and strength of the lithium amalgam, it is desirable to state more carefully the conditions under which consistent results were obtained. The amalgams contained 0.003–0.005% of lithium metal. The amalgam electrode from which amalgam flowed through the two capil-

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[5] Ref. 1, Equation 1.

lary exit tubes was filled in a vacuum with about 30 cc. of the amalgam. The exit tubes were of such a size and shape that the amalgam flowed in a stream from each tip for a distance of about 1 cm., where it broke into a spray of droplets. The rate of flow was about 3 cc. per minute from each tip.

The results are given in Table I. The temperature was 25°. Cols. 1 and 2 contain the molal concentrations (formula weights per 1000 g. of water) of the two cell compartments. Col. 3 contains the observed electromotive forces. The activity of water in the solutions was calculated by a method suggested by Lewis and Randall.[6] Approximate values of the activity of lithium hydroxide were computed, assuming that the activity of the water was the same in the two solutions. These values were then used for the graphical evaluation of the activity of the water for each concentration of lithium hydroxide. The results of this calculation are given in Col. 4, and the values of γ_2/γ_1 calculated by Equation 1, in Col. 5. Col. 6 contains the absolute values of the activity coefficients of lithium hydroxide at a concentration m_2, calculated on the assumption that the activity coefficient of 0.01 m lithium hydroxide is 0.900.[7]

Table I

Activity Coefficients of Lithium Hydroxide

m_2	m_1	E	$a_{H_2O(m_2)}/a_{H_2O(m_1)}$	γ_2/γ_1	γ_2
0.0505	0.01006	0.0761	0.999	0.877	0.788
.1008	.01005	.1088	.998	.828	.744
.1038	.01030	.1084	.998	.817	.734
.0235	.0978	− .0669	1.000	1.131	.839
.242	.1013	.0396	0.996	0.904	.668
.346	.0993	.0566	.995	.863	.639
.432	.0993	.0664	.992	.833	.617
.553	.1006	.0765	.990	.798	.590
.611	.1011	.0808	.989	.791	.586
1.063	.1006	.1069	.981	.748	.553
1.812	.0997	.1307	.966	.687	.509
2.386	.1043	.1412	.956	.668	.492
2.577	.1051	.1438	.952	.654	.482
3.926	.1051	.1651	.930	.652	.472

The Activity Coefficient of Lithium Hydroxide in Lithium Chloride Solutions

The activity coefficients of the hydroxide in the lithium chloride solutions were obtained from measurements of the cells, $H_2 \mid LiOH(m_2)$, LiCl-

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[6] Lewis and Randall, "Thermodynamics," McGraw-Hill Book Co., New York, **1923**, p. 269.

[7] The value at 0.01 M employed by Knobel and Harned for sodium and potassium hydroxide was 0.920. This is probably too high and we therefore have employed a lower value which agrees more closely with that obtained by Scatchard for potassium hydroxide [This Journal, **47**, 648 (1925)].

$(m) | Li_xHg | LiOH(m_1) | H_2$, where m_1 and m_2 had approximately the same value, and m was varied. Since the solutions had to be boiled, it was not convenient to attempt to make m_1 and m_2 exactly the desired value, and hence a correction to the observed electromotive force was necessary. This correction, ΔE, is obtained by the equation

$$\Delta E = 0.1183 \log \frac{m_1}{m_0} - 0.05915 \log \frac{m_2\,(m + m_2)}{m_0\,(m + m_0)} \tag{2}$$

in which m_0 is the desired concentration (0.1 or 0.01). The activity coefficient referred to an activity coefficient γ_0 at a concentration m_0 is given by the equation

$$\log \frac{\gamma}{\gamma_0} = \frac{E}{0.1183} - \log \sqrt{\frac{(m + m_0)}{m_0}} + \frac{1}{2} \log \frac{a_{H_2O(s)}}{a_{H_2O(m_0)}} \tag{3}$$

where $a_{H_2O(s)}$ and $a_{H_2O(m_0)}$ are the activities of the water in the solution containing salt and pure hydroxide, respectively.[1] The water activity ratios of p/p_0 were computed by the equation

