AN ABSTRACT OF THE THESIS OF

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Title: Mechanistic Studies of Chlorination

            of Aromatic Compounds Using Bleach

Abstract approved: _Kenneth F. Johnson_

The mechanism of chlorination of aromatic substances using a two-layer system with bleach as an aqueous phase and a catalyst to ensure phase transfer has been causing controversy for years. The kinetic studies of the chlorination of the α-position in toluene, the o,p-positions in anisole and the ipso-position in bromobenzene were conducted. The obtained reaction orders were mixed "0+1" for anisole and toluene and 2nd for bromobenzene, as well as 1st for chloride ion. The influence of a phase-transfer catalyst and a change of solvent on the rate of anisole chlorination and on its selectivity were also investigated. Based on the kinetic information available, a detailed mechanism is being proposed. The ratio of para- and ortho-isomers in anisole chlorination was calculated theoretically based on the proposed scheme and then compared to the experimental data. A definite correlation between the two series can be regarded as another evidence supporting the idea of a free-radical aromatic substitution mechanism involving both free $Cl\cdot$ and stabilized $Cl\cdot$ as the species reacting in a competitive manner.

MECHANISTIC STUDIES

OF CHLORINATION OF AROMATIC COMPOUNDS

USING BLEACH

———————

A Thesis

Presented to

the Division of Physical Sciences

EMPORIA STATE UNIVERSITY

———————

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by

Alexander M. Aronov

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Approved for the Major Division

Faye N. Vowell
Approved for the Graduate Council

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Dedicated to

Gregory M. Aronov

and

Yefim A. Dashevsky

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I. INTRODUCTION

1.1 Historical background

During the early growth of organic chemistry a major objective developed with the aim of isolating the parent organic "radicals" constituting the organic molecules in much the same way that it had been previously accomplished in inorganic chemistry. However, when these efforts proved to be fruitless, the idea was abandoned until 1900 when the first free radical ($Ph_3C\cdot$) was shown to be present in solution. Since that time the chemistry of free radicals has been rapidly advancing, thus catching up with its more developed "sisters", the chemistry of electrophiles and the chemistry of nucleophiles. Nevertheless, it has been unable to quite get out of their shadow - numerous textbooks still describe free radicals as "electrophile-like electron deficient species."

The term "radical" applies to any species that possesses an unpaired electron which is more or less localized on a non-metallic element (most often carbon). Typically, organic free radicals have a trivalent carbon with seven electrons in its valence shell (three bonds and an unpaired electron).

The standard methods of production of radicals include (1) thermolysis, e.g.:

(1.1) $\qquad$ R-O-O-R $\longrightarrow$ 2 R-O$\cdot$

(1.2) $\qquad$ R-COO-OOC-R $\longrightarrow$ 2 R$\cdot$ + 2 CO_2

(1.3) $$R-N=N-R \longrightarrow 2 \; R\cdot \; + \; N_2$$
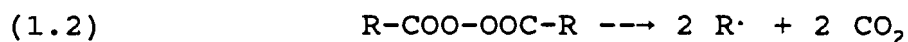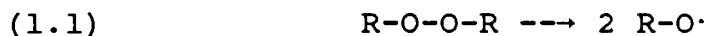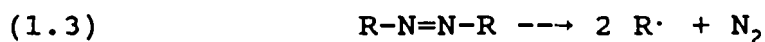
(2) photolysis - illumination of a substance with a high-energy (usually UV) lamp to cause homolysis of the bonds

(3) radiolysis - bombardment of a substance with an electron beam, which removes the valent electron

(4) redox reactions, the most famous of which is the Kolbe electrosynthesis of alkanes.

The most important characteristics of free-radical reactions are [1]:

(a) radical reactions are generally indifferent to the polar nature of the solvent;

(b) radical decompositions may be strongly influenced by solvents capable of forming a relatively stable radical by loss of hydrogen (induced decomposition);

(c) radical reactions often have chain mechanisms, which include initiation, propagation and termination steps;

(d) the rates of free-radical reactions are generally much less dependent on the substituent effects than those of ionic reactions: ρ values in Hammett plots are relatively small (Isaacs [1] cites the range from -1.4 to -0.5); and

(e) the rates of free-radical reactions are often extremely sensitive to the presence of certain impurities capable of scavenging the initiator radicals and inhibiting the chain reactions.

Important mechanistic information is often obtained from analyzing the minor reaction products which can give insights

into the reaction mechanism.

The three conversions that became the centerpiece of the current investigation are:

(1.4)   PhBr ---→ PhCl

(1.5)   PhOCH$_3$ ---→ p,o-CH$_3$OC$_6$H$_4$Cl

(1.6)   PhCH$_3$ ---→ PhCH$_2$Cl
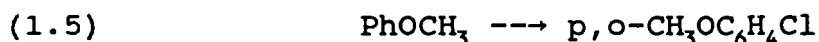
The reaction of bromobenzene has been under investigation for over a century. In 1890 Srpek [2] reported that chlorination of p-bromotoluene gave an inseparable mixture of chlorobenzyl and bromobenzyl chlorides and bromides. The chlorination of the α-position in toluene derivatives has been reported numerous times. In 1934 Asinger [3] confirmed the replacement of bromine attached to an aromatic ring by chlorine. Chlorodebromination in various substrates was cited in the 1950s and 1960s. The research of Miller and Walling [4] in 1957, establishing the free-radical character of the reaction based on the influence of light on accelerating the reaction and the inhibition studies, became the starting point for all further investigations in the field.

Echols et al. [5] undertook the first effort to establish the rate law for the ipso-chlorination of bromobenzene with chlorine gas. The reported rate law was:

$$d[PhCl]/dt = [PhBr][Br\cdot]\{k_1[Cl_2] + k_2[BrCl]\}.$$

Without adding bromine chloride to the mixture, which accelerated the reaction, $[Cl_2] > [BrCl]$ and

$$d[PhCl]/dt \approx k_1[PhBr][Br\cdot][Cl_2].$$

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This was the first work proposing the participation of a free radical in the abstraction of the leaving group; the authors regarded this as a rate-determining step. They also noted that an induction period followed by a sharp rise in product concentration was a feature of virtually every run. The following mechanism was postulated:

$$PhBr + Cl_2 \rightleftharpoons Br\text{-}C_6H_5\text{-}Cl_2$$

$$Cl\text{-}Cl\text{-}C_6H_5\text{-}Br + Br\cdot \rightarrow [Cl\cdots Cl \cdots Br\cdots Br] \rightarrow Cl\cdot + C_6H_5\text{-}Cl + Br_2$$

$$Br\text{-}Cl\text{-}C_6H_5\text{-}Br + Br\cdot \rightarrow [Br\cdots Cl \cdots Br\cdots Br] \rightleftharpoons Br\cdot + C_6H_5\text{-}Cl + Br_2$$
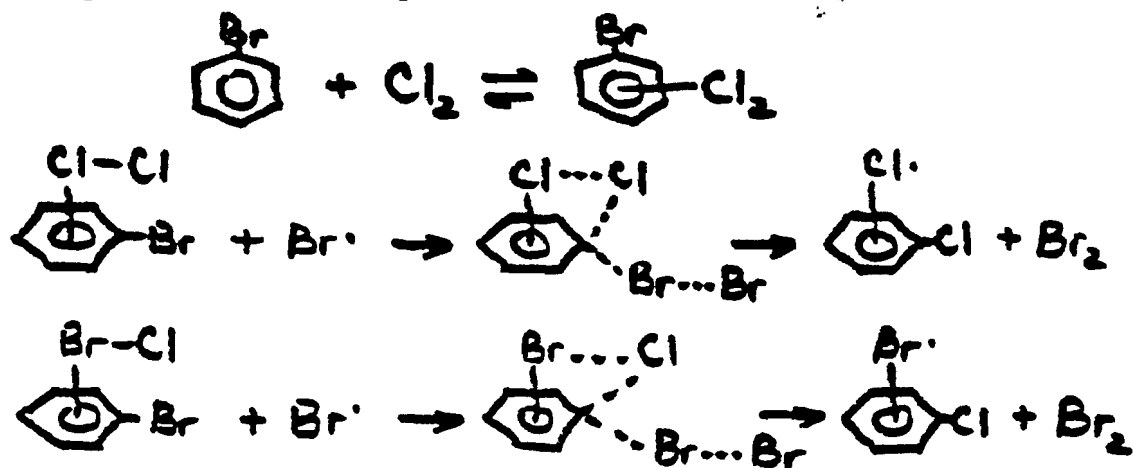
However, certain inconsistencies remained. First, the role of the bromine atom as the species abstracting the leaving group, could be questionable in case of low substrate concentrations. Second, along with the postulated first order in PhBr the authors acknowledged that the 100-fold reduction in the concentration of PhBr resulted in a roughly 10-fold reduction in the rate when all other factors were held constant. These results are examined in 4.5.

Other halodehalogenations were investigated as well. Bromodechlorination was achieved with Br_2 or even better with N-bromosuccinimide (NBS)[6]. Chlorodeiodination was compli-

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cated with the tendency of iodine to get involved in oxidative addition, going from I^{1-} to I^{1+}:

(1.7) $PhI + Cl_2 \dashrightarrow PhICl_2$
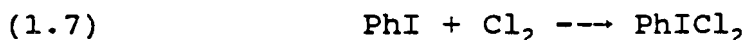
The substitution was accomplished by use of ICl or SO_2Cl_2/Bz_2O_2 [7].

Other examples of radical dehalogenations include the reactions of phenyl radicals (from benzoyl peroxide) and cyclohexyl radicals. Phenyl radicals replace both hydrogen and fluorine in fluorobenzene in 95:5 ratio [8]:

(1.8) $PhF + Ph\cdot \dashrightarrow PhC_6H_4F + Ph{-}Ph$
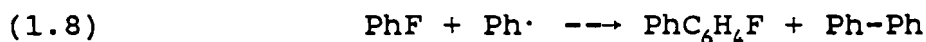
The reaction of PhBr with cyclohexyl radicals [9] is much less selective with a 1:5 statistical ratio of products:

(1.9) $PhBr + R\cdot \dashrightarrow Ph{-}R + RC_6H_4Br$
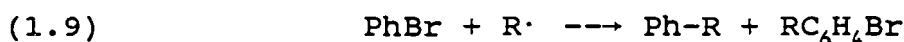
Nitro groups have been replaced through ipso-substitution with $Cl_2/h\nu$ [10].

An interesting feature of ipso-substitution is the possibility of the 1,2-migration of the attacking group following the formation of the σ-complex, which gives rise to a wider than expected variety of products A-E [11]:

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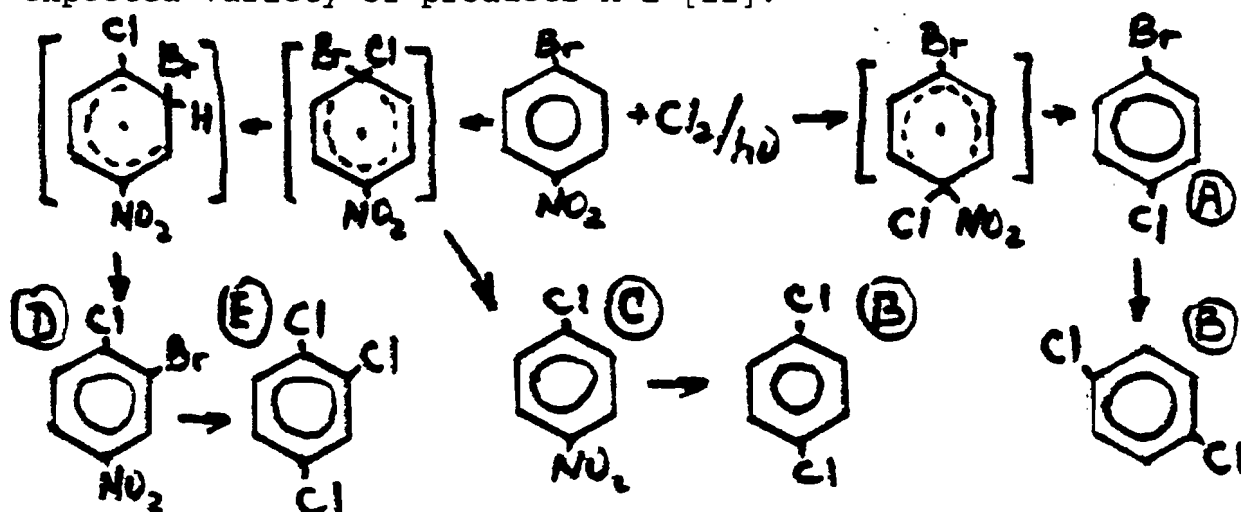
In 1974 Shelton and Lipman [12] obtained evidence that strongly favored abstraction of a leaving substituent by a free radical. The reaction of PhCl with cyclohexyl radicals produced cyclohexylbenzene and chlorocyclohexane. The latter can be formed only if another cyclohexyl molecule aids the previously formed σ-complex; otherwise HCl and no chlorocyclohexane would be observed. Based on several reports, Traynham [13] summarized it this way: the relative rates of σ-complex formation and decomposition depend on the identities of both the attacking radical and the leaving substituent and with different reactants both steps have been identified as rate-determining (slow).

In 1983 Hamilton with his group [14] used aqueous NaOCl solution in the presence of a PTC[1] to chlorinate toluene and anisole and to oxidize alkenes. They suggested ClO· as the reactive reagent in the system at pH 8-9. The postulated mechanism for the chlorination of anisole was:

$$(1.10) \qquad HOCl + ClO^- \longrightarrow Cl_2O + OH^-$$

$$(1.11) \qquad Cl_2O \longrightarrow Cl· + ClO·$$

$$(1.12) \qquad Cl· + ClO^- \longrightarrow Cl^- + ClO·$$

$$(1.13) \qquad PhOCH_3 + ClO· \longrightarrow PhOCH_3^+ + ClO^-$$

$$(1.14) \qquad PhOCH_3^{+·} + Cl_2O \longrightarrow CH_3C_6H_4Cl + H^+ + ClO·$$
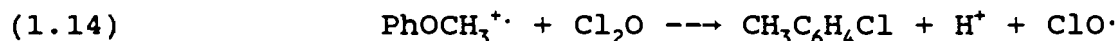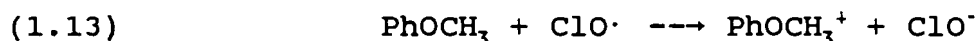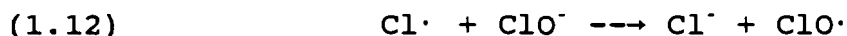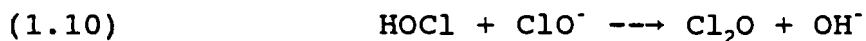
Toluene chlorination was explained with a prosaic chain mecha-

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[1] PTC (Phase Transfer Catalyst) - a compound used to ensure rapid transfer of substances between phases and to overcome the "surface barrier" in multiphase reactions. In the current project tetrabutylammonium bromide was used for these purposes.

nism (Bz = benzyl):

(1.15)   ClO· + BzH ---→ HClO + Bz·

(1.16)   Bz· + Cl$_2$O ---→ BzCl + ClO·
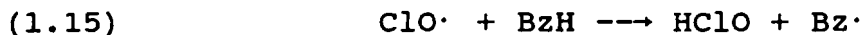
The Hammett correlation for toluene presented in the article gave ρ^+=-1.7, which is slightly out of the range cited by Isaacs, but is still sufficiently low to characterize the free-radical mechanism.   Acceleration, observed upon the addition of inorganic salts like sodium chloride and sodium nitrate, was explained as a 'salting out' effect, i.e. the decrease in the solubility of certain organic substances in water due to the increase in ionic strength of the solution [15]. On the basis of the selectivity studies of the chlorination of alkanes the conclusion was made that "none of the reactivity characteristics of the hypochlorite-PTC reagent are closely related to those of Cl· ".

In 1987 Hanzlik et al. [16] measured the deuterium isotope effect for a group of radicals. The isotope effect had a value of 6.37 for Br· and 5.90 for the radical in the hypochlorite-PTC system (CCl$_4$- solvent). The value for Cl· is significantly lower, e.g. Wiberg and Slaugh [17] obtained the values of 1.30 for Cl$_2$, 1.42 for SO$_2$Cl$_2$, 4.59 for Br$_2$, and 4.86 for NBS in CCl$_4$ at 77°C.

Later, Landolt and coworkers [18,19] converted PhBr to PhCl and PhI to PhCl (with iodylbenzene PhIO$_2$ as an intermediate) by using bleach in the presence of PTC at pH 7.5-9. Landolt, however, didn't propose any specific intermediate.

