then added, CO (1 atm) was bubbled through the solution for 5 min, and the resulting solution was analyzed by ^1H and ^{31}P NMR and IR spectroscopy as shown in Table III.

(ii) With PMe_3. To a solution of **4a** as above was added PMe_3 (2.16 μL, 0.021 mmol). Quantitative conversion to the PMe_3 adduct was observed.

(iii) With MeCN. To **4a** (40.0 mg, 0.042 mmol) in C_6D_6 (0.5 mL) in an NMR tube was added MeCN (2.2 μL, 0.042 mmol). Quantitative conversion to the MeCN adduct was observed.

(iv) With Alkenes. Low-temperature ^1H NMR experiments were carried out on a Bruker WM 250 (250-MHz) instrument with a probe precooled to 193 K (−80 °C). The temperature was maintained to within ±1 °C. In a typical experiment, $[IrH_2[(p\text{-}FC_6H_4)_3P]_2(\eta^2\text{-}O_2CCF_3)]$ (**4a**; 12.2 mg, 0.013 mmol) dissolved in CD_2Cl_2 at −80 °C (0.5 mL) was placed in an NMR tube under Ar, and *tert*-butylethylene (8.3 μL, 0.065 mmol) was then added at −80 °C. The ^1H NMR data indicated that an equilibrium between **4a** and the tbe adduct **5** had been established (Table IV). K_{eq} was determined by integration, or for more strongly binding olefins, which gave ≈100% of **5**, the equilibrium between cyclooctene and the olefin was studied.

(v) With D_2. To a solution of **4a** as in i above was added D_2. Early in the conversion, the mixed species $[IrHD(O_2CCF_3)L_2]$ was observed, but the final product was **4a**-d_2 (^1H, ^2H, and ^{31}P NMR; see text).

NMR Observation of Reactions of 9. (i) With CO. CO (1 atm) was bubbled through a solution of **9** (15 mg, 0.015 mmol) in CD_2Cl_2 (0.5 mL) to give **11a** and **11b**. ^1H NMR (resonances for **11a** given first; **11a**:**11b** intensity ratio 2:1): δ −20.15 (t, 14.5 Hz), −6.03 (t, 16 Hz), 7.02 and 7.14 (t, aryl CH ortho to F), 7.31 and 7.51 (c, aryl CH ortho to P). IR: 2038 and 2012 cm^{-1} (s, CO).

(ii) With I_2. Free C_6H_6 was observed when excess I_2 was added to a solution of **9** identical with that used above. The organometallic products were not characterized.

(iii) With H_2 and D_2. A solution of **9** as above was treated with H_2 (1 atm, 5 min), and free C_6H_6 was observed in the ^1H NMR. When D_2 was used, the C_6H_6 formed contained no D (GC/MS).

Acknowledgment. We thank the National Science Foundation (R.H.C.) and the Department of Energy (M.J.B.) for funding.

Relative Reactivities and Mechanistic Aspects of the Reactions of Organic Halides with Alkali Metals in Alcohol Environments

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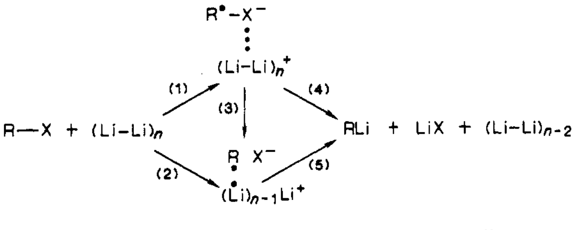
Abstract: The relative reactivities of organic halides over wide concentration ranges have been determined with limited amounts of lithium, sodium, and potassium in 2-ethoxyethanol (**1**) at 0 °C. Under these conditions the organometallics formed protonate to their hydrocarbons rather than undergo exchange, elimination, and simple or crossed coupling. In dilute solution in **1** the relative reactivities (r_1/r_2) of varied halides with lithium are essentially structure independent. However, as the concentrations of the halides increase, their relative reactivities become significantly different and depend on the total concentrations $\{C_T\ (M) = [R_1X] + [R_2X]\}$ of the organic halides. With lithium at increased halide concentrations (1) the reactivities are iodides > bromides > chlorides, (2) halides of lower molecular weight react more rapidly than their higher homologues, and (3) the reactivity orders of chlorides are (a) allyl > primary > secondary > tertiary > neopentyl, (b) 2-buten-1-yl > 1-buten-3-yl, (c) benzyl > phenyl, and (d) *p*-chlorotolyl > *o*-chlorotolyl > *m*-chlorotolyl. As examples, the relative reactivities of 1-chlorobutane/2-chloro-2-methylpropane (C_T = 5.83 M), 3-chloropropene/1-bromobutane (C_T = 4.60 M), bromobenzene/*p*-chlorotoluene (C_T = 4.37 M), and benzyl chloride/chlorobenzene (C_T = 4.02 M) are 6.71, 5.43, 24.1, and 22.1, respectively. Additions of aprotic solvents to 1-chlorobutane and 2-chloro-2-methylpropane in **1** decrease the relative reactivities of the halides. The effectiveness of cosolvents in lowering the relative reactivities of lithium with 1-chlorobutane and 2-chloro-2-methylbutane is tetrahydrofuran > dioxane ≅ 2-ethoxyethanol (**1**) > cyclohexene ≅ benzene. The relative reactivities of halides with sodium and with potassium in **1** at 0 °C are also total halide concentration (C_T) dependent. Under comparable concentrations the relative reactivity differences of halides are greater with lithium than sodium than potassium. The reactivities of halides under conditions of chemical control can be correlated with the ionization potentials of the alkali metals, and the kinetically controlling features of these systems are different from those with magnesium. The behavior of the alkali metals, the effects of concentration, and the roles of solvents on the reactivities of halides are discussed on the basis of (1) the active sites on the metal surfaces as modified by induction and (2) steric and electronic factors in the organic substrates. The kinetically controlled reactions of lithium with sp^3 halides may be interpreted to involve formation of lithio organohalide radical anions ($R^{\bullet}\text{–}X^-,Li^+$), electron transfer to the lithio radical anions on the metal surface, or unsymmetrical four-center carbanionic processes on the metal. In addition to incorporating an electron into the lowest unoccupied σ level of its C–X bond, an sp^2 halide offers the possibility for kinetically controlling electron transfer into the π system of its carbon–carbon double bond(s).

Varied organic halides react with alkali metals (Li, Na, and K) to yield alkali metal derivatives.[1] Although such organometallics are widely used, knowledge of their structures[2] and their mechanisms of formation is limited.[3] Of particular significance with respect to reaction mechanism is that lithium reacts with chiral 1-halo-2,2-diphenylcyclopropanes with major retention (68–85%) of stereochemistry and the stereospecificities of the conversions are related to the halide (Cl > Br > I) and to the sodium content and the particle size of the lithium.[3h–n,4] The reactions of lithium and halides have been discussed[3h,n] (Scheme I, processes 1–5) on the basis of transfer of a single electron to

Scheme I



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bonded halogen resulting in formation of a radical anion on the metal surface (process 1) or a loose radical pair (process 2). The radical anion can collapse (process 4) to the lithium reagent or dissociate (process 3) to a loose radical pair followed by formation of the lithium reagent (process 5) with racemization. Of further interest is that reactions of lithium with halides in dilute solution are first order relative to halide concentration[3g,m] and apparently first order with respect to surface area of lithium.[3g,m,5] The activation energies for reactions of lithium with bromoethane, 1-bromobutane, 1-bromo-2,2-dimethylpropane, and bromobenzene in dilute solution in ethyl ether are 4.5, 5.7, 13.8, and 9.1 kcal/mol, respectively,[3g] and for lithium with ethyl bromide ($C_{\text{R-X}} < 0.15$ M) and bromobenzene ($C_{\text{R-X}} = 0.01$ M) in tetrahydrofuran are 2.7 and 3.4 kcal/mol,[3m] respectively.

Results and Discussion

An aspect of metalation of halides that has not been studied is the effect of concentration of organic halides on the kinetics of reaction. Investigation has now been initiated of the *relative reactivities of various organic halide pairs*, over wide ranges of concentrations, with limited amounts of alkali metals. The competitive kinetic study was predicated on the basis that such information will be meaningful with respect to reaction mechanism, problems associated with differences in surface area and surface activity of the metals would be minimized or eliminated, and the organometallics produced (R_1M and R_2M; eq 1 and 2) could be

$$R_1X + 2M \xrightarrow[-MX]{k_1} R_1M \xrightarrow[-MOH]{H_2O} R_1H \qquad (1)$$

$$R_2X + 2M \xrightarrow[-MX]{k_2} R_2M \xrightarrow[-MOH]{H_2O} R_2H \qquad (2)$$

$$M = \text{Li, Na, and K}$$

determined conveniently upon hydrolysis of the reaction mixtures and gas chromatographic analysis of the hydrocarbon pairs (R_1H and R_2H) formed. The relative reactivities, r_1/r_2, for reactions of R_1X and R_2X in great kinetic excess could then be determined by eq 3, where R_2X/R_1X are the mole ratios of the organic halides

$$\frac{r_1}{r_2} = \frac{R_2X}{R_1X}\,\frac{R_1H}{R_2H} \qquad (3)$$

and R_1H/R_2H are the mole ratios of the hydrocarbons resulting from hydrolyses of the R_1M and the R_2M (eq 1 and 2) produced.