$$ln \frac{p_0}{p} = \frac{c}{55.5}\left(2 + 2.303\ \alpha c - \frac{4.606\ \beta\alpha'}{\alpha' + 1}\ c^{\alpha'}\right) \tag{4}$$

which has been shown by Harned[8] to be a close approximation throughout the concentration range from 0.01 to 3 M lithium chloride. α, α', and β are empirically determined constants. It was assumed that a_{H_2O} was the same in the solution $LiOH(m_0)$, $LiCl(m)$ as in a solution of $LiCl(m + m_0)$. From the electromotive-force measurements of MacInnes and Beattie,[4a] the constants α, β and α' were found to be 0.21, 0.323 and 0.463, respectively. The results are given in Table II. Cols. 1, 2 and 3 contain the concentrations in the cells, Col. 4 the observed electromotive force E, Col. 5 the correction and Col. 6 $(E + \Delta E)$. The ratio of the activity of the water in the solution containing the salt to that of the pure hydroxide is given in Col. 7, and Col. 8 contains the values of γ computed by Equation (3).

TABLE II

ACTIVITY COEFFICIENT OF LITHIUM HYDROXIDE IN LITHIUM CHLORIDE SOLUTION

$m_0 = 0.01$ $\gamma_0 = 0.900$

m_1	m_2	m	E	ΔE	$E + \Delta E$	$\frac{a_{H_2O(s)}}{a_{H_2O(m_0)}}$	γ
0.00996	0.01001	0.0121	0.01729	−0.00025	0.0170	1.0000	0.842
.01001	.01005	.0550	.03921	− .00019	.0391	0.998	.755
.01001	.01009	.1825	.06116	− .00020	.0609	.994	.669
.01001	.01003	.276	.06900	− .00001	.0690	.990	.641
.01001	.00995	.390	.07654	.00016	.0767	.987	.629
.00997	.01002	.635	.08645	− .00017	.0863	.978	.594
.01002	.00999	1.044	.09622	.00028	.0965	.962	.563
.01004	.01002	2.17	.11274	.00016	.1127	.911	.521
.01005	.01002	4.15	.13194	.00014	.1321	[.793]	[.515]
.00998	.01005	8.36	.16250	− .00021	.1624	[.486]	[.513]

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[8] Harned, THIS JOURNAL, **44**, 252 (1922).

TABLE II (*Concluded*)

$m_0 = 0.1$ $\gamma_0 = 0.739$

m_1	m_2	m	E	ΔE	$E + \Delta E$	$\frac{a_{H_2O(s)}}{a_{H_2O(m_0)}}$	γ
0.1003	0.0993	0.0175	0.00284	0.00049	0.0033	1.0000	0.727
.0988	.0998	.0916	.01388	− .00054	.0133	0.997	.691
.1003	.1000	.1941	.02218	.00015	.0223	.994	.663
.1003	.0985	.2109	.02307	.00067	.0237	.993	.662
.1003	.1008	.222	.02432	− .00009	.0242	.993	.656
.1003	.1000	.456	.03571	.00018	.0359	.984	.625
.1003	.0993	.469	.03588	.00039	.0363	.984	.622
.0997	.1012	.746	.04537	− .00020	.0447	.974	.598
.1003	.0984	.978	.04942	.00060	.0500	.964	.585
.1003	.1000	1.208	.05467	.00013	.0548	.955	.579
.0999	.1001	2.85	.07632	.00009	.0762	.872	.560
.1003	.1047	2.95	.07831	− .00109	.0772	.866	.560
.0999	.1000	4.84	.09577	− .00007	.0957	[.741]	[.576]
.0999	.1038	8.48	.12706	− .00102	.1260	[.472]	[.637]

The results in brackets are uncertain owing to the uncertainty of the values of the vapor pressures.