Nickel complexes are known for their ability to catalyze halogen exchanges [20]. Some Ni(II) square planar complexes were utilized in the nonphotochemical chlorination of PhBr [21,22]. A catalytic mechanism was proposed to explain the dramatic increase in the reaction rate (Fig.1). Pathway A leads to halodehalogenation in the nucleus; interestingly enough, the authors abandoned Hamilton's idea of the formation of a cation radical and used the σ-complex scheme proposed originally for the Cl_2/hv system. Pathway B was proposed to explain rapid epoxidation of olefins. While explaining the catalytic properties of Ni(II) complexes, the scheme doesn't answer the question of whether Cl· or ClO· is the true reactive species at pH < 9. According to the scheme, Ni(II) would catalyze the reaction:

(1.17) $Cl_2 \dashrightarrow 2\ Cl·$

as well as:

(1.18) $Cl_2O \dashrightarrow Cl· + ClO·$

1.2 Proposed investigation

The objective of the current investigation was to establish the true reactive species in the chlorination of aromatics at lower pH's using commercial bleach, to postulate a mechanism that would fit experimental data, and to explain the details of the mechanism using the facts and observations accumulated in the field of free-radical chemistry.

$$2 \, HOCl \rightleftharpoons Cl_2O$$
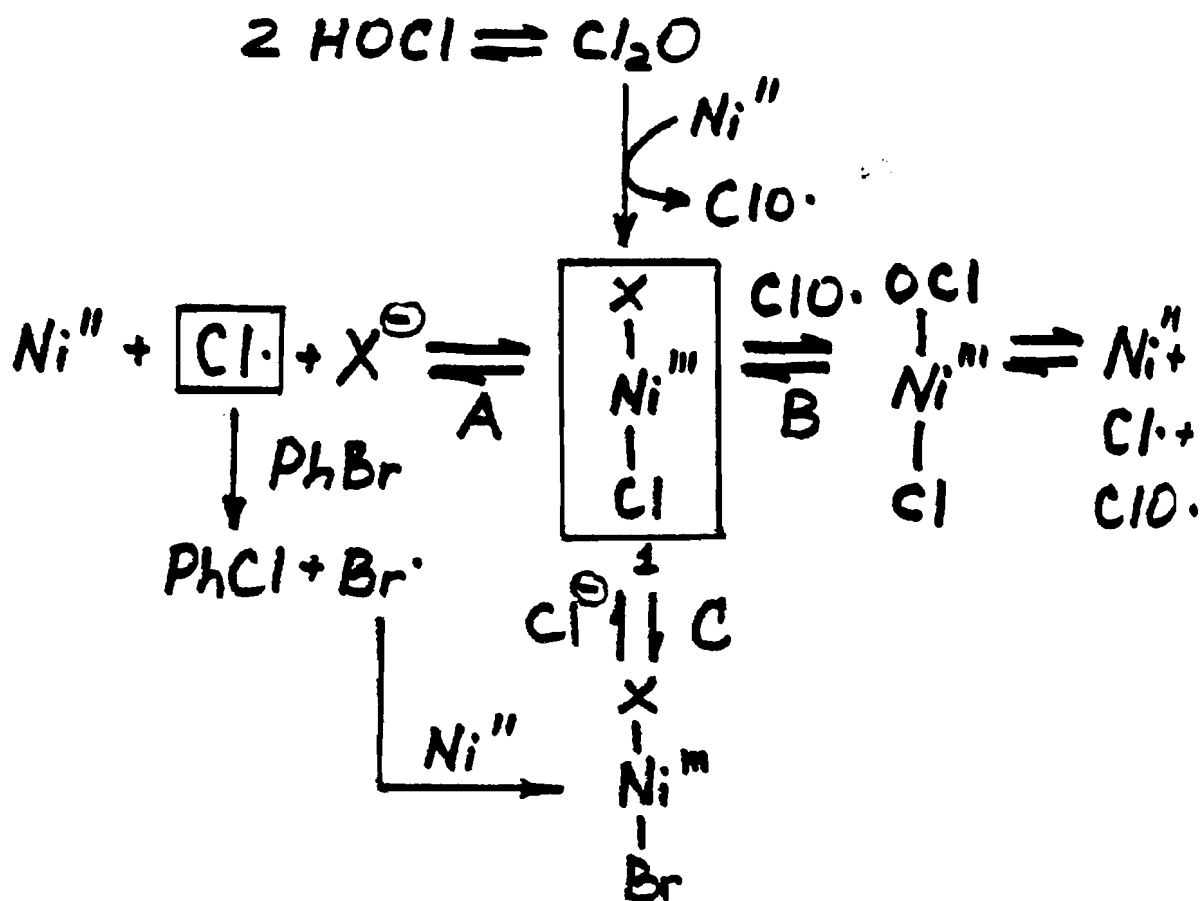
Figure 1   Catalysis of halogen exchange in aromatics by
nickel (II) complexes

II. EXPERIMENTAL

2.1 Equipment and Reagents

2.1.1 Chemical Equipment

(1) <u>Gas chromatograph</u>

A capillary Shimadzu GC-14A equipped with a capillary column 30 m x 0.25 mm with 0.25 μm film of SE30, a splitter, an FID detector and a temperature programmer was used to separate the components in the reacting mixture. Helium was used as a carrier gas. For experimental conditions see Section 2.2.4. The gas chromatograph was linked to a Hewlett-Packard 3390A integrator. Injections were made with a Hamilton Microliter® #701 syringe.

(2) <u>Stirring plate</u>

A Spinmark™ (Markson, Inc.) and a Magnestir (Scientific Products Co.) stirring plates were used to insure the uniformity of the mixture.

(3) <u>Glassware</u>

Pyrex beakers (e.g. 600 mL, 50 mL, etc.) and flasks (e.g. 500 mL, 125 mL, etc.), Kimax graduated pipets (e.g. 10 mL, 1 mL, etc.) and Corning 0.5 mL graduated pipets were employed.

(4) <u>Balance</u>

Sartorius 1265 MP (accurate to 1 mg) and Mettler H20T (accurate to 0.01 mg) balances were used.

(5) <u>pH pen</u>

A digital pH pen was used for measuring pH values of solutions.

2.1.2 Chemical Reagents

(1) Bromobenzene [C_6H_5Br]: Certified, Fisher Scientific

(2) Chlorobenzene [C_6H_5Cl]: Certified, Fisher Scientific

(3) Anisole [$C_6H_5OCH_3$]: Aldrich Chemical Company

(4) Toluene [$C_6H_5CH_3$]: J.T.Baker Chemical Company

(5) p-Dichlorobenzene [$C_6H_4Cl_2$]: Aldrich Chemical Company

(6) Benzene [C_6H_6]: Purified, Mallinckrodt Chemical Works

(7) p-Benzoquinone [$C_6H_4O_2$]: Eastman Organic Chemicals

(8) Tetrabutylammonium bromide [$(C_4H_9)_4N^+Br^-$]: Aldrich Chemical Company

(9) Galvinoxyl: Aldrich Chemical Company

(10) Cyclohexanone [$C_6H_{10}O$]: Purified Grade, Fisher Scientific

(11) Acetone [$(CH_3)_2CO$]: Certified A.C.S., Fisher Scientific

(12) Methylene Chloride [CH_2Cl_2]: Certified A.C.S., Fisher Scientific

(13) Chloroform [$CHCl_3$]: Certified A.C.S., Fisher Scientific

(14) Carbon Tetrachloride [CCl_4]: Gisco Company

(15) Clorox Regular Bleach [$NaOCl_{(aq)}+NaCl_{(aq)}$]: The Clorox Company

(16) Sulfuric Acid [H_2SO_4]: Reagent A.C.S., Fisher Scientific

(17) Hydrochloric Acid [HCl]: Reagent A.C.S., Fisher Scientific

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(18) Phosphoric Acid [H_3PO_4]: Mallinckrodt Chemical Works

(19) Sodium Hydroxide [NaOH]: Certified A.C.S., Fisher Scientific

(20) Potassium Hydroxide [KOH]: Certified A.C.S., Fisher Scientific

(21) Sodium Chloride [NaCl]: Fisher Scientific

(22) Sodium Nitrate [$NaNO_3$]: Mallinckrodt Chemical Works

(23) Potassium Iodide [KI]: Certified A.C.S., Fisher Scientific

(24) Potassium Dichromate [$K_2Cr_2O_7$]: Mallinckrodt Chemical Works

(25) Starch Indicator Solution: Fisher Scientific

2.2 Chemical Procedures

2.2.1 Reaction in the two-phase system

2.2.1.1 Kinetic studies

Certain amounts of substrates and internal standards (see Table I for the list of substrates and internal standards; amounts and masses are also given) were placed in a 600 mL beaker and dissolved in 25 mL of the solvent (methylene chloride unless stated otherwise). 1 mmol (320 mg) of a PTC

Table I

Substrates and internal standards used in the experiments

Compound	Amount, mmol	Mass, mg	Volume, ml (if applicable)	Molecular Weight	Density, g/ml(liq.)
Anisole (s.)	1.84	199	0.200	108.14	0.9961
Bromobenzene (s.)	1.00	157	0.105	157.01	1.4950
Chlorobenzene (i.s.)	1.00	113	0.100	112.56	1.1058
Dichlorobenzene (i.s.)	1.00	147	----	147.00	solid
Toluene (s.)	1.41	130	0.150	92.14	0.8669

(tetrabutylammonium bromide) was also dissolved.[2] 200 mL (unless specified) of an aqueous phase was added on top of the organic layer. The composition of the aqueous phase in this work is expressed by the percentile value of the concentration of inorganic salts in the given solution as compared to their concentration in the original Clorox bleach, which is referred to as a "100% bleach". The mixture was magnetically stirred. Additions of inorganic salts (in cases when they were used) were made and dissolved through stirring. After that the pH of the aqueous layer was adjusted by adding 50% H_2SO_4; a pH pen was used to monitor the pH throughout the experiment. At varying time intervals 1 μL aliquots of the organic layer were injected onto the GC column (see 2.2.4). The process of data acquisition continued until either an 80-90% conversion of the substrate (in case of bromobenzene) or until the disappearance of the substrate peak (in case of more reactive substrates: anisole or toluene). The data taken from chromatograms were later analyzed (see 2.3).

2.2.1.2 Selectivity studies

A 5-fold excess of a substrate was reacted in a system (the procedure was identical to the ones described in 2.2.1.1

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[2] If an amount of PTC other than 1 mmol was used, it will be specified by mass. In case of standard 1 mmol additions either a 'PTC' or '0.320 g' label will be used.

for two-phase and in 2.2.2 for single-phase systems, respectively). The organic mixture was analyzed by GC after 24 hrs. to determine the ratio of isomer products.

2.2.2 Reaction in the single-phase all-organic system

Two 500 mL flasks were used in these experiments. Flask #1 contained 50 mL of 100% bleach. The organic mixture (25 mL), composed of one or more of the following compounds: methylene chloride, bromobenzene, anisole, and benzene, was placed in Flask #2. The flasks were closed, using one- and two-hole stoppers, respectively, and connected with a thick rubber hose, capable of withstanding chlorine atmosphere. The glass tubing extended beneath the surface of the organic mixture in Flask #2. Flask #2 was also equipped with an outlet hose with a clamp, which was used to equalize pressures during the initial period of the reaction, before the system reached steady state with regard to chlorine formation.

To start the reaction, 15 mL of concentrated H_2SO_4 was added to bleach in Flask #1. The internal pressure was kept at approximately 1 atm by periodically releasing the excess of gas through the outlet hose. After 10 minutes the glass tubing in Flask #2 was raised out of the mixture to prevent it from drawing the substrate(s) back into Flask #1, which could be caused by the gradual cooling of the aqueous solution and

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a subsequent drop in pressure. The analysis of the substrates and the products was conducted using the capillary GC instrument (see 2.2.4).

2.2.3 Iodometric titration of bleach

A 0.1 N sodium thiosulfate solution was prepared (1.58 g of $Na_2S_2O_3$, diluted with distilled water to 1 L of solution and stabilized by adding 0.2 g of Na_2CO_3). The solution was left in a dark place for 5 days before standardization. The primary standard (0.1 N $K_2Cr_2O_7$) was prepared in a 100 mL volumetric flask (0.4904 g per 100 mL of aqueous solution).

Standardization of sodium thiosulfate

1 g of KI was dissolved in 20 mL of distilled water; 2 mL of H_2SO_4(conc.), 6 mL of HCl(conc.) and 10 mL of the primary standard were consecutively added to the solution. The flask was covered with a watch glass and the mixture was allowed to react for 5 minutes in the dark. Then the watch glass was removed, the drops of iodine were washed into the flask with distilled water, and the mixture was titrated with sodium thiosulfate to the pale greenish-yellow color. At that moment 4 drops of starch indicator solution were added and the solution was titrated further till it changed color from blue to light green. The procedure was repeated three times and the average value was used in the calculations of the concentration of sodium thiosulfate.

Titration of bleach

The titration procedure was practically identical to the one described above. 1 mL of bleach was diluted to 10 mL with distilled water and titrated to the pale yellow color. After the addition of the indicator the solution was titrated until it became colorless. The procedure was repeated three times and the average value was used [23].

2.2.4 GC Analysis

A capillary Shimadzu GC-14A was employed (see Section 2.1.1). The volume of aliquots in GC injections was maintained at 1 μL. The following isothermal operating conditions were established prior to conducting kinetic studies and were maintained throughout this project:

Column temperature: 100°C

Injector temperature: 200°C

Detector temperature: 300°C

The integrator regime used most frequently was:

Attenuation: 4

Area rejection: 10,000

The value for area rejection was decreased to 1,000 and below in an effort to detect traces of products in cases of unreactive substrates.

2.3 Data Analysis

The kinetic data obtained in the course of the experiments were computer-analyzed using Quattro Pro 3 software. The selectivity data were analyzed using a calculator.

The order of the reaction in the organic substrate was established through a reaction with an excess of other reactants. Integrated equations were used (see Discussion for derivations of some of them). The data were displayed in several different systems of coordinates; a straight line corresponded to the correctly chosen order. The order in inorganic reactants was determined by varying their concentrations while holding the concentration of the organic substrate approximately constant and comparing the rate constants for different runs.

All substrate concentrations used were in fact relative concentrations. A unit of relative concentration equals 0.0400 M for bromobenzene and 0.0343 M for anisole and toluene (the amount of internal standards was held at 1.00 mmol per 25.0 mL; the deviation is caused by the fact, that toluene and anisole have 7 carbons instead of 6, which makes a difference for the carbon counter).

III. KINETIC STUDIES

3.1 Bromobenzene

3.1.1 Reaction of bromobenzene at pH 9.0

The reaction of bromobenzene with 100% bleach (see 2.2.1.1) was conducted after the pH of the aqueous layer was adjusted to 9.0. The data from the chromatograms gave a linear graph in 1/[s] vs. t^3 coordinates (Fig.2) which characterizes a second order reaction:

$$(3.1) \qquad\qquad Rate = k * [s]^2$$

3.1.2 Reaction of bromobenzene at pH 6.5 and lower.

The reaction of bromobenzene with 100% bleach (see 2.2.1.1) was conducted after the pH of the aqueous layer was adjusted to 6.5. The linearity of the 1/[s] vs. t graph (Fig.3) suggests the reaction is second order in substrate

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[3] From here on the following abbreviations will be used:
[s]- relative substrate concentration, calculated as:

$$[s] = S_i/I_i,$$

where S_i is the area of the substrate peak, I_i- the area of the internal standard peak for the chromatogram #i;
t- time of the experiment;
$\tau_{1/2}$- halflife of the substrate
p/o- ratio of para- and ortho- isomers in the product mixture.

