While the elution maximum appeared within less than a column volume after change of conditions, the Au(III) band tailed significantly, presumably a rate effect. Similar separations were carried out in HCl solutions where tailing of the Au(III) band seemed to be slightly greater than in LiCl solutions.

In part B of Fig. 3, separation of Sn(IV) and In(III) from Sb(V) is illustrated. A mixture of Sn(IV), In(III) and Sb(V) in 12 *M* HCl was added to a small column. Sn(IV) and In(III) were eluted within a few column volumes with 12 *M* HCl while Sb(V) remained adsorbed. It was removed with 1 *M* NaOH, presumably as the antimonate ion.

Since Fe(III) is strongly adsorbed by Dowex-50, the technique may be used for separation of traces of iron from various concentrated electrolyte solutions. A similar method previously has been described for anion exchangers.[9] The cation exchange method may be considered an adjunct to the anion exchange method since the supporting electrolyte may be chosen from a different group of metals. In the one case the metals should be nonadsorbable (or weakly adsorbable) by cation exchangers, in the other by anion exchangers.

With cation-exchange technique we have, as an example, separated Fe(III) from 2 *M* $ZnCl_2$ solutions containing excess HCl (Zn(II) is strongly adsorbed by anion exchangers). Unfortunately, if the Fe(III) concentration is substantially larger than 10^{-4} *M*, breakthrough of iron occurs after only a moderate number of column volumes of solution have been passed through, presumably because of the rapid decrease of adsorbabilities with loading. Better retention of Fe(III) would probably occur with less acidic supporting electrolyte solutions, *e.g.*, using LiCl rather than HCl.

The technique is of course also adaptable to the isolation of other adsorbable elements, *e.g.*, to the isolation of Ga(III) from concentrated Zn(II) solutions as one might encounter, for example, in the isolation of Ga^{67} tracer from zinc targets. In all these cases removal of the adsorbed material from the column (regeneration) may be achieved by washing with dilute chloride solutions.

OAK RIDGE, TENNESSEE

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[9] K. A. Kraus and G. E. Moore, THIS JOURNAL, **72**, 5792 (1950).

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF TORONTO]

Equivalent and Ionic Conductances for Mixtures of Potassium and Lithium Chlorides in Aqueous Solution at 25°

BY G. S. KELL AND A. R. GORDON

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Conductances for KCl–LiCl mixtures at an ionic strength of 0.01 have been determined by the direct current method. Transference numbers for potassium and chloride ions were obtained by the electrometric method of following boundary movement and for lithium ion by the adjusted indicator technique. The deficit in the equivalent conductance of the mixture as compared with that calculated by the simple additivity rule was less than that predicted by the Onsager and Fuoss theory of the mixture effect, in agreement with other earlier results at higher concentrations. The cation conductances, while not of the highest precision, show a decrease in conductance for the potassium ion and an increase for the lithium in the mixture as compared with their values in binary solution, numerically greater than is predicted by the theory.

In 1911 Bray and Hunt[1] reported that the conductance of a mixture of HCl and NaCl (concentrations C_1 and C_2) was not additive in the contributions of the two electrolytes, *i.e.*, if Λ_1 and Λ_2 were the equivalent conductances of the single electrolytes in solution at the same ionic strength J as obtained in the mixture, the equivalent conductance Λ of the mixture was less than $(C_1\Lambda_1 + C_2\Lambda_2)/J$, where $\Lambda = 1000\kappa/J$, κ is the specific conductance of the mixture, and $J = C_1 + C_2$. Subsequently, Bennewitz, Wagner and Küchler[2] and Longsworth[3] measured transference numbers for HCl–KCl mixtures, the former by the Hittorf method and Longsworth by the moving boundary technique. It appeared that the conductance of the faster cation was decreased, and of the slower cation increased, the common anion being relatively unaffected—the increase (or decrease) being numerically the greater the less the particular ion fraction in the mixture. The net effect on the conductance of the mixture was that it was less than that predicted by the additivity rule in conformity with Bray and Hunt's findings.

