Internal Pressure's Role in the Selectivity of the Bromine Radical

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Abstract:

Internal Pressure's Role in the Selectivity of the Bromine Radical : an Interesting Entropic Effect

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The results of our studies demonstrate a link between selectivity in free radical brominations and the internal pressure of solvent. The study encompassed 18 solvents and/or mixtures over which the rate constant ratio for S_H2 ring opening of cyclopropylbenzene (k_C) vs hydrogen abstraction from toluene (k_H) by bromine atom was found to vary by nearly a factor of twenty as a function of solvent. The observed rate constant ratios correlate with the cohesive energy density of the solvent, which is approximately equal to the internal pressure (k_C/k_H is found to increase with increasing solvent pressure) for the solvents used in this study. A similar competition pitting cyclopropylbenzene against *p*-chlorotoluene shows the same solvent effect. It is suggested that the observed variation in selectivity with solvent pressure ocurrs because the volume of activation with k_C is more negative than for k_H by about 20 cm^3/mol.

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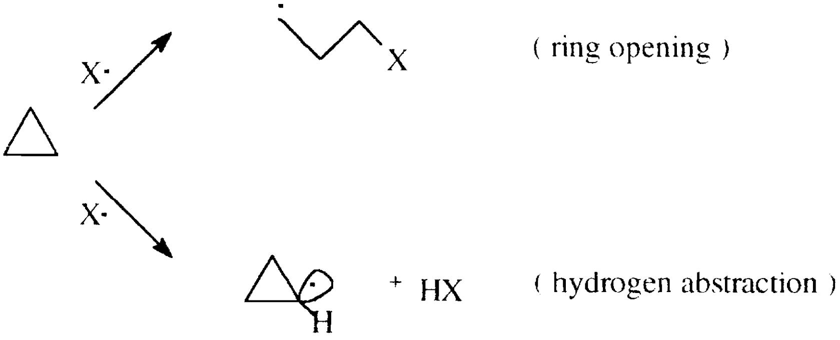
Background Part 1:

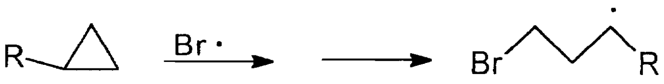
Free Radical Bromination Of Methyl and Cyclopropylarenes

The reactions of cyclopropane with free radicals have been well studied . The products of the reaction result from either hydrogen atom abstraction, ring opening or both . Which of these processes occurs depends on the nature of the attacking radical.[1] Scheme I illustrates the two possible reaction pathways.



Scheme I

In the case of the bromine atom, ring-opening to form the 3-bromopropyl radical is the predominant process. For substituted cyclopropanes, the reaction involves attack of the bromine atom at the least-hindered carbon of the cyclopropyl ring and yields the most stable radical (with inversion of configuration).[1]



Scheme II

Similar reaction of bromine atom with an arylcyclopropane involves several additional considerations. First the conformation of the cyclopropane relative to the aromatic system must be considered. The cyclopropane ring exists predominantly in one of two conformations, bisected and perpendicular (Figure 1).

Figure 1[2]

The two predominant conformations of cyclopropylarenes.

Of the two, the bisected conformation is usually the most stable due to the favorable alignment of orbitals which allows for interaction between the cyclopropyl HOMO and the aromatic LUMO (Figure 2). At the same time this conformation results in the C-H bonds of the cyclopropyl group being orthogonal to the aromatic pi-system. This greatly reduces the likelihood of hydrogen abstraction by the bromine radical because the transition state can not realize benzylic stabilization. On the other hand, in the perpendicular conformation, hydrogen atom abstraction is preferred due to the favorable alignment of the C-H bond with the aromatic pi-system which allows for stabilization of the transition state.[2]

For arylcyclopropanes, the ring opening process usually predominates. Hydrogen abstraction has only been observed when, because of steric constraints, the molecule is effectively locked into a perpendicular conformation.[2]

Figure 2

Orbital alignments for perpendicular and bisected cyclopropylarenes.

It is also unlikely that this predominant ring opening process is reversible. Formal carbon abstraction resulting in the opening of the cyclopropyl ring (an S_H2 process) is highly exothermic (-21kcal/mol).[2] This is due to the relief of cyclopropane's tremendous ring strain.

HBr Reversal

The possibility of reversal does however exist in the case of hydrogen abstraction fom toluene.[3] In this case the HBr formed by the hydrogen abstraction back reacts to regenerate the starting material (HBr reversal).

For the reaction of toluene with bromine the established mechanism is[3]

$$Br_2 \longrightarrow 2Br\cdot \qquad (1)$$

$$Br\cdot + C_6H_5CH_3 \rightleftharpoons HBr + C_6H_5CH_2\cdot \qquad (2)$$

$$C_6H_5CH_2\cdot + Br_2 \longrightarrow C_6H_5CH_2Br + Br\cdot \qquad (3)$$

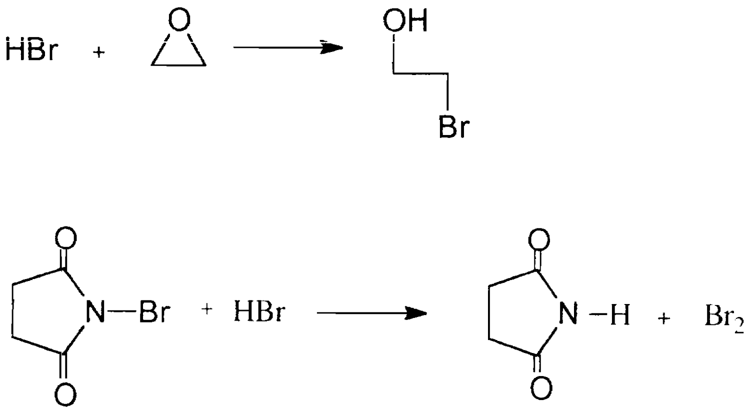
Scheme III

with the reverse process in step 2 representing the undesired regeneration of the toluene starting material.

In 1968 Timmons and Guzman, while studying kinetic isotope effects, found that the HBr reversal could be minimized by limiting the percent of conversion of their reactions to around 2%. They also discovered that by using a high ratio of toluene to bromine, any additional formation of HBr from further reaction of the benzyl bromide product could be avoided. In varying the toluene to bromine ratio from 5:1 to 20:1 they found that the steps taken to minimize HBr interference had little effect on their results and thus concluded that any complications from HBr reversal were eliminated.[4]

In 1975 Tanner suggested that diffusive HBr reversal (regeneration of the toluene by reaction of the benzyl radical with HBr) could be thoroughly eliminated by carrying out reactions in the presence of an HBr scavenger. Among the scavengers used were powdered potassium carbonate, N-bromosuccinimide and the epoxide ethylene oxide. The scavenging processes for three of these are shown in Scheme IV.

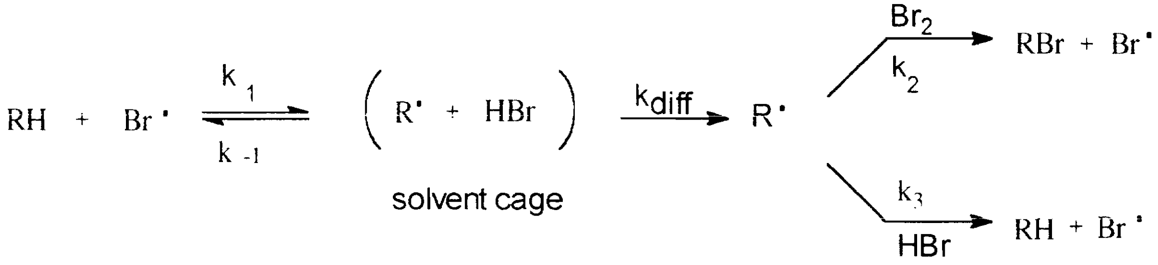
$$K_2CO_3 + HBr \longrightarrow KHCO_3 + KBr$$



Scheme IV[5]

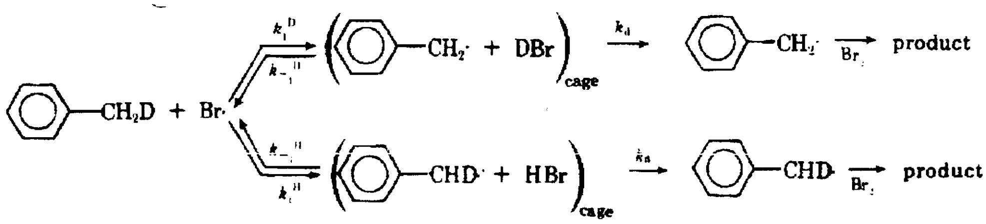
Among the findings by Tanner was that qualitatively the reactions run in the presence of potassium carbonate or ethylene oxide were five to ten times slower than identical reactions run without a scavenger. From these changes in rate, Tanner concluded that different mechanisms were governing the reactions occurring with and without HBr present. He further concluded that while potassium carbonate greatly reduced the presence of HBr, the use of ethylene oxide completely eliminated it, thus making the epoxide most efficient of the scavengers.[3]

In addition to diffusive HBr reversal, in 1975 Tanner proposed that HBr reversal could also occur internally within the solvent cage. According to Tanner the cage return to starting materials would likely occur at a rate competitive to diffusion from the solvent cage. In Scheme V the competition for the caged species between reversal (k_{-1}) and diffusion (k_{diff}) can be seen.



Scheme V[5]

Through isotopic studies Tanner proposed that the contribution of the cage effects could explain the difference in values for deuterium isotope effects observed for toluene in the liquid and gas phases. Scheme VI shows how the cage effect might alter solution phase results through in-cage reversal.