Figure 2 Bromobenzene chlorination
100% bleach/0.32 g PTC pH 9.0
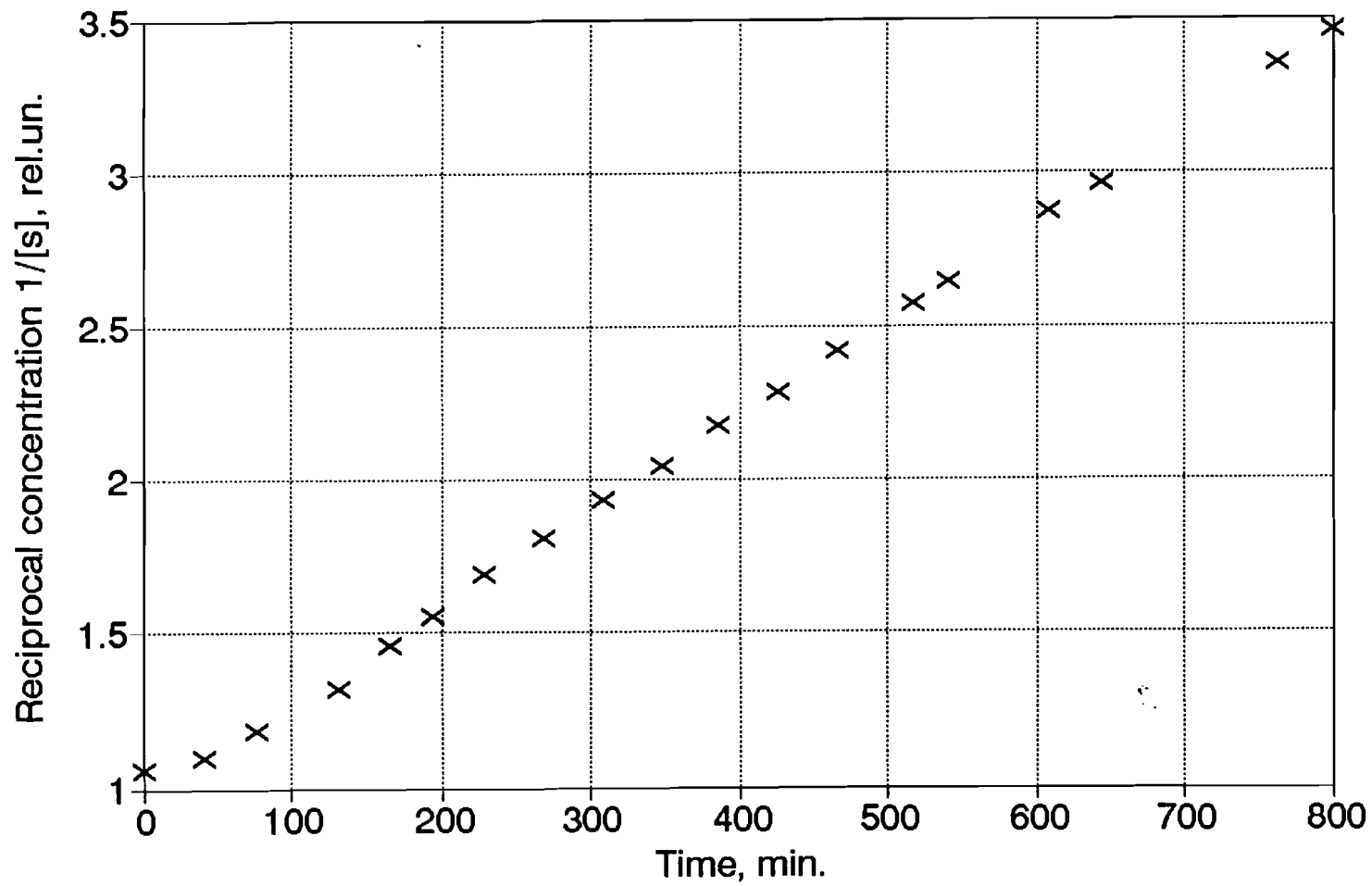
Figure 3 Bromobenzene chlorination
100% bleach/0.32 g PTC pH 6.5
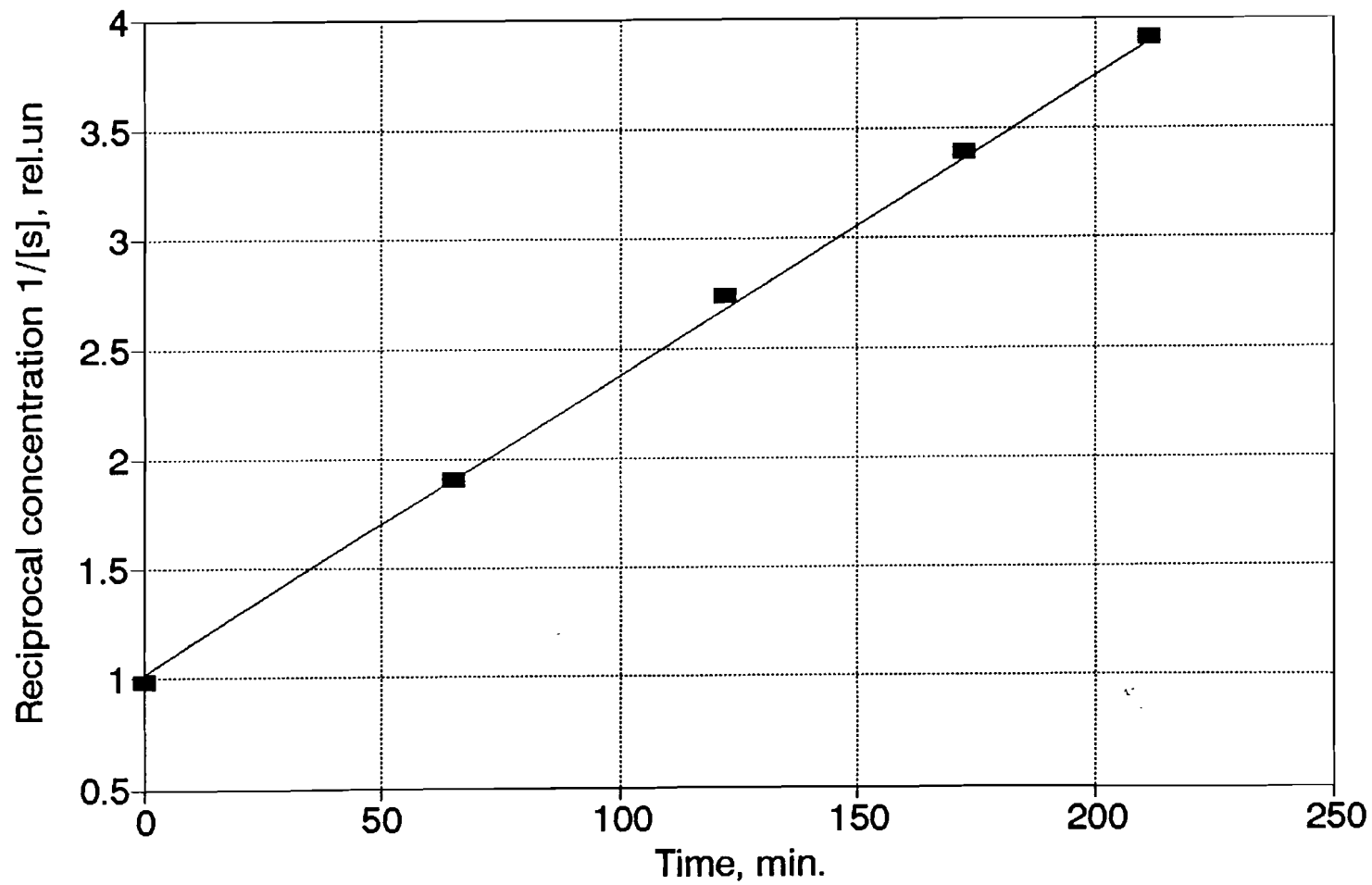
exactly like (3.1). However, the value of the rate constant is significantly ($\approx$ 4.40 times) greater at the lower pH.

The same reaction with 50% bleach, conducted at pH 3.0, proved to be also second-order in substrate, but the reaction was much faster. The halflife for this reaction was $\approx$ 32 min., while $\tau_{\frac{1}{2}}$ at pH 6.0-5.0 never fell below 50 min.(the amounts of substrate used were approximately equal).

An attempt to inhibit the reaction at pH 6.0 through trapping reactive radicals by adding p-benzoquinone failed when the peak of the inhibitor disappeared from the chromatogram in 10 min. (probably, it reacted with chlorine to produce isomer products with high boiling points).

3.1.3 Investigation of the relationship between the reaction rate and the concentration of bleach in the chlorination of bromobenzene.

The reaction of bromobenzene with different concentrations of bleach (2.2.1.1) was conducted. The values of reaction rate constants were obtained from regression analysis of 1/[s] vs. t. The average results are represented in Table II. They demonstrate that the rate of chlorination decreases upon the dilution of bleach (two-fold dilution leads to a decrease in the rate by approximately a factor of two for 50% to 100% bleach. At lower concentrations the rate becomes extremely slow and the experimental error in the rate appro-

aches the value of the rate. Thus, the lack of reproducibility in the bleach dilution trials prevents one from making any assumptions based on the collected data. Besides, the system is too complicated for a straightforward analysis. The changes in the bleach concentration affect not only the concentration of chlorine in the aqueous layer, but also, e.g. the distribution of PTC between the layers, etc. This will be examined in greater detail in section 4.2.

3.1.4 Investigation of the effect of NaCl on the reaction rate in bromobenzene chlorination.

3.1.4.1 Bleach dilutions and additions of NaCl

The two-phase reaction of bromobenzene (see 2.2.1.1) was conducted with the addition of various amounts of NaCl to 50% bleach. The second-order rate constants for these reactions are displayed in Table II. After that the rate constants for the reactions with 50% bleach and 0 g, 0.50 g, 1.00 g, and 2.00 g (## 1, 2, 3, and 5) of NaCl were analyzed. The values obtained fit a first-order plot (k vs. c_{NaCl}) but appeared to give a Y-intercept of 1.51 g[4]. Since the mixture that was used in those series of experiments contained 100 mL of bleach (200 mL of 50% bleach was used), the following explanation was

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[4] The concentration of NaCl was expressed by the mass of NaCl in 200 ml of the aqueous phase throughout the experiment in order to simplify the calculations.

Table II

Rate constants for chlorination of bromobenzene

#	Aqueous phase	Rate constant, $k*10^3$ unitless	Rate constant, $k*10^3$ L/(mol*min)
1	50% bleach	6.97	174
2	50% bleach + 0.50 g NaCl	8.00	200
3	50% bleach + 1.00 g NaCl	10.0	250
4	100% bleach	13.6	340
5	100% bleach + 1.45 g NaNO3	13.8	346
6	50% bleach + 2.00 g NaCl	14.4	361
7	100% bleach + 1.00 g NaCl	18.3	457

proposed. The key step in the industrial production of bleach includes bubbling chlorine gas through aqueous NaOH. The resulting concentrations of NaCl and NaOCl should be equal. However, in the process of storage, transportation, etc. reactions occur that alter the composition of bleach. The two most notable of them are (1) hypochlorite disproportionation and (2) hypochlorite decomposition. Reaction (1) is very slow at the pH of bleach:

(1) $3 \ ClO^- \rightleftharpoons ClO_3^- + 2 \ Cl^-$,

besides, it doesn't alter the normality of the oxidizing agent. Reaction (2) drops the normality of the oxidizing agent below the normality of chloride ion:

(2) $ClO^- \rightleftharpoons Cl^- + 1/2 \ O_2$,
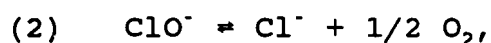
and it might be causing the regression line to move up. For more detailed examination of this phenomenon see 4.2. Based on the value of the Y-intercept, the rate constants were plotted against the corrected concentrations of NaCl:

$$c_{corr} = c_{exp} + 1.51$$

The linearity of the plot (Fig.4) supports the assumptions made. The reaction is first order in chloride.

3.1.4.2 Addition of $NaNO_3$

Since Fonouni et al. [14] mentioned both sodium chloride and sodium nitrate as rate accelerators for the chlorinations with bleach, it was decided to conduct comparative $NaCl/NaNO_3$

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Figure 4 Determination of NaCl excess through extrapolation
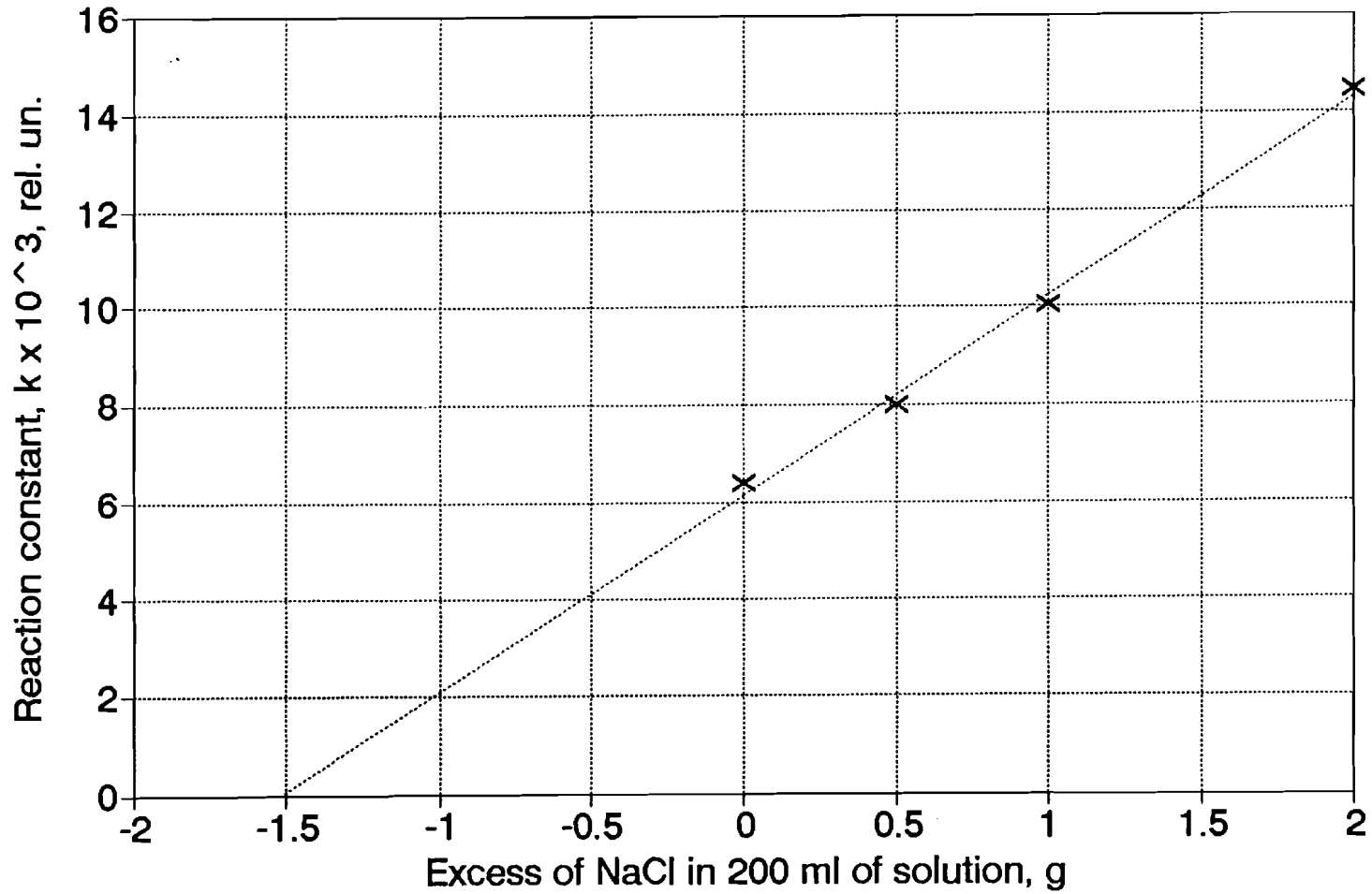
studies that seemed to be crucial in determining whether the "catalysis"[5] was caused by the "salting out" effect or Cl^- was a part of the chlorination mechanism. 100% and 50% bleach were used in the reactions (2.2.1.1); equimolar amounts of NaCl (1.00 g) and $NaNO_3$ (1.45 g) were added into two different vessels in both cases. The reactions were both second-order in the substrate and the corresponding rate constants are included in Table II. The results of this series of experiments are evidence of the basic difference in the nature of the effects caused by the two salts (to be discussed in 4.2).

As a result of these findings, a conclusion can be made that the value of 1.51 g, which was obtained previously, is in fact slightly higher than the true value due to the "salting out" effect of NaCl.

3.1.4.3 Iodometric titration of bleach

In order to determine the effect of chlorine formation the titration of bleach was carried out. Iodometric titration seemed to be the preferable method (sodium hypochlorite is an oxidizing agent, hence titration with permanganate would require the addition of an excess of a standardized reducing

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[5] The word "catalysis" in quotation marks signifies the fact of rate acceleration upon the addition of the mentioned salts without, however, stating any details of this effect or its catalytic nature. The nature and the mechanism proposed by Fonouni et al. are being questioned and the present findings are being discussed in 4.2 and 4.5.

agent prior to the titration). Besides, most sources referred to iodometric titration as a method of choice in chlorine chemistry [23,24]. The titration was conducted as described in 2.2.3. The calculations led to the value of $m_{NaOCl}(exp) =$ 4.99 g for 100 mL of bleach. Using the approximate value of 5.25% for the hypochlorite content in bleach and the measured value of 1.01 g/mL for the density of bleach, the equimolar amounts of NaCl and NaOCl in 100 mL of "ideal" bleach were established:

$m_{NaOCl}(id) = 5.30$ g; $m_{NaCl}(id) = 4.16$ g; $\Sigma m = 9.46$ g.

However, the value of 9.46 g wasn't used in further calculations. Instead a more accurate value was obtained by evaporating water from a bleach sample of known volume and weighing the salt: 9.98 g.

Next we calculate the m_{NaCl}:

$m_{NaCl} = 9.98 - m_{NaOCl} = 4.99$ g.

The excess of NaCl was then calculated as:

(3.2) $m_{NaCl}(excess) = [m_{NaCl}/M_{NaCl} - m_{NaOCl}/M_{NaOCl}]*M_{NaCl}$

The obtained value of $m_{NaCl}(excess) = 1.07$ g is compared to the value deduced kinetically in 3.1.4.1.

3.1.5 Single-phase bromobenzene chlorination

In order to study single-phase bromobenzene chlorination a special apparatus was assembled (see 2.2.2). No quantitative kinetic studies were performed, the main reason for that

being the problems with controlling constant pressure in the vessel throughout the experiment (for the purpose of accurate quantitative studies a chlorine tank would serve much better than a Cl_2 generator). However, the present study confirmed the presence of a significant induction period, as reported by Echols et al. [5]. Interestingly enough, the rapid increase in reaction rate after the induction period coincided with the reaction mixture turning red due to the formation of bromine, though no data was obtained to support any other consequential connection between the two events other than the obvious increase in $d[Br_2]/dt$ by the end of the induction period. In [5] the reverse connection is mentioned; it is discussed in 4.5.

