MECHANISMS OF AUTOXIDATION OF ORGANO-BORON COMPOUNDS

by

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ABSTRACT OF THESIS

The preparation of simple organoboron compounds has been considered, with particular mention of hydroboration and the reactions of trialkyl boranes. The autoxidation of hydrocarbons has been reviewed in detail, together with the modes of action of antioxidants. A historical survey of the autoxidation of organoboron compounds has also been made.

Extensive racemisation of configuration was shown to accompany the autoxidation of the optically active borane, diisopinocampheyl-s-butyl borane. In addition, the autoxidation was inhibited by the free radical galvinoxyl and by other commercial antioxidants. The autoxidation of other organoboron compounds could also be inhibited or retarded by a variety of phenols, amines, free radicals and sulphur or phosphorus containing antioxidants. A compound was isolated from the nitroxide inhibited autoxidation of triisobutyl borane, and shown to be an adduct between isobutyl radicals and the inhibitor.

It is concluded that a free radical chain mechanism governs the autoxidation, as in hydrocarbon systems.

The behaviour of antioxidants can satisfactorily be interpreted in terms of this conclusion, and the unique efficiency of galvinoxyl is attributed to its ability to scavenge alkyl, alkylperoxy and alkoxy radicals.

The great speed of autoxidation of some boranes is thought to be due to the very facile displacement of an alkyl radical from boron by alkylperoxy, and the importance of efficient stirring in these systems is emphasised. The rearrangement of peroxide, and the possible repercussions of this on the initiation of autoxidation, are discussed. Due mainly to the complexity of initiation processes, absolute rate constants for the individual steps in the chain reaction could not be evaluated. However, this preliminary study indicates the line for future investigations.

ACKNOWLEDGEMENTS

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Chapter 1

Preparation and Reactions of Compounds Containing Carbon-Boron Bonds

The preparation of compounds with direct carbon-boron bonds falls into two main classes. The first is transmetallation, where electrophilic cleavage of an organometallic bond results from its interaction with a boron halide or boron ester. The second is the reaction of diborane with an olefin, known as hydroboration.

a) Preparation of Carbon-Boron Bonds by Transmetallation:-

In the context of this chapter, transmetallation is taken to mean the reaction of an organometallic compound, R_nM, with a boron substrate, BX_3, where X = halogen or alkoxy.

$$3R_nM + nBX_3 \longrightarrow nBR_3 + 3MX_n \qquad (i)$$

i) Zinc. The use of dialkyl zincs to prepare boron compounds is largely of historical interest, as Frankland[1], in 1859, first made trimethyl borane from dimethyl zinc:

$$3ZnMe_2 + B(OEt)_3 \longrightarrow Me_3B + 3MeZn(OEt) \qquad (ii)$$

Ethyl diethylborinate was similarly prepared[2]:

$$2ZnEt_2 + B(OEt)_3 \longrightarrow Et_2B(OEt) + 2EtZn(OEt) \qquad (iii)$$

ii) Magnesium. The use of dialkyl zincs was superceded by Grignard reagents, which are much easier to prepare. Khotinsky and Melamed[3] obtained several alkyl and aryl boronic acids in this way:

$$B(OR)_3 + R'MgBr \longrightarrow R'B(OR)_2 \xrightarrow{H_2O} R'B(OH)_2 \qquad (iv)$$

A large number of different boron substrates have been used, but the principal ones are boron trifluoride[4], usually as the more convenient diethyl etherate[5], and trialkyl borates[6]. In both cases, one, two or three organic residues can be attached to boron by using the appropriate quantities of reactants.

iii) Other Metals. Although Grignard reagents have remained the most generally used in transmetallation, many other metals have also been investigated. Chief among these is lithium[7]:

$$B(OBu^n)_3 + LiBu^n \longrightarrow Bu^nB(OBu^n)_2 + LiOBu^n \qquad (v)$$

This method gave a higher yield (60%) than was obtained via the Grignard reagent (42%).

Michaelis et. al.[8] used diaryl mercury compounds to prepare aryl boron dihalides as long ago as 1880.

$$HgPh_2 + 2BCl_3 \xrightarrow[\text{sealed tube}]{\text{heat.}} 2PhBCl_2 + HgCl_2 \qquad (vi)$$

Recently, aryl mercury chlorides have also been shown to give fairly good yields (57-76%) of the aryl boron

dichloride[9]:

$$ArHgCl + BCl_3 \longrightarrow ArBCl_2 + HgCl_2 \qquad \text{(vii)}$$

Aluminium alkyls[10] have aroused interest, largely because their low price provides an economically attractive route to organoboron compounds on an industrial scale:

$$(RBO)_3 + 2AlR_3 \longrightarrow 3BR_3 + Al_2O_3 \qquad \text{(viii)}$$

Sodium[11], cadmium[12], tin[13], lead[14], antimony[15], and sulphur[16] compounds have also been used to prepare boron compounds.

b) Hydroboration:-

The reaction of diborane with olefins was first investigated in 1948 by Hurd[17], who heated ethylene, containing two per cent diborane, in a sealed tube at 100° for four days, after which he isolated a small quantity of triethyl borane:

$$6CH_2{=}CH_2 + B_2H_6 \longrightarrow 2Et_3B \qquad \text{(ix)}$$

This rather unpromising reaction attracted little attention until Brown and Subba Rao[18] discovered that a solution of sodium borohydride and aluminium chloride in diglyme rapidly converted 1-pentene and other simple olefins into trialkyl boranes:

$$3NaBH_4 + AlCl_3 + 9PrCH{=}CH_2 \longrightarrow 3(PrCH_2CH_2)_3B + 3NaCl + AlH_3 \qquad \text{(x)}$$

It was soon discovered that diborane itself, in ether solvents, reacted even faster[19], the reaction with simple olefins being complete within minutes at room temperature:

$$B_2H_6 + 6RCH{=}CH_2 \xrightarrow[\text{solvents}]{\text{ether}} 2(RCH_2CH_2)_3B \qquad \text{(xi)}$$

Largely due to subsequent work by Brown and co-workers, many useful reactions of trialkyl boranes have been uncovered. Coupled with these, hydroboration has now become an exceedingly valuable tool in synthetic organic chemistry.

i) Hydroboration Procedures. The experimental techniques for hydroboration have been described in detail by Brown[20,21]. Briefly, diborane, generated in situ or externally, is allowed to react with the appropriate olefin under nitrogen in a dry ether solvent, usually tetrahydrofuran (THF) or diethylene glycol, dimethyl ether (diglyme). In most cases, the reaction is over almost as fast as the reagents are brought together (Equation xi).

The reagents of choice, for either method of diborane generation, are sodium borohydride and boron trifluoride etherate:

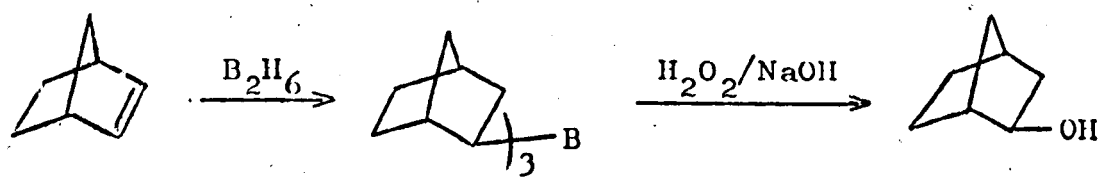
$$3NaBH_4 + 4BF_3{\cdot}OEt_2 \xrightarrow[\text{diglyme}]{\text{THF,}} 3NaBF_4 + 2B_2H_6 + 4Et_2O \qquad \text{(xii)}$$

Other reagent pairs have been used[22], and in particular, Long and Freeguard[23] have made a careful study of methods of producing diborane. Of the seventeen reactions which were investigated, the following gave a 98% yield of diborane, uncontaminated by side products:

$$HgCl_2 + 2LiBH_4 \xrightarrow{\text{diglyme}} Hg + 2LiCl + B_2H_6 + H_2 \qquad \text{(xiii)}$$

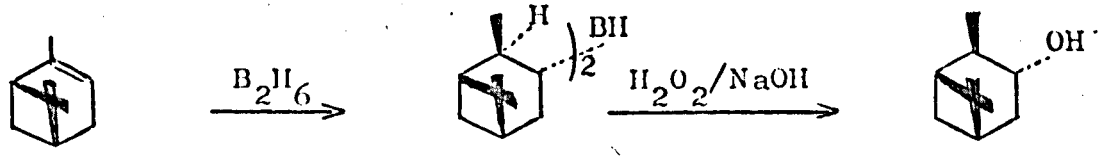
$$I_2 + \underset{(LiBH_4)}{2NaBH_4} \xrightarrow{\text{diglyme}} \underset{(LiI)}{2NaI} + B_2H_6 + H_2 \qquad \text{(xiv)}$$

ii) Hydroboration Mechanism. Brown and Zweifel[24] examined the products obtained by hydroboration of a large number of representative olefins, followed by oxidation with alkaline hydrogen peroxide:



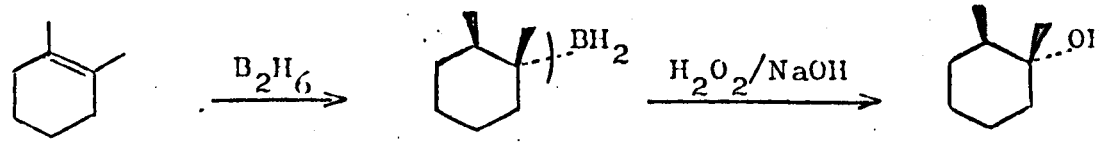
(xv)

Norbornene — exo-norborneol



(xvi)

α-pinene — isopinocampheol



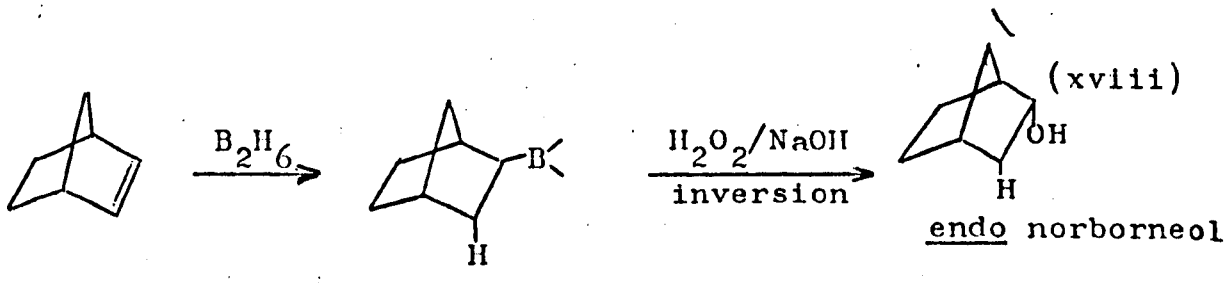
(xvii)

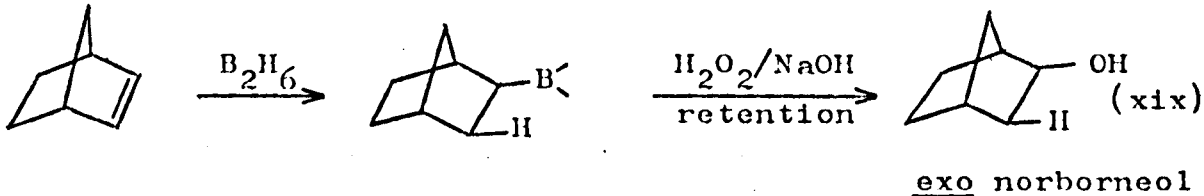
1,2 dimethyl cyclohexene — cis 1,2 dimethyl cyclohexanol

In each case the alcohol isolated corresponded to an overall

cis hydration of the double bond from the less hindered side of the molecule, even when this produced the thermodynamically less stable isomer (Equation xvi). The only two reports in the literature of an overall trans hydration of the double bond[25,26] have been subsequently explained in terms of other factors[27,28], and the results obtained by Brown and Zweifel appear to be generally applicable.

There are two possible routes to the observed products. Either the hydroboration step is a trans addition, followed by inversion of configuration on alkaline hydrogen perodide oxidation, (Equation xviii), or the hydroboration occurs cis, and oxidation involves retention of configuration(Equation xix).





Brown[24] argued that Equation xix represents the correct sequence. If the hydroboration were trans he would expect the boron atom to attach itself to the less hindered exo

position, when subsequent inversion on oxidation would give *endo* norborneol, (Equation xviii), whereas the *exo* alcohol is the product actually obtained.

Kuivila and co-workers[29] have made a kinetic study of the reaction between benzene boronic acid and hydrogen peroxide under a variety of conditions, and have concluded that the base-catalysed reaction follows the scheme depicted below:

$$OH^- + H_2O_2 \rightleftharpoons OOH^- + H_2O \qquad (xx)$$

$$C_6H_5\text{-}\overset{OH}{\underset{|}{B}}\text{-}OH + OOH^- \rightleftharpoons C_6H_5\text{-}\overset{OH}{\underset{|}{\overset{|}{B}}}^{-}\text{-}OH \qquad (xxi)$$

(I)

$$(I) \longrightarrow C_6H_5\text{-}\overset{OH}{\underset{|}{\overset{|}{B}}}^{-}\text{-}OH + OH^- \qquad (xxii)$$

(II)

$$(II) \longrightarrow \overset{OH}{\underset{OC_6H_5}{\overset{|}{B}}}\text{—}OH \xrightarrow[fast]{H_2O} B(OH)_3 + C_6H_5OH \qquad (xxiii)$$

The Hofmann, Curtius, Lössen and Beckmann reactions were found by Kenyon *et al.*[30] to proceed with complete retention of configuration of the migrating group, and by analogy, reaction (xxii) above would also be expected to do so. If this is so, and if this mechanism can be applied in three successive stages to the oxidation of trialkyl

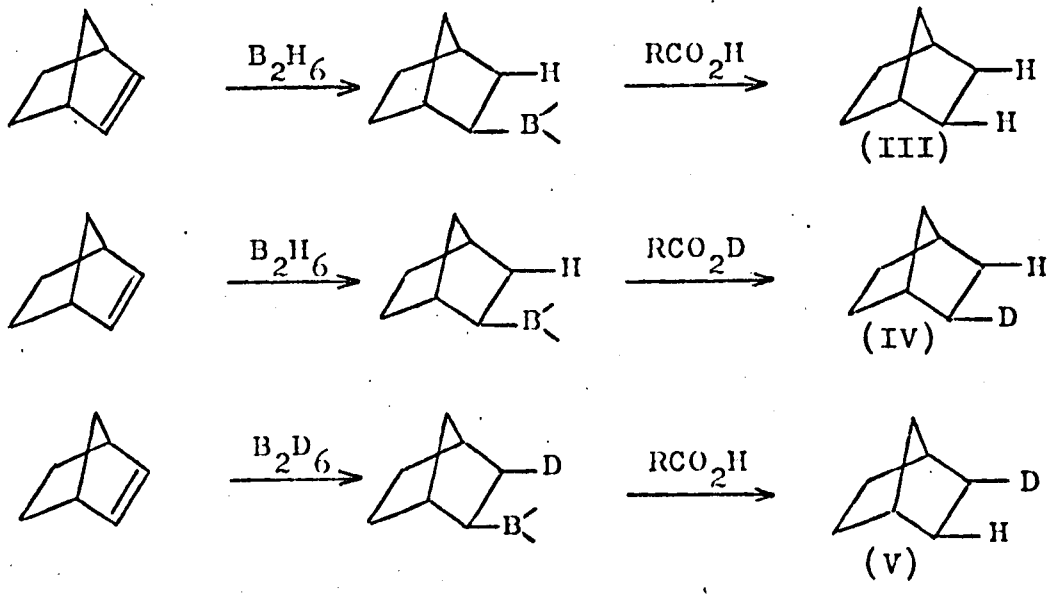
boranes, Equation (xix) is correct, and hydroboration involves a cis addition to the double bond.

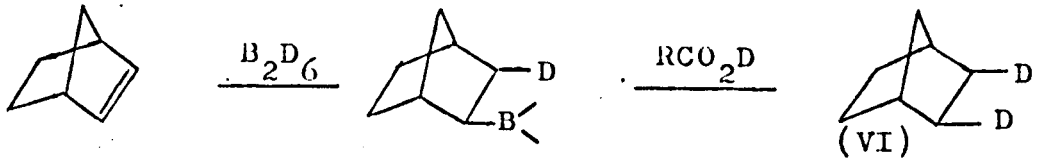
Further evidence comes from the protonolysis reaction. Although trialkyl boranes are resistant to attack by strong mineral acids, two of the three groups can be removed by excess acetic acid at room temperature, and all three are replaced by hydrogen on refluxing with propionic acid in diglyme for 2-3 hours[31].

$$R_3B \xrightarrow[\text{Room Temp}]{CH_3CO_2H} RB(OH)_2 + 2RH \qquad \text{(xxiv)}$$

$$R_3B \xrightarrow[160^\circ,\ 2\text{-}3\ \text{hrs}]{C_2H_5CO_2H} B(OH)_3 + 3RH \qquad \text{(xxv)}$$

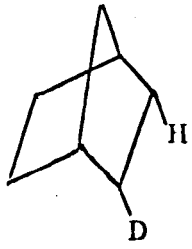
Brown and Murray[32] carried out the following series of isotopic reactions:



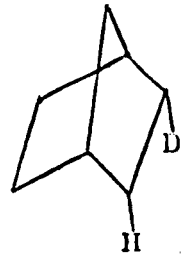


(VI) (xxix)

The products (III) - (VI) were examined by NMR spectroscopy. (IV) and (V) were identical, with one equatorial hydrogen replaced by deuterium relative to (III). In (VI), two equatorial hydrogens were replaced, thus confirming that the overall hydroboration-protonolysis occured cis, and from the less hindered side of the molecule. If the hydroboration were trans, with the boron atom attaching itself to the less hindered exo position, inversion of configuration on protonolysis would give the corresponding endo products (VII) and (VIII) below, instead of (IV) and (V) respectively:



(VII)

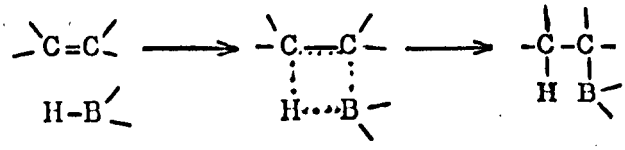


(VIII)

A trans hydroboration placing the boron at the more hindered endo position, followed by inversion of configuration

on protonolysis, would give the observed results (III) - (VI), but there is no evidence to favour this mechanism over that suggested by Brown. NMR spectroscopic examination of the borane intermediate would presumably distinguish absolutely between these two possibilities.

Brown and Zweifel[24,33] have proposed the existence of a four centre transition state which satisfactorily explains both the cis addition and the preferential attack from the less hindered side of the molecule:



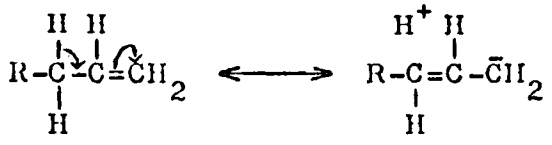
(xxx)

iii) Products of Hydroboration. In the hydroboration of unsymmetrical olefins, RCH=CHR', there are two possible positions of attachment of the boron atom, and its distribution between these is governed by steric and electronic factors. These usually act in the same direction to place most of the boron on the least sterically hindered carbon atom, giving anti-Markownikoff addition to the double bond.

1) Electronic Effects. The hydroboration of a straight-chain terminal olefin places 94% of the boron on the primary carbon atom and 6% on the secondary[33]:

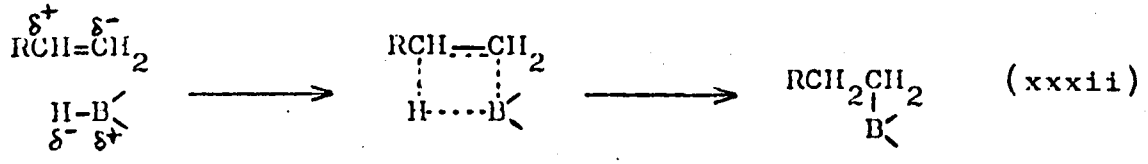
$$RCH{=}CH_2 \xrightarrow[2)\ H_2O_2/NaOH]{1)\ B_2H_6} \underset{(94\%)}{RCH_2CH_2OH} + \underset{OH\ \ (6\%)}{RCHCH_3} \qquad \text{(xxxi)}$$

This is complementary to the addition of hydrogen bromide to olefins, and can be explained by the polarity of the double bond. Contributions due to structures such as (IX) lead to a slight nett negative charge on



(IX)

the terminal carbon atom. Since the polarity of the boron-hydrogen bond is the reverse of that of hydrogen bromide, the primary borane would be expected (Equation xxxii)

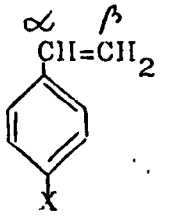


(xxxii)

The discrimination between the two positions is not so marked for internal olefins, in which the double bond is not so polar. While hydroboration-oxidation of 1-pentene gives 94% 1-pentanol and 6% 2-pentanol, 2-pentene yields 55% 2-pentanol and 45% 3-pentanol[33]. Trisubstituted olefins, $R_2C{=}CHR$ show considerable selectivity, but this is probably attributable mainly to steric effects, since reaction only goes to the dialkyl borane stage.

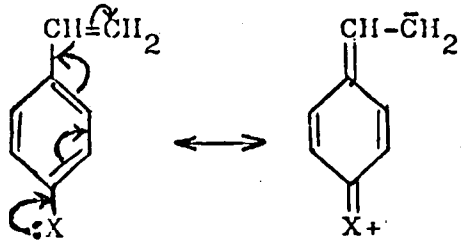
The presence of a phenyl group α to a double bond exercises a profound influence on the direction of addition.

Brown and Sharp[34] investigated the reaction between substituted styrenes and diborane by analysing the alcohols produced on oxidation. Whereas in styrene itself

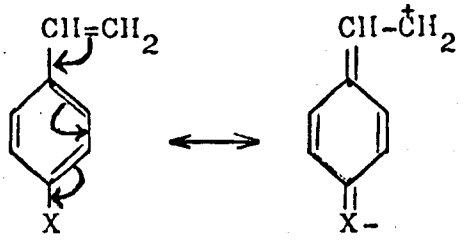


(X)

the distribution was 81%β, 19%α; for X=MeO, an electron releasing group it was 93%β, 7%α; and for X=Cl, an electron withdrawing group it was 73%β, 27%α. These results can be explained in terms of the ability of the phenyl group to donate or accept electrons, as in (XI) and (XII)



(XI)



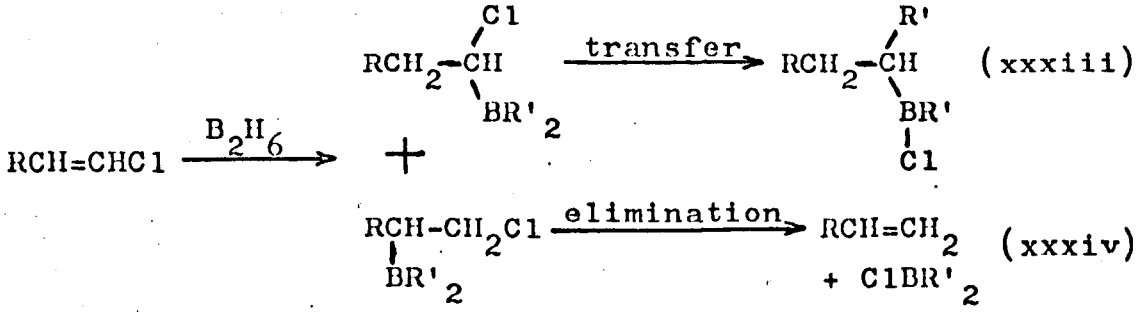
(XII)

A structure such as (XI) will be favoured by electron releasing groups such as methoxy, while p-chlorostyrene will have a greater contribution from structure (XII).

In agreement with this, it was found that a good straight line was obtained by plotting log (ratio β/α

alcohol) against the Hammett σ^+ constant of the substituent for several different groups X in the *meta* and *para* positions.

Electronic effects in the hydroboration of vinyl and allyl compounds, particularly halides, can lead to transfer and elimination reactions:



Many papers have been published in this field[35-40].

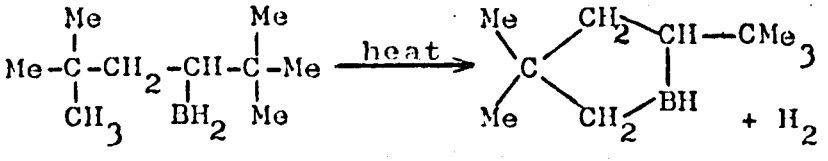
2) *Steric Effects*. The steric requirements of diborane itself are quite small. However, in the case of bulky olefins, the mono- and dialkyl boranes show much greater stereoselectivity. Table 1 indicates the degree of alkylation of borane achieved with different olefins. Disiamyl borane, thexyl borane and other partially alkylated boranes exhibit a far greater stereoselectivity towards reactive olefins than diborane itself. Thus with diborane, 1-pentene and styrene give 94% and 80% respectively of the terminal borane, whereas with disiamyl borane this figure is >98% for both these olefins[42]. Furthermore, the latter reagent is far more sensitive to the structure of the olefin

than the former, and it is possible to remove completely from a mixture the more reactive (i.e. less hindered) olefin by using a calculated quantity of disiamyl borane. For instance, cis-2-pentene was selectively removed from a mixture with trans-2-pentene in this way[42].

iv) Isomerisation and Displacement. If the organoborane from an internal olefin is heated for a few hours under reflux in diglyme, an equilibrium distribution of boron in the molecule is realised; i.e. the majority will be found at the least hindered position[43].

$$\text{3-hexene} \xrightarrow{B_2H_6} (\text{3-hexyl})_3B \xrightarrow[1\ hr]{150^\circ} \underset{(90\%)}{(\text{1-hexyl})_3B} + \underset{(6\%)}{(\text{2-hexyl})_3B} + \underset{(4\%)}{(\text{3-hexyl})_3B} \qquad \text{(xxxv)}$$

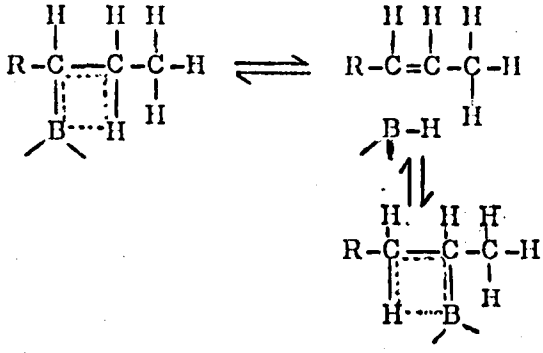
The boron atom will travel past a tertiary carbon atom, but not past a quaternary one. It will move out of a ring to an exocyclic position, though rather slowly. Attempts to force the boron to migrate past a quaternary carbon atom resulted in evolution of hydrogen and cyclisation[44]:



(xxxvi)

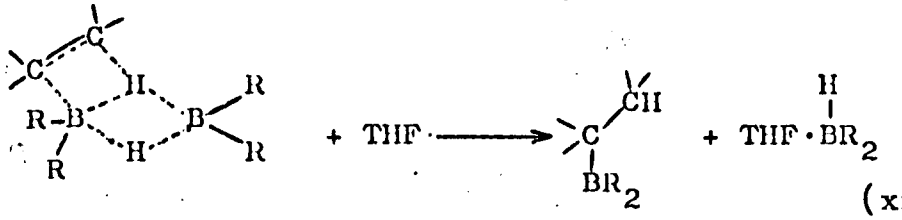
Brown envisaged the isomerisation as a series of

reversible additions, gradually giving an equilibrium distribution of the boron (ref.21 p.146):



etc. (xxxvii)

Subsequently he modified this scheme slightly to allow for the catalytic effect of B-H bonds by suggesting that the hydroborating entity was dimeric, and that the bridging B-H-B bond is the one which participates in the reaction:



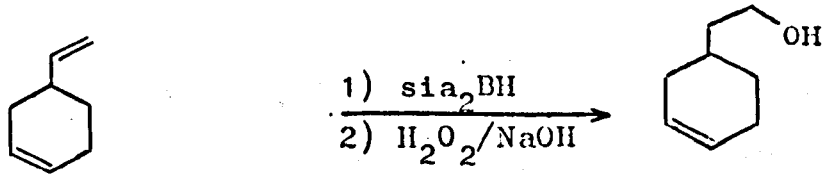
(xxxviii)

The ether solvent serves to lower the activation energy of the transition state by co-ordinating with the HBR_2 moiety.

It is a consequence of this mechanism that an added olefin would participate in the reaction. In fact, if an involatile olefin, such as decene, is added to tripentyl borane, refluxing in diglyme, pentene can be collected at the top of the fractionating column:

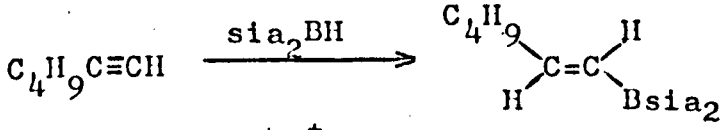
$$C_8H_{17}CH{=}CH_2 + (C_5H_{11})_3B \longrightarrow (C_{10}H_{21})_3B + C_3H_7CH{=}CH_2 \qquad \text{(xxxix)}$$

v) Dienes and acetylenes. Hydroboration of dienes and acetylenes offers few original features. Dienes frequently yield cyclic products[45-47], particularly with alkyl boranes, RBH_2 [48,49]. Of these, the six-membered boron heterocycles appear to be thermodynamically most stable[46]. Monohydroboration of conjugated dienes is not possible, even with disiamyl borane, sia_2BH, but this reagent is useful for addition to the more reactive of two unconjugated double bonds[42]:

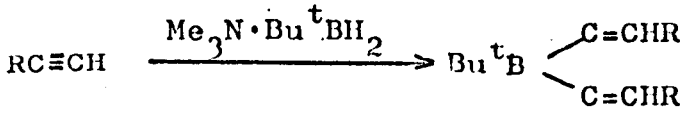


4 vinyl cyclohexene (xl)

The same reagent converts acetylenes into vinylboranes in high yield[50], as do other specialised hydroborating agents[49,51]



(xli)



(xlii)

Protonolysis of vinyl boranes gives the cis-olefin in high yield[50].

c) Reactions of Trialkyl Boranes:-

Prior to the discovery of hydroboration, trialkyl boranes, available only via transmetallation, were relatively

tedious to prepare, and appeared to offer little promise as reagents in synthetic organic chemistry. However, hydroboration occurs under very mild conditions, and this has stimulated considerable research into the reactions of organoboranes. Usually the organoborane can be converted into the desired product without the need to isolate it first.

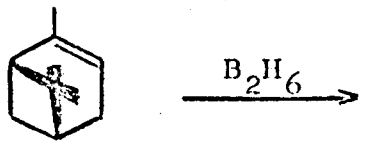
Oxidation with alkaline hydrogen peroxide replaces the boron atom by hydroxyl. This reaction is quantitative and has been used to estimate boronic acids by Snyder, Kuck and Johnson[6b] (Equations xx - xxiii). The oxidation proceeds with retention of configuration, as discussed earlier (p. 5) to place the hydroxyl group at the position formerly occupied by the boron atom.

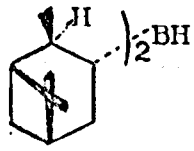
Other reactions of trialkyl boranes are summarised in Tables 2 and 3.

d) Diisopinocampheyl Borane:-

Diborane reacts with highly hindered olefins to give alkyl boranes, RBH_2 and R_2BH (Table1). Since the steric requirements of these compounds in subsequent reactions with olefins are very much greater than those of diborane, they are the reagents of choice when a very stereospecific reaction is desired.

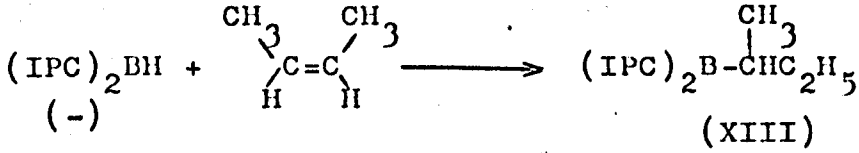
The reaction of (+) α-pinene and diborane gives (-)diisopinocampheyl borane, $(IPC)_2BH$:





(xliii)

In addition to showing the same stereoselectivity as disiamyl borane, this reagent can be used to achieve asymmetric syntheses. Thus <u>cis</u> olefins yield mixed boranes, which, on oxidation give the corresponding alcohols in high optical purity[72]:



(XIII) (xliv)

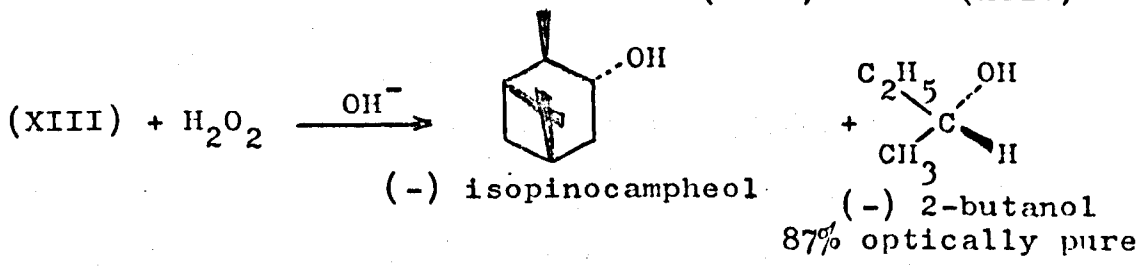
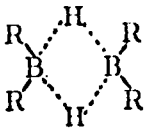


Table 4 shows some other syntheses achieved with $(IPC)_2BH$.

Brown has used a three-dimensional model[72] of diisopinocampheyl borane monomer to predict the configurations of the alcohols derived from <u>cis</u> olefins. However, this method fails with <u>trans</u> and hindered olefins, which only react very slowly, with displacement of pinene from the reagent. Streitweiser[78] has explained the configuration of

1-butanol-1-d (Table 4) on the basis of a 3-centre transition state, instead of the accepted 4-centre one. He also considers the borane monomer as reacting, although all experimental work on these systems points to the dimer,



as the reagent[72,79]. A recent communication[80] pointing out this fact postulates a mechanism involving overlap of the olefinic π orbital with the B-B antibonding orbital of the dimer. Using this model, correct chiralities are predicted for the alcohols derived from both cis and trans olefins.

e) Conclusion:-

Hydroboration is a reaction which requires mild conditions, is tolerant of many functional groups, and has a very clearly defined stereochemistry. The resulting organoboranes are readily converted into useful organic compounds by a wide variety of reactions, whose numbers are still growing. It is not surprising, therefore, that in the last ten years hydroboration has become a tool of major importance to the synthetic organic chemist.

Table 1 - Reactions of Olefins with Diborane

Olefin	Example	Borane	Ref.
Terminal	1-pentene	R_3B	33
Disubstituted terminal	2-methyl-1-butene	R_3B	33
Disubstituted internal	2-pentene	R_3B	33
Trisubstituted	2-methyl-2-butene	R_2BH disiamyl borane	33
Tetrasubstituted	2,3 dimethyl-2-butene	RBH_2 thexyl borane	41

Table 2 - Reactions of Organoboranes

Borane	Reagent	Conditions	Product	Yield	Ref.
R_3B	1)O_2 2)H_2O_2/Me_3COH	Hexane solution 0°C	ROOH	40-80	52
R_3B	$C_2H_5CO_2H$	Diglyme 160°C 3 hr	RH	80-90	31
R_3B	$ClNH_2$ or $HO\cdot SO_3NH_2$	Diglyme 100°C 3 hr	RNH_2	38-48	53
$(RR'CH)_3B$	1)$H_2O_2/NaOH$ 2)Cr^{VI}[a]		RR'CO		54
R_3B	$\bar{C}H_2\overset{+}{S}Me_2$	DMSO at -10°C	RCH_2OH[b]	ca 30	55
R_3B	$CH_2=CHCHO$	THF 25°C 1 hr	RCH_2CH_2CHO[c]	ca 30	56
R_3B	$CH_2=CHCO\cdot CH_3$	THF 40°C 2 hr	$RCH_2CH_2CO\cdot CH_3$	ca 30	57
R_3B	$BrCH_2CO_2Et$	Bu^tOH, Bu^tOK 25°C	$RCH_2CO_2C_2H_5$	28-33	58
R_3B	Br_2CHCO_2Et	Bu^tOH, Bu^tOK (1 mol.) 0°C	$RBrCHCO_2Et$	ca 30	59
R'_3B	$RBrCHCO_2Et$	Bu^tOH, Bu^tOK 50°C	$RR'CHCO_2Et$	ca 20	59
R_3B	$AgNO_3/NaOH$	H_2O added 0°C	R-R	35-80	60
$(RCH_2)_3B$	$:CCl_2$[d]	Benzene 70°C 40 min	$RCH=CH_2CH_2R$[e]	58-68	61
R_3B	HgO	$H_2O/NaOH$ 80°C ½ hr	R_2Hg	40-90	62
$(ClCH_2\overset{R}{\overset{\vert}{C}}HCH_2)_3B$	H_2O/Cl^-	Diglyme 50°C	R	43-71	63
RR'R''B	$CO/HOCH_2CH_2OH$	Diglyme 125°C 1 hr	RR'R''COH[b]	80-90	64, 65
R_3B	CO/H_2O	Diglyme 100°C 2½ hr	R_2CO[b], R_2CHOH[f]	53-60	66
Thexyl BRR'	CO/H_2O 70 atmos	THF 50°C 3 hr	RR'CO	45-84	67
R_3B	$CO/LiAlH(OMe)_3$	THF 0°C 30 min	RCHO[g], RCH_2OH[h]	ca 30	68

a Direct oxidation of the organoborane with chromic acid can lead to rearranged products[54].

b After oxidation with alkaline hydrogen peroxide.

c After hydrolysis.

d From thermal decomposition of (bromodichloromethyl) phenyl mercury in the reaction mixture.

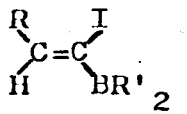
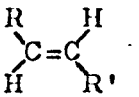
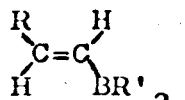
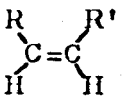
e Mixture of cis and trans olefins.

f Treatment with hot alkali prior to oxidation.

g Oxidation in the presence of a Na_2HPO_4/NaH_2PO_4 buffer.

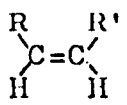
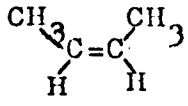
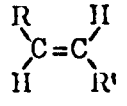
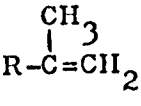
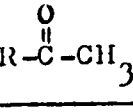
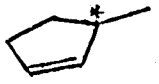
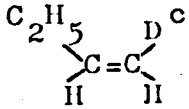
h After alkaline hydrolysis.

Table 3 - Reactions of Organoboranes

Unsaturated Compound	Hydro-borating Species	Intermediate	Subsequent Reagent	Product	Yield %	Ref.
$RCH=CH(CH_2)_nCH_3$ Internal olefin	B_2H_6	$(R(CH_2)_{n+3})_3B$ [a]	$R''CH=CH_2$	$R(CH_2)_{n+1}CH=CH_2$ Terminal olefin	80-90	69
Diene	Thexyl borane	 Bthexyl	H_2O/CO at 70 atmos.	 C=O	46-86	70
$RC\equiv CI$	R'_2BH		1) MeO^- 2) $H_2O_2/NaOH$ 1) MeO^- 2) $C_2H_5CO_2H$	RCH_2COR'  trans olefin	80-90	36
$RC\equiv CH$	R'_2BH		$NaOH/I_2$	 cis olefin	63-85	71

[a] After isomerisation of the initially formed organoborane.

Table 4 - Syntheses using Diisopinocampheyl Borane

Substrate	Product	Optical Purity, %	Remarks	Ref
	$RCH_2\overset{*}{C}H(OH)R'$ [a]	65-91		72
	$C_2H_5\overset{*}{C}H(NH_2)CH_3$ [b]	75		73
	$RCH_2\overset{*}{C}H(OH)R'$ [a]	13-22	Slow reaction: proceeds with elimination of pinene	74
	$R\overset{*}{C}H(CH_3)CH_2OH$ [a]	5-30		75
	$R\overset{*}{C}H(OH)CH_3$ [a]	11-30		76
 (racemic)		45	Olefin in excess; selective removal of one enantiomer	77
	$C_3H_7\overset{*}{C}H(OH)D$ [a]	56		78

[a] After oxidation with alkaline hydrogen peroxide.

[b] After amination with hydroxylamine-O-sulphonic acid.

[c] Prepared by the reaction path:

$$C_2H_5C{\equiv}CH \xrightarrow{C_2H_5MgBr} C_2H_5C{\equiv}CMgBr \xrightarrow{D_2O} C_2H_5C{\equiv}CD$$

$$C_2H_5C{\equiv}CD \xrightarrow[2)\ C_2H_5CO_2H]{1)\ \text{disiamyl borane}} \underset{Et}{\overset{H}{}}C{=}C\overset{H}{\underset{D}{}}$$

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Chapter 2

The Autoxidation of Carbon Compounds

This chapter covers the autoxidation of carbon compounds in some detail, because in the discussion section, a comparison is drawn between carbon and boron compounds under autoxidation conditions.

The discovery and development of rubber during the early part of the nineteenth century prompted research into the problem of ageing. Although it was some time before this was shown to be an oxidative process[1], considerable advances were made on an empirical basis with additives now known to be powerful antioxidants[2]. Natural rubber is a complex mixture, and Moureu and Douffraise[3] showed that the autoxidation of pure compounds was more likely to lead to an understanding of the mechanisms involved. Benzaldehyde was the first compound to be studied carefully[4,5] and it was shown by Bäckström[6,7] to autoxidise by a radical chain mechanism.

Many classes of organic compounds (e.g. hydrocarbons, aldehydes, ketones, ethers and amines) are attacked by atmospheric oxygen, and it is now known that the majority

of these follow the reaction scheme below[8-13]:

$$\text{Initiation} \quad RH \xrightarrow{r_i} R^{\cdot} \qquad \text{(i)}$$

$$\text{Propagation} \begin{cases} R^{\cdot} + O_2 \xrightarrow{k_2} RO_2^{\cdot} & \text{(ii)} \\ RO_2^{\cdot} + RH \xrightarrow{k_3} ROOH + R^{\cdot} & \text{(iii)} \end{cases}$$

$$\text{Termination} \begin{cases} 2R^{\cdot} \xrightarrow{k_4} & \text{(iv)} \\ R^{\cdot} + RO_2^{\cdot} \xrightarrow{k_5} & \text{(v)} \\ 2RO_2^{\cdot} \xrightarrow{k_6} & \text{(vi)} \end{cases} \Big\} \text{Non-radical products}$$

Termination by an inhibitor AH leads to two further equations:

$$\left.\begin{array}{ll} RO_2^{\cdot} + AH \xrightarrow{k_7} & \text{(vii)} \\ R^{\cdot} + AH \xrightarrow{k_8} & \text{(viii)} \end{array}\right\} \text{ inactive products}$$

The measurable quantity in most systems is the rate of disappearance of oxygen, $-\frac{d[O_2]}{dt}$ Using the stationary state approximation, it can be shown that, in the absence of inhibitor

$$-\frac{d[O_2]}{dt} = k_3 k_6^{-\frac{1}{2}} r_i^{\frac{1}{2}} [RH] \qquad \text{(ix)}$$

It is found that equation (ix) holds for a large number of hydrocarbons at oxygen pressures greater than 100 mm. (See Appendix for derivation of equations ix - xi.)

A more complex general expression can be derived for the case in which termination is brought about by an added inhibitor. Assuming that the inhibitor reacts by equation (vii) or (viii) exclusively (i.e. k_8 or k_7 is zero), the

expression simplifies to equation (x) or (xi) respectively:

$$-\frac{d[O_2]}{dt} = r_i k_3 k_7^{-1} [RH] [AH]^{-1} \qquad (x)$$

$$-\frac{d[O_2]}{dt} = r_i k_2 k_8^{-1} [O_2] [AH]^{-1} \qquad (xi)$$

a) Initiation:-

The main ways in which bond scission occurs to give radicals capable of starting autoxidation chains are (i) thermally induced homolysis; (ii) photolysis; (iii) redox reactions with metals of variable valence; (iv) induced decomposition of peroxides; (v) spontaneous initiation. The peroxidic products of autoxidation generally give rise to radicals, and as the reaction proceeds chain branching occurs. This autocatalytic effect gives rise to the familiar sigmoid curve observed in most autoxidations (Fig.1).

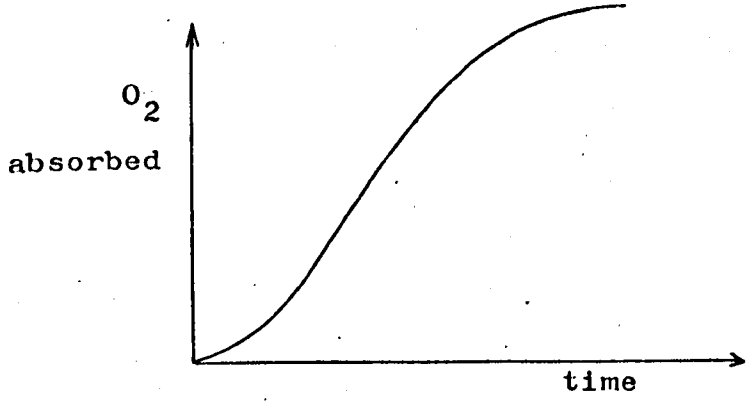


Fig.1

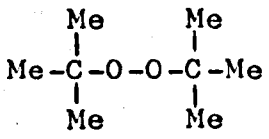
Auto-acceleration of Autoxidation.

i) Heat. Heat causes the homolysis of bonds, the ease of this depending inversely on the bond strength. This is

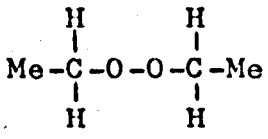
of great importance in the case of peroxides, since the O-O bond is relatively weak, and unimolecular fission occurs fairly readily (Equation xii). The stability depends on

$$ROOR' \longrightarrow RO^{\cdot} + R'O^{\cdot} \qquad \text{(xii)}$$

the groups attached to oxygen, and in general electron releasing groups strengthen the bond. Thus for di-t-butyl peroxide (I) the O-O bond energy is 37.5 k.cal./mole, whereas for diethyl peroxide (II) it is only 34.1 [14].



(I)



(II)

Hydroperoxides, although they can undergo unimolecular homolytic fission, are more susceptible to various induced decompositions described later.

ii) Light. Ultra-violet light is of sufficiently high energy to rupture many bonds, notably the O-O bond of peroxides, with the same consequences as thermal homolysis. Unsaturated compounds, particularly those containing the carbonyl structure, can often give a diradical (the triplet state). This is long-lived compared to the singlet state, and so has a greater chance of causing initiation (Equation xiii)

$$R_2C{=}O \underset{}{\overset{h\nu}{\rightleftharpoons}} R_2\dot{C}{-}\dot{O} \xrightarrow{R'H} R_2\dot{C}{-}OH + \dot{R}' \qquad \text{(xiii)}$$

Compounds in which the triplet state is particularly long-lived and easily formed, for example benzophenone, have

applications in other fields as photosensitisers.

It is difficult to avoid the presence of traces of carbonyl compounds in the original substrate, and since they are also formed during the decomposition of peroxides, they represent an important source of free radicals.

iii) Metal Ions. The decomposition of hydroperoxide is very markedly catalysed by metal ions of variable valence, according to equations (xiv) and (xv). If the metal is a powerful reducing agent, reaction (xiv) will predominate,

$$ROOH + M^{n+} \longrightarrow RO^{\cdot} + M^{(n+1)+} + OH^{-} \qquad (xiv)$$

$$ROOH + M^{(n+1)+} \longrightarrow ROO^{\cdot} + M^{n+} + H^{+} \qquad (xv)$$

and vice versa. In both these cases, nearly stoichiometric quantities of metal ion will be required to decompose the peroxide.

However, when the metal has two valence states of comparable stability, as have cobalt, copper and manganese for example, both (xiv) and (xv) occur with ease and a rapid chain reaction sets in. Since minute traces of these metals are effective, this is another major source of free radicals.

iv) Induced Decomposition of Hydroperoxides. In addition to homolytic fission, and occuring at a lower temperature,

three types of induced hydroperoxide decomposition can be distinguished, depending on whether a free radical, a solvent molecule or another hydroperoxide molecule is involved (Equations xvi - xviii).