The Activity Coefficient of Hydrochloric Acid in Lithium Chloride Solutions

In order to obtain the data necessary for the calculation of the dissociation of water in lithium chloride solutions, measurements of the cells, $H_2 \mid HCl(m_0), LiCl(m) \mid AgCl \mid Ag$, have been made, in which m_0 has a value of 0.1 and 0.01 molal and m was varied from 0 to 4 molal. The silver-silver chloride electrodes were made according to the directions of Noyes and Ellis.[9] The results are given in Table III. Col. 1 contains the lithium chloride concentration and Col. 2 the observed electromotive force, E, corrected to one atmosphere of hydrogen pressure. Col. 3 contains the values of $(E_0 - E)$ where E_0 is the electromotive force of the cell containing acid at a concentration m_0 and no lithium chloride. Col. 4 contains the activity coefficients of the acid calculated by the equation

$$\log \frac{\gamma}{\gamma_0} = \frac{E_0 - E}{0.1183} - \log \sqrt{\frac{(m + m_0)}{m_0}} \tag{5}$$

All values of γ are based on $\gamma_0 = 0.801$ at 0.1 M. The activity coefficients of the 0.1 M series are in good agreement with those obtained by Harned[10] from the cells, $H_2 \mid HCl(m), LiCl(m) \mid HgCl \mid Hg$, below 1 M lithium chloride concentration, but above this concentration the divergence becomes considerable and at 3 M corresponds to a difference of more than a millivolt. There are reasons to believe that these results are more accurate at the higher concentrations than those of Harned.

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[9] Noyes and Ellis, THIS JOURNAL, **39**, 2245 (1917).

[10] Harned, *ibid.*, **42**, 1808 (1920).

TABLE III

ACTIVITY COEFFICIENTS OF HYDROCHLORIC ACID IN LITHIUM CHLORIDE SOLUTIONS

$m_0 = 0.01$		$\gamma_0 = 0.914$	
m	E	$E_0 - E$	γ
0.0000	0.4645	0.0000	0.914
.00767	.4516	.0129	.884
.01028	.4485	.0160	.877
.01944	.4404	.0241	.852
.0422	.4268	.0377	.834
.0727	.4167	.0477	.805
.1209	.4063	.0582	.785
.1989	.3953	.0692	.769
.4024	.3785	.0860	.759
.699	.3634	.1011	.777
1.004	.3517	.1128	.816
2.189	.3180	.1465	1.088
3.987	.2756	.1889	1.807
$m_0 = 0.1$		$\gamma_0 = 0.801$	
0.000	0.3526	0.0000	0.801
.0705	.3407	.0120	.775
.1101	.3353	.0173	.774
.2004	.3271	.0256	.761
.3693	.3155	.0372	.763
.704	.3001	.0525	.783
1.007	.2890	.0636	.830
2.198	.2553	.0974	1.110
3.695	.2207	.1320	1.698

The Activity Coefficient Product and Dissociation of Water in Lithium Chloride Solutions

Harned[3] has shown that the activity coefficient product of water, $\gamma_H\gamma_{OH}/a_{H_2O}$ denoted K_γ may be calculated by means of the equation

$$E_1 + E_2 - E_3 = 0.05915 \log K_{\gamma(s)}/K_{\gamma(m_0)} \qquad (6)$$

where $K_{\gamma(s)}$ and $K_{\gamma(m_0)}$ are the values of this quantity in the salt solution and hydroxide, respectively. E_1, E_2, E_3 are the values of the electromotive forces of the cells

$Ag \mid AgCl \mid HCl(m_0), LiCl(m) \mid H_2 \mid HCl(m_0) \mid AgCl \mid Ag,$
$H_2 \mid LiOH(m_0), LiCl(m) \mid Li_xHg \mid LiOH(m_0) \mid H_2,$ and
$Ag \mid AgCl \mid LiCl(m + m_0) \mid Li_xHg \mid LiCl(m_0) \mid AgCl \mid Ag$

In order to compute $K_{\gamma(s)}$ by means of Equation 6, it is necessary to determine its value at 0.01 M concentration by some independent method. We have therefore, employed the approximate law[11] that in dilute solutions of mixtures of electrolytes the activity coefficient of one constituent depends only on the total ionic strength, or in the present instance, on the total concentration of the solution. The activity coefficient of the ions of

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[11] Ref. 6, p. 374.