Two more facts need to be mentioned about this experiment. Firstly, the introduction of benzene into the organic mixture practically stopped the reaction. In 24 hrs. a 100% conversion PhBr→PhCl was observed for the $PhBr/CH_2Cl_2$ mixture, but it dropped to 0.606% in the presence of benzene. The probable cause of this effect, the π-complex formation, is discussed in 4.1. Secondly, the introduction of galvinoxyl (a stable radical) to the otherwise inert PhBr, C_6H_6/CH_2Cl_2 mixture speeded up the reaction ($\approx$ 22.2 times).

3.1.6 Analysis of the isomeric content of the product mixture

The chlorination of bromobenzene by bleach (2.2.1.2)

gives a single dominant product - chlorobenzene. However, further chlorination of the deactivated aromatic ring gives small quantities of two isomers: o-dichlorobenzene and p-dichlorobenzene (the average yield in 24 hrs.is 0.93% with a Cl_2 generator and 0.37%[6] in a two-phase system with bleach at pH 5.0). The para/ortho ratio, or p/o was calculated to compare the two systems. The following results were obtained:

p/o = 1.35, when a Cl_2 generator was used;

p/o = 2.12, when a two-phase system was used.

The discrepancy in the value of the p/o ratio is evidence of the difference in the selectivities of the attacking species in the two cases.

3.2 Anisole

3.2.1 Reaction of anisole at pH 7.0 and lower

The reaction of anisole (2.2.1.1) at relatively low pH's proved not to be useful from the kinetic point of view. The substrate is so reactive that at pH 6.5 the 100% conversion is observed in less than a minute even with relatively dilute bleach phases (30% and lower). In this first minute the <u>only</u> product was p-chloroanisole. However, the reaction proceeds

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[6] The presented values were obtained through a correction factor, since the actual time of the experiment was 23 hrs. and 19.5 hrs., respectively. Because the rates of both reactions were extremely low and [s]≈const throughout the experiment, the (24/t) was accepted as a correction factor (t- time of the experiment).

further to yield a number of unidentified polychloroaromatics.

3.2.2   Reaction of anisole at pH 8.5

The reaction (2.2.1.1) at pH 8.5 was being used almost constantly in studying anisole chlorination.   The time for full conversion was generally around 23-25 min.   Different conditions were tried by changing the bleach concentration, the amount of PTC added to the system, and the solvent.   The data are presented in graphical form in Figures 5-9.   In almost all the cases the reaction order was hard to identify. It seemed to be a mixed "0+1" order, however, the exact relationship is yet to be determined.   Only three reactions gave results different from the ones described.   The reaction without PTC was first-order (Fig.8);   the reaction with a fivefold excess of PTC in comparison with the usual amount of 1 mmol gave a zero-order plot (Fig.9).   Finally, the reaction in CCl_4 gave only traces of the products after 1 hr.

The arguably most important feature of anisole chlorinations, the sliding downward p/o ratio of the isomers, is examined in section 3.2.6.
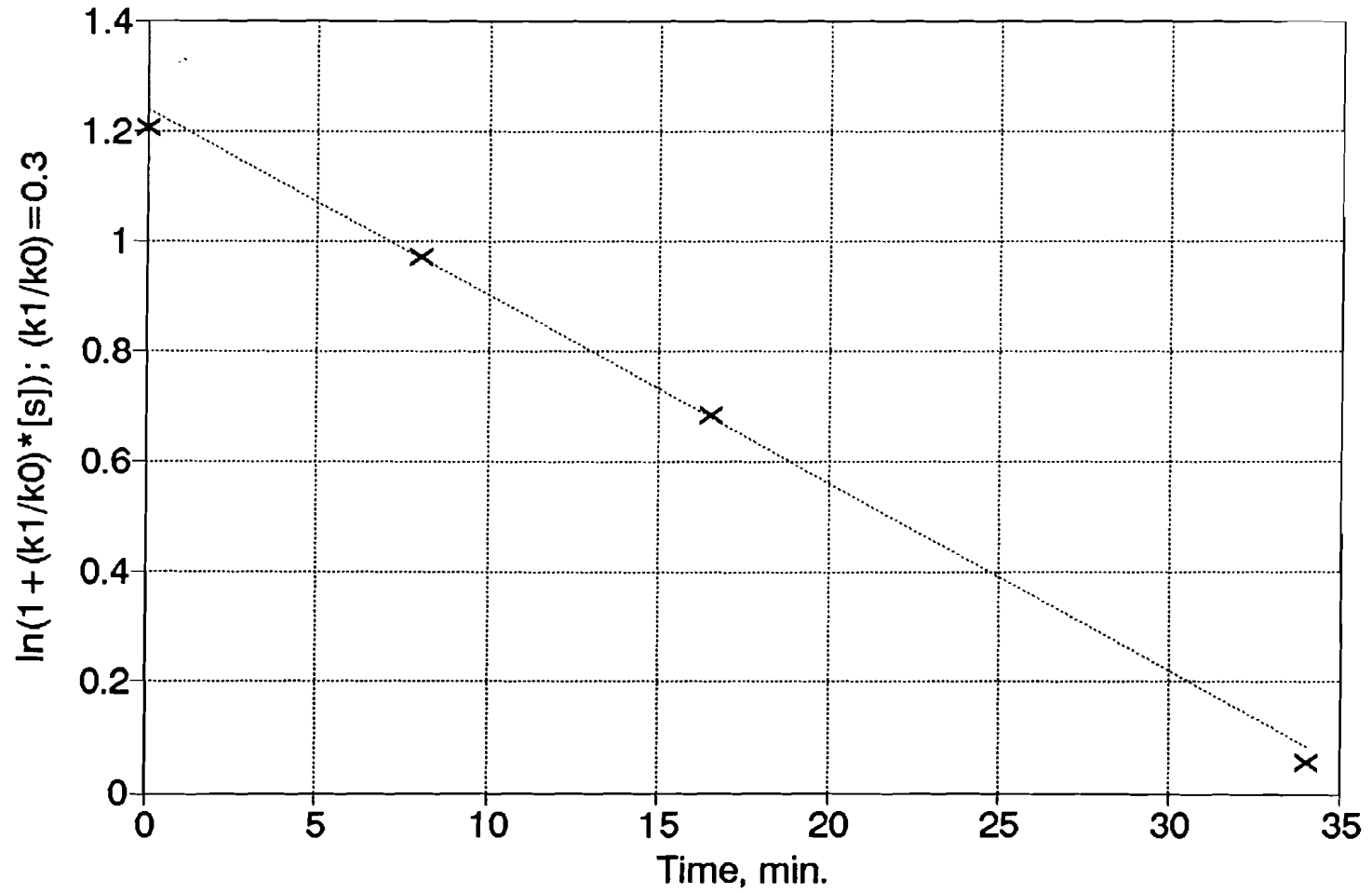
Figure 5 Anisole chlorination
100% bleach/0.32 g PTC pH 8.5

Figure 6 Anisole chlorination
50% bleach/0.32 g PTC pH 8.5
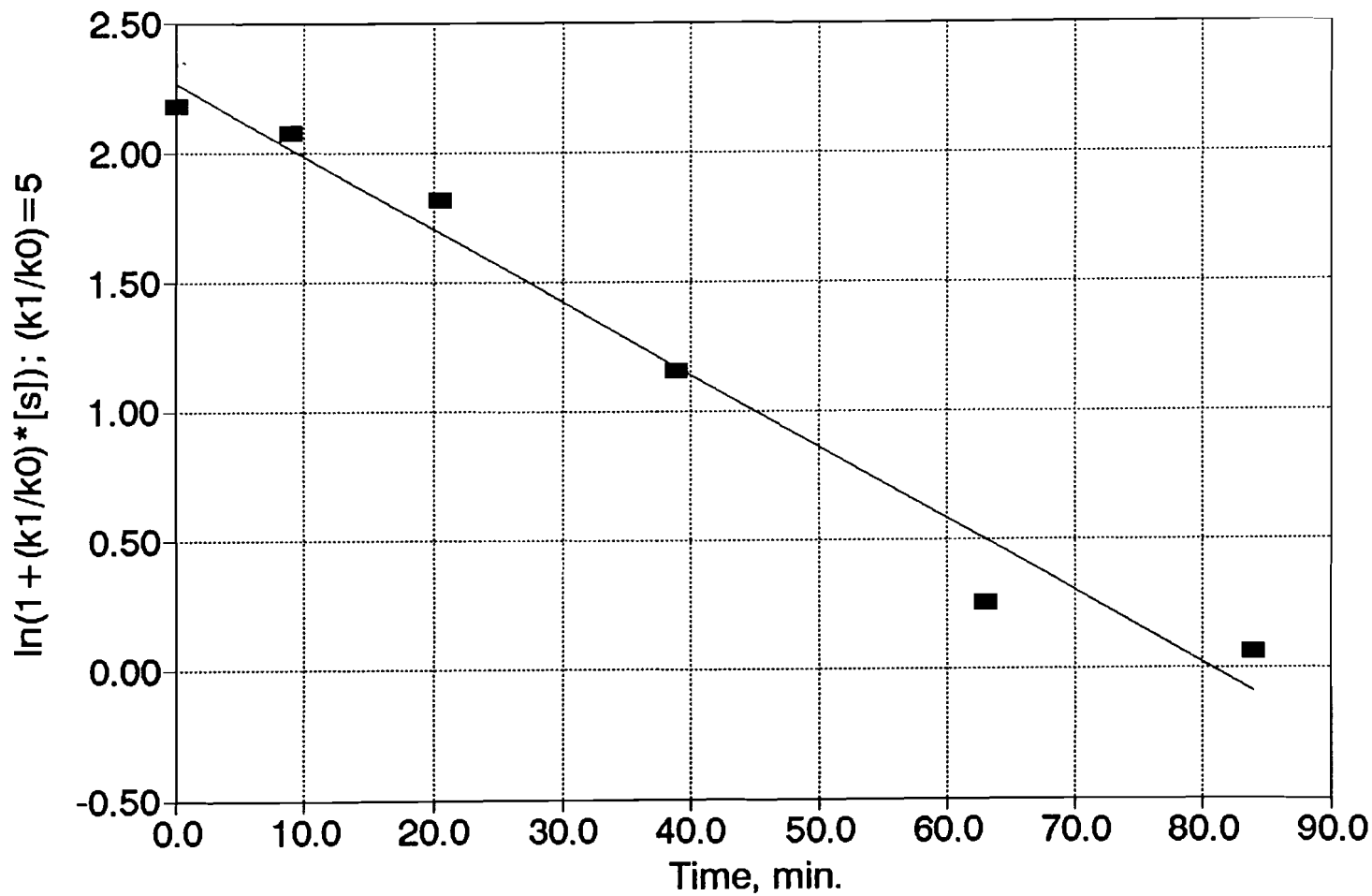
33

Figure 7 Anisole chlorination
50% bleach/0.16 g PTC pH 8.5

y-axis: ln(1+(k1/k0)*[s]); (k1/k0)=10
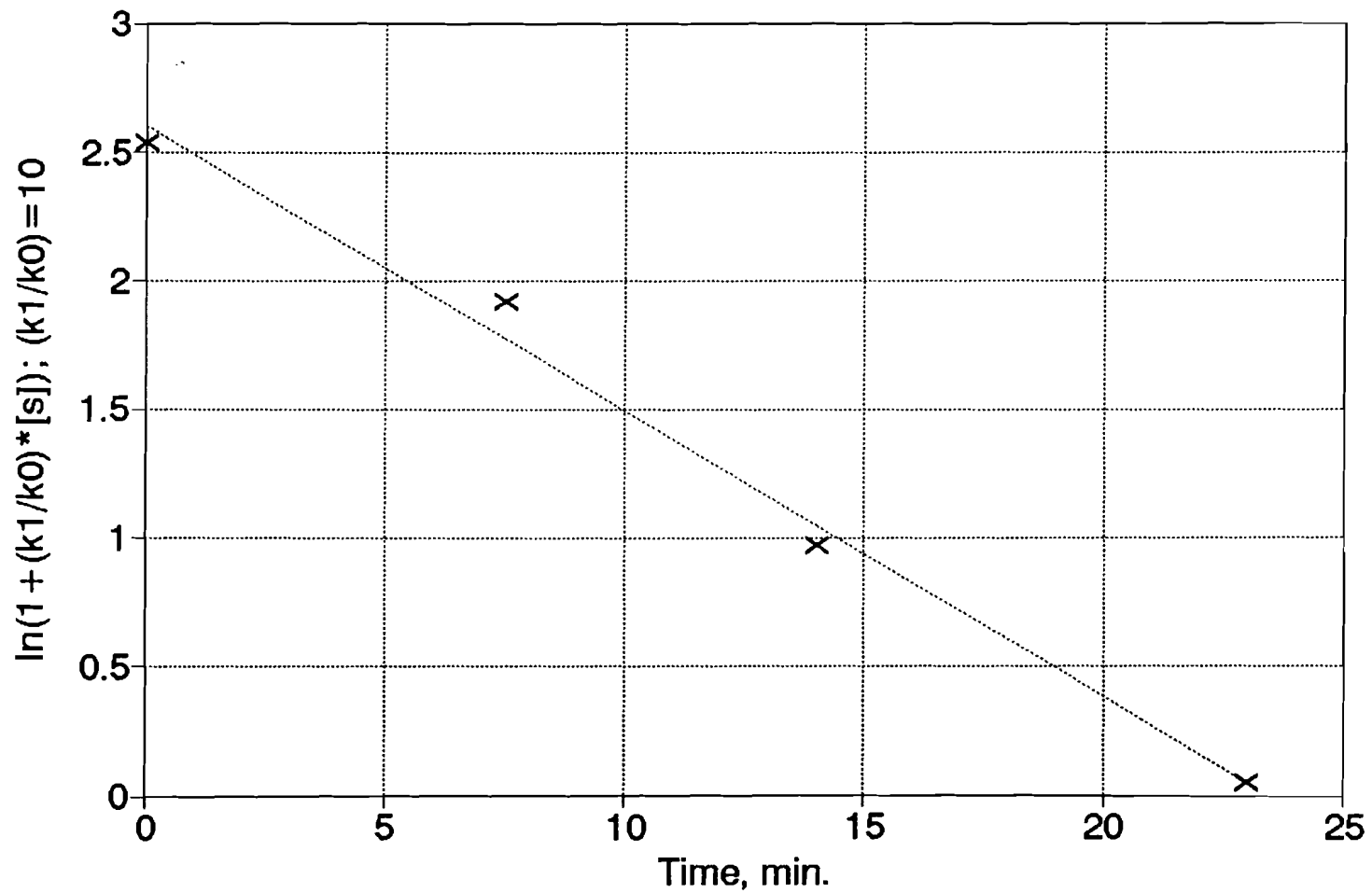
x-axis: Time, min.