Scheme VI[6]

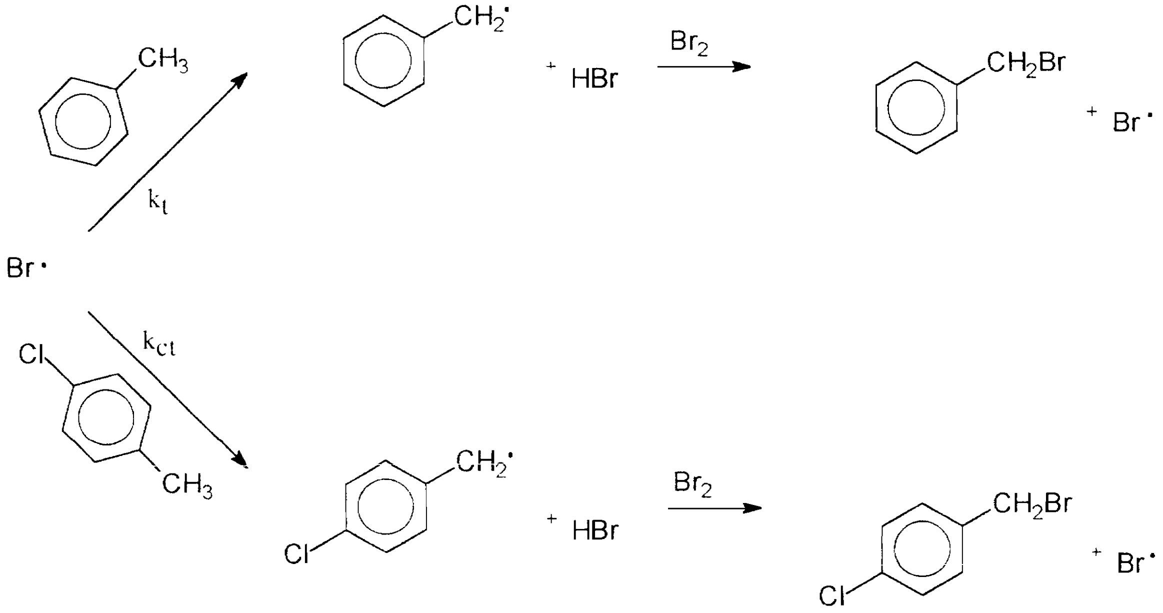
Later, through a series of experiments involving the bromination of cyclohexane the dependence of the intermolecular isotope effect on the concentration of bromine was shown with the relative rate constant k_{RH}/k_{RD}, calculated from the equation $k_{RH}/k_{RD} = \ln([C_6H_{12}]^o/[C_6H_{12}]^f)/\ln([C_6D_{12}]^o/[C_6D_{12}]^f)$ and used to measure the amount of intermolecular deuterium isotope effect occuring. (Table 1)

Table 1[6] The variation of k_{RH}/k_{RD} with Bromine concentration.

	Concentration, M x 10^3					
Conditions (temp, ^oC)	$[C_6H_{12}]^o$	$[C_6D_{12}]^o$	$[Br_2]^o$	$[C_6H_{12}]^f$	$[C_6D_{12}]^f$	k_{RH}/k_{RD}
Freon 113 (21)	23.2	22.8	10200	2.44	14.6	5.05
Freon 113 (21)	23.2	22.8	7710	4.03	15.9	4.86
Freon 113 (21)	23.2	22.8	3630	7.58	17.7	4.42
Freon 113 (21)	4.79	4.87	1030	2.24	4.05	4.12
Freon 113 (21)	4.91	4.99	149	2.85	4.42	4.48
Freon 113 (21)	4.75	4.83	68.6	2.52	4.17	4.31

Table 1 shows that at higher concentrations of molecular bromine the value of k_{RH}/k_{RD} increased due to increased cage scavenging of the radicals by the molecular bromine.Tanner also found that in running a reaction in 18 M bromine that any cage recombination had been completely eliminated (k_{RH}/k_{RD}= 5.38).[6] Unfortunately, in cases such as toluene, internal return can not be eliminated by scavenging due to the possible formation of a complex radical (Br_3 •) which is more selective than a bromine atom.[7]

Tanner discovered in 1987 that the internal cage return he had observed was dependent upon the viscosity of the solvent. In studying the relative rates (k_t/k_{ct}) of H-abstraction from toluene vs a similar abstraction from *p*-chlorotoluene (Scheme VII), Tanner noted that the relative rate k_{ct}/ k_t decreased with increasing viscosity.



Scheme VII

Similarly he found that the reaction rate increased with decreasing viscosity.[8] This result would seem to be consistent with the expected contributions from cage return since at higher viscosities the cage lifetime should increase, subsequently increasing cage return and slowing diffusion from the solvent cage. Figure 3 shows that at higher temperatures, the dependence of the relative rate k_{ct}/k_t on viscosity diminishes (with η_i representing the viscosity of the Freon solvent relative to Freon 11 at 23^oC.)

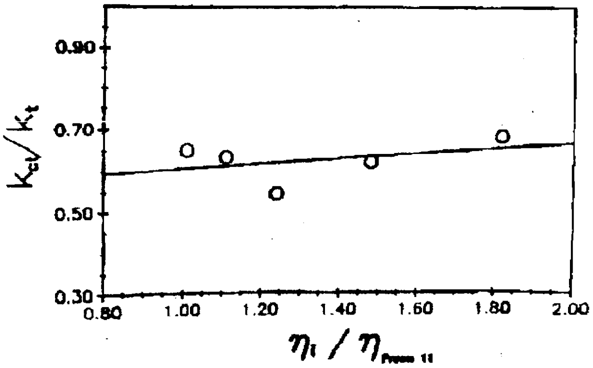


Figure 3[7]The relative rates of bromination 4-chlorotoluene/ toluene (k_{ct}/k_t) vs the relative viscosity of the Freon solvent ($\eta_i/\eta_{FREON\ 11}$).

Since viscosity is temperature dependent, reaction temperature must also be taken into account when studying the effect of internal return. The inverse relationship between viscosity and temperature suggests that temperature and hence viscosity would have its greatest effect on cage return at low temperatures (high viscosity). Figure 4 shows a plot of the relative rate k_{ct}/k_t vs temperature over a range of 90^o C at a constant viscosity with the open triangles representing values obtained at a constant viscosity of 1.253 relative to Freon 11 and the filled points representing extrapolated values.

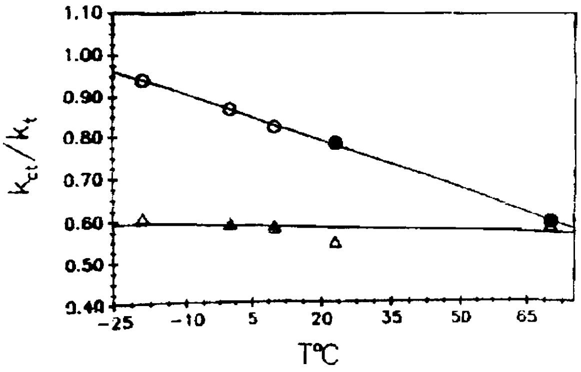


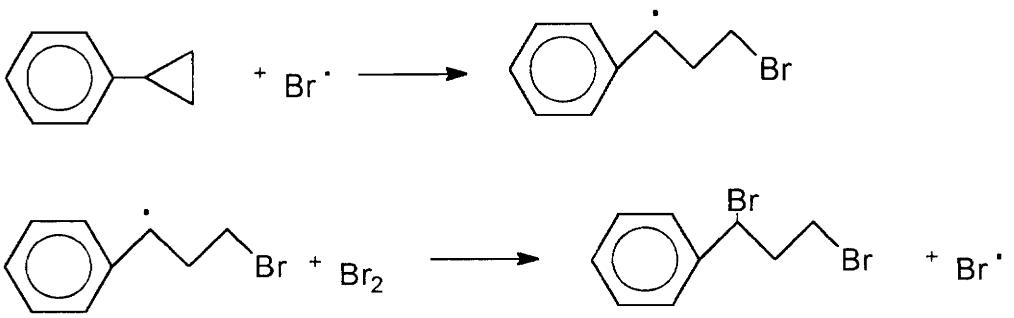
Figure 4[7]

Relative rates of bromination 4-chlorotoluene/ toluene vs temperature(°C).

The above plot reemphasizes that if the relative viscosity is held constant (at 1.253) "temperature itself has little effect " on the observed rate constant ratio k_{ct}/ k_t.[7]

Thus,while HBr reversal has been a major point of investigation in the reaction of bromine and toluene, several different areas have been explored concerning the reaction of bromine with an arylcyclopropane such as phenylcyclopropane. Among these are the formation of a thermal product via an electrophilic (dark) substitution reaction and complexation of the bromine atom to certain solvent molecules.

The conventional mechanism for the free radical reaction of phenylcyclopropane with bromine is illustrated in Scheme VIII.



Scheme VIII[9]

In 1972 LaLonde and Ferrara ran a series of reactions that showed the formation of a product resulting from a dark reaction of phenylcyclopropane and bromine in carbon tetrachloride at -20°C. Among their findings was that two separate mechanisms were controlling the light and dark reactions. They reported that the photoinitiated (light) reaction gave exclusively the 1,3-addition product (shown in Scheme VIII) and was 20 times faster than the dark reaction under their reaction conditions. The dark reaction was found to give only electrophilic aromatic substitution products (*p*-bromo phenylcyclopropane).

