Mechanistic studies of the interactions of hexamine and some derivatives with electrophiles

John Keith Scranage

MECHANISTIC STUDIES OF THE INTERACTIONS OF HEXAMINE AND SOME DERIVATIVES WITH ELECTROPHILES   by   John Keith Scranage. Submitted for the degree of Doctor of Philosophy, 1989.

ABSTRACT

Mechanistic studies of the nitration of hexamethylene-tetramine (1) and some derivatives are reported and are compared with acetylation reactions. Nitration reactions, with nitric acid, were carried out using mixtures of $[^{15}N_4]$- and $[^{14}N_4]$-compounds, and the destination of the nitrogen isotopes in the products was determined mass spectro-metrically. The results show that in nitration of (1) to give 3,7-dinitro-1,3,5,7-tetraazabicyclo[3.3.1]nonane (DPT) extensive ring cleavage occurs to give species containing amino-nitrogen fragments. However, the nitration of 3,7-diacetyl-1,3,5,7-tetraazabicyclo[3.3.1]nonane (DAPT) to 1,5-diacetyl-3,7-dinitro-1,3,5,7-tetraazabicyclo[3.3.1]-nonane (DADN) involves selective cleavage of the methylene bridge.

Kinetic studies are reported of the pH-dependence of the decomposition of DPT in aqueous media. The results show that at all acidities two stages are observed. The first stage, k_1, involves catalysis by protons and hydroxide ions, and it is suggested that reaction occurs via a low concentration, ring-opened structure which is in equilibrium with DPT. In acidic solution an intermediate is observed which is identified as nitramide (NH_2NO_2). Methylenedinitroamine is shown not to be an intermediate in the decomposition of DPT, since its rate of decomposition is too slow.

Mechanistic studies were also carried out on the synthesis and decomposition of 3,7-dinitroso-1,3,5,7-tetraazabicyclo[3.3.1]nonane (DNPT). Experiments using mixtures of $[^{15}N_4]$- and $[^{14}N_4]$-compounds showed that relatively little isotopic mixing occurs during nitrosation of (1) to DNPT, and hence the predominant reaction involves cleavage of a methylene bridge with the ring structure remaining intact. Data obtained for the decomposition of DNPT in acid solutions suggested that the rate-determining step of this process involves C—N bond cleavage, and were consistent with the intermediate formation of nitrosamine (NH_2NO).

Kinetic studies were conducted on the decomposition of 3,7-bis(arylazo)-1,3,5,7-tetraazabicyclo[3.3.1]nonanes in acid solution to the corresponding aniline. Evidence was found for the formation of an intermediate methyleniminium ion, the UV/visible spectra of which could be observed in acetonitrile solutions containing only minute quantities of water.

^{1}H NMR studies of solutions of hexamine in DMSO-d_6 containing electrophilic species suggested that under these conditions quaternary salts of hexamine were formed, rather than cleavage of the molecule to form N-substituted six- and eight-membered ring species.

MECHANISTIC STUDIES OF
THE INTERACTIONS OF HEXAMINE
AND SOME DERIVATIVES
WITH ELECTROPHILES

by

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CONTENTS

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'^{15}N Studies of the Mechanisms of Nitration of Hexamethylenetetramine and 3,7-Diacetyl-1,3,5,7-tetraazabicyclo[3.3.1]nonane', Tetrahedron, 1988, 44, 1679.

'Kinetic Studies of the pH Dependence of the Decomposition of 3,7-Dinitro-1,3,5,7-tetraazabicyclo[3.3.1]nonane (DPT) and Related Compounds', J.Chem.Soc. Perkin Trans.II, 1989, 77.

'Mechanistic Studies on the Synthesis and Decomposition of 3,7-Dinitroso-1,3,5,7-tetraazabicyclo[3.3.1]nonane', J.Chem.Res. (S), 1989, 72.

DECLARATION

The material in this thesis is the result of research carried out in the Department of Chemistry, University of Durham, between October 1986 and October 1989. It has not been submitted for any other degree, and is the author's own work, except where acknowledged by reference.

STATEMENT OF COPYRIGHT

For my Father and Mother,
who made it all possible.

CHAPTER 1

Introduction

1.1 General properties of hexamine

Hexamine, $C_6H_{12}N_4$, has structure (1.1), and is readily formed by condensation of ammonia and formaldehyde[1,2] as shown in Equation (1.1).