Figure 8 Anisole chlorination
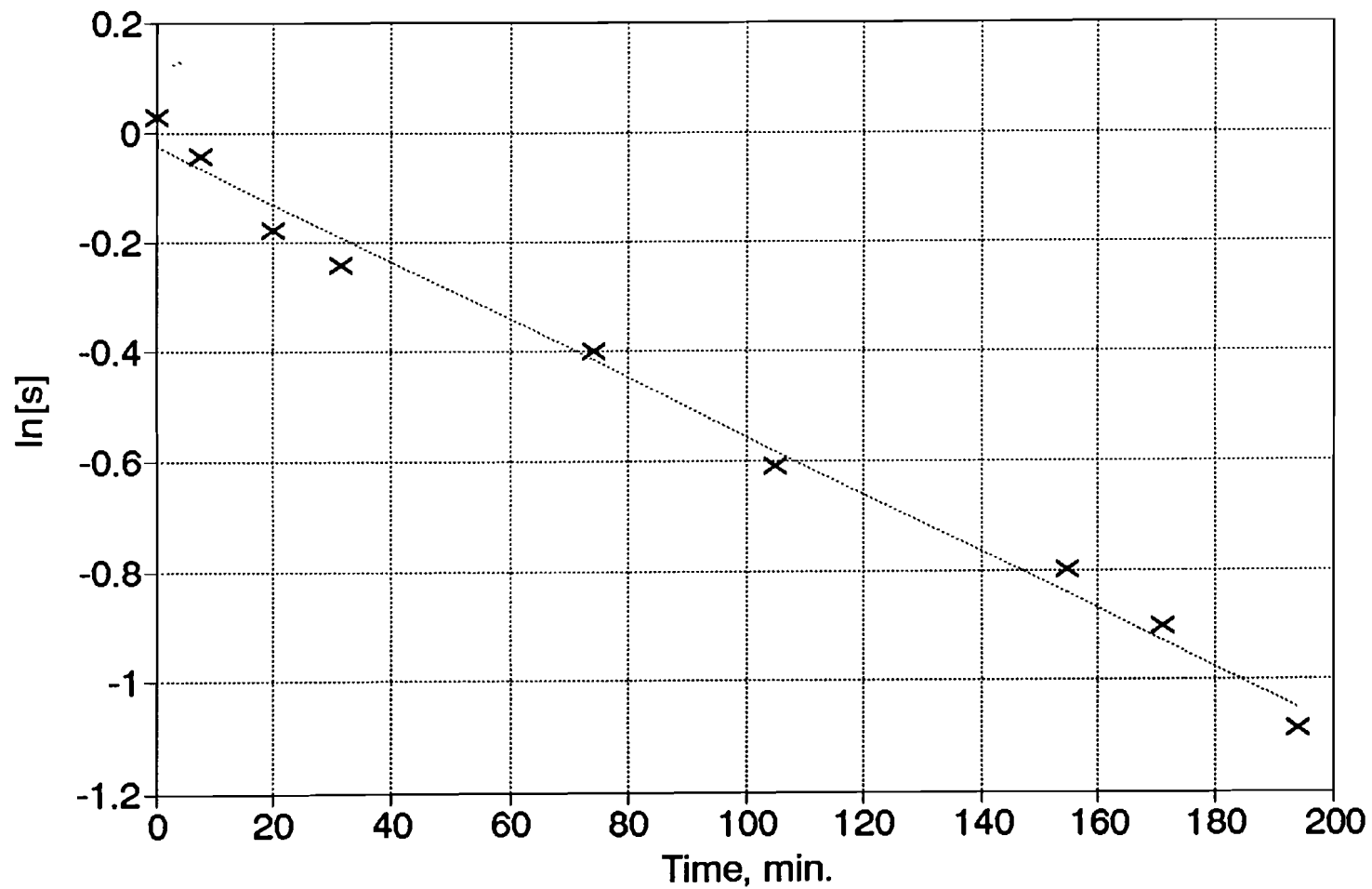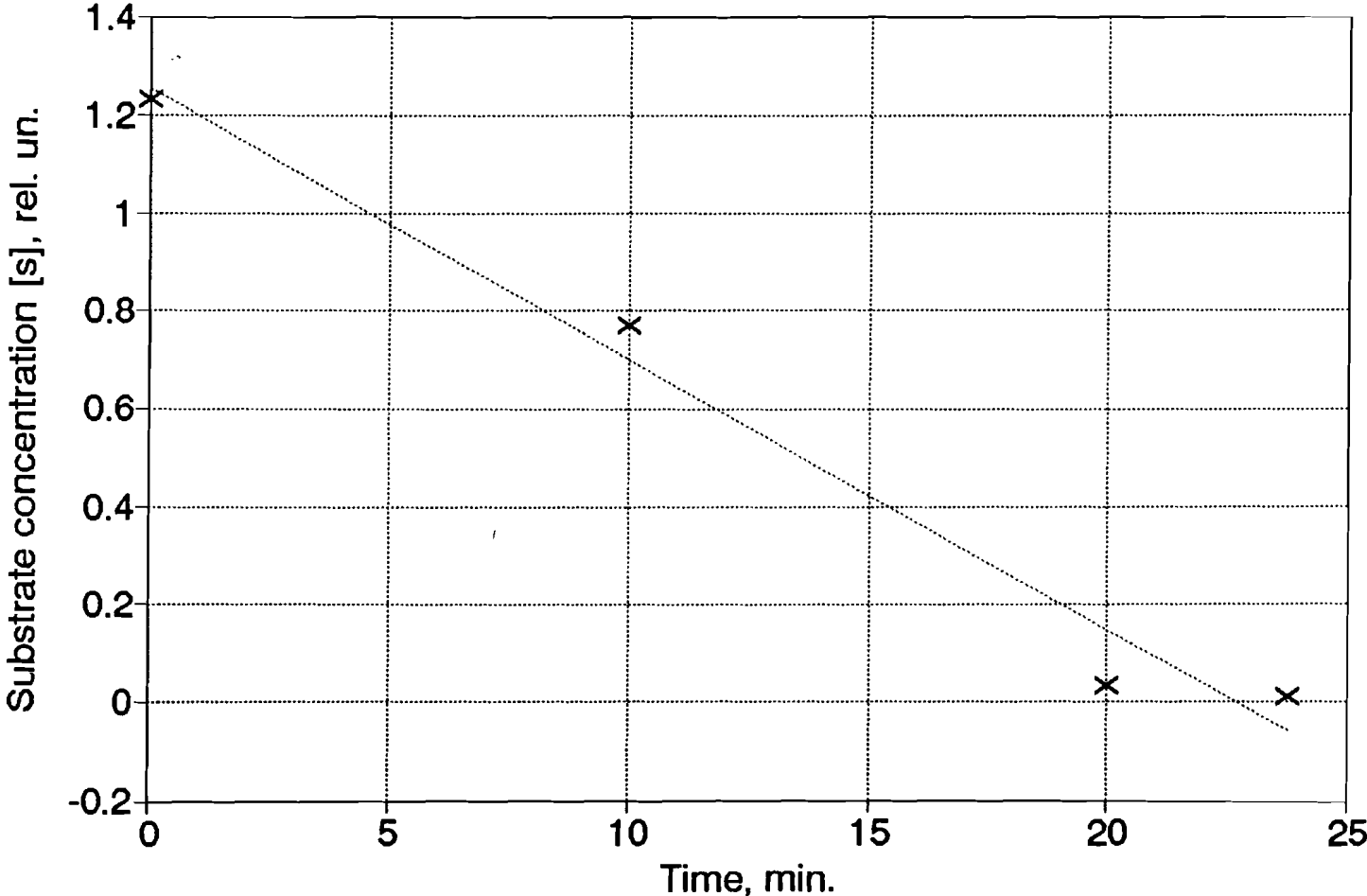50% bleach/no PTC pH 8.5

Figure 9 Anisole chlorination
50% bleach/1.60 g PTC pH 8.5

3.2.3 Investigation of the relationship between the reaction rate and the concentration of bleach in the chlorination of anisole.

The studies of the order of reaction in bleach for anisole chlorination were conducted as short runs that employed approximately the same concentration of anisole and various bleach concentrations. Two GC injections were made for every run (at the start and 4 min. after the start of each run). Since there was no definite order established for anisole, the concentrations of anisole were taken as close as possible (±2%) and then the reaction rates were compared: Rate(1)/Rate(2) = $k_1[s_1]^n/k_2[s_2]^n \approx k_1/k_2$ for $[s_1]\approx[s_2]$, where n is the reaction order in anisole, $n \in [0;1]$. The results show that a twofold dilution decreases the rate by $\approx$ 2.12. However, as it was in the case of PhBr, no plausible explanation for these results has been found due to the complexity of the system.

3.2.4 Additions of NaCl in anisole chlorination

Additions of sodium chloride to the reacting mixture anisole and bleach at pH 8.5 increased the rate considerably (in the 1.4-2.0 range). However, no reaction order was established for NaCl due to the fact that the mixed order in substrate brings in at least two rate constants (instead of

one in case of PhBr), which means that any effort to establish the order in chloride has to deal with the change in both constants at the same time.

3.2.5 Single-phase anisole chlorination

The chlorination of anisole using a Cl_2 generator proceeded faster than that of PhBr under the same conditions. The induction period was still present but it became considerably shorter due to the higher reactivity of anisole.

3.2.6 Analysis of the isomeric content of the product mixture

The arguably most intriguing feature of anisole chlorination was the downward sliding p/o ratio. At the same time no increase in the amounts of other byproducts was observed. This fact led to the conclusion about the presence of two competitive mechanisms differing from each other both in order and in selectivity (see 4.3 for the reasons behind the conclusion). Further observations suggested that the changes in the conditions of the experiment, such as the bleach concentration, the pH, the solvent, and the amount of PTC in the solution produced a shift of the range, in which p/o was sliding, up or down the scale. The increase in bleach concentration helped to bring the range down (from 22.2 → 17.0 with 50% bleach to 16.3 → 14.4 with 100% bleach at pH 8.5).

The drop in pH decreased the range dramatically (to 11.5 →
11.1 with 100% bleach at pH 5.8). The change in solvent (use
of chloroform instead of methylene chloride) pushed the range
down slightly (to 14.7 → 13.9 with 50% bleach at pH 8.5). The
amount of PTC used in the run appeared to be extremely
important. Table III provides the values of the limits of p/o
ratio observed in anisole chlorinations with different amounts
of PTC at pH 8.5 with 50% bleach. In the reactions without
PTC or with its excess the p/o ratio is approximately constant
or slightly rises, most probably due to the fact that o-isomer
is somewhat more reactive than p-isomer in subsequent polychl-
orination.

A decrease in pH influenced the reaction so significantly
that a run with 100% bleach at pH 5.0 provided p/o = 5.15 even
in the absence of PTC (a large excess of anisole with an
addition of a small amount of the solvent was utilized). It
is extremely close to the value of p/o = 4.66 for the chlori-
nation with a Cl_2 generator in a single layer (excess of
anisole used). The comparison of these values suggests that
probably the same attacking species is present in both cases.
It was also discovered that the amount of the solvent used
also influences the p/o range (in both cases the increase in
the amount of the solvent used makes the chlorination more
selective, increasing p/o ratio). A mechanism consistent with
this data will be proposed and the p/o ratio will be calculat-
ed theoretically. Then it will be compared with the experi-

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mental curves (p/o vs. % conversion of the substrate) in Figures 14-16 (see 4.3).

3.3 Toluene

The reaction with toluene (2.2.1.1) at pH 8.5 with 100% bleach gave a mixed "0+1" order so characteristic for anisole. However, the curve for toluene, as opposed to the anisole curves, fitted the order of 0.5 (Fig.10)! Benzyl chloride was the dominant product of the chlorination reaction.

3.4 Solvent

3.4.1 Chlorination of the solvent

The most notable thing about using methylene chloride or chloroform as a solvent was the chlorination of the solvent. E.g., in case of methylene chloride significant quantities of chloroform and carbon tetrachloride are being formed in the process of the chlorination of the substrate.

An interesting detail is that the yield of the chlorinated solvent was dependent upon the reactivity of the substrate. The reaction with bromobenzene gives low yields of $CHCl_3$ even at pH 5.0 ($\approx$ 0.85% in 30 min.). At pH 9.0 the yield dropped to 0.36%. At the same time chlorination of anisole was producing the yields of 2.60% in 30 min. at pH 8.5. This

observation will be critical later in 4.1, supporting the idea of the trapping of radicals by unreactive aromatic molecules through π-complexation.

3.4.2 Kinetic study

A kinetic study of CH_2Cl_2 in the presence of toluene gave an expected pseudo-zero order (the % conversion wasn't high enough to influence the rate of chlorination)(see Fig.11).

3.5 Chlorination of aliphatic compounds

The chlorination of iso-butyl alcohol (72 hrs.) gave a mixture of chlorination products. This observation suggests that the radicals present in the solution are reactive enough to chlorinate saturated hydrocarbon chains, where any mesomeric effects are absent.

Table III

Limits for p/o ratio in anisole chlorinations with

50% bleach at pH 8.5

Mass of PTC	p/o ratio	
added	Upper limit	Lower limit
0.000	23.21	23.64
0.160	23.67	23.61
0.320	22.20	17.00
1.600	12.40	13.55

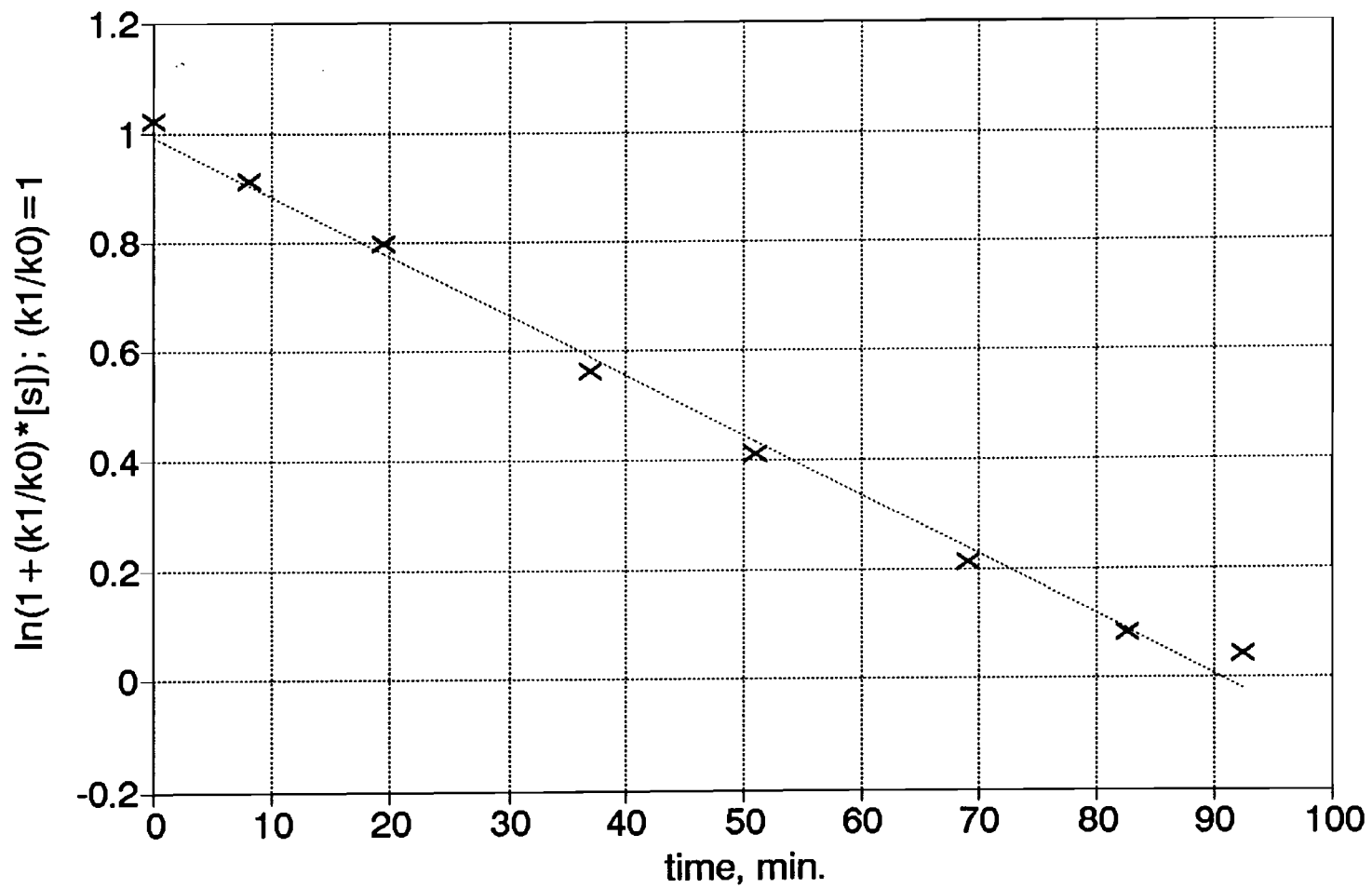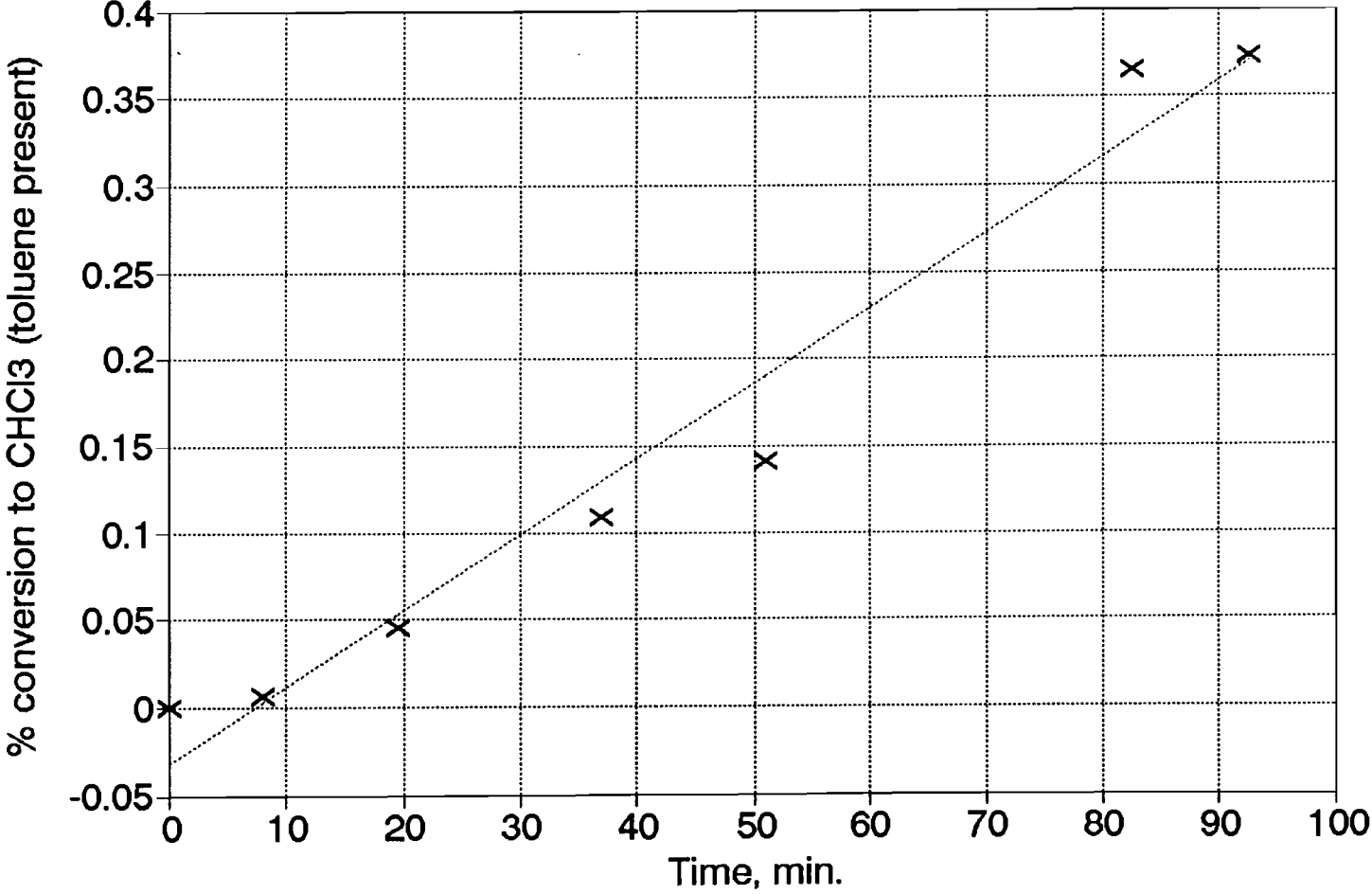Figure 10 Toluene chlorination
100% bleach/0.32 g PTC pH 8.5

Figure 11 Solvent chlorination
100% bleach/0.32 g PTC pH 8.5

IV. DISCUSSION

4.1 Free-radical nature of the mechanism

Since 1957, when Miller and Walling published the results of their investigation [4], the reaction of aromatic chlorination of the type that the current study is devoted to has been viewed as free-radical. Still almost every paper in the field traditionally starts with facts that help to classify the reaction as a case of free-radical aromatic substitution. This investigation follows this tradition and provides proof for the free-radical nature of the mechanism prior to attempting to postulate the mechanism itself based on the experimental data.