Radical induced $R^{\cdot} + R'OOH \longrightarrow RH + R'OO^{\cdot}$ (xvi)

Solvent induced[15] $RH + R'OOH \longrightarrow \begin{bmatrix} R'O^{\cdot} \\ R^{\cdot} \end{bmatrix} + H_2O$ (xvii)

Hydroperoxide induced[16] $2ROOH \rightleftharpoons \begin{bmatrix} ROOH\cdots OOR \\ \quad\quad\quad\quad\quad\quad H \end{bmatrix} \longrightarrow ROO^{\cdot} + H_2O + {}^{\cdot}OR$ (xviii)

v) Spontaneous Initiation. In the absence of metals and products of autoxidation, none of the above modes of initiation can occur. However, a slow reaction with oxygen still takes place with carefully purified materials and Robb[17] has shown, in a study of indene and tetralin, that this is a termolecular process, which he formulates as Equation (xix) in agreement with Denisov[18].

$$2RH + O_2 \longrightarrow 2R^{\cdot} + H_2O_2 \qquad (xix)$$

b) Propagation:-

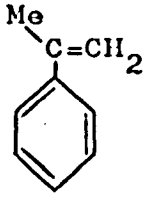
The steps by which chains are propagated are simple, and in the case of saturated hydrocarbons only reactions (ii) and (iii) can occur. With olefins, however, the alkylperoxy radical can add to the double bond as an alternative to reaction (iii), (Equation xx).

$$RO_2^{\cdot} + R'CH{=}CH_2 \longrightarrow ROOCH(R')-CH_2^{\cdot} \qquad (xx)$$

Hydrogen abstraction occurs preferentially at the α-methylenic position to give an allylic radical, and is strongly favoured by phenyl and alk-1-enyl substitution at this position[11,12,19]. This is because the resulting radical is stabilised by such resonance structures as (IV) and (V)

$$-CH{=}CH-\dot{C}H-CH{=}CH- \;(III) \longleftrightarrow -CH{=}CH-CH{=}CH-\dot{C}H- \;(IV) \leftrightarrow -\dot{C}H-CH{=}CH-CH{=}CH- \;(V)$$

Addition to the double bond is favoured in vinyl systems, or whenever there is no labile hydrogen atom. Thus 1-methyl styrene (VI) autoxidises predominantly in this way.



(VI)

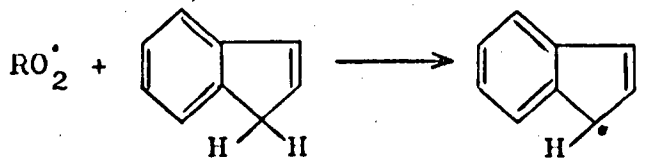
The polymer formed depends largely on oxygen pressure. In general, 1:1 copolymerisation occurs at high oxygen pressure (Equation xxi), whereas at lower pressures there is some competing homopolymerisation (Equation xxii).

$$ROOCH(R')-CH_2^{\cdot} + O_2 \longrightarrow ROOCH(R')-CH_2OO^{\cdot} \qquad (xxi)$$

$$ROOCH(R')-CH_2^{\cdot} + CH{=}CH_2(R') \longrightarrow ROOCH(R')-CH_2-CH(R')-CH_2^{\cdot} \qquad (xxii)$$

In certain cases epoxide may be formed[20].

Hydrogen abstraction and addition to the double bond may both occur in the same system. For instance, Russell[21] has shown that in the autoxidation of indene (VII) the kinetic chain length is about 430, while the copolymer formed is only about 10 units long. This suggests that a chain transfer reaction occurs, with the growing polymer abstracting from the highly active methylene group (Equation xxiii)



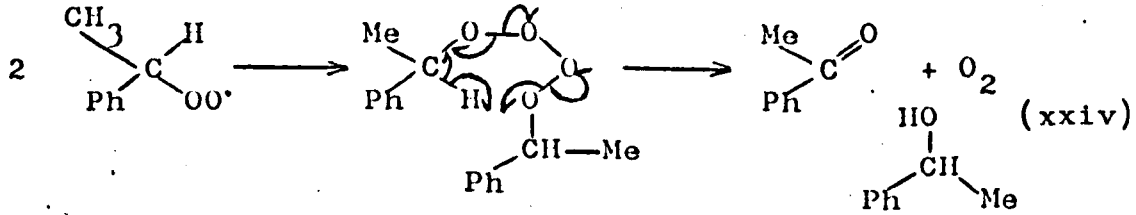
(VII) (xxiii)

c) Termination:-

The two propagating radicals, and therefore those most likely to be involved in termination processes, are R$^{\bullet}$ and $RO_2^{\bullet}$. However, k_2 is usually greater than k_3 by a factor of 10^6-10^8(ref.9), and so at high oxygen pressures the standing concentration of alkyl radicals is negligible, and termination is entirely between peroxy radicals (Equation vi).

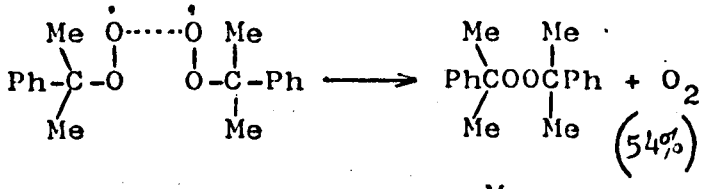
The nature of the products has been examined by Russell[22] who concludes from deuterium isotope studies that for primary and secondary peroxy radicals a cyclic transition state is

involved (Equation xxiv)

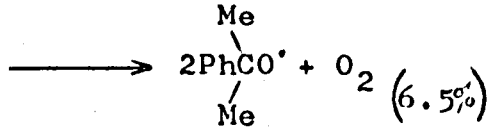


(xxiv)

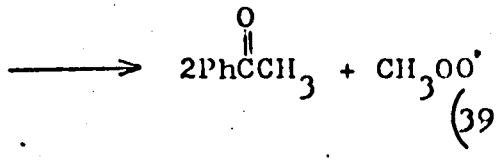
Cumylperoxy radicals, with no α-hydrogen atom, interact according to Equations (xxv) - (xxvii)[23],



(xxv)



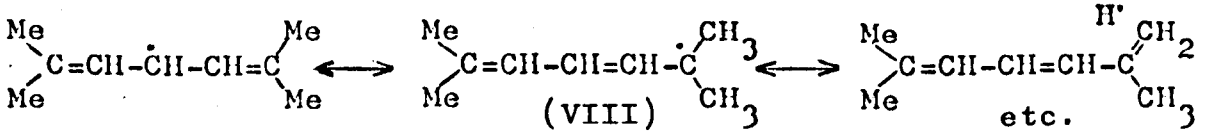
(xxvi)



(xxvii)

α-Methyl styrene and cumene hydroperoxide are also produced from cumylperoxy radicals.

As the oxygen pressure is lowered, reactions (iv) and (v) become important termination pathways. The pressure above which the autoxidation rate is oxygen pressure independent varies with the stability of the radical R$^{\cdot}$. While for 1-hexadecene it is less than 1 mm, for 2,6 dimethyl-hepta 2,5 diene, which gives rise to the resonance stabilised radical (VIII), reaction (v) still accounts for 50% of the termination at 800 mm[24].



(VIII)

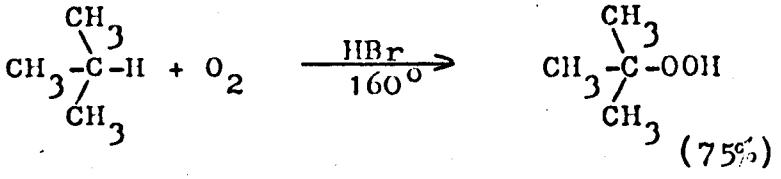
The products of reactions (iv) and (v) are much simpler than those due to combination of peroxy radicals:

$$ROO^{\cdot} + R^{\cdot} \longrightarrow ROOR \qquad \text{(xxviii)}$$

$$R^{\cdot} + R^{\cdot} \longrightarrow R_2 \qquad \text{(xxix)}$$

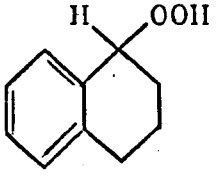
d) End Products of Autoxidation:-

In principle it should be possible to isolate the hydroperoxide, ROOH. However, the C-H bonds are relatively strong in saturated systems, and high temperatures are required to maintain the chains. Under these conditions the hydroperoxide decomposes, but at low conversions in special cases these compounds have been isolated[25]. The industrial preparation of t-butyl hydroperoxide from isobutane utilises hydrogen bromide as a catalyst[26].

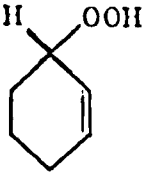


(75%) (xxx)

Olefins tend to give higher yields of hydroperoxides than saturated hydrocarbons. The identification of tetralin hydroperoxide[27] (IX) and cyclohexene peroxide[28,29] (X) were milestones in the understanding of the autoxidation mechanism.



(IX)



(X)

Where polymerisation occurs, the products have usually been isolated, except where electron-withdrawing substituents are attached α to the peroxide link[8,30,31].

The products of termination reactions between alkylperoxy radicals may predominate in conditions where the kinetic chain length is very short. Usually, however, the main sources of the observed end products are alkoxy radicals produced by decomposition of the first formed hydroperoxides. These can undergo many further reactions as shown in Equations (xxxi) - (xxxvi).

Abstraction

$$RO^{\cdot} + R'H(R'OOH) \longrightarrow ROH + R'^{\cdot}(R'OO^{\cdot}) \qquad \text{(xxxi)}$$

Disproportionation

$$2\ \begin{matrix} R \\ | \\ R'-C-O^{\cdot} \\ | \\ H \end{matrix} \longrightarrow \begin{matrix} R \\ | \\ R'-C=O \end{matrix} + \begin{matrix} R \\ | \\ R'-C-OH \\ | \\ H \end{matrix} \qquad \text{(xxxii)}$$

Elimination

$$\begin{matrix} R \\ | \\ R'-C-O^{\cdot} \\ | \\ R'' \end{matrix} \longrightarrow \begin{matrix} R' \\ | \\ C=O \\ | \\ R'' \end{matrix} + R^{\cdot} \qquad \text{(xxxiii)}$$

Association

$$RO^{\cdot} + R'O^{\cdot}(R') \longrightarrow ROOR'(ROR') \qquad \text{(xxxiv)}$$

Rearrangement and Dimerisation

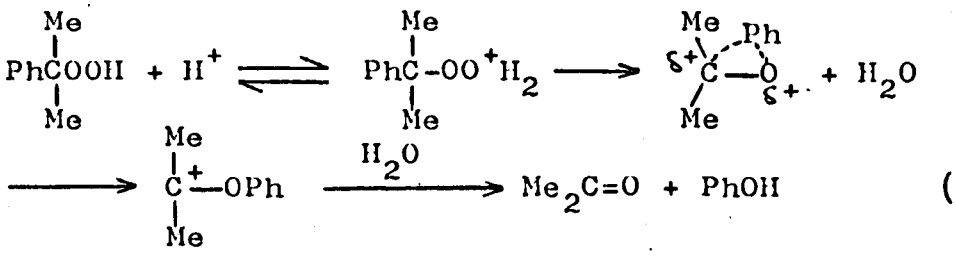
$$\begin{matrix} Ar \\ | \\ Ar'-C-O^{\cdot} \\ | \\ Ar'' \end{matrix} \longrightarrow \begin{matrix} Ar \\ | \\ C-OAr'' \\ | \\ Ar \end{matrix} \longrightarrow \begin{matrix} OAr'' & OAr'' \\ | & | \\ Ar-C & C-Ar \\ | & | \\ Ar' & Ar' \end{matrix} \qquad \text{(xxxv)}$$

Addition

$$RO^{\cdot} + \begin{matrix} | \\ -C=C- \\ | \end{matrix} \longrightarrow \begin{matrix} | \\ RO-C-C^{\cdot} \\ | \end{matrix} \qquad \text{(xxxvi)}$$

In addition, acid or base catalysed decomposition of peroxides leads to carbonyl compounds and alcohols by non-radical routes[8]. One of commercial importance is the acid

catalysed intramolecular decomposition of cumene hydroperoxide to phenol and acetone[32]:



(xxxvii)

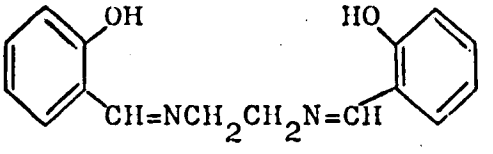
A single hydroperoxide can thus give rise to a wide variety of decomposition products. Since in many cases resonance stabilisation of the alkyl radical produces several different hydroperoxides, it is not surprising that the autoxidation of a pure substance results in a complex mixture of acids, alcohols, carbonyl compounds etc., which disguises the basic simplicity of the mechanism.

e) Antioxidants:-

Although in some circumstances, for example the drying of paints, atmospheric oxidation is a desirable thing, in the vast majority of cases it results in a deterioration of the starting material. Examples are rancidity in fats, due to the formation of acids and aldehydes; loss of colour in materials exposed to sunlight and the perishing of rubber. There are two main ways in which these effects can be suppressed - by removing the species responsible for the initiation of the chains, and by removing the chain carrying radicals.

i) Initiation Suppressors. Three different types of anti-oxidant are distinguished in this class - metal ion deactivators, ultra-violet light stabilisers and peroxide destroyers.

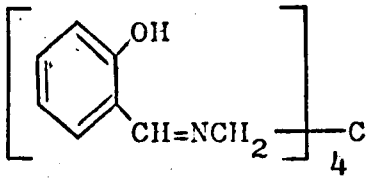
Metal ion deactivators operate either by complexing the ion to its maximum co-ordination number, or by altering the redox potential so that one valence state is preferred above the others. Frequently a chelate ligand which deactivates one metal will activate another. N,N' (disalicylidene) ethylene diamine (XI) is an effective inhibitor for copper,



(XI)

but the cobalt complex is one of the most powerful pro-oxidants known[33]. The reason for this type of behaviour lies in the different maximum co-ordination numbers of metals. A quadridentate ligand attached to a metal of co-ordination number six leaves possible points of attack by hydroperoxide, and the inhibitor is ineffective. In many cases the redox potential is actually rendered more favourable for the catalysis by incomplete co-ordination, as with the cobalt complex of (XI). One ligand which was found to be particularly effective for manganese, iron, cobalt, nickel and copper was the octadentate NN'N"N"' tetrasalicylidenetetra-

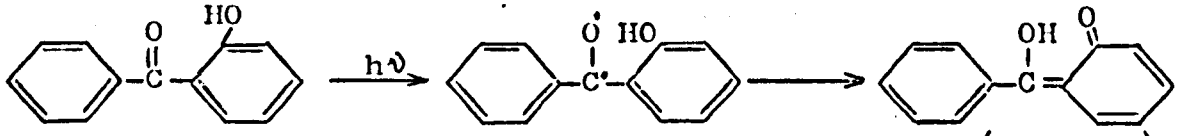
(aminomethyl) methane (XII)[33] which is presumably capable of



(XII)

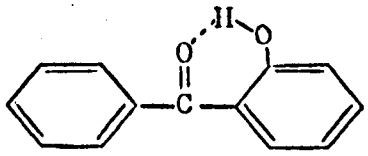
complexing all of them up to their maximum co-ordination number.

Ultra-violet light deactivators function by absorbing the radiation and disposing of it by non-radical generating processes. The most efficient stabiliser of all is carbon black, but unfortunately, like many other good ones, its colour renders it unsuitable for many purposes. Among non-coloured compounds, o-hydroxy benzophenones are in very wide use[34] in spite of the fact that benzophenone itself is a powerful photosensitiser. It is thought that hydrogen bonding is an important factor in getting rid of the acquired energy, possibly by hydrogen transfer as in equation (xxxviii)[35]. Evidence for the importance of hydrogen



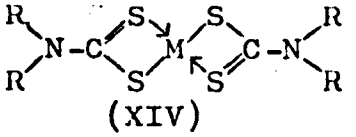
(xxxviii)

bonding comes from proton magnetic resonance studies on a number of hydroxy benzophenones. These show quite clearly that the stabilising efficiency is related to the degree of hydrogen bonding in the molecule[36].

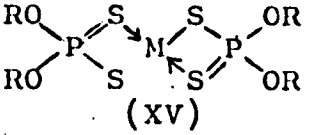


(XIII)

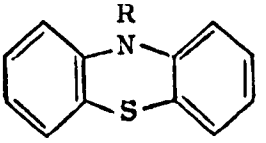
Hydroperoxide decomposers are mainly compounds of sulphur and phosphorus. Sulphides, metal dialkyldithiocarbamates (XIV) dialkyldithiophosphates(XV) and phenothiazine derivatives (XVI) are examples in common use[37].



(XIV)

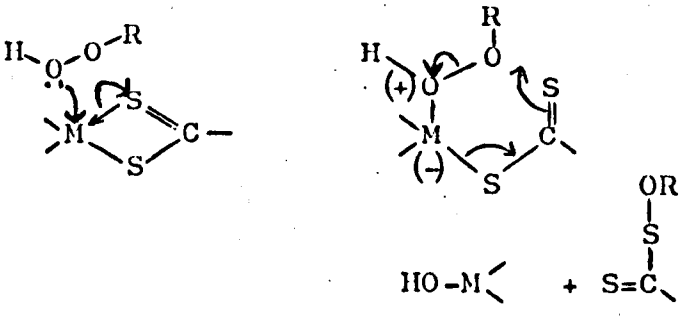


(XV)



(XVI)

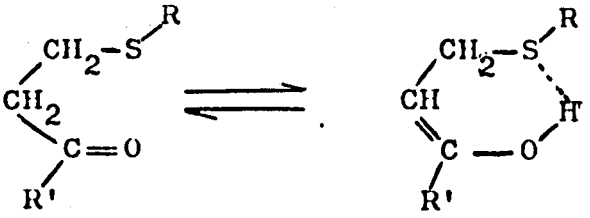
The mechanisms by which these compounds operate are complex. Dialkyldithiophosphates (XV), for instance, not only decompose hydroperoxides catalytically[38], but also act as propagation suppressors by removing the peroxy radical[39,40]. The same is true of phenothiazine, (XVI, R=H), which has been shown to catalyse the decomposition of cumene hydroperoxide by an ionic pathway[38], as well as acting as an amine antioxidant. It has been suggested that copper (II) di-n-butyl dithiocarbamate (XIV, R=butyl, M=Cu) acts according to the scheme below[41]:



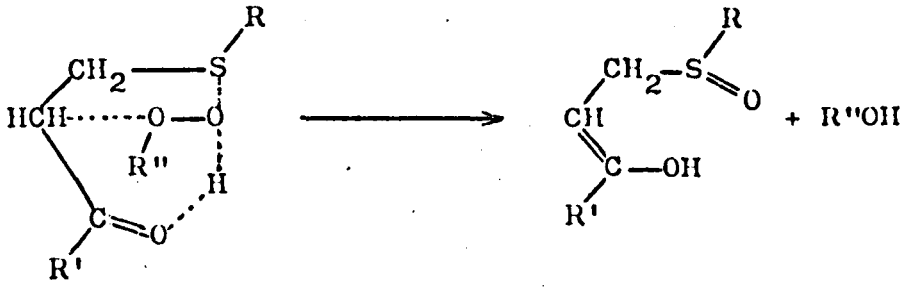
(xxxix)

Simple sulphides appear to decompose hydroperoxides

stoichiometrically, with the formation of sulphones and sulphoxides which are often themselves active[42]. The activity of keto sulphides depends very strongly on the length of the hydrocarbon chain, and it was suggested by Thompson[43] that the maximum activity of the β-keto sulphides was due to the hydrogen-bonded stabilisation of the enol form (XVII) which could then react with hydroperoxide (Equation xl).



(XVII)



(xl)

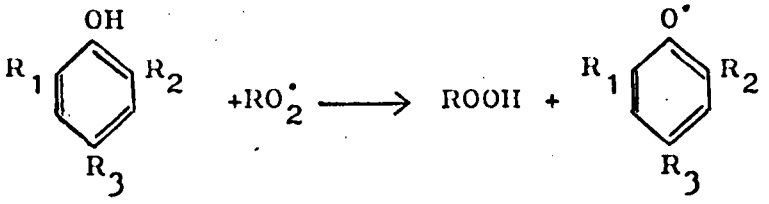
ii) Propagation Suppressors. Antioxidants in this class react directly with propagating radicals to give new radicals of much lower reactivity, or molecular products. Since alkyl radicals are present to a negligible extent in typical systems, most propagation suppressors are peroxy radical trappers.

1) Phenols and Amines. These operate by removing peroxy

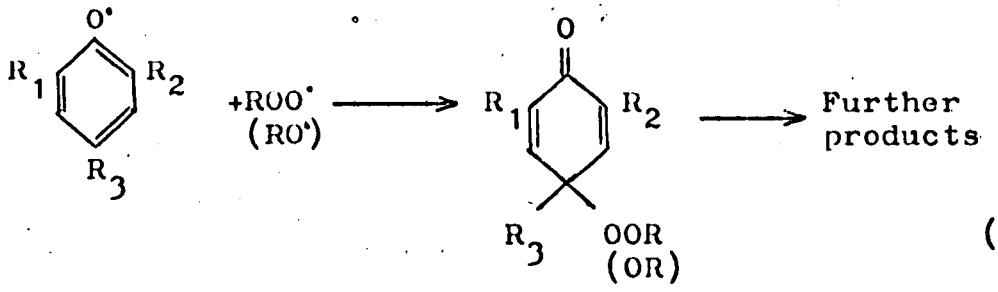
radicals by an alternative route to reaction (iii).

Two mechanisms have been suggested in the case of phenols.

The first, based on kinetic data[44-46], on the identification of end products[47,48] and on electron spin resonance spectra[49] is shown below:



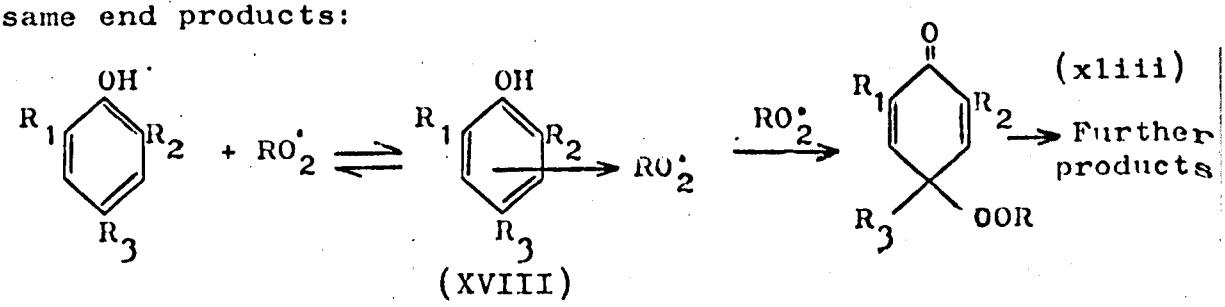
(xli)



(xlii)

The rate-determining step is the hydrogen abstraction by the peroxy radical.

The second is due to Boozer and Hammond[50], and postulates the reversible formation of a complex, followed by a rate-determining reaction with a second peroxy radical to give the same end products:



(XVIII)

(xliii)

Russell[51] has shown that complexes such as (XVIII) between free radicals and π-electron systems frequently occur. The main evidence for this mechanism lies in the absence of any kinetic isotope effect, which would be expected if hydrogen abstraction were rate-determining. However, Bickel and Kooyman[52] have pointed out that kinetic isotope effects are small for reactions of such low activation energy as (xli), and in support of this Ingold[53] has found a small isotope effect for some weak inhibitors, where reaction (xli) is not so facile. In a later paper[54] Ingold argues in favour of the second mechanism, and it is possible that either or both of these mechanisms may operate under specific reaction conditions.

The antioxidant activity of a given phenol, although somewhat dependent on the substrate, is very markedly influenced by the ring substituents. Electron releasing groups in the ortho and para positions increase the activity, whereas electron withdrawing groups in these positions decrease it. Boozer and Hammond[50] determined antioxidant activities relative to the parent compounds phenol, N-methylaniline and diphenylamine, using tetralin and cumene as substrates. They found that a fair correlation was obtained between log (relative efficiency) and the Hammett σ constants. Ingold[45] obtained a much better correlation,

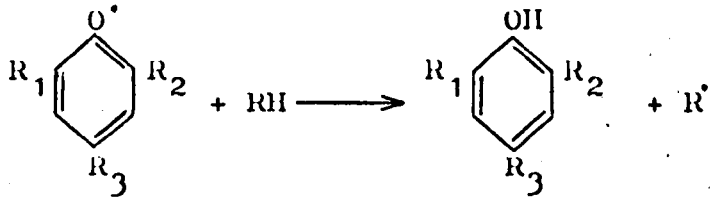
using σ^+ constants, for the reaction between t-butoxy radicals and substituted phenols. This system possesses the advantage of being free from complicating transfer reactions.

The effect of an electron releasing group in the ortho or para position is to weaken the O-H bond, facilitating hydrogen abstraction by the peroxy radical. This was demonstrated by deuterium isotope studies by Ingold and Puddington[53]. They showed that the isotope effect was smaller for more efficient antioxidants, indicating that the hydrogen atom was more labile in these compounds.

Oxidation-reduction potentials provide a better measure of the O-H bond strength, and Bolland and ten Have[46] showed the existence of a linear relationship between the normal oxidation-reduction potential and log (relative efficiency) for a number of phenols. A study by Penketh[55] of polarographic oxidation potentials revealed that compounds having a value greater than 0.80 volts (i.e. those with strong O-H bonds) were generally ineffective, whereas all compounds known to be good inhibitors had values below 0.70 volts.

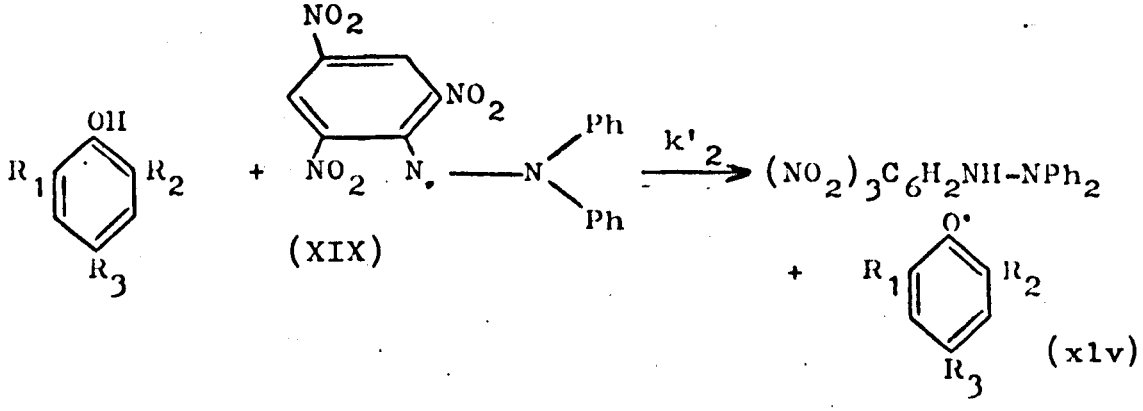
However, compounds with low values were not necessarily good antioxidants. This fact, as well as the deviants from the Hammett plot, can be accounted for in two ways. Firstly,

a very weak O-H bond may be susceptible to atmospheric oxidation, in which case it will actually initiate chains. Secondly, a lower redox potential may increase the tendency to chain transfer, in which the phenoxy radical abstracts from the substrate to start another chain:



(xliv)

This reaction is discouraged by a bulky ortho substituent, such as t-butyl, which allows approach of the alkylperoxy radical to the phenol, but prevents the phenoxy radical so formed from attacking the labile hydrogen atom of the substrate. The reaction between diphenylpicrylhydrazyl, DPPH, (XIX) and phenols, investigated by McGowan[56] (Equation xlv) brings out this point.

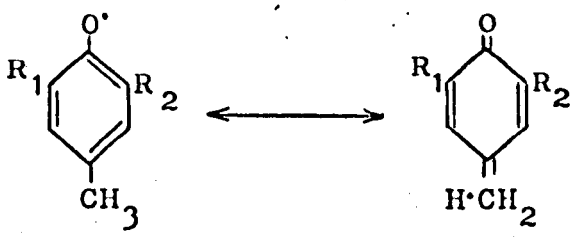


(xlv)

A linear relationship was found between log k'_2, the second order rate constant, and the relative oxidation potentials, for those phenols without bulky ortho substituents; those

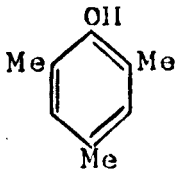
which had, gave anomalously low values for k'_2.

Conversely, α-branching in the group para to the phenolic OH decreases antioxidant activity, and this is attributed to loss of hyperconjugative stabilisation of the phenoxy radical due to structures such as (XX).



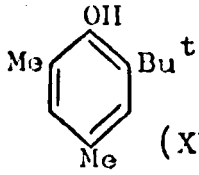
(XX)

The choice of antioxidant for a given system is a complex one, and accounts for the large number of commercially available compounds. To illustrate the confusing overall effect of substitution on activity, the six possible 2,4,6 methyl/t-butyl phenols are shown (XXI) - (XXVI), together with the relative values of their antioxidant activity in gasoline[8]. The most active is (XXII) in which the para-methyl gives the greatest hyperconjugative stability to the radical, and the ortho t-butyl group reduces the possibility of chain transfer. Two bulky ortho substituents shield the phenolic OH to such an extent that its reactivity towards peroxy radicals is also reduced.



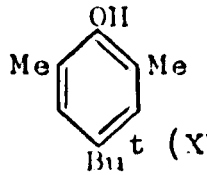
(XXI)

(118)



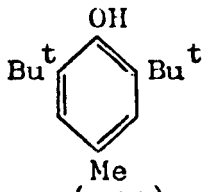
(XXII)

(170)



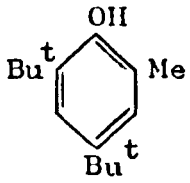
(XXIII)

(46)



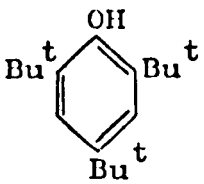
(XXIV)

(100)



(XXV)

(69)



(XXVI)

(36)

Amine antioxidants have not been studied to the same extent as phenols, but they appear to function by similar mechanisms[8,57]. However, the stoichiometries are usually higher, in the region of three chains destroyed per molecule, as opposed to two per phenolic nucleus[37]. Thomas[58] has shown that nitroxides are produced during inhibition by amines, and has accounted for these by the schemes:

$$Ar_2NH + RO_2^{\cdot} \longrightarrow Ar_2N^{\cdot} + ROOH \qquad \text{(xlvi)}$$

$$Ar_2N^{\cdot} + RO_2^{\cdot} \longrightarrow Ar_2NO^{\cdot} + RO^{\cdot} \qquad \text{(xlvii)}$$

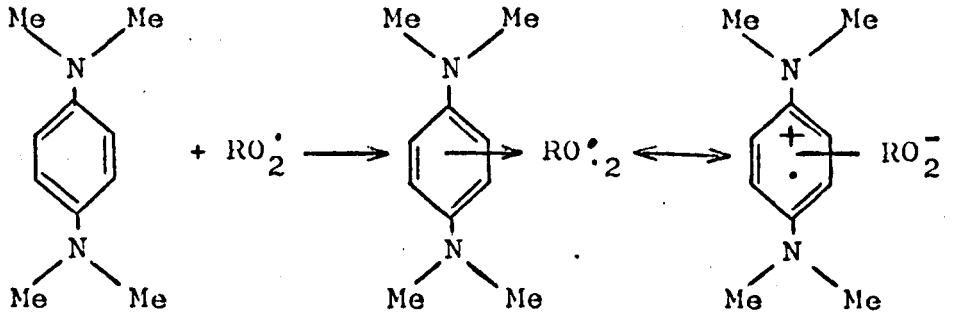
Nitroxides are themselves powerful inhibitors, explaining the high stoichiometry.

The results of Thomas's kinetic study appeared to favour the Boozer and Hammond[50] mechanism, involving initial complex formation, but Ingold[57] has found a significant kinetic isotope effect, indicating the importance of hydrogen abstraction, and has pointed out that Thomas's results are equally well explained by taking into account the reverse of the inhibition reaction (Equation xlviii).

$$A^{\cdot} + ROOH \longrightarrow AH + RO_2^{\cdot} \qquad \text{(xlviii)}$$

However, some fully alkylated amines are efficient

inhibitors[59], and these can only function by electron transfer.



(ii)

(XXVII)

Although the methyl groups would help to stabilise a structure such as (XXVII) by hyperconjugation, Pedersen [59] has suggested that their major function is steric protection of the nitrogen atoms.

2) Stable Free Radicals. Stable free radicals, for example DPPH, (XIX), and nitric oxide, have been used for a long time to inhibit radical polymerisations. However, their effectiveness is limited when oxygen is present.

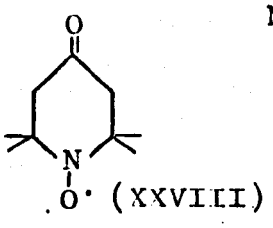
It was shown by Thomas[58,60] that when secondary aromatic amines are used as inhibitors, nitroxide radicals are produced. On the basis of kinetic evidence he proposed the existence of concurrent reactions leading to the catalytic and stoichiometric decomposition of peroxy radicals by diphenyl nitroxide:

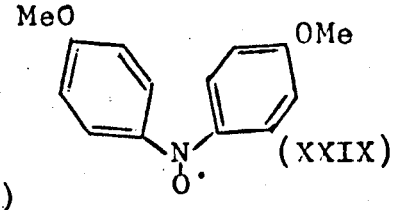
$$Ph_2NO^{\cdot} + RO_2^{\cdot} \longrightarrow \text{olefin} + O_2 + Ph_2NOH \qquad (1)$$

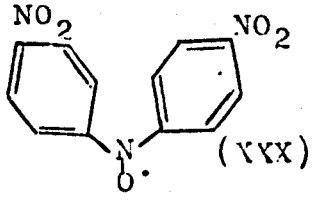
$$Ph_2NOH + RO_2^{\cdot} \longrightarrow Ph_2NO^{\cdot} + ROOH \qquad (ii)$$

$$Ph_2NO^{\cdot} + RO_2^{\cdot} \longrightarrow \text{amine oxide derivative} \qquad (iii)$$

A large number of nitroxides have been prepared in the last ten years, mainly by Russian workers[61-63]. They are stabilised either by electron delocalisation, as in diphenyl nitroxide, or by steric hindrance of the nitrogen atom, preventing dimerisation. Thus all the aliphatic ones which have been isolated have two tertiary carbon atoms attached to the nitrogen. The simplest member of this series, di-t-butyl nitroxide, was foundto be an efficient scavenger for cumyl radicals produced by the thermal decomposition of azocumene[64]. The aliphatic nitroxide 2,2,6,6 tetramethyl piperid-4-one 1-oxyl (XXVIII), used by Brownlie and Ingold[65] to inhibit the autoxidation of styrene, only reacted with alkyl radicals, whereas the aromatic nitroxides (XXIX) and (XXX) attacked both $R^{\cdot}$ and $RO_2^{\cdot}$ radicals.


(XXVIII)

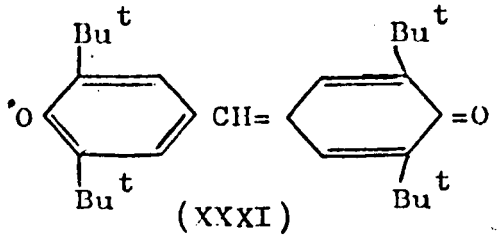

(XXIX)


(XXX)

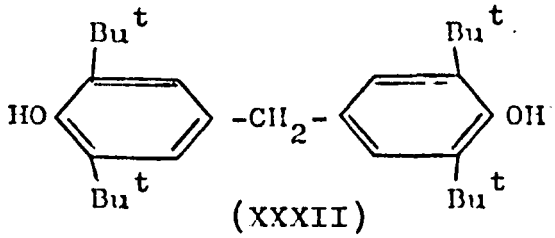
The authors concluded that nitroxides were not such efficient antioxidants as phenols and amines, and were unlikely to be of much commercial use.

Another stable free radical relevant to this work is galvinoxyl (XXXI) which is derived from the commercial

antioxidant Ethyl 702 (XXXII) by oxidation with lead dioxide.



(XXXI)

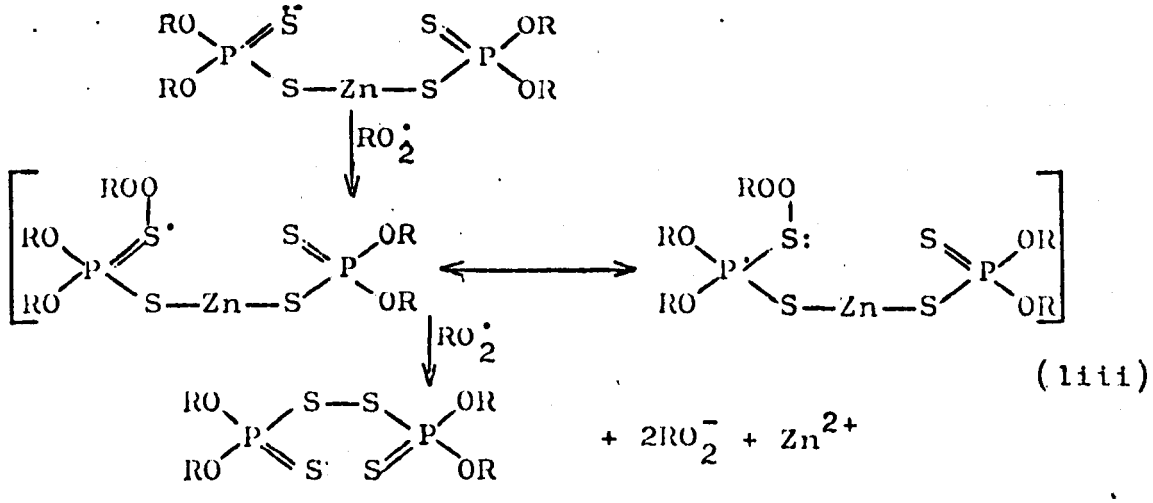


(XXXII)

It was found by Bartlett[66] to be a very efficient scavenger for both the cyanoisopropyl and t-butoxy radicals.

3) Compounds containing Sulphur and/or Phosphorus. The action of certain sulphur and phosphorus containing anti-oxidants involves propagation suppression as well as hydroperoxide decomposition. Thus it has been shown by Bateman et al[67] and Cunneen et al[68] that sulphoxides, derived from sulphides, act as peroxy radical trappers.

Dialkyldithiophosphates have also been shown to react with peroxy radicals, and Burn[40] has proposed the following reaction scheme:

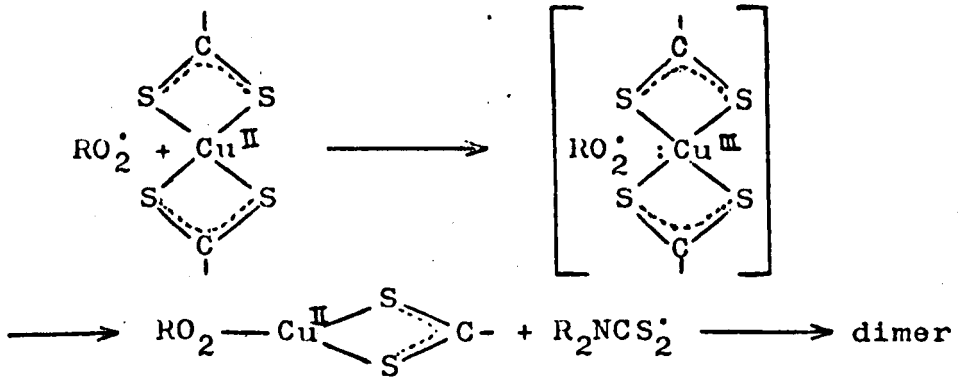


(liii)

The very great activity of the copper (I) salt was attributed to its ability to transfer an electron to the peroxy radical, providing another route for its removal:

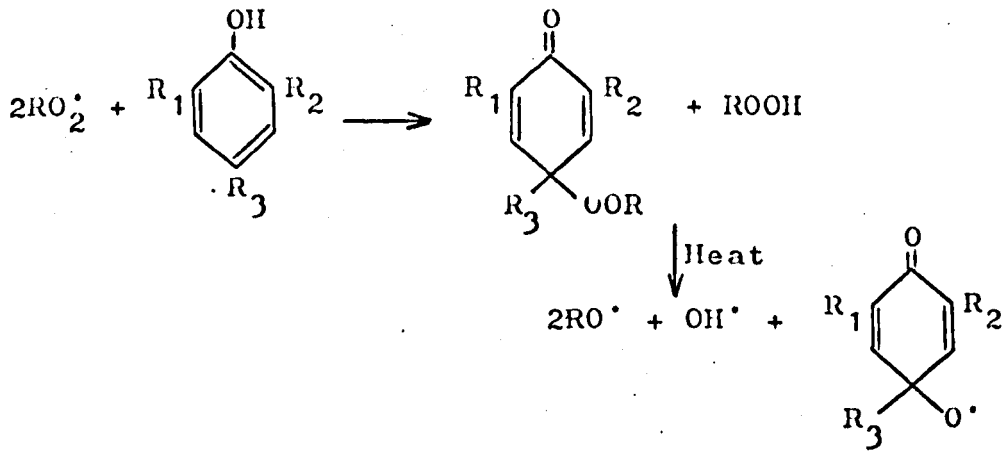
$$Cu^{+} + RO_2^{\bullet} \longrightarrow Cu^{2+} + RO_2^{-} \qquad (1iv)$$

Di-N-alkyldithiocarbamates also remove peroxy radicals, possibly by a mechanism similar to Equation (liii) above. The copper (II) salt is unusually active, and it has been suggested that the reason for this lies in the ability of the dialkyldithiocarbamate anion to stabilise copper (III)[41].



(1v)

iii) Synergism. No single class of antioxidant can give complete protection under all conditions. Thus a metal ion deactivator will prevent one form of chain initiation, but the action of heat or ultra-violet light may produce radicals against which it is powerless. Similarly, phenolic antioxidants, while very efficient at low temperatures, give rise to peroxides which themselves are a source of free radicals when heated:

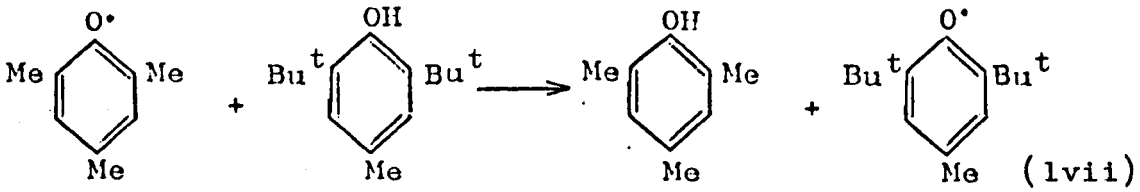


(1vi)

The addition of a peroxide decomposer in the second case may produce an effect which is greater than that of either of the antioxidants alone. This phenomenon is known as synergism. Two types can be distinguished: in the first, as above, antioxidants from two different classes mutually reinforce one another; whereas in the second, an efficient antioxidant is backed up by another compound which may not necessarily have any activity on its own.

The first type is self-explanatory - for maximum effect, a propagation suppressor is normally used with an initiation suppressor. The second kind is often useful industrially, where a large amount of an expensive antioxidant can be replaced by a smaller quantity, which when augmented by a cheaper compound gives the same activity. Examples exist in phenol systems, and it has been suggested[8] that the more efficient antioxidant, which would normally give rise to

chain transfer, is instead regenerated by reaction with the second, and less efficient.



(lvii)

In support of this, it is found that the best results are obtained with mixtures which are complementary with regard to bulky ortho substituents. The synergism observed between bisphenols and dialkyl phosphites has also been attributed to regeneration of the phenol[69].

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Chapter 3

The Autoxidation of Boron Compounds

a) Early History:-

It has been known for over one hundred years that simple alkyl boron compounds take up atmospheric oxygen very readily. Thus in 1862, Frankland[1] reported that triethyl and trimethyl boranes, prepared by the action of triethyl borate on the appropriate zinc alkyls, were spontaneously inflammable. By more controlled oxidation he was able to prepare several boronic esters and acids[1,2]:

$$Et_3B + O_2 \longrightarrow EtB(OEt)_2 \xrightarrow{H_2O} EtB(OH)_2 \qquad (i)$$

This technique was later employed by Krause and co-workers[3,4]. Trialkyl boranes, prepared from the action of boron trifluoride on the Grignard reagent, were allowed to stand in the air for some days in a loosely stoppered bottle. At the end of this time the product was recrystallised from water, yielding the boronic acid (Equation i).

Johnson et al[5-7], in 1938, carried out a careful study on the autoxidation of boranes and boronic acids. Dry n-butylboronic acid was slowly converted to boric acid and n-butanol on exposure to air over a period of five days (Equation ii), whereas the moist acid took up no oxygen in eighteen months[6].

$$Bu^nB(OH)_2 + \tfrac{1}{2}O_2 \longrightarrow \left[Bu^nOB(OH)_2\right] \xrightarrow{H_2O} Bu^nOH + B(OH)_3 \qquad \text{(ii)}$$

Water also had a marked effect on the autoxidation of tri-n-butyl borane[5]. When dry, this compound took up 0.97 mol. oxygen to give pure di-n-butyl n-butylboronate (Equation iii)

$$Bu^n{}_3B + O_2 \longrightarrow Bu^nB(OBu^n)_2 \qquad \text{(iii)}$$

However, in the presence of excess water, only half the quantity of oxygen was absorbed, to give n-butyl di-n-butylborinate in high yield (Equation iv)

$$2Bu^n{}_3B + O_2 \xrightarrow{H_2O} 2Bu^n{}_2BOBu^n \qquad \text{(iv)}$$

b) Organoperoxyboranes:-

The mechanism put forward by Johnson and van Campen[5] to account for the autoxidation product of tri-n-butyl borane involved an intermediate "borine-peroxide" (I).

$$R_3B + O_2 \longrightarrow R_3B{-}O{=}O \quad \text{(I)} \qquad \text{(v)}$$

$$R_3B{-}O_2 + R_3B \longrightarrow 2R_2BOR \qquad \text{(vi)}$$

Evidence for peroxidic intermediates was found by Grummitt[8] in the autoxidation of neat liquid trialkylboroxines. The partially autoxidised liquids liberated iodine from acidified potassium iodide and initiated the polymerisation of vinyl acetate, reactions characteristic of peroxides. However, it was another fourteen years before a peroxide was positively identified as an autoxidation product. The reaction of trimethyl borane with oxygen at low pressures in a gaseous flow system gave a 1:1 liquid adduct, Me_3BO_2[9].

This had previously been obtained by Bamford and Newitt[10], who formulated it as dimethyl methylboronate, $MeB(OMe)_2$. Petry and Verhoek[9] showed iodometrically, by mass spectrometry and by other physical measurements that it was in fact methylperoxy dimethylborinate, Me_2BOOMe.

$$Me_3B + O_2 \longrightarrow Me_2BOOMe \qquad \text{(vii)}$$

About the same time, and in subsequent publications, Davies and co-workers reported on the preparation of organoperoxyboranes[11-18]. These are available by two routes, nucleophilic substitution and autoxidation.

The reaction of t-butyl hydroperoxide with boron trichloride at room temperature gave tri-t-butylperoxyborate[11] (Equation viii). This was a colourless liquid monomeric

$$3Bu^tOOH + BCl_3 \longrightarrow (Bu^tOO)_3B + 3HCl \qquad \text{(viii)}$$

in boiling benzene. It was stable for months at room temperature, and for short periods at 100°, but was rapidly hydrolysed by water to give the original hydroperoxide and boric acid. Other triperoxy borates, as well as mono-(II) and di-(III) peroxyborates have also been prepared by nucleophilic substitution.

