plane surface and the vapor phase, become identical—which is just what Donnan has postulated.

Is it not more than likely that both these hypotheses present phases of the truth and that, as Smoluchowski has pointed out, local dense spots of varying volume and density are formed in a nearly saturated solution and that some of these, when once formed, are stable, for the reasons which Thomson and Donnan have given? As to the size of these drops, it seems possible, since it is related to the surface tension, that the size may vary with different liquids and also with the temperature for the same liquid. The amount of supersaturation possible is regulated by the size of these drops.

Summary.

1. The degrees of supersaturation of four solutions of liquids in water have been measured. In the cases of nitrobenzene and carbon disulfide, the existence of supersaturation has been definitely established.
2. The solubility of nitrobenzene in water has been measured at three different temperatures.
3. The solubility of water in nitrobenzene has been measured at six different temperatures.
4. Various important theories have been correlated and an attempt made to use them to explain the phenomena observed.
5. A change in the density and other physical properties of small drops with the radius has been suggested as fulfilling one of the conditions in Donnan's theory of the critical opalescence.

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[CONTRIBUTION FROM THE RESEARCH LABORATORY OF PHYSICAL CHEMISTRY OF THE MASSACHUSETTS INSTITUTE OF TECHNOLOGY, NO. 103.]

THE CONDUCTANCE OF SOLUTIONS OF CERTAIN IODIDES IN ISOAMYL AND PROPYL ALCOHOLS.

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1. Introduction.

The study of the conductance of largely ionized solutes in aqueous solution has shown that the results over the range of concentration from 0.001 to 0.1 normal can be expressed fairly well by the equation $\frac{(c\gamma)^n}{c(1-\gamma)} = K$, first proposed by Storch.[1] In this expression c denotes the concentration, and γ the conductance ratio Λ/Λ_0; K is a constant, and n is another constant which has a value which varies somewhat with the nature of the solute, but almost always lies between 1.40 and 1.55.[2] In the case

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[1] *Z. physik. Chem.*, 19, 13 (1896).

[2] Noyes, THIS JOURNAL, 30, 343 (1908).

of non-aqueous solutions a large amount of experimental material has been collected which seemed to show that electrolytes in these solvents obey a law very different from that applicable to aqueous solutions; thus in certain solvents the equivalent conductance has been found even to increase with increasing concentration. A satisfactory way of reconciling this discrepancy between the behavior of aqueous and certain non-aqueous solutions was recently described in an article published by Kraus and Bray,[1] in which it is shown that the conductance of almost any uniunivalent solute in any solvent from the concentration zero up to a fairly high concentration (one where the change in the viscosity of the solution becomes an important factor), can be expressed by an equation of the form

$$\frac{(c\gamma)^2}{c(1-\gamma)} = K + D(c\gamma)^m.$$

In this equation K, D, and m are empirical constants which vary with the nature of the solute and of the solvent and with the temperature. This expression evidently requires that any solute in any solvent conform to the mass-action law at sufficiently small concentrations (when the term $D(c\gamma)^m$ becomes negligible); and Kraus and Bray have shown that uniunivalent salts do in fact conform to this law at small concentrations in the solvent liquid ammonia, in which reliable conductance measurements have been made in very dilute solutions by Franklin and Kraus,[2] and also in certain alcoholic solvents which have been studied by Dutoit and Duperthuis.[3] Computations made by Arrhenius[4] have shown that the mass-action law is probably also conformed to by solutes in aqueous solution of salts at concentrations below 0.0002 normal.

It seemed nevertheless desirable to confirm and extend Dutoit's measurements with alcohols, as these were confined to a few concentrations and as they have recently been questioned by Walden.[5] A study of the conductance of sodium iodide in isoamyl alcohol and in propyl alcohol and of ammonium iodide in isoamyl alcohol was therefore undertaken. As will be seen, the results of this investigation are in substantial agreement with those of Dutoit and Duperthuis.

This investigation was carried out with the aid of a grant made to Professor A. A. Noyes by the Carnegie Institution of Washington, for which we wish to express our great indebtedness. We also wish to thank Prof. C. A. Kraus and Prof. A. A. Noyes for many helpful suggestions.