In changing solvents to the more polar chloroform, LaLonde discovered that the 1,3-addition became competitive in the dark reactions. He also noticed increased 1,3-addition as the temperature was increased to 25°C. From this he concluded that two mechanisms were responsible for the 1,3-addition product. Under dark conditions , a temperature- enhanced slow electrophilic addition was occurring, while under light-induced conditions a faster free radical process was taking place.[9]

In 1958 Russell described the complexation of the chlorine atom to certain solvents. Due to this complexation he found that the selectivity of chlorine for primary , secondary and tertiary hydrogens was altered.[10] While this selectivity effect has not been found for the bromine atom, LaLonde has found some evidence of molecular bromine-phenylcyclopropane complexes. He found in dark reactions that new UV bands appeared at 296 nm in carbon tetrachloride and at 299 nm in n-hexane. LaLonde proposed that these new bands could be the result of charge-transfer complexes similar to those previously observed with other halogens in simple aromatic hydrocarbons. An additional observation of LaLonde was that upon addition of methanol to the solutions, the new bands disappeared and were replaced by stronger bands from a methanol-bromine complex. He supported this finding by noting the greater stability and lower lambda max of the methanol-bromine complex thus making it energetically more favorable than the phenylcyclopropane-bromine complex. Subsequently, he found that the methanol addition retarded the rate of bromine addition as the bromine-phenylcyclopropane complex disappeared. This led him to suspect that the bromine-phenylcyclopropane complex may play some role in the rate of electrophilic addition under dark conditions.[9]

Background Part 2:

THE EFFECT OF PRESSURE ON REACTION RATES

The effect of pressure on rate constants can be seen through the thermodynamic expression

$$(\partial G/\partial P)_T = V$$

with V representing the molar volume. From this the free energy of activation can be defined as $\Delta G^{\neq} = G^{\neq} - \Sigma G_R$ with ΣG_R as the sum of the molar free energies of the reactants and $G^{\neq}$ as the free energy of the transition state. This leads to the expression

$$(\partial \Delta G^{\neq}/\partial P)_T = V^{\neq} - \Sigma V_R = \Delta V^{\neq}$$

with $\Delta V^{\neq}$ representing the activation volume and $V^{\neq}$ representing the volume of the transition state. Combining this with an expression from transition state theory ($\Delta G^{\neq} = -RT\ln k + \text{const}$), leads to

$$(\partial \ln k/\partial P) = -\Delta V^{\neq}/RT$$

in which the activation volume is itself pressure dependent. If $\Delta V^{\neq}$ is negative, an increase in pressure will increase the reaction rate.[11]

$\Delta V^{\neq}$ can be estimated by looking at the mechanism of a given process. Table 2 shows the effect of several mechanistic features on $\Delta V^{\neq}$.

Table 2[12]: The contributions of mechanistic features to $\Delta V^{\neq}$.

Mechanistic Feature	Contribution,cm^3/mol
Bond cleavage	+10
Bond formation	-10
Displacement	-5
Diffusion control	>+20
Cyclization	0
Charge dispersal	+5
Charge concentration	-5

Background Part 3:

A REVIEW OF SOLVENT PARAMETERS

One of the basic assumptions used in studying solvent parameters is that there is some correlation between cohesive energy density and solubility.Cohesive energy density (c.e.d.) is generally defined as potential energy per unit volume. This energy is derived from three fundamental sources; dispersion forces resulting from a fluctuating atomic dipole, polar interactions such as dipole-dipole (Keesom) and dipole-induced dipole (Debye) interactions and chemical interactions such as hydrogen bonding. Equation 1 shows the integral expression for cohesive energy density. The $\Delta^{g}_{l}U$ term represents the energy required to vaporize a liquid to its saturated vapor, while the $(\partial U/\partial V)_T\, dV$ term represents the energy required to isothermally expand the saturated vapor to infinite volume.

Equation 1[13]

$$-E = \Delta^{g}_{l}U + \int_{V=V_{vap}}^{V=\infty} (\partial U/\partial V)_T \, dV$$

At sub-boiling temperatures cohesive energy density may be defined as **c.e.d.= -E/V** if ideal gas behavior can be assumed.

A related solvent parameter is internal pressure (P_i) which was described by Hildebrand as "the cohesive force which is the resultant of forces of attraction and forces of repulsion between molecules of a liquid."[13] Internal pressures can be calculated by several methods. In cases in which the internal pressure is much greater than the external pressure , internal pressure may be derived from equations of state (equation 2) where $P_i = (\partial U/\partial V)_T$.

Equation 2[13]

$$(\partial U/\partial V)_T = T(\partial P/\partial T)_V - P$$

Alternatively internal pressure can be calculated from coefficients of thermal expansion (α) and compressibility (β). (equation 3)

Equation 3[13]

$$P_i = T(\alpha/\beta)$$

A third method for determining internal pressure is to use the assumption that the internal molar latent heat of vaporization represents the work required against internal pressure to vaporize one mole of liquid occupying a volume V. (equations 4 & 5)[14]

Equations 4 & 5[14]

$$L_i = (\partial U/\partial V)_T V$$

$$P_i = (\partial U/\partial V)_T = L_i/V$$

"For common organic solvents the internal pressures range from 1800 to 5000 atm at 25°C."[14]

While cohesive energy density (c.e.d.) and internal pressure (P_i) are often mistakenly used interchangeably, the two parameters are different. Cohesive energy density represents an integral quantity while internal pressure provides an instantaneous isothermal volume derivative of internal energy.[13] Figure 5 shows internal pressure and cohesive energy density for a series of n-alkanes plotted as a function of (1/N) with N representing the number of carbon atoms. The plot clearly shows deviations as the number of carbon atoms and the molecular weight increases (hence 1/N approaches 0).

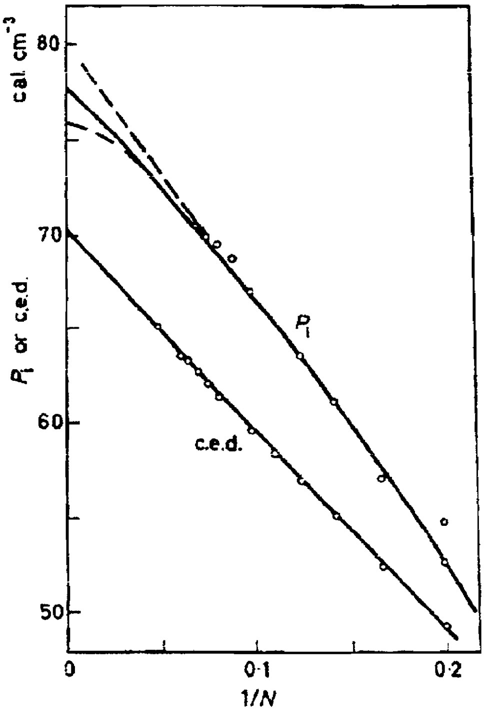


Figure 5[15]

Internal pressures and cohesive energy densities of n-alkanes plotted as a function of 1/N.

An alternative way of viewing the relationship between cohesive energy density and internal pressure can be seen in Figure 6.

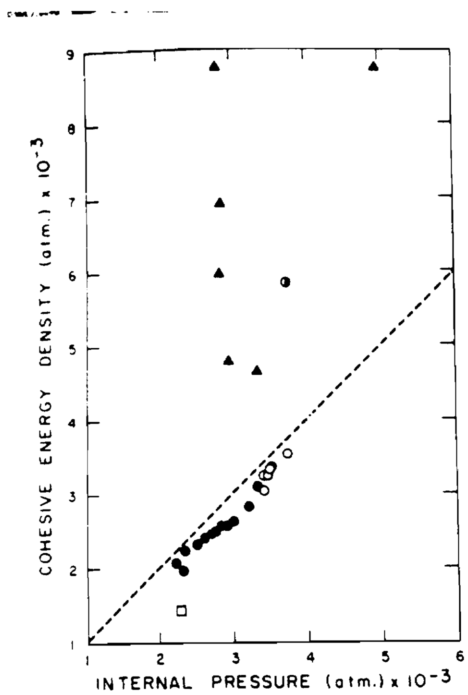


Figure 6[16]-- The relationship between cohesive energy density ($\Delta E_{vap}/V_T$) and internal pressure ($(\partial E/\partial V)_T$ for simple alkanes (•) , aromatic hydrocarbons (o) and alcohols (Δ) .

Although Figure 6 displays both cohesive energy density and internal pressure in atm. , it is more common to express internal pressure in units of pressure such as pascals while cohesive energy density is typically expressed in units such as Jcm^{-3}.[13]

In practice the parameter δ, known as the Hildebrand parameter, is more commonly used than c.e.d. or P_i . Hildebrand defined δ as $\delta = (\Delta E_{vap}/V_m)^{1/2}$ which is the square root of cohesive energy density and serves to approximate internal energy. While δ is known for a wide variety of solvents, it applies best to regular solutions (solutions which lack specific interactions and molecular polarity). For nonregular solutions the Hildebrand parameter retains some use but has generally been replaced by multicomponent parameters such as Hansen parameters.[13]

Determination of the Hildebrand parameter for many low molecular weight liquids can be accomplished through the expression

Equation 6[13]

$$\delta = \left[(\Delta H_v - RT)/ V_m \right]^{1/2}$$

with values for ΔH_v coming from heat-capacity temperature measurements.[17]

Another useful method for obtaining ΔH_v is through the use of Hildebrand's equation (equation 7) which relates the enthalpy of vaporization to the liquid's normal boiling point (T_b).

Equation 7[17]

$$\Delta H_v = T_b^2 + 23.7T_b - 2950$$

The basis of this treatment is Hildebrand's Rule which states that "the molar entropy of vaporization is the same for all regular liquids if measured at temperatures such that their vapors have equal volumes." This method provides a more simplistic calculation although corrections may be required if hydrogen bonding is present.[13]

Occasionally, provided enough structural information is available, values for ΔH_v can be approximated by summing atomic and group contributions at a specified temperature. Equation 8 shows the expression for this treatment developed by Dunkel in 1928.

Equation 8[17]

$$\delta = [\ \Sigma_i\ \Delta h_i/V - RT/V\ ]^{1/2}$$

The term Δh_i of the expression represents the individual contribution of a particular group or atom.[17] Table 3 shows some representative group and atomic contributions.

Table 3[17] Atomic and Group contributions to the Heat of Vaporization.