(1.1) hexamine

$$6 \ CH_2O \ + \ 4 \ NH_3 \ \longrightarrow \quad \text{[structure]} \qquad \text{Eqn. (1.1)}$$
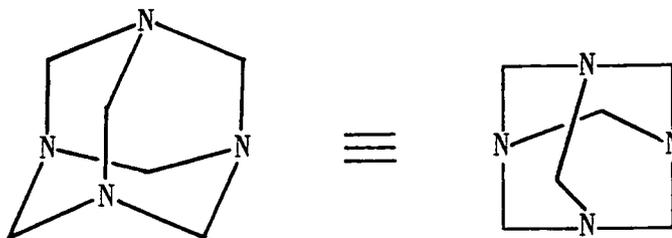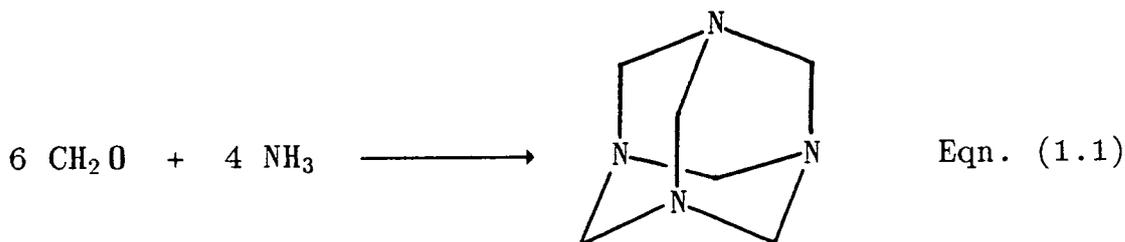
It is also known as hexamethylenetetramine (a descriptively more accurate trivial name), 1,3,5,7-tetraazatricyclo[3.3.1.1$^{3.7}$]decane, 1,3,5,7-tetraazaadamantane, methanamine, aminoform, formin, urotropine, and food additive E239[3]. Pure hexamine is a colourless, odourless, crystalline solid with a sweet taste[1] although commercial samples invariably have a mild 'fish-like' smell due to the occurrence of slight decomposition. It crystallises in regular rhombic dodecahedrons[4] which are stated to show piezoelectric properties[5]. Hexamine is soluble in water, somewhat soluble in alcohols, slightly soluble in ether and aromatic hydrocarbons, and, with the exception of

chloroform, in which it is fairly soluble, it is only slightly soluble in chlorinated aliphatics[1,2].

Hexamine is commercially employed as a special form of anhydrous formaldehyde[1] in the manufacture of synthetic resins, and in the hardening of proteins. It reacts as formaldehyde only in the presence of catalysts, when heated, or when brought into contact with an active formaldehyde acceptor, and thus, in many cases, its reactions are more readily controlled than those of formaldehyde itself[1]. A small quantity is used for medicinal purposes as a urinary antiseptic. As food additive E239, hexamine is used as a preservative for marinated herring and mackerel[3].

Hexamine was the first organic compound to be subjected to X-ray crystal structure analysis, by Dickinson and Raymond[6] in 1923. Previously this analytical method was only applied to elements and polar inorganic compounds. Dickinson chose to study hexamine since it is one of the few known organic compounds without salt characteristics which have cubic symmetry.

In addition to being an ammono-formaldehyde, hexamine is also a tertiary amine, and as such forms salts, addition compounds, and complexes. In this respect it resembles pyridine, triethylamine, and other tertiary bases. Hexamine was found to have a pK_a value of 4.89[7], which is similar to that of pyridine at 5.25[8], but is much lower than that of triethylamine at 11.04[8]. The similarity with pyridine is attributed to the electron-withdrawing effect of neighbouring N-atoms on any particular N-atom, in much the same way as the aromatic ring withdraws electron charge from the N-atom of pyridine. No such effect exists for

triethylamine, and so the higher basicity is reflected in a higher pK_a value.

Hexamine sublimes in a vacuum at 230-270°C.[9] This high temperature is clearly associated with the considerable symmetry and rigidity of the cage structure. The corresponding all carbon structure, adamantane, also has a high sublimation point (268°C[8]). Various physico-chemical methods[10,11] have been employed to demonstrate the chemical and steric equivalence of the four nitrogen atoms. On protonation of one of the nitrogen atoms the molecule loses its symmetry, and various acid-catalysed fragmentation processes may then occur.[2] Depending on the conditions employed, two, three or more carbon-nitrogen sub-units can be formed, or the reagent can serve as a source of formaldehyde and ammonia. At ordinary temperatures a pure water solution of hexamine is comparitively stable, showing only very slight hydrolysis to formaldehyde and ammonia,[1] and the degree of hydrolysis at elevated temperatures in a neutral aqueous solution remains minimal.[12]

Although hexamine does not normally occur in a hydrated form, it is possible[1] to form a hexahydrate, $C_6H_{12}N_4 \cdot 6H_2O$, when a saturated aqueous solution is cooled to slightly above 0°C.

1.2 Nitration of hexamine

Hexamine is of military interest, since nitration was found to yield the high explosive cyclotrimethylene-trinitramine,[13] also called cyclonite, hexogen, and more recently designated as RDX (research department explosive). It has structure (1.2).
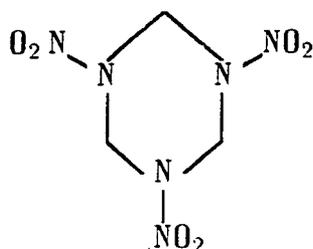
(1.2) RDX

RDX is a white crystalline substance of melting point 202-207°C,[14] and was first prepared by Henning[15] in 1879 by the action of nitric acid on hexamine dinitrate. In 1920 Herz[16] clarified its chemical structure as that shown above. He also recognised its nature as an explosive. In 1925 Hale[13] reported an improved method of preparation of RDX involving gradual addition of hexamine to an excess of 99.8% nitric acid at about 20-30°C. Hale represented the reaction by the following equation:

$$C_6H_{12}N_4 + 4\ HNO_3 \longrightarrow RDX + 3\ CH_2O + NH_4NO_3$$
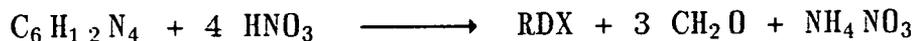
Eqn. (1.2)

although Schnurr[18] suggested that the reaction proceeded thus:

$$C_6H_{12}N_4 + 6\ HNO_3 \longrightarrow RDX + 6\ H_2O + 3\ CO_2 + 2\ N_2$$

Eqn. (1.3)

Since ammonium nitrate, formaldehyde, carbon dioxide, nitrogen and water can all be detected in the products,[18] it is reasonable to suggest that the reaction can occur via both pathways. Later it will be seen to be significant that some of the methylene groups and nitrogen atoms of hexamine are not utilised for the production of RDX. Indeed, the nitration of hexamine to RDX requires from four to eight times the theoretical amount of nitric acid.