2. Preparation of the Substances and Solutions.

A good grade of isoamyl alcohol was fractionated repeatedly. The

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[1] This Journal, 35, 1315–1434 (1913).

[2] *Am. Chem. J.*, 23, 277 (1900).

[3] *J. chim. phys.*, 6, 699 (1908).

[4] *Meddelanden fr. k. Vetenskapsakademiens Nobelinstitut*, 2, No. 42, 1–12 (1913).

[5] *Z. physik. Chem.*, 78, 257 (1911).

portion distilling between 130.5° and 131.5° was boiled with potassium hydroxide for several hours with a return-cooler. It was then distilled, metallic sodium was added to the distillate, and the alcohol was again distilled. By this treatment alcohol was obtained which had a specific conductance of 1.5×10^{-8}.

The pure alcohol was preserved in an amber-colored bottle from which the air was exhausted. Preserving the alcohol in this way lessened considerably the difficulties met with in the earlier part of the work, where yellow solutions of the iodides were obtained. The conductivity of these solutions also decreased rapidly, especially at the low concentrations. Treating the alcohol with potassium hydroxide as mentioned above removed this difficulty; but it was necessary to repeat this treatment frequently in order to keep the alcohol in good condition.

The propyl alcohol was dried with metallic sodium and fractionated, that portion distilling within the limits 97.3° and 97.5° being used for the measurements. This had a specific conductance of 2×10^{-8}. The propyl alcohol solutions also showed a slight tendency to become colored; but the change was very slow, and by carefully purifying the alcohol all trouble was avoided.

The sodium iodide was prepared by neutralizing purest sodium hydrocarbonate with gaseous hydrogen iodide; the ammonium iodide, by passing this gas into pure dry liquid ammonia. Very pure ammonium iodide was also obtained by subliming the commercial salt in a vacuum. Solutions of ammonium iodide prepared from the salt obtained by these two methods had identical conductances. All the salts were analyzed for their iodine content by precipitation of the iodine as silver iodide. The analyses in every case showed the theoretical amount of iodine within 0.1%. The hydrogen iodide was prepared according to a method described by Travers.[1]

All the solutions were made as follows: The salt in the form of pellets, thoroughly dried in a vacuum at 150°, was weighed and placed in a weighed container; alcohol was quickly added, and the container was exhausted, sealed off, and again weighed. From the weight of salt and alcohol, the volume concentrations were calculated with the aid of density values. The density measurements, on account of the large coefficient of expansion of isoamyl alcohol, could not be conveniently made with the Ostwald-Sprengel type of pycnometer. A small narrow-neck flask of about 25 cc. capacity fitted with a ground glass stopper was found to give sufficiently accurate results and was accordingly employed.

The following table contains the values of the density measured at 25° and referred to that of water at 4°.

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[1] Travers "Study of Gases," p. 51.

TABLE I.—DENSITY OF THE IODIDE SOLUTIONS AT 25°.

In isoamyl alcohol.			In propyl alcohol.		
Salt.	G. salt in 1000 g. solution.	Density.	Salt.	G. salt in 1000 g. solution.	Density.
None	0	0.8110	None	0	0.8043
NaI	35.80	0.8346	NaI	13.83	0.8141
NaI	84.96	0.8712	NaI	14.26	0.8141
NH_4I	0.845	0.8119	NaI	80.70	0.8626
NH_4I	23.41	0.8254	NaI	92.50	0.8715
NH_4I	32.34	0.8308	NaI	159.9	0.9256

3. Apparatus and Method of Measurement.

For the study of the conductance of non-aqueous solutions the methods hitherto employed are not well adapted, partly owing to the difficulty experienced in keeping the solutions free from moisture or other atmospheric impurities and partly owing to the oxidizing effect which the oxygen of the air may have upon organic liquids. Moreover, it is generally necessary to deal with small amounts of solvent, and consequently the method used must be such as will insure reasonable economy of material.