(1) The reaction is slightly accelerated in illuminated conditions [14].

(2) When air isn't evacuated from the vessel (as has been the case in the present study), no appreciable reaction slowdown is observed, but the amount of the products coming from the disintegration of the peroxy-radicals (like benzaldehyde and benzyl alcohol in the chlorination of toluene) increases significantly [14].

(3) The chlorination of aromatic rings with two or more substituents leads to the products, the formation of which can be explained only with group migration in the σ-complex; this supports the hypothesis about the free-radical substitution
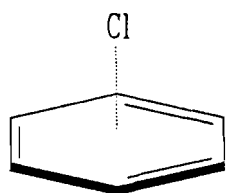
mechanism [11].

(4) The reaction gives a number of products, which can result from a free radical reaction only, such as benzyl chloride from toluene, chloroform and carbon tetrachloride from methylene chloride, etc.

(5) Electron-donating substituents (Type I) accelerate the reaction, but the influence of the electron-withdrawing groups is not quite clear.

(6) The addition of BrCl to the reacting mixture promotes the reaction [5]. Galvinoxyl, a stable free radical, was also found to speed up the chlorination (3.1.5).

(7) Benzene inhibited the single-phase bromobenzene chlorination reaction (3.1.5). It is extremely unreactive, although the π-electron cloud of the aromatic ring is known to form stable π-complexes with free radicals [25-27]. Depending upon the concentration of benzene, it can form either 'umbrellas' or 'sandwiches' or both (Fig.12a-b). Thus, the reacting species is trapped in a thermodynamically stable complex and the reaction is effectively inhibited. The fact that conversion of methylene chloride to chloroform was significantly higher in the process of anisole chlorination rather than in case of PhBr (3.4.1) supports this conclusion.

(a) "Umbrella"

(b) "Sandwich"

(Bu = butyl)
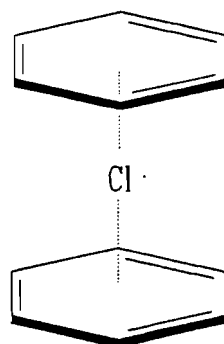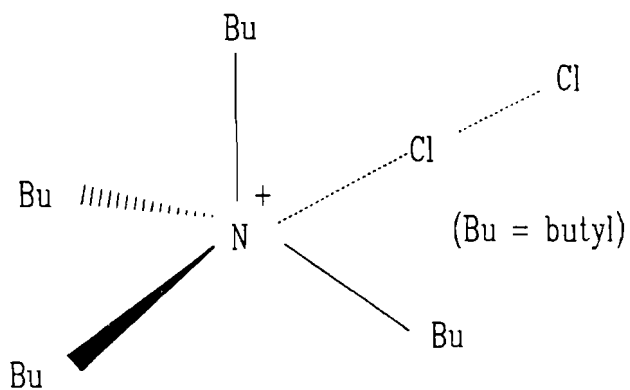
(c) PTC–Cl$_2$ complex

(b) Ar–Cl$_2$ complex

Figure 12   Structures of complex intermediates

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4.2 Reaction orders and the approach to postulating a common mechanism

A. Bromobenzene
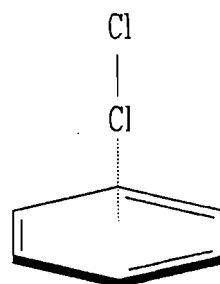
The collected data (3.1) suggests second order in PhBr for all the runs. This result comes as a surprise, since the overwhelming majority of organic reactions of this type are first-order in substrate, but it narrows the number of possibilities to a minimum.

As was mentioned in 3.1.3, the trials with diluted bleach may not be regarded as a usual study of the order in the reactant present in an excess. The obvious relationship:

(4.1) $K_{eq} = [Cl_2(aq)]/[NaCl][NaOCl]$

is not the only criteria governing the reaction rate. There are several other relationships influencing the rate, e.g. the fact that dilution decreases the ionic strength of the aqueous layer, thus changing the distribution of PTC between the layers (lower ionic strength decreases the magnitude of the 'salting out' effect, thus increasing the concentration of PTC in water), the relationship between the ionic strength and the solubility of chlorine in water, and others, some of them extremely hard to foresee. These complications prevented use of these data in deducing the reaction mechanism.

In order to get an insight into the role of chlorine (or, for that matter, any other inorganic reactive species present in the mixture) the studies using NaCl were conducted. It was supposed also to test Hamilton's hypothesis of NaCl and other

inorganic salts 'salting out' PTC from water into the organic layer. Most of the complications encountered in the previous study are avoided: the ionic strength of the aqueous layer remains practically unchanged as compared to the dilution (the amount of salts is increased by a maximum of $\approx 20\%$ or less; the volume can be considered constant, assuming that:

$(\partial V_{sol'n}/\partial n_{NaCl})_{p,t,n}$ is small and $\Delta n_{NaCl} \to 0$).

The study revealed the reaction is first order in chloride given that the aqueous phase (200 ml of 50% bleach) contains 1.51 g of NaCl prior to the additions. At the same time the addition of an equimolar amount of $NaNO_3$ increased the rate constant by 1.8%. This evidence confirms the basic difference in the influence of the two salts on the reaction rate, thus refuting Hamilton's hypothesis. Sodium nitrate increases the ionic strength of the solution and increases the ratio PTC(org)/PTC(aq) solely through the 'salting out' effect, while sodium chloride is present in the rate law for this reaction. If so, then we have Cl_2 as a reacting species (not Cl_2O, as proposed by Fonouni et al. in [14]), and:

(4.2)   Rate $\sim [Cl_2]^n \sim (K_{eq}*[NaOCl][NaCl])^n$.

Since from the experiment n=1,

(4.3)   Rate $\sim [Cl_2]$.

Then with an addition of NaCl: c = [NaCl](new)/[NaCl](old)

(4.4)   Rate(new) $\sim K_{eq}*[NaOCl]*c*[NaCl](old) \sim c*[Cl_2](old)$,

which is exactly what was observed (3.1.4.1). However, one more point needs to be made. Since [NaCl](old) is an equilib-

rium concentration and [NaCl](new) is not, these equations
will only be true if reestablishing a new equilibrium would
not change the values of [NaCl] and [NaOCl], i.e.

(4.5) [NaCl]>[Cl$_2$]<[NaOCl] or, in other words, K_{eq} < 1.
This means that the increase in [NaCl] will be compensated by
the same increase in [Cl$_2$] with [NaCl] and [NaOCl] practically
not altered by this self-adjustment of the system. The values
for K_{eq} at certain pH's calculated from the reduction poten-
tials data [28] are put together in Table IV. Indeed, they
are small enough for the Equations (4.2)-(4.5) to hold.

Iodometric titration put the excess of NaCl at 1.07 g per
100 ml of 100% bleach, which means that the amount of NaOCl
present is 7.5 mmol per 200 ml, and of NaCl - 25.8 mmol per
200 ml. The probable answer to the question, how with a very
low [Cl$_2$(aq)] the content of NaCl could fall from 4.99 g per
200 ml (3.1.4.3) to only 1.51 g, is that for the process:

Cl$_2$(aq) $\longrightarrow$ Cl$_2$(g) $\Delta G° = -6.90$ kJ/mol at 298 K (m(Cl$_2$aq)=1).
Kinetics for this process is very fast: each addition of acid
was followed by numerous bubbles of evolving gas, which was
allowed to escape from the vessel to prevent the build-up of
pressure. The adjustment of pH to ≈6.5 led to the establish-
ment of the equilibrium concentrations mentioned earlier.
Further, since

(4.6)   [Cl$_2$](org) = K_{eq}*[Cl$_2$](aq),
the data collected in the trials with NaCl provides additional
evidence for the first order in Cl$_2$.

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Table IV

Equilibrium constants for formation of chlorine
from hypochlorite and chloride at several pH's

pH	K(eq)*1000
6.0	41.00
6.9	5.19
7.6	1.33
8.5	0.01

B. **Anisole and toluene**

The data (3.2; 3.3) suggests mixed "0+1" order for chlorination of both anisole and toluene. Since, judging by the products, the mechanism is the same in both cases, we would assume that first order for Cl_2 in case of anisole and toluene. In fact, the data collected in the trials with NaCl is somewhat consistent with that; however, the linear correlation in case of anisole, for example, involved significantly greater deviations than in the case of PhBr.

The challenge from this point on lies in trying to integrate the second order in PhBr, the "0+1" order in anisole and toluene, and the first order in chlorine into one common mechanistic scheme. Quite obviously, these facts do not allow us to narrow the scheme to one particular mechanism, though they help to eliminate a large number of pathways from consideration. It is necessary to extract other relationships from the experimental data: a good mathematical analogy would be a need for N linear algebraic equations to solve for N independent variables in order to get unique values.

4.3 Search for additional kinetic evidence

The search for additional kinetic information started with an attempt to explain the sliding of the p/o ratio for anisole derivatives throughout the reaction. If the two isomers are being formed in two competitive, superselective

(leading to one isomer each) reactions of the same order in substrate, then we can record the rate laws as:

(4.7) $\qquad dc_1/dt = k_1 * c^n \qquad ; \qquad dc_2/dt = k_2 * c^n,$

where c_1, c_2- isomer concentrations; c- substrate concentration. From (4.7) we have:

(4.8) $\qquad dc_1/dc_2 = k_1/k_2,$

and integrating:

(4.9) $\qquad c_1/c_2 = k_1/k_2 = const,$

which excludes any possibility of a sliding ratio c_1/c_2. For two competitive reactions of the same order in substrate we will have a similar equation:

(4.10) $\qquad c_1/c_2 = (k_{11} + k_{21})/(k_{12} + k_{22}),$

where $k_{ij} = k_i * j$ (j is the selectivity of the reaction #i towards isomer #j, $j \epsilon [0;1]$).

The only way one can account for the sliding p/o ratio is to postulate at least two competitive reactions of different selectivities, e.g.:

(4.11) $\qquad c_1/c_2 = k_1 * c^m/k_2 * c^n.$

Remembering the mixed "0+1" order in anisole and toluene, we assign m=1 and n=0. Hence the rate law for anisole chlorination will be:

(4.12) $\qquad -dc/dt = k_1 * c + k_0,$

where k_1 and k_0 are the rate constants of the first- and zero-order reactions, respectively. The zero-order reaction is less selective and its term is independent of time, while the first-order reaction is more selective and its term decreases

with time. Other rate laws, that are frequently used to characterize the mixed "0+1" order, like $-dc/dt = k*c^{\frac{1}{2}}$ and $-dc/dt = kc/(1+kc)$, are inapplicable in this case, since they do not explain the described phenomenon. In order to apply the technique of establishing the rate law, that has been used previously, we would have to integrate (4.12) to a mathematical form $f(c) = kt + const$:

(4.13) $-dc/dt = k_0 + k_1 c = k_0 * [1 + (k_1/k_0)*c]$

Integrating,

(4.14) $\int dc/[1 + (k_1/k_0)*c] = -\int k_0 dt,$

from where:

(4.15) $(k_0/k_1)*\ln[1+(k_1/k_0)*c] = -k_0 t + const.$

Cancelling out k_0, we get:

(4.16) $\ln[1+(k_1/k_0)*c] = -k_1 t + const.$

The derived equation (4.16) was used in regression calculations. The value of k_1/k_0[7] was varied in order to yield the largest value of R^2. Obviously, a certain error is involved in this type of calculations, since the described approach fails to take into account the error of the experiment, however, the results of this curve-fitting proved to be extremely informative.

The k_1/k_0 ratio, similarly to p/o ratio, is a measure of selectivity of a given reaction: the former is a kinetic

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[7] The ratio of k_1/k_0 is calculated as a unitless value. It is derived from relative concentrations of the reactants and does not take into account the difference in units for the constants of two different orders.

criteria, and the latter is a yield criteria. Lower k_1/k_0
value signifies lower selectivity, and vice versa. Curve-
fitting provided an opportunity to assess the influence of
different factors on the overall selectivity of the reaction.
pH depression drastically decreased k_1/k_0, and indeed, the
observed p/o ratio fell (3.2.6). The use of a more concen-
trated bleach also decreased k_1/k_0, and again there was a
corresponding experimental evidence (3.2.6). The change of
solvent from CH_2Cl_2 to $CHCl_3$ did not in any way influence k_1
(±2.5%), but increased k_0 by ≈1.9. The amount of PTC present
in the mixture had the most significant influence on the k_1/k_0
ratio. With a fivefold increase in the amount of PTC $k_1/k_0 \to 0$,
i.e. the R^2 for the zero-order c vs. t plot became the
limiting value for $\ln[1+(k_1/k_0)*c]$ vs. t. With a two-fold
decrease in the amount of PTC the ratio increased from 5 to
10. When no PTC was used, $k_1/k_0 \to \infty$, i.e. the R^2 for the first-
order ln(c) vs. t plot became the limiting value for
$\ln[1+(k_1/k_0)*c]$ vs. t. The described relationship between
k_1/k_0 and the amount of PTC used is shown in Figure 13. Table
V contains k_1/k_0 values for different systems.

In order to test the hypothesis of two competitive
mechanisms the following calculations were carried out. The
p/o ratio for anisole was determined theoretically as a
function of per cent conversion and compared to the values
established experimentally. These calculations are discussed
below.

The constant in (4.16) can be determined given the conditions for the start of the experiment:

(4.17) $t = 0 \rightarrow c_0 = a$

Then:

(4.18) $const = \ln[1+(k_1/k_0)*a]$,

which leads to the final formula:

(4.19) $\ln\left\{[1+(k_1/k_0)*a]/[1+(k_1/k_0)*c]\right\} = k_1 t$.

Inserting initial concentration a and current concentration c, we get $k_1 t = const$. Hence $t = const/k_1$. Knowing this, the amount of the product obtained through the zero-order pathway can be calculated: $y = k_0 t = const/(k_1/k_0)$. Given that o- and p-chloroanisole are by far the dominant products, we have the amount of the product obtained through the first-order pathway as: $x = a - c - y$. The selectivities for these two processes were determined experimentally. For the zero-order pathway it was the p/o ratio of the reaction with a fivefold excess of PTC (≈ 12); for the first-order pathway it was the p/o ratio of the reaction without PTC (≈ 23). Using these selectivity values, the theoretical p/o ratio was computed as:

(4.20) $p/o = [x_p + y_p]/[x_o + y_o]$.

As a result, $p/o = f(k_1/k_0; a; c)$. The calculated p/o ratio is being compared to the experimental data in Figures 14-16. The sliding pattern of the theoretical and the experimental curves in all the cases is similar, which supports the earlier hypothesis. The shift along the vertical axis can be explained with the errors in determining the p/o ratio of the

two limiting cases. The conclusion is that with methylene chloride as a solvent and 50% bleach at pH 8.5, the window of p/o ratios is from 23 to 12; the sliding curve can be shifted up or down by changing the amount of PTC in the mixture.