water, $\sqrt{\gamma_H\gamma_{OH}}$, in a 0.01 M lithium hydroxide solution will accordingly be nearly the same as that of other strong electrolytes at 0.01 M total concentration. From the theory of Debye and Hückel[12] it follows that the activity coefficient below 0.01 M may be obtained from the equation

$$\log \gamma = -0.495 \sqrt{m}/1 + A \sqrt{2m} \tag{7}$$

where A is a constant which depends on the mean distance of approach of the ions. If we assign to A a value equal to that obtained for other strong uniunivalent electrolytes, namely, 0.79, then $\gamma_s = \sqrt{\gamma_H\gamma_{OH}} = 0.901$; and $K_{\gamma(s)} = 0.812$ at 0.01 M. The dissociation of water, or the molal concentration of dissociated water molecules per 1000 g. of water, m_w, may be computed by

$$m_w = m_H = m_{OH} = \sqrt{K_w/K_{\gamma(s)}} \tag{8}$$

where K_w is the ionic product of water, which was taken to be 1.005×10^{-14}. The calculation of these quantities was made and the results are given in Table IV. The values of E_1 and E_2 were obtained from the graphs of the data in Tables II and III. For the calculation of E_3, activity coefficients based on the measurements of MacInnes and Beattie[4a] were employed. Col. 5 contains $E_1 + E_2 - E_3$ and Col. 6 $K_{\gamma(s)}$. Since γ_s =

TABLE IV

THE ACTIVITY COEFFICIENT PRODUCT, ACTIVITY COEFFICIENT AND DISSOCIATION OF WATER IN LITHIUM CHLORIDE SOLUTIONS

$m_0 = 0.01$

$m+m_0$	E_1	E_2	E_3	$E_1+E_2-E_3$	$K_{\gamma(s)}$	a_{H_2O}	γ_s	$m_w \times 10^7$
0.01	0.0000	0.0000	0.0000	0.0000	0.812	1.000	0.901	1.112
.02	.0158	.0149	.0338	− .0031	.719	0.999	.848	1.182
.03	.0246	.0234	.0531	− .0051	.666	.999	.816	1.228
.05	.0367	.0339	.0781	− .0075	.606	.999	.778	1.288
.1	.0521	.0478	.1113	− .0114	.520	.996	.720	1.390
.2	.0682	.0668	.1449	− .0149	.454	.993	.672	1.487
.3	.0777	.0704	.1643	− .0162	.433	.990	.654	1.524
.5	.0909	.0811	.1906	− .0186	.393	.983	.622	1.599
.75	.1027	.0895	.2122	− .0200	.372	.974	.602	1.643
1.	.1123	.0955	.2277	− .0199	.374	.964	.600	1.638
2.	.1425	.1108	.2730	− .0197	.387	.920	.589	1.632
3.	.1665	.1216	.3066	− .0185	.395	.866	.585	1.594

$m_0 = 0.1$

$m+m_0$	E_1	E_2	E_3	$E_1+E_2-E_3$	$K_{\gamma(s)}$	a_{H_2O}	γ_s	$m_w \times 10^7$
0.1	0.0000	0.0000	0.0000	0.0000	0.520	0.996	0.720	1.390
.2	.0161	.0142	.0336	− .0033	.458	.993	.675	1.481
.4	.0329	.0289	.0680	− .0062	.409	.986	.635	1.568
.6	.0440	.0374	.0890	− .0076	.387	.979	.616	1.611
.8	.0524	.0435	.1045	− .0086	.372	.972	.602	1.631
1.1	.0634	.0505	.1221	− .0082	.378	.960	.603	1.630
2.1	.0924	.0667	.1653	− .0062	.409	.915	.612	1.568
3.1	.1163	.0777	.1980	− .0040	.445	.860	.619	1.502

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[12] Debye and Hückel, *Physik. Z.*, 9, 185 (1923).

$\sqrt{\gamma_{\mathrm{H}}\gamma_{\mathrm{OH}}} = \sqrt{K_{\gamma(s)}\, a_{\mathrm{H_2O}}}$, it is necessary to evaluate $a_{\mathrm{H_2O}}$. For this purpose Equation 4 was employed and the values obtained are given in Col. 6. Col. 7 contains γ_s and Col. 8, $m_w \times 10^7$.