$ROOB(OR')_2$ (II) $\qquad\qquad$ $(ROO)_2BOR'$ (III)

Peroxyborinates (IV), peroxyboronates (V) and

peroxyboronic anhydrides (VI) have been prepared by autoxidation of dilute solutions of the appropriate borane

$ROOBR_2$ (IV) $(ROO)_2BR$ (V) $(ROOBO)_3$ (VI)

in an inert solvent[12-17]. For instance, tri-s-butyl borane in ether, stirred under oxygen in a gas burette, rapidly absorbed two molecules of gas to form the peroxyboronate[13]:

$$Bu^s{}_3B + 2O_2 \longrightarrow (Bu^sOO)_2BBu^s \qquad (ix)$$

By restricting the oxygen supply during the autoxidation of diisobutyl-t-butyl borane in ether, the monoperoxide, (IV), was formed[15]. It was not isolated, but the product after absorbtion of 1.08 mol. oxygen and oxidation with peroxy-octanoic acid yielded 1.08 mol. hydroperoxide and 1.92 mol. alcohol (Equation x).

$$ROOBR_2 + 2C_7H_{15}CO_3H \longrightarrow ROOB(OR)_2 + 2C_7H_{15}CO_2H$$

$$\xrightarrow{H_2O} ROOH + 2ROH + B(OH)_3 \qquad (x)$$

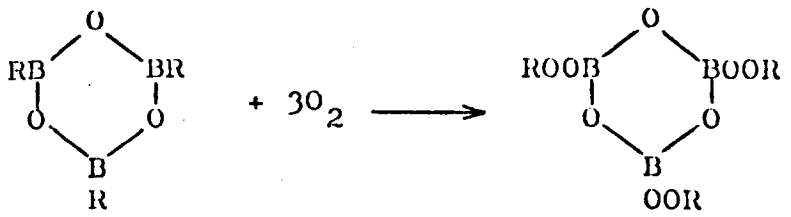
The diperoxyboronates (V) have been isolated in a number of cases[12-15]. Like the peroxyborates they are readily hydrolysable oils, but they are much more unstable, and quite rapidly rearrange to give monoperoxyborates (II):

$$(ROO)_2BR \longrightarrow ROOB(OR)_2 \qquad (xi)$$

This reaction is reviewed later (Section e).

The reaction of oxygen with dilute solutions of

trialkylboroxines gives peroxides of type VI (Equation xii).



(xii)

but these have only been detected iodometrically[17].

c) Rate of Autoxidation:-

The relative rates of autoxidation of different organoboranes have been reported by various authors and are summarised in Table 5. The order of reactivity in the last row is due to steric factors. Trimesitylborane is completely inert to autoxidation, and tri-α-naphthylborane only absorbed 0.5 mol. oxygen in 50 days[19]. Triphenylborane absorbed the same amount of oxygen in one hour at 25°C. Ramsey and Leffler[22] have recently drawn attention to the great ease of oxidation of triphenylborane, which is contrary to the earlier impression given in reviews and textbooks. In this reference is a claim to the first correct UV spectrum of triphenylborane, previous spectra[23,24] having been contaminated by oxidation products.

d) Mechanism of Autoxidation:-

Johnson and van Campen[5] were the first to attempt an explanation of the formation of esters during the autoxidation

of trialkylboranes. (Equations (v) and (vi).

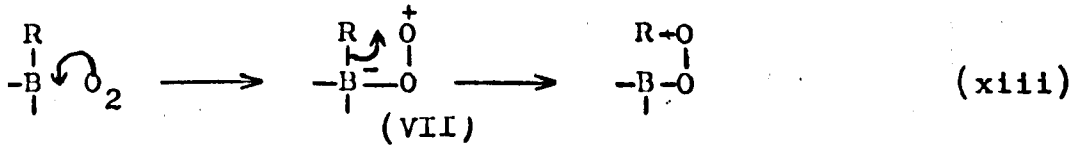
$$R_3B + O_2 \longrightarrow \underset{(I)}{R_3B\text{-}O\text{=}O} \qquad (v)$$

$$R_3B\text{-}O_2 + R_3B \longrightarrow 2R_2BOR \qquad (vi)$$

Evidence for a peroxidic intermediate, presumed to be (I), was obtained by Grummitt[8] in the autoxidation of neat liquid trialkylboroxines. Although the reaction was inhibited by 0.1% phenyl-β-naphthylamine, and was compared by Grummitt to the autoxidation of benzaldehyde, (known to proceed by a radical chain mechanism[25],) he favoured the scheme postulated by Johnson[5].

The idea of a chain mechanism in borane oxidations was first put forward by Bamford and Newitt[10], in their classical kinetic study of the gas phase oxidation of trimethyl and tri-n-propylboranes. The detailed kinetics of the reaction for trimethylborane were interpreted as a chain mechanism beginning and ending on the walls of the vessel; however it was not possible to decide whether this was of a radical or a polar nature. In either case, the rate of such a reaction should depend on the wall surface area and the pressure of the gas, and Brokaw, Badin and Pease showed that this was so for mixtures of triethylborane and oxygen at low pressures[26].

In contrast to these findings, the known facts about the liquid phase oxidation of organoboranes did not favour a chain mechanism. The reaction did not appear to have an induction period, and was not inhibited by small quantities of hydroquinone [8,12], unlike hydrocarbon autoxidations. It was, however, suppressed by amines[13] in excess, or by water[6], both of which can co-ordinate with the boron atom and render it less susceptible to nucleophilic attack by oxygen. The following reaction scheme was therefore suggested[12,20]:

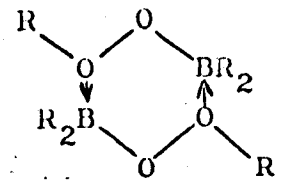


(xiii)

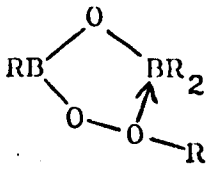
(VII)

The formation of an initial complex with oxygen which subsequently rearranged to the peroxide was reported by Zutty and Welch[27]. Oxygen was bubbled through a dilute solution of tri-n-butyl borane in benzene, hexane or octane, which was then thoroughly purged with nitrogen; the peroxide concentration went up from 30% to 60% in three hours. However, Davies[16] was not able to repeat this, nor could any such intermediate be detected spectroscopically during the formation of methylperoxy dimethylborinate at cryogenic temperatures[28]. Moreover, Hansen and Hamann[29] report that in the autoxidation of triethyl borane, all peroxide is formed within fifteen seconds. (See Discussion, p.202).

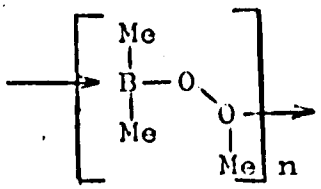
In 1966 Davies and Coffee[17] found, in the autoxidation of trialkylboroxines, that two facts were not compatible with the mechanism previously proposed (Equation xiii). Small amounts of boric acid were formed during the reaction, and non-stoichiometric quantities of oxygen were absorbed. Previously[13,16] the latter effect had been attributed to inhibition of the reaction by inter- or intra- molecularly co-ordinated products such as (VIII) or (IX), and in support



(VIII)



(IX)



(X)

of this, a polymeric peroxide of suggested structure (X) was isolated from the reaction of oxygen with trimethyl borane at low temperatures[28]. Davies now pointed out that the evidence did not exclude a free radical mechanism of the form:

$$R^{\cdot} + O_2 \longrightarrow RO_2^{\cdot} \qquad \text{(xiv)}$$

$$RO_2^{\cdot} + \begin{matrix} R \\ | \\ B{-}R \\ | \\ R \end{matrix} \longrightarrow \left[\begin{matrix} R \\ | \\ RO_2{-}B^{\cdot}{-}R \\ | \\ R \end{matrix} \right] \longrightarrow \begin{matrix} R \\ | \\ RO_2{-}B \\ | \\ R \end{matrix} + R^{\cdot} \qquad \text{(xv)}$$

(See Discussion, p 151).

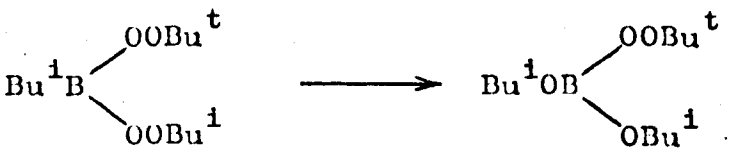
e) Rearrangement of Peroxides:-

The borinic, boronic and borate esters observed as autoxidation products by earlier workers[1-8] must be formed by rearrangement of the initial peroxide. Davies, Hare

and White[14,15] have followed this reaction using proton magnetic resonance techniques. The spectrum of (XI), ten minutes after isolation, showed only one doublet at low field, due to the $-CH_2OO-$ grouping[14]; but after 2.5 hrs. at room temperature a second doublet had appeared at slightly higher field, corresponding to the appearance of the $-CH_2O-$ group:

$$(Me_2CHCH_2OO)_2BCH_2CHMe_2 \longrightarrow Me_2CHCH_2OOB(OCH_2CHMe_2)_2 \qquad (xvi)$$
(XI)

The peroxide from diisobutyl-t-butyl borane was studied in a similar way[15]. It was found for this compound that the isobutylperoxy group was reduced more readily than the t-butylperoxy group (Equation xvii).



(xvii)

This would be expected for either an inter- or an intramolecular rearrangement, but Davies suggested that the latter was more likely, by analogy with the Baeyer-Villiger reaction and other rearrangements to electron-deficient centres[30,31].

(xviii)

The intermolecular mechanism (Equation (xix) was preferred by Mirviss[32], who investigated the air oxidation of trialkyl boranes and found that the redox reaction was

favoured by increasing concentration of borane. Thus the

$$-\overset{|}{B}-R + -\overset{|}{B}O_2R \longrightarrow 2-\overset{|}{B}OR \qquad \text{(xix)}$$

peroxide content during oxidation was lower, and the alcohol yield after hydrolysis higher, for neat liquid boranes than for 50% by weight solutions in heptane. In a recent publication[33] he has produced more convincing evidence for the intermolecular reaction (See Discussion p. 198).

f) Polymerisation with Boron Compounds:-

In 1942 Grummitt[8] observed that vinyl acetate polymerisation was initiated by partially air oxidised trialkyl boroxines. Subsequently Furukawa[34] and Kolesnikov[35] found that trialkyl boranes were equally effective, the rate of polymerisation being markedly accelerated by traces of oxygen or peroxides[36]. The number of papers published in this field has increased greatly in the last few years. Boranes have been used for homo- and co-polymerisation either alone (probably with traces of oxygen present) or with various co-catalysts. Oxygen, peroxides, halides and ligands such as ammonia have all been successfully combined with trialkyl boranes to initiate polymerisation[37]. The mechanism has been interpreted by some authors as anionic[34], but in most cases the evidence favours a free radical reaction.

The co-polymerisation of acrylonitrile and methyl methacrylate, initiated by metal alkyls, was studied by Zutty and Welch[38]. Monomer reactivity ratios, defined as $r_1 = \frac{k_{11}}{k_{12}}$ and $r_2 = \frac{k_{22}}{k_{21}}$ (Equations xx - xxiii), were measured for each initiator.

$$\sim M_1^{\bullet} + M_1 \xrightarrow{k_{11}} \sim M_1M_1^{\bullet} \qquad \text{(xx)}$$

$$\sim M_1^{\bullet} + M_2 \xrightarrow{k_{12}} \sim M_1M_2^{\bullet} \qquad \text{(xxi)}$$

$$\sim M_2^{\bullet} + M_1 \xrightarrow{k_{21}} \sim M_2M_1^{\bullet} \qquad \text{(xxii)}$$

$$\sim M_2^{\bullet} + M_2 \xrightarrow{k_{22}} \sim M_2M_2^{\bullet} \qquad \text{(xxiii)}$$

[$\sim M^{\bullet}$ represents a polymer chain terminating in a radical derived from monomer $\dot{M}$.]

These ratios have different values for free radical and polar chain mechanisms, and showed quite clearly that while alkyls of lithium, sodium, magnesium and beryllium catalysed an anionic polymerisation, those of boron, aluminium, cadmium and zinc catalysed a free radical process.

It was further shown by Bawn, Margerison and Richardson[39] that although a partially autoxidised trialkyl borane would initiate the polymerisation of methyl methacrylate, neither borane nor peroxide alone were effective. To account for this they proposed the following mechanism:

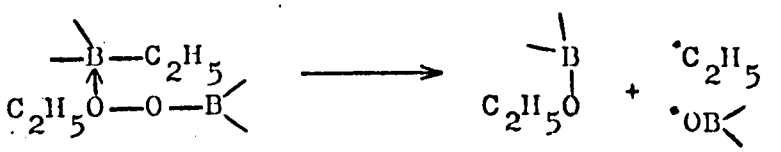
$$R_3B + M \rightleftharpoons R_3B{\cdot}M \qquad \text{(xxiv)}$$

$$R_3B{\cdot}M + R_2BOOR \longrightarrow R_3B{\cdot}MRO^{\bullet} + R_2BO^{\bullet} \qquad \text{(xxv)}$$

$$\left.\begin{matrix} R_3B\cdot MRO^{\cdot} \\ R_2BO^{\cdot} \end{matrix}\right\} + M \longrightarrow M^{\cdot} \quad \text{etc.} \qquad \text{(xxvi)}$$

A very similar scheme was put forward later by Welch[40].

Unsuccessful attempts were made to detect ethoxy radicals in the polymerisation of methyl methacrylate with a mixture of triethyl borane and the peroxide derived from it[29]. In the presence of iodine as a radical trap, the main by-product was ethyl iodide, and it was suggested that interaction of peroxide and borane gave ethyl radicals, which were the initiating species (Equation xxvii). Careful kinetic

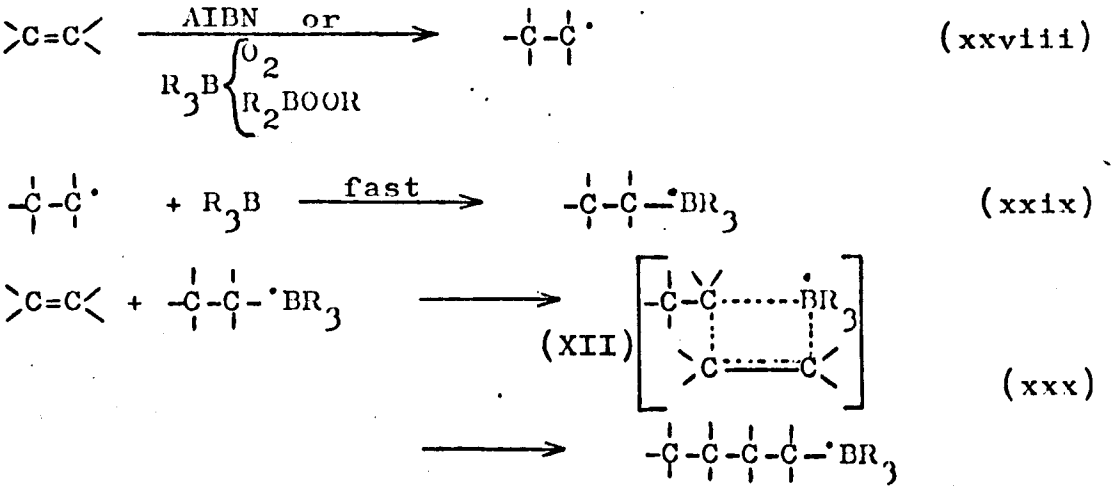


(xxvii)

studies on this system were interpreted in favour of this mode of initiation[41], although the very slow polymerisation initiated by triethyl borane in the absence of oxygen was thought to be ionic in nature.

Recently Arimoto[42] has proposed that the growing polymer radical is bound in some way to the boron atom. He was able to polymerise various monomers in the presence of inhibitors, using trialkyl borane and either oxygen, alkylperoxyborane or bis-azoisobutyronitrile (AIBN) as initiators for the reaction. Since polymerisation took place even with a

large excess of phenothiazine present (16.7 mol. per mol. of borane), he concluded that the latter must be competing very effectively for the growing radical, and postulated the following mechanism:



(xxviii)

(xxix)

(xxx)

g) Autoxidation of Other Organo-metallic Compounds:-

The autoxidation of other organometallic compounds shows certain distinct trends. In general, the more polar the metal-carbon bond, the more readily autoxidation occurs. At one extreme, the lower members of the alkali metal alkyls ignite on contact with air while at the other, organosilicon compounds are of comparable stability to those of carbon. The rate of aerial oxidation of trimethylphenyl silane, for instance, is reported to be very similar to that of t-butyl benzene[43].

Peroxides have been detected in many cases[44], and may be

the first product of reaction of all metal alkyls. The stability of these peroxides decreases with decreasing electronegativity of the metal. Thus those of lithium and magnesium are immediately reduced to alkoxides by unoxidised substrate, while cadmium alkyls can be converted to the monoperoxide by a deficiency of oxygen, with rearrangement to the dialkoxide only occurring very slowly[45].

$$R_2Cd \xrightarrow{O_2} RCdOOR \xrightarrow{slow} (RO)_2Cd \qquad \text{(xxxi)}$$

The more reactive peroxides have been shown in many cases to exist by adding a dilute solution of the metal alkyl to rapidly stirred oxygen saturated ether at -78°[46,47]. The products were not isolated, but hydrolysed to the corresponding hydroperoxide:

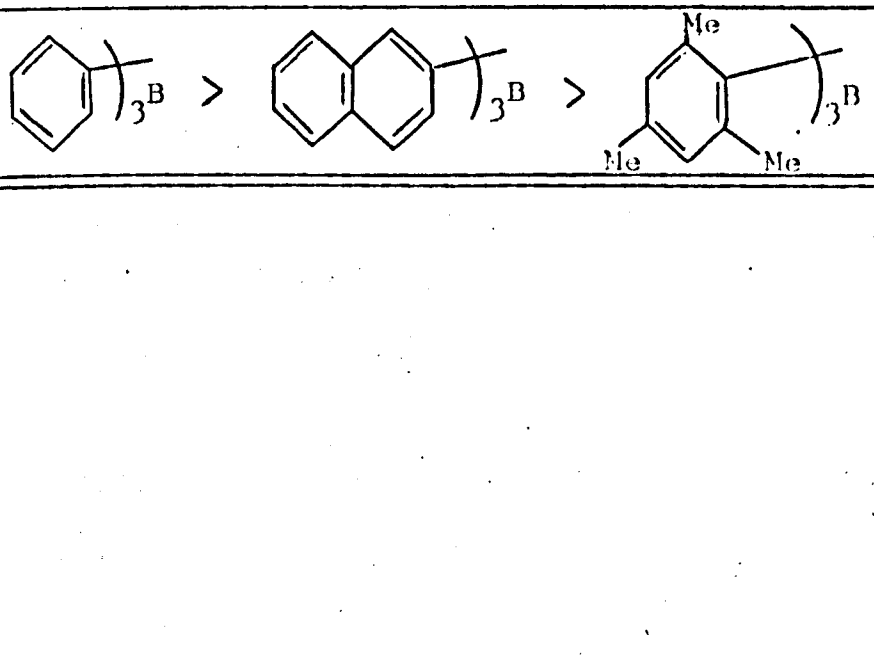
$$RMgCl \xrightarrow{O_2} ROOMgCl \xrightarrow{H_2O} ROOH + HOMgCl \qquad \text{(xxxii)}$$

The autoxidation of metal aryls does not result in significant quantities of peroxide. This may be due to their great instability once formed, or to the lower reactivity of aryl radicals towards oxygen, resulting in coupled products[48].

Compounds of groups II and III in the periodic table appear to be similar in the pattern of their autoxidation

to organoboranes. Group I compounds are far more reactive, and may autoxidise by a different route due to their ionic nature. In Group IV a full octet of electrons results in greatly reduced activity, whereas the trivalent metallic compounds in Group V, which readily expand their covalencies to five, react vigourously with oxygen.

Table 5 - Relative Rates of Autoxidation of Organoboranes

Relative Rates of Autoxidation	Reference
$R_3B > R_2BX > RBX_2$ (X=halogen, OR, OOR, OH)	5,8
$R_2BF > R_2BCl > R_2BBr > R_2BI$	18
$R_2BOH > R_2BOR > R_2BONa$	8
Alkyl—B< > Aryl—B< ~ Vinyl—B<	7,20
t—RB< > s—RB< > p—RB< > CH_3B<	7,8,17,21
	19

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Chapter 4

Preparation of Solvents and Reagents and Methods of Analysis

a) Solvents and Reagents:-

i) Diglyme. The dimethyl ether of diethylene glycol (diglyme) ($2\frac{1}{2}$ l.) was stood over calcium hydride (10 g.) for 24 hr., then flash distilled from it, collecting the fraction b.p. 160-165°C/760 mm. Lithium aluminium hydride (ca 5 g.) was added cautiously under nitrogen, and the diglyme distilled under reduced pressure, collecting the fraction b.p. 65°/16 mm[1]. It was kept under nitrogen.

ii) Tetrahydrofuran. Tetrahydrofuran (THF) ($2\frac{1}{2}$ l.) was stood over sodium wire for two days, then distilled from lithium aluminium hydride[1]. The fraction b.p. 66°/760 mm. was stored under nitrogen.

iii) Boron Trifluoride Etherate. Ether (15 ml./l.) was added to crude boron trifluoride diethyl etherate, which was then distilled at reduced pressure, collecting the fraction b.p. 46°/10 mm[1]. This was stored under nitrogen.

iv) Sodium Borohydride. In most cases this was used as supplied by the manufacturers, after analysing it for active

hydride by the method of Lyttle, Jensen and Struck[2]. Where this analysis was low, the sodium borohydride was recrystallised from hot diglyme as described by Brown and Subba Rao[3], and stored in a desiccator.

v) α-Pinene. (+) α-Pinene was distilled at 73°/50 mm. and had $[\alpha]_D^{19^\circ}$ + 47.7°. (-) α-Pinene, similarly treated, had $[\alpha]_D^{22^\circ}$ - 42.4°. Both were stored under nitrogen in the dark. GLC analysis on Apiezon L at 110° showed that small quantities of β-pinene (less than 5%) were present.

vi) Butenes. cis-2-Butene (ex. Phillips Petroleum) and 1-butene were analysed by GLC on dimethyl sulpholane at 23° and shown to be better than 99% pure.

vii) Chlorobenzene. Chlorobenzene (general purpose reagent) was distilled at atmospheric pressure, collecting the fraction b.p. 130.5°C. This was stored over 5A molecular sieve.

viii) Other Solvents. Ether, benzene (analar) and pentane were dried by standing over sodium wire.

ix) Ethanolamine. Ethanolamine was distilled at reduced pressure, b.p. 59°/4 mm. and was stored under nitrogen.

x) Diethanolamine. This was treated as ethanolamine, and had b.p. $110^{o}/0.2$ mm.

xi) Acetylacetone. Acetylacetone was washed with 2N sodium hydroxide until the washings were alkaline to litmus, and then with water[4]. After drying with molecular sieve it was distilled under reduced pressure, collecting the fraction b.p. $45^{o}/30$ mm., and stored under nitrogen in the dark.

xii) Butyl Bromides. After drying with molecular sieve, butyl bromides were fractionally distilled at atmospheric pressure and stored over molecular sieve, n-butyl bromide b.p. 102^{o} s-butyl bromide b.p. 90.5^{o} isobutyl bromide b.p. 91^{o}.

xiii) Methyl Borate. This was prepared by the method of Schlesinger and Brown[5] and had b.p. $68.5\text{-}69^{o}$. (Yield = 38% based on boric acid.) Care was taken to exclude moisture at all stages during this preparation.

xiv) 5% Palladium on Barium Sulphate. This hydrogenation catalyst was prepared as detailed by Augustine[6].

xv) Molecular Sieves. Linde molecular sieves, when required for drying oxygen-sensitive compounds, were twice degassed and brought back to atmospheric pressure under nitrogen.

xvi) Nitrogen. White spot nitrogen (British Oxygen Co.) was freed from possible traces of other gases by passage over B.A.S.F. catalyst (finely divided copper) to remove oxygen, followed by sodium hydroxide and finally magnesium perchlorate.

xvii) Standard Solution of Diborane. Diborane, generated by the dropwise addition of iodine (130g., 0.51 mole) in diglyme (400 ml.) to a stirred solution of sodium borohydride (38g., 1.0 mole.) in diglyme (200 ml.)[7] over a period of ten hours, was passed in a slow stream of nitrogen through a trap at -78° and into well-stirred THF (500 ml.) in a flask fitted with a serum cap and nitrogen outlet. The generator was then heated to 70° (2 hrs.) cooled and disconnected. The flask containing the THF solution was stoppered under nitrogen and analysed for boron by withdrawing a sample by syringe (1 ml.), injecting this below the surface of some acetone (10 ml.), diluting with water, saturating with mannitol and titrating with $\frac{N}{20}$ sodium hydroxide (phenolphthalein). The diborane solution was stored at 0°C.

b) Organoboranes for Autoxidation:-

Since these compounds were exceedingly oxygen sensitive, elaborate precautions were taken to exclude air during their preparation and in subsequent manipulations. All solvents

were degassed by prolonged bubbling with nitrogen, and all operations were conducted under nitrogen. In cases where it was also important to exclude moisture, apparatus was dried in an oven at 140°, assembled rapidly and thoroughly flushed with nitrogen. Ground glass joints were lubricated with silicone grease and spring-loaded.

i) Diisopinocampheylbutyl Boranes. $(IPC)_2BBu^s$ and $(IPC)_2BBu^n$ were prepared by in situ generation of diborane[1], using cis-2-butene or 1-butene, and most of the diglyme removed by prolonged pumping at 10^{-2} mm. Benzene was added and the solution filtered into a flask which was then closed with a serum cap and stored at -20° until required.

ii) s-Butylboronic Acid and Anhydride. s-Butyl magnesium bromide was prepared on a 2 mole scale and transferred to a dry, nitrogen-filled, graduated funnel. Two samples were removed for analysis, using the apparatus of Fig.2. With tap A closed, the pressure inside the apparatus was reduced slightly by means of tap B. With B closed, A was then opened cautiously to admit about 3 ml. of solution into the apparatus. A was then closed, and the pressure reduced further via B, thus transferring most of the solution into the test tube. Nitrogen was admitted, and the test tube rapidly changed so that a second sample could be taken.

The serum cap enabled the fine nozzle C to be unblocked with a needle if necessary.

The Grignard solution (1.7 mole) was added dropwise to methyl borate (200 ml., 1.84 mole) in ether (750 ml.) at -74° ($4\frac{1}{2}$ hr.). After allowing to warm overnight, 4 N sulphuric acid (900 ml.) was added (2 hr.), the mixture stirred ($\frac{1}{2}$ hr.) and the aqueous layer separated and extracted with two 700 ml. portions of ether. The combined ether extracts were distilled to low bulk, then toluene (250 ml.) added and water removed as an azeotrope, using a 12" fractionating column packed with Fenske helices, and a Dean and Starke distillation head. After removing the toluene at $33-35^{\circ}/50$ mm., s-butyl boronic anhydride was collected over the range $62.5-66^{\circ}/1$ mm. (55g., 38% yield based on Grignard). The main fraction had b.p. $65^{\circ}/1$ mm. (Found B=12.0% lit[8] b.p. $79.5-80.5^{\circ}/2.5$ mm., calc. for $C_{12}H_{27}B_3O_3$ B=12.9%.)

The residue, mainly polymeric boronate, was treated with water, and recrystallised from toluene under nitrogen, yielding the crude boronic acid (42 g., 24% yield based on Grignard; total yield 62%). This was treated with toluene, water azeotroped off, and more boronic anhydride recovered, the process being repeated until the majority of the yield

was in the form of the anhydride which, being a liquid, was easier to handle.

Several distillations failed to improve the boron analysis, and the compound was eventually purified via the diethanolamine derivative.

To crude s-butylboronic anhydride was added a slight excess of diethanolamine in ethanol, and the mixture heated on a steam bath for an hour. Ethanol was removed on the rotary evaporator, the solid washed with carbon tetrachloride, and recrystallised from acetonitrile. On exposure of the hot solution in acetonitrile to the air, it became orange, and five recrystallisations were necessary to get a white product, m.p. 155° (lit[8] m.p. 156°).

The diethanolamine derivative (30g) was shaken for two hours with 5% sulphuric acid (154 ml.) and benzene (200 ml.). The benzene layer was washed with water, the water azeotroped off, and the boronic anhydride distilled, b.p. 73°/1.5 mm. (Found C, 56.9; H, 11.3; B, 12.6%. Calc. for $C_{12}H_{27}B_3O_3$ C, 57.2; H, 10.7; B, 12.9%). Yield: 8g. (54%).

iii) Diethyl s-Butylboronate. In a small scale distillation apparatus with a 3" column packed with Fenske helices

were placed s-butylboronic anhydride (5 ml.) and ethanol (25 ml.). Excess ethanol was removed by slow distillation at atmospheric pressure, and then the ester at reduced pressure (3 ml.) b.p. $43^{o}/9$ mm., $n_D^{25^o}$ 1.3979. (Found: B, 7.2%; $C_8H_{19}BO_2$ requires B 6.9%.) Mass spectrometry showed a molecular ion of 158 and the NMR spectrum had a peak centred at 6.25τ, split 1:3:3:1 of total intensity 4, together with a complex system centred at 9τ of total intensity 15-16. There was also a small unexplained peak at 6.03τ

iv) Di-n-butyl s-Butylboronate. This was prepared as above from s-butylboronic anhydride (4 ml.) and n-butanol (25 ml.). There was obtained di-n-butyl s-butylboronate (5 ml.) b.p. $98^{o}/10$ mm, $n_D^{25^o}$ 1.4133. (Found: B, 5.0%; $C_{12}H_{27}BO_2$ requires B 5.1%.) Mass spectrometry showed no molecular ion at 214; the highest m/e peaks were at 157 and 100, corresponding to the loss of one and two butyl groups, respectively. The NMR spectrum had a peak centred at 6.3τ and split 1:2:1, which would be expected for an $-OCH_2-$ group split by an ajacent methylene.

v) Tri-s-butyl Borane. This was prepared from s-butyl magnesium bromide and boron trifluoride etherate by the method of Hennion, McCusker et al.[9] with the modification

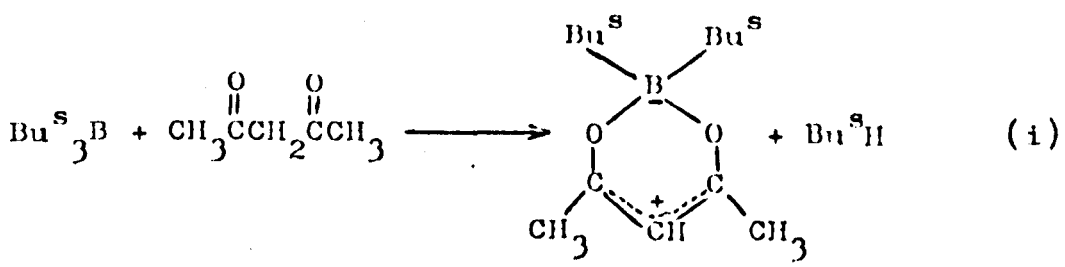
that the Grignard solution was maintained at -10° to -5° during the addition. After extraction and fractional distillation of the product there was obtained tri-s-butyl borane (0.82 mole, 80% yield based on boron trifluoride) b.p. 61-62°/3 mm., $n_D^{25°}$ 1.4352; found: B, 5.8% (lit[9] b.p. 59.7-60°/2.5 mm; $n_D^{25°}$ 1.4349; Calc. for $C_{12}H_{27}B$, B 5.9%).

vi) Tri-n-Butyl Borane. This was prepared as above from the Grignard and boron trifluoride etherate, yielding tri-n-butyl borane (0.33 mole, 66% yield based on boron trifluoride) b.p. 84°/5 mm; found: B, 5.9% (lit[10] b.p. 90-91°/9 mm; calc. for $C_{12}H_{27}B$, B 5.9%).

vii) Tri-isobutyl Borane. This was prepared as above on a one mole scale, but after completing the addition of boron trifluoride the solution was only allowed to stand for two hours before workup, instead of overnight. After distillation a fraction was obtained (175g.) b.p. 68°/6.5 mm; $n_D^{21°}$ 1.4212; found: B, 4.8%, (lit[9] b.p. 68°/7 mm; $n_D^{25°}$ 1.4203; calc. for $C_{12}H_{27}B$, B 5.9%). Mass spectrometry showed peaks at mass number 209 and 152, corresponding to loss of one and two butyl groups respectively from diisobutylborinic anhydride. Accordingly, ethanolamine (61g., 1 mole) was added, and a liquid distilled at 82°/17 mm. (Found; B 4.3%.) This was shaken with excess 10% sulphuric acid for one hour

to remove ethanolamine, washed with water, dried and distilled at atmospheric pressure. Three fractions were obtained (40g., 24% yield) of which fraction two had b.p. 189-189.5°, B 5.3%, $n_D^{20°}$ 1.4220 and fraction three had b.p. 189.5-191°, B 5.8% $n_D^{20°}$ 1.4208 (lit[11] b.p. 189-190°/751 mm., $n_D^{25°}$ 1.4200). All experiments were conducted using fraction three.

viii) Di-s-Butylborinic Anhydride. This was prepared by the method of Hawthorne and Reintjes[12]. Acetyl acetone (30g., 31ml., 0.30 mole) and tri-s-butyl borane (36.4g., 47.5 ml., 0.20 mole) were heated at 75° under nitrogen for five days. The bright yellow liquid was distilled, yielding acetyl acetone, a fraction b.p. 29°/0.15 mm. (unreacted tri-s-butyl borane, 25 ml.) and finally s-butyl boronium acetylacetonate (I) (20g. 0.089 mole, 45% yield) b.p. 67°/0.12 mm., $n_D^{23°}$ 1.4709 (lit[12] b.p. 71°/0.15 mm., $n_D^{20°}$ 1.4709), (Equation i).



(i)

(I)

The boronium compound (I) (17.7g., 0.079 mole) was

hydrolysed by refluxing under nitrogen with potassium hydroxide (14.0g., 0.25 mole) and water (125 ml.) (24 hr.). After neutralising with 15% hydrochloric acid, the liquid was extracted twice with ether, the extracts dried with magnesium sulphate and the ether removed under reduced pressure. On distillation there was obtained a small fraction b.p. 29°/0.2 mm., and di-s-butylborinic anhydride (5 ml., 40% yield) b.p. 76°/0.2 mm. $n_D^{20°}$ 1.4400 (lit[12] b.p. 84-86°/0.15mm; $n_D^{20°}$ 1.4392. Found: B 8.4%; calc. for $C_{16}H_{36}B_2O$, B 8.15%.).

ix) s-Butyl Di-s-Butylborinate. This was prepared in an analogous manner to diethyl s-butylboronate above, from di-s-butylborinic anhydride (2.5 ml.) and s-butanol (26 ml.). There was obtained s-butyl di-s-butylborinate (3 ml.) b.p. 77-78°/10 mm., $n_D^{24°}$ 1.4133. (Found: B 5.0%; $C_{12}H_{27}BO$ requires B 5.45% A second distillation, taking a middle cut. failed to improve the boron analysis.

x) 1-Phenylethylboronic Acid. (PEB) Crude n-butyl 1-phenylethylboronate[13] (28.5g., 0.092 mole assuming 85% pure) was converted to the diethanolamine ester (20g., 72% yield) as described for s-butylboronic anhydride. Two recrystallisations from acetonitrile gave a product m.p. 194° (lit[14] m.p. 204-205°). The ester, hydrolysed

as previously described, gave a solution of the free acid in benzene. This was shaken for a few minutes with 5A molecular sieve, and filtered into a flask which was then sealed with a serum cap. After analysing the solution for boron, it was used for autoxidation runs without isolating the free acid.

c) Inhibitors:-

With the following exceptions, inhibitors were used without purification.

i) Diphenylpicrylhydrazyl. The IR spectrum of this compound was measured. The band at 3250 $cm.^{-1}$ corresponding to the N-H stretch of the parent amine was found to be absent.

ii) Phenothiazine. The dark green powder was boiled with 80-100° petroleum ether, and filtered. On cooling, the filtrate deposited pale yellow crystals which were filtered, dried and stored under nitrogen, m.p. 182-3° (lit[15] 185.5°). Exposure to air over a period of a few days caused the crystals to turn green.

iii) Diphenylbenzidine. The black powder was recrystallised from toluene, using the same technique as for phenothiazine. The silver-grey flakes had m.p. 233° (lit[16] 245-247°) and

were still slightly impure.

iv) 2,2,6,6 Tetramethyl Piperid-4-one Nitroxide (Triacetonamine Nitroxide).

1) Triacetonamine. This was prepared from acetone and ammonia by the method of Hall[17]. After working up the product and subjecting it to careful fractional distillation through a 3 ft. column packed with Fenske helices, there was obtained triacetonamine, 2,2,6,6,tetramethyl piperid-4-one (258g., 8% yield), b.p. 97-100°/18 mm; m.p. 36°(lit[17] b.p. 102-105°/18 mm; m.p. 34-36°). The pale yellow crystals were kept under nitrogen in a refrigerator; however, they gradually became orange, with a lowering of the melting point, even under these conditions.

2) Oxidation to the Nitroxide[18]. To a solution of sodium tungstate (0.1 g.) and disodium ethylene diamine tetraacetate (2.5 g.) in water (50 ml.) were added triacetonamine (37.5 g., 0.24 mole) and 30% hydrogen peroxide (50 ml.). The mixture was stirred magnetically for 4½ days in the dark, then a further 25 ml. 30% hydrogen peroxide added and stirring continued for another 24 hours in the dark. The liquid was then heated to boiling for a few minutes to decompose the peroxide, and on cooling, saturated with anhydrous potassium carbonate. The red oil which separated was

extracted with 2 x 150 ml. ether, and the ether removed at reduced pressure on a rotary evaporator, down to 0.25 mm. Pentane (10 vol.) was added to the oil, then just enough ether to make the two miscible. The solid which came down on cooling to -78° was filtered under nitrogen and recrystallised in the same way, yielding dark orange crystals of triacetonamine nitroxide (23 g., 56% yield). A further recrystallisation from hexane gave the pure compound (9.6 g.) m.p. 42-44° (lit[19] m.p. 36°); light orange crystals with IR spectrum identical to that of an authentic specimen[20], and shown iodometrically to be 99% pure.

v) Galvinoxyl. A solution of 4,4' dihydroxy 3,5,3',5' tetra-t-butyldiphenylmethane (Ethyl 702) (43.2g., 0.102 mole.) in anhydrous ether (550 ml.) was stirred with lead dioxide (111g., 0.46 mole) under nitrogen at room temperature (4 hr.). After filtering under nitrogen, a further 111g. of lead dioxide
under nitrogen for a further 4 hours. The solution was filtered again, and the solvent removed under reduced pressure. The solid residue (42.5g.) was recrystallised from benzene, yielding dark blue crystals of galvinoxyl (40g., 93% yield) m.p. 152°, (lit[21] m.p. 153.2-153.6°). A very intense maximum was observed in the visible spectrum at 431 mμ in agreement with the literature[21]. The ESR

spectrum (Fig.3) has two quintets, corresponding to minor splitting by the four equivalent ring protons superimposed on the major splitting by the methine proton.

vi) Hydrogalvinoxyl. A solution of Ethyl 702 (0.425g., 1 mmole) and galvinoxyl (0.844g., 2 mmole) in benzene (30 ml.) was degassed and left under nitrogen for three days at 25°. The solvent was removed under vacuum, leaving a yellow solid, hydrogalvinoxyl. This had an absorbtion maximum at 385 mμ in the UV, and bands at 2.75μ, 6.20μ, 6.33μ and 6.45μ in the IR region. (Lit[22], UV maximum at 385 mμ, OH at 2.74μ and 3 intense bands between 6.1μ and 6.5μ)

d) Analyses:-

i) Boron. Weighed samples containing ca 0.005g. boron were treated with 30% hydrogen peroxide (10 ml.) at 60° (1 hr.). The solutions were then titrated with $\frac{N}{20}$ sodium hydroxide solution (bromothymol blue) after saturating with mannitol. A blank was run concurrently with each titre[14], care being taken to use exactly the same amount of hydrogen peroxide as for the sample, and to treat both in exactly the same way.

The end point is unsatisfactory in this method, and

several attempts were made to find a better one. In the absence of added hydrogen peroxide, potentiometric and conductometric titrations of sodium hydroxide against standard solutions of boric acid gave sharp end points. However, the effect of hydrogen peroxide in the mixture was two-fold: it increased the blank titration considerably, and it prevented a sharp change in pH at the end point. A conductometric titration was not possible, as local decomposition of hydrogen peroxide by the platinum electrodes made readings erratic.

At first, solutions were boiled to remove excess hydrogen peroxide, but this did not give consistent results. Several decomposition catalysts were tried, namely platinum oxide, cobalt chloride, manganese dioxide and the enzyme catalase. The first-named was the most effective, decomposition being complete in 1 hr. at 60°. However, even in this case the accuracy was not improved sufficiently to justify the extra step, and it was concluded that the method first described[14] was the best simple one available.

ii) Peroxide. The peroxide solution (usually in benzene) was heated under reflux under nitrogen with isopropyl alcohol (50 ml.) glacial acetic acid (1 ml.) and sodium iodide (1 g.) for about 10 minutes. After cooling, the

condenser was washed down with isopropyl alcohol (10 ml.) and the liberated iodine titrated with $\frac{N}{10}$ sodium thiosulphate. A blank was run on new bottles of sodium iodide; this was always zero.

iii) Free Radicals. The sample (ca 0.1 g.) was shaken under nitrogen for a few minutes with glacial acetic acid (25 ml.), benzene (25 ml.) and sodium iodide (1 g.), then diluted with water (100 ml.) and the liberated iodine titrated with $\frac{N}{10}$ sodium thiosulphate.

iv) Elemental Analyses. Carbon, hydrogen and boron micro-analyses were performed by Beller, W. Germany.

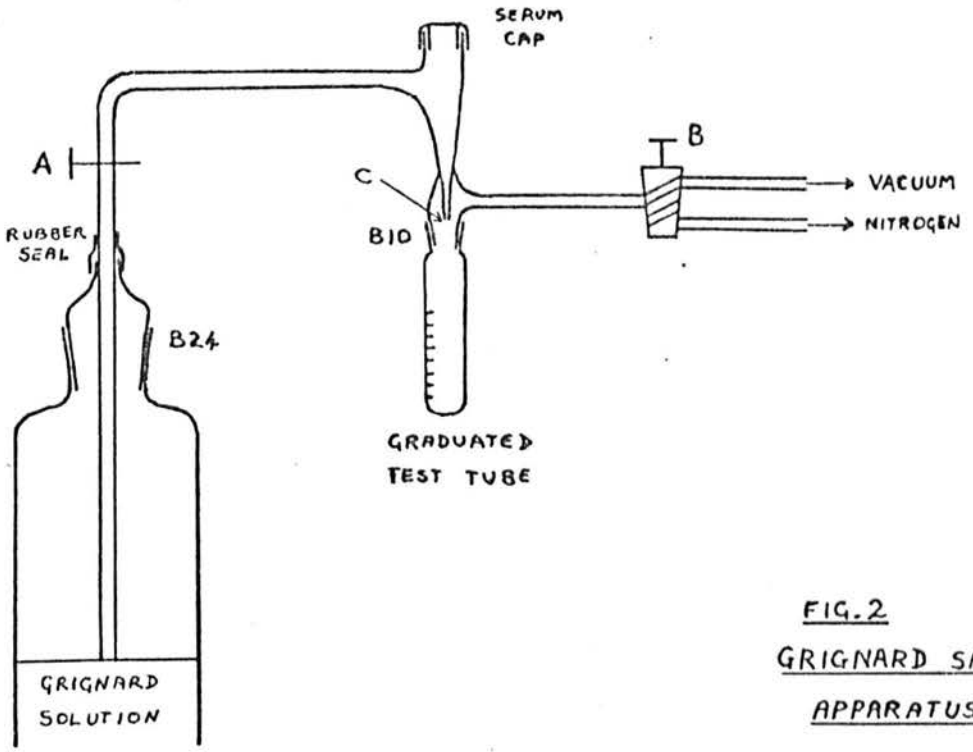
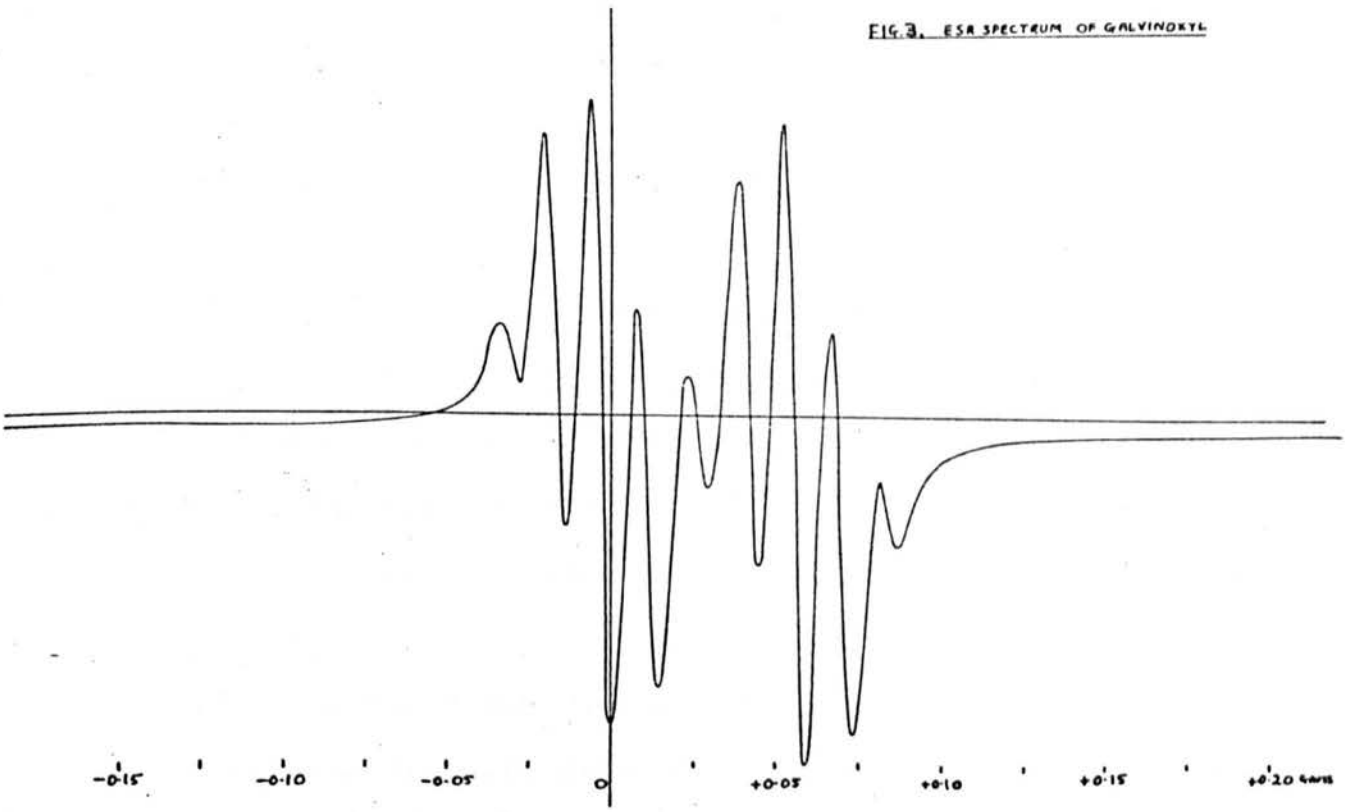


FIG. 2
GRIGNARD SAMPLING APPARATUS.

FIG. 3. ESR SPECTRUM OF GALVINOXYL



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Chapter 5

Studies on Diisopinocampheyl-s-butyl Borane

a) Preparation and Isolation of Diisopinocampheyl-s-butyl Borane $(IPC)_2BBu^s$:-

i) Early Attempts at Isolation. Early attempts to prepare $(IPC)_2BBu^s$ by external generation of diborane, using the method described by H.C. Brown for the preparation of trinorbornyl borane[1], resulted in low yields of $(IPC)_2BBu^s$, contaminated with large amounts of unreacted α-pinene. The mixed borane was therefore prepared by the in situ method originally utilised by Brown[1], although the isolation involved the removal of sodium borofluoride as an additional step.

For example, a diglyme solution of $(IPC)_2BBu^s$ (0.025 mole) was pumped exhaustively at 10^{-2} mm. to remove solvent. Pentane (60 ml.) was added, and after filtering, the filtrate evaporated to dryness under reduced pressure. There was obtained a waxy solid (11.5g, 134% of theory, probably containing diglyme. Found: B 3.7%).