Table I. Coupling of Halides by Lithium in Ethyl Cellosolve (**1**) and in Ethyl Cellosolve (**1**)/Tetrahydrofuran

halide[a] R–X	molar concn of R–X	**1**, mL	THF, mL	Li, mmol	R–R[b] mol × 10^5	% R–R
1-bromobutane	1.92	20		2.16	9.0	4.2
2-bromobutane	2.21	20		2.16	5.8	2.7
	1.94	20		2.16	2.4	1.1
1-chlorobutane	2.28	20		2.16	2.9	1.3
2-chlorobutane	2.20	20		2.32	<0.5	<1.0
1-iodopropane	2.00	20		2.16	11.4	5.3
	2.00	20		2.60	11.5	4.4
	2.00	15	5	2.45	8.7	3.6
	2.00	10	10	2.60	3.3	1.3
	2.00	5	15	2.16	17.6	8.1
	2.00	1	19	2.60	47.3	18.0
2-iodopropane	2.00	20		2.16	7.4	3.4
	2.00	15	5	2.45	6.2	2.5
	2.00	10	10	3.03	2.1	0.7
	2.00	5	15	2.60	7.6	2.9
bromobenzene	1.91	20		3.31	0	0
	1.91	20		3.45	0	0

[a] All experiments were conducted at 0 °C with constant rapid stirring. [b] Moles of coupling product were determined by gas chromatographic analysis using an internal standard (hexane, heptane, 2,3-dimethylbutane, or naphthalene).

The relative reactivities of halide pairs were first studied in ethyl ether with limited amounts of lithium.[6] The systems could not be investigated advantageously because of (1) exchange of the organolithium derivatives with the organic halides present (eq 4)

$$R_1Li + R_2X \rightleftarrows R_1X + R_2Li \qquad (4)$$

to yield lithium products of different compositions than initially formed, (2) simple and crossed coupling of the resulting organolithium compounds with the organic halides present to give R_1R_1, R_2R_2, and R_1R_2, and (3) elimination of the organic halides where possible by the lithium compounds generated to yield unsaturated [R_1(–H) and R_2(–H)] and saturated [R_1H and R_2H] products. Complications 2 and 3 could be minimized or eliminated by use of larger quantities of lithium, shortening the reaction times, and detailed analyses of the reaction products. The complication arising from exchange of the organolithium products as in eq 4 could not be avoided, however. The relative rates of reaction of halides with lithium were then investigated in solvents which effect protonative destruction of the organolithium products to hydrocarbons, R_1H and R_2H, at much greater rates than the reactions having the complications described earlier.

A study was first made of the behavior of 1-bromobutane and lithium in alcohols. With primary alcohols (methyl, ethyl, and *n*-propyl) and 1-bromobutane (1.5 M) in the presence of lithium, the predominant reactions are between the alcohols and lithium to give lithium alkoxides and hydrogen. Mixtures of secondary or tertiary alcohols (isopropyl, *sec*-butyl, *tert*-butyl, and *tert*-amyl) with 1-bromobutane and lithium form coatings on the metal. 2-Ethoxyethanol (ethyl Cellosolve (**1**)), although a primary alcohol, reacts much more slowly with lithium than does *n*-propyl alcohol and hydrogen evolution nearly ceases when 1-bromobutane (0.5–1.0 M) is present.[7] No detectable reaction occurs between **1** and lithium at 1-bromobutane concentrations above 1.0 M and

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(1) (a) Wakefield, B. J. *The Chemistry of Organolithium Compounds*; Pergamon Press: New York, 1974; pp 21–26. (b) Schöllkopf, U. In *Methoden Der Organischen Chemie, Metalloorganische Verbindungen*, 4th ed.; Müller, E., Ed.; G. Thieme Verlag: Stuttgart, 1970; Vol. 13, Part 1, pp 1–25.

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(3) (a) Tarbell, D. S.; Weiss, M. *J. Am. Chem. Soc.* **1939**, *61*, 1203–1205. (b) Braude, E. A.; Coles, U. A. *Ibid.* **1951**, *73*, 2078–2084. (c) Bordwell, F. D.; Landis, P. S. *Ibid.* **1957**, *79*, 1593–1597. (d) Curtin, D. Y.; Crumpt, J. W. *Ibid.* **1958**, *80*, 1922–1926. (e) Allinger, N. L.; Herman, R. B. *J. Org. Chem.* **1961**, *26*, 1040–1042. (f) Seyferth, D.; Vaughn, L. G. *J. Organomet. Chem.* **1963**, *1*, 201–204. (g) Yao, C.-Y. Ph.D. Dissertation, The Ohio State University, Columbus, OH, 1963; *Diss. Abstr.* **1964**, *24* (11), 4414. (h) Walborsky, H. M.; Aronoff, M. S. *J. Organomet. Chem.* **1964**, *4*, 418–420. (i) Hsieh, H. L. *Ibid.* **1967**, *7*, 1–10. (j) Dewar, M. J. S.; Harris, J. M. *J. Am. Chem. Soc.* **1969**, *91*, 3652–3653. (k) Grovenstein, E., Jr.; Cheng, Y. M. *Chem. Commun.* **1970**, 101–102. (l) Walborsky, H. M.; Aronoff, M. S. *J. Organomet. Chem.* **1973**, *51*, 55–75. (m) Cameron, G. G.; Milton, A. J. S. *J. Chem. Soc., Perkin Trans. 2* **1976**, 378–382. (n) Walborsky, H. M.; Banks, R. B. *Bull. Soc. Chim. Belg.* **1980**, *89*, 849–868.

(4) Retention of stereochemistry in metalation of 1-halo-2,2-diphenylcyclopropanes increases with the sodium content and decreases with the particle size of the lithium.[3h,l]

(5) References 3g and 3m report for formation of organolithium compounds from halides and lithium that the reactions are first order with respect to concentration of the halide and to the initial surface area of the metal. In an actual kinetic experiment lithiation of a halide is therefore zero order with respect to the metal.[3g] The zero-order dependence has been explained on the basis of (1) maintenance of constant surface area in which there is compensation of the decrease in surface due to consumption of metal by an increase in surface due to pitting processes or (2) constancy of the total number of active sites on the lithium surface throughout reaction even though there may be significant changes in the surface area.

(6) The alkali metals used in this research were reagent grade. The percentages of other metals, particularly sodium, in these commercial products are unknown. In the discussions of the present results, the possible contributions of trace metals to the mechanistic processes are ignored. Studies are to be initiated of the relative reactivities of various halides in **1** with lithium containing known percentages of sodium.

(7) Slack and Phelan (Slack, G. C.; Phelan, J. C., private communication, State University of New York, Potsdam, NY) have found that the activation energy for reaction of lithium with pure **1** is 10.6 kcal/mol. The activation energies for reactions of lithium with **1** containing halides at varied concentrations are presumably somewhat larger.

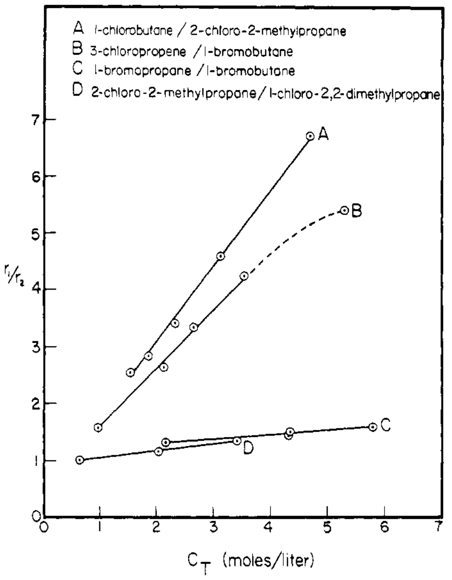


Figure 1. Effect of total halide concentration [C_T (M), eq 5] on the relative reactivities (r_1/r_2) of the halide pairs (A) 1-chlorobutane/2-chloro-2-methylpropane, (B) 3-chloropropene/1-bromobutane, (C) 1-bromopropane/1-bromobutane, and (D) 2-chloro-2-methylpropane/1-chloro-2,2-dimethylpropane with lithium in **1** at 0 °C.

Table II. Relative Reactivities (r_1/r_2) of Halide Pairs (R_1X/R_2X) with Lithium in **1** at 0 °C

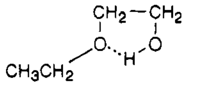
halide pair, R_1X/R_2X	R_2X/R_1X[a]	C_T[b]	r_1/r_2[c]
1-iodobutane/1-bromopropane	0.90	2.46	1.89
2-bromobutane/2-chloropropane	1.24	3.42	3.22
	1.82	3.42	3.28
1-bromobutane/2-chloropropane	1.18	5.13	3.20
	1.65	5.18	3.28
1-chlorobutane/2-chloropropane	0.78	3.53	1.88
	1.03	3.53	1.82
2-chloropropane/2-chloro-2-methylpropane	1.36	3.01	2.94
	1.40	3.01	3.00
1-chlorobutane/2-chloro-2-methylpropane	1.05	1.55	2.54
	1.15	1.86	2.83
	1.15	2.34	3.42
	1.15	3.11	4.58
	1.15	4.68	6.71
2-chloro-2-methylpropane/1-chloro-2,2-dimethylpropane	0.99	2.17	1.33
	1.21	4.32	1.46
	0.99	4.35	1.51
1-bromopropane/1-bromobutane	1.42	0.64	1.02
	1.11	2.40	1.17
	1.28	3.40	1.37
3-chloropropene/1-bromobutane	1.18	0.96	1.58
	1.18	2.12	2.65
	1.18	2.65	3.35
	1.18	3.53	4.24
	1.18	5.30	5.43
1-chloro-2-butene/1-bromopropane	1.01	3.36	1.65
1-bromobutane/2-bromopropene	0.99	3.36	1.21
2-bromobutane/2-bromopropene	0.74	3.30	1.01
2-bromopropane/1-chlorobutane	1.50	2.02	1.37
	1.50	2.52	1.60
	1.50	5.06	2.20
3-chloro-1-butene/1-bromopropane	0.65	3.33	1.18
benzyl chloride/bromobenzene	0.96	1.06	2.01
	1.02	2.10	4.25
	1.00	3.08	8.19
	0.95	4.16	11.6
	0.98	5.09	14.0
benzyl chloride/chlorobenzene	1.12	1.07	2.58
	0.90	2.07	6.61
	1.03	3.06	15.4
	0.96	4.02	22.1
bromobenzene/*p*-chlorotoluene	0.99	0.47	0.80
	0.99	0.89	1.16
	0.98	1.27	1.88
	1.99	2.60	8.13
	1.99	3.49	16.0
	1.99	4.37	24.1
bromobenzene/*m*-chlorotoluene	0.99	0.55	1.06
	0.99	1.06	2.00
	0.99	2.16	9.30
	0.99	3.01	21.1
	1.92	4.20	31.9
bromobenzene/*o*-chlorotoluene	0.98	1.10	1.64
	0.98	2.72	10.9
	0.98	3.06	12.6
	1.93	3.59	19.5
	1.97	4.33	22.5