Atom or Group	Δh_i (in cal/mole)
CH_3	1780
$=CH_2$	1780
CH_2	990
O	1630
CHO	4700

By the same method molar volume can also be estimated from group contributions. In 1968 Rheinbeck and Lin, in their study of homologous series of n-alkanes , found that molar volumes could be approximated using the molar volume contributions of the end methyl group (ΔV_{CH3}), the z substituent (ΔV_z) and the methylene groups (Δ_{CH2}).

Equation 9[17]

$$V = \Delta V_{CH3} + \Delta V_z + (n-1)\ V_{CH2}$$

Some typical group contributions are given in Table 4.

Table 4[17]Group contributions to molar volumes.

Group	Δv_i (in cm^3/mole)
CH_3	34.0
OH	8.7
CHO	26.0
COOH	27.0
Phenyl	75.0

Once an acceptable value for δ has been obtained it can be used to predict solvent-related changes in reaction rates by using a relationship devised by Glasstone in 1941 (equation 10) where A and B represent reactants in solvent S with δ_A ,δ_B, δ_S and $\delta_{\neq}$ representing their Hildebrand values and V^A,V^B and $V^{\neq}$representing the volumes of reactants A&B and the transition state .

Equation 10[18]

$$RT\ln k = \text{const.} + V^A(\delta_A - \delta_S)^2 + V^B(\delta_B - \delta_S)^2 - V^{\neq}(\delta_{\neq} - \delta_S)^2$$

Equation 10 shows that reaction rate is dependent upon the difference in the molar volumes of the reactants and the activated complex ($\Delta V^{\neq} = V^{\neq} - (V^A + V^B)$). This quantity ($\Delta V^{\neq}$) known as the " volume of activation " is critical in relating solvent pressure to reaction kinetics and is dependent on molecular structure and the ability of the solvent to compress this volume by electrostriction.[18]

Equation 10 also shows the relationship between k and δ_S and can be rearranged as a quadratic equation in δ_S.

Equation 11

$$RT \ln k = A\delta_S^2 + B\delta_S + C$$

If $A\delta_S^2 > B\delta_S$ the expression reduces to an expression similar to that shown for external pressure ($\Delta V^{\neq} = V^{\neq} - V_A - V_B$) with $A = \Delta V^{\neq}/RT$, $B = -2(V_A\delta_A + V_B\delta_B - V^{\neq}\delta^{\neq})/RT$ and $C = (V_A\delta_A^2 + V_B\delta_B^2 - V^{\neq}\delta^{\neq 2})/RT$.

Equation 10 can be derived from the energy of mixing.

Equation 11

$$\Delta E_M = (n_1 n_2 V_1 V_2)/(n_1 V_1 + n_2 V_2)\,[(\Delta E_1/V_1)^{1/2} - (\Delta E_2/V_2)^{1/2}]^2$$

substituting in the Hildebrand term $\delta = \text{c.e.d.}^{1/2} = (\Delta E_{vap}/V)^{1/2}$ leads to the expression $\Delta E_M = (n_1 n_2 V_1 V_2)/(n_1 V_1 + n_2 V_2) * (\delta_1 - \delta_2)^2$. Differentiating this expression with respect to n_2, gives the partial molar energy of mixing 2 in 1:

Equation 12

$$(\partial \Delta E_M/\partial n_2)_{n1} = \Delta E_2 = V_2\phi_1^2(\delta_1 - \delta_2)^2$$

For an ideal solution, the entropy of mixing is $\Delta S_2 = -R\ln x_2$. Where x_2 is the mole fraction of component 2. Writing $\Delta G_2 = \Delta H_2 - T\Delta S_2$ and replacing ΔH_2 with ΔE_2 (equating Gibbs free energy and the Helmholtz free energy) leads to

Equation 13

$$\Delta G_2 = V_2\phi_1{}^2(\delta_1 - \delta_2)^2 + RT\ln x_2$$

Since $\mu_2 - \mu_2{}^0 = \Delta G_2 = RT\ln a_2$, where μ_2 is the chemical potential and a_2 is the activity, this expression becomes $\ln(a_2/x_2) = \ln\gamma_2 = [V_2\phi_1{}^2(\delta_1 - \delta_2)^2]/RT$ where γ_2 represents the activity coefficient of component 2. If the solution of component 2 is very dilute, $\phi_1 \approx 1$, then the expression becomes $RT\ln\gamma_2 = V_2(\delta_1 - \delta_2)^2$. Applying this to the kinetic expression $k = k_0(\gamma_A\gamma_B/\gamma^{\neq})$, leads to equation 11.[11]

Thermodynamically activation volumes can be derived from the equation of absolute rates (equation 14).

Equation 14[12]

$$\Delta G^{\neq} = -RT\ln k + \text{const.}$$

This leads to the equation (equation 15).

Equation 15[12]

$$(\partial \ln k/\partial P)_T = -\Delta V^{\neq}/RT$$

It can be seen that the activation volume can only be determined by measuring the effect of hydrostatic pressure on the rate constant. Equation 11 shows that internal pressure should influence reaction rates in the same direction as external pressure.[18] This implies that reactions which have a negative volume of activation will tend to have accelerated rates while reactions with a positive activation volume should show a decrease in rate.

In extending considerations on activation volume to competition reactions the term $\Delta\Delta V^{\neq}$ must be introduced. $\Delta\Delta V^{\neq}$ is taken to be the difference in activation volumes between the 1st and nth reaction (equation 16).

Equation 16[12]

$$\Delta\Delta V^{\neq} = \Delta V^{\neq}{}_{n} - \Delta V^{\neq}{}_{1}$$

The effect of pressure on inhibitting the formation of radical pairs is relatively small. When the effects of pressure on product distributions are studied, formation within the solvent cage is barely suppressed when compared to escape product yields.

In 1975 Zhulin noticed a systematic effect of internal pressure on a series of reactions in which substituted toluenes competed for the N-bromosuccinimide derived bromine radical. By using the σ constants of the substituents the Hammond postulate was used to accurately predict the direction of solvent effects.[12] Generally the direction of competition reactions of aromatic substrates for radicals can be predicted by assuming that a more crowded or highly-branched product will be more highly preferred as the internal pressure increases.[12]

Research Objectives

The primary objective of our research was to examine a "solvent effect" we noted several years ago on the relative rate constant k_C/k_H. This was approached by looking at the competition of cyclopropylbenzene and toluene for the bromine atom. The competitions were carried out over a very large range of solvents varying in polarity, viscosity and internal pressure in order to better understand the molecular basis for the observed effect.

Discussion Of Results:

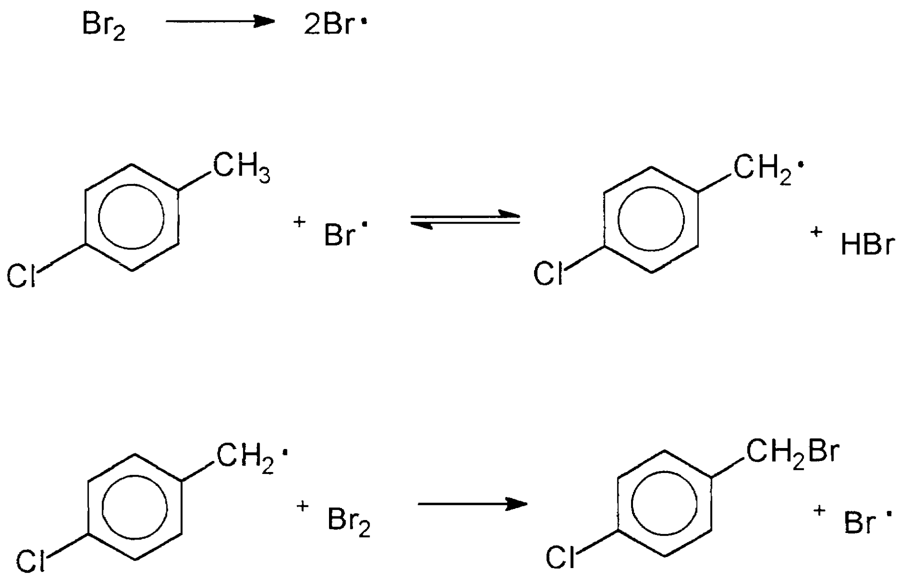
Reaction of the Bromine Atom with Cyclopropylbenzene

As previously mentioned, cyclopropylbenzene is known to react with bromine to yield almost exclusively a 1,3-addition product (1,3 dibromophenyl propane) upon photoinitiation (scheme VIII). However, the presence of electrophilic aromatic substitution under dark conditions has been shown by LaLonde and others.[9] This thermal electrophilic process has also been shown to be kinetically slower than the photoinitiated free radical process.

While our research has been conducted entirely under light (free radical) conditions, the reactants were under dark conditions during our equilibration period (5-10 minutes). However, past work has shown that no polar product is formed during this length of time.[2] Furthermore, our results clearly show that the major product of all reactions (with the exception of pentane) was due to 1,3-addition. This is in agreement with the findings of LaLonde and others.[9]

Reaction of Methylarenes with the Bromine Atom

The mechanism for the reaction of bromine with toluene has been shown previously (Scheme II). A similar mechanism exists for the substrate *p*-chlorotoluene (Scheme IX).



Scheme IX

While the mechanisms are essentially the same, the rate of hydrogen abstraction from the para chloro substituted toluene is slower. Tanner demonstrated this for three Freon solvents in 1990.