Discussion

In Fig. 1, the activity coefficients of the alkaline hydroxides have been plotted against the molal concentrations. The activity coefficients of lithium hydroxide (Curve 3) have been taken from Table I of this paper, and those of sodium (Curve 2) and potassium hydroxide (Curve 1) from the data of Knobel[4b] and of Harned.[1] At a given concentration, it is seen that the activity coefficients of these three hydroxides increase in the order, lithium, sodium and potassium, whereas with the chlorides the reverse is true. From the thermodynamic point of view, potassium hydroxide is the strongest of these three, and sodium hydroxide is stronger than lithium hydroxide. This fact may lead to the conclusion that sodium and lithium hydroxides and perhaps potassium hydroxide are not completely dissociated.

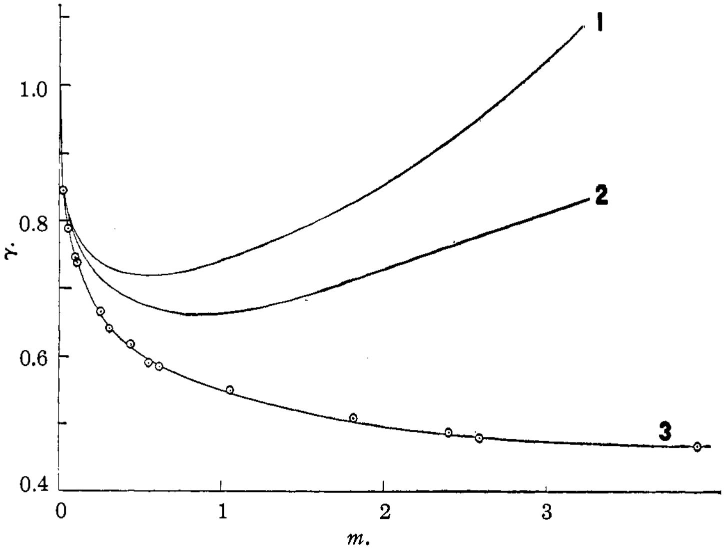

Fig. 1.—Activity coefficients of the alkaline hydroxides. 1. KOH; 2. NaOH; 3. LiOH.

In Fig. 2, the activity coefficients of lithium hydroxide in aqueous solutions (Curve 1) as well as those of 0.1 *M* lithium hydroxide (Curve 2) and 0.01 *M* lithium hydroxide (Curve 3) in lithium chloride solutions taken from Tables I and II are plotted against the total molal concentration. If these results are correct, then, at a given total concentration, the activity coefficient of lithium hydroxide first increases, passes through a maximum, and then decreases with decreasing hydroxide concentration. This is contrary to the behavior of other uni-univalent electrolytes in salt solutions. Although we are not certain, we are inclined to attribute this apparent anomaly to experimental error in the 0.01 *M* hydroxide-salt series which was found to be the most difficult to measure.

In Fig. 3, the ionic concentration of water in lithium chloride solutions (Curve 3) from the last column in Table IV has been plotted against the square root of the molality of this salt. Similar curves in potassium (Curve 1) and sodium (Curve 2) chloride solutions are also given.[3] The results obtained from the 0.1 *M* and 0.01 *M* series superimpose up to and including 1 *M* concentration. Above this concentration, the results from the two series differ somewhat. At 3 *M*, this difference corresponds to 2 mv. As mentioned above we think that this discrepancy is due to experimental error which is most likely to occur in the hydroxide-chloride cells at high concentrations.