In a classic paper,[4] Onsager and Fuoss gave a general treatment of the problem and for the three-ion case obtained a solution in closed form.[5] Their development, however, involves essentially the assumptions underlying the Debye–Onsager Limiting Law for single binary electrolytes, and it was therefore not to be expected at an ionic strength of 0.1 (as obtained in Longsworth's measurements) that more than qualitative agreement with theory could be obtained. Later, Krieger and Kilpatrick[6] studied KCl–LiCl mixtures over a range of ion fractions at varying ionic strengths and also found decreases in conductance in the mixtures as compared with the additivity rule; these were in general less numerically than those predicted by the Onsager–Fuoss theory, although the discrepancy tended to decrease with decreasing ionic strength. Since they made no transference measurements,

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[1] W. C. Bray and F. L. Hunt, THIS JOURNAL, **33**, 781 (1911).

[2] K. Bennewitz, C. Wagner and K. Küchler, *Physik. Z.*, **30**, 623 (1929).

[3] I. G. Longsworth, THIS JOURNAL, **52**, 1897 (1930).

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

[5] For the polyion case, see L. Onsager and S. K. Kim, *ibid.*, **61**, 215 (1957).

[6] K. A. Krieger and M. Kilpatrick, THIS JOURNAL, **59**, 1878 (1937).

they were unable to obtain ion conductances, and it should be noted that these are the quantities explicitly dealt with by Onsager and Fuoss.

The recent development in this Laboratory of an electrometric method of following boundary movement,[7] applicable at concentrations far below those accessible to the conventional optical procedure, suggested a study of the mixture effect at an ionic strength low enough to permit a possibly significant test of the Onsager–Fuoss theory. Moreover, by the use of the adjusted indicator technique,[8] it would be possible to measure the transference number of the third ion (which Longsworth was unable to do) and thus obtain an independent check on the internal self-consistency of the data. We report here the results of such measurements for the system KCl–LiCl at $J = 0.01$. Our reason for selecting this ionic strength is that the mixture effect (which is proportional to $J^{1/2}$) must be large enough to be measurable and at the same time the concentrations must be sufficiently low for the independent ionic mobility rule to apply in solutions of the single salts; moreover, the "higher order terms" in C, $C \log C$, etc., should be as small as possible. It must be admitted at once that the precision of the transference measurements (and consequently the conductances) of the cations is not what one has come to expect in studies of single electrolytes, but the results (we believe) are nevertheless significant.

Experimental

The conductances were determined by the direct current method[9] previously used in this Laboratory and are based on the Jones and Bradshaw 0.01 D standard.[10] The transference number of the chloride ion was measured in a cell with rising boundary with iodate as indicator and of the potassium in a cell with falling boundary with LiCl as indicator; in both cases boundary movement was followed electrometrically as described by Lorimer, Graham and Gordon,[7] and the volume calibration of the cells is based on the value $t_+ = 0.4902$ for 0.01 N KCl at 25°.[11,12] Silver–silver chloride and cadmium electrodes served as cathode and anode on the closed side of the cells. The concentration of the adjusted indicator solution behind the LiCl/LiCl + KCl boundary was determined as described by Muir, Graham and Gordon,[8] the only difference in technique being that the conductiometric analysis was effected at 5000 c.p.s. instead of at 1000 c.p.s. as in the earlier work. The actual indicator solutions used were those formed in the measurement of t_K; thus two runs, one to measure t_{Cl} and one other to measure t_K and the indicator concentration, served to fix all three transference numbers.

Since the effect of interest is at best small and difficult to measure, preliminary calculation showed that the most promising solution for the transference studies was equinormal in the two salts. It was found that with such a mixture the potassium ion boundary (while not as sharp as that for chloride) was reasonably well defined, and the measured t_K or t_{Cl} was current independent for a 100% variation of current. In contrast, a solution 0.0025 N in KCl and 0.0075 N in LiCl gave a potassium ion boundary too ill-defined for meaningful results.

The solutions were prepared gravimetrically from the purified salts and CO_2 stabilized conductivity water (sp. cond. 7×10^{-7}) as previously described.[12,13] In converting to volume concentrations, it was assumed that the salts made additive contributions to the density of the solution, these being obtained from the density data for the solutions of the individual salts.

Results

The conductance measurements are summarized in Table I. Here μ is the ion fraction of potassium in the mixture, *i.e.*, 2μ is the electrolyte mole fraction of KCl. Hundredth normal LiCl and KCl solutions were prepared for each of the four series of the table, and the conductance of the LiCl solution used for a particular series is given at the head of each column. In all four, the equivalent conductance of the KCl solution was taken as 141.271, being in fact the cell calibration[14] for the series. For convenience the entries have been corrected to $J = 0.01$ by means of the known values of $d\Lambda/dC$ (-38_0 and -36_0) for this concentration for KCl and LiCl, on the assumption that $d\Lambda/dJ$ for the mixture is linear in μ. Since in only one measurement did J differ from 0.01 by as much as 0.2%, the resulting corrections in Λ amount in general to a few thousandth's of a conductance unit.