The Hale process was further improved by Bachmann and Sheehan[19] who used a nitrating mixture of nitric acid, acetic anhydride and ammonium nitrate:

$$C_6H_{12}N_4 + 4\ HNO_3 + 2\ NH_4NO_3 + 6\ Ac_2O \longrightarrow 2\ RDX + 12\ AcOH$$

Eqn. (1.4)
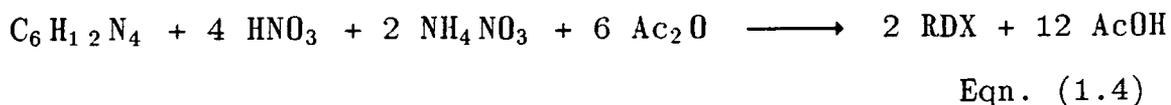
RDX was used extensively in World War II after much successful research into the development of a practical method of manufacture. This led to a spate of publications in the years following the war[20-29] detailing the chemistry of the process.

A by-product of the production of RDX, studied by Wright and co-workers[20] and also Bachmann and Sheehan,[19] is 1,3,5,7-tetranitro-1,3,5,7-tetraazacyclooctane, also known as octogen, and more commonly HMX (high melting explosive).

This compound, with structure (<u>1.3</u>), has a melting point of 276-277°C.[30]

$$\begin{array}{c}
NO_2 \\
| \\
-N- \\
\end{array}$$

O₂N—N   N—NO₂

—N—
|
NO₂
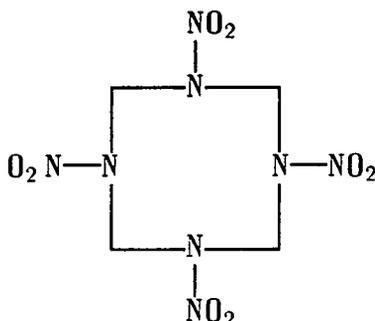
(<u>1.3</u>) HMX

HMX is a white crystalline substance which occurs in four polymorphous forms, three of which are stable at room temperature, and one which transforms very rapidly. Like RDX, HMX is insoluble in water and non-hygroscopic, and also has a similar chemical reactivity. They differ only in that HMX is more resistant to attack by sodium hydroxide solution, this reaction forming the basis of one method of separating HMX and RDX. The ratio of such a mixture, formed during nitration of hexamine, is dependent upon the reaction conditions employed. HMX is a powerful high explosive, being superior to RDX in having a higher ignition temperature, and being chemically more stable.[31] However, due to the high cost of manufacture (approximately three to four times that of RDX) it finds use only in specialised ordnance, such as shaped charges, where maximum explosive performance is needed.[32]

HMX can be prepared by nitration of 3,7-dinitro-1,3,5,7-tetraazabicyclo[3.3.1]nonane, also known as 3,7-dinitro-pentamethylenetetramine, abbreviated to DPT (<u>1.5</u>), which

Wright and co-workers[20] synthesised by the reaction of methylenediamine with dimethylol nitramide (**1.4**) :

$$O_2N{-}N \begin{matrix} CH_2OH \\ | \\ | \\ CH_2OH \end{matrix} \quad + \quad CH_2 \begin{matrix} NH_2 \\ | \\ | \\ NH_2 \end{matrix} \quad + \quad N{-}NO_2 \begin{matrix} HOCH_2 \\ | \\ | \\ HOCH_2 \end{matrix}$$

(**1.4**) dimethylol
       nitramide

$$\downarrow$$

$$O_2N{-}N \underset{\boxed{}}{\overset{{-}N{-}}{\underset{{-}N{-}}{CH_2}}} N{-}NO_2 \quad + \quad 4\ H_2O$$