[a] The ratio R_2X/R_1X represents the initial mole ratio of the two reactant halides in great excess relative to lithium. This ratio is thus essentially equal to that for the two halides after reaction with lithium. [b] The total molar concentration of the halide pair initially present (eq 5). [c] The relative reactivities of R_1X and R_2X calculated as in eq 3.

the lithium surface is ***bright*** and ***shiny*** throughout the entire reaction of the halide.[7] The contrast in the rates of reaction of **1** and *n*-propyl alcohol with lithium may well be the result of the internal hydrogen bonding and steric factors in **1** as **2**. Ethyl Cellosolve (**1**) also greatly or totally eliminates coupling of halides

$CH_3CH_2OCH_2CH_2OH$

1



2

by lithium (Table I) and exchange of organolithium products with halides (eq 4). Although it is not clear whether the coupling processes involve (1) radical pairing ($2R^\bullet \rightarrow RR$) on lithium and/or (2) bimolecular displacement (S_N2; $RLi + RX \rightarrow RR + LiX$), such reactions are unimportant (<6%) for 1- and 2-iodopropanes, 1- and 2-bromobutanes, and 1- and 2-chlorobutanes and undetectable for bromobenzene at halide concentrations up to 2.5 M upon exposure to lithium (0.1 equiv) in **1**. Coupling is also minor for halides and lithium in tetrahydrofuran when the systems contain appreciable **1**. Further, exchange of 2-iodopropane and 1-bromopropane (total concentration = 2.5 M) to 2-bromopropane and 1-iodopropane does not occur upon reaction with lithium (0.1 equiv) in **1**.

Investigation was then made of the relative reactivities (r_1/r_2) of various halides with lithium (0.05–0.10 equiv) in rapidly stirred **1** at 0 °C. The competitive method was used throughout. There were no major complications involving reactions of lithium with **1** in the presence of the halides, the experiments are reproducible, the lithium remains clean throughout the reactions, and the relative reactivities are independent of stirring rate. The experimental data were treated as in eq 3 and the results for various halides are summarized in Table II. In relatively dilute solutions of initial halide pairs (Table II, Figures 1 and 2), the effects of structure on the relative rates of reactions of halides are small. As revealed, however, in Table II (Figures 1 and 2), as the concentrations of halide pairs in **1** are increased, the reactivity differences between halides become significantly larger. The rate ratios (Table II) at increased halide concentrations lead to the following initial conclusions: (1) iodides are more reactive than bromides than chlorides; (2) lower molecular weight halides react more rapidly than their higher homologues; (3) the reactivity order of chlorides is (a) allyl > primary > secondary > tertiary > neopentyl, (b) 2-buten-1-yl > 1-buten-3-yl, and (c) benzyl > phenyl; (4) bromobenzene lithiates more rapidly than chlorobenzene; and (5) the relative reactivities of the following halorenes are bromobenzene > *p*-chlorotoluene ≳ *o*-chlorotoluene > *m*-chlorotoluene.

A striking aspect of the relative reactivities, r_1/r_2 (eq 3), in **1** is that they depend systematically on the total concentrations, C_T (M, eq 5), of the organohalide pairs. For the systems in Table II and as illustrated in Figures 1 and 2, the rate ratios generally increase linearly or sigmoidally, approaching an upper limit at higher halide concentrations. Of major interest is that the rate

$$C_T = [R_1X] + [R_2X] \qquad (5)$$

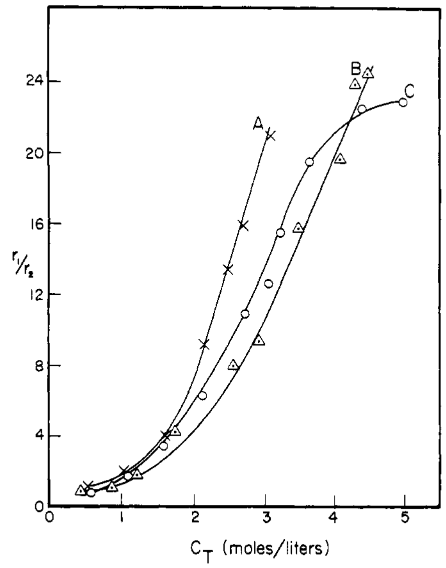


Figure 2. Effect of total halide concentration [C_T (M), eq 5] on the relative reactivities (r_1/r_2) of the halide pairs (A) bromobenzene/*m*-chlorotoluene, (B) bromobenzene/*p*-chlorotoluene, and (C) bromobenzene/*o*-chlorotoluene with lithium in **1** at 0 °C.

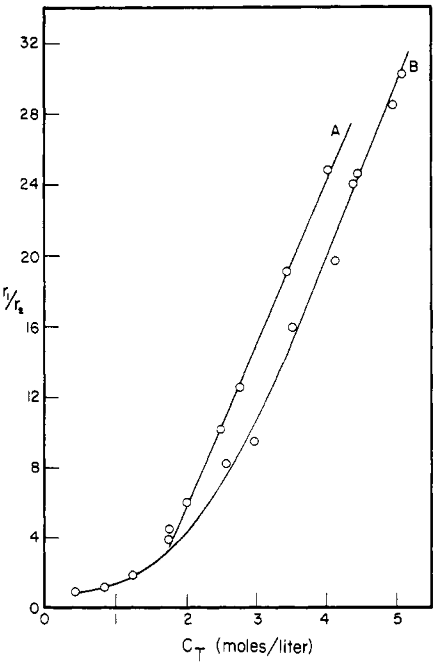


Figure 3. Curve A: effect of total halide concentration [C_T (M), eq 5] on the relative reactivities (r_1/r_2) of the halide pair bromobenzene/*p*-chlorotoluene with lithium in **1** at 0 °C upon addition of increasing amounts of 1-chloro-2-ethylbenzene. Curve B: effect of increasing total halide concentration [C_T (M), eq 5] on the relative reactivities (r_1/r_2) of the halide pair bromobenzene/*p*-chlorotoluene with lithium in **1** at 0 °C.

ratios can become quite large (2 to >20). An additional feature of these results is that the large C_T effects are inconsistent with overall reduction reactions involving dissolving-lithium processes in which there is kinetically important transfer of solvated electrons in solution in **1** to the halides, halogen loss, and protolysis.[8]

Investigations were then extended to a halide trio in which there is competition for lithium. Thus, the amounts of bromobenzene

Table III. Relative Reactivities (r_1/r_2) of Bromobenzene (R_1X) and *p*-Chlorotoluene (R_2X) with Lithium at Fixed Concentration (C_T) in **1** at 0 °C upon Addition of 1-Chloro-2-ethylbenzene (R_3X)

$R_1X + R_2X$, mmol	R_3X, mmol	$C_T{}^a$	$C_T{}^b$	$r_1/r_2{}^c$
88.7	0	1.77	2.65	3.90
88.5	12.4	1.77	2.02	6.04
88.5	34.9	1.77	2.47	10.1
88.5	48.1	1.77	2.73	12.5
88.5	80.2	1.77	3.37	19.2
88.5	108.4	1.77	3.94	24.8

[a] The total millimoles of R_1X and R_2X (eq 5). [b] The total millimoles of R_1X, R_2X, and R_3X. [c] The relative reactivities of bromobenzene (R_1X) and *p*-chlorotoluene (R_2X) with lithium in the presence of 1-chloro-2-ethylbenzene (R_3X).

Table IV. Relative Reactivities (r_1/r_2)[a] of 1-Chlorobutane (R_1X) and 2-Chloro-2-methylpropane (R_2X) in **1** and Various Cosolvents at 0 °C

1, mL	cosolvent	cosolvent vol, mL	$C_T{}^b$	r_1/r_2
5	tetrahydrofuran	5	4.66	3.70
10		10	3.11	1.38
10		20	2.34	1.21
10		30	1.86	1.09
10		50	1.33	1.07
10	dioxane	10	3.11	4.41
10		20	2.33	3.43
10		30	1.86	2.83
10		50	1.34	2.06
10	cyclohexane	10	3.26	5.55
10		20	2.44	5.26
10		30	1.96	5.15
10		50	1.39	4.92
5		50	0.75	4.71
10	cyclohexene	10	3.10	5.47
10	benzene	10	3.10	4.87

[a] The relative reactivities (eq 3) of 1-chlorobutane (R_1X) compared to 2-chloro-2-methylpropane (R_2X). [b] Total concentration (M) of R_1X and R_2X (eq 5) in the cosolvent mixture.