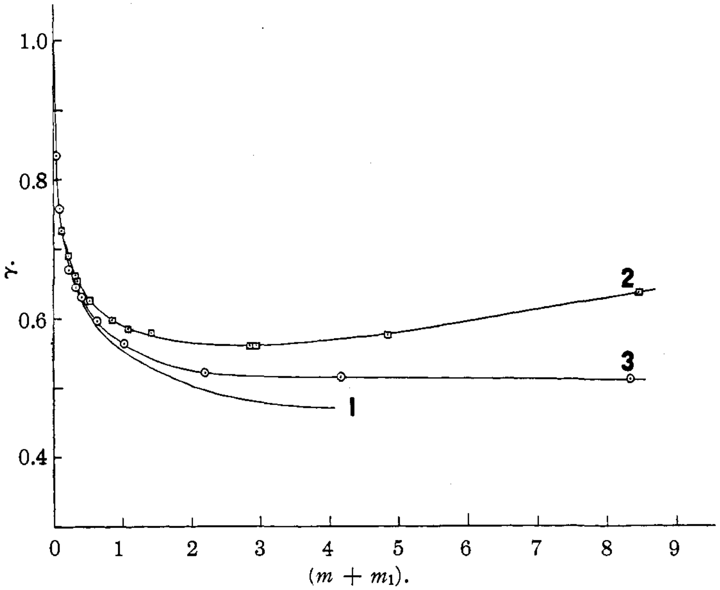

Fig. 2.—Activity coefficients of lithium hydroxide in lithium chloride solutions. 1. LiOH. 2. LiOH(0.1), LiCl. 3. LiOH(0.01), LiCl.

As a result of this experimental study, we find that up to a concentration of about 0.03 *M* there is a very slight difference between the effects of these chlorides on the dissociation of water. Above this concentration, lithium chloride causes a considerably greater dissociation than do the other salts. All the curves pass through a maximum.

It should be borne in mind that these results are all referred to a value at 0.01 *M* salt concentration which was somewhat arbitrarily selected and which may be in error to an extent of approximately 1%. This affects the absolute but not the relative values of the results.

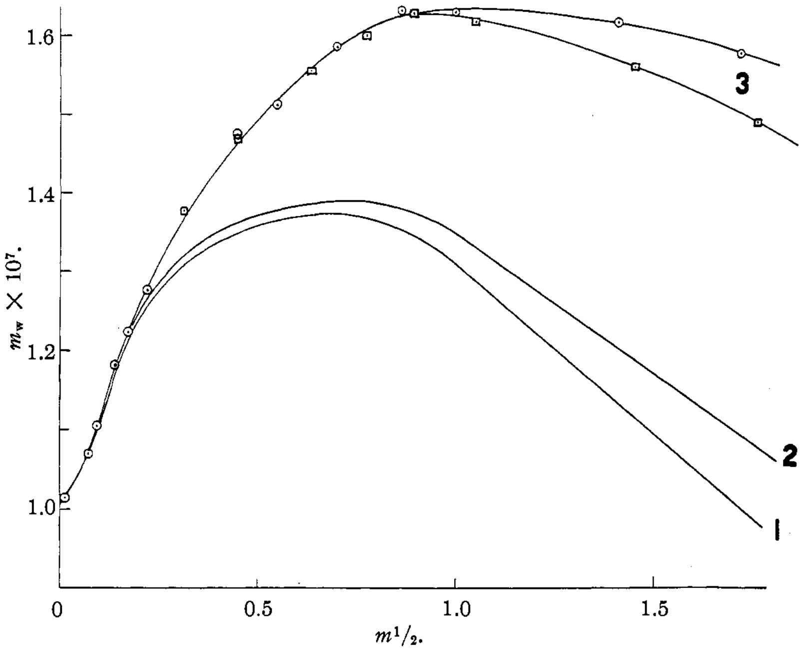

Fig. 3.—Hydroxyl- and hydrogen-ion concentrations in aqueous KCl, NaCl and LiCl solutions. 1. KCl; 2. NaCl; 3. LiCl. ⊡. 0.1 *M* series. ⊙. 0.01 *M* series.

Summary

1. Measurements of the following types of cells were made: $H_2 \mid LiOH(m_2) \mid Li_xHg \mid LiOH(m_1) \mid H_2$, $H_2 \mid LiOH(m_0), KiCl(m) \mid Li_xHg \mid LiOH(m_0) \mid H_2$, $H_2 \mid HCl(m_0), LiCl(m) \mid AgCl \mid Ag$.

2. From the results the activity coefficients of lithium hydroxide in aqueous solution and in lithium chloride solutions of various strength, and of hydrochloric acid in lithium chloride solutions were calculated.

3. A calculation of the ionic activity coefficient and the ionic concentration of water in lithium chloride solutions was made.

4. These results show that at the higher concentrations, lithium chloride produces a considerably greater increase in the dissociation of water than do sodium and potassium chlorides. The dissociation in the lithium chloride solutions reaches a maximum at 1 *M* salt concentration.

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