(**1.5**) DPT
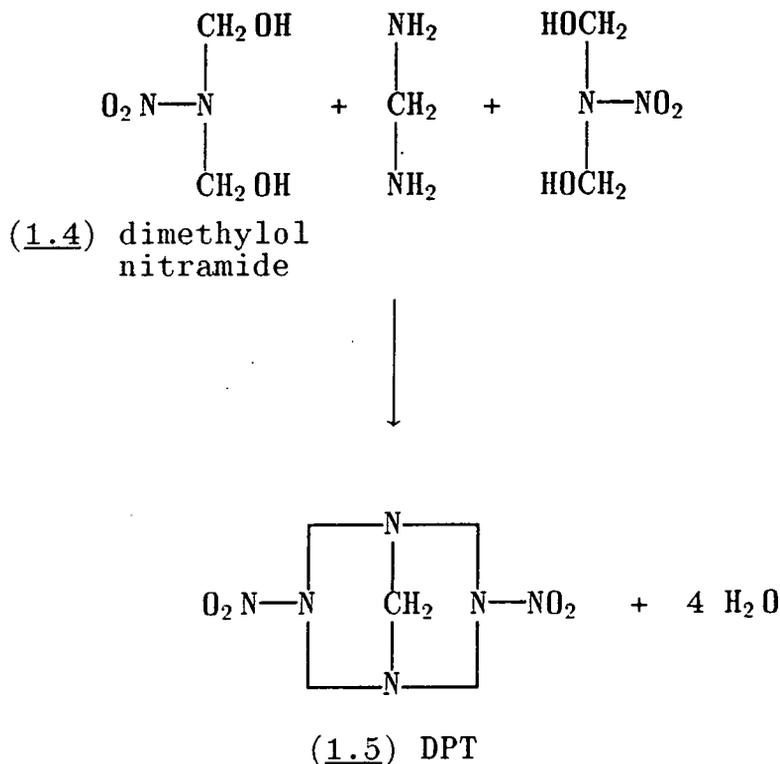
Scheme 1.1

DPT can also be obtained from the preparation of RDX by removing the RDX and neutralising the mother liquor to effect precipitation of DPT,[13,20,33,34] this being formed from the fragments not utilised for the formation of RDX. It has been suggested[20] that dimethylol nitramide is the direct precursor to DPT in this process, but work by the author, described in Chapter 4, shows that this is not necessarily the case.

Mechanisms postulated[19,24,36-38] for synthesis of RDX and HMX from hexamine include the selective cleavage of bonds within the hexamine molecule, or total cleavage of the molecule to smaller fragments, followed by nitration and

recombination. Work done by Castorina and co-workers[39,40] involving ^{14}C and ^{15}N tracer studies provides evidence for the latter mechanism. The ^{14}C tracer experiments showed that nitration to RDX and HMX involved total non-selective breakdown of the hexamine molecule to fragments containing chemically equivalent methylene groups. Addition of paraformaldehyde to the nitration mixture showed that the final products contained methylene groups from a common source comprised of those from hexamine itself, and from paraformaldehyde. ^{14}C exchange did not occur between unreacted hexamine and paraformaldehyde during the nitration process. Work done using $^{15}NH_4NO_3$ in the nitrating medium[36,39] demonstrated the possibility of exchange between the amino nitrogen of the salt with hexamine during nitration, and also the possibility of exchange between the amino nitrogen of the salt with the trimethylene substituted nitrogens in DPT prior to nitration to HMX. The tracer studies also showed that the formation of DPT from hexamine involved extensive decomposition of the hexamine molecule rather than simple, selective cleavage of the methylene bridge. The formation of DPT as an isolatable intermediate of the nitration of hexamine was re-investigated as described in Chapter 3.

1.3 Acetylation of hexamine

It has been shown[24,41-46] to be easy to acetylate
hexamine to form the eight-membered 3,7-diacetyl-1,3,5,7-
tetraazabicyclo[3.3.1]nonane (diacetylpentamethylene-
tetramine, DAPT) (1.6), or the six-membered 1,3,5-triacetyl-
1,3,5-triazacyclohexane (triacetyltrimethylenetriamine,
TRAT) (1.7).

$(CH_3CO)_2O$
5-10°C

(1.6) DAPT

$(CH_3CO)_2O$
90-100°C

(1.7) TRAT
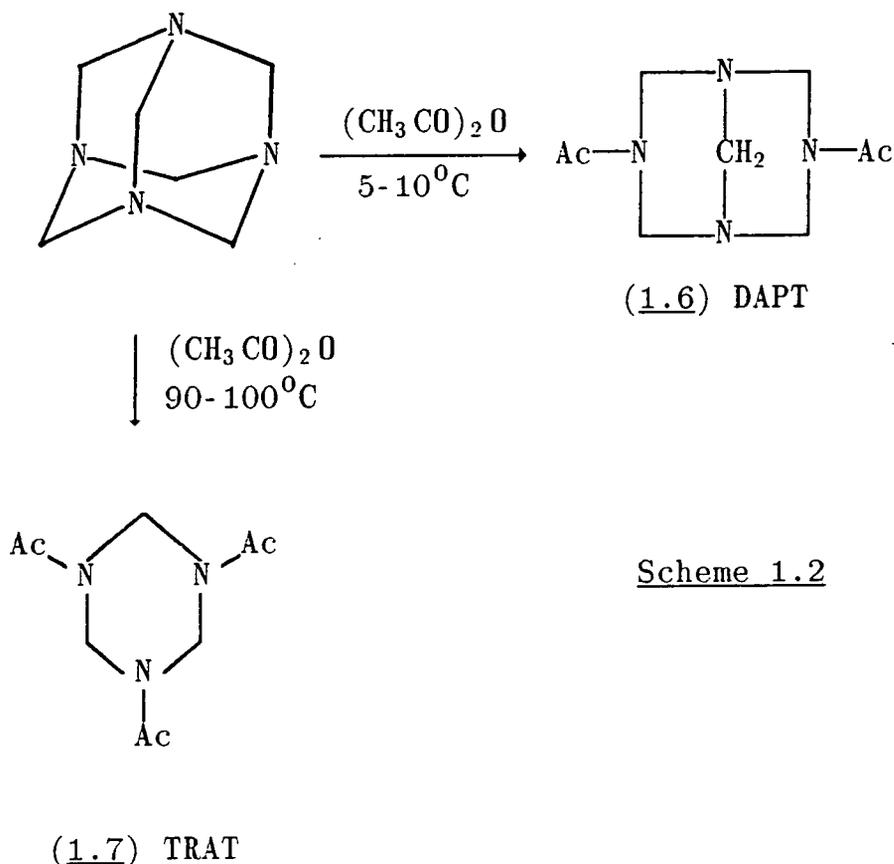
Scheme 1.2

DAPT is the main product at low temperature, TRAT being
formed only at elevated temperatures.