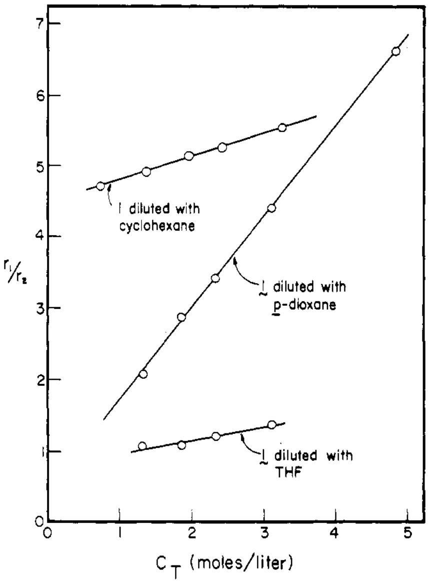


Figure 4. Effect of total halide concentration [C_T (M), eq 5] on the relative reactivities (r_1/r_2) of the halide pair 1-chlorobutane/2-chloro-2-methylpropane with lithium at 0 °C in **1** diluted with the indicated cosolvents.

(R_1X) and *p*-chlorotoluene (R_2X) were kept constant and the C_T values ($R_1X + R_2X + R_3X$) of the system were increased by adding 1-chloro-2-ethylbenzene (R_3X). As seen in Table III and Figure 3, increasing the amount of 1-chloro-2-ethylbenzene (R_3X) markedly increases the selectivity of the system. The above results

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[8] For discussion of dissolving-metal reductions and electron-capture reactions of halides, see: (a) House, H. O. *Modern Synthetic Reactions*, 2nd ed.; W. A. Benjamin: Menlo Park, CA, 1972; pp 145–227. (b) Hart, E. J.; Anbar, M. *The Hydrated Electron*; Wiley-Interscience: New York, 1970; pp 138–141.

Table V. Relative Reactivities (r_1/r_2) of Halide Pairs (R_1X/R_2X) at Various Concentrations (C_T) in **1** with Sodium and Potassium at 0 °C

halide pair, R_1X/R_2X	R_2X/R_1X[a]	C_T[b]	r_1/r_2[c]
3-chloropropene/1-bromobutane			
sodium	1.14	1.76	1.89
	1.14	2.64	2.85
	1.14	5.28	5.01
potassium	0.92	2.64	2.04
	0.92	3.46	5.28
benzyl chloride/chlorobenzene			
sodium	1.12	1.07	1.75
	0.90	2.07	3.55
	1.03	3.06	4.79
	0.96	4.02	8.94
	2.63	4.85	10.7
potassium	0.90	2.07	1.42
	1.03	3.06	1.66
	0.96	4.02	2.03
	2.63	4.85	4.23
bromobenzene/*p*-chlorotoluene			
sodium	0.97	1.04	0.91
	0.97	1.66	2.68
	0.97	2.68	6.91
	0.97	3.54	9.87
	1.95	4.94	14.3
potassium	0.98	1.34	0.78
	0.98	2.14	1.19
	0.98	3.01	1.72
	0.97	3.88	2.78
	0.97	5.03	3.89
bromobenzene/*o*-chlorotoluene			
sodium	1.09	0.57	1.08
	1.65	2.12	4.30
	2.00	3.21	8.87
	1.97	4.94	17.9
	2.00	3.21	1.94
	1.80	4.33	4.25
	2.04	5.23	5.82

[a] The mole ratio R_2X/R_1X of the two reactant halides. [b] The total molar concentration of the halide pair initially present (eq 5). [c] The relative reactivities of R_1X and R_2X calculated as in eq 3.

are thus of practical interest in that, under preparative conditions, there can be large differences in the relative rates of conversions of organic halides by lithium to organolithium compounds.

Study was then made of the effects of aprotic solvents on the relative reactivities of 1-chlorobutane (R_1X) and 2-chloro-2-methylpropane (R_2X) upon admixture with **1**. The systems were investigated by keeping the quantities of halides (R_1X and R_2X) and **1** constant and varying the total molar concentration, C_T, by adding an aprotic solvent.

Dilution with tetrahydrofuran, dioxane, or cyclohexane (Table IV, Figure 4), results in linear decreases in the relative rate constants (r_1/r_2). Also, the effects of cyclohexene and benzene (Table IV) are similar to that of cyclohexane. Further, reductions in the reactivity ratios by dioxane are essentially identical with that resulting solely by dilution with **1**. The overall order of effectiveness of cosolvents in lowering the reactivities of 1-chlorobutane (R_1X) relative to 2-chloro-2-methylpropane (R_2X) is tetrahydrofuran > dioxane $\simeq$ **1** > cyclohexane $\simeq$ cyclohexene $\simeq$ benzene and apparently reveals the significance of electron-donor effects by environments on the activation processes for reactions of halides with lithium.

Investigation was then made of the relative reactivities of halide pairs with lithium, sodium, and potassium, respectively, in **1** at 0 °C. The halide pairs were 3-chloropropene/1-bromobutane, bromobenzene/*p*-chlorotoluene, bromobenzene/*o*-chlorotoluene and benzyl chloride/chlorobenzene. The results (Table V) reveal that, as C_T increases, the reactivity ratios for pair members with sodium and potassium increase (Figure 5) much as with lithium. Under comparable conditions, however, the relative selectivities of these metals are Li > Na $\gg$ K. In fact, the relative reactivities of pair members at fixed C_T concentrations correlate linearly (Figure 6) with the ionization energies of the reactant metals (Li > Na $\gg$ K).[9] Further, the C_T effects with the different alkali metals also argue against reduction of the halides by solvated electrons in the hydroxylic solvent systems.[8]

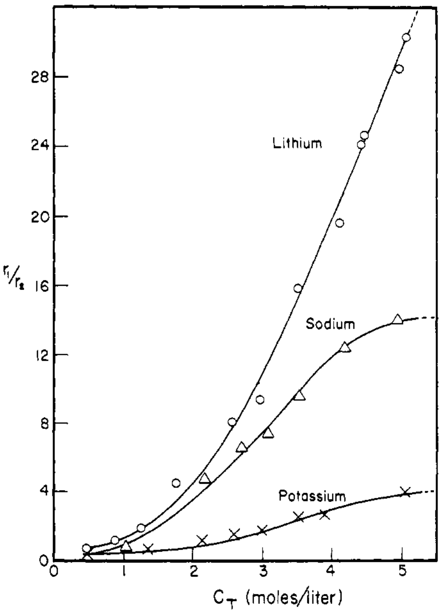

Figure 5. Effect of total halide concentration [C_T (M), eq 5] on the relative reactivities (r_1/r_2) of the halide pair bromobenzene/*p*-chlorotoluene with lithium, sodium, and potassium, respectively, in **1** at 0 °C.

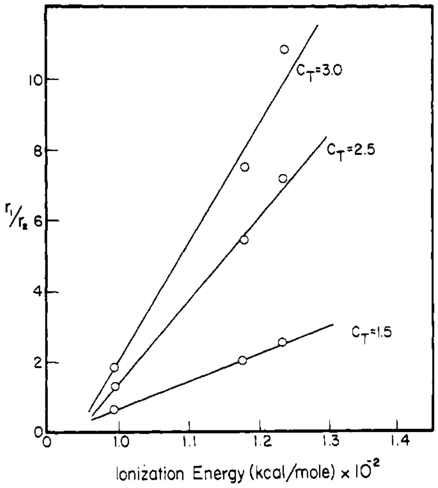

Figure 6. Correlation of the ionization energies of lithium (123.60 kcal/mol), sodium (118.06 kcal/mol), and potassium (99.62 kcal/mol) with the relative reactivities (r_1/r_2) of the metals with the halide pair bromobenzene/*p*-chlorotoluene at the indicated total halide concentrations [C_T (M), eq 5] in **1** at 0 °C.

The above results raise significant questions with respect to reaction mechanism and to practice in preparative chemistry. Among the points of interest are the following: (1) the reactions of organic halides and alkali metals in **1** and various cosolvents are heterogeneous; (2) with various halides at *increasing concentrations* in **1**, an alkali metal becomes a significant kinetically discriminating reactant; (3) the effects of structure on the reactivities of halides in relatively concentrated solution in **1** are greater with lithium than sodium than potassium; (4) the reactivity differences of various halides with lithium are less in ether than in hydrocarbon solvents; (5) the kinetic behavior of organic halides with alkali metals does not correlate with the electrochemical reduction potentials of the halides;[10] and (6), as will be seen,

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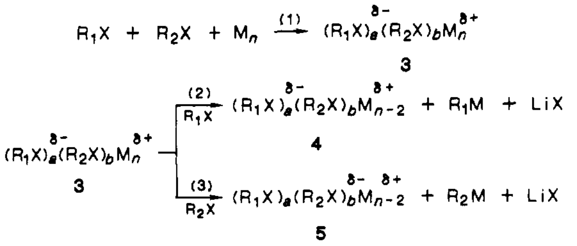
[9] Pauling, L. *The Nature of the Chemical Bond*; Cornell University Press: Ithaca, NY, 1960.

Scheme II



mechanistic principles can explain the differences in the reactivities of alkali metals with varied halides in concentrated solution.

In the present experiments the reactions between solid alkali metals and organic halides in **1** and cosolvents presumably occur at surface–liquid interfaces. The overall reaction sequences thus have the usual complications of solid–liquid systems in that they involve (1) diffusion of organic halides to the metal,[11] (2) physical adsorption of the halides on the surface of the metal, (3) chemical processes on the metal resulting in metalation, (4) physical desorption of the products from the surface, and (5) diffusion of the products into solution plus (6) protolysis of the products which might occur during or after processes 4 or 5. Diffusion of organic halides to and adsorption on the alkali metals (processes 1 and 2) in the presence of **1** are expected to contribute to the relative reactivities presently reported.[11] As will be amplified greatly, the kinetic behavior of metals with halides in **1** can be interpreted on the basis of mechanism paths in their conversions to organometallics by direct reactions with the metals (process 3). Alternative questions that arise immediately then are whether (a) desorption or diffusion as in processes 4 and 5[12] or (b) protolyses of the products (process 6) kinetically control the present reduction systems. A response to the first question is that desorption and diffusion as in processes 4 and 5 do not lead to comprehensive rationalization of the present kinetic data.[12] As for the second question, selective protolytic destruction of the organometallics on the metal surface appears highly unlikely and also fails to explain significant portions of the relative reactivity results observed.[13] Chemical rationalizations of the behavior of halides with lithium, sodium, and potassium are now to be described and discussed.