Finally, the computations described provide crucial information about the true role of PTC in the chlorination reaction: it is not a phase-transfer catalyst (if it were, it would not change the k_1/k_0, and consequently the p/o ratio, but it would catalyze both pathways by transporting more chlorine into the organic phase), but simply a chemical catalyst in a zero-order reaction. Anisole is extremely reactive, and the formation of Cl· from Cl_2 without a photochemical impact becomes a rate-determining conversion:

(4.21) $Cl_2 + PTC \dashrightarrow 2\ Cl\cdot + PTC$
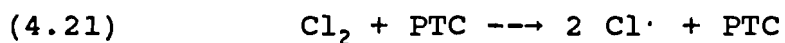
In the light of this idea, the fact that only the zero-order reaction appeared to be significantly accelerated by a change of solvent becomes much more meaningful. As was mentioned in the introduction, indifference to the change of solvents (as in case of the first-order pathway) is one of the most notable features of free-radical reactions. The significant rate acceleration observed in chloroform signifies the somewhat ionic character of the interaction. Indeed, since the nitrogen atom of PTC has a partial positive charge, the substrate-catalyst complex will most probably approximate the species described in Figure 12c. The slight difference in the polarities of methylene chloride and chloroform would then be

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beneficial for a better solvation of the complex with $CHCl_3$!
When CCl_4 was used as a solvent, none of the radical species
present in the solution were solvated, and as a result, only
traces of the products were detected.

The first-order pathway generating $Cl\cdot$ can be explained
with a following scheme:

(4.22) $Ar \dashrightarrow Ar\text{--}Cl_2 \dashrightarrow Ar\text{--}Cl\cdot + Cl\cdot$

In this scheme the arene substitutes the PTC in its role as a
catalyst in the nonphotochemical homolysis of the Cl-Cl bond.
The π-complex $Ar\text{--}Cl_2$ (Fig.12d) was proposed by Echols et al.
[5] as an intermediate. In fact, it seems plausible, that the
aromatic ring serves as an electron density donor; the
chlorine molecule accepts that density which fills the
antibonding orbitals of the Cl_2 molecule and brings its bond
order below the usual value of 1, thus weakening the bond,
increasing the distance between chlorine atoms, and aiding the
homolysis of the bond.

Altogether, the information obtained from the k_1/k_0 and
p/o ratio computations provides an insight into some of the
important details of the reaction. Now we have sufficient
data to move forward with postulating the mechanism.

TABLE V

k1/k0 values for different systems

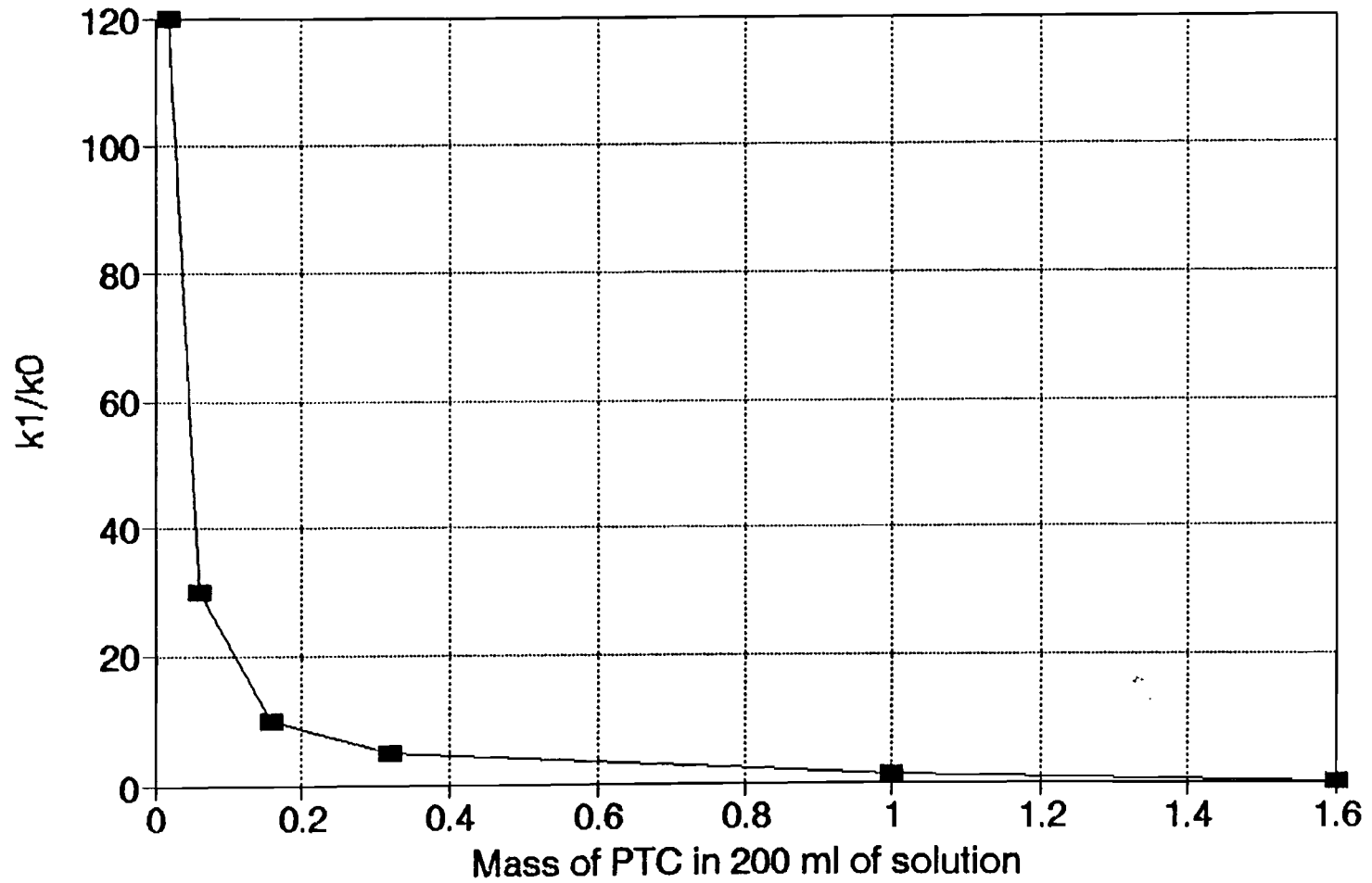| System	k1/k0
100% bleach/PTC pH 8.5	0.3
50% bleach/0.320 g PTC pH 8.5	5
50% bleach/0.180 g PTC pH 8.5	10
50% bleach/1.800 g PTC pH 8.5	< 0.001
50% bleach/no PTC pH 8.5	>200
50%bleach/PTC pH 8.5 (chloroform as a solvent)	2.7

Figure 13 Relationship between
the addition of PTC and k1/k0

Figure 14 Para/ortho ratio
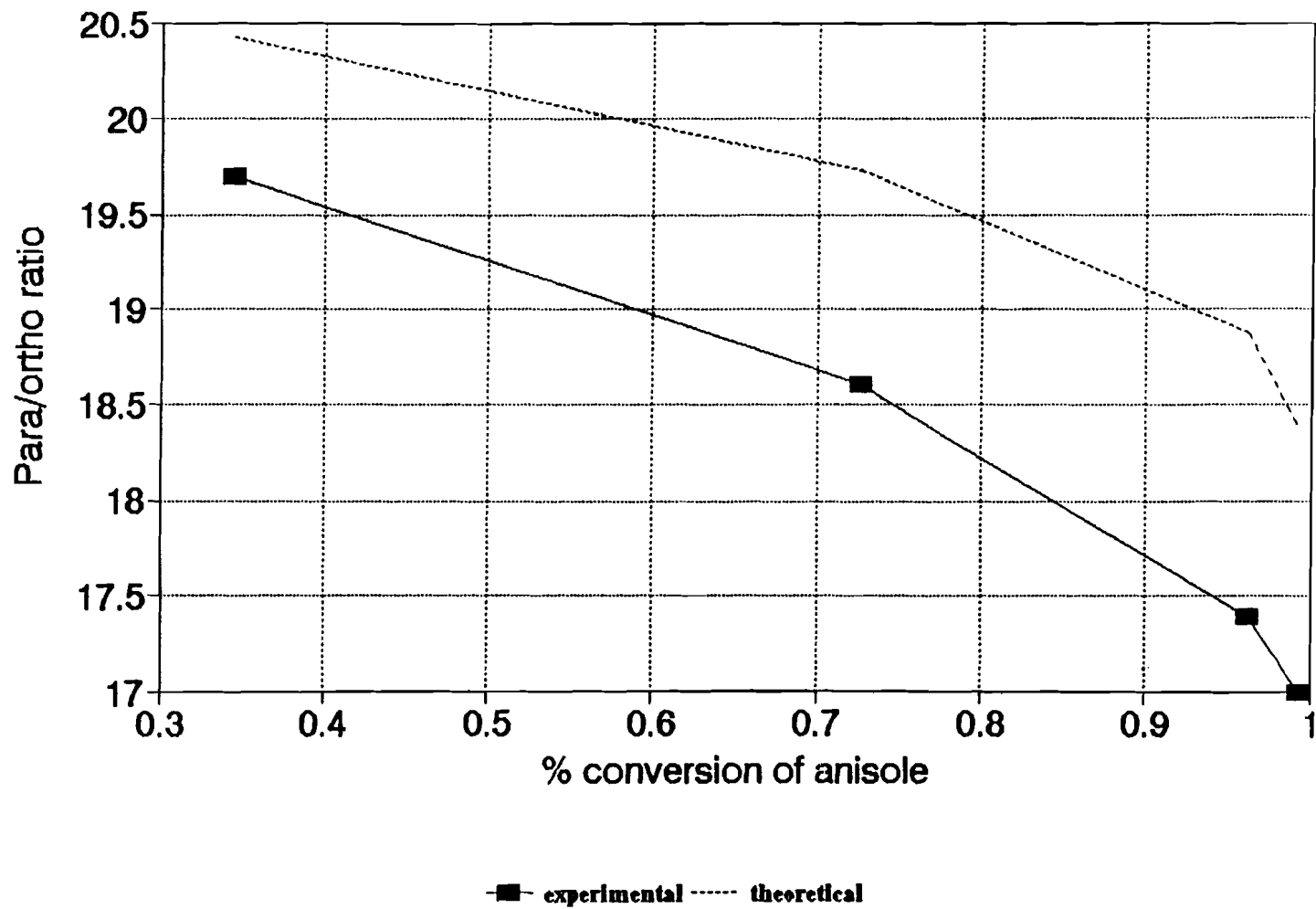50% bleach/0.32 g PTC pH 8.5

Figure 15 Para/ortho ratio
100% bleach/0.32 g PTC pH 8.5

% conversion of anisole
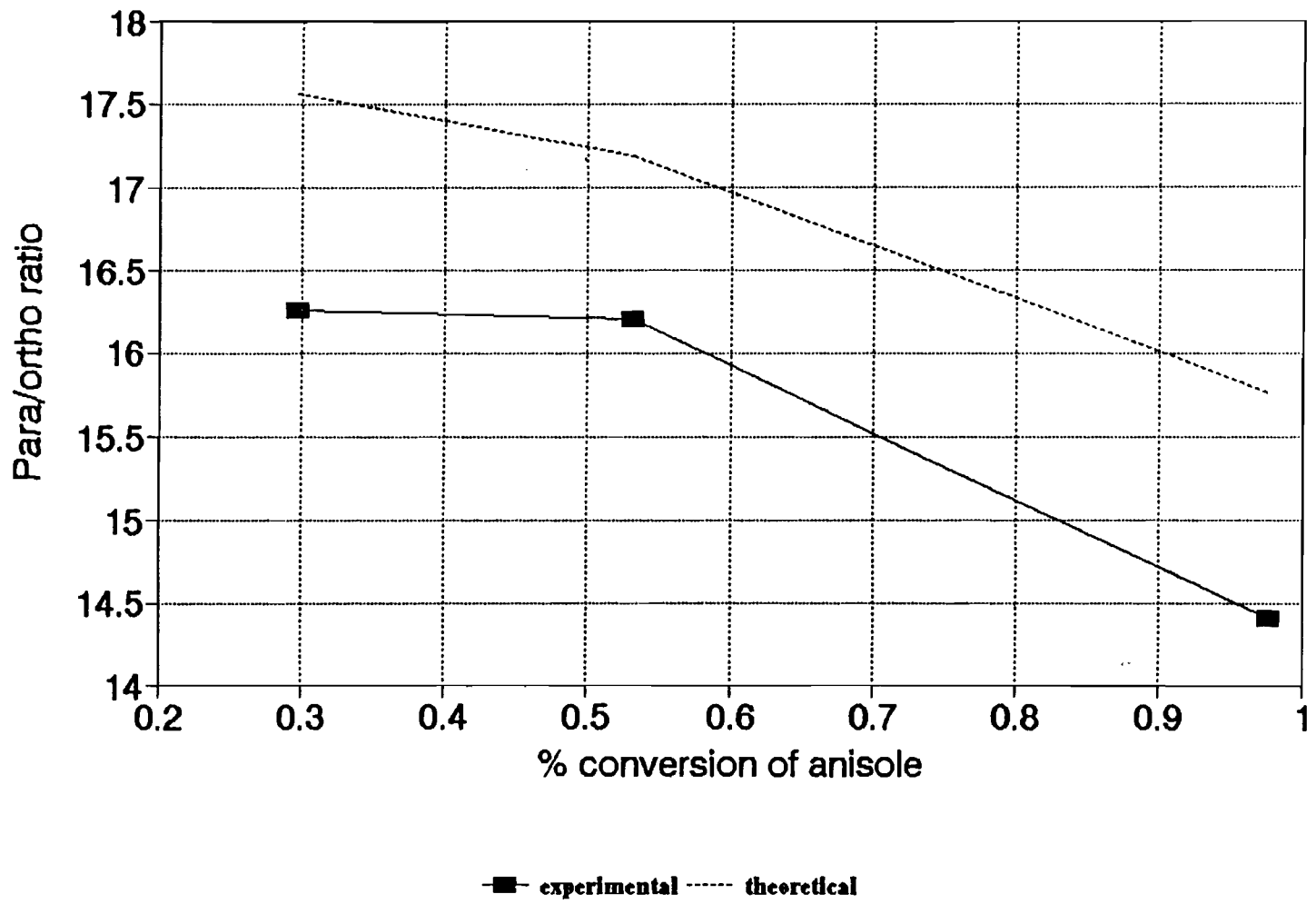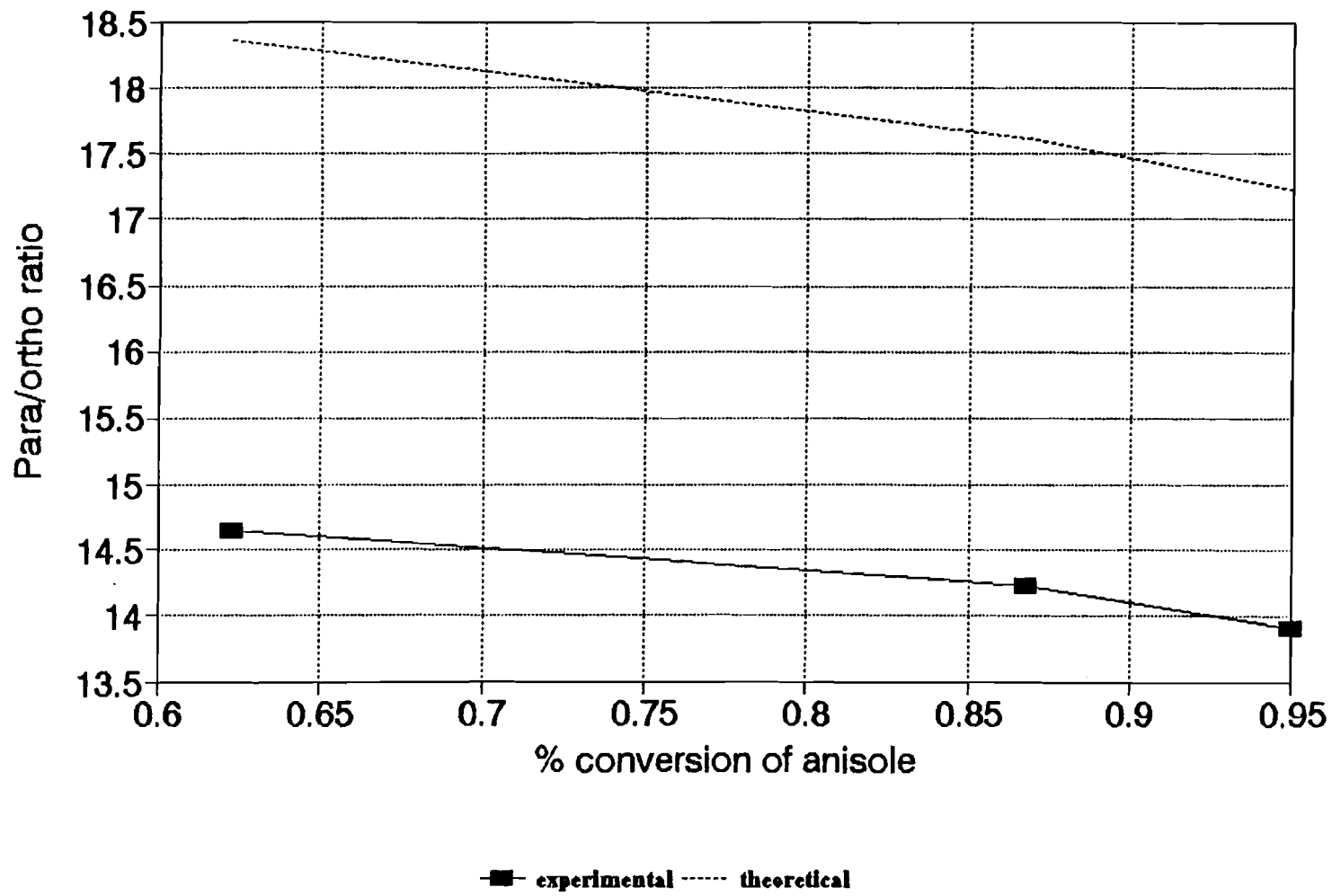
Para/ortho ratio

—■— experimental ------ theoretical

Figure 16 Para/ortho ratio
50% bleach/PTC pH 8.5 (solvent-CHCl3)

4.4 Postulating the mechanism

Summarizing the conclusions from the previous paragraph, we can identify the reaction pathways and combine them into a single scheme (Fig.17). The reactions of the central intermediate, Ar--Cl·, are being proposed based on a view that it may react with any of the reactive species present in the solution, namely Cl·, Ar--Cl· or form a 'sandwich' complex Ar--Ċl--Ar which can further react with Cl·. The fact that both free chlorine atom and π-complexed chlorine atom can act as leaving group abstractors, explains the difference in selectivities between the two outlined processes [25,29]. The complexed chlorine atom is stabilized, has lower energy and is thus more selective (the pathway without PTC). The scheme of reaction pathways logically leads to the mechanism in Figure 18. The kinetic derivations that will follow, are based on the outlined reactions. It should be noted that reaction (1) can be described as $v = k[N]*([Cl_2]/A+[Cl_2])$, where A is a constant (it is denoted K_M and called Michaelis constant in biochemistry). If A is significantly greater than chlorine concentration, which seems plausible from the K_{eq} values calculated in 4.2, then we have $v = k'[N][Cl_2]$. This expression is denoted v_1 and used in derivations below.

k_i is the rate constant for reaction #i;

v_i is the rate for reaction #i;

K_i is the equilibrium constant for reaction #i;

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[A] is the concentration of compound A;

$\sigma = \sigma$-complex $(\text{H-Ar-Cl})\cdot$;

$SCl\cdot = $ 'umbrella'; $SCl\cdot S = $ 'sandwich' (Fig.12b);

$S\cdot = C_6H_5CH_2\cdot$.