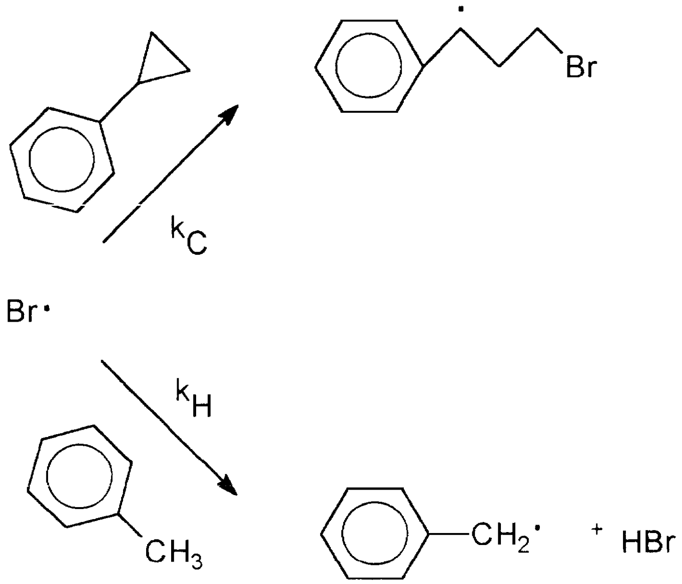
Table 5[5]: The relative rates of hydrogen abstraction for toluene and *p*-chlorotoluene in Freon solvents.

Solvent	Temp (o C)	Tol	PCTol
Freon 11	-18	1.0	0.66
Freon 113	23	1.0	0.545
Freon 112	70	1.0	0.583

It has further been suggested by Russell that benzylic abstraction by the bromine atom is sensitive to polar effects.[10]

Competition of Phenylcyclopropane and Toluene

The major competition used in our work was between phenylcyclopropane and toluene. In this competition series, the reaction of the phenylcyclopropane represents a S_H2 (substitution, homolytic, bimolecular) formal carbon abstraction process (k_c). Reaction of the toluene represents the hydrogen abstraction process (k_H). An overview of this competition is shown by scheme X.



Scheme X

From the reaction of this competition series in fourteen different solvents the following relative rates of the processes were obtained.

Table 6: Relative rate constants observed for CPB versus toluene competitions at 21°C.

Solvent	k_C/k_H
CCl_4	1.75±0.2
CS_2	2.47±0.1
PhH	2.91±0.2
Freon 113	1.01±0.1
pentane	0.558±0.1
$CHCl_3$	3.25±0.1
$ClCH_2CH_2Cl$	4.17±0.5
$BrCH_2CH_2Br$	4.30±0.1
PhCl	3.55±0.1
PhBr	4.00±0.1
CH_2Cl_2	3.90±0.1
Freon 11	1.17±0.1
Freon 112	1.51±0.1
1-Bromonaphthalene	9.69±1.0

Another point to be addressed in our work was the role of "in cage" HBr reversal as proposed by Tanner. According to previous work , the extent of "in cage" reversal should be partially dictated by the cage lifetime. This in turn should be related to the viscosity of the solution. This means that if "in cage" reversal was responsible for the variation in rate constants there shown be a strong correlation between solvent viscosity and the relative rate observed.

A model developed by Noyes {using the equation $k_C/k_H=k_C/k^0_H\ (1+F^{-1})= k_C/k^0_H+k_C/k^0_H(F^{-1})$} predicts an inverse relationship between $F = 1/\beta-1$(with β representing the cage efficiency) and η. If it is assumed that k_C and k^0_H are viscosity independent, the Noyes model anticipates that a plot of k_C/k_H vs η would be linear if geminate HBr reversal was the source of the observed solvent effect. From Noyes work, Koenig developed a model for the diffusion of geminate radical pairs which predicts that F is proportional to $\eta^{-1/2}$. Based on Noyes equation. this should predict a linear relationship between k_C/k_H and $\eta^{1/2}$.

The results of our work showed little if any correlation between the square root of solvents viscosities and the observed rate constants. The actual viscosities (calculated relative to Freon 11 at 17°C) are shown in table 7.

Table 7: Viscosities observed for competition solvents relative to Freon 11 at 17°C.

Solvent	Rel. Viscosity
CCl_4	0.82
CS_2	0.38
PhH	0.61
Freon 113	0.66
pentane	0.27
$CHCl_3$	0.59
$ClCH_2CH_2Cl$	0.74
CH_2Cl_2	0.47
$BrCH_2CH_2Br$	1.35
PhCl	0.73
PhBr	1.10
Freon 11	0.42
Freon 112	1.10
1-Bromonaphthalene	3.00

In plotting the relative rate constants versus the obtained viscosity some correlation was found (fig.7).

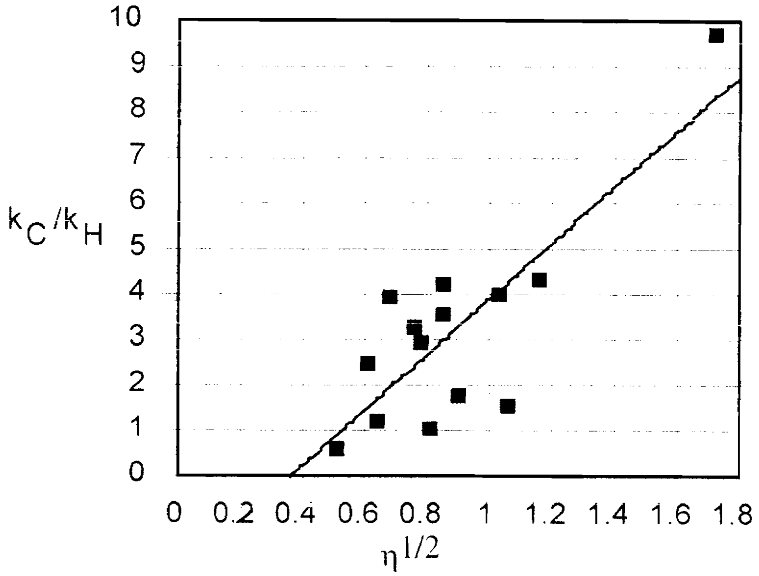


Figure 7: Plot of k_C/k_H versus the square root of the solvents' viscosities

However, this correlation with viscosity (R = 0.8124) is highly dependent on the 1-bromonaphthalene data point. removing this point reduces R to 0.0586. This lack of correlation among the bulk of our data led us to believe that some effect other than internal return must be responsible for our observed variation in k_c/k_H .

Solvent polarity was next studied as the possible cause of variation among our relative rates. The basis of this study was the Clausius-Mosotti equation

Equation 17

$$P_m = M/d * (\varepsilon-1)/(\varepsilon+2) = 4\pi N/3 * [\alpha+ (\mu^2/kT)]$$

which relates ε (the dielectric constant), μ (the dipole moment) and P_m (the molar polarization) to the rate constant k. Additionally, the equation uses the molecular weight (M), density (d) andthe polarizability (α) of the solvent.

First a plot of the relative rate constants versus the dipole moment of the solvents' was examined (Figure 8).

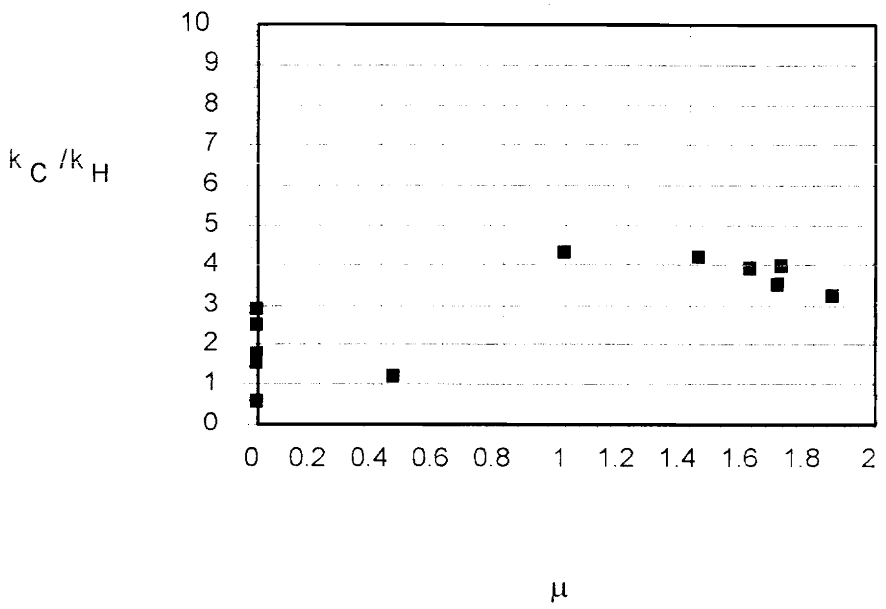


Figure 8: Plot of k_C/k_H versus the solvents' dipole moments

This plot also showed only a modest correlation (R= 0.7653). Next the relative rate constants were plotted versus the dielectric constant of each solvent employing the Kirkwood function. This treatment descibes the free energy of transfer for a dipole moment μ in changing form a medium having ε=1 to a medium of dielectric constant ε. It predicts an increase in rate with increasing solvent polarity and also predicts a linear relationship between the log of the rate constants and (ε - 1) / (2ε + 1).[11]

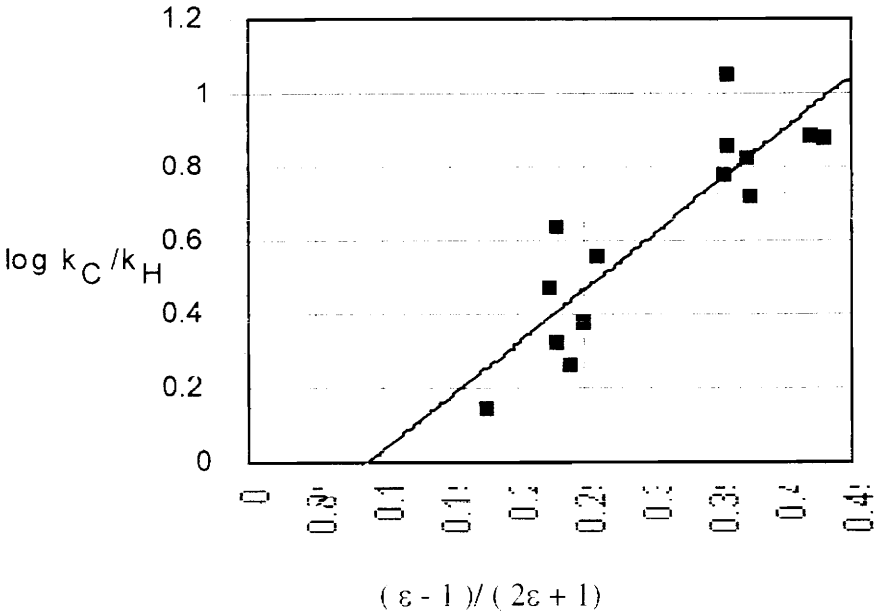


Figure 9: Plot of log k_C/k_H versus the Kirkwood function

This did show some correlation but not enough to explain the observed changes in relative rate (R = 0.8740).