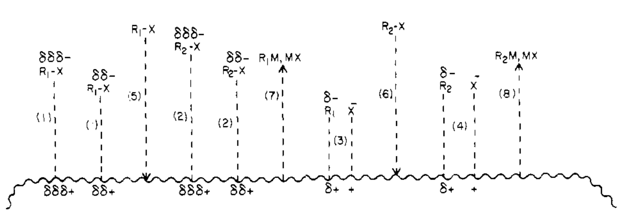


Figure 7. Representation of the behavior of a metal undergoing competitive reactions with two halides at any given time as in Scheme II.

In the initial studies of the reactions of lithium with halide pairs in *dilute solution* in **1**, the reactivities of varied halides were found to be similar and small. Under such conditions lithium is an indiscriminate reagent, possibly because the rate-determining steps involve diffusion (mass transport) of the halo organics to the metal surface.[11] This interpretation is consistent with the small activation energies (usually less than 6 kcal/mol)[3g,m,14] for reactions of lithium with varied halides in dilute solutions.

Reactions of lithium with halide pairs in relatively concentrated solution lead to decided kinetic responses in the systems. As the concentrations are increased, the reactivities of members of halide pairs respond differentially, the kinetic behavior of a halide is related to its structure, chemical processes apparently control the rates of reaction, and *most importantly lithium becomes progressively a more discriminating reactant.* Although such effects are not as great with sodium and potassium, the relative chemical responses of halides with these alkali metals parallel those with lithium.

The above results for reactions of metals (M = Li, Na, and K) with halides are to be analyzed initially on the basis of the overall sequences in Scheme II and in particular to the observations that the reactivity of a metal is halide concentration dependent. Certain aspects of the heterogeneous reactions of halides with metals are described in detail because the specific features and the mechanistic disciplines are quite different from those usually experienced in homogeneous systems. The hydrolyses of the organometallics (R_1M and R_2M) to their hydrocarbons (R_1H and R_2H) are not illustrated or elaborated because these conversions are assumed to be much faster than formation of their precursors.[13]

Step 1 in Scheme II represents a metal (M_n) undergoing irreversible dissociative chemisorption with a halide pair to give **3**. Competition between R_1X and R_2X for reaction with the metal surface occupied as in **3** occurs with loss of two atoms of metal to give **4** and **5** (steps 2 and 3). These processes continue until all of the metal is consumed. Subscripts *a* and *b* are the numbers of R_1X and R_2X molecules engaged as such *per unit area of metal.* Thus, since R_1X and R_2X are in *large excess* in the present experiments, the number of molecules, $a + b$, involved on a unit surface area of the metal is *constant* throughout reactions at given halide concentrations. When the concentrations of R_1X and R_2X are increased or decreased, $a + b$ for the surface area of metal increases or decreases. The new values of $a + b$ are effectively *constant*, however, throughout reactions of the unit surface under

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[10] (a) The general ease of electrochemical reductive dehalogenation of alkyl halides is tertiary RX > secondary RX > primary RX and thus is opposite to the present results for reactions of such halides with alkali metals in **1**. (b) Lambert, F. L.; Kobayashi, K. *J. Am. Chem. Soc.* **1960**, *82*, 5324–5328. (c) Lambert, F. L. *J. Org. Chem.* **1966**, *31*, 4184–4188. (d) Hawley, M. D. In *Encyclopedia of Electrochemistry of the Elements*; Bard, A., Ed.; Marcel Dekker: New York, 1980; Vol. 14, Chapter 1, p 14.

[11] (a) Rogers et al. (Rogers, H. R.; Hill, C. L.; Fugiwara, Y.; Rogers, R. J.; Mitchell, H. L.; Whitesides, G. M. *J. Am. Chem. Soc.* **1980**, *102*, 217–226) report that with magnesium in ethyl ether (1) all organic iodides and many secondary alkyl bromides react at mass-transport or diffusion-controlled rates, (2) the rates of metalation of less reactive bromides (neopenty, phenyl, and cyclopropyl) contain contributions both from mass transport and from chemical reaction, and (3) most chlorides react at rates that are not mass transport limited. For further important discussion of mixing, mass transport, and diffusion on selectivity in chemical reactions, see: (b) Rys, P. *Acc. Chem. Res.* **1976**, *9*, 345–351. (c) Roberts, G. W. In *Catalysis in Organic Synthesis*; Rylander, P. N., Greenfield, H., Eds.; Academic Press: New York, 1976; pp 1–48.

[12] (a) If the rates of reactions of halides with lithium were controlled by desorption of the organolithium products (R–Li) generated, the observed reactivity order of allyl–X and benzyl–X > R_p–X > R_s–X > R_t–X leads to a conclusion that the adsorptivities are R_t–Li > R_s–Li > R_p–Li > benzyl–Li and allyl–Li. Such an adsorption order would appear to reflect the electron-donor or the nucleophilic (carbanionic) character of the lithium reagents generated. The reactivities iodides > bromides > chlorides of halides having identical carbon structures disagree, however, with desorptivity control since the organolithiums formed are presumably identical unless there are some remarkable desorptive effects from the different lithium salts (LiCl, LiBr, LiI) produced as coproducts. Although the relative reactivities for structurally identical organic halides are RI > RBr > RCl, this generalization fails as follows (Table II): 3-chloropropene > 1-bromobutene; 1-chloro-2-butene > 1-bromopropane; benzyl chloride > bromobenzene. The above exceptions appear to negate significant desorptive effects relative to the lithium salts. Finally, nucleophilicities (or steric assistance) leading to desorptive kinetic control do not account for the following reactivity order (Table II) for lithiation of related chlorides: 1-chlorobutane > 2-chloro-2-methylpropane > 1-chloro-2,2-dimethylpropane. (b) Since within the same types of compounds large molecules usually diffuse more slowly than similar smaller homologues, control of the rates of lithiation by diffusion of the organolithiums after desorption apparently disagrees with the following reactivity orders (Table II): 1-iodobutane > 1-bromopropane; 2-bromobutane, 1-bromobutane, and 1-chlorobutane > 2-chlorobutane; 1-chloro-2-butene > 1-bromopropane; R–I > R–Br > RCl for halides having identical carbon structures.

[13] The following reactivity orders indicate that lithiations of halides in **1** are not generally controlled kinetically by protolyses of the organolithium products on lithium: (1) 1-iodobutane > 1-bromobutane > 1-chlorobutane, (2) 1-iodobutane > 1-bromopropane > 1-bromobutane, (3) 1-bromobutane > 2-bromopropene > 1-chlorobutane, (4) benzyl chloride > bromobenzene > chlorobenzene, and, of particular note, (5) 1-chlorobutane > 2-chloropropane > 2-chloro-2-methylpropane > 1-chloro-2,2-dimethylpropane.

[14] For most liquids the activation energies of diffusion, E_D, are ~3–5 kcal/mol. See: Glasstone, S.; Laidler, K. J.; Eyring, H. *The Theory of Rate Processes*; McGraw-Hill: New York, 1941; pp 522–525. Benson, W. S. *The Foundations of Chemical Kinetics*; McGraw-Hill: New York, 1960; p 499. Shooter, D. In *Comprehensive Chemical Kinetics*; Bamford, C. H., Tipper, C. F. H., Eds.; Elsevier: Amsterdam, 1969; p 253. Evans, D. F.; Chan, C.; Lamartine, B. C. *J. Am. Chem. Soc.* **1977**, *99*, 6492.

the indicated conditions until the metal is depleted. Loss of area does not change the chemical nature of the metal and thus **3**, **4**, and **5** (Scheme II) are chemically equivalent.

Figure 7 represents a metal reacting with two halides at any given time as in Scheme II. Since many events occur on the surface which result in removal of layer after layer of atoms, the site reactivity of the metal is assumed to be uniform throughout. Processes 1 and 2 (Figure 7) portray early to late transitional stages of dissociative adsorption of the halides and processes 3 and 4 (Figure 7) represent completion of chemical interactions with the metal prior to desorption. Processes 5 and 6 (Figure 7) emphasize that an external halide must compete for metal atoms on a surface that is also undergoing varied interactions with other halide molecules. Chemical relaxation of the organometallic and metal halide products from the surface constitutes processes 6 and 7 (Figure 7).

In the dissociative interactions (Scheme II and Figure 7) of an organic halide with a metal the processes are expected to be structure, concentration, and environment dependent. Of particular interest is that the surface functions as an electron donor and becomes relatively positive during the chemisorptive processes. Since alkali metals are excellent conductors it might be expected that, *by induction*, the kinetic electronic demands at a reacting site will *lower* the activities of other sites. Therefore, as interactions of halides occur essentially simultaneously at more than a single location on a highly active surface, the metal becomes even less electron rich. Further, when the concentration of a halide is increased, more chemical events occur coincidentally, the surface becomes more positive per unit area, and the metal behaves as a more discriminating—a less reactive—electron donor. For metallic potassium to be less sensitive to concentration (C_T) effects than sodium than lithium comes from its lower ionization energy[9,15] and thus its greater ability to donate electrons from its surface (even) under conditions of intense attack.