Interest in DAPT has been aroused recently[47,48] due to
its potential usefulness in a synthetic route to the eight-
membered HMX exclusively from hexamine. It has been shown[42]
that the acetylation of hexamine to DAPT proceeds via
selective cleavage of a methylene bridge, with retention of

- 10 -

the eight-membered structure, to an exclusively eight-membered product.

Further acetylation of DAPT yields 1,3,5,7-tetraacetyl-1,3,5,7-tetraazacyclooctane (TAT)[47,49] (1.8).

Ac
|
N
Ac—N   N—Ac
|
N
|
Ac
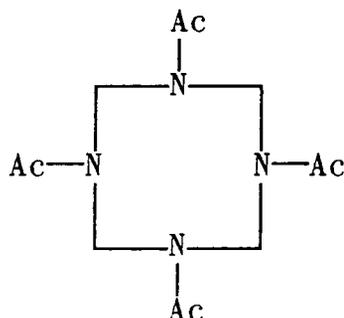
(1.8) TAT

TAT can be converted to HMX by heating with a mixture of nitric acid and phosphorus pentoxide[47].

Nitration of DAPT using a mixture of nitric and sulphuric acid yields 1,5-diacetyl-3,7-dinitro-1,3,5,7-tetraaza-cyclooctane[47,48] (DADN) (1.9), though DADN can be formed in a 'one-pot' synthesis without isolation of DAPT[32,47].

NO_2
|
N
Ac—N   N—Ac
|
N
|
NO_2
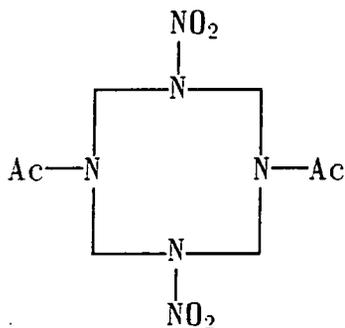
(1.9) DADN

A potentially useful method of nitrating DADN to HMX has been developed[47] which involves the use of a mixture of nitric acid and dinitrogen pentoxide (N_2O_5). N_2O_5 is thought to be the nitrating agent in nitric acid/phosphorus pentoxide mixtures, which were initially used.

It has also been shown[47,48] to be possible to convert DAPT to 1,5-diacetyl-3-nitro-7-nitroso-1,3,5,7-tetraaza-cyclooctane (DANNO) (1.10), using a mixture of nitric acid and N_2O_4 (this mixture being known as red fuming nitric acid).
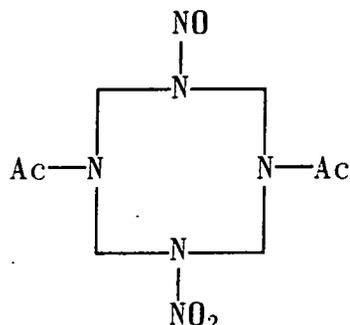
(1.10) DANNO

It was thought that it might be possible to convert DANNO directly to HMX, but initial attempts were unsuccessful, giving much lower yields than the route via DADN. It has been shown to be possible to convert DANNO to DADN in 95% yield.

1.4 Nitrosation of hexamine

Degradative nitrosation of hexamine in aqueous solution occurs by the simultaneous addition of hydrochloric acid or acetic acid and a solution of sodium nitrite.[50]

The pH of the solution determines the nature of the products. At pH 1-2, trimethylenetrinitrosamine (TMTN) (1.11) is formed, whereas at pH 3-6, dinitrosopenta-methylenetetramine (DNPT) (1.12) is formed.
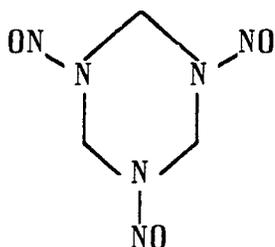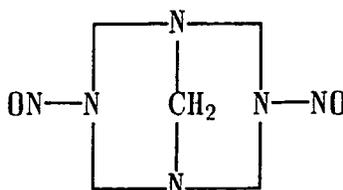
TMTN (1.11)   DNPT (1.12)

TMTN was first described by Mayer[51] in 1881, who suggested that the structure was that given as (1.11). Duden and Scharff,[52] Bachmann and Deno,[50] and Auberstein[53] gave detailed descriptions of its preparation and chemical properties. Finally, Fischeroulle and Kovache[54] worked out the mechanism of its production on a semi-commercial scale. TMTN decomposes explosively under the action of concentrated sulphuric acid at room temperature. Oxidation of (1.11) yields RDX (1.2). According to Brockmann, Downing and Wright[55] oxidation with a solution of hydrogen peroxide (30%) in nitric acid (99%) in the ratio of 1 mole of TMTN to 82 moles of nitric acid, 3 moles of hydrogen peroxide and 3.7 moles of water at -40°C gives the dinitro-nitrosamine (1.13) as an intermediate.