TABLE I

EQUIVALENT CONDUCTANCE OF KCl–LiCl MIXTURES

$J = 0.01$; 25°; $\Delta_{KCl} = 141.271$

2μ	Λ	2μ	Λ
(a)		(b)	
0	107.214	0	107.226
0.24954	115.668 (0.045)	0.24914	115.665 (0.043)
.50121	124.217 (.067)	.50190	124.266 (.047)
.75003	132.722 (.036)	.75011	132.727 (.036)
(c)		(d)	
0	107.212	0	107.237
0.25083	115.713 (0.042)	0.24770	115.628 (0.039)
.50132	124.240 (.046)	.49685	124.092 (.055)
.75050	132.736 (.037)	.74927	132.695 (.043)

TABLE II

TRANSFERENCE NUMBERS OF KCl–LiCl MIXTURES

$10^4 \times J$	2μ	$\Lambda(J = 0.01)$	t_{Cl}	λ_{Cl}
99.942	0.50171	124.23	0.5797	72.02
99.942	.50171	124.23	.5797	72.02
100.102	.50048	124.21	.5800	72.04
			Mean	72.03
			$\Delta\lambda_{Cl}$	+0.02

$10^4 \times J$	2μ	$\Lambda(J = 0.01)$	t_K	t_{Li}	λ_K	λ_{Li}
100.102	0.50048	124.21	0.2777	0.1435	68.92	35.68
100.102	.50048	124.21	.2775	.1427	68.87	35.48
99.961	.50088	124.24	.2770	.1431	68.71	35.62
99.961	.50088	124.24	.2772	.1424	68.76	35.45
				Mean	68.81	35.56
				$\Delta\lambda_i$	−0.45	+0.35

The average of the four results for LiCl solutions is 107.22, the corresponding value of the Shedlovsky function[15] Λ_0' being 115.89. This result and those of Jervis, *et al.*,[13] for concentrations less than 0.005

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[7] J. W. Lorimer, J. R. Graham and A. R. Gordon, THIS JOURNAL, **79**, 2347 (1957).

[8] D. R. Muir, J. R. Graham and A. R. Gordon, *ibid.*, **76**, 2157 (1954).

[9] H. E. Gunning and A. R. Gordon, *J. Chem. Phys.*, **10**, 126 (1942).

[10] G. Jones and B. B. Bradshaw, THIS JOURNAL, **55**, 1780 (1933).

[11] L. G. Longsworth, *ibid.*, **54**, 2741 (1932).

[12] R. W. Allgood, D. J. LeRoy and A. R. Gordon, *J. Chem. Phys.*, **8**, 418 (1940).

[13] R. E. Jervis, D. R. Muir, J. P. Butler and A. R. Gordon, THIS JOURNAL, **75**, 2855 (1953).

[14] The 0.01 D standard, for KCl = 74.557, corresponds to C = 0.0099713, Λ = 141.282. Since $d\Lambda/dC$ for 0.01 N KCl solutions = −380, Λ for 0.01 N is 141.271.

[15] T. Shedlovsky, THIS JOURNAL, **54**, 1405 (1932).

N, can be represented within 0.01 conductance unit by $\Lambda_0' = 115.03 + 86C$, a relation which fits all the data as closely as their $\Lambda_0' = 115.03 + 88C$. Our value is, however, lower than that interpolated from Shedlovsky's measurements[16] (107.30) and that of Krieger and Kilpatrick[6] (107.27). Since we are concerned here primarily with conductance differences, we shall use our result for self-consistency.