Solvents also have important effects on the kinetic behavior of metals with halides. As indicated previously, lithium is a more indiscriminate reactant with halide pairs in ethers than in hydrocarbons. Ethyl ether and tetrahydrofuran also expedite reactions of lithium with organic halides. The relative reactivity and the solvent effectiveness for reactions of halides with lithium are now interpretable on the basis that ethers facilitate removal of lithio organic and lithium halide products from the lithium surface such that the *number* and the *reactivities* of its reaction sites will be increased or even maximized. Under such circumstances the surface is kinetically cleaner, the activity and the reactivity of the lithium per unit area are enhanced, and the metal is less discriminating as an electron donor to halides.

There are additional factors to be considered with respect to the kinetic effects of solvents on the reactions on lithium. Earlier it has been discussed that, in the chemically controlling transition states for conversions of halides to lithiated products, lithium surface sites become positive and these electrical effects are transmitted by *induction* along the surface of the metal. Ethers may thus coordinate and dissipate positive charge more effectively than hydrocarbons at various positions on the lithium surface. Further, the ability of an ether to facilitate reactions of halides with lithium can be related to the electron density and steric availability of the coordinating oxygen centers in the solvent. Perhaps even more significantly, the solvent coordination effects on the reactive lithium surfaces may be of such magnitudes that the reactivities of halides with the metal become essentially independent of the halide structure.

In the present study, emphasis has been given to the roles of the metal, concentration of halide reactants, solvent effects, and transport phenomena on the lithiation processes. Attention is now directed to the mechanisms and structural effects in the reactions of lithium with halides.

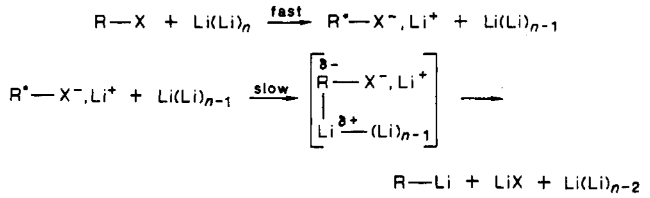
Scheme III

$$\mathrm{R{-}X} + \mathrm{Li(Li)}_n \xrightarrow{\text{slow}} \mathrm{R^{\bullet}{-}X^-,Li^+} + \mathrm{Li(Li)}_{n-1}$$

$$\mathrm{R^{\bullet}{-}X^-,Li^+} \xrightarrow{\text{fast}} \mathrm{R^{\bullet}} + \mathrm{LiX}$$

$$\mathrm{R} + \mathrm{Li(Li)}_{n-1} \xrightarrow{\text{fast}} \mathrm{R{-}Li} + {}^{\bullet}\mathrm{Li(Li)}_{n-2}$$

Scheme IV



$$\mathrm{R{-}X} + \mathrm{Li(Li)}_n \xrightarrow{\text{fast}} \mathrm{R^{\bullet}{-}X^-,Li^+} + \mathrm{Li(Li)}_{n-1}$$

$$\mathrm{R^{\bullet}{-}X^-,Li^+} + \mathrm{Li(Li)}_{n-1} \xrightarrow{\text{slow}} \left[\begin{array}{l} \overset{\delta-}{\mathrm{R}}\mathrm{{-}X^-,Li^+} \\ \mid \\ \mathrm{Li} \overset{\delta+}{-} \mathrm{(Li)}_{n-1} \end{array}\right] \longrightarrow \mathrm{R{-}Li} + \mathrm{LiX} + \mathrm{Li(Li)}_{n-2}$$

Among the important considerations in halide/lithium systems are (1) under conditions that are not transport limited, the rates of lithiation are a function of the halide concentration and the surface area of the metal, (2) reactions of 1-halo-2,2-diphenylcyclopropanes and lithium occur with major retention of stereochemistry, (3) the reactivities of alkyl halides are RI > RBr > RCl,[16] (4) acyclic chlorides follow the reactivity order allyl > primary > secondary > tertiary > neopentyl,[16] (5) allyl chloride lithiates more rapidly than 1-bromoalkanes, (6) halogen bonded to sp^2 carbon is only slightly less reactive than in a halogen/sp^3-carbon bond,[16] and (7) steric effects on the reactivities of halides are small.[16] Generalizations 2, 4, 6, and 7 immediately allow the conclusions that the lithiation reactions do not involve (1) S_N1-like ionizations and subsequent reduction of intermediate carbocations and (2) back-side S_N2-like displacements of the halides by the metal. Further, the ease of electrochemical reduction (half-wave potentials) of bromides (R–Br) and chlorides (R–Cl) to their corresponding hydrocarbons (R–H, an overall two-electron-transfer process) is allyl > tertiary > secondary > primary > vinyl and aryl.[10] Interestingly, the reactivities of chlorides (R–Cl) with magnesium to form Grignard reagents are also allyl > tertiary > secondary > primary > vinyl and aryl.[17] It is thus clear that there are significant differences in the rate-determining steps for lithiation compared to electrochemical reduction[10] and Grignard formation reactions of organic halides under conditions that do not involve mass transport or diffusion control.

As will be developed later, there may be substantial differences in the mechanisms of reactions of lithium with allyl, alkyl, and alkenyl halides. To facilitate discussion of these systems, two mechanisms that might be operational in lithiation of halides bonded to sp^3 carbon are outlined in Schemes III and IV. Both sequences meet the requirement that the reaction rates show first-order dependencies on the concentrations of the halides and the surface areas of the metal. The two mechanisms are of particular interest in that they avoid rate-determining transition

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(15) (a) Hammond, G. S. *J. Am. Chem. Soc.* **1955**, *77*, 334–338. (b) That the kinetic reactivities correlate with the ionization energies is perhaps to be expected since the metalation processes are low activation energy, exothermic, and reactant-like.[15a]

(16) In a prior study of the kinetics of reactions of halides with lithium in ethyl ether, the conclusions of ref 3g are similar to that of the present investigation.

(17) For further important research on the mechanisms of formation of Grignard reagents from organic halides and magnesium in ether solvents, see ref 3n and 10 and: (a) Kharasch, M. S.; Reinmuth, O. *Grignard Reactions of Non-Metallic Compounds*; Prentice-Hall: Englewood Cliffs: NJ, 1954. (b) Coates, G. E.; Green, M. L. H.; Wade, K. *Organometallic Compounds*; Methuen: London, 1967; Vol. 1, Chapter 2. (c) Bodewitz, H. W. H. J.; Blomberg, C.; Bickelhaupt, F. *Tetrahedron Lett.* **1972**, 281–284. (d) Walborsky, H. M.; Aronoff, M. S. *J. Organomet. Chem.* **1973**, *51*, 31–53. (e) Bodewitz, H. W. H. J.; Blomberg, C.; Bickelhaupt, F. *Tetrahedron* **1973**, *29*, 719–726. (f) Bodewitz, H. W. H. J.; Blomberg, C.; Bickelhaupt, F. *Tetrahedron* **1975**, *31*, 1053–1063. (g) Bodewitz, H. W. H. J.; Blomberg, C.; Bickelhaupt, F. *Tetrahedron Lett.* **1975**, 2003–2006. (h) Vogler, E. A.; Stein, R. L.; Hayes, J. M. *J. Am. Chem. Soc.* **1978**, *100*, 3163–3166. (i) Rogers, H. R.; Deutch, J.; Whitesides, G. M. *Ibid.* **1980**, *102*, 226–231. (j) Rogers, H. R.; Rogers, R. J.; Mitchell, H. L.; Whitesides, G. M. *Ibid.* **1980**, *102*, 231–238. (k) Barber, J. J.; Whitesides, G. M. *Ibid.* **1980**, *102*, 239–243. (l) Okuhara, K. *Ibid.* **1980**, *102*, 244–252. (m) Lawrence, L. M.; Whitesides, G. M. *Ibid.* **1980**, *102*, 2493–2494. (n) Hill, C. L.; Vander Sande, J. B.; Whitesides, G. M. *J. Org. Chem.* **1980**, *45*, 1020–1028.

states in which there is significant free radical character on carbon.

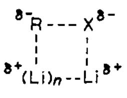
Scheme III is a kinetic adaption of the lithiation mechanism by Walborsky et al.[3h,l] and involves rate-controlling electron transfer from surface lithium to the halide to give the lithio organohalide radical anion ($R^\bullet{-}X^-,Li^+$). Homolysis of the radical anion then occurs with rapid loss of lithium halide to yield the corresponding free radical ($R^\bullet$), which is then converted by lithium to the lithio product (RLi). Homolytic decomposition of radical anions, though not rate limiting, explains the stereochemistry of reactions of chiral 1-halo-2,2-diphenylcyclopropane with lithium.[3h,l] The immediate concerns therefore are the structural and kinetic factors that control formation of halide radical anions such as $R^\bullet{-}X^-,Li^+$.

Short-lived alkyl halide radical anions $(R^\bullet{-}X)^-$ have been presumed to be intermediates from electron capture by halides in pulse radiolysis, polarography, and reactions with various metals.[18,19] As has been found for varied halides in the present reduction systems, the reactivities of alkyl halides with lithium are iodides > bromides > chlorides. The bond strengths of iodides are less than bromides than chlorides[20] and thus a reactivity order of the type RI > RBr > RCl is interpretable for Scheme III on the basis of rate-controlling electron transfer from lithium into the lowest unoccupied σ orbital in the carbon–halogen bond of the halide.[21] The fact that a chloride, allyl chloride, undergoes lithiation faster than a 1-bromoalkane can therefore be rationalized on the basis that an allylic C–Cl bond is weaker than a primary C–Br bond in conjunction with the precepts of Scheme III.[20a]

Other observations are consistent with Scheme III. A major effect is that, in the transition state leading to a radical anion, the lithium surface becomes relatively positive. An electron-donor solvent such as tetrahydrofuran or ethyl ether is thus expected to expedite reaction of a halide with the metal. Further, production of halide radical anions ($R^\bullet{-}X^-$) should be facilitated by electron-withdrawing substituents and retarded by electron-donating substituents (R) on the C–X moieties. The observed lithiation reactivities of halides, primary > secondary > tertiary and short-chain primary > long-chain primary > neopentyl, can thus be attributed (in part) to the abilities of various substituents (R) to release electrons into carbon–halogen bonds. Of additional importance is that lithiation occurs slower with bulky than with less hindered halides.[3g] The ease of formation of a halide radical anion can therefore also be related (in part) to the steric effects of substituents in reaction (including solvation) of its precursor halide chemically adsorbed on an active lithium site on a lithium surface.[3g] It is noted that electronic and steric factors which retard formation of halide radical anions can accelerate subsequent homolytic decompositions.