$$\text{(\underline{1.13})} \qquad\qquad \text{(\underline{1.2}) RDX}$$
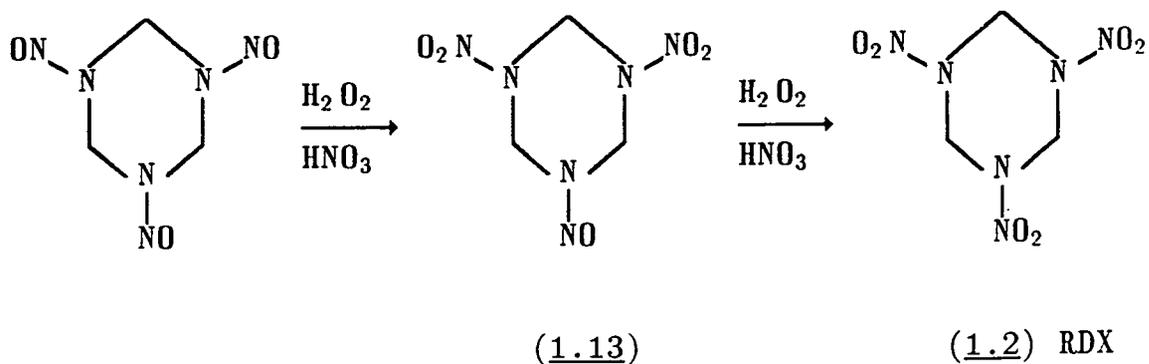
Scheme 1.3

TMTN is a powerful explosive with a sensitiveness to impact of the same order as that of trinitrotoluene.[56]

DNPT was first obtained by Griess and Harrow.[57] Formulae for both TMTN and DNPT were proposed by Cambier and Brochet,[58] and Duden and Scharff.[52] DNPT is used as a gasifiable product for the production of porous plastics and rubber[59] (described as a 'blowing agent').

1.5 Reaction of diazonium ions with hexamine

The 3,7-bis(arylazo)-1,3,5,7-tetraazabicyclo[3.3.1]-nonanes (1.14) are formed by diazonium coupling with hexamine.[60,61]

$$Ar-N=N-N \underset{N}{\overset{N}{\begin{array}{c} CH_2 \end{array}}} N-N=NAr$$
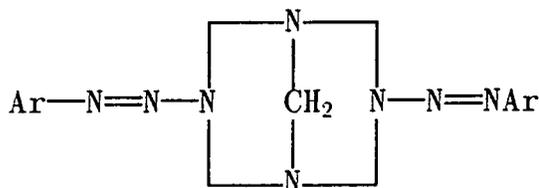
(1.14)

In recent years several new examples of this novel class of bicycloheterocycle have been reported,[62] arising out of their preferential formation during attempted diazonium coupling with ammonia-formaldehyde mixtures (from which hexamine can be formed) to yield compounds of type (1.15). Triazenes such as (1.15) are thought to be potential anti-tumour agents, structures of type (1.16) having already been shown to be significantly active against mouse tumour models in vivo.[62]

$$ArN=N-N \overset{H}{\underset{CH_2 OH}{<}} \qquad ArN=N-N \overset{Me}{\underset{CH_2 OH}{<}}$$
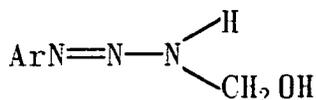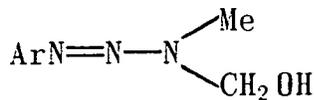
(1.15)   (1.16)

A proposed mechanism[62] of formation of bis(arylazo)tetra-azabicyclononanes by diazonium coupling with hexamine is shown in Scheme 1.4.

(1.1)

(1.17)

(1.18)

(1.19)