The quantities in parentheses after the entries are the values of $-\Delta\Lambda$, the deficit in the equivalent conductance as compared with that computed by the additivity rule

$$\Delta\Lambda = \Lambda - \Lambda_{\mathrm{LiCl}} - 2\mu(141.271 - \Lambda_{\mathrm{LiCl}}) \qquad (1)$$

the value of Λ_{LiCl} appropriate for the series being used in each case. If the third entry in series a be ignored, the mean values of $-\Delta\Lambda$ (0.042, 0.050 and 0.038 for 2μ = 0.25, 0.50 and 0.75, respectively) are surprisingly precise, the apparent uncertainty being of the order of ± a few thousandths of a conductance unit; they are consistent with Krieger and Kilpatrick's results[6] for this ionic strength within the much larger apparent experimental uncertainty of their data. The corresponding theoretical values of $-\Delta\Lambda$ (see eq. 5, *et seq.*) are 0.064, 0.084 and 0.061, the experimental results thus being roughly two-thirds the theoretical; just how misleading this can be in terms of the ionic quantities is discussed below.

Four determinations of t_{Cl} for 0.01 N LiCl gave 0.6715 ± 0.0001, *i.e.*, $t_{\mathrm{Li}} = 0.3285$; this is somewhat below the value for this concentration interpolated from Longsworth's data, *viz.*, 0.3289,[11] but once again for consistency we shall use our result. It should be noted that λ_{Cl} is 72.00 for $0.01N$ LiCl ($t_{\mathrm{Cl}} = 0.6715$, $\Lambda = 107.22$) and 72.02 for 0.01 N KCl ($t_{\mathrm{Cl}} = 0.5098$, $\Lambda = 141.271$); thus the independent ionic mobility rule holds within the accuracy of the transference numbers, and we shall adopt $\lambda_{\mathrm{Cl}} = 72.01$, $\lambda_{\mathrm{K}} = 69.26$, $\lambda_{\mathrm{Li}} = 35.21$ for this concentration in binary solution.

The transference numbers for the mixture are given in Table II. For t_{K} and t_{Cl}, they are computed by the familiar relation

$$t_{\mathrm{i}} = VC_{\mathrm{i}}/1000 + \text{cor. terms} \qquad (2)$$

where V is the volume swept out by the boundary per faraday, C_{i} is the concentration of the leading ion in equivalents per liter, and the last item stands for the usual solvent and volume corrections.[17] For lithium, the conservation relation[18] is used

$$T_{\mathrm{Li}}^* - T_{\mathrm{Li}} = V(C_{\mathrm{Li}}^* - C_{\mathrm{Li}})/1000 = (t_{\mathrm{K}}/C_{\mathrm{K}})(C_{\mathrm{Li}}^* - C_{\mathrm{Li}}) \qquad (3)$$

Here the asterisk denotes the adjusted LiCl indicator solution, and t_{Li} and C_{Li} refer to the leading solution. Thus a measurement of C_{Li}, which fixes[19] t_{Li}^*, yields t_{Li}, since t_{K} is determined from the observed boundary movement in the same experiment. The correction given in eq. 4 of ref. 8 has not been applied to the lithium results since it is negligible in view of the reproducibility of the data.

The table gives the value of Λ (adjusted to J = 0.01) for the actual solutions used in the transference measurements, and the resulting ionic conductances in the mixture[20]

$$\lambda_{\mathrm{i}} = t_{\mathrm{i}}\Lambda(J/C_{\mathrm{i}}) \qquad (4)$$

The cation conductances show a much greater spread than one could wish—approaching ±0.1 unit—but the mean values give a value of Λ of 124.22, which is in excellent agreement with the "best" value (124.20 for J = 0.01, 2μ = 0.5) from Table I. In fact, if the mean values be used with Λ = 124.20, the corresponding transference numbers (0.2770, 0.1432 and 0.5799) add to 1.0001. While this must be to some extent fortuitous, it nevertheless suggests that the mean values are reliable within narrower limits than one might infer from the scatter of the individual results. The values of $\Delta\lambda_{\mathrm{i}}$ in the table are the differences (λ_{i} in the mixture) − (λ_{i} in binary solution at J = 0.01).