Although Scheme III explains lithiation of sp^3-carbon halides, there is yet no direct proof for rate-determining generation of halide radical anions and their homolysis to discrete radicals ($R^\bullet$) as discussed. Further, if the mechanism is totally or partly operational in the present systems, the two-electron-transfer path might not be followed by sp, sp^2, highly strained, or angularly distorted carbon halides.

Scheme IV, an alternate to Scheme III, involves rapid reaction of a halide with lithium by electron transfer to yield the lithio organohalide radical anion ($R^\bullet{-}X^-,Li^+$).[21] Subsequent rate-determining attack on the radical anion by surface lithium (a further extension of the mechanism proposal by Walborsky et al.)[3h,l,n] then takes place with loss of lithium halide to give the lithio organic product. As a major further consideration is that with only slight modification in the timing, the mechanism processes encompass rate-determining unsymmetrical concerted four-center transition states for lithiation with anionic character as in **6**. The mechanistic sequence of Scheme IV or/and transition state **6** explains all of the kinetic structural and solvent effects as does Scheme III and bypasses decomposition of halo radical anions to organic radicals ($R^\bullet$). To be emphasized is that Scheme IV or transition state **6** is not excluded by lithiation of chiral 1-halo-2,2-diphenylcyclopropanes with retention and *some* racemization.[3h,l,n] Homolytic front-side displacement of lithium halide from adsorbed chiral lithio 1-halo-2,2-diphenylcyclopropane radical anions by surface lithium will give 1-lithio-2,2-diphenylcyclopropane with retention as does the intimate collapse process of Scheme I. Reaction of lithium with lithio 1-halo-2,2-diphenylcyclopropane radical anions on a lithium surface as in Scheme III, however, may be sterically resisted by the phenyl groups. Adsorbed optically active lithio 1-halo-2,2-diphenylcyclopropane radical anions can then give racemic 1-lithio-2,2-diphenylcyclopropane by (1) collapse with loss of lithium halide to give the achiral 2,2-diphenyl-1-cyclopropyl radical in solution, which then tumbles and reacts with surface lithium, and/or (2) back-side surface displacement of lithium halide by lithium ion or lithium reagents in solution. As a further, possibly highly significant, point the four-centered processes as in **6** will be retarded with sterically hindered halides such as 1-halo-2,2-diphenylcyclopropanes and thus allow intervention of the above racemization processes.



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Halo olefin bonds, though considerably stronger than their haloalkyl counterparts, lithiate readily. Thus the reactivities and the mechanisms of lithiation of sp^2-carbon halides are intriguing subjects. In dilute solution in **1** the relative reactivities (r_1/r_2) of lithium with the halide pairs 2-bromobutane/2-bromopropene, bromobenzene/*p*-chlorotoluene, and bromobenzene/*m*-chlorotoluene are ~1.00 (Table II). Under such conditions the reactions of sp^2- as well as the sp^3-carbon halides are apparently mass transport controlled.[11] Also, at halide concentrations leading to structural control of kinetic behavior, lithium reacts only slightly more rapidly with 1-bromobutane than 2-bromopropene ($r_1/r_2 = 1.21$ at $C_T = 3.36$ M; Table II). Further, the relative reactivities of lithium with 2-bromopropene are greater than with 1-chlorobutane ($r_1/r_2 = 1.37–2.20$ at $C_T = 2.02–5.06$ M; Table II). Questions then arise as to the origins of the impressive reactivities of lithium with sp^2-carbon–halogen bonds.

As for sp^3-carbon halides, a single electron has been calculated to be transferred electrochemically into the lowest unoccupied σ rather than the corresponding π level of a C–X bond of a conjugated halide.[21] On the basis that lithiation on a metal surface occurs from a direction perpendicular to the C–X bond, steric resistance to attack on a conjugated halide or its lithium alkenyl halide radical anion (**7**) will be small. The mechanistic sequences of Schemes III and IV, including **6**, can then be extended to avoid generation of alkenyl radicals ($R_2C{=}\dot{C}{-}R$) in any rate-determining step in lithiation of a conjugated halide. Thus the kinetic, steric, and stereochemical requirements in the conversions of alkenyl halides to their corresponding alkenyllithium reagents can be met by formation or lithiation (**8**) of **7** or by reactions involving concerted unsymmetrical four-center processes (**9**). Lithiation of an alkenyl halide as in **8** or **9** also offers the possibility for advantageous electron transfer into the π system of its carbon–carbon double bond.

The results and the mechanisms of the reactions of alkali metals with benzyl chloride and with aryl halides (sp^2-carbon halides) remain for discussion. As expected, (1) in dilute solution the reactivities of individual alkali metals, (Li, Na, and K) with benzyl chloride are greater than with bromobenzene and chlorobenzene,

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[18] See ref 8a and: (a) Fleischmann, M.; Pletcher, D.; Vance, C. S. *J. Electroanal. Chem.* **1971**, *29*, 325–330. (b) Thomas, J. K. *J. Phys. Chem.* **1967**, *71*, 1919–1925. (c) Anbar, M.; Hart, E. J. *Ibid.* **1967**, *71*, 3700–3702. (d) Mishra, S. P.; Symons, M. C. R. *J. Chem. Soc., Perkin Trans. 2* **1973**, 391–394. (e) Infelta, P. P.; Schufer, R. H. *J. Phys. Chem.* **1972**, *76*, 987–999.

[19] Peters et al. (Peters, D. G.; Clark, J. A.; Mubarak, M. S.; Vieira, K. L.; Anderson, M. R. *J. Electroanal. Chem. Interfacial Chem.* **1986**, *198*, 107–124), however, find no evidence to support the existence of radical anions in the electrochemical reduction of alkyl halides.

[20] (a) Benson, S. W. *J. Chem. Educ.* **1965**, *42*, 502–518. (b) Kerr, J. A.; Trotman-Dickenson, A. F. In *Handbook of Chemistry and Physics*, 61st ed.; Weast, R. C., Ed.; CRC Press: Boca Raton, FL, 1980; pp F233–F234.

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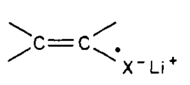

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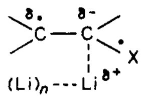

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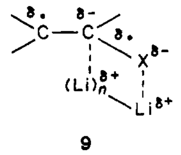

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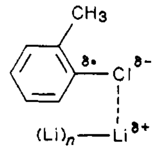
(2) upon increasing their concentrations, benzyl chloride becomes considerably more reactive than bromobenzene than chlorobenzene, and (3) potassium is less selective than sodium than lithium. The reactivity of benzyl chloride is expected from the weakness of its C–Cl bond,[20a] the mechanism profiles of Scheme III or the various versions of Scheme IV, and the electron-donor capacities of the alkali metal surfaces[9] while under intense attack.

The kinetic behavior of bromobenzene, chlorobenzene, and *o*-, *m*-, and *p*-chlorotoluene with alkali metals in **1** is summarized in Tables II and V. Although these aromatic derivatives have strong sp^2-carbon–halogen bonds, their reactivities in dilute solution with alkali metals are structure independent and presumably mass transport controlled. As the concentrations are increased in the kinetic experiments, bromobenzene becomes the most reactive of the above aryl halides. Whatever are the mechanistic details (Schemes III and IV) in chemical control of metalation of these aromatic halides, the reactivity of bromobenzene can be rationalized on the basis of the electron-withdrawing effect of the phenyl group and on the ease of electron transfer into the lowest unoccupied σ orbital and the strength of its C–Br bond.

An intriguing series of results is that, at higher concentrations, the rates of metalations of *o*-, *m*-, and *p*-chlorotoluenes (Tables II and V) in **1** are similar. More precisely, *p*-chlorotoluene is slightly more reactive than *o*-chlorotoluene and *m*-chlorotoluene is somewhat less reactive than *o*-chlorotoluene. For metalation of *o*-chlorotoluene to occur only slightly slower than for *p*-chlorotoluene implies that steric resistance plays a minor role in the reactions of the ortho-substituted halide on the metal surfaces. That steric interference in metalation of *o*-chlorotoluene is small is consistent with the postulate that the rate-controlling reactant-like transition state is formed by a low activation energy process[15] in which electron transfer occurs upon adsorption and reaction as in **10** or **11**.



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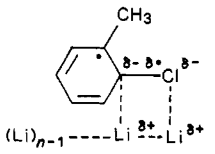

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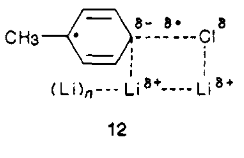
The final feature of this series of experiments is that lithiation of *m*-chlorotoluene is slower (25–28%) than for its ortho and para isomers. On the basis of electron release of a methyl group into the lowest unoccupied σ orbitals of the C–Cl bonds of the chlorotoluenes as in Scheme III, it would be predicted that *o*- and *p*-chlorotoluenes would lithiate more slowly than their metal isomer. For the opposite to be the case, the question is raised that lithiation occurs by the four-center unsymmetrical mechanistic version of Scheme IV or as in transition state **6**. Thus rate-limiting formation of transition states **11** and **12** for lithiation of *o*- and *p*-chlorotoluenes will be accelerated electronically by their methyl groups as compared to **13** as derived from *m*-chlorotoluene. Further study of substituent effects on the reactivities of halobenzenes with metals, the kinetics of reaction of various halides with metals, and the mechanistic details of the reactions of alkali metals with alcohols is in progress.