(1.14)
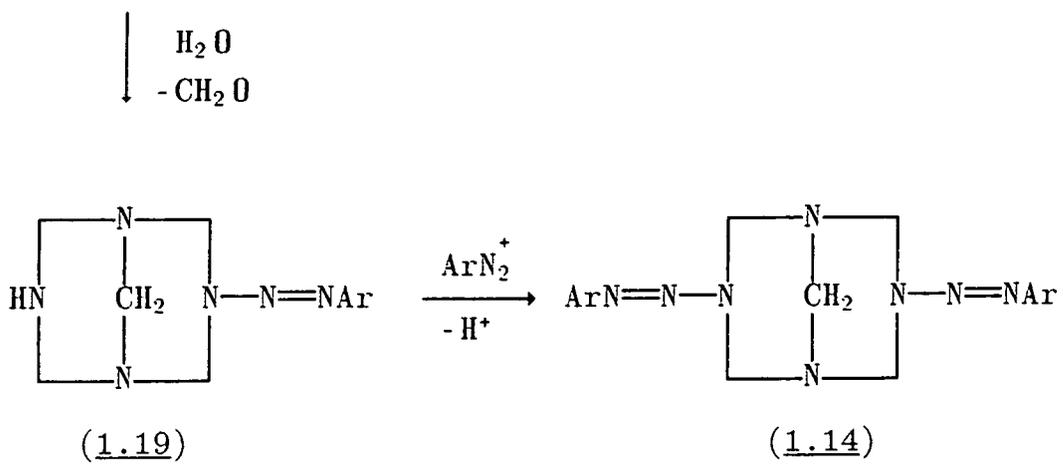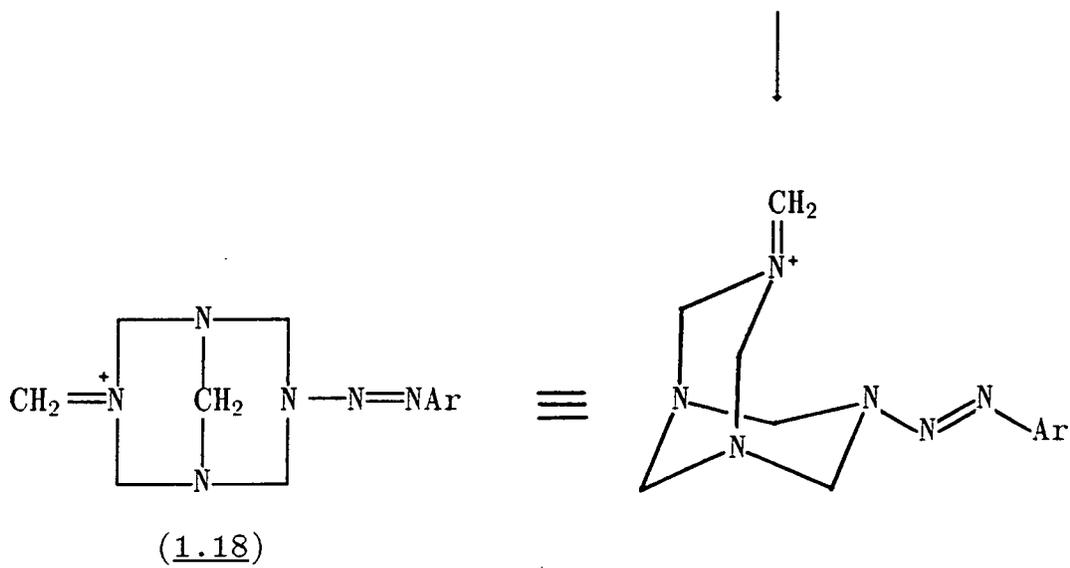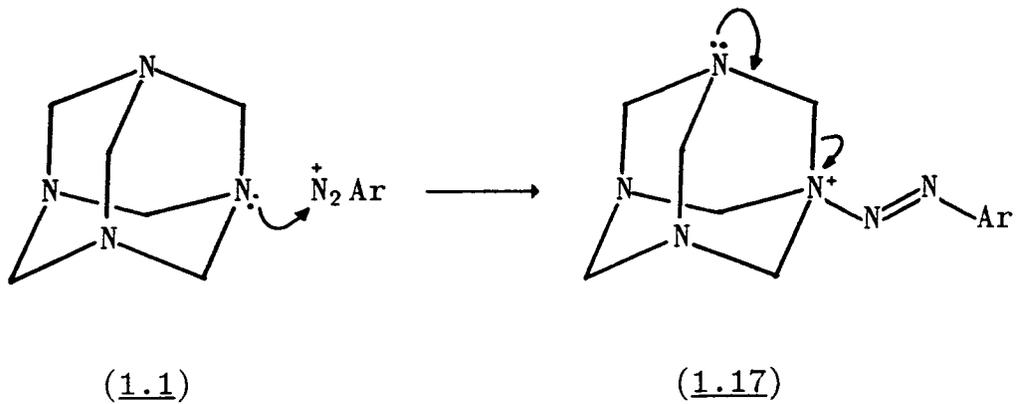
Scheme 1.4

Attachment of the electrophilic diazonium ion at one of the four equivalent nitrogen atoms of hexamine (<u>1.1</u>) gives the triazenium ion (<u>1.17</u>) and initiates ring cleavage, typical of an aminal,[63,64] to give the iminium ion (<u>1.18</u>). Hydrolysis of (<u>1.18</u>) and loss of formaldehyde[65,66] leads to the mono(arylazo)tetraazabicyclononane (<u>1.19</u>), which undergoes further diazonium coupling at the secondary amine position to give the product (<u>1.14</u>). This mechanism assumes selective cleavage of the hexamine molecule during diazotisation, and, although not proven, this is consistent with observations made by the author, to be discussed in detail in Chapter 3.

1.6 Salt and complex formation

1.6.1 Salts with acids

When hexamine is reacted with dilute organic and inorganic acids, salts are the primary product formed, and may be isolated in many circumstances.[1,2] In its interactions with mineral acids, hexamine generally behaves as a monobasic compound, forming salts with stoichiometry $C_6H_{12}N_4 \cdot HX$ which are best isolated by preparation in non-aqueous solvents (e.g. ethanol or chloroform) or, in some cases, cold aqueous solution. The monohydrochloride of hexamine, $C_6H_{12}N_4 \cdot HCl$, can be prepared by addition of aqueous hydrochloric acid to an alcoholic solution of hexamine, or by the action of hydrogen chloride gas on a hot solution of hexamine in absolute alcohol.[67] It is possible to form a compound of composition $C_6H_{12}N_4 \cdot 2HCl$ by the action of excess hydrogen chloride.[68] X-ray studies[69] on the mono-hydrochloride have shown the structure to be that of (1.20).
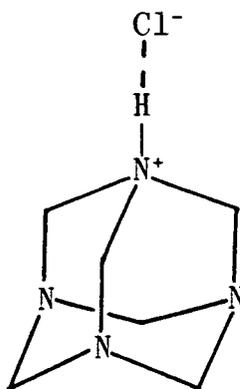
(1.20)