A new form of apparatus was therefore devised for the purpose. This is shown in Fig. 1. It consists of two parts, the cell proper M and the part N used for adding known amounts of solution to the solvent in the cell. The cell had a capacity of about 300 cc. The electrodes are shown at A, electrical connection being made with these through the small tubes B by means of mercury. The electrodes were made of bright platinum. Two different cells were used. In the larger cell the electrodes had an area of about 10 sq. cm. and were placed about $^1/_2$ cm. apart; in the smaller cell they had an area of about 1.5 sq. cm. and were placed about 1 cm. apart. Since it was necessary to shake the liquid in the cell, these electrodes were firmly tied together by means of glass crossbars connecting the corners. This gave reasonable assurance that the conductance capacity of the cell would not change during the measurements. However, this was measured frequently to make certain of this. The side tube C was merely for emptying the cell, and was sealed off before carrying out the measurements. The side tube at E was used in exhausting the cell, the opening at D being closed, either by the dropper N or by a rubber stopper. The cell was exhausted with a rotary oil vacuum pump capable of giving a pressure as low as 0.001 mm. After thoroughly exhausting it, the large stopcock, having a bore of 1 cm., could be closed and the cell handled at will.

The dropper N was a bulb of about 50 cc. capacity. The delivery tube L consists of a very fine capillary tube having a diameter of about 0.75 mm. This was further obstructed by a fine platinum wire. The dropper was closed at the top by the ground glass valve F. This valve was kept tightly closed by a piece of rubber tubing, as shown at H. The enlargement at G is part of a ground glass joint which fits into the cell M at D. The part shown at P is a small glass cap which fits on the end of the capillary tube. This was provided with four glass prongs. These served to catch the drops of liquid just before they fell. On falling from these tips they fell freely without touching the sides of the cell.

In addition to the cell just described, an apparatus was constructed for preserving liquids of low vapor pressure in a vacuum. This is shown in Fig. 2. The container is closed by the small valve *a*, which can be opened by raising the plug with an electromagnet acting on D. This was a flat piece of iron fixed in place in the glass tube by means of a little cement. The valve could be held open by suspending the plug, by means of the lugs *c*, from two small hooks attached at *d* (not shown in the figure). When closed the valve was tightly sealed by surrounding it with mercury from the reservoir B. In transferring solvent the cell could be sealed to the container at *f* and the whole apparatus exhausted through a side tube. The valve *a* could then be opened and the solvent distilled into the cell. The advantage of this container is that there are no stopcocks to contaminate the solvent with grease.

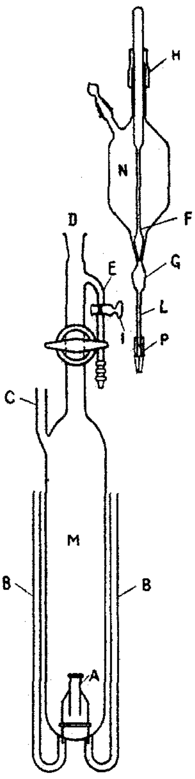

FIG. 1.

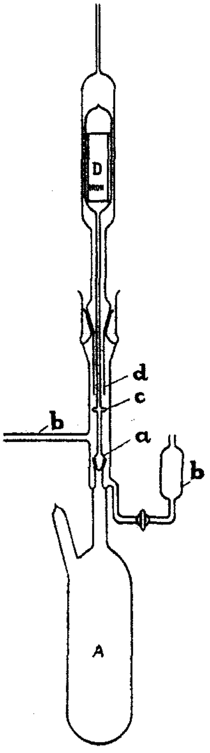

FIG. 2.

To make a measurement the cell was thoroughly dried, exhausted, and weighed. The required amount of alcohol (generally about 50 cc.) was then poured into the cell, and the cell was again exhausted and weighed, after which the conductivity of the solvent was immediately measured. The weighed dropper, in which was a solution of known concentration (by weight), was then placed in position, and the small space above the large stopcock was exhausted. The stopcock I was then closed and the large stopcock opened. By opening the valve F, solution varying from one to two drops to any desired volume could be added. The large stopcock was next closed and air was admitted to the chamber, after which the dropper was removed and weighed. The solution in the cell was then thoroughly mixed and the conductance measured. This process could be repeated until the desired number of concentrations had been obtained. In this way solutions were easily and accurately made of any desired concentration up to the limit which could be measured in the cell.