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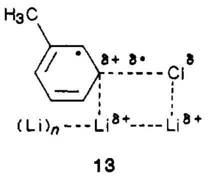

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Experimental Section

Organic Halides. The organic halides (reagent grade) were passed through an alumina column and used directly if they gave a single peak on a gas chromatogram (a 5-ft column of 30% silicone). If more than one peak was observed in the gas chromatogram, the halide was washed several times with cold concentrated sulfuric acid, water, sodium bicarbonate solution, and then water until neutral to Alkacid test ribbon. (The aromatic halides were not washed with sulfuric acid; possible sulfonation was thus avoided.) The halides were dried through anhydrous magnesium sulfate and fractionated through a helix-packed column (1 ft). The final product gave only one peak on a gas chromatogram.

Solvents. Ethyl Cellosolve (comm) was fractionated through a helix-packed column (1 ft), stored under nitrogen, and used in less than 24 h. Tetrahydrofuran and ethyl ether were each distilled (under nitrogen) over lithium aluminum hydride and used immediately. Benzene, cyclohexane, and cyclohexene (all reagent grade) were distilled (under nitrogen) over lithium hydride through a Newman condenser and stored under nitrogen until used. *p*-Dioxane was passed through an alumina column (2 × 14 in.), fractionated in a helix-packed column (1 ft), and stored under nitrogen.

General Procedure. The equipment and the stepwise sequence for effecting reactions of halides with lithium, sodium, or potassium were identical in all experiments and are described as follows.

The reaction equipment consisted of a round-bottom (100–250 mL), ground-glass-jointed (24/40), one-neck flask with a tubular opening (8-mm diameter) located just below the neck of the flask. The tubule was fitted with a serum cap through which a hypodermic needle connected to a nitrogen inlet was inserted. Purified nitrogen was passed via the hypodermic needle through the reactor flask. The halide and the solvent(s) were then added by syringe through the serum cap. The 24/40 joint was fitted with a Claisen adapter, which allowed attachment of a motor stirrer and stirrer shaft by way of a Tru-Bore bearing. The Claisen adapter was also attached to a reflux condenser, fitted with a drying tube connected to a 1-L water displacement tower containing a tubular (8 mm) sampling port.

The procedure for a typical competitive metalation experiment is as follows. The reactor vessel was immersed in an ice bath and continuously flushed with nitrogen. A halide pair (5–10 mL) of predetermined (by weight) mole ratio was added by syringe. Freshly distilled solvent or solvents were also added by syringe. The stirring motor was started and the system allowed to equilibrate under nitrogen in the ice bath. The flow of nitrogen was then discontinued and the water displacement tower connected. Freshly cut alkali metal (0.1–0.3 mmol) was then quickly added to the reactor through the tubule. The experiment was terminated when all the metal had been consumed (15–45 min for Li).

Since the relative rate constants are independent of time in these systems, reaction times were not rigorously recorded. The reactions in ethyl Cellosolve (**1**) or **1** containing tetrahydrofuran or *p*-dioxane were complete in ≈15 min at high halide concentrations ($C_T \sim 5.0$ M). The reaction times increased (30–45 min) when **1** was diluted with cyclohexane, cyclohexene, or benzene. In all of the systems reported the alkali metal remains shiny throughout the entire reaction.

There is no appreciable induction period in the aforementioned systems in that bubbling of gaseous hydrocarbons into the displacement tower commences within 1–2 min after addition of a metal. No bubbles were formed during reactions in which the hydrocarbons produced are nongaseous (benzene and toluene).

Handling of Metals. The alkali metals used were regular grade lithium from Lithium Corp. of America and reagent grade sodium and potassium from J. T. Baker Co. In each case the metal was cut into a rectangular parallelepiped piece (~1 × 0.6 × 0.1 cm) with all faces freshly cut. The metal was added to a preweighed vial (2 mL) flushed with nitrogen, weighed, and then transferred directly to the reaction mixture. The amount of metal added was enough to react with only 5–10 mol % of the halides present in a reaction mixture.[22]

Analyses of Products. After completion of the reactions of a halide pair and an alkali metal, the mixture was refluxed and the volatile material was collected in the displacement tower (200–250 mL was displaced at this point). The reaction system was then flushed slowly (10–15 min) with nitrogen by way of a hypodermic needle inserted through the serum cap and below the surface of the refluxing liquid. When the gas volume in the displacement tower was ~1000 mL, the flow of nitrogen was

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(22) The facts that nitrogen reacts with lithium and that sodium and potassium react with oxygen became concerns during this work. As it turns out, the present systems are apparently self-cleansing, and the initial condition of the surface or the presence of oxygen has no measurable effect on the results. Lithium coated with nitride or sodium and potassium coated with oxides give the same results as the procedure presently described and used throughout this work.

discontinued. After the reaction mixture had been purged with nitrogen, samples taken directly over the refluxing reaction mixture showed no detectable hydrocarbons in the reactor vessel. The gaseous hydrocarbon mixture was sampled and then analyzed by gas chromatography on Dowtherm (15%). The analytical method was found to be accurate upon preparing various mixtures of isobutane and *cis*-2-butene of known compositions in cyclohexane/ethyl Cellosolve/1-chlorobutane (1–5 M), separating the gaseous components by the method described above and gas chromatography.

The products from reactions of aryl halide pairs with alkali metals in **1** are benzene and toluene (nongaseous). The hydrocarbons were concentrated in the halide reaction mixture upon removing the ethyl Cellosolve by extraction with saturated sodium chloride solutions (3–5 extractions). The hydrocarbon ratios were determined by gas chromatography (20% SF-96 column) of the hydrocarbon/halide mixtures.

The nongaseous coupling products (Table I) were analyzed by a procedure identical with that used for benzene and toluene.

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Reactions of *N*-(Arylsulfonoxy)-*N*-benzylmethylamines with Sodium Methoxide in Methanol[1]

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Abstract: Reactions of *N*-(arylsulfonoxy)-*N*-benzylmethylamines (**1**) with MeONa in MeOH have been investigated. The eliminations are quantitative, producing only benzylidenemethylamines. The reactions are first order in substrate and first order in base, and an E2 mechanism is evident. Comparison of the rate data reveals that the arenesulfonates are better leaving groups than halogens in imine-forming eliminations. For eliminations from **1**, $\rho = 1.14$, $k_H/k_D = 3.6$, $\rho_{lg} = 1.30$, and $\beta_{lg} = -0.43$ have been calculated. Differences in the transition-state characters for imine-forming eliminations from $ArCH_2N(Cl)CH_3$ and **1** have been interpreted in terms of changes in the leaving-group propensities and base strength.

Extensive studies of olefin-forming β-elimination reactions have led to a qualitative understanding of the relationship between the structure of E2 transition state and changes in reactants and reaction conditions.[2,3a] Results from these studies indicate that the transition state shifts from central E2 to the E1cb borderline in the More O'Ferrall–Jencks reaction coordinate diagram as the β-phenyl substituent is made more electron withdrawing or the leaving group becomes poorer. In contrast, relatively little is known about factors that influence the shift of a central E2 transition state to the E1 borderline.

Earlier we reported that the reactions of *N*-chloro-*N*-benzylmethylamines with MeONa–MeOH proceeded via an E2-central transition state with appreciable C_β–H and N_α–Cl bond cleavage and significant double-bond character.[4-8] On the other hand, for reactions of *N*-(arylsulfonoxy)-*N*-benzylamines with benzylamine in MeOH, Hoffman and co-workers[9-12] found an E1-like transition state with extensive N_α–OSO_2Ar rupture, but limited C_β–H bond scission and carbon–nitrogen double-bond development. It appears that the transition-state differences for these two reactions may provide information regarding a shift in transition-state character from central E2 to the E1 borderline.

Table I. Rate Constants for Eliminations from $ArCH_2N(OSO_2Ar)CH_3$ Promoted by MeONa–MeOH[a]

entry	compd[b]	temp, °C	k_2,[c] $M^{-1}\,s^{-1}$
1	**1a**	25.0	1.98
2	**1a**	35.0	3.98
3	**1a**	45.0	7.68
4	**1b**	25.0	0.544
5	**1c**	25.0	1.49
6	**1d**	25.0	4.48
7	**1e**	25.0	5.44
8	**1f**	25.0	14.7
9	**1g**	25.0	0.439
10	**1h**	25.0	1.02
11	**1i**	25.0	3.64

[a] [MeONa] = 1.92×10^{-3} to 7.68×10^{-2} M. [b] [Substrate] = $(1.0\text{–}2.0) \times 10^{-4}$ M. [c] Estimated uncertainty, ±3%.

Hoffman suggested that the transition-state differences for these two reactions may be attributed to the large difference in the energies of N–Cl and N–OSO_2Ar bonds.[10] However, the validity of this interpretation is questionable because these two systems differ not only in the nature of the leaving group but also in the base strength and charge type. Recently, we reported that the

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(1) This research was supported by grants from the Korea Science and Engineering Foundation and the Korea Research Foundation.

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(12) Hoffman, R. V.; Shankweiler, J. M. *J. Am. Chem. Soc.* **1986**, *108*, 5536–5539. We thank Professor R. V. Hoffman for a preprint of this article.

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