This was redissolved in the minimum volume of pentane and cooled to -78°, but no precipitation occurred. The replacement of pentane by methylene chloride or THF gave

the same negative result, and by this time the solid had deteriorated to a semi-liquid mush.

Since the last traces of diglyme appeared difficult to remove, the lower boiling THF was used as the solvent in one preparation of $(IPC)_2BBu^s$. (50% excess α-pinene was used to prevent the dissociation of the intermediate tetraisopinocampheyldiborane[2].) However, on removal of the last traces of solvent, the hitherto white solid suddenly deteriorated to a viscous brown oil which could not be purified.

ii) Isolation of $(IPC)_2BBu^s$. It was found in later experiments that high yields of $(IPC)_2BBu^s$ could be obtained from diborane generated externally, by using an excess of sodium borohydride and boron trifluoride etherate over that required by the stoichiometry of the reaction (Discussion p.164).

Diborane was generated by the dropwise addition of boron trifluoride etherate (28.75 ml., 0.23 mole) in diglyme (25 ml.) to a magnetically stirred solution of sodium borohydride (4.25g., 0.112 mole) in diglyme (110 ml.) (2 hr.). This was bubbled in a slow steam of nitrogen into α- pinene (32 ml., 0.20 mole, $[\alpha]_D^{22^\circ} -42.4^\circ$) in

diglyme (175 ml.) maintained at 0° and stirred magnetically. The usual precautions to exclude oxygen and moisture were taken, and in addition the exit gases were passed through acetone to destroy escaping diborane.

The generator was heated to 80° (1 hr.), allowed to cool and removed. After four hours at 0°, during which time a thick white precipitate appeared in the reaction flask, cis-2-butene (12.6 ml., 0.163 mole) was evaporated into the flask below the liquid surface in a slow stream of nitrogen ($2\frac{1}{2}$ hr.). The flask was then allowed to warm to room temperature (2 hr.) and the ether removed under reduced pressure. A white waxy solid settled out on cooling at -78° (12 hr.), this partially redissolved when warmed to room temperature, with crystallisation of the remainder. The solution was cooled to -78° again, filtered rapidly and the solid dried by sucking a stream of nitrogen through it. The white crystals of $(IPC)_2BBu^s$ were transferred to a weighed bottle in a nitrogen-filled glove box (13.0g., 38% yield. Found: B 3.07%, $C_{24}H_{43}B$ requires B 3.16%).

b) Oxidation to s-Butanol:-

A solution of $(IPC)_2BBu^s$ (0.1 mole) in diglyme, prepared by in situ generation of diborane[1], was reduced to low bulk by pumping at 10^{-2} mm. (8 hr.). Pentane (200 ml.)

was added, and the solution filtered under nitrogen to remove sodium borofluoride. With the exception of the first experiment, the filtrate was then divided in halves which were treated in different ways as detailed below. In the case of Expt.1, the original diglyme solution was divided in two, and only for the half to be autoxidised was the diglyme replaced by pentane.

i) Oxidation with Alkaline Hydrogen Peroxide. Pentane was removed under vacuum from one half of the $(IPC)_2BBu^s$ solution (except for Expt.1), and the residue was decanted under nitrogen into a flask fitted with a reflux condenser, and a dropping funnel. Three normal sodium hydroxide (20 ml.) was added from the funnel, followed by 30% hydrogen peroxide (20 ml.). The latter was added dropwise so that the temperature did not rise above 50°C, and the flask was then maintained at 50° for another hour, the contents being stirred magnetically. The liquid was extracted with ether (3 x 50 ml.), the combined extracts washed with saturated sodium chloride solution and dried over magnesium sulphate[3]. The ether was distilled off using a 6" column packed with Fenske helices, and was shown by GLC to contain no butanol. Water (ca 10 ml.) was added, and the mixture distilled until the temperature reached 99-100°, adding more water when necessary. (s-Butanol/water azeotrope b.p.87.5°, 27.3% water[4]).

The fractions were dried, analysed by GLC on diethyl hexyl sebacate at 125° and those richest in s-butanol purified by preparative GLC on a Wilkins A705 using a 20 ft x ¼" diameter Igepal column at 125°. (For Expt.1 a Wilkins A770 was used, with a Carbowax PC20M20 column at 100°. This failed to purify the s-butanol completely, and it was passed through a second column (Silicone Nitrile at 68°)). The s-butanol was diluted with racemic s-butanol by a known factor, and the optical rotation measured. The results are shown in Table 6.

The residue after azeotroping out the s-butanol in Expts.1 and 3 was diluted with ether, washed with cold water to remove diglyme and dried over 5A molecular sieve. The ether was removed under vacuum, and the solid remaining was recrystallised from pentane at -78°, yielding white needles of isopinocampheol (11.2g., 68% yield) m.p. 49-50°, improved to 53-54° by a second recrystallisation from nitromethane, (lit[5] m.p. 57°). Table 6 shows the observed optical rotations.

It was not possible to purify the s-butanol by distillation, as it was always contaminated by varying amounts of α-pinene, as well as by other products of the reaction. Typical GLC analyses are shown in Figs. 4 and 5.

Column chromatography was used in early attempts to separate s-butanol and α-pinene. A 50 cm. column packed with a mixture of sodium sulphate and alumina, eluted with pentane containing increasing proportions of acetone, gave a reasonable separation, but better results were obtained more easily using preparative GLC.

A polar column was required to give a symmetrical peak without tailing, for s-butanol. Carbowax (polyethyleneglycol) was very good, but was only efficient above ca 125°, at which temperature separation of the alcohol and pinene was poor. A similar compound, Igepal (nonylphenoxypolyethyleneoxyethanol) had the advantage of a lower working temperature, and a preparative scale column was made up in the following manner.

Chromosorb W (250g., 60/80 mesh), an inert support, was suspended in a solution of Igepal CO-880 (50g.) in acetone (500 ml.) and shaken vigorously while the acetone was removed under reduced pressure. An annealed copper tube (20 ft. x $\frac{1}{4}$" diameter), with suitable compression unions at the ends, was wound around a former into the required shape, and then re-annealed by heating to redness and quenching in cold water, while a stream of nitrogen was passed through it to prevent oxide formation. A glass wool

plug was then placed in one end, and the coated support introduced at the other. Packing was accomplished by applying suction at the plugged end, and tapping the tube throughout the operation. Before use the column was fitted into a Wilkins Autoprep A705 and purged with nitrogen at 150° (48 hr.).

ii) Autoxidation and Hydrogenation. The other half of the pentane solution was autoxidised to form the peroxyborane, and this was then hydrogenated over a catalyst to give s-butanol. Stringent safety precautions were taken in these experiments, following an unexplained explosion during a small-scale autoxidation of a solution of $(IPC)_2BBu^s$ in a gas burette. Face masks and gloves were used, and all reactions were carried out behind safety screens. The high speed stirring apparatus of Fig.6 was also earthed to prevent the possible buildup of static electricity.

The experimental variables in each of the three runs are shown in Table 6. The solution of $(IPC)_2BBu^s$ (0.05 mole) in pentane was autoxidised by blowing dry air (oxygen in Expt.1) through it, when the temperature rose to the b.p. of pentane (35°). When the temperature of the solution began to fall, the system was flushed with nitrogen, platinum oxide catalyst added, and hydrogen passed until the amount

of peroxide, estimated iodometrically, had fallen to a negligible value. The solution was filtered to remove catalyst, and concentrated by distilling off most of the pentane through a 12" Fenske helices packed column. The residue was then worked up for s-butanol by azeotropic distillation with water as previously described. The purification was difficult due to the multiplicity of products, and yields of the alcohol were correspondingly low. Table 6 gives the optical rotations observed for s-butanol.

Experiment 1 was conducted in a 2 l. flask fitted with a sintered glass gas bubbler, a thermometer, another gas inlet serving to admit nitrogen rapidly should it become necessary, and a 24" spiral condenser acting also as the gas outlet. A safety valve, which was a tube dipping 3-4 cm. below the surface of some mercury, was incorporated into the inlet system, and in addition the reacting gas could be cut off and replaced by nitrogen at the turn of a tap.

Experiments 2 and 3 were carried out in the apparatus of Fig. 6. This was capable of stirring the solution at high speed (ca 3,000 r.p.m.) and the hollow stirrer recycled the gas through the liquid, ensuring a very high effective

surface area. Both autoxidation and hydrogenation were achieved much more rapidly than in Expt.1.

c) Attempted Preparation of Optically Active s-Butylboronic Acid:-

To recrystallised $(IPC)_2BBu^s$ (7.3g., 21.4 mmoles) in THF (50 ml.) in a flask fitted with nitrogen inlet and outlet and a serum cap, was added by syringe a solution of diborane in THF (53 ml., 0.43 M; 46.5 mmoles BH_3). The solution was stirred magnetically at 50-55° (5 hr.) under a static nitrogen atmosphere after which time methanol (5 ml., 4.8g., 150 mmole) was added dropwise by syringe. After stirring overnight (12 hr.) the solution was distilled, yielding three fractions containing boron (total : 27.7 mmoles). These were combined, and water (1.5 ml., 83 mmoles) was added to hydrolyse the methyl esters. THF was removed at the pump, leaving a white solid (1.0g.), mainly boric acid. It was leached with ether (10 ml.), filtered, the ether removed, leaving a very small residue of an optically inactive white powder, probably boric acid.

The residue, containing 16.9 mmoles boron, was cooled to -78°, warmed and cooled a second time, and after filtering under nitrogen gave recovery of $(IPC)_2BBu^s$ (2.9g., 40%. Found: B 3.16%; calc. for $C_{24}H_{43}B$, B 3.16%.

$[\alpha]_D^{19^\circ}$ + 99.5° ± 1.5° (C26, benzene)).

Several earlier attempts also failed to achieve the desired reaction. The experimental details are summarised in Table 7.

Table 6 - The Autoxidation/Hydrogenation of Diisopinocampheyl-s-butyl Borane

Expt. No.	Pentane Used, ml.	Time[a] Air Passed, hrs.	Peroxide at end of Autox., Mols. (Approx.)	PtO_2 Used, g.	Time Hydrogen Passed, hrs.	$[\alpha]_D^{(T^o)}$ of α-pinene[f] used.	$[\alpha]_D^{(T^o)}$ of s-BuOH[g] via Oxidation	$[\alpha]_D^{(T^o)}$ of s-BuOH via Autox./ Hydrog.	$[\alpha]_D^{(T^o)}$ of Isopinocampheol[h] via Oxidation
1	1,500	6	—	0.5[b]	17[c]	+ 47.7° (19°C)	- 7.7° (22°C)	- 2.8° (22°C)	-32.5°[j] (19°C)
2	700	1½	1.8	0.85	2	- 42.4° (22°C)	+ 10.3° (24°C)	- 0.44° (24°C)	—
3	400	2½	1.1	0.88[d]	4[e]	- 42.4° (22°C)	+ 11.4° (24°C)	+ 1.54° (24°C)	+ 32°[i] (20°C)

[a] Oxygen in Expt.1

[b] A further 0.5g. added after 10 hrs hydrogenation.

[c] Over a period of four days.

[d] A further 0.12g. added after 2 hrs hydrogenation.

[e] Solution allowed to stand overnight after 2 hours hydrogenation.

[f] Optically pure, $[\alpha]_D$ + 51.1° 6

[g] Optically pure, $[\alpha]_D^{20^o}$ - 13.51° 7

[h] Optically pure, $[\alpha]_D^{20^o}$ - 32° 5

[i] C4, benzene

[j] C31, benzene

Table 7 - The Reaction of Diborane with Diisopinocampheyl-s-butyl Borane

Diborane Generation	Scale Moles[a]	Temp. and Duration of Treatment	Reagent added for work-up	Desired Product
External[b]	0.025	0^{o}, 3 hr.; 20^{o}, 1 hr.	Diethanolamine	$Bu^{s}B(OBu^{n})_{2}$[d]
In situ	0.025[c]	0^{o}, 2 hr.; 20^{o}, 2 hr.	Water	$(Bu^{s}BO)_{3}$[e]
External	0.050[c]	0^{o}, 4 hr.	Water	$Bu^{s}B(OBu^{n})_{2}$[f]
External	0.050[c]	0^{o}, 2½ hr.; 20^{o}, 1 hr.	H_2O_2/NaOH	$Bu^{s}OH$ [g]
External	0.10	30^{o}, 5½ hr.	Methanol	$Bu^{s}B(OMe)_{2}$[h]

[a] Starting quantity of $(IPC)_2BBu^s$. Enough diborane added to convert $R_3B \longrightarrow 3RBH_2$.

[b] Diborane, previously condensed into ampoules from a cylinder, was allowed to evaporate into the solution

[c] 50% excess of α-pinene used.

[d] After pumping off solvent, hydrolysing the diethanolamine derivative and treating with n-butanol, a very small optically inactive fraction b.p. $70-80^{o}$/6 mm. was obtained. (Found: B, 2.7%; $C_{12}H_{27}BO_2$ requires B 5.06%)

[e] No boron containing fraction was obtained on distillation. The ether soluble part of the residue was pumped at 100^{o}/0.01 mm. when a very small two layer distillate was obtained, the lower layer of which had B 6.6% (Calc. for $C_{12}H_{27}B_3O_3$, B 12.9%).

[f] Addition of water gave a white ppt; found after filtration to consist of boric acid, and an ether soluble part which on treatment with n-butanol and distillation gave three fractions low in boron and of high optical activity. The filtrate, after removal of solvent at the pump, treatment with n-butanol and three distillations, yielded a clear liquid (1 ml.) b.p. 46^{o}/0.02 mm., negligible optical rotation. (Found: C 65.8, H 12.5, B 4.9%; $C_{12}H_{27}BO_2$ requires C 67.2, H 12.6, B 5.05%).

[g] Half the $(IPC)_2BBu^s$ was worked up for s-butanol before treatment with excess diborane and had $[\alpha]_D^{20^o} + 8.8^{o} \pm 0.4^{o}$ After treatment with diborane and oxidation, s-butanol from the other half had $[\alpha]_D^{20^o} + 8.1^{o} \pm 0.4^{o}$.

[h] The only boron containing fraction distilled out after addition of methanol was methyl borate. The residue was worked up for s-butanol (4.5 ml.) which had $[\alpha]_D + 3.1^{o} \pm 0.4^{o}$.

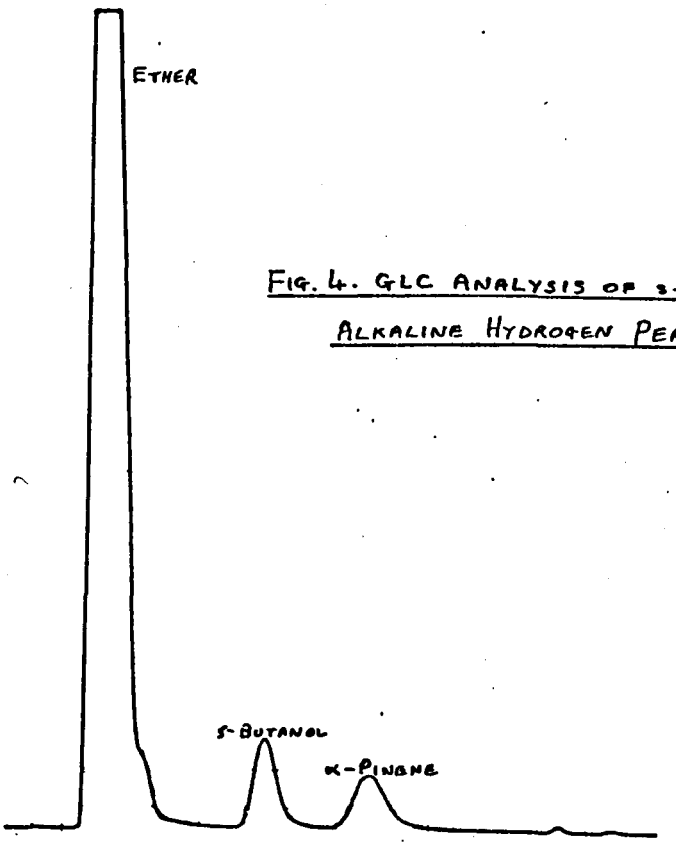


FIG. 4. GLC ANALYSIS OF S-BUTANOL VIA ALKALINE HYDROGEN PEROXIDE OXIDATION.

FIG. 5. GLC ANALYSIS OF S-BUTANOL VIA AUTOXIDATION AND HYDROGENATION.

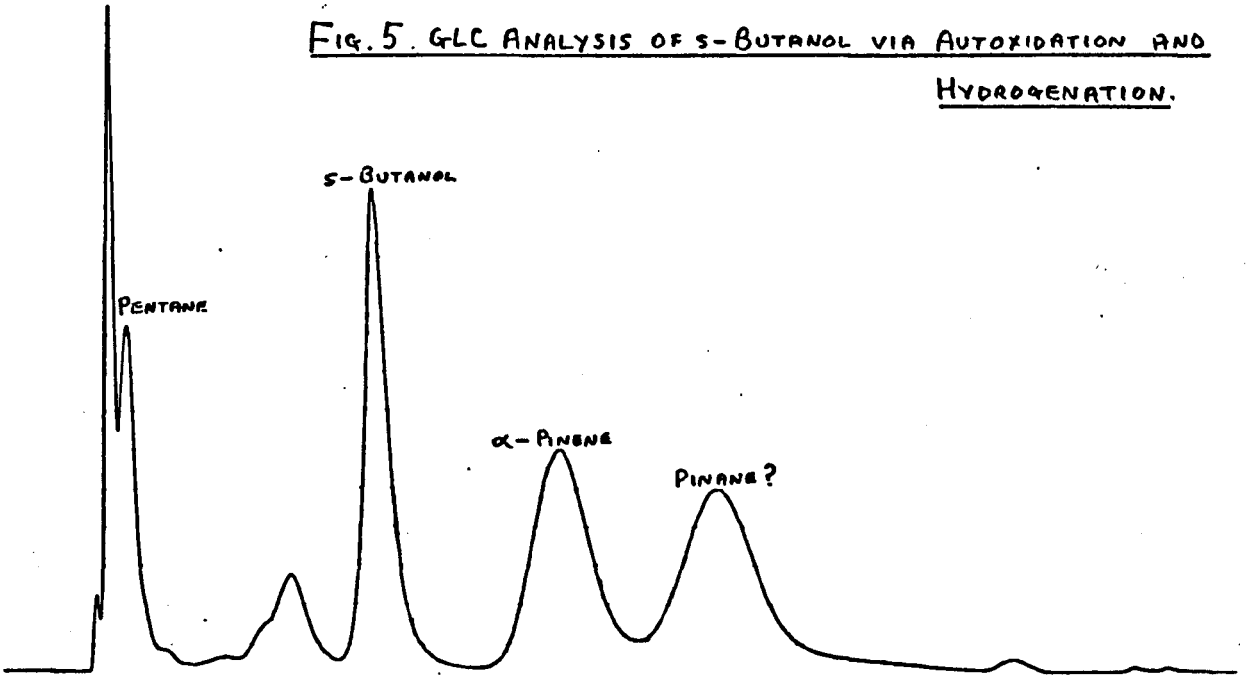
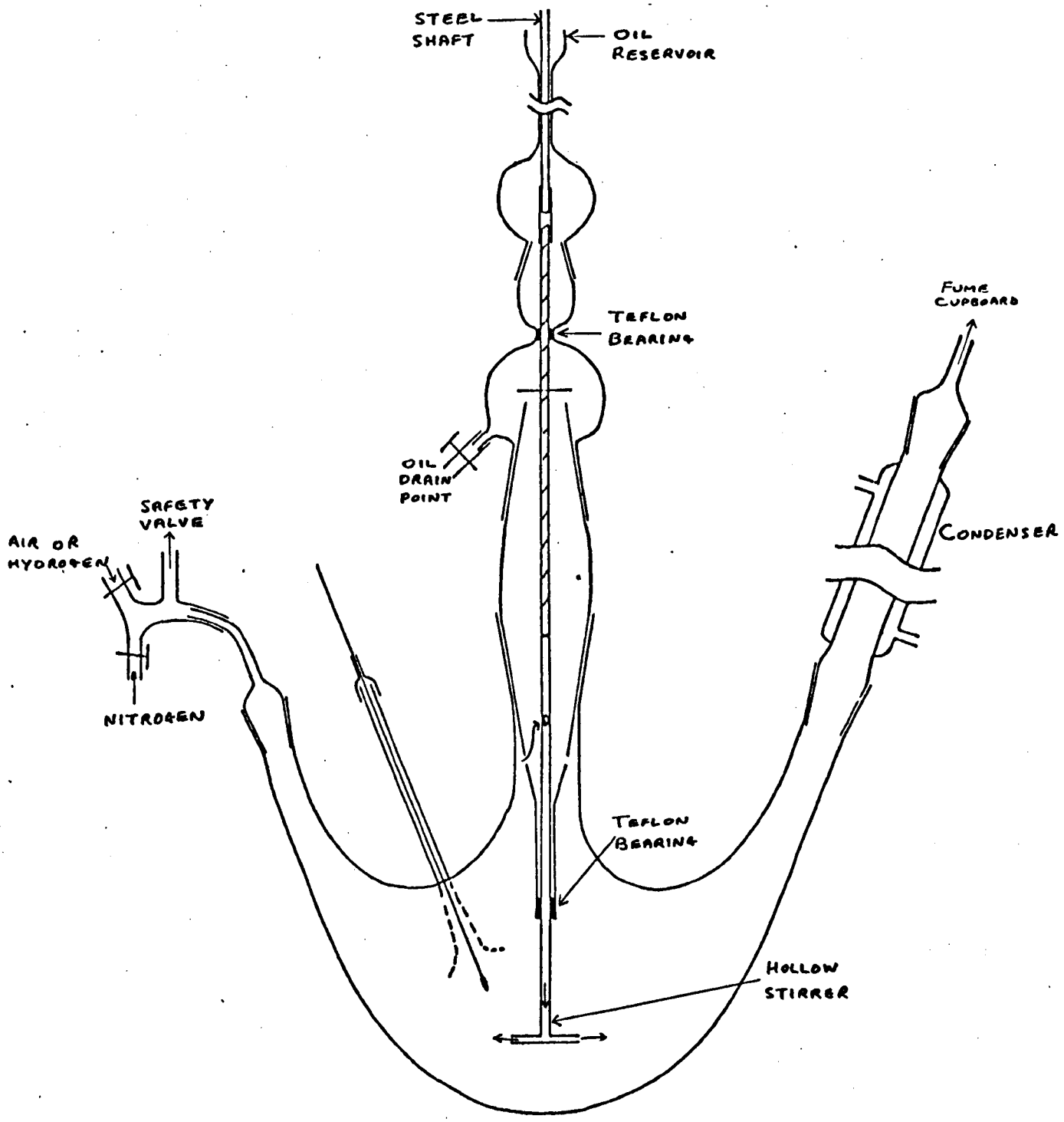


FIG. 6. HIGH-SPEED STIRRING REACTION FLASK.



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Chapter 6

Gas Burette Autoxidation of Organoboron Compounds and Reactions with Inhibitors

a) Gas Burette Autoxidations:-

i) Standard Solutions. Solutions of $(IPC)_2BBu^s$, $(IPC)_2BBu^n$, s-butylboronic acid and PEB in benzene were prepared as described in Ch. 4 (pp. 84, 86, and 91). In all other cases the organoborane was added by syringe to degassed benzene in a flask closed by a serum cap. The quantities used for all solutions were such that 1 ml. took up ca 20 ml. oxygen at STP. A 1 ml. aliquot was analysed for boron, using the same syringe both for analyses and for injection of samples into the gas burette. The syringe was fitted with a stop so that an accurately reproducible quantity of liquid was delivered (0.986 ± 0.003 ml.).

Inhibitors were usually dissolved in benzene to give solutions of known concentration, and injected into the reaction flask with a micrometer syringe just prior to the borane. In some cases, a weighed quantity of solid inhibitor was added to the reaction flask before equilibriating the burette.

ii) Apparatus and Technique. The apparatus is shown in

Fig. 7. The stirrer in the reaction flask A consisted of a shaft running in two "Teflon" bushes and bearing on a similar disc at the top. A small glass-encased magnet was attached to the top of the shaft as shown, and was activated from above the flask by an "Eclipse" horseshoe magnet, centrally mounted and driven at up to 6,000 r.p.m. by an Anderman multi-speed stirrer. The lower part of the shaft was hollow, and thus gas was recycled through the liquid as in the high speed stirring apparatus of Fig. 6.

The burette, B, was graduated from 0-50 ml. by 0.1 ml. divisions. Readings of volume were made by adjusting the height of the mercury reservoir by means of a "Lab-Jack" until the pressure was the same in the burette and the compensating tube, C. At this point, the mercury in the small U-tube was just making contact with the platinum terminal D, and the light bulb was flickering. Using this method, readings could be obtained quickly, and were reproducible to better than 0.1 ml. under equilibrium conditions.

A typical run was performed as follows: Benzene (75 ml. was placed in flask A and the apparatus assembled, all joints being lubricated with silicone grease and spring loaded. The flask was clamped in a thermostat at 25°, and set

stirring at the desired speed, usually 3,000 r.p.m. (measured with a stroboscope). Water from the thermostat was circulated through a glass jacket surrounding the burette and compensating tube.

Oxygen, saturated with benzene vapour at 25°, was passed in through tap E and out through tap F until the apparatus was thoroughly flushed. These taps were then closed. Taps G and H were left permanently open.

When the system was equilibriated, i.e. when the burette reading had not changed by 0.1 ml. in 5 min., the borane solution (1 ml.) was injected and a stop clock was started. Readings were taken at the shortest possible time intervals (ca 6 seconds), or at every ca 1 ml. absorbed, depending on the rate of oxygen uptake. At the end of each run, the solution was analysed for peroxide.

iii) Modifications for Pressure and Temperature Variation. When the temperature was to be varied, flask A was clamped in a vacuum flask containing a cooling mixture at the required temperature. Chlorobenzene was the reaction solvent down to -40°, and below that, ether was used. The borane solutions were similarly made up either in chlorobenzene or ether.

When the gas pressure was varied from atmospheric, the apparatus was first set stirring and flushed with oxygen as usual. Taps E and G were connected to a common source of vacuum, and to a manometer. The mercury reservoir was lowered as the pressure was slowly reduced to the required level. Tap E was then closed, leaving G open to the manometer to indicate any leak.

The pressure differential made injection of a sample difficult, and although a 5" 26 gauge needle was used, the apparent total amount of oxygen taken up varied considerably, even between runs at the same pressure.

iv) Runs at the Natural Rubber Producers' Research Association

The reaction flask already described was used in conjunction with a burette capable of more rapid and accurate response than that of Fig. 7. A U-tube containing a conducting liquid (Lockheed brake fluid) was open to the burette at one end and to the atmosphere at the other, and was just touching a contact to complete a circuit at equilibrium. A clock, built into the apparatus, was started on injection of the borane, and every time the brake fluid broke contact due to gas absorbtion, a motor was activated, advancing a piston along the burette by a known amount. The times at which this happened were punched out on paper tape. It

was found that only at -74^{o} did the boranes under consideration react slowly enough for automatic operation.

b) Results:-

The majority of the results are shown in Tables 8 and 9, and Graphs 1-23. Oxygen absorbtions are given as mols. per boron atom. At least twenty readings were taken for the majority of the curves; however, for compounds which took up oxygen very rapidly it was not possible to obtain a first reading until almost 1 mol. had been absorbed (ca 10 secs.). All the fast reactions were shown to be independent of stirring speed at the speeds used.

c) Reactions with Inhibitors in the Presence of Oxygen:-

i) Reaction with Galvinoxyl. In a flask fitted with a sintered glass gas bubbler, a serum cap and a double surface condenser serving as gas outlet were placed galvinoxyl (1.00g., 2.37 mmoles) and benzene (500 ml.). After flushing thoroughly with nitrogen, tri-s-butyl borane (5.6 ml., 4.3g., 23.6 mmoles) was injected into the solution, which was stirred magnetically. Oxygen was passed through the solution.

After 90 minutes the very dark red colour was perceptibly lighter, and during the next ten minutes the solution became

bright orange, with considerable oxygen absorbtion. The benzene was removed at the pump, and the residual red oil pumped hard at 10^{-3} mm and 30^{o}. A pale yellow distillate (2.2g.) was obtained, which rapidly hydrolysed on exposure to air, and corresponded on the basis of peroxide and boron titrations to 90% $(BuO)_{2.6}B(OOBu)_{0.4}$ with 10% of a non-peroxidic product containing no boron; this was probably benzene.

The remaining oil was hydrolysed and extracted with ether to remove any product from boric acid. On drying, and removal of the ether, a clear red oil was obtained. This was shown by column chromatography to contain at least seven components, but it was not possible to obtain a separation by this method. Of the stationary phases tried, only two (activated alumina and silica gel) gave a good separation, and both of these reacted with the components, so that they could not be eluted from the column.

ii) Reaction with a Nitroxide. Triacetonamine nitroxide (1g., 5.85 mmoles) and triisobutyl borane (1.45 ml., 1.07g., 5.88 mmoles) in dry ether (250 ml.) were stirred under a static oxygen atmosphere in a flask fitted with a condenser and a serum cap. At intervals, samples were withdrawn by syringe and the disappearance of colour followed in an E.E.L.

absorbtiometer. After 9, 23 and 34 hours further samples of borane (1.45 ml.) were added, and after 48 hours the orange colour of the nitroxide had completely disappeared.

The ether was removed at the pump, leaving an almost colourless oil. Semicarbazide hydrochloride (1.0g.) and crystalline sodium acetate (1.6g.) in water (10 ml.) were added. A few drops of ketone-free methanol were added to make the liquids miscible, and the solution vigorously shaken. The resulting white precipitate was filtered and recrystallised from aqueous ethanol, yielding white flakes (0.62g., 37% of theory, m.p. 219° (d)). The mass and NMR spectra are shown in Figs. 8 and 9. (See discussion p.184).

d) Reactions with Inhibitors in the Absence of Oxygen:-

i) Reaction with Galvinoxyl. Galvinoxyl (0.4 mg., 0.001 mmole) was placed in a UV cell (10 mm. path length) which was sealed with a serum cap and thoroughy flushed with nitrogen, using two hypodermic needles. Benzene (2 ml.), degassed three times at 10^{-3} mm, was introduced to the cell by syringe, followed by a solution of tri-s-butyl borane in benzene (0.12 mmoles). The rate of disappearance of colour was followed with an E.E.L. absorbtiometer, using filter 601. A control containing no borane was also

prepared. The results are shown on Graph 24. A similar reaction between tri-n-butyl borane (0.1 mmole) and galvinoxyl (0.0095 mmole) is plotted on Graph 25.

ii) Reaction with DPPH. The reaction between DPPH (0.6 mg., 0.0015 mmole) and tri-s-butyl borane (0.06 mmole) was followed as above, and the results are shown in Graph 26.

iii) Reaction with Triacetonamine Nitroxide. A mixture of triacetonamine nitroxide (0.1g., 0.59 mmole) and triisobutyl borane (0.15 ml., 0.62 mmole) in benzene (2 ml.) prepared as above in the absence of oxygen did not show any reaction, measured colourimetrically, over 45 hours.

iv) Reaction with Phenothiazine. Phenothiazine (0.4g., 2.0 mmole) in benzene (14 ml.) was degassed three times at 10^{-3} mm. Tri-s-butyl borane (1.75 ml., 7.2 mmole) was added, causing partial precipitation of the phenothiazine. The yellow solution remained unchanged for 48 hours. On exposure to oxygen after this time, the solution rapidly became green, and over a period of two days turned deep pink.

Table 8 - Induction Periods with Galvinoxyl (G·)[a]

Borane	Conc. of G·, moles % [b]	Age of stock soln. of G·, hr. [c]	Injection prior to borane, min.	Induction period, min.
$(IPC)_2BBu^s$	0.1	d	<90	17
$(IPC)_2BBu^n$	1.4	d	<90	5.7
$(Bu^sBO)_3$	0.25	0	<1	126
Bu^i_3B	5.00	0	<1	57
Bu^i_3B	5.00	0	<1	58
Bu^n_3B	5.00	0	<1	retards
Bu^n_3B	10.00	0	<1	retards
Bu^s_3B	1.00	0 [e]	<1	21.3
Bu^s_3B	0.50	19 [e]	<1	11.5
Bu^s_3B	0.25	23 [e]	<1	5.0
Bu^s_3B	0.10	41 [e]	<1	2.0
Bu^s_3B	2.00	44 [e]	20	30
Bu^s_3B	2.00	66 [e]	1.0	33
Bu^s_3B	0.50	69 [e]	1.0	4.3
Bu^s_3B	1.00	0 [f]	<1	18
Bu^s_3B	1.00	21 [f]	<1	20.5
Bu^s_3B	1.00	26 [f]	60	15
Bu^s_3B	1.00	144 [f]	<1	7.8
Bu^s_3B	1.00	240 [f]	<1	8.0
Bu^s_3B	1.00	336 [f]	<1	13.5
Bu^s_3B	5.00	0	<1	54
Bu^s_3B	5.00	0	<1	50.5
Bu^s_3B	10.00	0	<1	78

[a] Those results obtained with solutions of galvinoxyl left in contact with oxygen for more than a few minutes were irreproducible and are not included in the table.

[b] Per boron atom. [c] At time of injection into reaction flask.

[d] Solid added prior to equilibriation of burette.

[e] Solution made up under nitrogen and stored at -20° under nitrogen.

[f] Solution containing 6% added hydrogalvinoxyl made up under nitrogen and stored at -20° under nitrogen.

Table 9 - Quantity of Peroxide Obtained

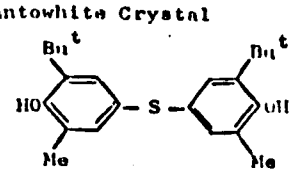
Borane	Inhibitor, moles %	O_2 uptake, mol.	Peroxide, mol. (after time, hrs.)
$(IPC)_2BBu^s$	None	1.45	1.45 (1.5)
$(IPC)_2BBu^n$	"	1.45	1.43 (1.0)
$Bu^sB(OH)_2$	"	0.83	0.76 (32)
$(Bu^sBO)_3$	"	0.96 [a]	0.96[a] (2.2)
$(Bu^s{}_2B)_2O$	"	1.92	1.72 (14)
$Bu^s{}_2BOBu^s$	"	1.00	0.93 (1.3)
$PhCH(Me)B(OH)_2$	"	0.84	0.77 (0.5)
$Bu^n{}_3B$	"	1.91	1.90 (1.5)
$Bu^s{}_3B$	"	1.91	1.91 (0.4)
$(IPC)_2BBu^s$	Galvinoxyl, 0.1	1.43	1.46
$(IPC)_2BBu^n$	" 1.4	1.42	1.38
$(Bu^sBO)_3$	" 0.25 [a]	0.70 [a]	0.70 [a]
$Bu^s{}_3B$	" 1.0	1.91	1.90
$(IPC)_2BBu^s$	DPPH, 0.3	1.44	1.41
$Bu^s{}_3B$	" 5.0	1.90	1.85
"	$Bu^t{}_2NO^\bullet$, 10.0	1.91	1.85
$(IPC)_2BBu^s$	Phenothiazine, 0.3	1.36	1.31
$(IPC)_2BBu^n$	" 2.8	1.49	1.36
$Bu^s{}_3B$	" 10.0	2.00	1.53
"	$PhCH_2SH$, 5.0	1.98	1.81
"	$[(Pr^iO)_2PS_2]_2Zn$, 5.0	1.92	1.66
$(IPC)_2BBu^s$	$(Et_2NCS_2)_2Zn$, 0.75	1.40	0.95
"	$(Bu_2NCS_2)_2Cu$, 0.1	1.41	1.25
$(IPC)_2BBu^n$	" 1.3	1.54	1.07
$Bu^s{}_3B$	" 2.0	1.92	1.36
"	$(Pr^iO)_2PS_2Cu$, 5.0	1.90	1.23
"	Cu naphthenate, 5	1.91	1.34
"	Cu E.E.A.[b], 5.0	1.89	1.22
$Bu^s{}_3B$	Ph_2NH, 5.0	1.90	1.88
"	Ethyl 702, 10.0 [e]	1.89	1.72
"	BHT, 5.0	1.93	1.56
"	246Bu^t Phenol, 5.0	1.89	1.22
"	2OH4Bu^tPhenol[c], 11.0 [e]	1.91	1.84
"	BHA, 5.0	1.91	1.89
"	26MeO Phenol, 5.0	1.91	1.93
"	Ethyl 720, 10.0 [e]	1.91	1.80
"	Cyanamid 2246, 10.0 [e]	1.95	1.78
"	S. Crystal[d], 10.0 [e]	1.81	1.49
$Bu^i{}_3B$	NaOH [f]	0.98	Nil
$Bu^n{}_3B$	"	0.95	Nil
$Bu^s{}_3B$	"	1.80	0.76 (0.1)

[a] Per boron atom.

[b] Copper ethyl acetoacetate complex.

[c] p-t-Butyl catechol.

[d] Santowhite Crystal



[e] Of OH groups.

[f] Excess powered solid.

FIG. 7. GAS BURETTE.

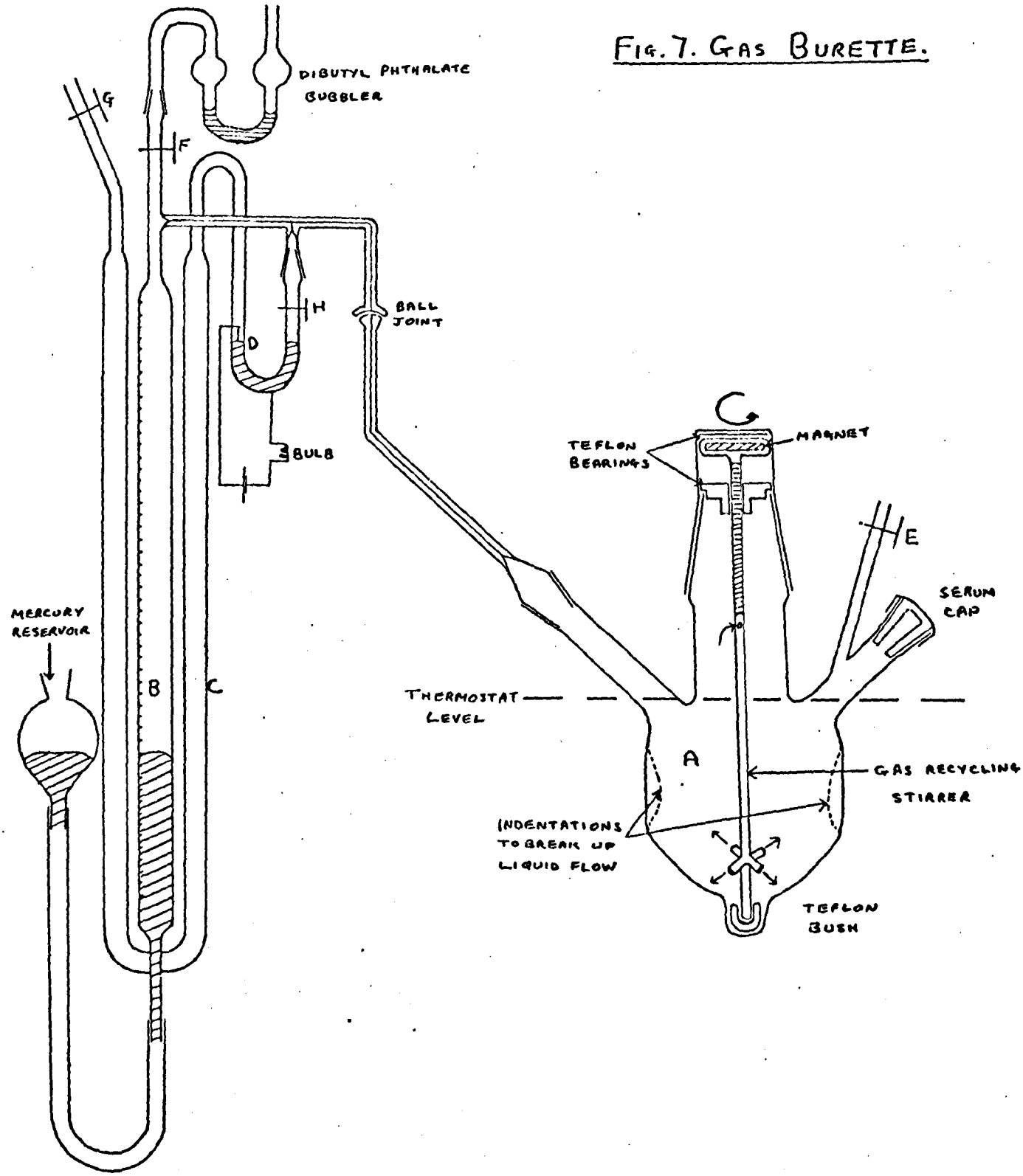


FIG. 8 MASS SPECTRUM OF $Bu^{i}_{3}B$ - NITROXIDE ADDUCT

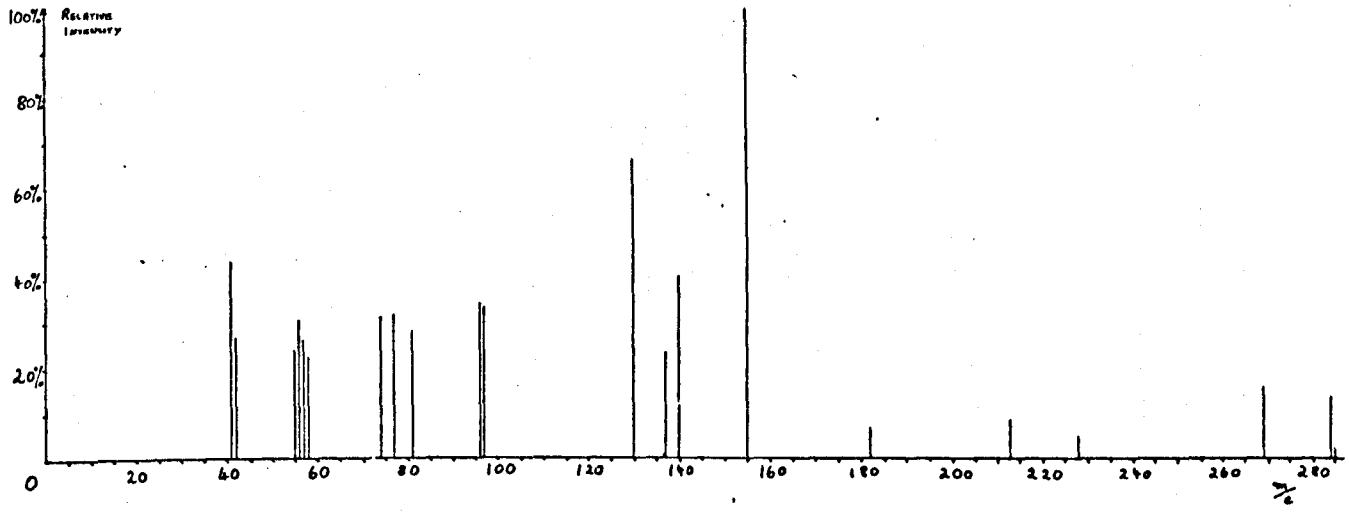
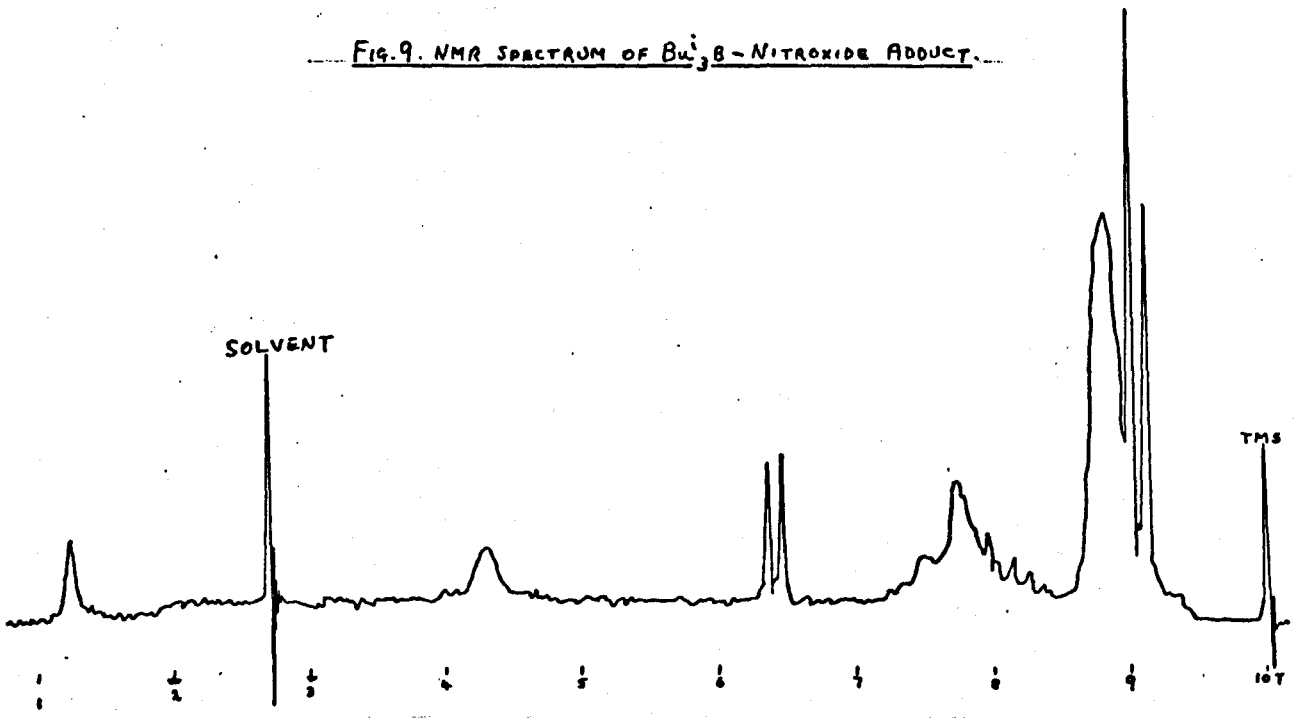
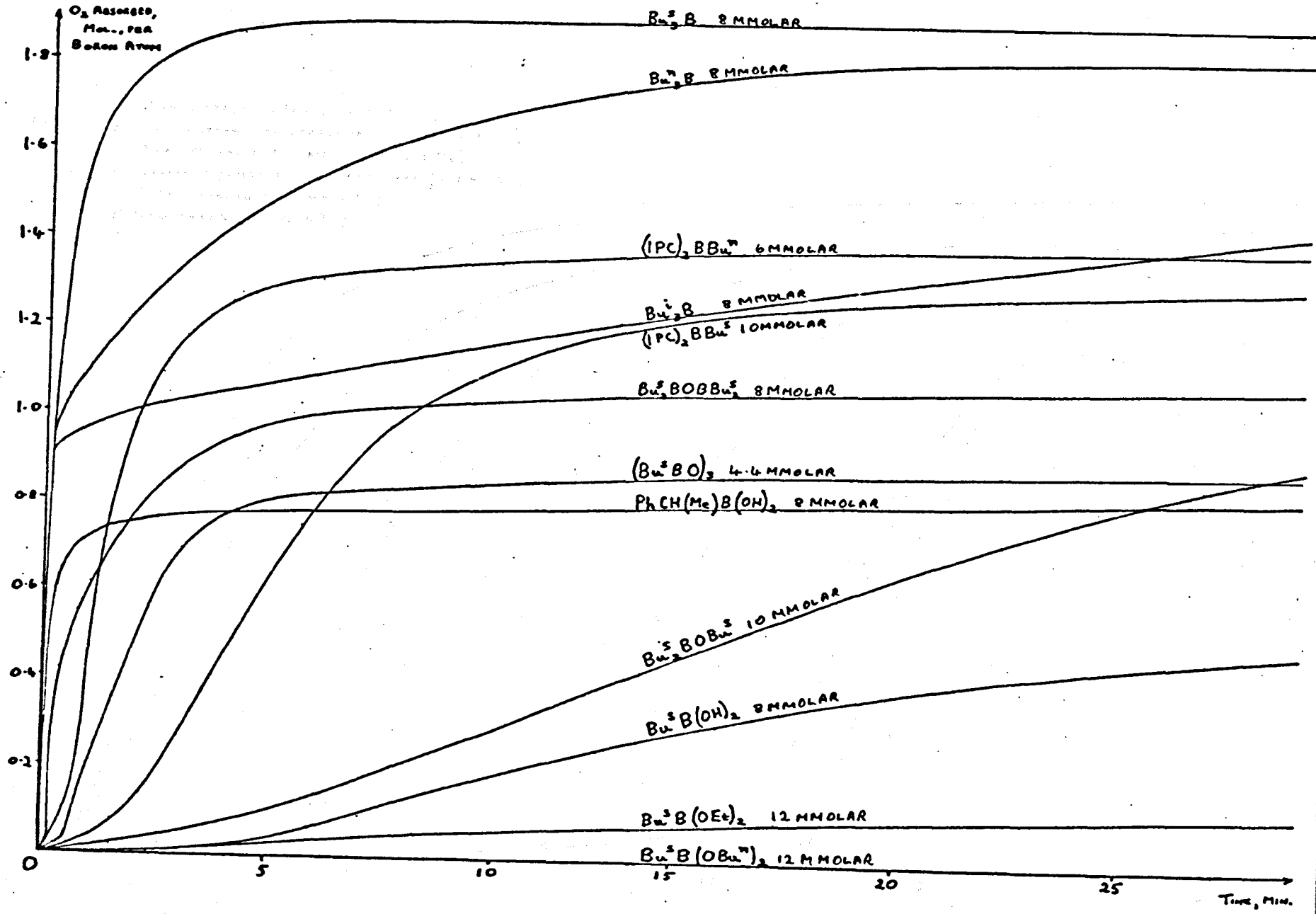
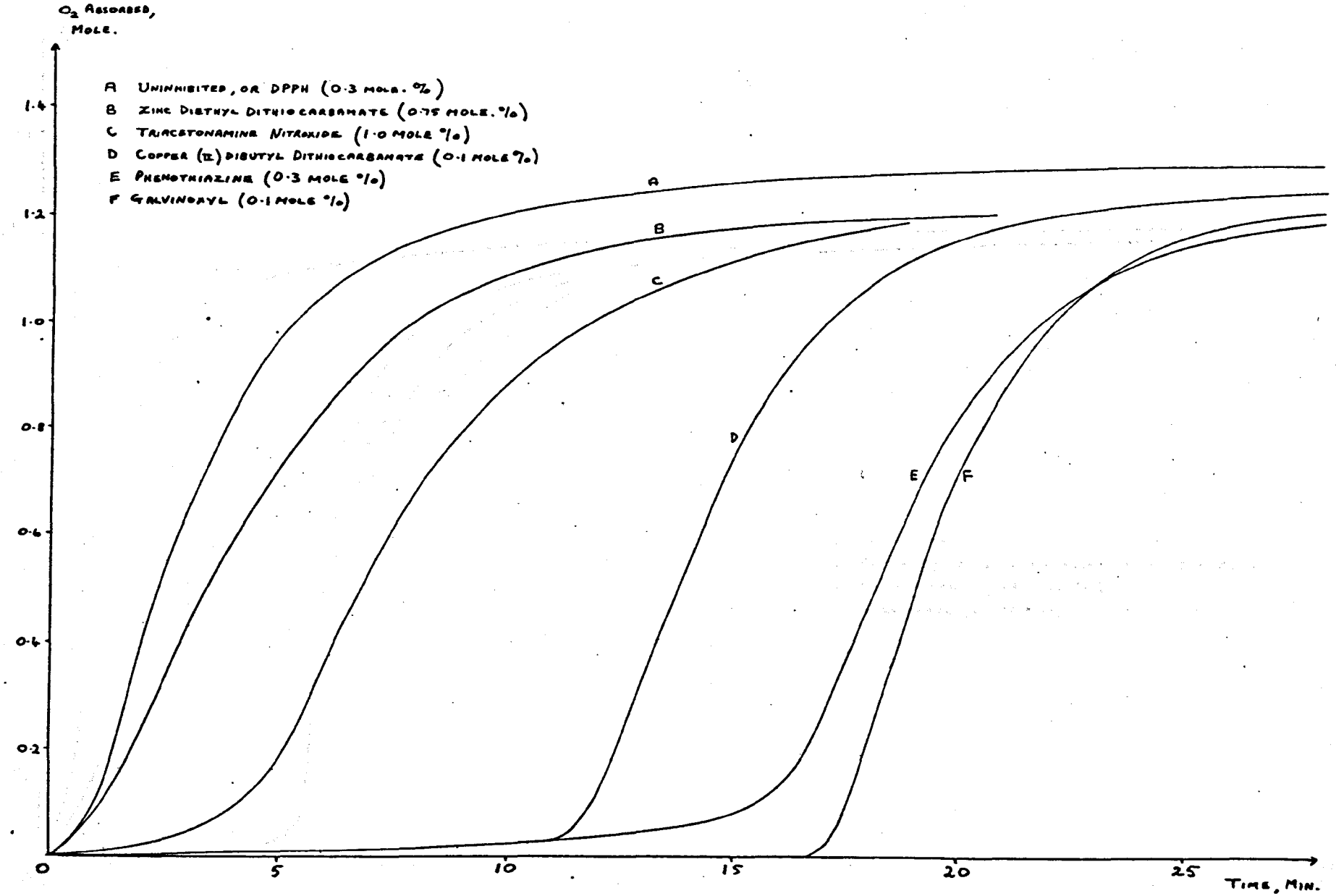


FIG. 9. NMR SPECTRUM OF $Bu^{i}_{3}B$ - NITROXIDE ADDUCT.