A. Anisole

$$\text{Rate} = V = d[SCl]/dt = v_5 + v_6$$

$$d[Cl\cdot]/dt = 2v_1 + v_2 - v_3 - v_6$$

$$d[SCl\cdot]/dt = v_2 + v_3 - v_4 - v_5 - v_7$$

$$d[SCl\cdot S]/dt = v_7 - v_8$$

$$d[\sigma]/dt = v_4 - v_5 - v_6 + v_8$$

Based on the assumption that in case of anisole all these species are extremely reactive, we can implement the Bodenstein steady state principle, which suggests a rapid establishment of a steady state for very reactive species, at which $dc/dt = 0$. Adding the four equations together, we get:

$$2v_1 + 2v_2 - 2v_5 - 2v_6 - v_7 = 0$$

$$V = v_1 + v_2 - 0.5v_7$$

For $v_7 \ll v_{1,2}$ we have the final equation:

(4.23) $V = k_1[Cl_2][PTC] + k_2[Cl_2][S]$.

The rate law in (4.23) is first order in chlorine and "0+1" in anisole. The concentration of PTC affects the zero-order term, as it was deduced from the experimental data.
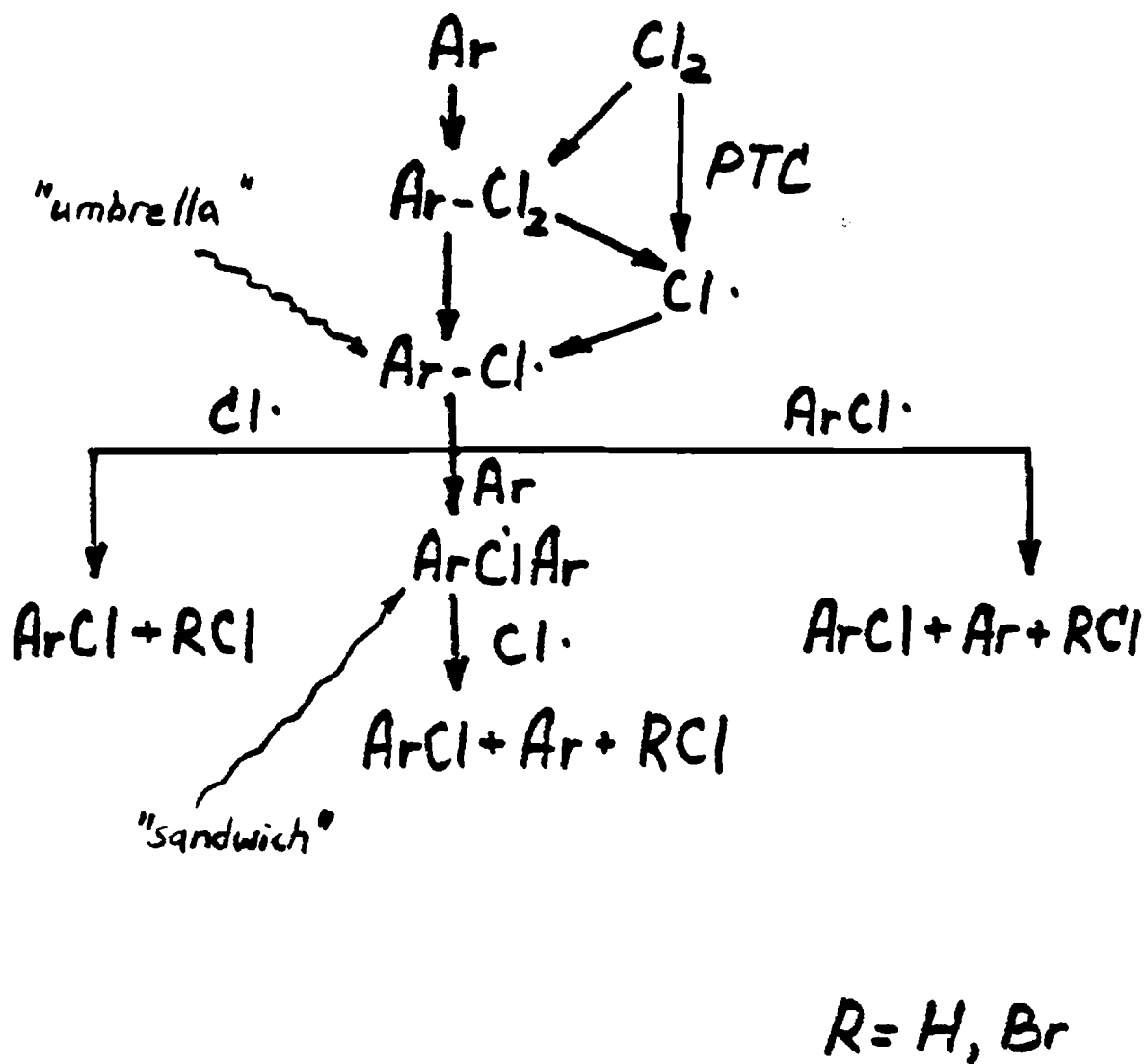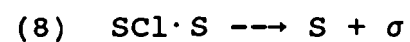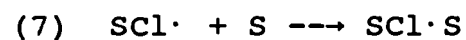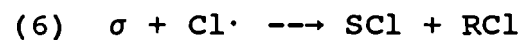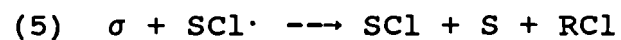
Figure 17   Scheme of the mechanism for anisole or bromobenzene
chlorination

(1) Cl_2 + PTC $\longrightarrow$ PTC + 2 Cl·

(2)   S + Cl_2 $\longrightarrow$ SCl· + Cl·

(3)   S + Cl· $\longrightarrow$ SCl·

(4)   SCl· $\longrightarrow$ σ

(5) σ + SCl· $\longrightarrow$ SCl + S + RCl

(6) σ + Cl· $\longrightarrow$ SCl + RCl

(7)   SCl· + S $\longrightarrow$ SCl·S

(8)   SCl·S $\longrightarrow$ S + σ
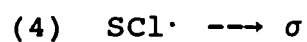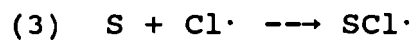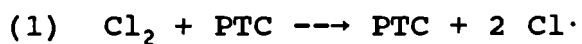
R = H, Br

σ = σ-complex (Cl-Ar-R)·

Figure 18   Mechanism for anisole or bromobenzene chlorination

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B. Bromobenzene

For bromobenzene the steady state principle is inapplicable, because, e.g. the presence of bromine as a leaving group instead of hydrogen makes the σ-complex disintegration the rate determining step, as opposed to the σ-complex formation being the slower step for anisole. When for two consecutive reactions $k_1 > k_2$, the steady state is being established extremely slowly, and using it as an approximation throughout the reaction would be inaccurate. The same applies to SCl· and SCl·S due to their inertness. As a result, we would be using the 'rapid equilibration' approximation as a basis for the calculations.

$$d[SCl]/dt = k_5[\sigma][SCl\cdot] + k_6[\sigma][Cl\cdot] = [\sigma]*\left(k_5[SCl\cdot] + k_6[Cl\cdot] \right)$$

$$[\sigma] = K_4[SCl\cdot] \; ; \; [SCl\cdot] = K_3[S][Cl\cdot] \; ; \; [Cl\cdot] = K_1^{\frac{1}{2}}[Cl_2]^{\frac{1}{2}}.$$

The resulting equation will be:

$$V = K_1 K_3 K_4 [S][Cl_2]*\left(k_5 K_3[S] + k_6 \right)$$

With $k_6 \ll k_5 K_3[S]$ we have:

(4.24) $$V = k_5 * K_1 (K_3)^2 K_4 [s]^2 [Cl_2],$$

which is in agreement with the earlier observations.

C. Toluene

The mechanism for toluene chlorination (Fig.19) is based on the previous mechanism (A-B), but is designed to take into account the α-chlorination, which outweighs the chlorination in the aromatic ring.

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(1) Cl_2 + PTC --→ PTC + 2 $Cl\cdot$

(2) S + Cl_2 --→ $SCl\cdot$ + $Cl\cdot$

(3) S + $Cl\cdot$ --→ $S\cdot$ + HCl

(4) S + $SCl\cdot$ --→ $S\cdot$ + HCl + S

(5) $S\cdot$ + Cl_2 --→ SCl + $Cl\cdot$
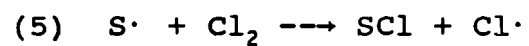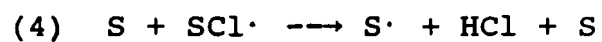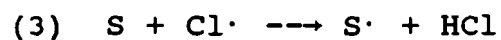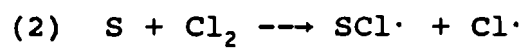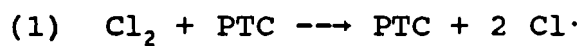
S = toluene

$S\cdot$ = benzyl radical

Figure 19   Mechanism for toluene chlorination

$$V = d[SCl]/dt = v_5$$

$$d[Cl\cdot]/dt = 2v_1 + v_2 - v_3$$

$$d[SCl\cdot]/dt = v_2 - v_4$$

$$d[S\cdot]/dt = v_3 + v_4 - v_5$$

Using the steady state principle (the species mentioned are all extremely reactive) we add up the three equations:

$$2v_1 + 2v_2 - v_5 = 0$$

(4.25) $V = 2*[v_1 + v_2] = 2k_1[Cl_2][PTC] + 2k_2[S][Cl_2]$

The reaction orders are identical with those in A.

As we have seen, the postulated mechanism fits the experimental data and explains the phenomena observed.

4.5 Notes

The present notes contain comments to some of the observations and conclusions made by scientists in the field of the current investigation in the light of the newest information brought to the readers' attention in this thesis. (1) In the process of postulating the mechanism (Fig.18) the scheme proposed by Hamilton and his group [14] was being looked upon from the thermodynamic point of view. As a result, the value of $\Delta G° = 657$ kJ/mol was obtained for the rate-determining step in the chlorination of anisole (thermodynamic values from [30] were used):

$$ClO\cdot + S \longrightarrow ClO^- + S\dotplus.$$
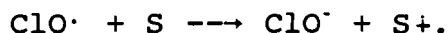
Naturally, it means that the activation energy E > 657 kJ/mol. This value is inconsistent with the observation of a fast

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reaction.

(2) The mechanism postulated by Echols et al. [5] is slightly similar to the one that is being proposed in the current project. The inconsistencies in their data can be explained by a concentration effect, which becomes sizable when the concentration of the aromatic substrate is dramatically changed. It may influence the stoichiometry of the arene-radical interaction [25]. Besides at low substrate concentrations bromine atom can no longer be considered a bromine-abstracting species.

(3) The hypothesis of ClO· as the major reacting species is being rejected for bleach at pH < 9, as a result of a current investigation. The qualitative test for chlorate ion $(MnSO_4+H_3PO_4)$ [31] confirmed its presence immediately after the acidification which suggests that hypochlorite rapidly disproportionates. The formation of Cl_2O would thus be extremely unlikely. However, when NaOCl solution is being used instead of bleach, the formation of small quantities of chlorine monoxide seems feasible. Besides, the fastest PhBr chlorination was observed at pH 3, which strongly supports the idea of chlorine as the reactant.
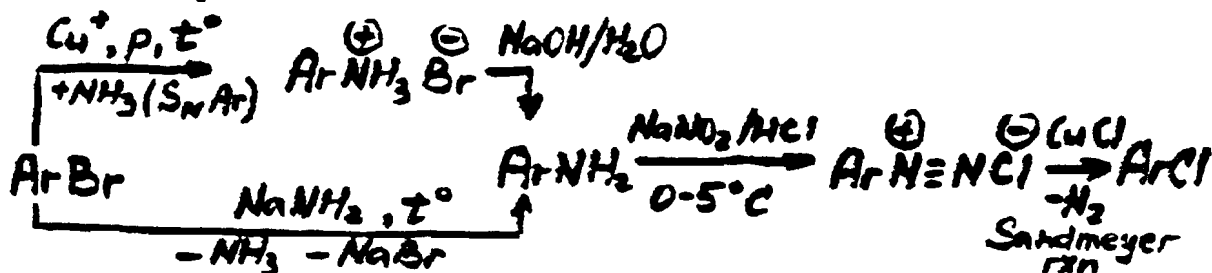
(4) Ni(II) complexes used by Burrows and O'Connor to accelerate the reaction [22], can catalyze Cl· formation from both Cl_2O and Cl_2. The obstacles for the direct investigation include the possible decomposition of the complex and thus the impossibility of working with a broad range of pH's.

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(5) The NaCl studies disproved Hamilton's mechanism, since according to his scheme the addition of chloride should decrease the reaction rate (according to the LeChatelier principle). In fact, the addition of chloride significantly facilitated the reaction.

(6) The present study gives a possible explanation to the failure of Landolt's group to find a Hammett correlation with either σ or σ+ values. The order in substrate depends on its reactivity, and this will make it impossible to compare the rate constants.

4.6 Synthetic applications

The reactions under investigation in the present project can be used for synthetic purposes. Chlorination of toluene is very fast and selective. Chlorination of anisole gives more byproducts, but the conditions can be adjusted to increase the p/o ratio based on the results of this project. The ipso-substitution in PhBr may prove to be the most useful of the three. This direct halodehalogenation helps to avoid a multistep chain:

$$Ar-Br \xrightarrow[+NH_3 (S_N Ar)]{Cu^+, p, t^\circ} \quad Ar\overset{\oplus}{N}H_3 \, \overset{\ominus}{Br} \xrightarrow{NaOH/H_2O}$$

$$Ar-Br \xrightarrow[-NH_3 \ -NaBr]{NaNH_2, \ t^\circ} Ar-NH_2 \xrightarrow[0-5^\circ C]{NaNO_2/HCl} Ar\overset{\oplus}{N}\equiv N\overset{\ominus}{Cl} \xrightarrow[-N_2]{\ominus CuCl} Ar-Cl$$

Sandmeyer rxn

This reaction can be used for substituting chlorine for

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bromine in complex substrates carrying numerous substituents.

4.7 Prospects for future investigations

One of the topics for future research could be the mechanistic study of bromodechlorination of chlorobenzene with NBS. The reaction with respect to NBS might be investigated with a UV instrument. Usually the brominations with NBS are said to proceed through the intermediate formation of bromine. However, NBS was reported [6] to give higher yields of brominated products in bromodechlorination than bromine itself. Monitoring NBS and bromine concentrations with a UV spectrometer could provide explanations to this phenomenon.

Another possible avenue for future investigations could be the study of the PTC-Cl_2 complex by using FTIR and ^{1}H NMR. By subtracting the IR spectrum of the organic mixture without PTC from the analogous spectrum with PTC one might be able to observe the Cl-Cl vibration in the 100-300 cm^{-1} range due to the dipole moment induced by the positively charged nitrogen atom of PTC. The ^{1}H NMR spectrum might show the shift of the two peaks in the "alkyl range" of $\approx$ 0.9 - 2.4 δ, or a third peak, which would indicate a distortion of the tetrahedral structure of the ammonium salt. This could provide a insight into the actual catalytic mechanism, kinetic evidence for which was found in the current investigation.

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