The conductance-capacity of the cell was determined in the usual way, with the aid of a 0.005 normal potassium chloride solution.

The measurements were made with an ordinary slide-wire bridge provided with a platinum-iridium wire, which had been previously calibrated.

The conductance of the solvent, owing to its small value, could not be measured on the bridge. In order to obtain this value, the cell was attached directly to a 220-volt circuit and the current passing through the cell was measured. This was accomplished indirectly by measuring the voltage-drop across a shunt of 0.1 ohm, by means of the swing of a galvanometer. The galvanometer had previously been calibrated with the aid of a standard cadmium element in a circuit of 10,000 ohms resistance.

4. The Conductance Data.

The conductance data are presented in Tables II, III, and IV. They are given in the order in which they were originally obtained. Each series consists of six or more measurements which were all made with one filling

TABLE II.—CONDUCTANCE DATA FOR SODIUM IODIDE IN ISOAMYL ALCOHOL.

$10^3 c$.	10^6 L.	Λ.	$100\Lambda/\Lambda_0$.
0.00	....	7.79	100.00
0.0763	0.5306	6.96	89.35
0.2153	1.299	6.03	77.38
0.502	2.528	5.04	64.63
0.942	4.030	4.28	54.94
1.58	5.779	3.67	47.11
2.21	7.274	3.29	42.26
0.067	0.475	7.04	90.37
0.130	0.850	6.54	83.98
0.232	1.381	5.95	76.35
0.386	2.061	5.35	68.65
0.802	3.57	4.45	57.15
1.280	4.996	3.90	50.09
1.533	5.654	3.69	47.33
1.617	5.86	3.63	47.66
6.714	15.41	2.30	29.45
15.93	27.93	1.75	22.50
37.26	53.36	1.43	18.37
58.55	78.32	1.34	17.18
71.38	93.72	1.31	16.85
2.296	7.48	3.26	41.82
7.106	15.91	2.24	28.74
20.43	33.39	1.63	20.98
38.77	54.93	1.41	18.19
55.12	74.36	1.33	17.11
73.28	94.67	1.29	16.59
10.30	20.68	2.01	25.76
22.02	35.51	1.61	20.69
48.80	66.94	1.37	17.59
86.82	112.64	1.30	16.65
146.5	192.7	1.32	16.88
9.36	19.33	2.067	26.53
17.85	30.37	1.70	21.82
34.95	50.49	1.45	18.55
74.43	97.15	1.31	16.75
130.0	169.14	1.30	16.69
179.4	237.5	1.32	16.98
201.5	270.1	1.34	17.20
270.5	373.8	1.38	17.74
322.2	452.6	1.40	18.02
374.1	526.2	1.41	18.05
502.7	702.8	1.40	17.95
174.8	231.5	1.32	16.98
248.0	336.6	1.36	17.42
308.0	430.1	1.40	17.92
393.0	554.8	1.41	18.13
499.5	697.9	1.40	17.93

TABLE III.—CONDUCTANCE DATA FOR AMMONIUM IODIDE IN ISOAMYL ALCOHOL.