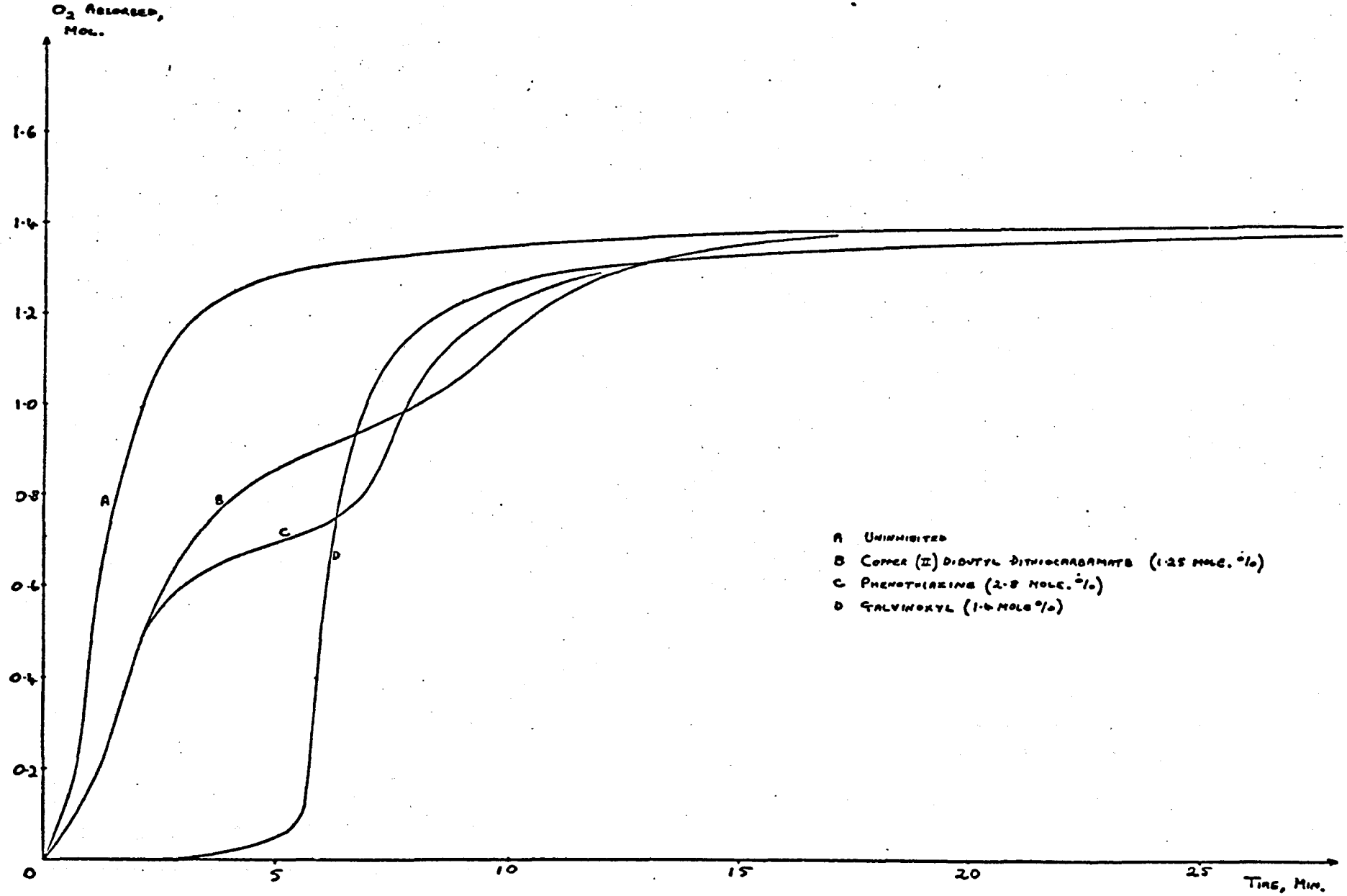




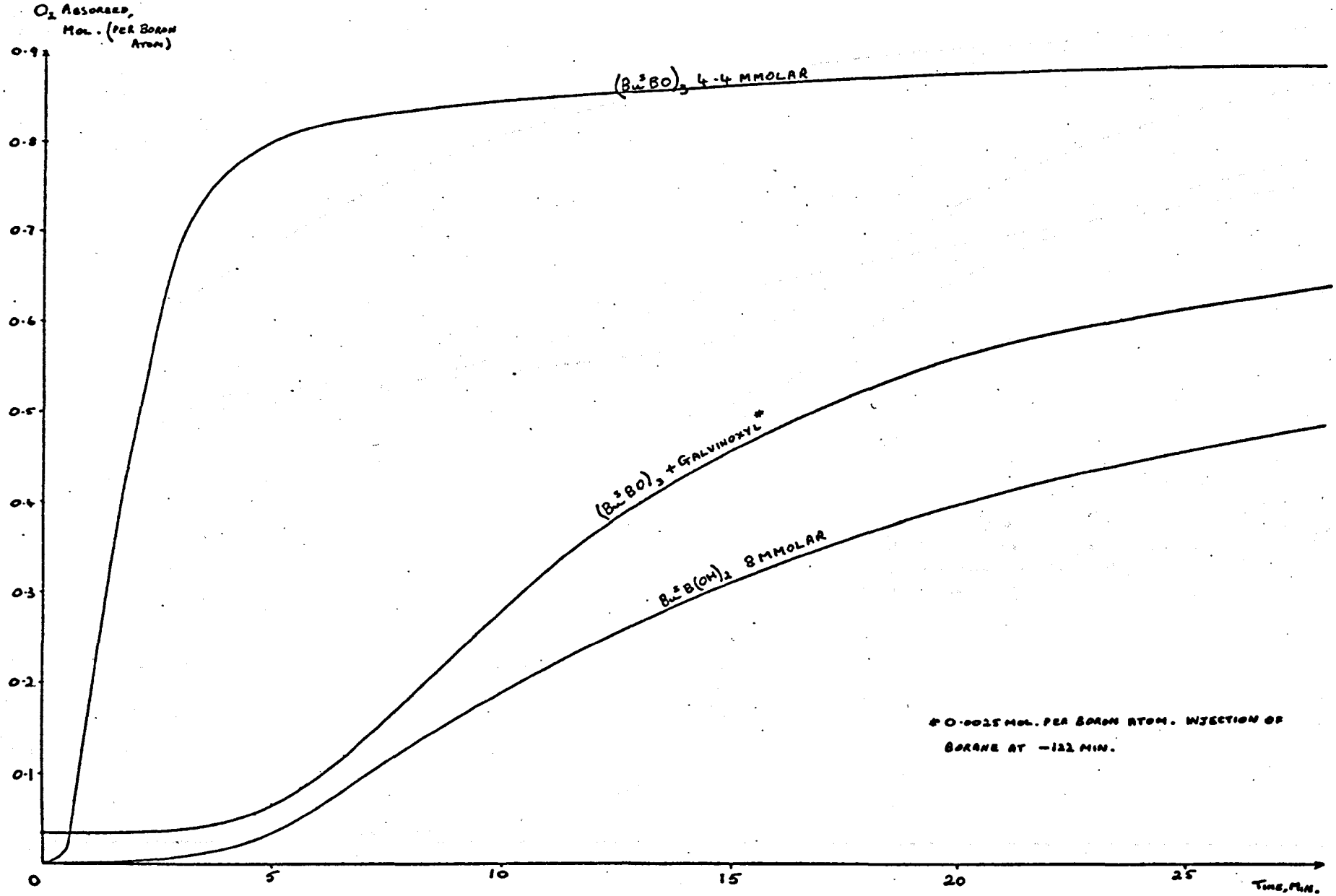
GRAPH 1. THE AUTOXIDATION OF ORGANOBORON COMPOUNDS IN BENZENE AT 25°C.



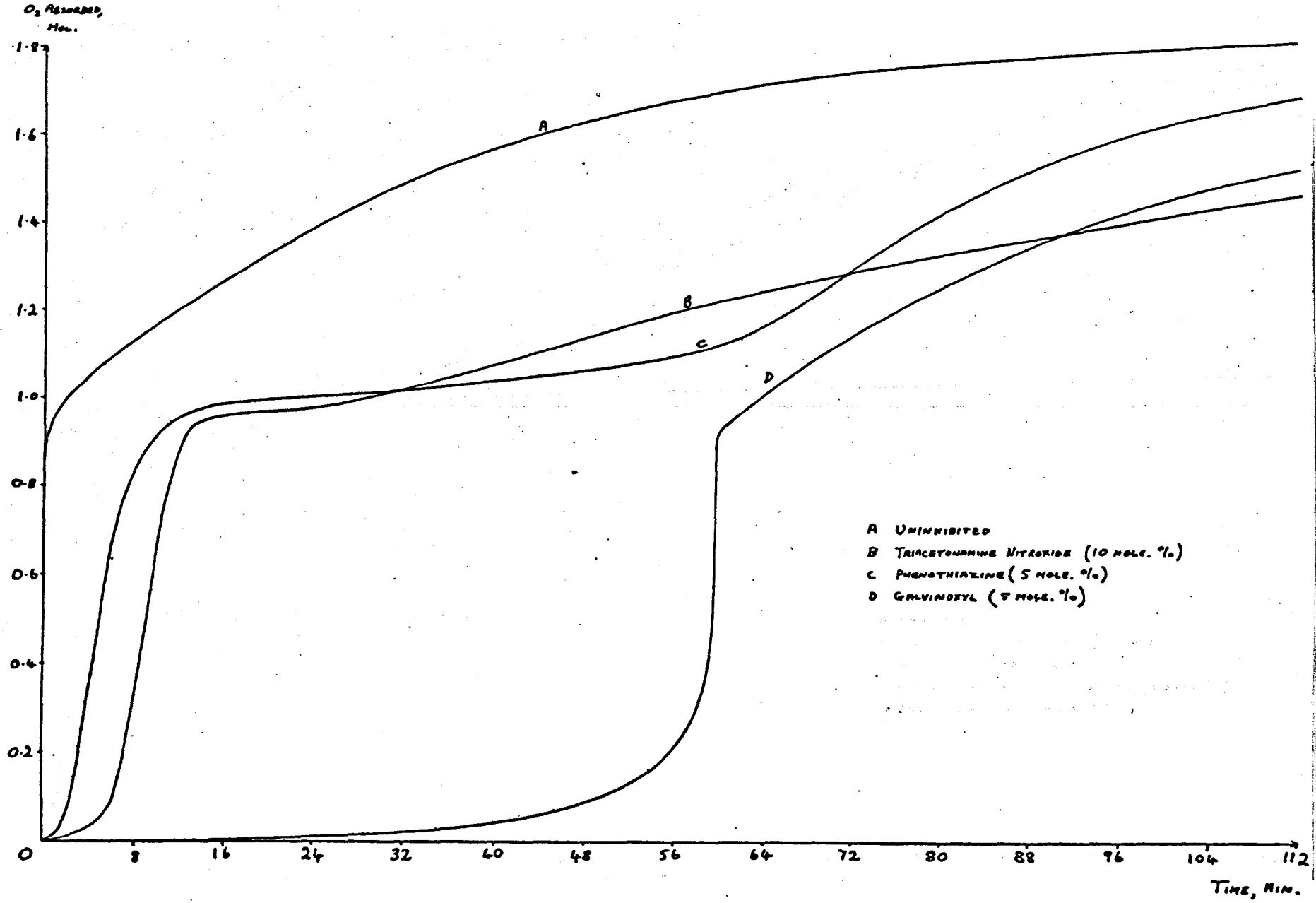
Graph 2. The Autoxidation of $(IPC)_2 BBu^s$ in Benzene (10 mmolar) at 25°C in the Presence of Inhibitors.



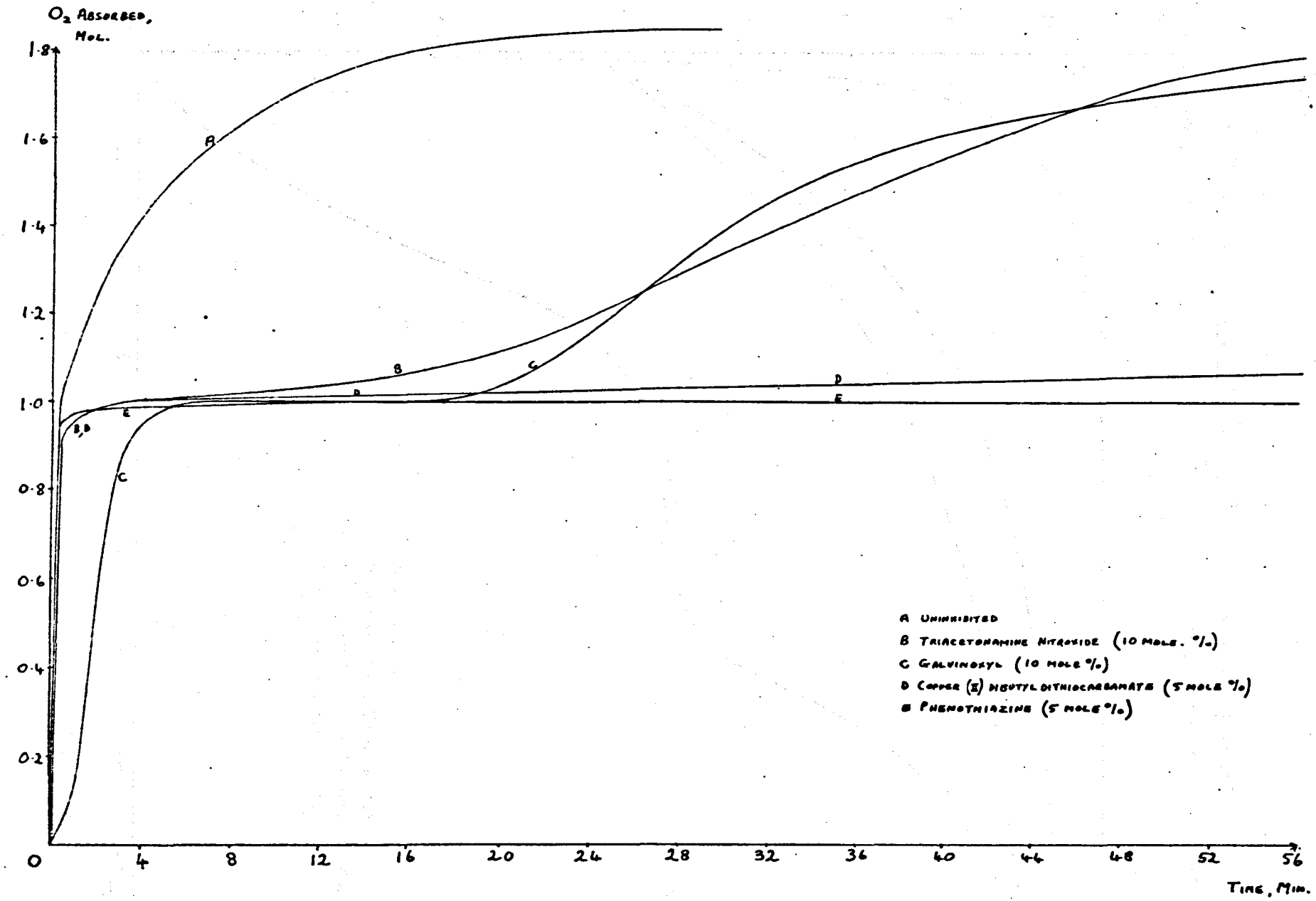
GRAPH 3. THE AUTOXIDATION OF $(IAc)_2$ B Bu^n IN BENZENE (6MMOLAR) AT 25°C IN THE PRESENCE OF INHIBITORS.



GRAPH 4. THE AUTOXIDATION OF s-BUTYLBORONIC ACID AND ANHYDRIDE IN BENZENE AT 25°C.

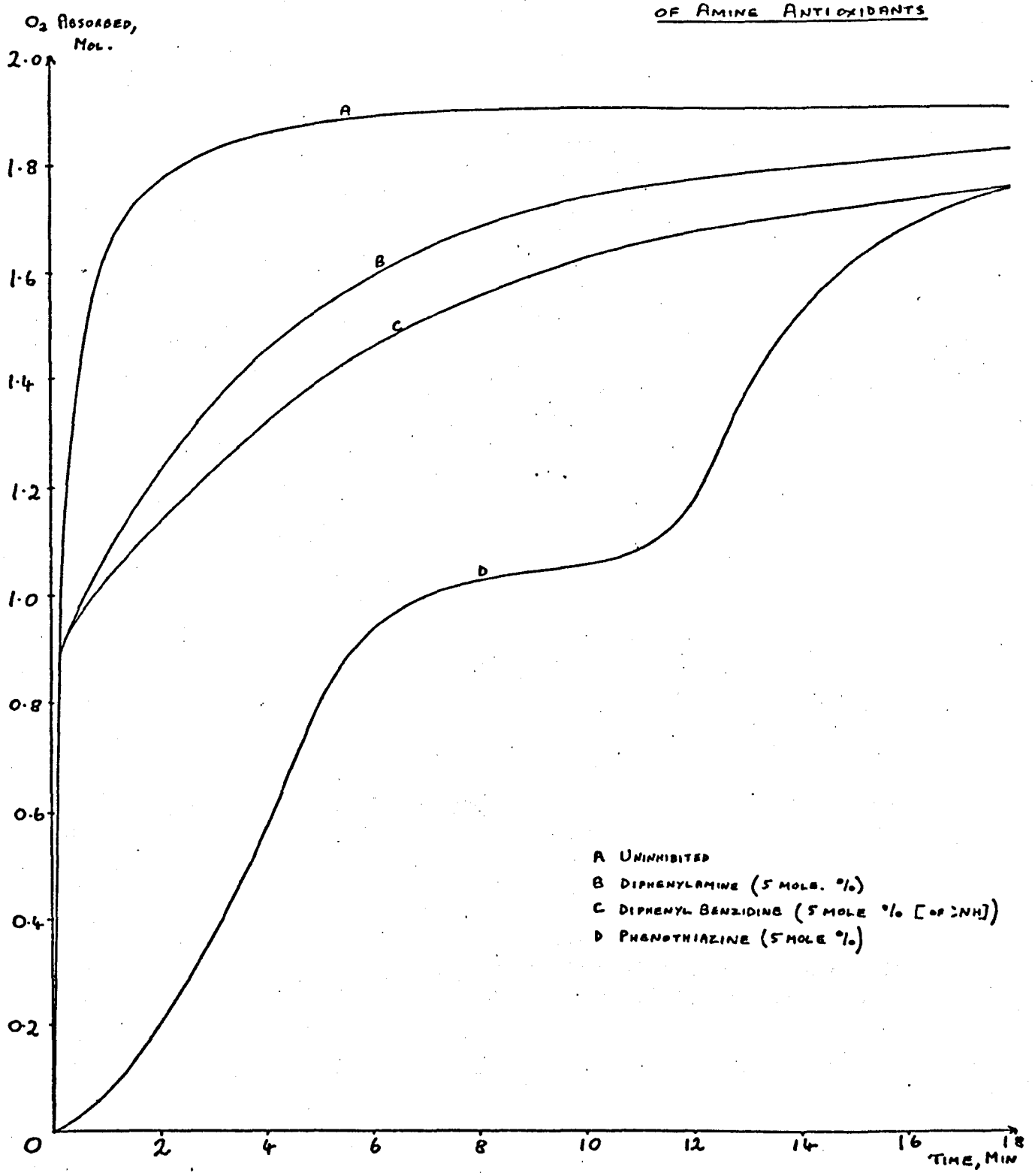


Graph 5. The Autoxidation of Triisobutyl Borane in Benzene (8 mmolar) at 25°C in the Presence of Inhibitors.

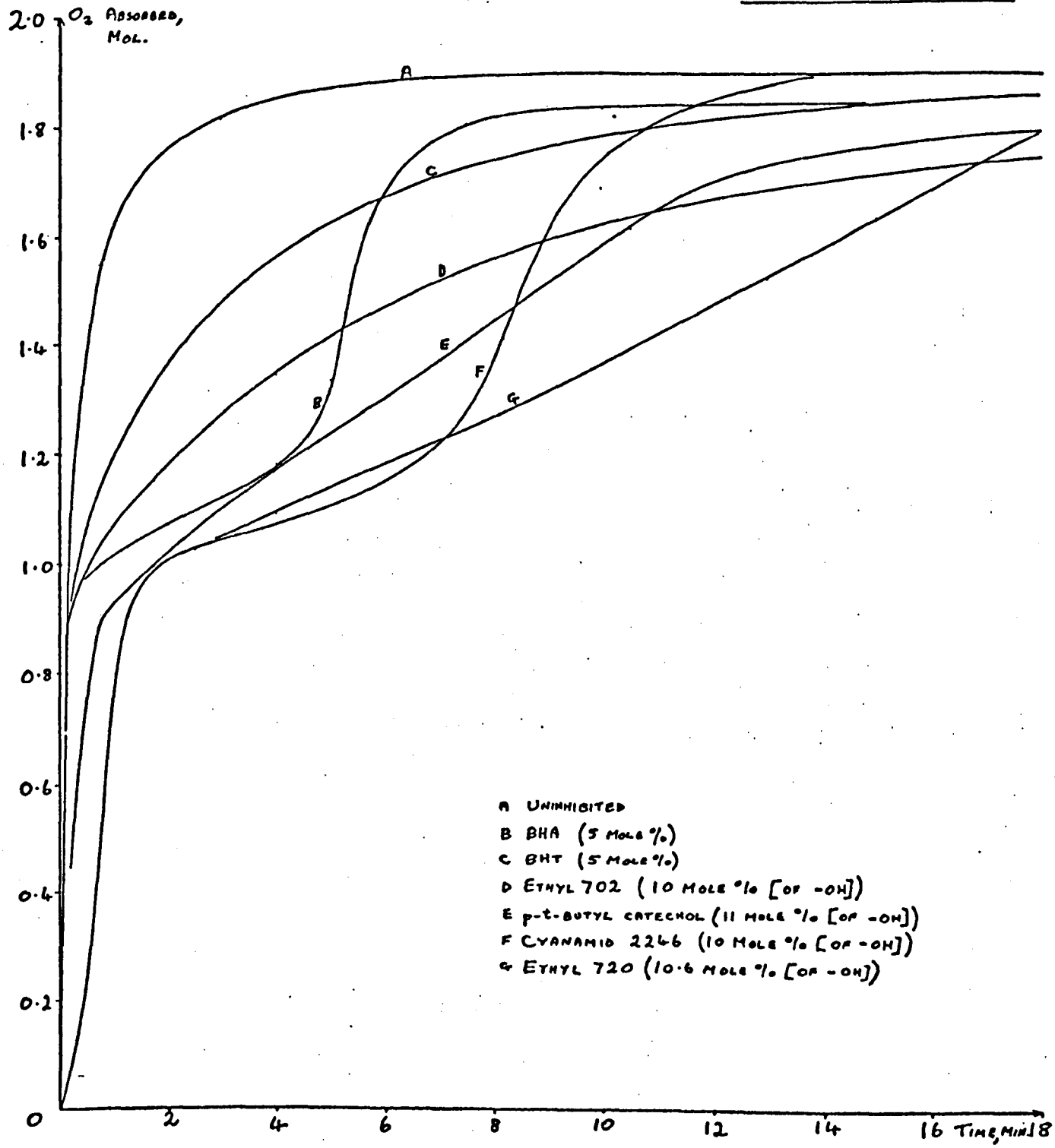


GRAPH 6. THE AUTOXIDATION OF TRI-n-BUTYL BORANE IN BENZENE (8 MMOLAR) AT 25°C IN THE PRESENCE OF INHIBITORS.

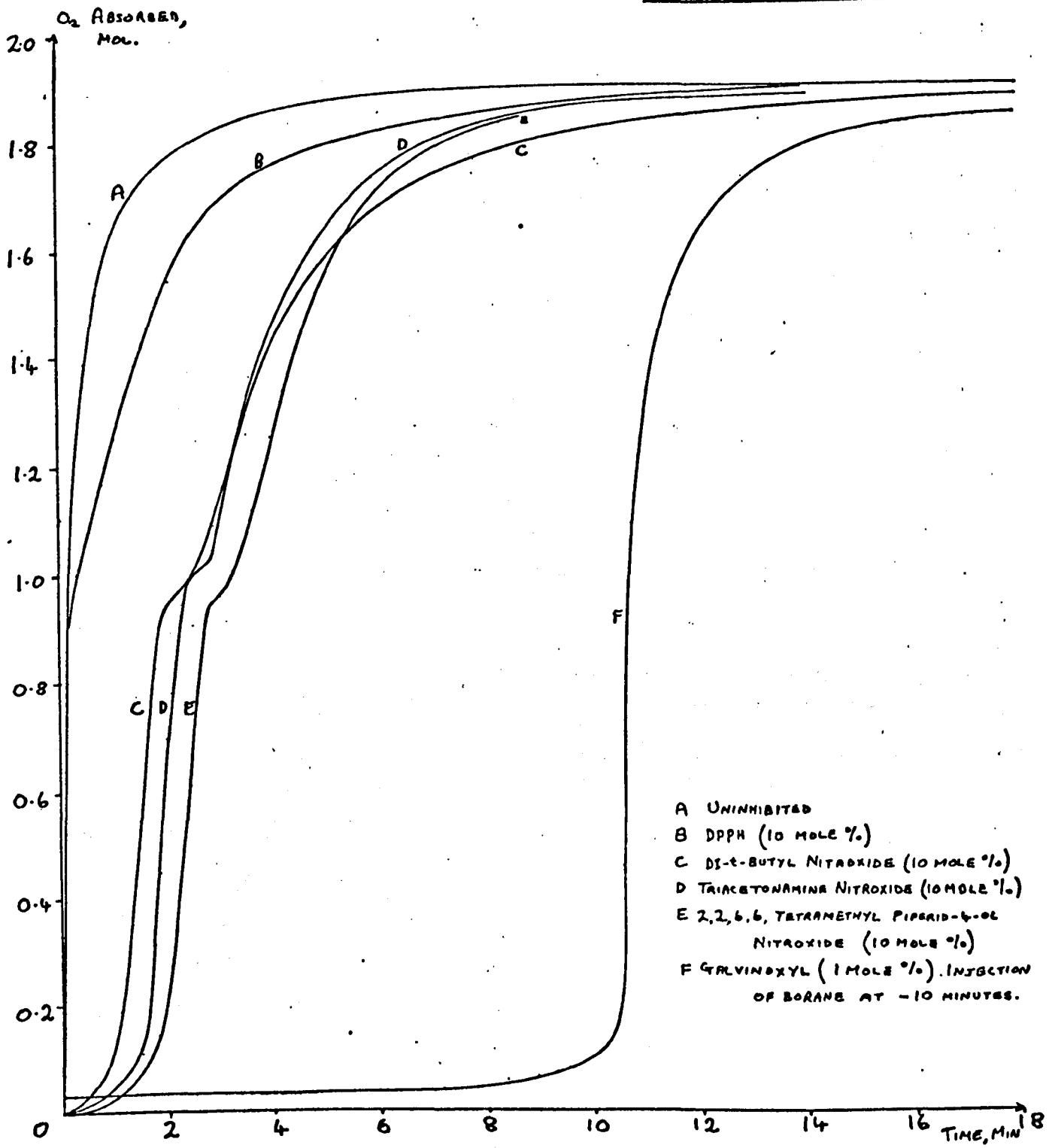
GRAPH 7. THE AUTOXIDATION OF Bu_3^sB IN BENZENE (6 MMOLAR) AT 25°C IN THE PRESENCE OF AMINE ANTIOXIDANTS



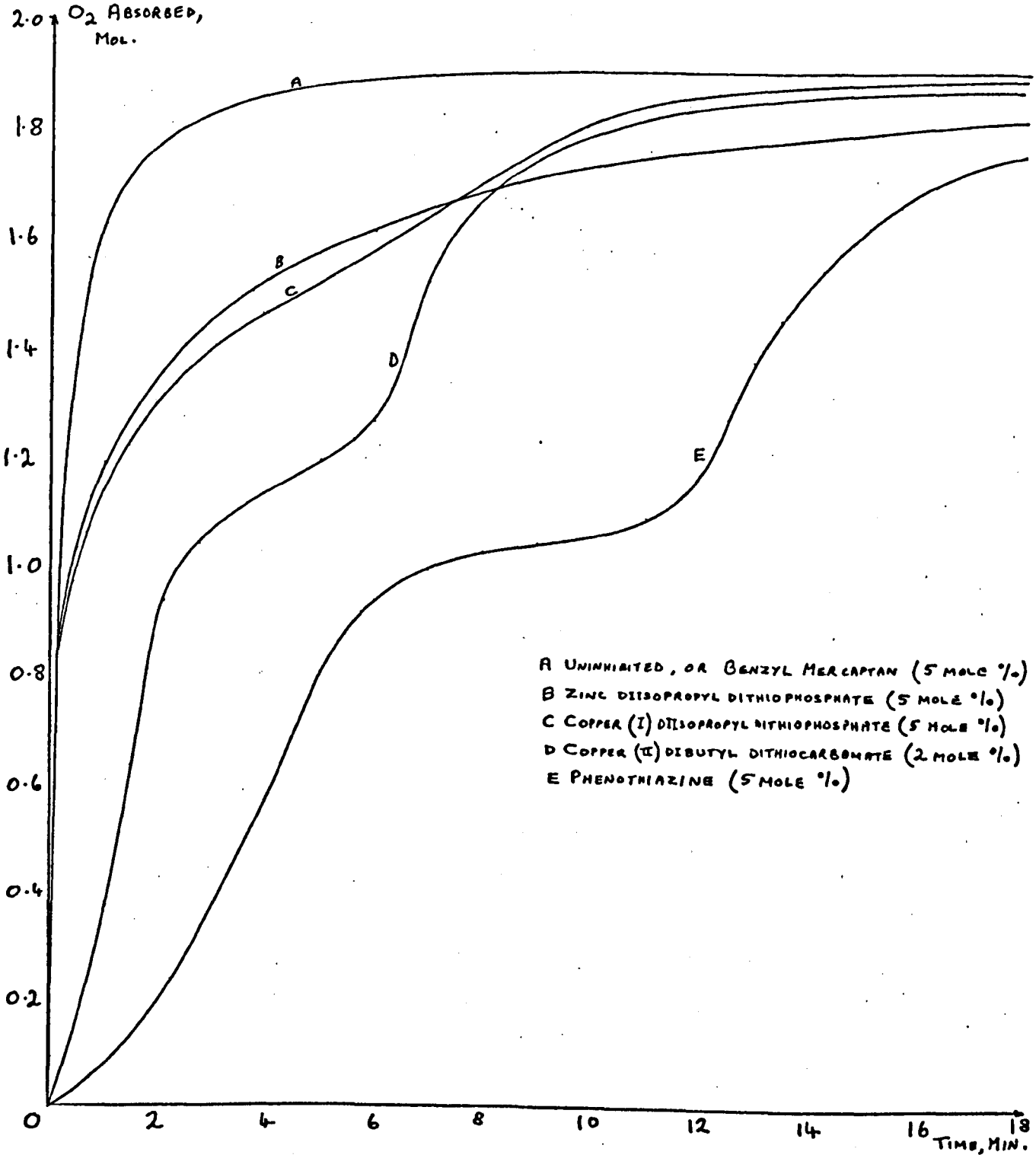
GRAPH 8. THE AUTOXIDATION OF Bu^s_3B IN BENZENE (6 MMOLAR) IN THE PRESENCE OF PHENOLIC ANTIOXIDANTS.



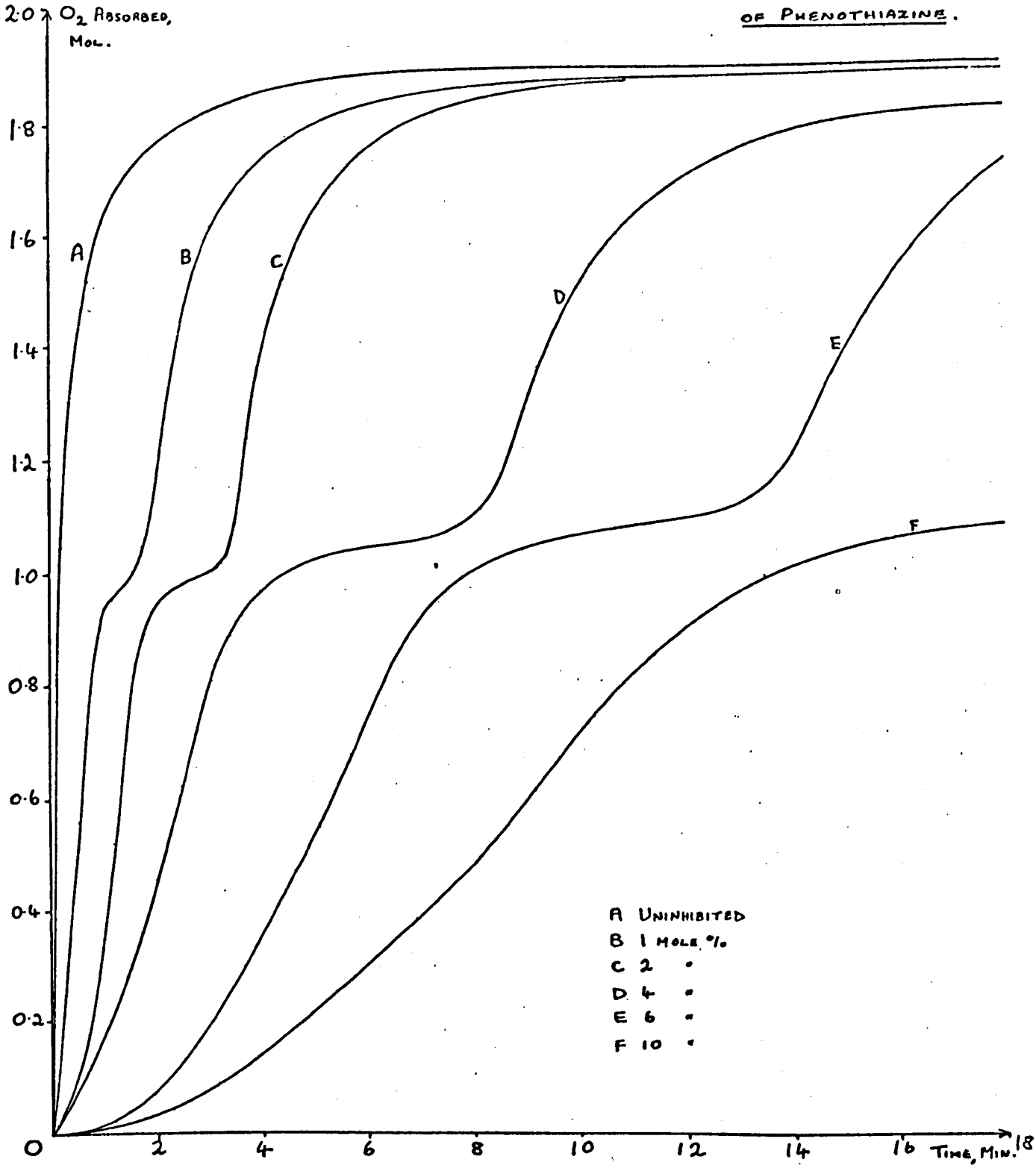
GRAPH 9. THE AUTOXIDATION OF $Bu^{s}_{3}B$ IN BENZENE (6MMOLAR) AT 25°C IN THE PRESENCE OF FREE RADICAL ANTIOXIDANTS.

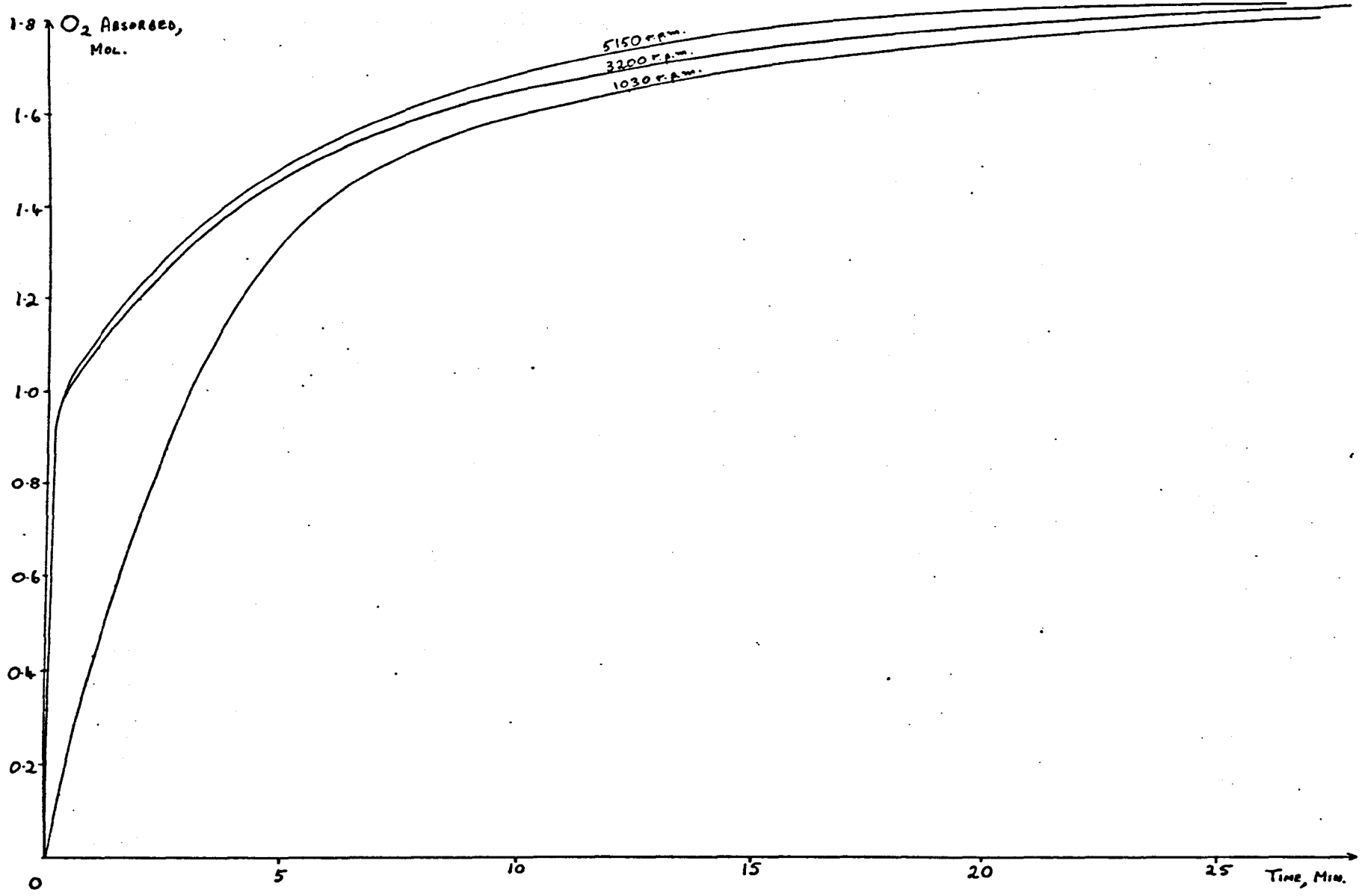


GRAPH 10. THE AUTOXIDATION OF Bu^s_3B IN BENZENE (6MMOLAR) AT 25°C IN THE PRESENCE OF SULPHUR AND PHOSPHORUS CONTAINING ANTIOXIDANTS.



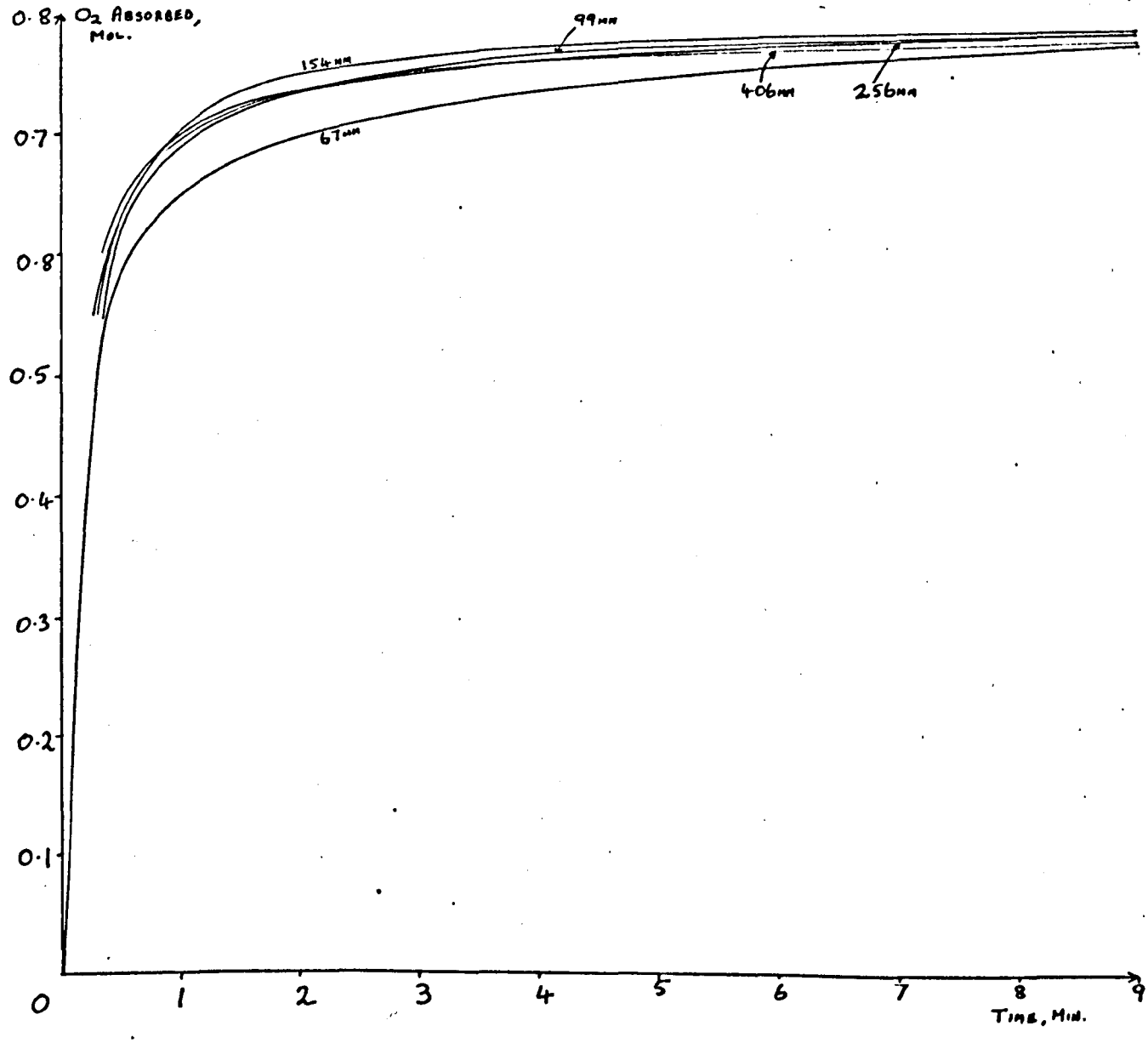
GRAPH II. THE AUTOXIDATION OF Bu^s_2B IN BENZENE (8 MMOLAR) AT 25°C IN THE PRESENCE OF PHENOTHIAZINE.