Discussion

For a solution containing three univalent ions, the Onsager–Fuoss theory gives

$$\lambda_{\mathrm{i}} = \lambda_{\mathrm{i}}^{\circ} - \vartheta_{\mathrm{i}}\lambda_{\mathrm{i}}^{\circ}J^{1/2} - \sigma J^{1/2} \qquad (5)$$

where the coefficients ϑ_{i} and σ are defined by

$$\vartheta_{\mathrm{i}} = 1.980 \times 10^6(2^{1/2})z_{\mathrm{i}}(1 - H^{1/2})_{\mathrm{is}}r_{\mathrm{s}}/(DT)^{3/2} \qquad (6)$$

$$\sigma = 29.15(2^{1/2})/(DT)^{1/2}\eta \qquad (7)$$

Here z_{i} is the sign of the charge of the ion, D and η are the dielectric constant and viscosity of the solvent at temperature T, the subscript s indicates summation over the three ions and the vectors r are +1 for the cations and −1 for chloride. For a single binary electrolyte, the matrix product $z_{\mathrm{i}}(1 - H^{1/2})_{\mathrm{is}}r_{\mathrm{s}}$ becomes 0.29289—the familiar $q/(1 + q^{1/2})$ where $q = 1/2$, *i.e.*, for aqueous binary solutions at 25°, $\vartheta_{\mathrm{i}} = 0.2289$. The change of conductance of the ion in the mixture as compared with its conductance in binary solution is thus ascribed to the change in the time of relaxation coefficient ϑ_{i}.

To compute for a given μ, the matrix elements h_{ji}, one requires the nine quantities $\omega_{\mathrm{ji}} = \lambda_{\mathrm{j}}^{\circ}/(\lambda_{\mathrm{j}}^{\circ} + \lambda_{\mathrm{i}}^{\circ})$ where $\lambda_{\mathrm{K}}^{\circ}$, $\lambda_{\mathrm{Li}}^{\circ}$ and $\lambda_{\mathrm{Cl}}^{\circ}$ are 73.50, 38.68 and 76.35, respectively.[21] From these, one computes the matrix invariants I_1 and I_2, and from the invariants, the scalars a and b which serve to evaluate the matrix product

$$z_{\mathrm{i}}(1 - H^{1/2})_{\mathrm{is}}r_{\mathrm{s}} = z_{\mathrm{i}}(ar_{\mathrm{i}} + bh_{\mathrm{is}}r_{\mathrm{s}}) \qquad (8)$$

These quantities are given for reference in Table III. The difference between the entry in the table and 0.29289 when multiplied by (0.2289/0.2929). $\lambda_{\mathrm{i}}^{0}J^{1/2}$ gives the change in ionic conductance in the

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(16) T. Shedlovsky, THIS JOURNAL, **54**, 1411 (1932).

(17) D. A. MacInnes and L. G. Longsworth, *Chem. Revs.*, **11**, 171 (1932).

(18) L. G. Longsworth, THIS JOURNAL, **67**, 1109 (1945).

(19) Since $\lambda_{\mathrm{Li}}^{\circ}$ is 38.68 (R. A. Robinson and R. H. Stokes, "Electrolyte Solutions," Butterworth's Scientific Publications, London, 1955, p. 452; H. S. Harned and B. B. Owen, "The Physical Chemistry of Electrolytic Solutions," 3rd Ed., Reinhold Publ. Corp., New York, N. Y., 1958, p. 231; and ref. 13) t_{Li}° is 0.3363; with our value of t_{Li} for 0.01 N LiCl, the value of the Longsworth function $t_+^{\circ\prime}$ (see ref. 11) for this concentration is 0.3377. Since $t_+^{\circ\prime}$ is linear in C for the alkali chlorides, t_{Li}^* can be obtained by means of the relation $t_{\mathrm{Li}}^{\circ\prime} = 0.3363 + 0.14C_{\mathrm{Li}}^*$.

(20) Note that (J/C_{i}) is unity for chloride, $1/(2\mu)$ for potassium and $1/(1 - 2\mu)$ for lithium.

(21) See references in footnote 19, and G. C. Benson and A. R. Gordon, *J. Chem. Phys.*, **13**, 473 (1945).

TABLE III

VALUES OF $z_i(1 - H^{1/2})_{is}r_s$

2μ	0	0.25	0.50	0.75	1.00
K^+	0.40695	0.37672	0.34772	0.31981	0.29289
Li^+	.29289	.26603	.24010	.21504	.19072
Cl^-	.29289	.29370	.29391	.29362	.29289

mixture as compared with that in binary solution at the same J. From the resulting ionic quantities, the calculated $\Delta\Lambda$ referred to earlier may be obtained at once.