$10^3 c$.	10^6 L.	Λ.	100 Λ/Λ_0.	$10^3 c$.	10^6 L.	Λ.	100 Λ/Λ_0.
0.00	....	7.92	100.00	2.290	7.994	3.49	44.06
0.0662	0.471	7.13	90.00	2.204	7.42	3.62	45.68
0.151	0.988	6.54	82.56	5.97	15.40	2.58	32.57
0.366	2.035	5.56	70.18	12.70	25.67	2.02	25.50
0.904	4.073	4.51	56.88	27.23	45.07	1.65	20.88
1.54	5.976	3.88	48.95	50.28	73.07	1.45	18.34
2.11	7.446	3.53	44.54	71.05	98.48	1.39	17.50
0.0575	0.4204	7.32	92.35	3.23	10.13	3.13	39.54
0.1707	1.099	6.440	81.29	6.57	16.39	2.49	31.48
0.3043	1.752	5.76	72.65	13.81	27.47	1.99	25.10
0.7909	3.679	4.650	58.70	29.75	48.17	1.62	20.42
1.39	5.583	4.015	50.68	47.27	71.34	1.46	18.44
1.81	6.713	3.714	46.88	67.35	94.10	1.40	17.63
0.051	0.377	7.47	94.30	62.88	87.93	1.397	17.63
0.117	0.800	6.87	86.72	88.35	118.8	1.346	16.99
0.263	1.556	5.94	74.98	112.7	150.2	1.332	16.81
0.644	3.180	4.94	62.36	137.0	182.8	1.335	16.85
1.365	5.520	4.04	51.00	171.7	229.6	1.337	16.88
2.208	7.720	3.49	44.09	58.08	82.03	1.412	17.82
0.0577	0.4286	7.43	93.79	81.30	110.2	1.356	17.12
0.1316	0.8947	6.80	85.79	102.6	137.1	1.336	16.87
0.2617	1.598	6.11	77.13	127.9	169.7	1.326	16.74
0.6980	3.395	4.87	61.41	161.8	215.9	1.334	16.84
1.376	5.567	4.05	51.06				

TABLE IV.—CONDUCTANCE DATA FOR SODIUM IODIDE IN PROPYL ALCOHOL.

$10^3 c$.	10^6 L.	Λ.	100 Λ/Λ_0.	$10^3 c$.	10^6 L.	Λ.	100 Λ/Λ_0.
0.00	....	20.10	100	53.40	448.6	8.40	41.8
0.0447	0.9057	20.25	...	75.08	585.2	7.79	38.7
0.1693	3.26	19.26	95.8	2.303	35.57	15.42	76.7
0.4553	8.276	18.17	90.4	5.092	69.28	13.61	67.7
1.074	18.12	16.87	83.9	12.37	141.65	11.44	56.9
2.083	32.57	15.62	77.7	29.81	283.6	9.510	47.3
0.0623	1.244	19.94	99.2	53.85	451.4	8.38	41.7
0.1581	3.060	19.36	96.3	77.45	601.9	7.77	38.6
0.2261	4.311	19.06	94.8	28.53	272.6	9.55	47.5
0.3902	7.167	18.36	91.3	55.4	459.2	8.29	41.2
0.6591	11.69	17.72	88.2	106.1	767.0	7.23	36.0
1.055	17.90	16.96	84.4	187.6	1219.5	6.5	32.3
1.498	24.41	16.30	81.1	315.6	1872.0	5.93	29.5
2.260	35.04	15.5	77.1	35.18	321.1	9.14	45.4
2.310	35.61	15.40	76.7	129.7	903.4	6.97	34.8
5.89	77.93	13.23	65.8	242.9	1506.5	6.21	30.9
13.26	149.7	11.28	56.1	362.0	2108.0	5.82	28.9
27.77	272.6	9.815	48.8				

of the cell. In the first column the concentrations $10^3 c$ are given expressed in milliequivalents of salt per liter of solution. In the second column are given the specific conductance L expressed in reciprocal ohms and multiplied by 10^6. In the next column are given the equivalent conductances Λ expressed in reciprocal ohms; and in the last column are given the conductance-ratios Λ/Λ_0 multiplied by 100.

5. Treatment of the Results.

If the mass-action law is conformed to in dilute solutions, the equation $c\Lambda = (1/\Lambda)\, K\Lambda_0{}^2 - K\Lambda_0$ holds true, and a straight line would be obtained by plotting $1/\Lambda$ against $c\Lambda$. Extrapolation to zero concentration would then give the value of $1/\Lambda_0$. From this value of Λ_0 that of K can be calculated directly from the above equation by introducing known values of c and Λ, or it can be obtained by determining the slope of the line on the plot.