Upon further examination of solvent parameters, it was proposed that internal pressure was responsible for the observed effect. The Hildebrand solubility parameter , a factor of the solvent's energy of vaporization and its molar volume, was used to approximate internal pressure. It was predicted that the relative rate constant (k_C/k_H) would vary with Hildebrand's parameter as long as the volume of activation was not equal to zero.

A difference in the activation volumes of two processes can be rationalized by considering the # moles of reactant per mole of product in each case. In the case of hydrogen abstraction , two moles of reactants (Br• and toluene) yield two moles of product (benzyl bromide and HBr). For the S_H2 case , two moles of reactant (phenylcyclopropane and Br•) react to yield just one mole of product (Ph-CH•-CH_2CH_2Br). A plot of k_C/k_H versus δ is shown in figure 10.

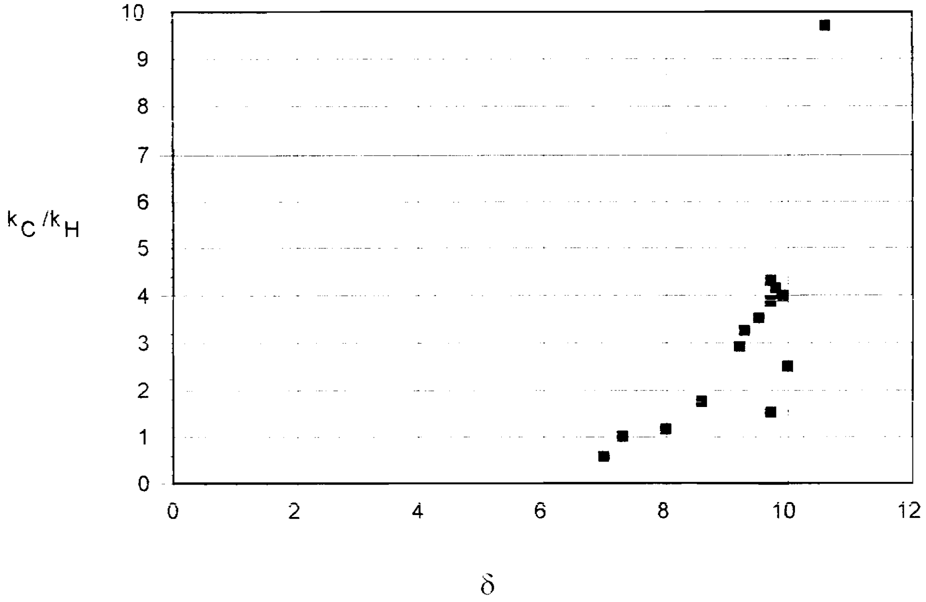


Figure 10: Plot of k_C/k_H versus δ

A better fit of the data was obtained by plotting $\ln(k_C/k_H)$ versus δ^2 (figure 11). This plot allows us to correlate the observed changes in rate constant with the cohesive energy density of the solvent which serves as an approximation for internal pressure.

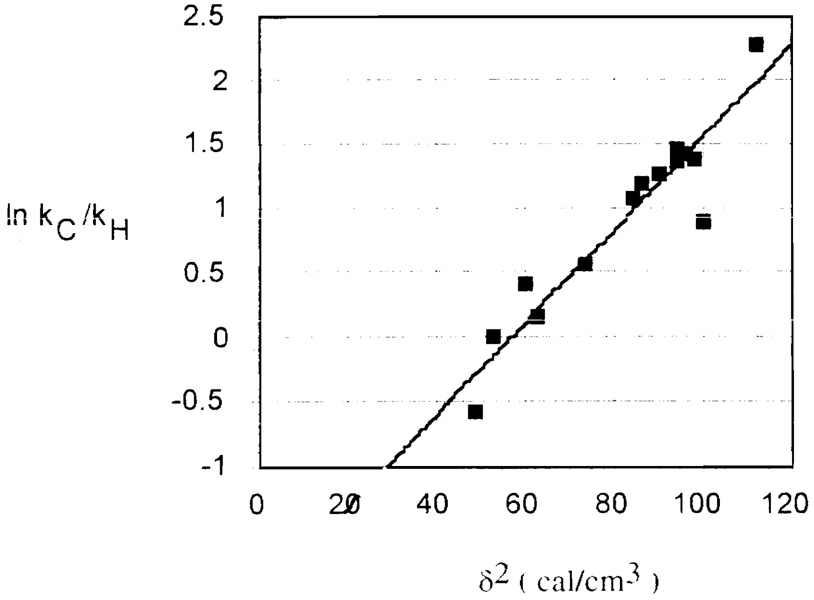


Figure 11: Plot of ln k_C/k_H versus δ^2

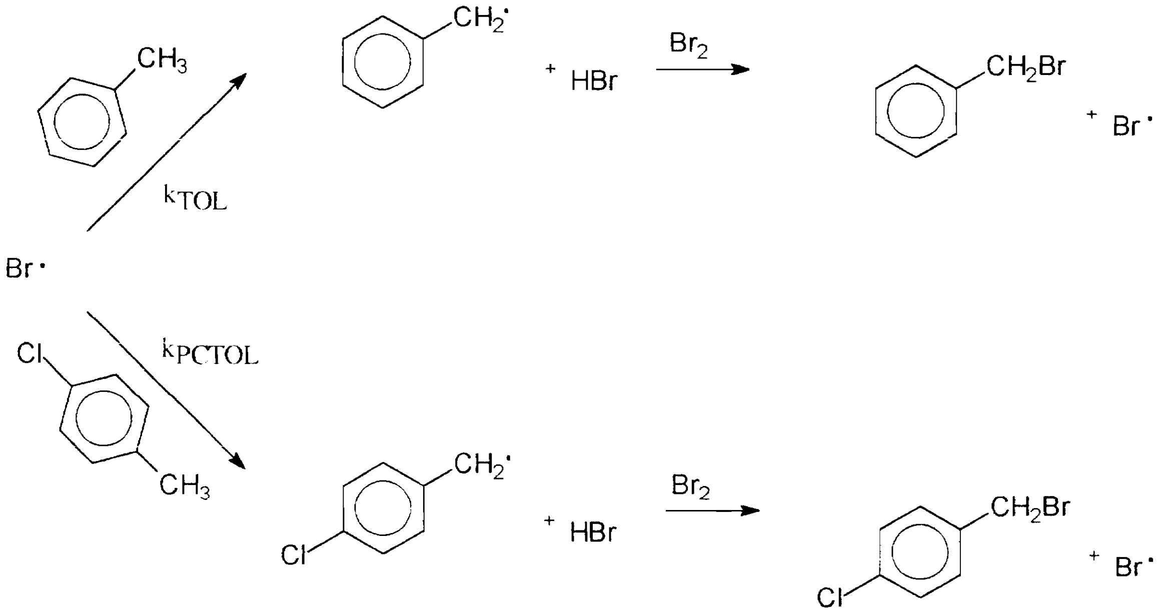
Linear regression analysis of this fit gave a correlation coefficient of R= 0.9463. An explanation for this correlation can be made by assuming that $\Delta V^{\neq}$for theS_H2 process is more negative than for the hydrogen abstraction process. With the δ^2 term in equation 11 assumed to be dominant, a least square fit of figure 11 gives a slope of m=0.038±0.003 cm^3/cal. This corresponds to a $\Delta\Delta V^{\neq}$=-22±1 cm^3/mol=$\Delta V^{\neq}_C$-$\Delta V^{\neq}_H$. This correlation suggests that it is indeed internal pressure which is causing the observed changes in our relative rates.

It is important to note that several assumptions must be made in order the explain the observed solvent pressure effect through use of equation 11. The first of these is that the solutions must be treated as "regular solutions" meaning that there is a random distribution of molecules and a lack of any specific solute-solvent interactions(hydrogen bonding, dipole-dipole, etc...). Secondly, it must be assumed that V_A,V_B,$V_{\neq}$,δ_A,δ_B and $\delta_{\neq}$ are independent of solvent and can therefore be treated as constants. Finally the use of

δ^2(c.e.d.) in place of internal pressure(P_i) is only permissible in the absence of strong intermolecular interactions such as hydrogen bonding. These assumptions limit the use of this treatment primarily to reactions of non-polar species in non-polar solvents.

Competitions of Toluene and *p*-Chlorotoluene

Another competition series was run with *p*-chlorotoluene competing with toluene for the bromine atom to examine the relative reactivity of the species.



Scheme VII

As previously mentioned the polar effect of the para chloro substituent reduces reactivity. Thus values of k_{PCTol}/k_{Tol} would be expected to be less than one. Indeed this is the case in all solvents. Table 8 shows the observed relative rates for each solvent.Of particular interest are the values for the three Freon solvents, these values are in close agreement with values determined by Tanner in 1990(see Table5,p.28).

Table 8: Observed relative rate constants for *p*-chlorotoluene versus toluene competitions in selected solvents.