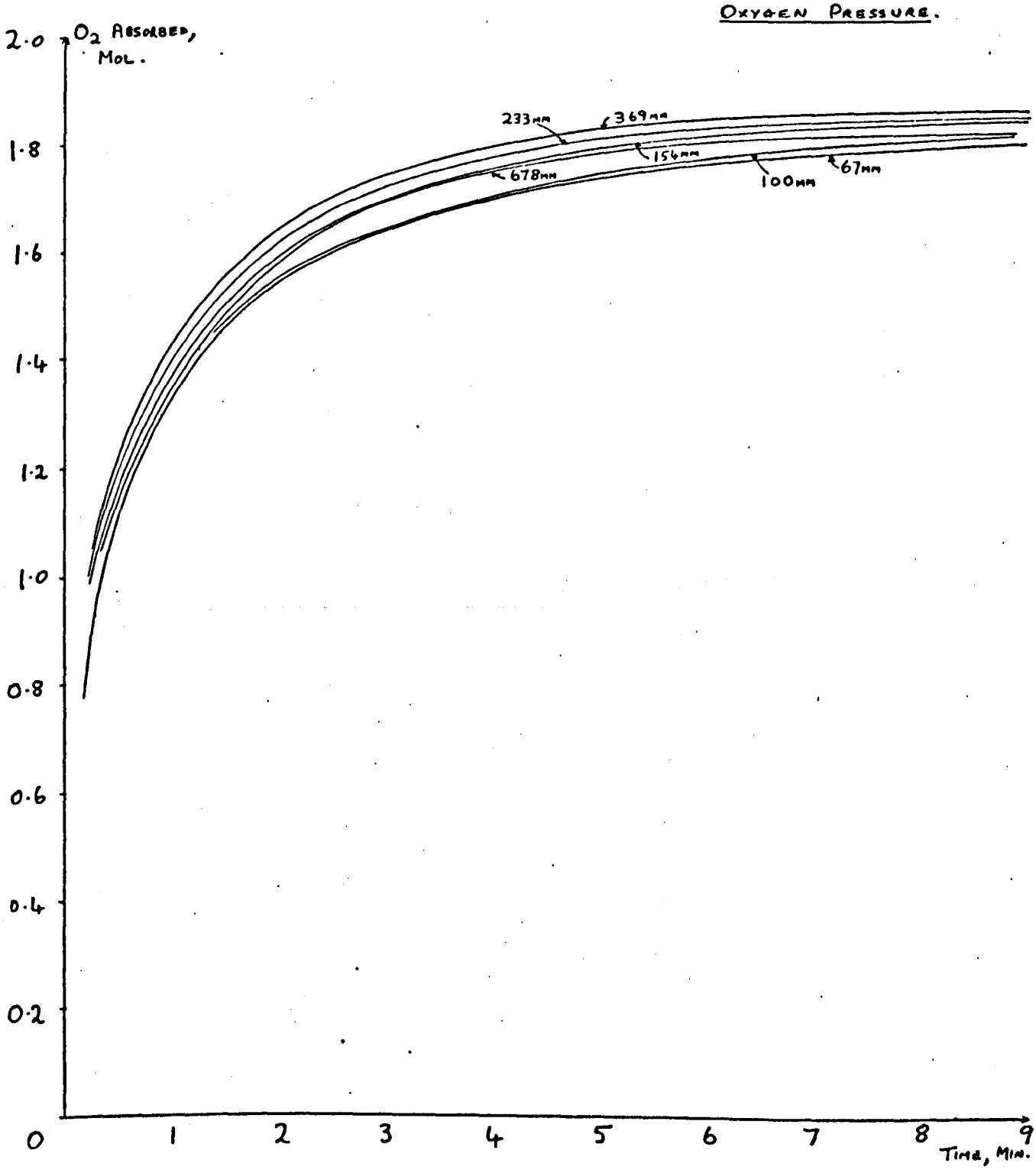


Graph 12. The Autoxidation of Bu^n_3B in Benzene (8 mmolar) at 25°C at Varying Stirrer Speeds.

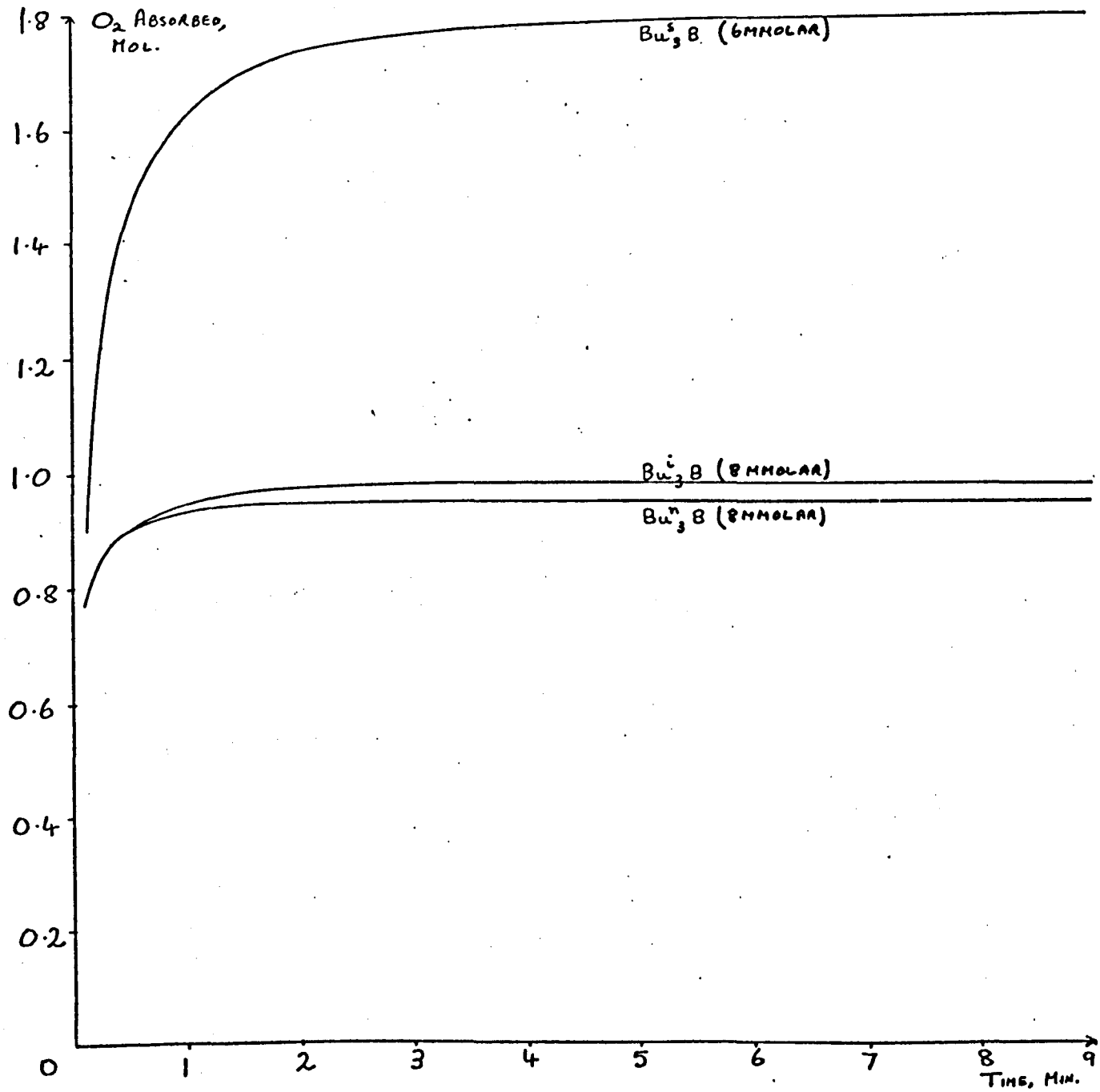
GRAPH 13. THE AUTOXIDATION OF 1-PHENYLETHYLBORONIC ACID IN BENZENE (4 MMOLAR) AT 25°C AT VARYING OXYGEN PRESSURES.

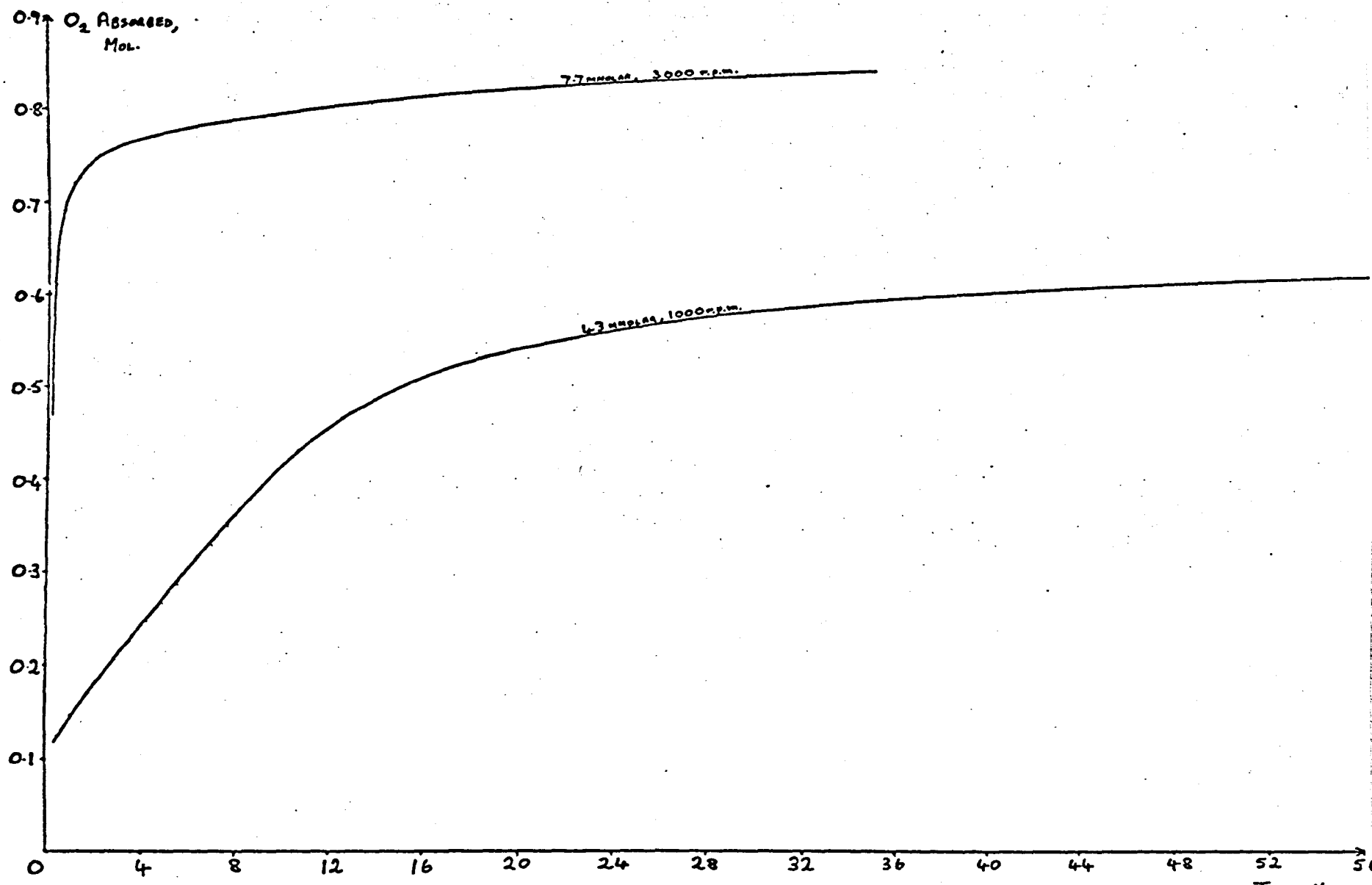


GRAPH 14. THE AUTOXIDATION OF Bu^s_3B IN BENZENE (2 MMOLAR) AT 25°C AT VARYING OXYGEN PRESSURE.



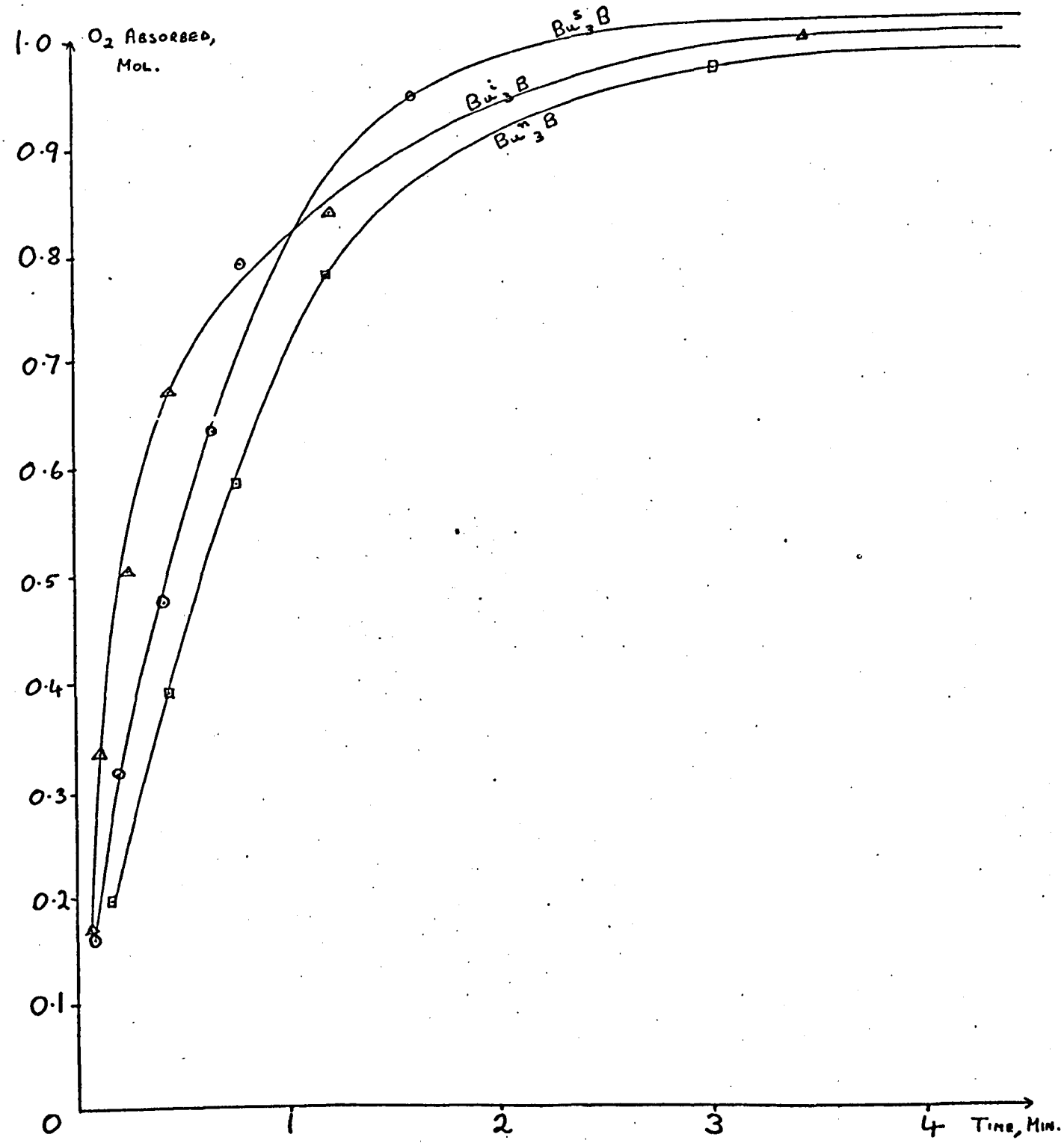
GRAPH 15. THE AUTOXIDATION OF Bu_3B IN BENZENE AT 25°C IN THE PRESENCE OF SODIUM HYDROXIDE.

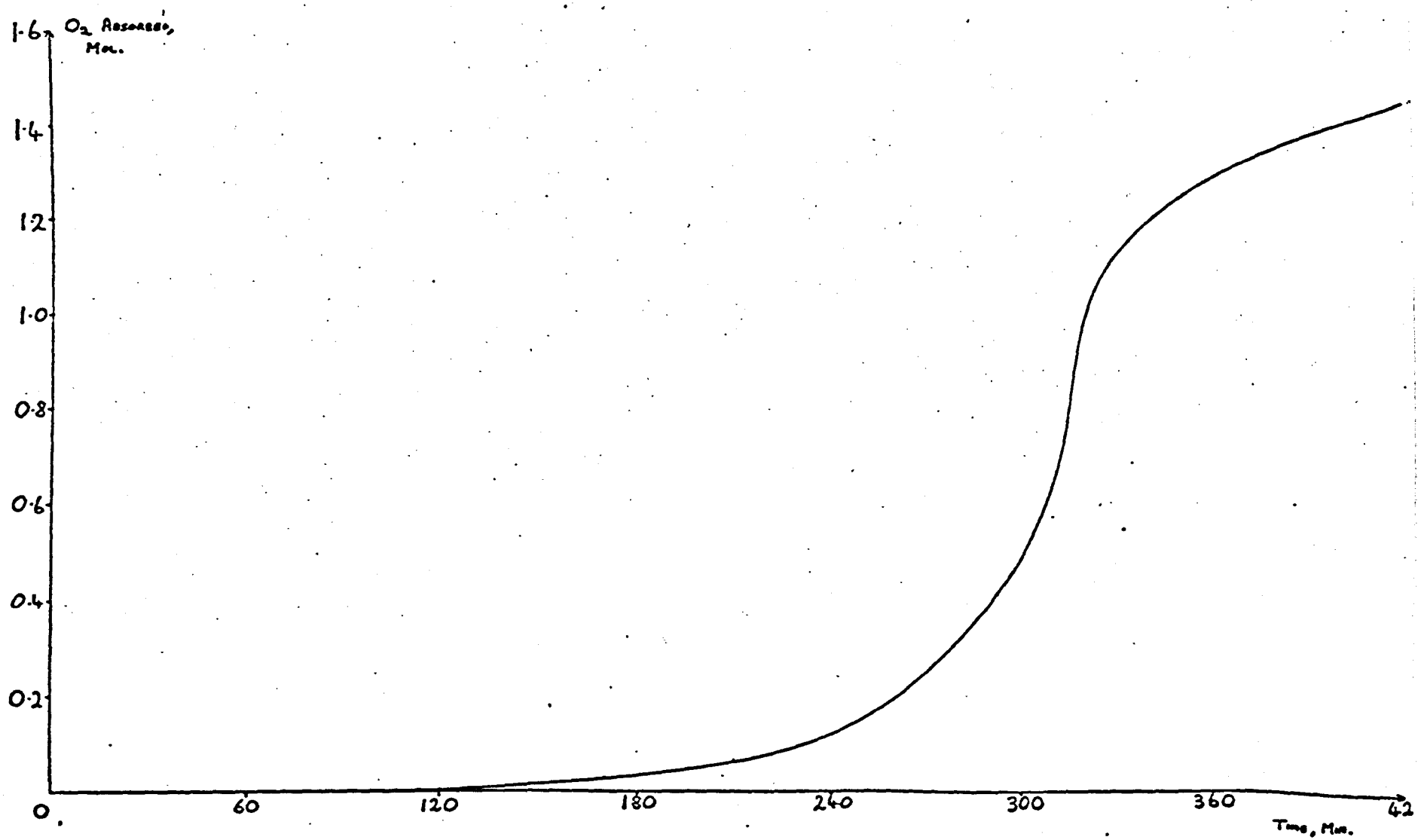




Graph 16. The Autoxidation of 1-Phenylethylboronic Acid in Benzene at 25°C.

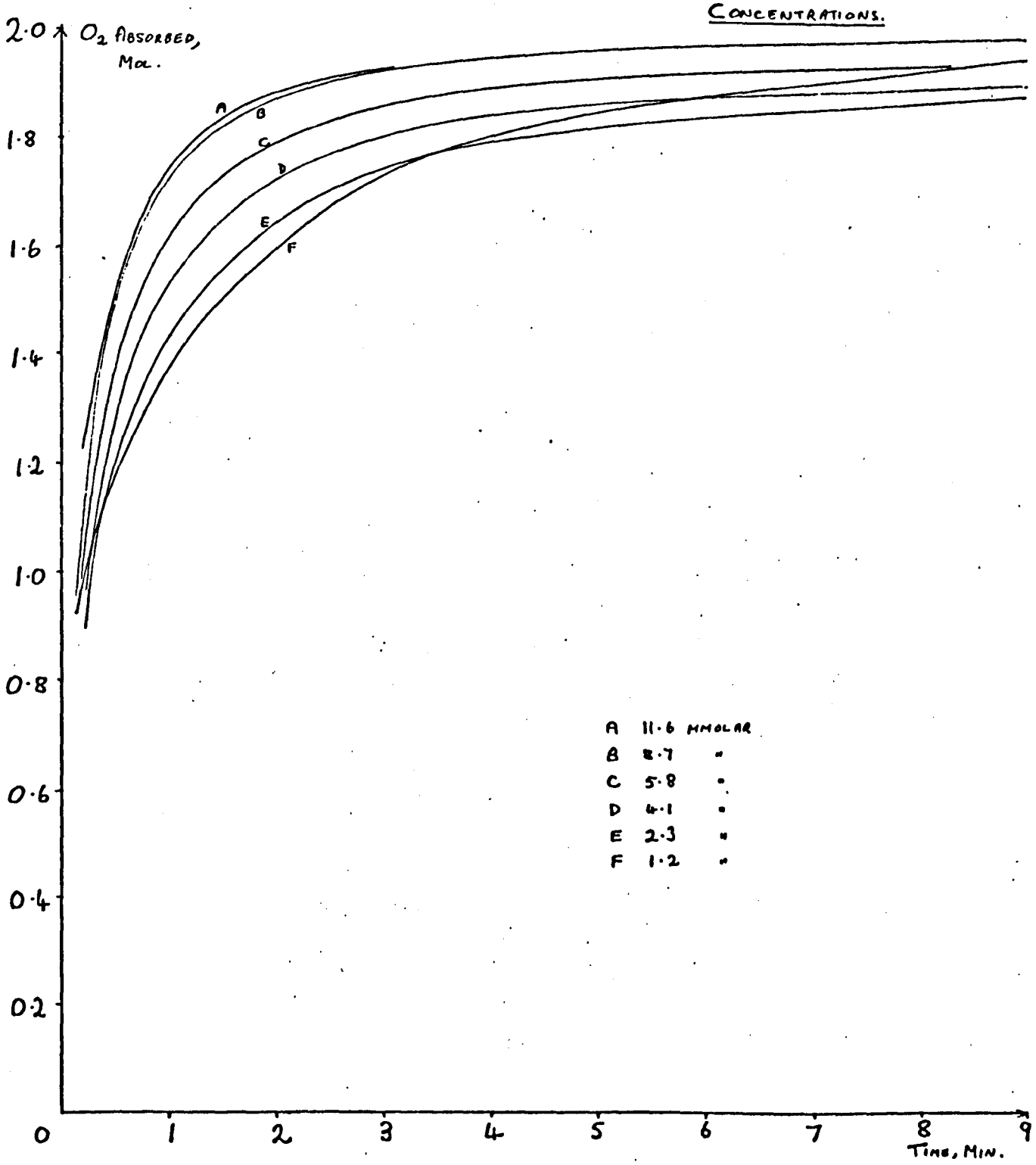
GRAPH 17. THE AUTOXIDATION OF Bu_3B IN ETHER (2 MMOLAR) AT −74°C.

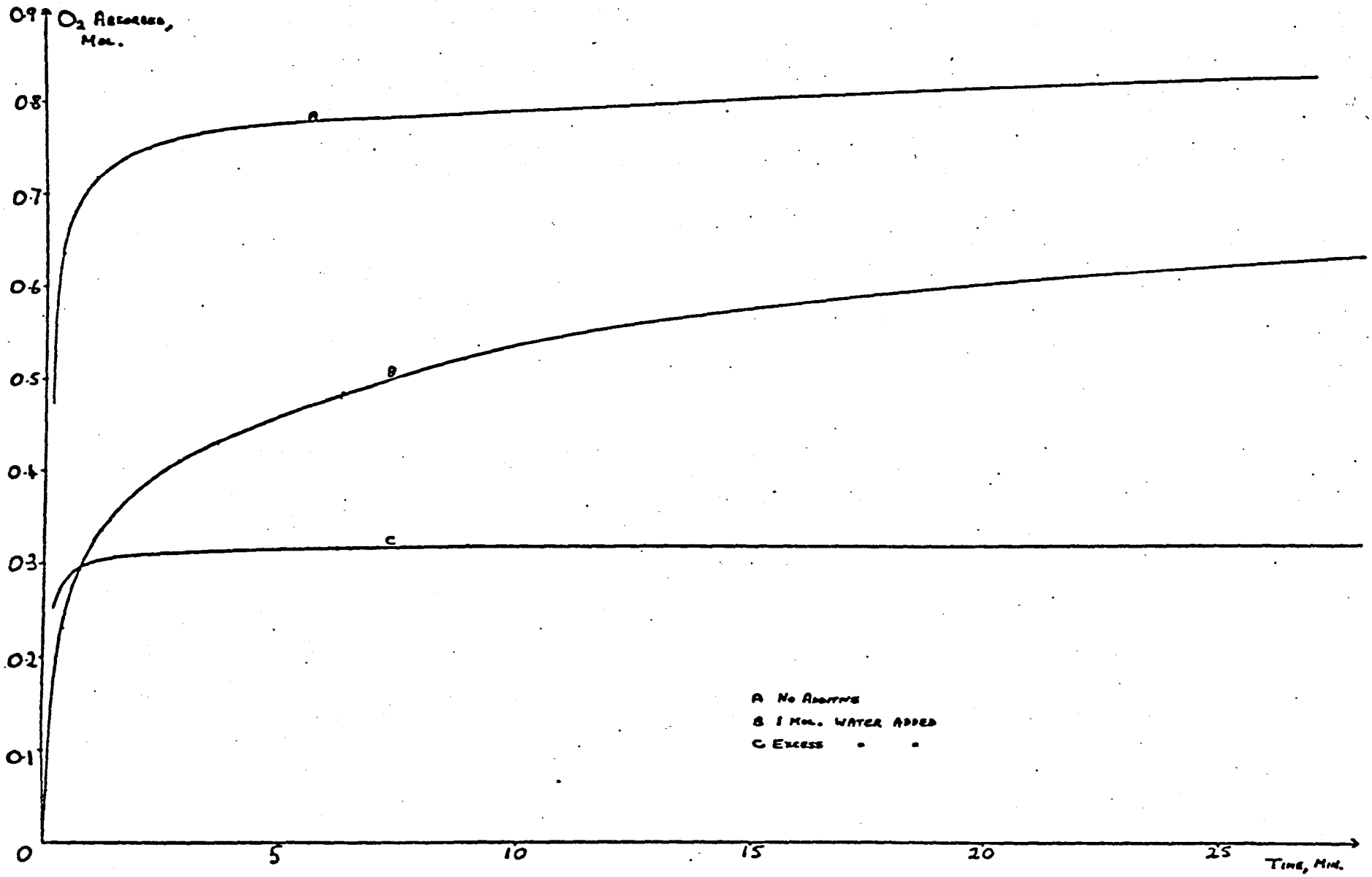




GRAPH 18. THE AUTOXIDATION OF GALVINOXYL IN BENZENE (13 MMOLAR) AT 25°C.

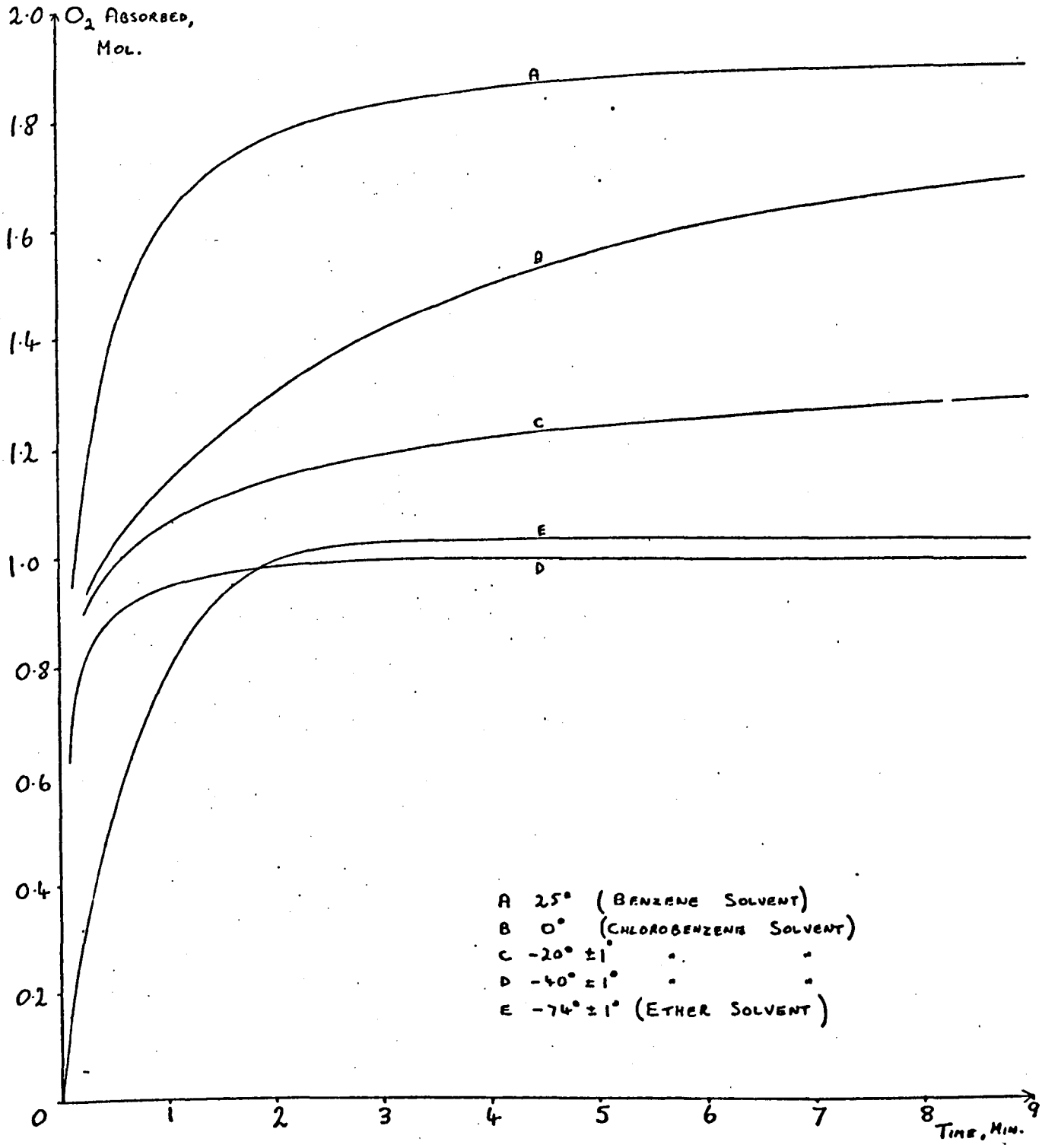
GRAPH 19. THE AUTOXIDATION OF Bu^s_3B IN BENZENE AT 25°C AT VARYING CONCENTRATIONS.



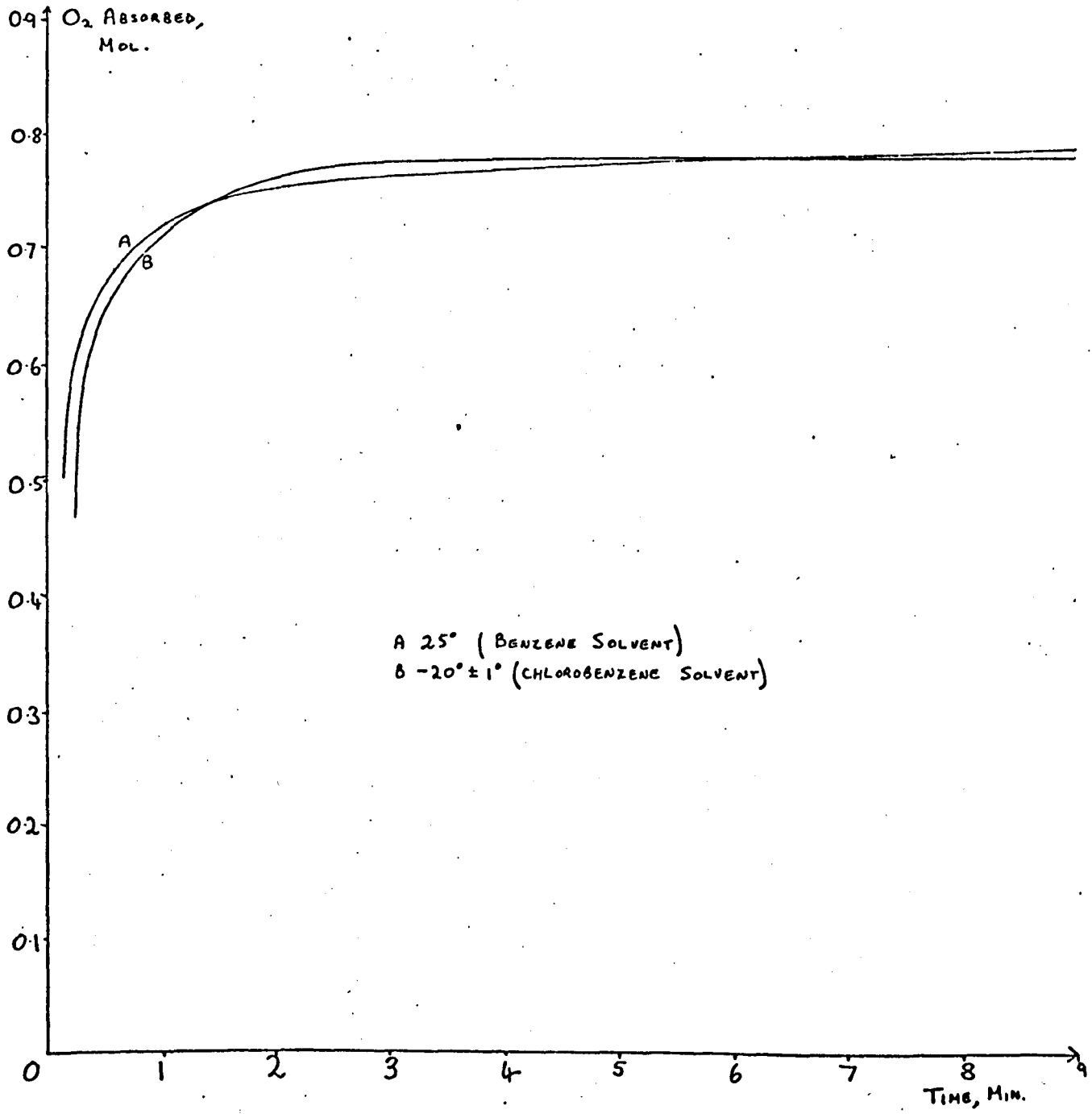


GRAPH 20. THE AUTOXIDATION OF 1-PHENYLETHYLBORONIC ACID IN BENZENE (8MMOLAR) IN THE PRESENCE OF WATER.

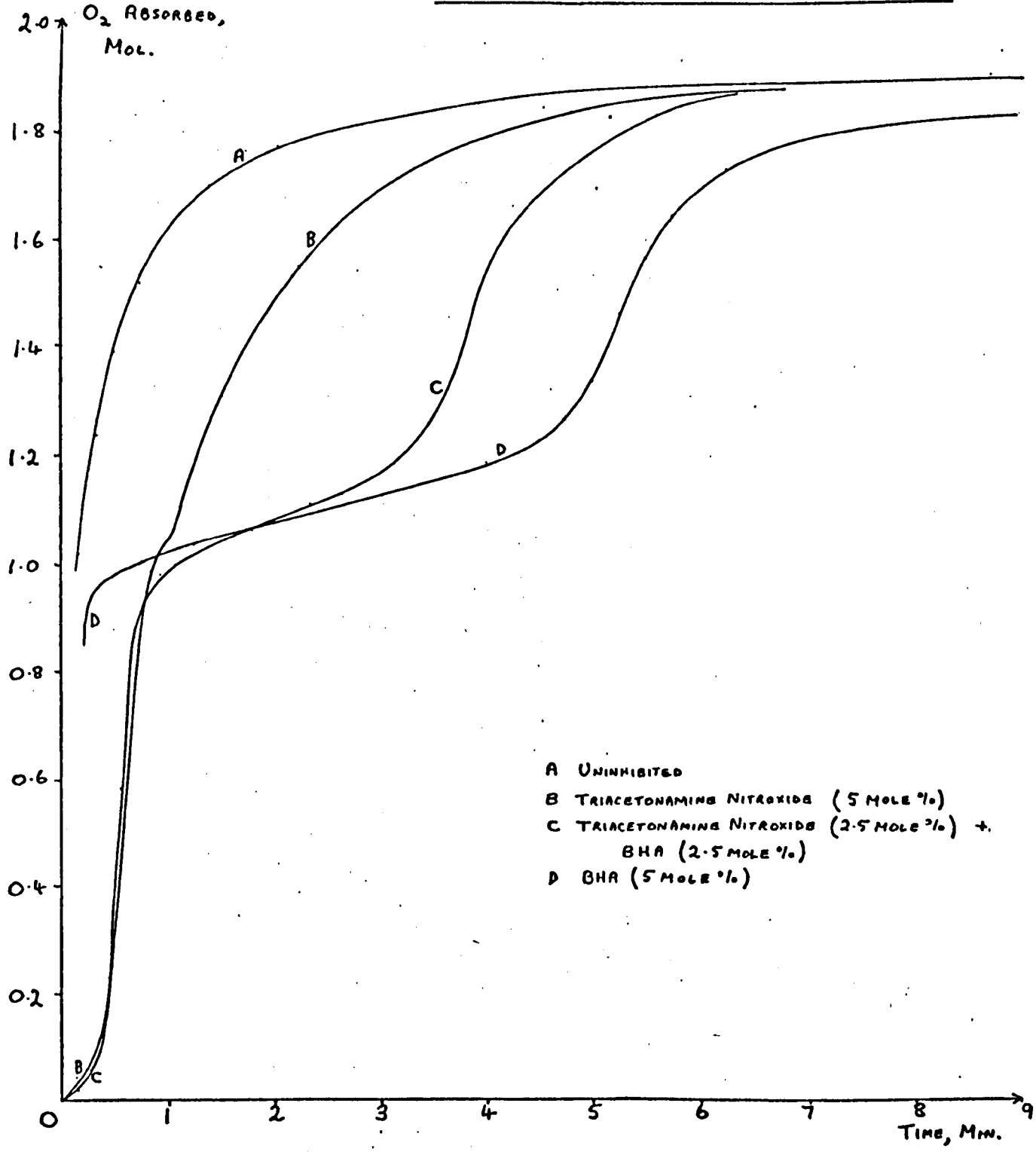
GRAPH 21. THE AUTOXIDATION OF Bu^s_3B (8 MMOLAR) AT VARYING TEMPERATURES.

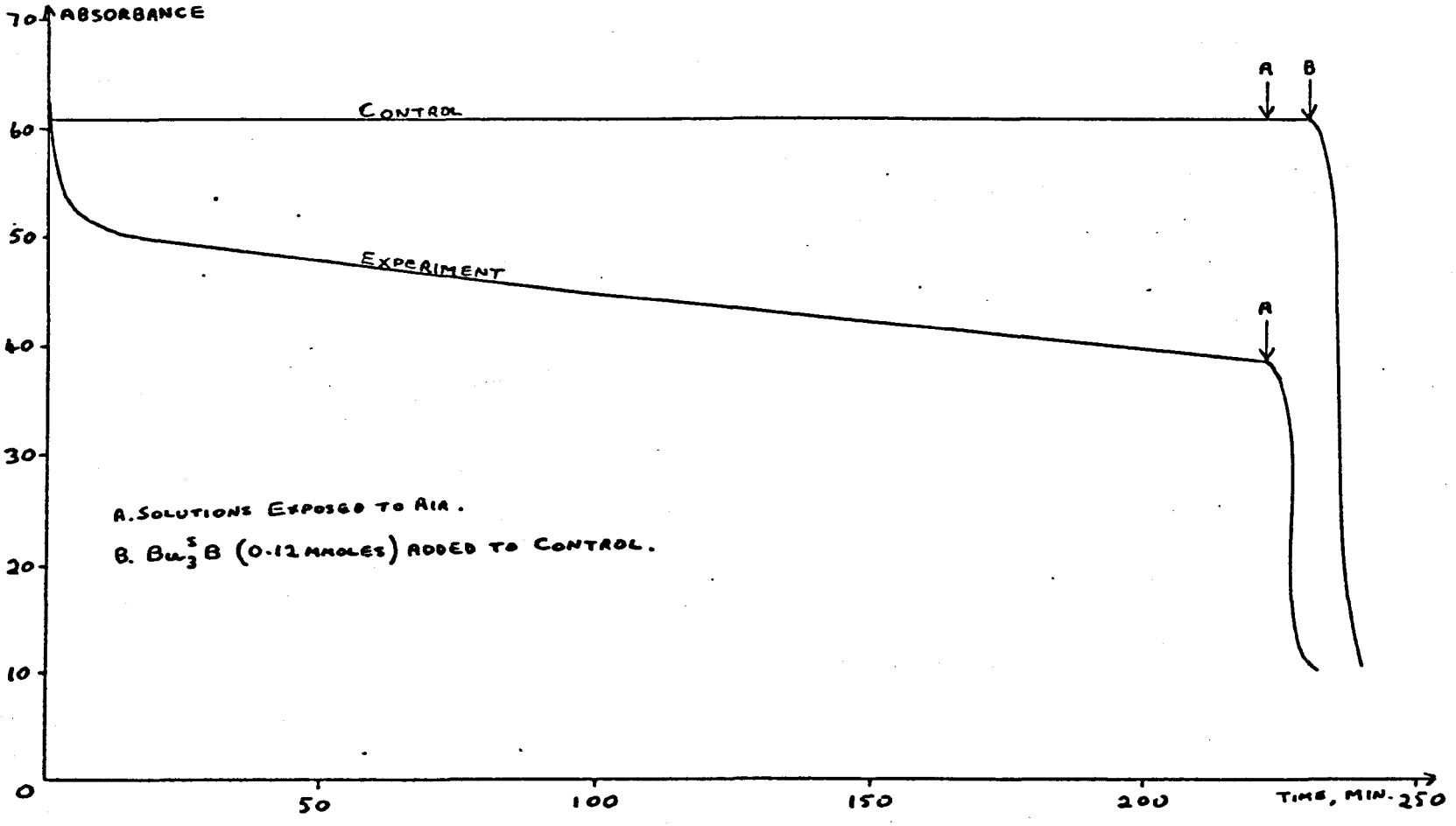


GRAPH 22. THE AUTOXIDATION OF 1-PHENYLETHYLBORONIC ACID (8MMOLAR) AT VARYING TEMPERATURE.

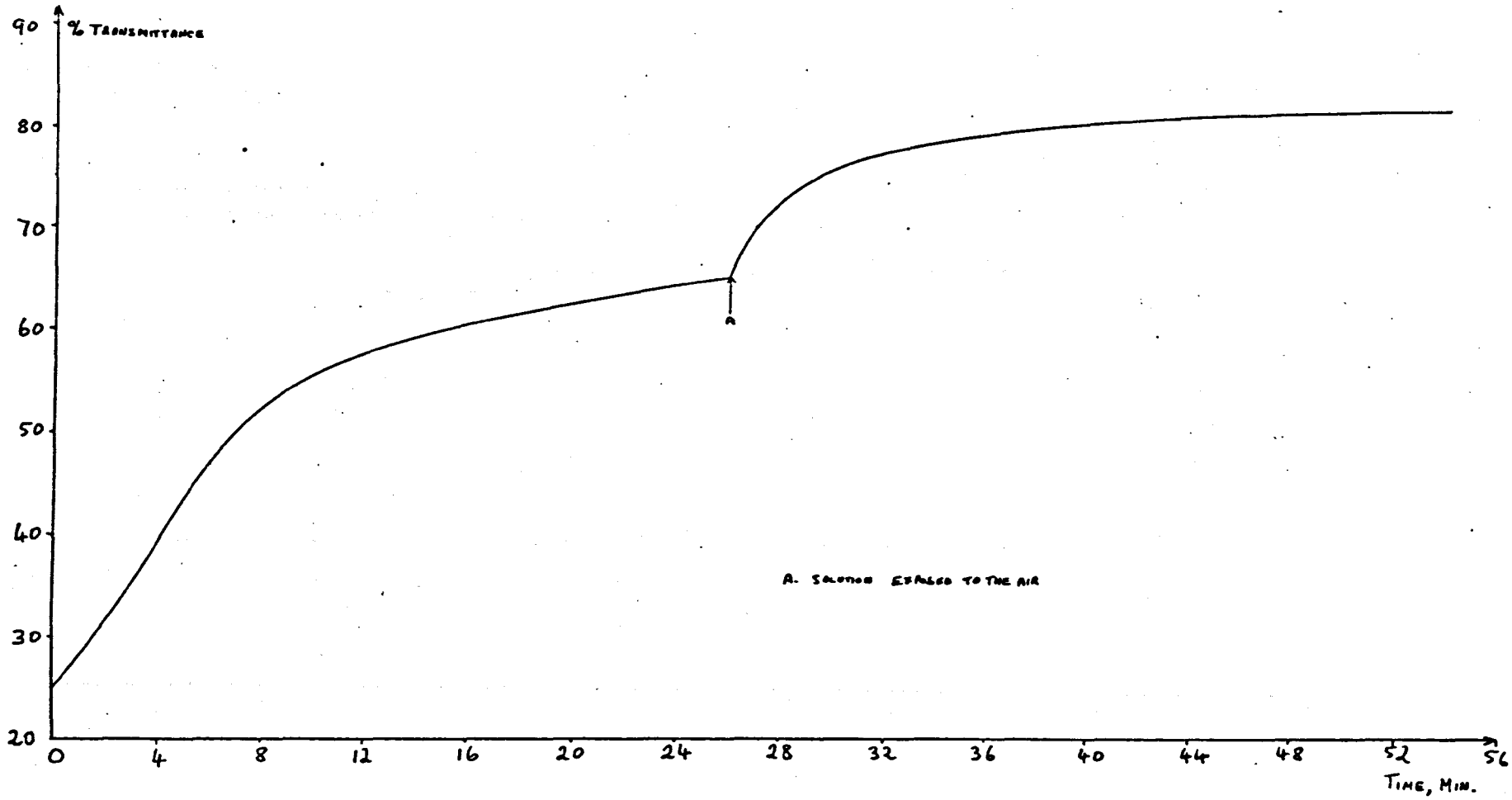


GRAPH 23. THE AUTOXIDATION OF Bu^s_3B IN BENZENE (6MMOLAR) AT 25°C IN THE PRESENCE OF A MIXED INHIBITOR.