For $J = 0.01$, $2\mu = 0.5$, the calculated $\Delta\lambda_{K}$, $\Delta\lambda_{Li}$ and $\Delta\lambda_{Cl}$ are thus -0.315, $+0.160$ and -0.006 unit, respectively, which may be compared with the observed values of Table II. Not much stress can be placed on the difference in sign for chloride, although the discrepancy is slightly greater than the apparent precision of the chloride data. For the cations, however, the differences between the calculated and observed $\Delta\lambda_i$ are greater than the most pessimistic estimate of the experimental error. What is unexpected is that the observed $\Delta\lambda_i$ are numerically greater than the calculated rather than less. The assumptions underlying the theory of the mixture effect are essentially those inherent in the familiar Limiting Law for binary electrolytes, and for the alkali chlorides at these concentrations, the Limiting Law over—rather than underestimates—the effect of ionic interaction. It is of course possible that the assumption implicit here that the higher order terms are the same in the mixture as in binary solution is seriously in error, but at the moment no theory of a mixture effect for the higher order terms is available. However, these terms for all three ions in binary solution at $J = 0.01$ are of the order of 0.4 unit—roughly one-half the numerical value of the relaxation term of eq. 5 for Li^+ and one quarter that for K^+; hence, to account in this way for the discrepancy would require rather drastic changes percentually in their values.

One obvious conclusion is that further study, both experimental and theoretical, of the mixture effect is in order; a second is that when testing any theory dealing with ionic conductances, equivalent conductances *per se* can give little if any explicit information which is pertinent. Finally, we wish to express our thanks to the National Research Council of Canada for a grant in aid of this research and to Canadian Industries Limited for the award to G. S. K. of a fellowship.

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[CONTRIBUTION FROM THE GENERAL ELECTRIC RESEARCH LABORATORY AND BROOKHAVEN NATIONAL LABORATORY]

A Comparison of the Decomposition of Hexane and Cyclohexane by Different Types of Radiation

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Hexane and cyclohexane have been irradiated with radiations having energy loss parameters from 0.02 to 5 e.v./Å. and the liquid products have been examined by gas chromatographic and infrared methods. In this preliminary study no change in over-all decomposition is observed which can be attributed directly to a variation in the linear energy transfer of the radiation.

At the present time there is a paucity of comparative data on the decomposition produced in liquid organic materials by different types of radiation. This situation is in marked contrast to the reasonably detailed understanding now available of the dependence of radiation effects on linear energy transfer (LET) in aqueous systems. The present studies represent a preliminary investigation into the effect of variation with LET of the over-all decomposition produced in two of the more simple hydrocarbons which have been the subject of previous work with fast electrons.[3,4] Except for minor differences which appear to be attributable to changes in intensity, very similar yields are observed for the various types of radiations employed here.

Early experiments by Sworski and Burton[5] on aromatic systems which compared the relative yields of the formation of hydrogen and methane for reactor radiations and for γ-rays have indicated that the ratio of these products increased with increase of LET. McDonell and Gordon[6] have compared the γ-ray radiolysis of methanol with previous studies of the decomposition produced by 28 mev. helium ions and have observed a difference in the yield for formation of ethylene glycol. More recent studies of the yield of hydrogen produced in the decomposition of cyclohexane by fast electrons and by cyclotron radiations have shown little difference between the effect of light and heavy particles (less than 10%).[7] Investigations currently in progress on the determination of radical production by the iodine scavenging technique have shown a somewhat larger effect.[8] All of these studies emphasize that meaningful interpretation of the effect of LET on radiation chemical reactions in hydrocarbon systems will result only from studies of the quantitative details of relatively small

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(1) General Electric Research Laboratory, Schenectady, New York.

(2) Mellon Institute Radiation Research Laboratories, Pittsburgh, Pennsylvania; work performed while at Brookhaven National Laboratory and supported by the U. S. Atomic Energy Commission.

(3) H. A. Dewhurst, *J. Chem. Phys.*, **24**, 1254 (1956); *J. Phys. Chem.*, **61**, 1466 (1957); **62**, 15 (1958).

(4) H. A. Dewhurst and E. H. Windlow, *J. Chem. Phys.*, **26**, 969 (1957).

(5) T. J. Sworski and M. Burton, THIS JOURNAL, **73**, 3790 (1951).

(6) W. R. McDonell and S. Gordon, *J. Chem. Phys.*, **23**, 208 (1955).

(7) R. H. Schuler and A. O. Allen, THIS JOURNAL, **77**, 507 (1955).

(8) R. H. Schuler, to be published.