Solvent	k_{PCTol}/k_{Tol}
Ph-Cl	0.610
Freon 11	0.643
Freon 113	0.611
Freon 112	0.670
CCl_4	0.689
Ph-H	0.547
Ph-Br	0.554
1-Bromonaphthalene	0.551
CS_2	0.450

While the relative rates were similar to those expected, a plot of ln k_{PCTol}/k_{Tol} versus δ^2 did not show a correlation similar to that obtained with the phenylcyclopropane vs toluene competitions (R = 0.5078).

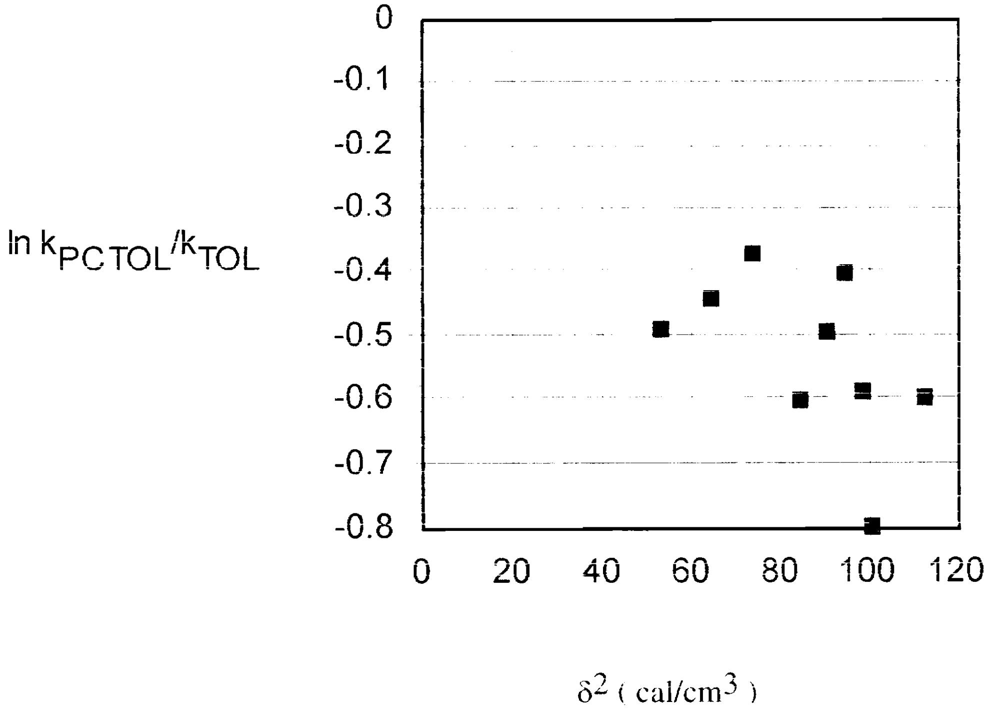


Figure 12: Plot of the natural log of k_{PCTol}/k_{Tol} versus δ^2

These results along with Figure 13 further support the idea that viscosity and consequently internal return was not the reason for the observed changes in relative rate.

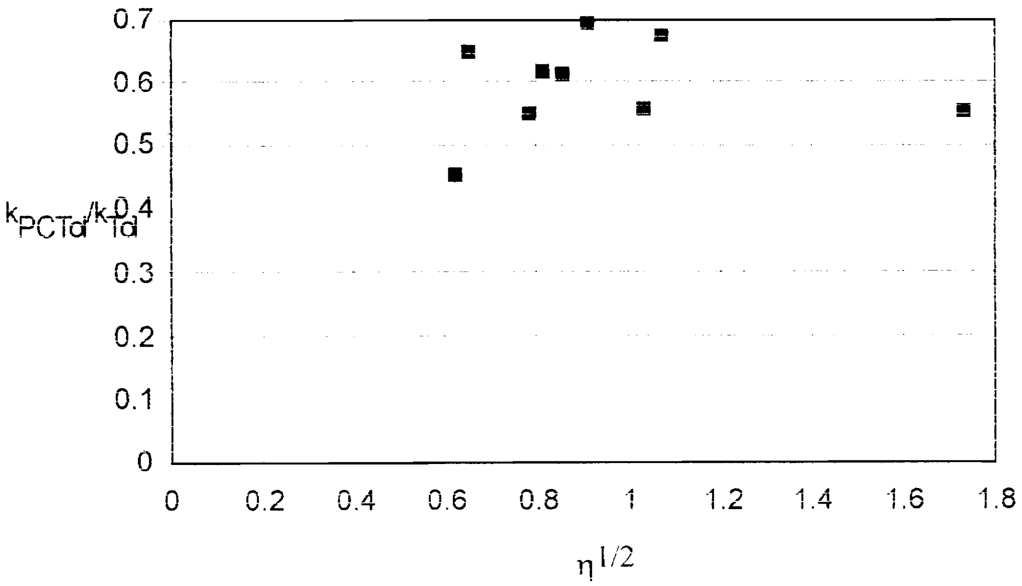


Figure 13: Plot of k_{PCTol}/k_{Tol} versus the square root of the solvents' viscosities

Had viscosity been the primary factor, the relative rate k_{PCTol}/k_{Tol} should have varied linearly with the square root of the viscosity. Our results show little change in k_{PCTol}/k_{Tol} with changes in solvent (R=0.2503).

Similarly, a plot of k_{PCTol}/k_{Tol} versus the solvents' dielectric constants (Figure 14) showed little correlation (R=0.5304).

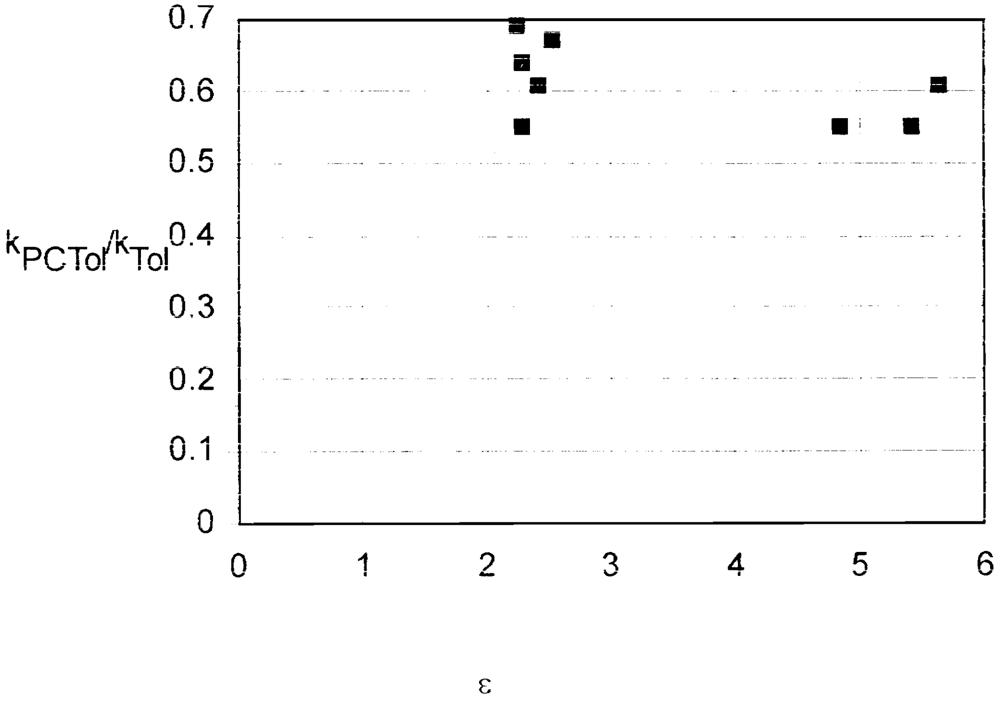
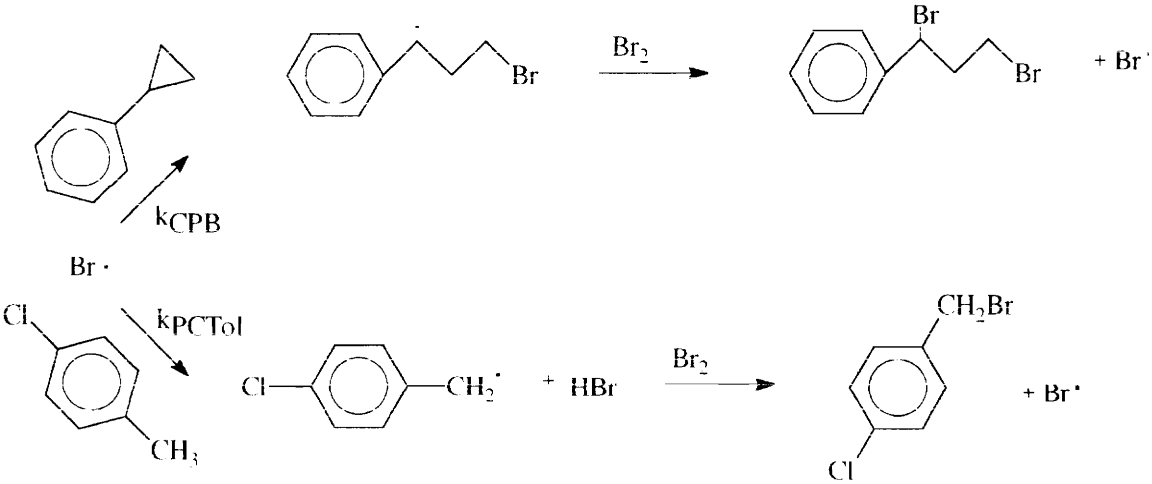


Figure 14: Plot of k_{PCTol}/k_{Tol} versus the solvents' dielectric constants

Competitions of Phenylcyclopropane and *p*-Chlorotoluene

A third series of competitions for the bromine atom was run with phenylcyclopropane and *p*-chlorotoluene to test the generality of the solvent effect. (Scheme XI).