Plots made in this way showed that the mass-action law is conformed to in very dilute solution; for up to an ion-concentration about 0.0004 normal the points fall nearly on a straight line. Thus in the case of ammonium iodide in isoamyl alcohol the average deviation from the best representative straight line of the twelve points corresponding to the values of $1/\Lambda$ determined between the ion concentrations 0.00005 and 0.0004 is 0.8%; and in the case of sodium iodide in isoamyl alcohol the average deviations of the corresponding ten points is 0.4%. A plot showing these results is presented in Fig. 3. The results with sodium iodide in propyl

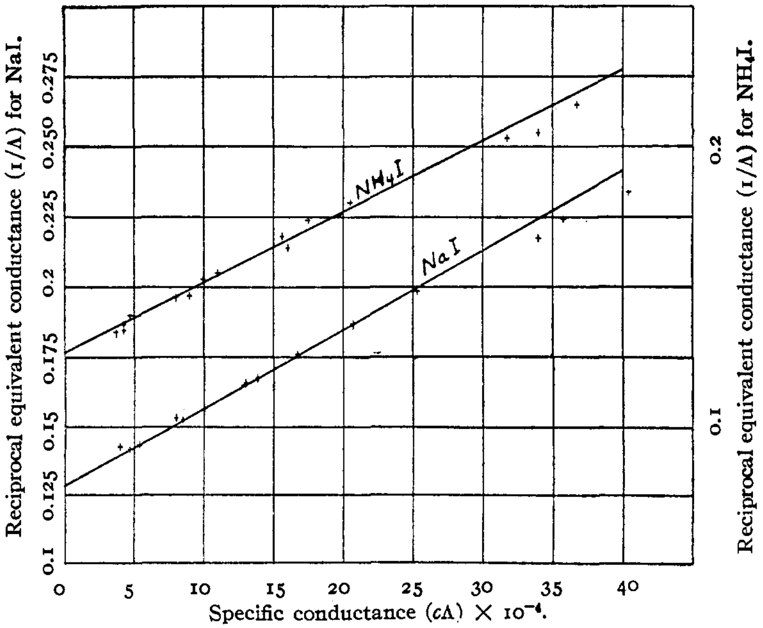

Fig. 3.

alcohol show a correspondingly good agreement with the mass-action law: when plotted, only the two measurements at the two lowest concentrations (45 and 62 × 10^{-6} normal) show an appreciable deviation from the best representative straight line.

The Λ_0 values obtained from such plots[1] are given at the beginning of Tables II, III and IV. The value given for ammonium iodide is only approximate.[2] The value for sodium iodide (7.79) is somewhat lower than the value (7.87) found by interpolation from Dutoit's results.

By transposing the Kraus equation and taking the logarithm of both sides we get the equation $\log\left(\frac{c\gamma^2}{1-\gamma} - K\right) = m \log c\gamma + \log D$. Hence, a straight line should be obtained by plotting the first member of this equation as ordinates against $\log c\gamma$ as abscissas, and the slope of this line should give m and its intersection with the axis of ordinates $\log D$.

Plots made in this way (with the aid of a provisional value of K derived from the dilute-solution plots referred to above) showed a very good agreement with the Kraus equation, especially in the more concentrated solutions. Variations in the dilute solutions were to be expected, since the equation is so sensitive in this region to the value assumed for K. Having determined from such plots the value of the constant m, the constants D and K were more accurately determined by plotting $c\gamma^2/(1-\gamma)$ against $(c\gamma)^m$. For the equation $c\gamma^2/(1-\gamma) = K + D(c\gamma)^m$ requires that this graph should be a straight line, the slope of which corresponds to the value of D and its intersection on the axis of ordinates to the value of K. Using this new value of K in a logarithmic plot of the kind first described, new values of m and D were obtained.

The final values of the constants so computed are given in Table V.

TABLE V.—CONSTANTS OF THE KRAUS EQUATION.

Alcohol.	Salt.	Λ_0.	$10^4 K$.	m.	D.
Isoamyl........	NaI	7.79	5.85	1.2	0.374
Isoamyl........	NH_4I	8.00	6.93	1.17	0.324
Propyl.........	NaI	20.10	38.3	0.75	0.208

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[1] Walden (*Z. physik. Chem.*, **78**, 257 (1911)) has taken exception to this method of determining Λ_0, involving the assumption that the mass-action law is obeyed in dilute solution, and has attempted to show that Λ_0 is better determined by using the Kohlrausch extrapolation formula. This, however, is a purely empirical relation, and in view of the present facts its use hardly seems to be justified.