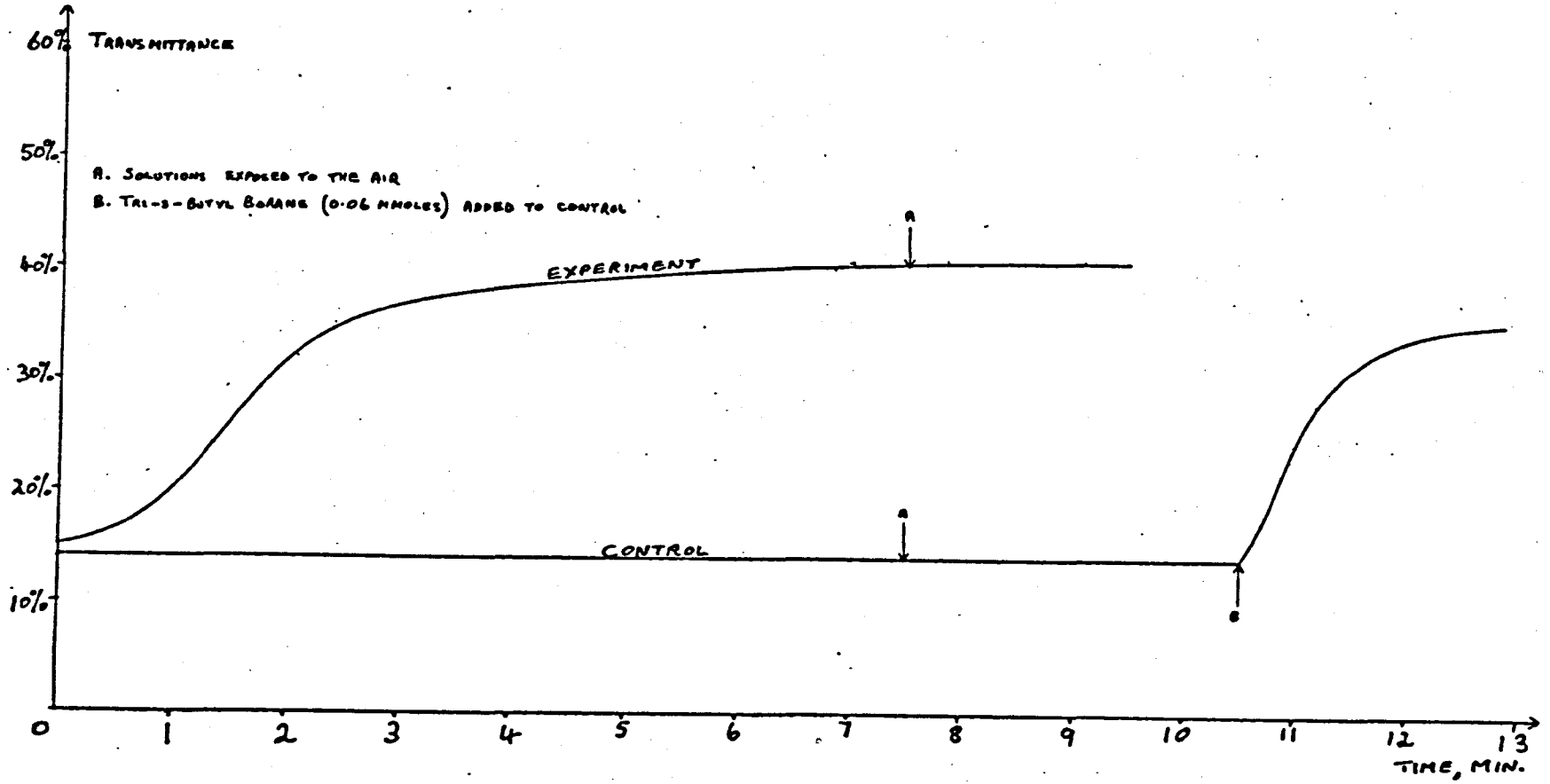




GRAPH 24. THE REACTION OF Bu_3^sB WITH GALVINOXYL (1 MOLE%) IN BENZENE IN THE ABSENCE OF OXYGEN.



GRAPH 25. THE REACTION OF Bu^n_3B WITH GALVINOXYL (10 MOLES%) IN BENZENE IN THE ABSENCE OF OXYGEN.



GRAPH 26. THE REACTION OF $Bu^{s}_{3}B$ WITH DPPH (2·6 MOLES %) IN BENZENE IN THE ABSENCE OF OXYGEN.

Chapter 7

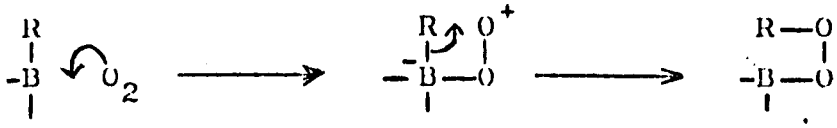
Autoxidation Mechanism and Studies on Diisopinocampheyl-s-butyl Borane

a) Theoretically Possible Mechanisms:-

The autoxidation of boron-carbon bonds can be represented by the stoichiometric equation (i). The mechanism of this

$$\text{>B-R} + O_2 \longrightarrow \text{>BOOR} \qquad \text{(i)}$$

reaction has been discussed by several authors, and until recently the scheme (ii) was widely accepted[1]. (Ch.3 p.67.)



(ii)

The evidence for this ionic mechanism was threefold. Firstly, quinol[2,3] and iodine[4] were not effective inhibitors for the reaction. This was taken to mean that, unlike hydrocarbon autoxidations, a free radical chain was not involved. Secondly, inter- or intramolecular donation of electrons into the vacant orbital of boron reduced the reactivity of the >B-R group. Finally, in the autoxidation of diisobutyl-t-butyl borane the t-butyl group was more reactive than the isobutyl group[5]. This last observation bears some resemblance to the results of Minato, Ware and Traylor[6] for the reaction between hydrogen peroxide and boronic acids (Equation iii). Both reactions fall into

the $S_E C$ classification of Abraham and Hill[7], and show the predicted order of migratory aptitude,

R=t-Bu > s-Bu > n-Bu

$$RB(OH)_2 + \bar{O}OH \rightleftharpoons HO-\overset{\overset{R}{|}}{\underset{\underset{OH}{|}}{B}}-O-OH \xrightarrow{slow} (HO)_2BOR + OH^- \qquad (iii)$$

However, Grummitt's[3] inhibition of the autoxidation of trialkylboroxines and the polymerisation studies of Zutty and Welch[8] had the characteristics of a free radical process (Ch.3 p.70). This work was therefore undertaken to resolve the conflicting evidence for a heterolytic or homolytic process.

b) Stereochemical Considerations:-

The mechanism (ii) involves rearrangement of an alkyl group to an electron deficient centre. Similar processes, such as the Beckmann and Baeyer-Villiger reactions, occur with retention of configuration of the migrating group[9], and by analogy reaction (ii) might be expected to do likewise. However, in a free radical chain reaction with propagating steps such as (iv) and (v), the alkyl radical would be expected to become planar, with subsequent racemisation of configuration.

$$R^{\cdot} + O_2 \longrightarrow RO_2^{\cdot} \qquad (iv)$$

$$RO_2^{\cdot} + BR_3 \longrightarrow RO_2BR_2 + R^{\cdot} \qquad (v)$$

A study of the steric course of this reaction would thus

provide a valuable, though not conclusive, piece of evidence in favour of one mechanism or the other.

The preparation by H.C. Brown[10] of an optically active borane provided a suitable compound for this investigation. Diisopinocampheyl-s-butyl borane, $(IPC)_2BBu^s$, was oxidised by alkaline hydrogen peroxide to s-butanol which was 86% optically pure, (route vi), and it has been argued that this step involves retention of configuration (Ch.1 p.5.). Therefore $(IPC)_2BBu^s$ was autoxidised and hydrogenated, and the activity of s-butanol produced by routes (vi) and (vii) compared.

$$(IPC)_2BBu^s \xrightarrow{H_2O_2/NaOH} (IPC)OH + B(OH)_3 + Bu^sOH \qquad (vi)$$

$$(IPC)_2BBu^s \xrightarrow{1)\ O_2;\ 2)\ H_2/Pt} (IPC)OH + Bu^sOH \text{ and other products} \qquad (vii)$$

Since the hydrogenation of the peroxide does not affect the asymmetric centre, this comparison must throw light on the stereochemical changes involved in the autoxidation step.

However the situation is complicated by the possibility of peroxide rearrangement (Equation viii), which may occur with retention of configuration (see p.197).

$$(IPC)_2BOOBu^s + (IPC)_2B\overset{*}{B}u^s \longrightarrow (IPC)_2BOBu^s + (IPC)_2BO\overset{*}{B}u^s \qquad (viii)$$

This reaction, which is relatively slow in the absence of an added nucleophile[1], would become more important in cases where the autoxidation was carried out slowly.

The results of three experiments are shown below and in Table 6.

Expt. No.	Initial α-pinene	Time O_2 passed, hr.	s-Butanol via route vi	s-Butanol via route vii
1	+ 47.7°	6	- 7.7°	- 2.8°
2	- 42.4°	1.5	+ 10.3°	- 0.44°
3	- 42.4°	2.5	+ 11.4°	+ 1.54°

The variation in optical purity of the s-butanol produced by route (vi) is probably a reflection of an improvement in experimental technique. More important is the ratio of activities of the alcohols produced by the two routes. In Expt.1, where stirring was inefficient and the surface area of liquid in contact with gas was low, there was 36% retenton. However, this value fell to 12% in Expt.3, for which the high speed stirring apparatus of Fig.6 was used. This gave a very large surface area of contact between gas and liquid phases, with a much shorter autoxidation time. The results indicate that racemisation has occurred to a significant extent during the autoxidation stage, and the residual activity may be due to the oxidation of some s-butyl groups by peroxide (Equation viii).

The small negative value obtained via route (vii) in

Expt.2 is puzzling. It is not due to the presence of ∝-pinene, since the s-butanol was shown by GLC to contain much less than the 1% required to give this rotation. The s-butanol used to dilute the sample was checked and found to be racemic. However, the value of -0.44^{o} was calculated from an observed rotation of -0.0007^{o}, which is near the limit of accuracy of the polarimeter[11] ($\pm 0.0002^{o}$), and this result must be viewed with caution.

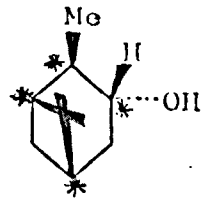
If a free radical mechanism were operating, products such as methyl ethyl ketone, 3,4 dimethyl hexane etc. would be expected due to side reactions of the s-butyl and s-butylperoxy radicals (Ch.2). Figs. 4 and 5 are typical GLC analyses of the ether/pentane extracts of s-butanol formed by routes (vi) and (vii) respectively. It can be seen that there are many additional impurities in the latter, and it is possible that some of these arise in this way. However, excess ∝-pinene was present during the autoxidation, and this has been shown to react rapidly with oxygen at 100^{o} to give a variety of products[12]. These are mainly terpenes, but include compounds of lower molecular weight which could equally well account for the extra peaks of Fig. 5.

The optical activity of isopinocampheol (I) produced by routes (vi) and (vii) could also in theory have been used to

help decide between the two mechanisms discussed above, but this was complicated in two ways.

In the first case, the isopinocampheol would be difficult to isolate. $(IPC)_2BBu^s$ only takes up about 1.5 mol. oxygen (Table 9) and since the s-butyl group appears to be preferentially autoxidised, the yield of isopinocampheol by route (vii) must only be about 25%. It would also be contaminated, not only by boronic and borinic acids, but by other by-products of autoxidation. An attempt was made to isolate the alcohol by prolonged steam distillation of the autoxidation/hydrogenation residue, but without success.

Secondly, it would not be easy to interpret the optical activity of the alcohol, even when it had been isolated. Isopinocampheol has four asymmetric carbon atoms (shown asterisked in I) only one of which could suffer racemisation on autoxidation.



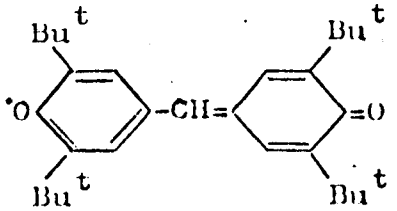
(I)

The results obtained show that a considerable degree of racemisation accompanies the autoxidation step. This alone

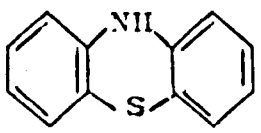
is not proof of the involvement of free radicals, but merely indicates that a symmetrical intermediate must be present. The rearrangement mechanism (ii) is therefore unlikely, but further evidence is required before a definite alternative can be postulated.

c) Effect of Inhibitors:-

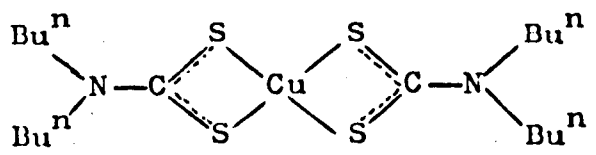
Compared with simple trialkyl boranes, $(IPC)_2BBu^s$ autoxidises relatively slowly (Graph 1), and it was thought that although earlier workers had been unsuccessful in their attempts to inhibit borane autoxidations, a sufficiently powerful antioxidant might prove effective in this case. In fact, the presence of small quantities of certain compounds was found to have a profound influence on the autoxidation of $(IPC)_2BBu^s$, Thus galvinoxyl (II) in a concentration of 0.1 mole % of the borane, completely suppressed the reaction for 17 minutes (5 half lives). Phenothiazine (III), used extensively as an industrial antioxidant[13], and copper dibutyl dithiocarbamate[14] (IV) a good stabiliser for natural rubber at high temperature, were also very effective (Fig.10/ Graph 2).



(II)



(III)



(IV)

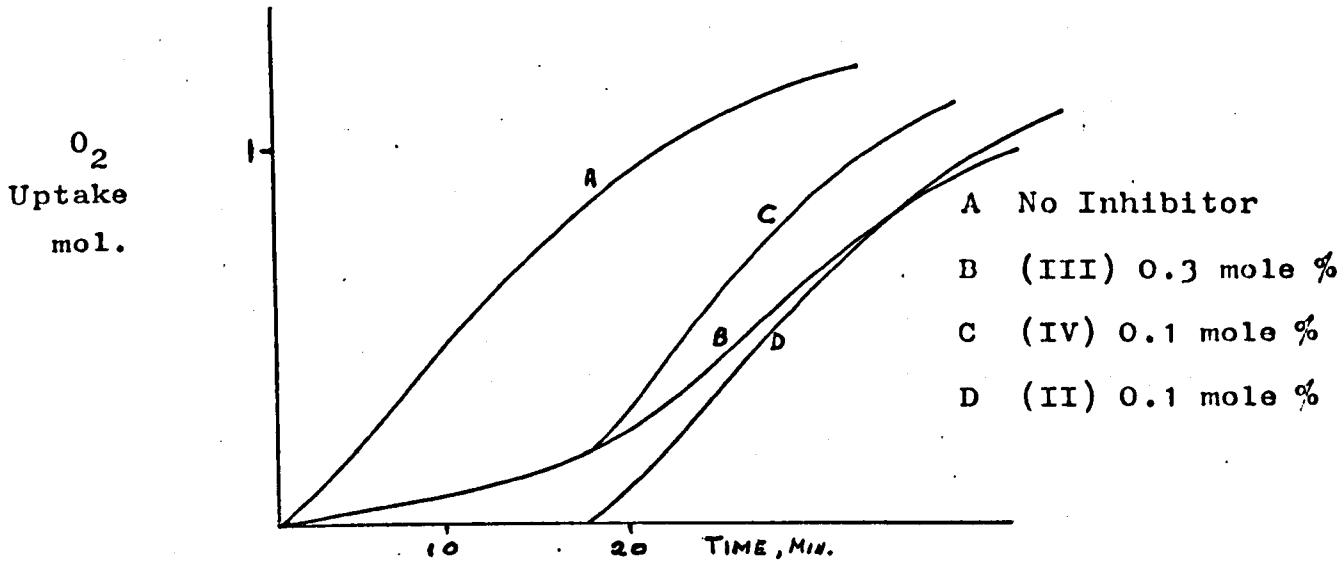
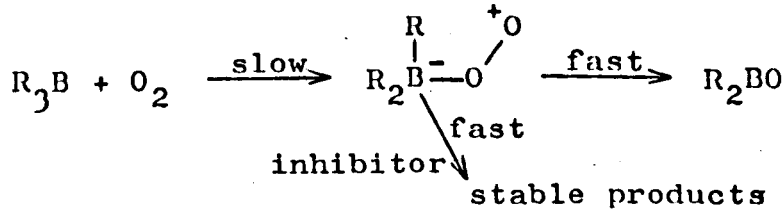


Fig.10 - Autoxidation of $(IPC)_2BBu^s$ in Benzene(10 mmolar)at 25°C.

The autoxidation of many other organoboron compounds could also be inhibited, and these results are discussed fully in Chapter 8.

The explanation for the effectiveness of small quantities of an additive in stopping a reaction is that it must be removing in some way a key intermediate present in very low concentration. This still does not exclude the ionic mechanism (ii) since it could be argued that the formation of the intermediate $R_3B^-O_2^+$ is a slow step, and it is this which reacts with the inhibitor.



$$R_3B + O_2 \xrightarrow{\text{slow}} R_2B{-}O{-}\overset{R}{\underset{|}{C}}{-}O^{+} \xrightarrow{\text{fast}} R_2BOOR \qquad \text{(ix)}$$

$$R_2B{-}O{-}\overset{R}{\underset{|}{C}}{-}O^{+} \xrightarrow{\text{fast}} \text{inhibitor} \rightarrow \text{stable products}$$

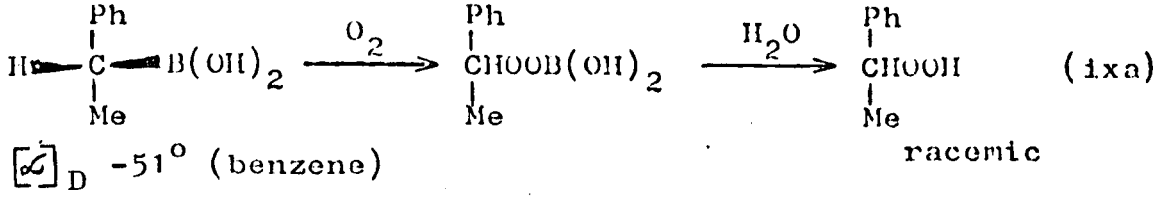
However, since the autoxidation is completely inhibited, the rate of formation of this intermediate cannot be greater than its rate of removal by galvinoxyl (0.001 mol. in 17 min., assuming a 1:1 reaction between intermediate and galvinoxyl). The overall rate of oxidation is much greater than that of this individual step. ($RB< \longrightarrow RO_2B<$; $t_{\frac{1}{2}}$ = 4 min. at 25°C.) This can only be explained in terms of a chain reaction, with slow initiation and rapid propagating steps. Therefore, the mechanism (ii) must be ruled out, together with any other theoretically possible mechanisms not involving chains.

The inhibitors used in this study are known to stop free radical chains (Ch.2) and it would be difficult to explain their effectiveness if the chain-carrying species were ionic in nature. This is particularly true of galvinoxyl, which is itself a free radical, and is the most powerful inhibitor so far discovered for organoborane autoxidation. Further proof of the free radical nature of the intermediates comes from the isolation of an adduct between an inhibitor and an alkyl radical (see p184).

Thus it is concluded that the autoxidation of organoboranes must proceed by a free radical chain mechanism.

d) Evidence from other Authors:-

While this work was in progress a paper was published[15] suggesting that the autoxidation of organoboron compounds could involve a free radical chain, and a subsequent paper by Davies and Roberts[16] established that this was so. Optically active 1-phenylethylboronic acid was found to give racemic 1-phenylethyl hydroperoxide on autoxidation and hydrolysis:



(ixa)

A slight residual activity was found in the 1-phenylethyl hydroperoxide, which was attributed to peroxide rearrangement. This effect was much less than was observed for the autoxidation of $(IPC)_2BBu^s$ (Table 6), and this is possibly due to the dilute solutions and consequently short reaction times ($t_{\frac{1}{2}}$ ca. 6 min.) used in the above work[16].

The autoxidation of 1-phenylethylboronic acid was found to be inhibited by galvinoxyl, as was that of triisobutyl borane, di-s-butylborinic anhydride and s-butylboronic

anhydride. The authors postulated propagating steps such as (xi) and (xii) below.

e) Autoxidation Mechanism:-

The evidence for a free radical chain mechanism can be summarised as follows:

(i) The autoxidation proceeds with racemisation of configuration of the alkyl group concerned.

(ii) Small quantities of known free radical trappers inhibit the reaction for many half lives

(iii) An adduct between an alkyl radical and an inhibitor has been isolated from the reaction mixture.

It is concluded therefore, that a free radical chain mechanism governs the autoxidation of boron-carbon bonds. This is probably similar to the autoxidation of carbon compounds reviewed in Chapter 2, and can be represented by an analogous set of equations:

Initiation $R_3B \longrightarrow R^{\bullet} +$ (x)

Propagation
$$R^{\bullet} + O_2 \longrightarrow RO_2^{\bullet} \quad \text{(xi)}$$
$$RO_2^{\bullet} + R_3B \longrightarrow \left[RO_2\dot{B}R_3\right] \longrightarrow RO_2BR_2 + R^{\bullet} \quad \text{(xii)}$$

Termination $R^{\bullet}, RO_2^{\bullet} \longrightarrow$ non-radical products (xiii)

The individual steps are discussed in greater detail in Chapters 8 and 9.

The evidence previously put forward in favour of the ionic mechanism (ii) is also compatible with a free radical chain reaction. The ineffectiveness of quinol and iodine previously reported is not surprising, since they are not particularly efficient inhibitors, and might not be expected to stop such a facile reaction. Secondly, donation of electrons into the boron atom, giving partial double bond character, will reduce the tendency of the alkylperoxy radical to attack it. This will however, only be important if reaction (xii) is rate determining, that is, if reaction (xi) is very much faster, and if chains are long so that initiation can be ignored. The great effect of small amounts of galvinoxyl, and the almost quantitative yield of peroxide, show that the chains are indeed very long, and the virtual independence of reaction rate on oxygen pressure (Graphs 13 and 14) justifies the first assumption (see p.205).

The third point advanced in favour of the ionic mechanism, namely the relative migratory aptitude of the t-butyl and isobutyl groups, is the same as would be predicted for reaction (xii). The ejection of the more stable t-butyl radical from diisobutyl-t-butyl borane would lower the overall activation energy of the reaction, and hence this group would be expected to autoxidise more readily. Hansen and Hamann[4] autoxidised triethyl borane in a mixture of

benzene and methyl methacrylate, and failed to observe polymerisation. They therefore concluded that the reaction did not involve free radicals. However, in the same paper they claim that almost all peroxide was formed within 15 seconds. It follows that very few chains would be initiated after this time, and so polymerisation need not occur to an observable extent.

Other authors have used boranes in the presence of oxygen to initiate polymerisation (see Ch.3 p. 70), and in a subsequent paper Hansen[17] did get polymerisation of methyl methacrylate in the presence of triethyl borane and oxygen. The concentrations he used were similar in both papers, and the significant difference was the rate of stirring. In the first[4], this was in excess of 1,000 r.p.m., which accounts for the very rapid reaction of the borane. In the second, the rate was only 350 r.p.m., and the autoxidation was reported to take 3 - 10 minutes[17]. The importance of efficient stirring is discussed later (Ch.9).

f) Isolation of Diisopinocampheyl-s-butyl Borane:-

The initial work of Brown and Zweifel on $(IPC)_2BBu^s$ was carried out in solution, and the pure borane was not isolated[18]. The possibility that this compound would be a suitable substrate for a detailed kinetic study, with a view to

substantiating further the free radical chain mechanism, led to unsuccessful attempts to isolate it. Removal of the last traces of diglyme by prolonged pumping at room temperature caused decomposition of the borane., and at the time it was believed that the complexed solvent was essential to its stability.

It was only at a much later stage in this research that the borane was isolated as a crystalline solid by suitably modifying the published method of preparation[19]. Experiments suggested that the yield of diborane from sodium borohydride and boron trifluoride etherate was less than would be expected from Equation (xiv). Thus by using a large excess

$$3NaBH_4 + 4BF_3{\cdot}OEt_2 \longrightarrow 3NaBF_4 + 2B_2H_6 + 4Et_2O \quad (xiv)$$

of these reagents, $(IPC)_2BBu^s$ could be prepared satisfactorily from diborane generated externally.

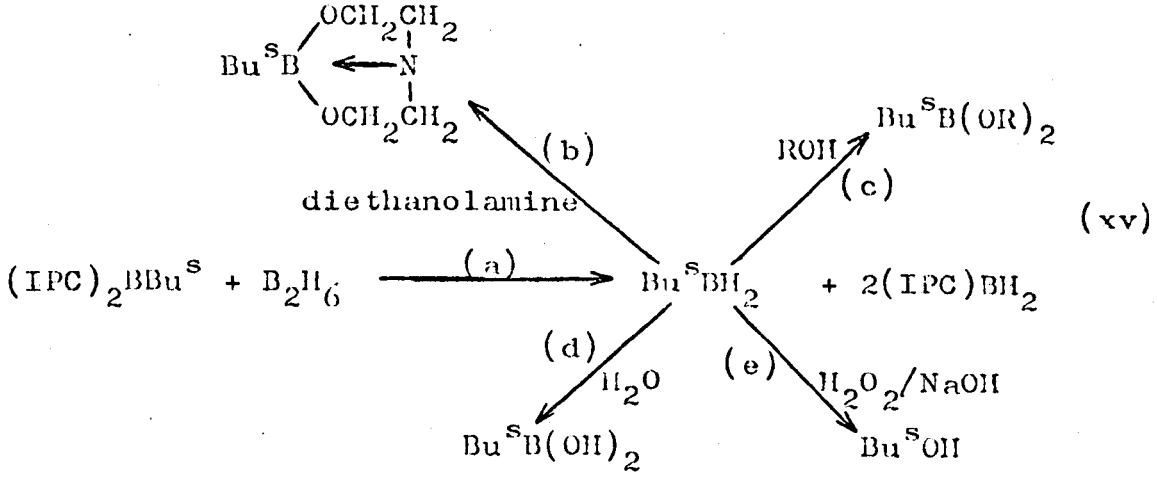
Cooling of the diglyme solution to -78^o gave a white gum which could not be filtered. On warming, this partially redissolved, and the remainder crystallised. These crystals acted as seeds, and a second cooling to -78^o caused a heavy crystalline precipitate to come down.

g) Attempted Preparation of Optically Active s-Butylboronic Acid:-

In the light of the work of Davies and Roberts[16], on

optically active 1-phenylethylboronic acid, it was considered desirable to try and convert $(IPC)_2BBu^s$ to active s-butylboronic acid. This simple boronic acid and its derivatives could be used to study any reaction in which the boron-carbon bond is broken, to establish the stereochemistry and provide information about the mechanism. Comparison with the behaviour of active 1-phenylethylboronic acid could produce useful information about the relative effects of β-phenyl and β-ethyl groups on the reactivity of the boron atom.

Attempts were therefore made to bring about the disproportionation of $(IPC)_2BBu^s$ with excess diborane and then to isolate the boronic acid or an ester according to the scheme below (results summarised in Table 7):



(xv)

Reaction (xv) (a) might be expected to proceed with retention of configuration if it involved the four centre

transition state postulated by Brown[19]. In fact it was eventually found that the borane could not be disproportionated according to route (xv) (a), even under the forcing conditions used (p.108). This is in marked contrast to the discovery by Brown et al.[20] that simple trialkyl boranes readily undergo redistribution to mono- and dialkyl boranes in the presence of excess diborane in THF at 50-55°.

The inertness of $(IPC)_2BBu^s$ under these conditions is probably due to shielding of the boron atom by the bulky groups surrounding it. If this is the correct explanation, the n-butyl isomer would be expected to be more reactive. Although no attempt was made to react $(IPC)_2BBu^n$ with diborane, the reactivity of the two boranes towards oxygen is in the same order n-Bu > s-Bu (Graph 1). This is the reverse of the order for simple trialkyl boranes (Table 5) and can be attributed to steric factors.

An attempt was made to correlate the optical activity of s-butanol produced by alkaline hydrogen peroxide oxidation of $(IPC)_2BBu^s$ before and after treatment with excess diborane (Table 7, and route (xv) (e)). The high value for the specific rotation in the latter case (+ 8.1° compared with + 8.8° in the former) was originally taken to mean that disproportionation had occurred with retention of configuration.

Although this conclusion was invalid, the experiment indicates that $(IPC)_2BBu^s$ is optically stable in contact with diborane at 20^oC.

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Chapter 8

The Autoxidation of Organoboron Compounds in The Presence of Inhibitors

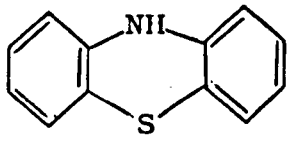
One of the main criteria for the existence of a free radical chain mechanism is the effectiveness of small quantities of added agents in stopping the reaction, either by direct interaction with the radicals or by prevention of their formation in some way (Ch.2). The apparent similarities between trialkyl borane and hydrocarbon autoxidations suggested a systematic investigation of the effect of typical hydrocarbon antioxidants on the former. Tri-s-butyl borane was chosen for detailed study, and a comparison was made with certain other organoboron compounds, *viz*. $(IPC)_2BBu^s$, $(IPC)_2BBu^n$, s-butylboronic anhydride, triisobutyl and tri-n-butyl boranes. The range of reactivities thus embraced was extremely wide, and varied from an absorbtion of the first molecule of oxygen for tributyl boranes which was too fast to measure ($t_{\frac{1}{2}}$ *ca*. 2 sec.) to the slow second mol. uptake for triisobutyl borane ($t_{\frac{1}{2}}$ 31 min.) (Graph 1).

Since initiation processes in organoboron systems are likely to be significantly different to those in hydrocarbon autoxidations, this work was confined to a study of

antioxidants known to be propagation suppressors. In general these compounds had far less effect than in hydrocarbon autoxidations, and only with galvinoxyl was true inhibition observed (Graphs 2-11).

a) Phenols and Amines:-

It can be seen from Graphs 7 and 8 that the autoxidation of tri-s-butyl borane was only slightly affected by the presence of phenols and amines commonly used as antioxidants. The exception was phenothiazine (I) which is a special case and is discussed on p.177.



(I)

i) Phenols. In hydrocarbon autoxidation, three factors are likely to influence phenol activity. These are (a) the ease of hydrogen abstraction:

$$AH + RO_2^{\bullet} \longrightarrow A^{\bullet} + ROOH \qquad (i)$$

(b) the readiness of the phenoxy radical $A^{\bullet}$ to undergo chain transfer:

$$A^{\bullet} + RH \longrightarrow AH + R^{\bullet} \qquad (ii)$$

and (c) the number of peroxy radicals that subsequently react with the phenoxy radical. K.U. Ingold and co-workers[1-4] have demonstrated that it is the first which is most important.

Since Hammett correlations involve σ^+ constants (e.g. with styrylperoxy radicals) it seems that charge separation in the transition state (Structure II) is more important than steric factors. However, the magnitude of electronic effects

$$ROO^{\cdot}\ H{:}A \longrightarrow \left[ROO^{-}{:}\ \overset{+}{H}{\cdot}A\right]_{II} \longrightarrow ROO{:}H\ {}^{\cdot}A \qquad \text{(iii)}$$

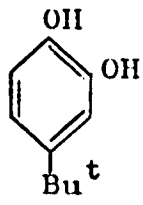
is not great, since only small negative ρ values are involved. Therefore the most effective phenols in hydrocarbon autoxidation are those with low steric hindrance of the OH group coupled with electron donating groups in the 2 and 4 positions, provided that the oxidation potential is not so low that reaction (ii) or air oxidation of the phenol occurs.

It is reasonable in the absence of evidence to the contrary to suggest that these factors still apply in organoboron autoxidations. The competition between phenol and borane for peroxy radicals (Equations i and iv) is much keener, resulting in retardation rather than inhibition,

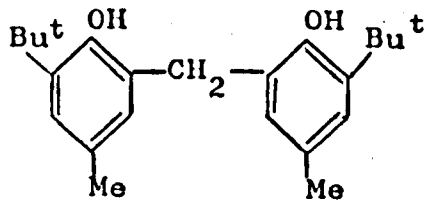
$$RO_2^{\cdot} + BR_3 \longrightarrow RO_2BR_2 + R^{\cdot} \qquad \text{(iv)}$$

and the differences in phenol efficiency became less easily measured.

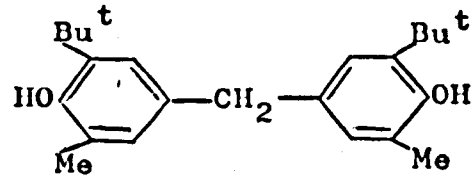
Only three phenols showed any detectable retardation of the first mol. oxygen uptake. These were p-t-butyl catechol (III), Cyanamid 2246 (IV) and Ethyl 720 (V) (Graph 8, curves E,F,G). The significance of this result is



(III)



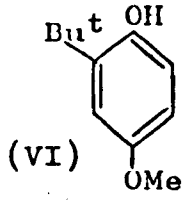
(IV)



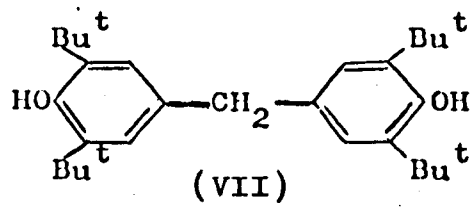
(V)

reduced by the fact that all three are dihydric phenols. Approximately the same molar concentration of each inhibitor was used, which gives dihydric phenols double the concentration of OH groups.

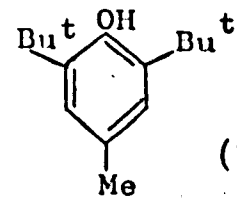
The rate of autoxidation of the second stage, being a much slower process due to reduced acceptor power of the boron atom, was retarded by all the phenols used. Two different types of behaviour could be distinguished (Graph 8). Cyanamid 2246 (IV) and butylated hydroxy anisole (BHA) (VI) gave relatively strong retardation followed by an increase in rate almost to that of the uninhibited reaction. On the other hand, Ethyl 702 (VII) and butylated hydroxy toluene (BHT) (VIII) gave smooth curves resulting from slight retardation over the whole range. These results can be



(VI)



(VII)

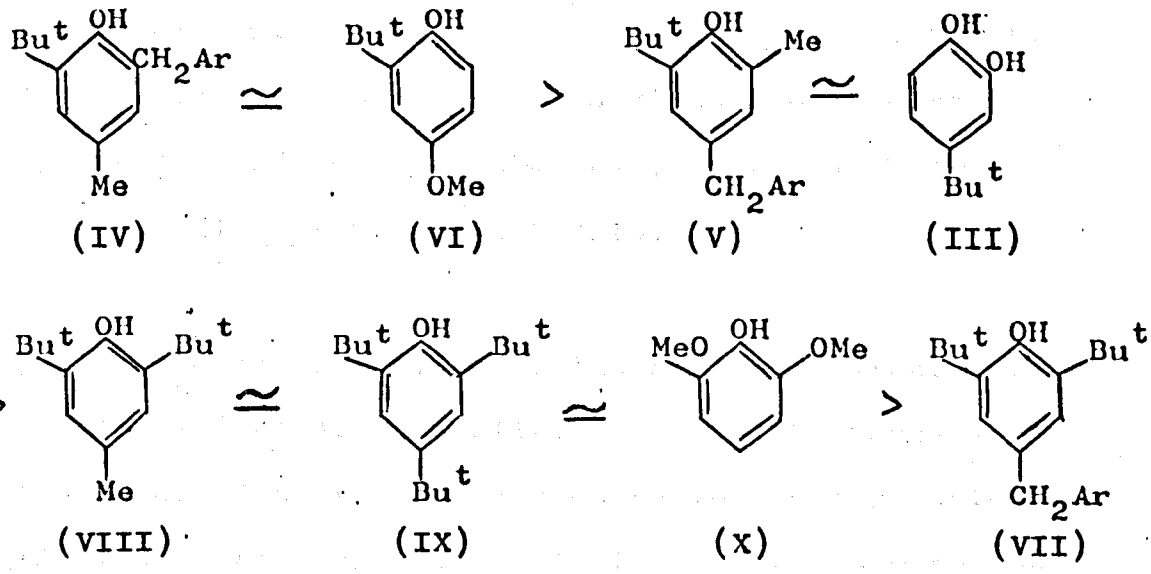


(VIII)

explained by Howard and Ingold's[1] observation that phenols with a small substituent in one ortho position such as

(IV) and (VI) are more efficient at peroxy radical trapping, and hence more rapidly used up. Weaker inhibitors, which do not participate to such a great extent in the reaction, are still not completely utilised by the end of the autoxidation.

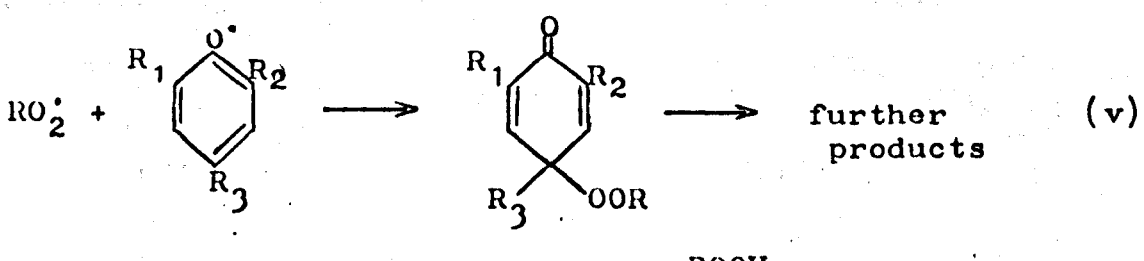
The order of effectiveness of substituted phenols in retarding the autoxidation of tri-s-butyl borane during the uptake of the second molecule of oxygen is shown below. 2,4,6 Tri-t-butyl phenol (IX) and 2,6 dimethoxy phenol (X), which gave almost identical curves to BHT at the same concentration, are not shown on Graph 8.



Since these phenols only retard, it is not possible to give a numerical comparison. Even so it is quite clear that the presence of two bulky ortho substituents, giving steric

protection to the OH group, adversely affects the inhibiting power of the phenol. Thus the dominating factor is ease of abstraction of the hydrogen atom, in agreement with Ingold's work on hydrocarbon systems[1-4], and this further substantiates the proposed free radical mechanism of autoxidation. Bearing in mind the extreme facility of the propagation step (iv) compared to that in hydrocarbons, it is not surprising that the prime requisite of a good phenolic inhibitor is that it should react as rapidly as possible with peroxy radicals. Certain phenols with low oxidation potentials are not effective in hydrocarbon systems, because the hydrogen atom is so easily removed that a slow chain-initiating reaction occurs with oxygen. Such compounds might be expected to retard the autoxidation of tri-s-butyl borane even more efficiently than BHA or Cyanamid 2246, provided they do not have two bulky ortho substituents.

The subsequent reactions of the aryloxy radical have not been investigated. Usually, however, the benzene solution acquired a strong green or yellow colour during the autoxidation of the borane. This is compatible with the formation of quinones (Ch.2 p.45).



(v)

It was also noted that the quantity of peroxide formed was often significantly less than the oxygen uptake (Tables 9 and 10), and it is possible that induced decomposition has occurred. There are reports that some phenols, particularly those which give rise to stable phenoxy radicals, can cause the decomposition of hydroperoxide. Thus 2,4,6 tri-t-butyl phenoxy reacts rapidly with t-butyl hydroperoxide[5], and in this work it gave the lowest peroxide titre of all the phenols.

ii) Amines. Here again the established mode of action of secondary aromatic amines can be interpreted into the organo-borane system. The mechanism resembles phenolic inhibition in that there is a rate determining removal of hydrogen (shown by a deuterium effect[6]). The amino radical then reacts with another peroxy radical to form the nitroxide in fairly high concentration[7] (Ch.2, pp.50 and 51).

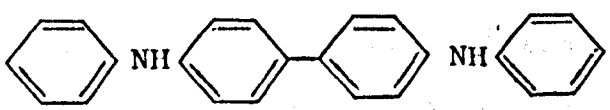
$$Ar_2NH + RO_2^{\bullet} \longrightarrow Ar_2N^{\bullet} + ROOH \qquad \text{(vi)}$$

$$Ar_2N^{\bullet} + RO_2^{\bullet} \longrightarrow Ar_2NO^{\bullet} + RO^{\bullet} \qquad \text{(vii)}$$

$$Ar_2NO^{\bullet} + \begin{cases} R^{\bullet} \\ RO_2^{\bullet} \end{cases} \longrightarrow \text{stable products} \qquad \text{(viii)}$$

Three amine antioxidants were studied: diphenylamine (XI), diphenylbenzidine (XII) and phenothiazine (I).

(XI)



(XII)

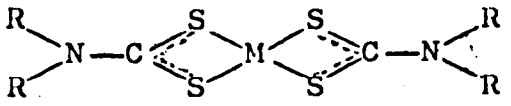
The similarity between (X) and (XI) is reflected in their retardation of the second mol. oxygen uptake of tri-s-butyl borane.

A strong orange-pink colour appeared during the reaction for all three amines. (In the case of phenthiazine this occurred during the retardation following the first mol. oxygen uptake.) This could be attributed to the formation of the nitroxides observed by Thomas[7].

In conclusion therefore, the low efficiency of phenol and amine antioxidants in organoborane autoxidations is thought to be due to unsuccessful competition with the substrate for alkylperoxy radicals. The other observations can be interpreted satisfactorily in terms of the same mechanisms which govern their action in inhibiting hydrocarbon autoxidations, and are therefore quite compatible with a free radical chain mechanism for organoborane autoxidations.

b) Sulphur and Phosphorus Containing Antioxidants:-

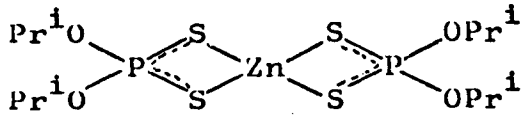
Of the five compounds in this class which were investigated, two proved second only to galvinoxyl in efficiency. These were copper (II), di-N-butyl dithiocarbamate (XIII) and phenothiazine (I), and their influence on the autoxidation of different boranes is shown in Graphs 2,3,5,6,10 and 11.



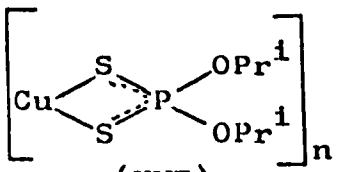
(XIII) R=Bu^n, M=Cu

(XIV) R=Et, M=Zn

Less effective were zinc diethyl dithiocarbamate (XIV) and zinc and copper (I) diisopropyl dithiophosphates (XV) and (XVI). (Graphs 2 and 10.)



(XV)



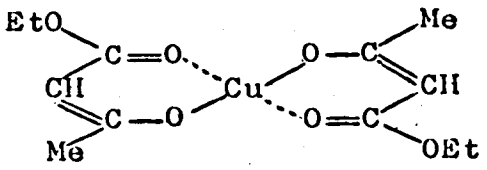
(XVI)

In hydrocarbon systems these compounds operate both by removal of peroxy radicals and by the catalytic decomposition of hydroperoxides (Ch.2). The first of these two reactions does not involve the substrate directly, and would also be expected to occur during borane autoxidations. However, little is known about their activity as regards the decomposition of boron peroxides. In Tables 9 and 10, the yield of peroxide is compared with the oxygen uptake for a large number of runs, and it is clear that decomposition has occurred whenever sulphur or phosphorus containing antioxidants have been used.

The titre was particularly low for copper compounds. Even copper (II), ethyl acetoacetate complex (XVII) and cupric naphthenate* (not shown on any graph), which at

*Naphthenic acid is a complex mixture of carboxylic acids derived from petroleum.

5 moles % had no effect on the autoxidation of tri-s-butyl borane, caused the decomposition of peroxide. At least



(XVII)

10 mols. of peroxide were destroyed per copper atom, indicating a catalytic reaction. Furthermore, both valence states were effective (Table 10). This is reminiscent of the decomposition of hydroperoxide by copper salts (Ch.2 p.33).

$$ROOH + Cu^{2+} \longrightarrow ROO^{\cdot} + H^{+} + Cu^{+} \qquad (ix)$$

$$ROOH + Cu^{+} \longrightarrow RO^{\cdot} + OH^{-} + Cu^{2+} \qquad (x)$$

However, it is unlikely that a similar mechanism is operating here, for two reasons: (a) if the decomposition generated radicals, cupric naphthenate and (XVII) would be expected to accelerate the autoxidation, whereas in fact they had no effect; (b) there is no apparent correlation in Table 10 between the amount of peroxide destroyed and the efficiency of sulphur and phosphorus containing inhibitors.

It thus seems probable that the metal catalyses the rearrangement of peroxide by a non-radical pathway (Equation xi). This in turn implies that the observed effects on the rate of autoxidation must be due mainly to peroxy radical trapping.

$$(ROO)_2BR \xrightarrow{M^{n+}} ROOB(OR)_2 \qquad (xi)$$

The same considerations apply here as with phenols and amines. The inhibitor is competing for the peroxy radical with the extremely facile displacement of an alkyl group from boron, instead of with a slow hydrogen abstraction step (Equations iv and xii).

$$RO_2^{\bullet} + BR_3 \longrightarrow RO_2BR_2 + R^{\bullet} \qquad \text{(iv)}$$

$$RO_2^{\bullet} + RH \longrightarrow ROOH + R^{\bullet} \qquad \text{(xii)}$$

The only two really good retarders, the copper salt (XIII) and phenothiazine (I) were surprisingly similar in effect, considering the difference in structure, in that they showed parallel behaviour with most of the substrates studied. (Graphs 2,3,5,6 and 10.) An exception was the inhibited autoxidation of tri-n-butyl borane in the second mol oxygen uptake. With 5 mole% phenothiazine, the reaction restarted at approximately the uninhibited rate after 140 min. (not shown on Graph 6), whereas in the presence of the same concentration of the copper salt only 1.13 mol. oxygen had been absorbed after 18 hr. This is consistent with the observation that copper catalyses the peroxide rearrangement more efficiently than phenothiazine, and has converted all the substrate to the boronic ester:

$$Bu^nOOBBu^n_2 \longrightarrow (Bu^nO)_2BBu^n \quad \text{(XVIII)} \qquad \text{(xiii)}$$

By analogy with di-n-butyl s-butylboronate (Graph 1), the ester (XVIII) would not be expected to autoxidise

appreciably, although Mirviss[8] has shown that it does take up oxygen slowly. The diethyl ester of s-butylboronic acid, synthesised as a potential substrate for kinetic studies, autoxidised more rapidly ($t_{\frac{1}{2}}$ 12 hr.) although not fast enough for the original purpose (Graph 1).

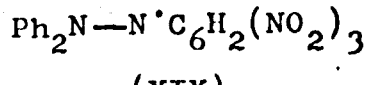
It is interesting to note that although triisobutyl borane reacted with oxygen more slowly than tri-n-butyl borane in the second stage (Graph 1), the same concentration of phenothiazine inhibited the former for less than half the time (Graphs 5 and 6). This is because, unlike the n-butyl isomer, with triisobutyl borane, some of the phenothiazine was utilised in retarding the first mol. uptake.

The similarity in effect between copper (II) dibutyl dithiocarbamate and phenothiazine must have been fortuitous, since these two inhibitors do not remove peroxy radicals in the same way. Other copper compounds, and zinc diethyl dithiocarbamate, were poor retarders, and so (XIII) must have a unique reaction, possibly that described earlier (Ch.2 p.54). Phenothiazine, known to be a very efficient peroxy radical trapper, must in part function as an aromatic amine antioxidant along the lines given on p.50. The role of the sulphur atom is obscure, but dialkyl sulphides are known to decompose hydroperoxide catalytically, as well as

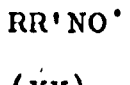
functioning as peroxy radical trappers via the sulphoxide[9], and it is possible that phenothiazine acts in the same way.

c) Free Radicals:-

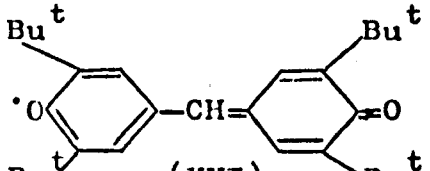
The number of free radicals stable in the presence of oxygen is very limited. Among those which show a useful degree of stability are DPPH (XIX), nitroxides (XX) and galvinoxyl (XXI).



(XIX)



(XX)



(XXI)

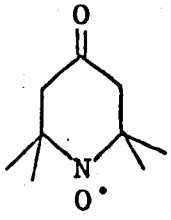
i) DPPH. This compound was almost completely ineffective in these systems (Graphs 2 and 9). It must therefore be concluded that it cannot compete with oxygen or the substrate for either $R^{\bullet}$ or $RO_2^{\bullet}$ radicals. However, the very intense colour was destroyed within seconds of injecting the borane into the gas burette, indicating some reaction between DPPH and borane in the presence of oxygen.

Graph 26 shows the reaction of DPPH with tri-s-butyl borane in the absence of oxygen. The colour of the radical was slowly discharged, and was not changed further on admitting air after $7\frac{1}{2}$ min. A sample of DPPH added to fully autoxidised tri-s-butyl borane was decolourised almost instantaneously, and this is presumable the same reaction as

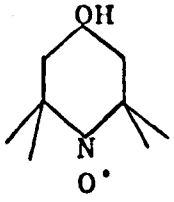
was observed in the gas burette.