Scheme XI

This series yielded rather high values for k_{CPB}/k_{PCTol} due to the differences in reactivity of the two species. In an extreme case, 1-bromonaphthalene , values for k_{CPB}/k_{PCTol} were very high and irreproducible. The phenylcyclopropane was much more reactive than the *p*-chlorotoluene therefore only a small amount of *p*-chlorobenzyl bromide was produced. Subsequently the amount of phenylcyclopropane was reduced to half its normal amount to allow a more measurable amount of product from the *p*-chlorotoluene to form.

The relative rates observed for this competition series are summarized in Table 9.

Table 9: Relative rate constants observed for CPB versus *p*-chlorotoluene competitions for bromine atom at 21°C.

Solvent	k_{CPB}/k_{PCTol}
CCl_4	2.95
CS_2	3.59
Ph-H	4.31
Freon 113	1.82
pentane	1.40
$CHCl_3$	5.94
$ClCH_2CH_2Cl$	7.58
CH_2Cl_2	7.68
$BrCH_2CH_2Br$	7.21
Ph-Cl	5.22
Ph-Br	6.64
Freon 11	2.10
Freon 112	2.38
1-Bromonaphthalene	11.20

A plot of ln k_C/k_H versus δ^2 for this series yields a correlation (R = 0.9301) similar to that found with phenylcyclopropane and toluene (figure 15).

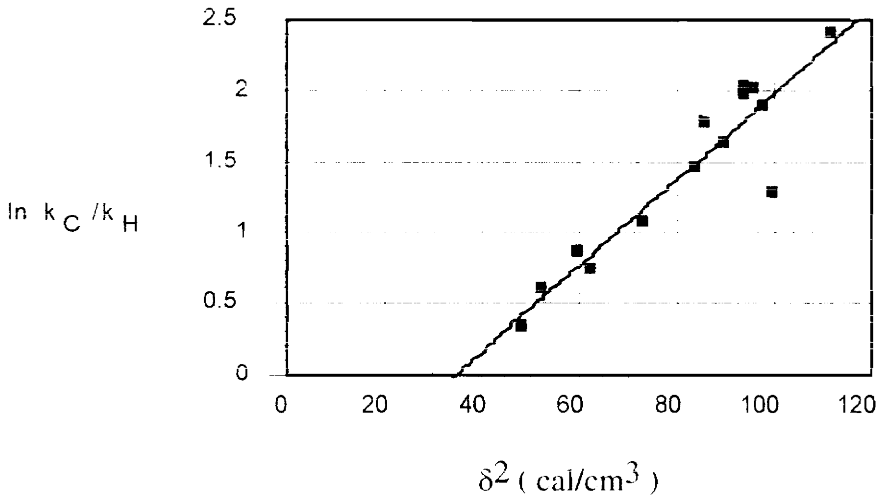


Figure 15: Plot of the natural log of k_{CPB}/k_{PCTol} versus δ^2

The most notable difference in the two plots is that due to the decreased reactivity of the *p*-chlorotoluene , the plot for this series fails to pass below log $k_C/k_H = 0$ ($k_C/k_H = 1$). Thus the rate of formal carbon abstraction exceeds the rate of hydrogen abstraction for the entire pressure range studied.

Making the same assumption that the δ^2 term of equation 11 is dominant (as made with figure 11) gives a slope of $m=0.031\pm0.004$ cm^3/cal that is almost identical to that obtained for the cyclopropylbenzene versus toluene competition.

Conclusions :

Through our research a selectivity efffect due to solvent for bromine atom has been found. In examining several solvent parameters internal pressure was found to be the cause of the observed effect. More convenentional solvent parameters such as polarity and viscosity were unable to explain the cause of the observed effect.

Our study focused on systems that closely approximate regular solutions with exclusively non-polar solvents being used. This assured that van der Waals forces were dominant within the solutions. The use of non-polar solvents also allowed us to use cohesive energy density in place of internal pressure since the two are approximately equal for these solutions.

The results of our studies show excellent reproducibility and allow us to estimate a difference in the volumes of activation ($\Delta\Delta V^{\neq}$= -20cm^3/mol) based on the observed pressure dependence of k_C/k_H. Future experiments in this area should pit different mechanistic processes against each other to see if similar trends can be found. Gas phase experiments should also provide additional information involving the effects of internal pressure.

Experimental:

General Considerations

All gas chromatographic analyses were performed on a Hewlett-Packard HP5890A instrument equipped with dual flame ionization (FID) detectors. Results were reported on an HP 3393A integrator. Analyses were run through either a Supelco SE-54 (30m X 0.25mm) capillary column with a 0.25μm film thickness or on an Alltech SE-54 (15m X 0.25mm) with a 0.25μm film thickness. All relative viscosities taken were measured on an Oswald microviscometer using 4ml of solution.

General Competition Experiments

A stock solution was prepared for each competition series with the solvent under investigation being used. Appropriate amounts of the competing species, along with an internal standard (*o*-dichlorobenzene) were added to a clean, dry volumetric flask. A Mettler analytical balance was used to measure added volumes by actual weight. The solution was then diluted with the chosen solvent and then stoppered and shaken to insure a homogeneous solution.

5ml of the stock solution was then pipeted into a 30-ml pressure tube along with 100μl of 1,2-epoxybutane , serving as an HBr scavenger , and then fitted with an O-ringed Teflon valve and sealed. The solution was then degassed 3-4 times via (freeze-pump-thaw) under high vacuum at -196 C to remove any oxygen present. 20μl (0.389mmoles) of molecular bromine was then degassed and transferred into the pressure tube via vacuum distillation.

The sealed tube was then placed in a water bath (T_{H_2O} = 21+/- 1 o C) and allowed to equilibrate for 5 minutes. The solution was then irradiated with 1-2 400-W medium pressure mercury arc lamps (at a distance of 7-10 inches) until the bromine color was completely discharged.

The products were then analyzed by GLC using predetermined response factors with o-dichlorobenzene as an internal standard.

CPB vs Toluene Competitions

A typical competition consisted of approximately 2.2 mmoles of cyclopropylbenzene and about 3.5 mmoles of toluene as the competing species. Additionally 1.7-1.8 mmoles of o- dichlorobenzene was added as an internal standard.

CPB vs p-chlorotoluene Competitions

Competitions of this series consisted of 2.2 mmoles of cyclopropylbenzene and about 3.5 mmoles of p-chlorotoluene per reaction. 1.7-1.8 mmoles of o-dichlorobenzene were also added as an internal standard.

p-chlorotoluene vs Toluene Competitions

Each competition consisted of 3.5 mmoles of p-chlorotoluene and 3.7 mmoles of toluene along with 1.7-1.8 mmoles of o-dichlorobenzene.

Quantitation by GLC

The peak area counts of the competition products along with those of the internal standard were used to calculate product yields by the following equation.

Equation 18

n_f Product = Area product/ Area standard X (#moles std.) X (response factor)

All analyses were performed in triplicate.

Calculation of Relative Rates

All relative rates were calculated using the equation $k_A/k_B = A_f \text{ X } B_0 / B_f \text{ X } A_0$ (with A_0 and B_0 representing the initial amounts of the two competitors and A_f and B_f representing the product yields obtained by G.C. analysis). This equation was derived from the rate laws:

$$d[A]_f/ dt = k_C [Br\bullet] [A]_i$$

$$d[B]_f/dt = k_H [Br\bullet] [B]_i$$

Calculation of Relative Viscosities

Viscosities for all prepared solutions were done relative to trichlorofluoromethane (Freon 11) on an Oswald microviscometer. The viscosities were calculated using the equation

Equation 19

$$\eta_{sol'n} = p \times (t_a - t_b)$$

where ρ represents the solution's density and $(t_a - t_b)$ represents the time required for the solution to travel between the two fiducial marks.

This is derived from the viscosity equation for an Oswald viscometer

Equation 20

$$\eta/\eta^* = t/t^* \times \rho/\rho^*$$

in which the relative viscosity is proportional to the drainage times after correcting for the different densities.

Materials

The following materials were purchased from the Aldrich chemical company and used as received (unless otherwise noted) methylene chloride (HPLC grade) , 1,2-epoxybutane , 1,1,2-trichlorotrifluoroethane (Freon 113 , HPLC grade) , chlorobenzene (HPLC grade) , cyclopropylbenzene and bromine. 1-Bromonaphthalene, p-chlorotoluene and carbon disulfide (HPLC grade) were fractionally distilled to remove any radical inhibitors. Chloroform (HPLC grade) was fractionally distilled over phosphorous pentoxide and passed through alumina. Benzene (A.C.S. grade) was repeatedly washed with sulfuric acid and fractionally distilled over calcium carbonate. Carbontetrachloride (A.C.S. grade) was slurried in potassium hydroxide for 24 hours and then fractionally distilled from phosphorous pentoxide.

Toluene (A.C.S. grade) and ethylene dichloride were purchased from the Baker chemical company with the toluene being distilled from potassium metal with benzophenone used as an indicator.

Ethylene bromide and pentane were purchased from the Fisher chemical company and were used as received.

Bromobenzene was purchased from the MCB chemical company and used as received.

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Vita

Glenn A. Hulvey was born on May 21, 1964 in Waynesboro, Virginia to Glenn and Doris Hulvey. He received his Bachelor of Science degree in Chemistry from Bridgewater College in 1990.

In the fall of 1990, he enrolled in the Chemistry graduate program at Virginia Polytechnical Institute and State University. There he worked under the guidance of Dr. James M. Tanko in researching solvent pressure effects. This effort resulted in the completion of this thesis. While at Virginia Tech, he held a Graduate Teaching Assistant position and taught General Chemistry Laboratory.

He is a member of the American Chemical Society.

Glenn A. Hulvey