[2] This value is in fact about 1% higher than the value (7.92) obtained by extrapolation on the dilute-solution plot. The reasons for the choice of this value were that variations took place in the cell-constant during the measurement of the dilute solutions which made the extrapolation somewhat uncertain and that this value fitted the results in concentrated solutions much more accurately than the extrapolated value. In other words, the Λ_0 value was really obtained by extrapolation from the concentrated solutions.

In Table VI are given side by side the "observed" equivalent conductance values accurated interpolated on a large-scale plot of Λ against log c and the "calculated" equivalent conductance values computed by the Kraus equation $c\gamma^2/(1 - \gamma) = K + D(c\gamma)^m$ with the aid of the constants given in Table V. The column heading $10^3\,c$ gives the concentration in milliequivalents of salt per liter of solution.

TABLE VI.—OBSERVED AND CALCULATED VALUES OF THE EQUIVALENT CONDUCTANCE.

	NaI in isoamyl alcohol.			NH_4I in isoamyl alcohol.			NaI in propyl alcohol.		
$10^3\,c$.	Λ obs.	Λ calc.	% diff.	Λ obs.	Λ calc.	% diff.	Λ obs.	Λ calc.	% diff.
0.00	7.79	...	..	7.92	...	..	20.10	...	..
0.05	7.182	7.194	+0.17	7.396	7.476	+1.08	(20.11)	...	..
0.10	6.636	6.658	+0.33	7.032	7.105	+1.04	19.65	19.53	—0.61
0.20	6.115	6.142	+0.43	6.256	6.352	+1.54	19.17	19.13	—0.21
0.40	5.307	5.335	+0.53	5.464	5.598	+2.46	18.42	18.49	+0.38
0.80	4.451	4.460	+0.20	4.672	4.794	+2.62	17.38	17.56	+1.03
1.00	4.184	4.183	—0.03	4.408	4.542	+3.03	17.00	17.21	+1.24
2.0	3.394	3.374	—0.61	3.588	3.663	+2.23	15.73	16.03	+1.90
5.0	2.560	2.534	—1.02	2.656	2.676	+0.75	13.62	13.56	—0.44
10.0	2.024	1.995	—1.43	2.168	2.164	—0.186	12.04	11.80	—1.99
20.0	1.649	1.635	—0.69	1.776	1.763	—0.73	10.44	10.31	—1.25
40.0	1.408	1.404	—0.28	1.448	1.456	+0.55	8.90	8.82	—0.90
100	1.294	1.290	—0.31	1.336	1.339	+0.22	7.33	7.26	—0.95
150	1.312	1.309	+0.023	1.332	1.335	+0.22	6.85	6.79	—0.88
200	1.339	1.336	—0.22						
300	1.393	1.400	+0.50						
400	1.413	1.459	+3.20						
500	1.396	...	..						

It will be seen that the agreement of the observed values with those calculated by the Kraus equation is fairly close (except at the highest concentration in the case of sodium iodide in isoamyl alcohol) concentration.

6. Summary.

In this article has been described an accurate method of determining conductances in an evacuated apparatus.

Results have been presented on the equivalent conductance of sodium iodide and ammonium iodide in isoamyl alcohol and of sodium iodide in propyl alcohol from a concentration of about 0.00005 normal up to saturation (about 0.5, 0.2, and 0.4 normal), respectively.

It has been shown that up to an ion concentration of 0.0004 normal these electrolytes conform to the mass-action law; and that their conductance is expressed throughout the whole range of concentration by the equation $c\gamma^2/(1 - \gamma) = K + D(c\gamma)^m$ proposed by Kraus, in which c is the concentration, γ is the conductance-ratio Λ/Λ_0, and K, D, and m are empirical constants. Thus the corresponding generalizations drawn by Kraus and Bray are confirmed by these results.

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