The loss of colour in the absence of oxygen recorded on Graph 26 may thus be due either to a direct reaction between borane and DPPH, or to interaction with small amounts of peroxidic material in the borane. (Stock solutions may have slowly peroxidised by diffusion of oxygen through the serum caps, but this was a very small effect, as shown by a comparison of autoxidation experiments on old and new samples).

ii) Nitroxides. Three aliphatic nitroxides were studied, di-t-butyl nitroxide[10], (XX, R=R'=Bu^t), 2,2,6,6 tetramethyl piperid-4-one nitroxide (triacetonamine nitroxide, XXII) and the corresponding piperidol[11] (XXIII).



(XXII)



(XXIII)

Ingold has shown that (XXII) only reacts with alkyl radicals (Ch.2 p.52), and since all three were very similar in effect (Graph 9) the same is probably true for the other two as well.

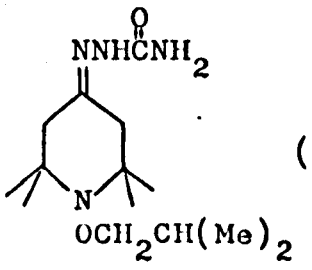
They therefore provide an interesting comparison with phenols, which trap peroxy radicals. Contrary to expectation they were more efficient than phenols during the first

mol. oxygen uptake (Graphs 8 and 9) which indicates that they were competing with some success with oxygen for the alkyl radicals. The effect of (XXII) on the autoxidation of other boranes was much as would be expected from the results with tri-s-butyl borane. (Graphs 2,5 and 6). Measured by its ability to retard the first stage of tributyl boranes, it was the fourth most successful inhibitor for these systems.

A mixture of (XXII) and a peroxy radical trapper would be expected to show synergism (p.54), and an attempt was made to detect this phenomenon with the most efficient of the phenols, BHA (VI) (Graph 23). Since the inhibiting efficiency of the nitroxide alone was found to be proportional to its concentration over a 10-fold range, it is clear that the retarding power of the nitroxide has been approximately doubled in the first stage by the presence of the phenol. However, the retardation of the second stage was only that expected for the amount of BHA used. A much greater synergistic activity would be expected for a mixture of phenothiazine and the nitroxide, since, unlike phenols, the former is capable of retarding the first mol. oxygen uptake when used alone.

Triacetonamine nitroxide was a particularly suitable

compound for the preparation of an isolatable adduct with radicals in the system, as it showed no reaction with borane in the absence of oxygen, or with peroxide. Unlike galvinoxyl it was also quite stable to oxygen. The main disadvantage was that it was only utilised to about 20% (estimated colourimetrically) during an autoxidation, whatever the initial concentration. This problem was overcome in the adduct preparation by adding successive quantities of triisobutyl borane at intervals until the colour was completely discharged (48 hr.), and a compound was then isolated from the reaction mixture as its semicarbazide.



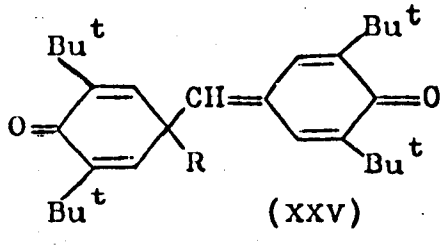
(XXIV)

This was shown to have structure (XXIV) by its mass and NMR spectra (Figs. 8 and 9). Accurate mass measurement of the parent ion gave a molecular formula $C_{14}H_{28}N_4O_2$, and the integrated NMR spectrum can be interpreted satisfactorily in terms of structure (XXIV).

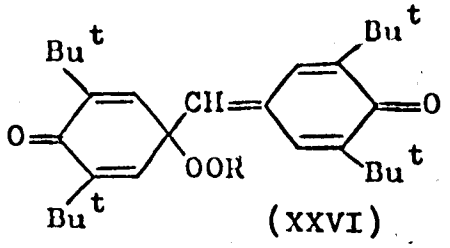
No reaction was observed over 44 hours between triacetonamine nitroxide and triisobutyl borane in the absence of oxygen. This indicates that spontaneous dissociation of

the borane does not occur to a significant extent. It is clear that isobutyl radicals, formed during the autoxidation process, have coupled with the nitroxide to give (XXIV), and this is an important piece of evidence in favour of the free radical chain mechanism (Ch.7).

iii) Galvinoxyl. Galvinoxyl was unique in its ability to give a clean induction period followed by oxygen uptake at the same rate as the uninhibited reaction for all but one of the boranes studied (Graphs 2-6 and 9). This means that it must scavenge at least one of the radicals present with very high efficiency. If alkyl and alkylperoxy radicals react in the same way as cyanoisopropyl and t-butoxy are known to do[12], the most probable products are the p-cyclohexadienones (XXV) and (XXVI).



(XXV)



(XXVI)

In general, the more slowly autoxidising boranes had longer induction periods, but this was not always the case. The most dramatic exception was tri-n-butyl borane, which would be expected from Table 5 and from the low temperature runs of Graph 17 to be the slowest of the three butyl boranes. Even in a concentration of 10 mole%, galvinoxyl only retarded

the first stage slightly, and this is evidence of a very facile propagation reaction (Ch.9).

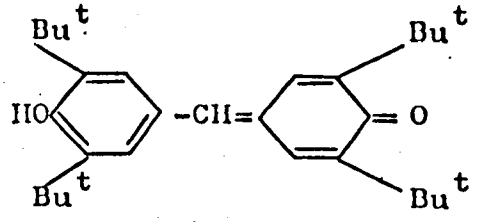
Induction periods with galvinoxyl were not very reproducible (Table 8) and there are at least two points to consider here. Firstly, there is some reaction between borane and galvinoxyl (Graphs 24 and 25) and secondly, galvinoxyl is itself subject to autoxidation[13].

As with DPPH there was a relatively slow reaction between galvinoxyl and borane in the absence of oxygen* (Graphs 24 and 25), but addition of a solution to fully autoxidised tri-s-butyl borane caused a fairly rapid discharge of the colour. Variable small amounts of peroxide in the borane could thus account for the inconstancy of the induction period. A simple calculation bears this out. If two molecules of galvinoxyl are destroyed per >BOOR group, and 0.08% of the bonds are peroxidised (equivalent to 0.025 c.c. oxygen in a typical experiment) then the maximum induction period of about 20 min. will be halved.

* The slow reaction may be accounted for by a process such as $R'^{\bullet} + R_3B \longrightarrow R'BR_2 + R^{\bullet}$, where R' = DPPH or galvinoxyl. This is analogous to chain transfer with phenolic antioxidants.

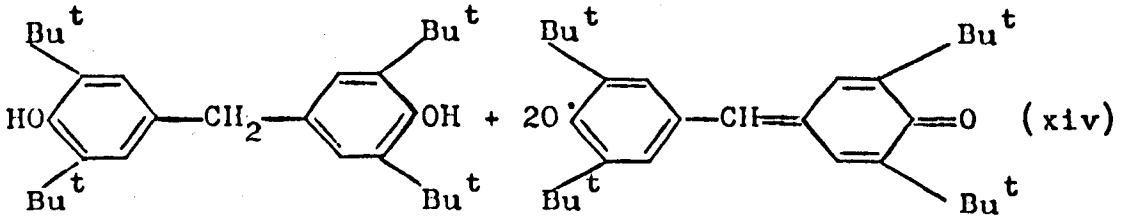
Attempts to get reproducible results by making up and keeping solutions of galvinoxyl under nitrogen, and injecting it into the reaction flask immediately before the borane were only partially successful, since the stock solution still deteriorated with time (Table 8 superscripts e).

Greene[13] has studied the autoxidation of galvinoxyl itself, and has observed an induction period which was due to hydrogalvinoxyl (XXVII), usually present as an impurity.



(XXVII)

This arises because it is difficult to remove from galvinoxyl traces of the precursor, Ethyl 702, with which it reacts.



(xiv)

⟶ 3 (XXVII)

The induction period was reported to be about 15 mins. per cent added hydrogalvinoxyl.

A solution of galvinoxyl was prepared under nitrogen with 6% added hydrogalvinoxyl, and this failed to increase the induction period in the autoxidation of tri-s-butyl borane.

In fact although the value obtained (ca. 20 min. for 1 mole %) was in agreement with other freshly made solutions of galvinoxyl, the reproducibility was no better than previously (Table 8, superscripts f).

A sample of the galvinoxyl used in all the experiments was autoxidised and shown to have an induction period of about 180 minutes (Graph 18). From Greene's data[13], this means that it fortuitously contained about 12% hydrogalvinoxyl. It is therefore reasonable to assume that when the inhibitor was injected immediately before the borane, induction periods less than 180 minutes were not affected by autoxidation of the inhibitor.

It was also observed that when degassed solutions of galvinoxyl in "spectrosol" hexane were stored for three days under nitrogen, the characteristic band at 429 mμ had disappeared. This band was still present in a control solution where oxygen had not been completely excluded. This result is difficult to explain, but it has certain features in common with the stabilisation of styrene by small amounts of oxygen.

The variable induction periods resulting with galvinoxyl are not readily explained and a careful experimental

investigation is necessary if the effects of extraneous impurities and chain transfer reactions (increasingly important as the concentration of inhibitor is increased) are to be evaluated.

It was concluded that the most reproducible inhibition was obtained using freshly prepared solutions (less than 24 hours old) of galvinoxyl containing hydrogalvinoxyl, the solution being stored under nitrogen at -20^{o} and injected into the burette immediately before the borane.

The significance of the induction period with galvinoxyl is discussed further in Chapter 9.

d) Sodium Hydroxide:-

In one particular experiment, a solution of tr-n-butyl borane only took up just over 1 mol. oxygen, instead of the usual 1.85. In an investigation to find out why this was so, various agents which might have been accidentally introduced were added to the pure benzene. The presence of excess powdered sodium hydroxide caused the extraneous result, and hence the use of sodium wire to dry the benzene probably gave rise to this phenomenon.

The benzene was about 0.8 mmolar in sodium hydroxide after shaking with the powdered solid, (about 10% of the

borane), and at this concentration there was no detectable change in oxygen uptake or peroxide titre. Graph 15 shows the autoxidation of the three tributyl boranes in the presence of excess powered sodium hydroxide. While both triisobutyl and tri-n-butyl boranes took up 1 mol. oxygen and had negligible peroxide titres, tri-s-butyl borane took up 1.8 mol. oxygen at the usual rate, but with a reduced peroxide titre of 0.76 mol. (Table 9).

It is very likely that the hydroxide ion coordinates with the borane (Equation xv).

$$R_3B + OH^- \rightleftharpoons R-\underset{\displaystyle R}{\overset{\displaystyle R}{\overset{|}{\underset{|}{B^-}}}}-OH \qquad \text{(xv)}$$

(XXVIII)

A structure such as (XXVIII) could not participate in the free radical chain autoxidation postulated. However, the first mol. oxygen uptake is not slowed down detectably, which argues that the equilibrium in reaction (xv) cannot lie completely to the right. The absence of peroxide can be explained by assuming that sodium hydroxide causes the rearrangement of boron peroxides.as do water and other bases[14]. This rearrangement to boronic esters must be a rapid process for the iso- and n- isomers. The second mol. oxygen uptake is quite fast for tri-s-butyl borane, however, and must occur more rapidly than the rearrangement.

The peroxide titre obtained in this case is close to that required by the reaction:

$$R_nB(OOR)_{3-n} \longrightarrow (RO)_{2n}B(OOR)_{3-2n} \qquad (xv)$$

e) Conclusion:-

Most antioxidants were found to inhibit or retard the autoxidation of all the organoboron compounds studied, thus establishing the general applicability of the free radical chain mechanism postulated in Chapter 7.

i) The efficiency of inhibitors was related to their ability to compete with oxygen or the substrate for alkyl or peroxy radicals. The observed order of scavenging ability was:

Galvinoxyl > copper dibutyl dithiocarbamate $\simeq$ phenothiazine > nitroxides > DPPH, phenols and amines.

The adduct between the intermediate isobutyl radicals and triacetonamine nitroxide was isolated and identified.

ii) Sulphur, phosphorus and copper containing compounds, sodium hydroxide and some phenols catalysed the decomposition of peroxide, by non-radical mechanisms.

Table 10 - The Destruction of Peroxide by Inhibitors

Borane	Inhibitor (Moles %)	Mol. peroxide / mol. O_2 uptake	Perox. destroyed per mol. inhibitor
Bu^s_3B	246 Bu^t Phenol, 5.0	0.65	13
"	S. Crystal[a], 10.0[b]	0.82	3
"	BHT, 5.0	0.81	7
"	Ethyl 702, 10.0[b]	0.91	1.5
"	Ethyl 720, 10.0[b]	0.94	1
"	Cyanamid 2246, 10.0[b]	0.91	1.5
"	20H 4Bu^t Phenol, 11.0[b]	0.96	< 1
"	BHA, 5.0	0.99	< 1
"	26MeO Phenol, 5.0	1.01	< 1
"	$(Bu_2NCS_2)_2Cu$, 2.0	0.71	29
$(IPC)_2BBu^s$	" 0.1	0.89	≃ 160
$(IPC)_2BBu^n$	" 1.3	0.70	36
Bu^s_3B	$(Pr^iO)_2PS_2Cu$, 5.0	0.65	13
"	Cu^{II} Naphthenate, 5	0.70	11
"	Cu^{II} EEA[c], 5.0	0.65	13
$(IPC)_2BBu^s$	$(Et_2NCS_2)_2Zn$, 0.75	0.68	60
Bu^s_3B	$[(Pr^iO)_2PS_2]_2Zn$, 5.0	0.87	5
"	Phenothiazine, 10.0	0.77	5
$(IPC)_2BBu^s$	" 0.3	0.96	≃ 10
$(IPC)_2BBu^n$	" 2.8	0.91	5

[a] Santowhite Crystal (Table 9)

[b] Of OH Groups

[c] Copper ethyl acetoacetate.

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Chapter 9

Autoxidation of Organoboron Compounds under Conditions of High-Speed Stirring

a) Importance of Efficient Gas-Liquid Mixing:-

In any experiment to measure the rate of a chemical reaction, it is important to achieve thorough mixing of the reagents in order to avoid diffusion control. This condition is easily met in many cases, but where reaction is very rapid more stringent precautions must be taken to ensure that the observed rate is not dependent on stirring speed. Where one reagent has to cross a phase boundary, as in autoxidations in solution, this problem takes on a new form, since here the liquid must always be saturated with gas. This requires that the surface area of the interface must be large enough to allow the gas to dissolve as fast as it is reacting.

Usually in autoxidations this effect can be achieved using a relatively slow rate of stirring. However, many organoboron compounds autoxidise very rapidly, and special precautions are necessary to avoid obtaining pseudo kinetics due to the solution being starved of oxygen. In this work the problem was overcome by using the reaction flask A of Fig.7, which embodied several features designed to give

thorough gas-liquid mixing. (a) The powerful external magnet permitted stirring speeds of up to 6,000 r.p.m. (b) Indentations in the side of the flask ensured turbulent flow in the liquid. (c) The rate of gas recycling through the hollow stirrer depended on stirrer speed, and consequently, at higher speeds the bubbles of gas were much smaller. The net result of these three factors was the attainment of a very large gas-liquid interface, which was continuously increasing with stirrer speed.

Even with this apparatus, stirring speeds in excess of 2,000 r.p.m. were usually required before the reaction was proceeding at its true rate, illustrating the fact that published rates of autoxidation of organoboron compounds err on the slow side. Hansen and Hamann[2] appear to be the only authors who have noted this phenomenon. They reported that the autoxidation of triethyl borane was still stirrer controlled at 1,000 r.p.m.

Graph 16 shows the autoxidation of 1-phenylethylboronic acid in benzene at two concentrations and stirring speeds. Curve B, 43 mmolar and 1,000 r.p.m., obeys first order kinetics over the middle part of the range, k_1 (observed) = 0.51 $min.^{-1}$ at 25°. This compares with a quoted value[1] of 0.13 $min.^{-1}$ for a solution 50 mmolar in benzene at 20°.

Curve A, 7.7 mmolar and 3,000 r.p.m., shows the true rate of autoxidation of this compound. No accurate kinetic measurements could be made as most of the oxygen was taken up within thirteen seconds.

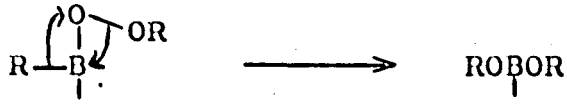
Many other organoboron compounds also autoxidise considerably faster than has previously been supposed. The results obtained in this study, together with previously published values, are shown in Table 11. Tri-n-butyl and triisobutyl boranes were particularly interesting, since for these compounds there was a distinct change in rate between the first and second mol. oxygen uptake (Graph 1). This has not been observed before, and the reason for this is made clear in Graph 12, which shows the autoxidation of tri-n-butyl borane at various stirrer speeds. Above 3,000 r.p.m. the reaction was independent of stirring speed, but at 1,000 r.p.m. a smooth curve was obtained. At this lower speed, the rate of dissolving of oxygen was being measured, and not the true rate of autoxidation of tri-n-butyl borane.

b) Rearrangement of Peroxide:-

Apart from the relative rates, the most striking difference between hydrocarbon and borane autoxidations is in the end products. While peroxide is formed initially in both cases, it can rarely be isolated in large amounts from

saturated hydrocarbon autoxidations (Ch.2). However, dilute solutions of boranes in inert solvents give practically quantitative yeilds of peroxide, based on the oxygen uptake. The reason for this lies in the ease of formation of the latter, rather than in their inherent stability. Organoboron compounds can be autoxidised at room temperature or below in a few minutes, whereas relatively high temperatures and long reaction times are required to autoxidise most hydrocarbons, and under these conditions the hydroperoxide breaks down (Ch.2.). Boron peroxides seem, in fact, to be less stable than many purely organic peroxides, particularly in the presence of unoxidised boron-carbon bonds. The triperoxy compounds, $(ROO)_3B$ are reported to be considerably more stable[7], but these are not available via autoxidation.

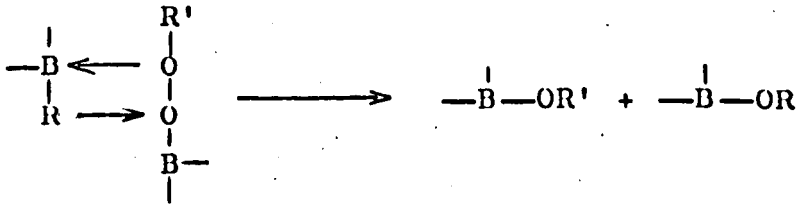
Three mechanisms have been put forward to account for the rearrangement of peroxide. Davies[8] has suggested an intramolecular redox reaction analogous to other 1,2 nucleophilic migrations:



(i)

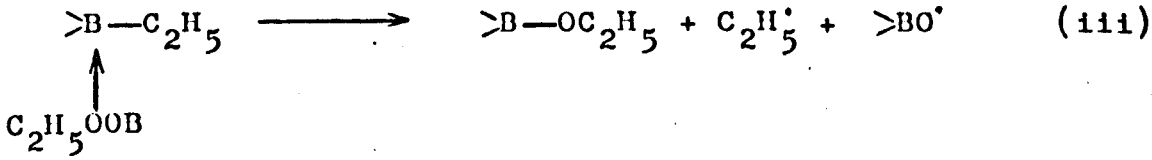
However, Mirviss[9] found that the addition of excess tri-n-hexyl borane to a partially autoxidised dilute solution of tri-n-butyl borane in heptane led to a marked drop in peroxide content, and to a corresponding amount of n-hexanol

on saponification. This, with other less convincing evidence, led him to propose an intermolecular rearrangement:



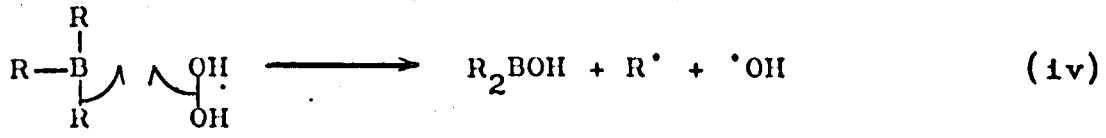
(ii)

Finally Hansen[2] has explained the ability of a mixture of triethyl borane and the peroxide to initiate free radical polymerisations by the reaction:



(iii)

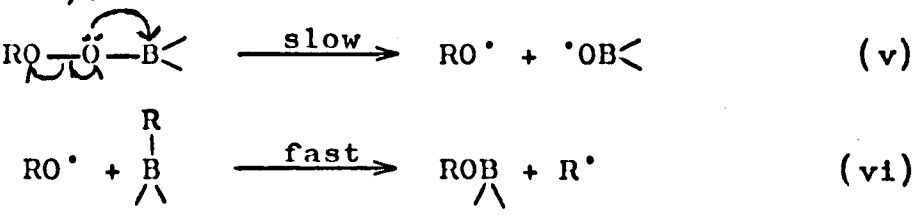
His scheme was based on the failure of the borane or the peroxide alone to initiate polymerisation, and on the formation of substantial quantities of ethyl iodide when borane and peroxide were allowed to react in the presence of iodine. To account for the borate ester usually formed as a result of this rearrangement, Hansen suggested that the majority of the radical pairs recombined within a solvent cage. A similar reaction has been suggested by Bigley and Payling for the attack of neutral hydrogen peroxide on trialkyl boranes:



(iv)

Homolysis of the O—O bond would give a fourth

mechanism, also involving free radicals. This might be expected to occur more readily than with hydroperoxides, due to $p\pi - p\pi$ donation of electrons from oxygen into boron, and hence may be significant at room temperature. The alkoxy radical thus produced could then displace an alkyl radical to give the same overall reaction as was suggested by Hansen[2], but with a unimolecular rate determining step, (Equations v and vi).



$$RO{-}\ddot{O}{-}B< \xrightarrow{\text{slow}} RO^{\bullet} + {}^{\bullet}OB< \qquad (v)$$

$$RO^{\bullet} + \underset{\diagup\diagdown}{\overset{R}{|}B} \xrightarrow{\text{fast}} \underset{\diagup\diagdown}{ROB} + R^{\bullet} \qquad (vi)$$

There is thus the possibility of intra- or intermolecular rearrangement of peroxide, and this may take place by ionic or free radical mechanisms. It may be that all of these operate to differing extents, and it should be possible to verify this by careful experimental work. For instance, the reaction of an optically active borane with peroxide may indicate the relative importance of ionic and free radical processes, while a study of the kinetics of decomposition of a single peroxy borane would distinguish between an inter- and an intramolecular reaction.

Bearing in mind the occurrence of this rearrangement, the total oxygen uptake and peroxide formed must depend on

the concentration of the organoboron compound, and the length of time over which it is autoxidised. The presence of peroxide was not detected at all by Johnson and van Campen[11], who allowed air to diffuse into neat tri-n-butyl borane and found that only 1 mol. of oxygen was taken up:

$$Bu^n_3B + O_2 \longrightarrow (Bu^nO)_2BBu^n \qquad \text{(vii)}$$

The autoxidation of a dilute solution of the same compound in an inert solvent, however, leads to the rapid uptake of 1.6-1.8 mol. oxygen and an almost quantitative yield of peroxide[5]. This difference can be attributed to the low availability of oxygen in the former case, due partly to concentration and partly to lack of stirring, resulting in rearrangement of the peroxide as fast as it was formed.

One of Mirviss's[9] arguments for an intermolecular peroxide rearrangement, the inverse dependence of peroxide content on borane concentration, is invalid for much the same reason. From the quantities of borane and flow rates of air quoted, it is clear that the autoxidations were oxygen-starved in the early stages, even for dilute solutions. Since this problem would become more acute at higher concentrations, peroxide rearrangement by any mechanism would become more significant, and this would give the observed results.

A fact which is not so readily explained is the evidence of Zutty and Welch[12] for a borane-oxygen complex which slowly rearranged to peroxide, (Ch.3 p.67). Mirviss[9] has observed a similar effect, and has substantiated the suggestion of Davies[5] that traces of ferric ion made a profound difference to the peroxide titre. After passage of air through a 2% solution of triethyl borane in heptane at 0^{o} (87 min.) the peroxide content was 0.438 mol. and the total oxygen uptake was 1.68 mol. The system was purged with nitrogen (21 min.) during which time the peroxide titre increased to 0.605 mol. However, when the titrations were carried out in the presence of traces of ferric ion, little or no increase was obtained[9]. The role of the ferric ion is obscure, but a long-lived borane-oxygen complex in such high concentrations is incompatible with the proposed free radical mechanism of autoxidation, and its existence must be ruled out.

c) Detailed Mechanism of Autoxidation:-

The mechanism put forward in Chapter 7 to account for the autoxidation of organoboron compounds is analogous to that discussed in Chapter 2 for hydrocarbons.

Initiation $R_3B \xrightarrow{r_i} R^{\cdot}$ (viii)

$$\text{Propagation} \begin{cases} R^{\bullet} + O_2 \xrightarrow{k_2} RO_2^{\bullet} & \text{(ix)} \\ RO_2^{\bullet} + R_3B \xrightarrow{k_3} RO_2BR_2 + R^{\bullet} & \text{(x)} \\ \left[RO_2^{\bullet} + RH \xrightarrow{k_3'} ROOH + R^{\bullet}\right] & \text{(x')} \end{cases}$$

$$\text{Termination} \begin{cases} 2R^{\bullet} \xrightarrow{k_4} & \text{(xi)} \\ R^{\bullet} + RO_2^{\bullet} \xrightarrow{k_5} & \text{(xii)} \\ 2RO_2^{\bullet} \xrightarrow{k_6} & \text{(xiii)} \end{cases} \text{Non-radical products}$$

The mathematical treatment of these equations is simplified considerably by neglecting termination reactions involving the alkyl radical (Equations xi and xii). This is justified in hydrocarbon systems because the propagation reaction (ix) is so much faster than (x') that the standing concentration of alkyl radicals is negligible compared to that of alkylperoxy radicals. However, borane autoxidations are rapid, and in some cases show little dependence on temperature. This indicates a low overall activation energy and implies that reactions (ix) and (x) may be of comparable speeds. Thus the assumption that chain ending occurs only between $RO_2^{\bullet}$ radicals cannot be made automatically.

Fortunately, in the two cases investigated, (Graphs 13 and 14) the autoxidation rate was oxygen pressure independent down to about 100 mm. This shows that reaction (ix) was not rate determining, and hence must have been significantly faster than (x) in these systems. Thus for tri-s-butyl

borane and 1-phenylethylboronic acid, stationary state consideration lead to the same simple kinetic equation as for hydrocarbons (See Appendix for derivation).

$$\frac{-d[O_2]}{dt} = k_3 k_6^{-\frac{1}{2}} r_i^{\frac{1}{2}} [R_3B] \qquad \text{(xiv)}$$

For convenience, it is assumed in the following discussion that this equation can also be applied to the other organoboron compounds studied. Strictly speaking the oxygen pressure independence should be verified for each case, but within the limits of this preliminary study, the assumption is felt to be justified.

The two figures which would made an interesting comparison with hydrocarbon systems are the initiation rate constant, k_i, and the propagation rate constant k_3. Unfortunately a complete evaluation of these was not possible due to several complicating factors. In hydrocarbon autoxidations, k_3 can usually be studied independently of k_i. This is because the normal rate of autoxidation is very slow, and can be accelerated by the addition of an initiator which decomposes at a far greater rate than the chance production of radicals by the substrate. Thus r_i will have a constant, known, value depending only on the added initiator. This simplifying feature is absent from

organoboron autoxidations, since these normally react so fast that the natural initiation processes must be considered. These are probably complex (see below) and must depend on the concentration of the species involved.

A second problem arises from the amount of oxygen consumed. Where this is more than one molecule, it is necessary to distinguish clearly between the different stages, as no clear cut kinetics would be expected if these stages overlapped. $(IPC)_2BBu^s$ and $(IPC)_2BBu^n$ therefore proved unsuitable - no change in rate was discernible on completion of the first mol. uptake. On the other hand, where such a distinct break did occur, as with triisobutyl and tri-n-butyl boranes, the first molecule was taken up too rapidly to measure the rate at 25° (Graph 1). For these compounds most calculations were made assuming a starting concentration of monoperoxy borane, $ROOBR_2$, equivalent to the initial concentration of tributyl borane.

In spite of these difficulties, several interesting facts emerge from the mathematical treatment of results, and the initiation mechanism in particular would amply repay further study.

i) Initiation. Although galvinoxyl suffers from

disadvantages (Ch.8 p.186), under carefully controlled conditions fairly reproducible induction periods could be obtained (Table 8). These can be converted into rates of initiation, r_i, from the relation,

$$r_i = \frac{n[AH]_o}{\tau} \qquad (xv)$$

where $[AH]_o$ is the initial concentration of inhibitor, n is the number of radicals destroyed per inhibitor molecule, and τ is the induction period measured by the tangent intercept method. Values of r_i are shown in Table 12, assuming n=1 for galvinoxyl. These must be regarded as maximum rates, due to the possibility of galvinoxyl being destroyed by alternative routes. Table 12 also includes an approximate value for the rate of initiation of autoxidation of di-n-butyl n-butylperoxy borane, $Bu^n_2BOOBu^n$, in the presence of phenothiazine. This calculation is valid, since (a) the first mol. uptake was not retarded, and so it is unlikely that any inhibitor was utilised, and (b) after 140 minutes the reaction took off at approximately the uninhibited rate (not shown on Graph 6).

Values of k_i, the initiation rate constant, can only be calculated from the data of Table 12 if the mechanism of initiation is known. Assuming that this only involves the boron substrate and oxygen, i.e. there is no impurity present,

r_i and k_i are related by the equation:

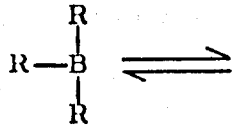
$$r_i = k_i \left[BR_3\right]^x \left[O_2\right]^y \qquad \text{(xvi)}$$

where x and y must be determined. This can be done by consideration of equation (xiv), which contains a term in $r_i^{\frac{1}{2}}$, and therefore simplifies to

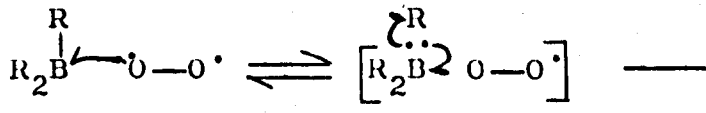
$$-\frac{d\left[O_2\right]}{dt} = K \left[BR_3\right]^{1+\frac{x}{2}} \left[O_2\right]^{\frac{y}{2}} \qquad \text{(xvii)}$$

where $K = k_3 k_6^{-\frac{1}{2}} k_i^{\frac{1}{2}}$. Thus the overall reaction order in borane and oxygen should give a powerful clue to the initiation mechanism. Table 13 shows the overall reaction order for a number of different substrates, determined by plots of $\log -\frac{d\left[O_2\right]}{dt}$ against $\log \left[BR_3\right]$.

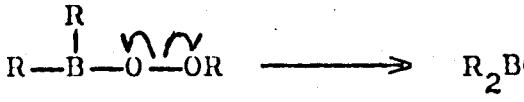
It is clear from this table that several mechanisms must be operating. Neglecting for the moment the possibility of initiation by an impurity, values of x from 1.0 to 2.0 can be interpreted as due to combinations of the suggested mechanisms (xviii)-(xxi).



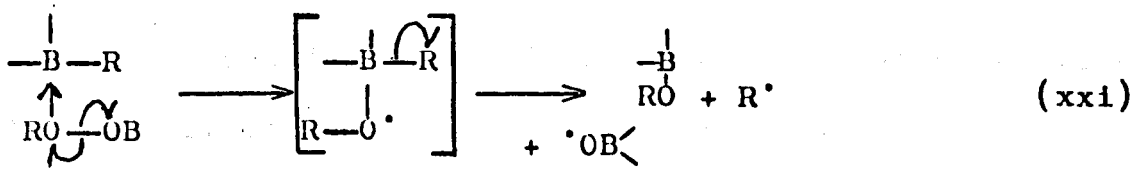
(xviii)



(xix)



(xx)



(xxi)

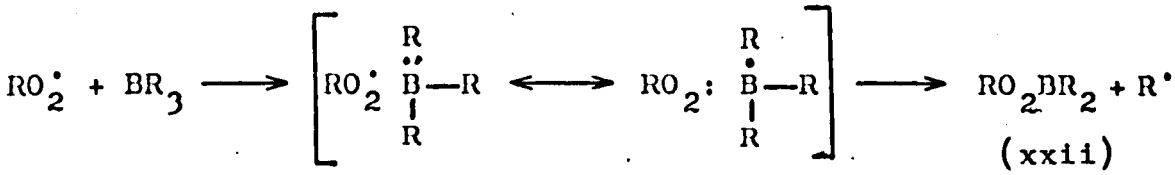
Reaction (xxi) is the only one which is second order in the boron substrate, and would be expected to be favoured by higher concentrations of peroxide. Although this is born out by the first three rows of Table 13 (see also Graph 19), at higher concentrations still the order starts to fall again. It is tempting to attribute this effect to the multiplicity of mechanisms for the peroxide rearrangement, but it is dangerous to speculate too far at this stage. In the first case the figures in Table 13 are calculated from single experiments, and secondly, the substrates were not rigourously purified. The assumption that after one mol. oxygen uptake the concentration of monoperoxy borane, R_2BO_2R, was equal to the initial concentration of trialkyl borane, is also open to question. For carefully purified unoxidised borane, initiation pathways (xx) and (xxi) are not available, and spontaneous dissociation of the borane into radicals seems unlikely. There remains only reaction (xix), for which an overall reaction order in borane of 1.5 and in oxygen of 0.5 would be predicted from equation (xiv). However, the implication from Table 13 is that when there is no peroxide present, the overall order in borane is 1, suggesting an initiation step independent of boron substrate. This applies not only to tributyl boranes at -74^{o}, but to

s-butylboronic anhydride at 25°, in agreement with Grummitt's results[13]. It is difficult to postulate an initiation mechanism to account for this, and the results provide some evidence for the presence of an impurity, possibly peroxide, which supplies the initiating radicals.

The conclusion, therefore, is that several different modes of initiation must operate in the organoboron autoxidations reported here, one of which probably involves bimolecular decomposition of the peroxide (Equation xxi). It is highly likely that impurities also give rise to free radicals, and for these reasons initiation rate constants, k_i, cannot be calculated from the induction periods in the presence of galvinoxyl.

ii) Propagation. The reaction of the alkylperoxy radical with organoboron compounds is interesting because of its great facility. This is illustrated by the low overall activation energy of the autoxidation, E, which is obtained by considering the temperature dependence of the reaction rate. For 1-phenylethylboronic acid (Graph 22) and for the first mol. oxygen uptake of tri-s-butyl (Graph 21) and tri-n-butyl boranes (not shown) the rate varied so little with temperature that an estimate of E could not be made.

The main reason for the facility of the propagation step must be the availability of the vacant π orbital of boron for co-ordination. Although the alkylperoxy radical is normally considered to be electrophilic, this is only a relative effect, and it would be expected that such a reactive species would readily attack the boron atom. The boron-oxygen bond thus formed would be considerably stronger than the boron-carbon one broken[14], and a plausible mechanism for this step is shown below:



(xxii)

A similar displacement of an alkyl radical from boron has been proposed by other authors[1,14,15].

The rate constant for this reaction, k_3, will depend on three factors: (a) the steric protection afforded to the boron atom in BR_3, (b) the ability of the groups R to release electrons into the boron atom, and (c) the stability of the radical R$^{\cdot}$. Steric hindrance to the approach of alkylperoxy would produce the order of reactivity R = $Bu^n > Bu^i > Bu^s > Bu^t$ in the borane. However, the reverse order would be predicted for the release of steric strain in going from sp^2 to sp^3 in the transition state, or if electronic effects (b) and (c) were dominant. The

observed overall rate of reaction must be due not only to the influence of these factors on k_3, but to the way they affect the initiation and termination processes. Great care must therefore be taken in interpreting gross reactivities in terms of individual rate constants, but bearing this in mind, certain observations can be made.

For instance, the efficiency of peroxy radical trappers in different systems gives a guide to the relative values of k_3. Graphs 5,6 and 10 show the effect of phenothiazine at the same relative concentration on the autoxidation of iso-, n- and s-butyl boranes at 25°. The retarded rate of the first mol. uptake is in the order $Bu^n > Bu^s > Bu^i$, indicating the same order in the magnitude of k_3. Similarly, the initiation rate for $(IPC)_2BBu^n$ is three times as great as for triisobutyl borane at the same concentration, in spite of the very much greater overall autoxidation rate of the latter. This illustrates the fact that a relatively slow propagation step is responsible for the reduced reactivity of $(IPC)_2BBu^n$.

The quoted order of reactivity in the overall autoxidation, R = t-alkyl > s-alkyl > p-alkyl (Table 5), must be questioned in view of the hitherto unsuspected great speed of the first mol. oxygen uptake for tributyl boranes. The reactivity sequence at room temperature could not be

established, but at -74°C the reaction $R_3B \longrightarrow ROOBR_2$ was slow enough to allow the rates to be measured, and they fell in the order $Bu^i > Bu^s > Bu^n$, (Graph 17). The reactivity of triisobutyl borane is surprising, since the isobutyl radical is less stable than the s-butyl, while the steric hindrance due to three such groups around boron must be considerably greater than for n-butyl. However, the difference in rates was only slight, even at this temperature, and hence the above sequence would be influenced by quite small changes in structure. Confirmation of this comes from the autoxidation of the monoperoxides, Bu_2BO_2Bu, at 25°, for which the order is $Bu^s > Bu^n > Bu^i$, (Graph 1). Yet another sequence, $Bu^n > Bu^s$, was observed for $(IPC)_2BBu^s$ and $(IPC)_2BBu^n$, and here steric protection of the boron atom is clearly the dominant factor.

It is concluded that it is not possible to give a generalised order of reactivity of trialkyl boranes towards oxygen, since the overall rate of autoxidation depends on several rate constants which are themselves influenced to differing extents by changes in temperature and in the structure of the alkyl group.

During the course of this work, a search was made for a better substrate, and the borinic esters may fulfil this

function, since s-butyl di-s-butyl borinate, $Bu^s_2BOBu^s$, was found to take up only one molecule of oxygen, with a reaction half life of 16 minutes. However, Graph 1 reveals that an increase in rate occurs during the early part of this reaction, suggesting autocatalysis. This in turn implies that peroxide decomposition may be increasing the rate of initiation, even at this low conversion.

The conclusion to be drawn from this section is that the complexity of the initiation processes prevents absolute rate constants from being evaluated. An essential preliminary to a detailed kinetic investigation would be the preparation of substrates free from traces of impurity. A study of peroxide rearrangement and other radical producing processes would then be necessary before the magnitude of k_3 could be determined for different organo-boron compounds.

Table 11 - Rates of Reaction of Organoboron Compounds with Oxygen[a]

Compound	Previous Work				This work		
	Conc.	T°C	$t_{\frac{1}{2}}$	Ref.	Conc.	T°C	$t_{\frac{1}{2}}$
$PhCH(Me)B(OH)_2$	56 mmolar	20°	6 min.	1	7.8 mmolar	25	ca. 8 sec.
$Bu^i{}_3B$	32[b] "	20°	15 "	1	7.5 "	25	ca. 10 sec.
"	33 "		50[d] "	3			108 min.[d]
$(Bu^sBO)_3$	23 "	21°	5[e] "	1	4.4 "	25	1.2 min.[e]
"	23 "		26[e] "	4			
$Bu^n{}_3B$	12 "		60[f] "	5	7.3 "	25	ca. 10 sec. (9.5 min.[f])
$Bu^s{}_3B$	16[c] "		45[g] "	6	5.8 "	25	ca. 10 sec. (15 min.[g])

[a] Benzene solvent
[b] Bromobenzene solvent
[c] Ether solvent
[d] Time for 1.82 mol. oxygen uptake
[e] Time for 0.5-1.5 mol. oxygen uptake (based on trimer)
[f] Time for 1.67 mol. oxygen uptake
[g] Time for 1.91 mol. oxygen uptake

Table 12 - Rates of Initiation of Autoxidation of Organoboron Compounds at 25°C

Substrate	Conc. moles $l.^{-1}$	Inhibitor AH	$[AH]_o$ moles $l.^{-1}$	γ min.	r_i moles $l.^{-1}$ $min.^{-1}$
Bu^s_3B	5.75×10^{-3}	Galvinoxyl	2.88×10^{-5}	11.7	2.5×10^{-6}
Bu^i_3B	7.5×10^{-3}	"	3.75×10^{-4}	58	6.5×10^{-6}
$(IPC)_2BBu^s$	10.1×10^{-3}	"	1.01×10^{-5}	17	0.60×10^{-6}
$(IPC)_2BBu^n$ [b]	7.5×10^{-3}	"	1.05×10^{-4}	5.6	19×10^{-6}
$(Bu^sBO)_3$ [b]	4.4×10^{-3} [c]	"	3.38×10^{-5}	126	0.26×10^{-6}
$Bu^n_2BO_2Bu^n$ [b]	9.2×10^{-3}	Phenothiazine	4.6×10^{-4}	140	$3.3n^a\times10^{-6}$

[a] The stoichiometry is uncertain for phenothiazine.

[b] Based on a single experiment.

[c] Concentration of trimer.

Table 13 - Orders of Reaction in the Autoxidation of Organoboron Compounds in Benzene

Substrate	Initial Conc., moles $l.^{-1}$	Temp. $^{o}C.$	Overall order in Borane, $1 + \frac{x}{2}$	Overall order in Oxygen, $\frac{y}{2}$	Order in Borane in Initiation Step, x
$Bu^{s}{}_{2}BO_{2}Bu^{s}$	1.2×10^{-3}	25	1.5		1.0
"	2.3 "	"	1.8		1.6
"	4.1 "	"	2.0		2.0
"	5.8 "	"	1.9		1.8
"	8.7 "	"	1.7		1.4
"	11.6 "	"	1.6		1.2
"	1.6 "	"		0	
"	7.7[a] "	0	1.8		1.6
$Bu^{n}BO_{2}Bu^{n}$	7.4 "	25	1.7		1.4
$Bu^{i}BO_{2}Bu^{n}$	7.5 "	25	b		
$Bu^{s}{}_{3}B$	2.2[d] "	-74	0.9[c]		0
$Bu^{n}{}_{3}B$	1.8[d] "	-74	0.9[c]		0
$Bu^{i}{}_{3}B$	2.5[d] "	-74	b		
$(Bu^{s}BO)_{3}$	4.4[e] "	25	1.0		0
$PhCH(Me)B(OH)_{2}$	3.9 "	25		0	

[a] Chlorobenzene Solvent.

[b] Reaction order varied throughout run.

[c] Approximate values from scattered points.

[d] Ether Solvent

[e] Concentration of trimer.

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APPENDIX

Derivation of Kinetic Equation for Free Radical Chain Reactions:-

Consider the chain reaction represented by Equations (i) - (viii) below:

$$RH \xrightarrow{r_i} R^{\bullet} \qquad \text{(i)}$$

$$R^{\bullet} + O_2 \xrightarrow{k_2} RO_2^{\bullet} \qquad \text{(ii)}$$

$$RO_2^{\bullet} + RH \xrightarrow{k_3} ROOH + R^{\bullet} \qquad \text{(iii)}$$

$$2R^{\bullet} \xrightarrow{k_4} \qquad \text{(iv)}$$

$$R^{\bullet} + RO_2^{\bullet} \xrightarrow{k_5} \qquad \text{(v)}$$

$$2RO_2^{\bullet} \xrightarrow{k_6} \qquad \text{(vi)}$$

(iv), (v), (vi): Non-radical Products

$$RO_2^{\bullet} + AH \xrightarrow{k_7} \qquad \text{(vii)}$$

$$R^{\bullet} + AH \xrightarrow{k_8} \qquad \text{(viii)}$$

(vii), (viii): Inactive Products

where RH is the hydrocarbon substrate and reactions (vii) and (viii) represent chain ending by an inhibitor AH.

Assuming that the chains are long, i.e. that the oxygen consumed in the initiation step (i) is negligible, the rate of uptake of oxygen is given by:

$$-\frac{d[O_2]}{dt} = k_2[R^\cdot][O_2] \qquad (ix)$$

The "stationary state" approximation uses the assumption that the rate of removal of a reactive intermediate is equal to its rate of formation, i.e. in this case,

$$\frac{d[R^\cdot]}{dt} = \frac{d[RO_2^\cdot]}{dt} = 0 \qquad (x)$$

In the absence of inhibitor, equation (x) leads to the relations

$$\frac{d[R^\cdot]}{dt} = r_i + k_3[RO_2^\cdot][RH] - k_2[R^\cdot][O_2] - k_4[R^\cdot]^2 - k_5[R^\cdot][RO_2^\cdot] = 0 \qquad (xi)$$

and

$$\frac{d[RO_2^\cdot]}{dt} = k_2[R^\cdot][O_2] - k_3[RO_2^\cdot][RH] - k_5[R^\cdot][RO_2^\cdot] - k_6[RO_2^\cdot]^2 = 0 \qquad (xii)$$

from Equations (i) - (vi).

Equations (ix), (xi) and (xii) can be solved to give a complex general expression not involving $[R^\cdot]$ or $[RO_2^\cdot]$. In most cases, however, the simplifying assumption can be made that the concentration of alkyl radicals, $[R^\cdot]$, is very much less than $[RO_2^\cdot]$ due to the great speed of reaction (ii)

relative to (iii). Hence termination reactions involving $R^{\bullet}$ can be ignored, and Equations (xi) and (xii) become (xiii) and (xiv) respectively.

$$r_i + k_3 \left[RO_2^{\bullet}\right]\left[RH\right] - k_2\left[R^{\bullet}\right]\left[O_2\right] = 0 \qquad \text{(xiii)}$$

$$k_2\left[R^{\bullet}\right]\left[O_2\right] - k_3\left[RO_2^{\bullet}\right]\left[RH\right] - k_6\left[RO_2^{\bullet}\right]^2 = 0 \qquad \text{(xiv)}$$

whence

$$r_i = k_6\left[RO_2^{\bullet}\right]^2 \qquad \text{(xv)}$$

From (ix) and (xiii)

$$\frac{-d\left[O_2\right]}{dt} = r_i + k_3\left[RO_2^{\bullet}\right]\left[RH\right] \qquad \text{(xvi)}$$

Ignoring r_i and replacing for $\left[RO_2^{\bullet}\right]$ from (xv) gives

$$\frac{-d\left[O_2\right]}{dt} = k_3 k_6^{-\frac{1}{2}} r_i^{\frac{1}{2}}\left[RH\right] \qquad \text{(xvii)}$$

In the case where an added inhibitor, AH, scavenges $RO_2^{\bullet}$ radicals sufficiently well to prevent termination by reaction (vi), Equation (xiv) becomes

$$k_2\left[R^{\bullet}\right]\left[O_2\right] - k_3\left[RO_2^{\bullet}\right]\left[RH\right] - k_7\left[RO_2^{\bullet}\right]\left[AH\right] = 0 \qquad \text{(xviii)}$$

$$\therefore\ r_i = k_7\left[RO_2^{\bullet}\right]\left[AH\right] \qquad \text{(xix)}$$

and hence

$$\frac{-d\left[O_2\right]}{dt} = k_3 k_7^{-1} r_i\left[AH\right]^{-1}\left[RH\right] \qquad \text{(xx)}$$

Similar treatment of the situation where termination is exclusively between alkyl radicals (reaction iv) or between

an alkyl radical and an inhibitor (reaction viii) leads to Equations (xxi) and (xxii) respectively:

$$-\frac{d[O_2]}{dt} = k_2 k_4^{-\frac{1}{2}} r_i^{\frac{1}{2}} [O_2] \qquad \text{(xxi)}$$

$$-\frac{d[O_2]}{dt} = k_2 k_8^{-1} r_i [AH]^{-1} [O_2] \qquad \text{(xxii)}$$

These last two equations apply at low oxygen pressures.

The assumptions made in deriving Equation (xvii) were:

(a) That chains are long so that oxygen consumed in the initiation step (i) can be ignored,

(b) that the oxygen pressure is high enough to ensure that termination occurs exclusively between alkylperoxy radicals (Equation vi),

(c) that a standing concentration of alkyl and alkylperoxy radicals is rapidly attained.

If the autoxidation of organoboron compounds conforms to these conditions, an equation identical in form to (xvii) also governs this process.

$$-\frac{d[O_2]}{dt} = k_3 k_6^{-\frac{1}{2}} r_i^{\frac{1}{2}} [BR_3] \qquad \text{(xxiii)}$$