The molecule possess full 3m symmetry, and is isostructural with $C_6H_{12}N_4 \cdot HBr$, the monohydrobromide salt of

hexamine. The monohydroiodide salt is also known. With hydrogen fluoride, complexes containing 1-4 molecules of HF per molecule of hexamine are formed.[70]

The sulphate salt of hexamine, $(C_6H_{12}N_4)_2 \cdot H_2SO_4$, is precipitated by the action of sulphuric acid in cold alcoholic solution.[71-73] Other salts of hexamine with inorganic acids that have been reported are the phosphates, $C_6H_{12}N_4 \cdot H_3PO_4$ and $5C_6H_{12}N_4 \cdot 6H_3PO_4 \cdot 10H_2O$, the perchlorate, $C_6H_{12}N_4 \cdot HClO_4$, and the exlosive chromates, $2C_6H_{12}N_4 \cdot H_2Cr_2O_7$ and $2C_6H_{12}N_4 \cdot H_2Cr_4O_{13}$.

It is possible to precipitate a mononitrate salt, $C_6H_{12}N_4 \cdot HNO_3$, by the action of cold, dilute nitric acid on an aqueous solution of hexamine at 0^oC.[71-73] A dinitrate salt, $C_6H_{12}N_4 \cdot 2HNO_3$, can be formed with more concentrated acid.[13,19] As mentioned earlier, the first route to RDX was via nitration of hexamine dinitrate. Both the mono- and dinitrate salt have been studied by NMR.[74,75]

The primary reaction of hexamine with organic acids is salt formation. In general, these salts may be isolated by combining base and acid in the theoretical proportions in concentrated aqueous solution, and subjecting the product to vacuum evaporation, although it has been shown[2] that even in the presence of a large excess of acid less than four molecules of acid co-ordinate with hexamine. Heating should be avoided when preparing the salts of stronger acids,[1] since hydrolytic reactions may take place. Hexamine formate may be prepared by the action of the pure acid on hexamine, or by employing a slight excess of aqueous (e.g. 50%) acid, and removing water and excess acid under reduced pressure. If heated with formic acid, hexamine is reduced with evolution

of carbon dioxide and formation of methylamines.

It has been mentioned earlier that reaction of hexamine with dilute nitrous acid leads not to salt formation, but cleavage of the hexamine molecule and formation of cyclic nitrosamines.

1.6.2 Quaternary salt formation

Alkyl halides react with a solution of hexamine in chloroform to give the quaternary salts of structure (1.21). The starting materials are soluble, whereas the products are not and precipitate out.

(1.1)   (1.21)

Eqn. (1.5)

For instance, methyl iodide adds to hexamine in an absolute alcohol solution to give lustrous, needle-like crystals of the product, $C_6H_{12}N_4 \cdot CH_3I$.[76] The N-methyl-hexamine mono- and dinitrate, and N-dimethylhexamine dinitrate have been prepared[77] and studied by NMR techniques.[74,75]

α-halogenated tertiary amines react with hexamine to give quaternary salts of type (1.22).

$$(\underline{1.1}) \quad + \quad X-CH_2NR_2 \quad \xrightarrow[CH_2Cl_2/N_2]{dry} \quad$$

(1.22)

Eqn. (1.6)

In dilute aqueous acid the quaternary salts decompose to give the secondary amine, formaldehyde, and the ammonium salt of the acid.

$$\xrightarrow{H^+/H_2O} \quad NHR_2 \quad + \quad 4\ NH_4X \quad + \quad 7\ CH_2O$$

Eqn. (1.7)

(1.22)
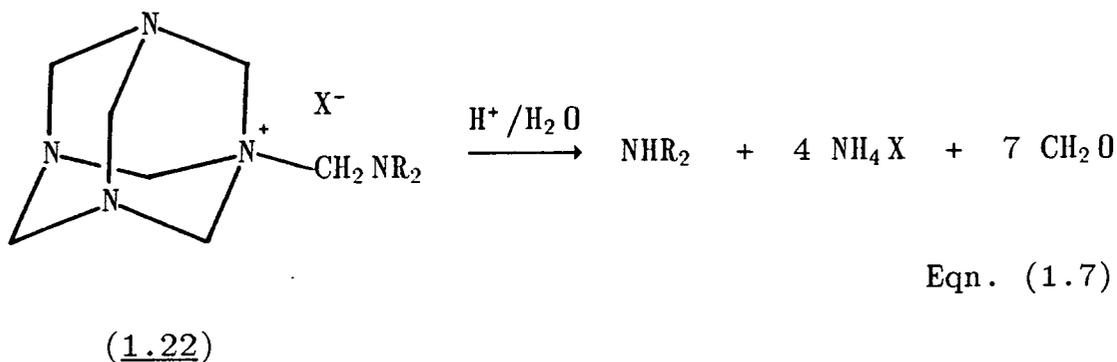
Quaternary salts of the type (1.23), prepared by reaction of hexamine with haloacetates or haloacetonitriles in tetrachloromethane solution, and reaction of hexamine with haloalkylnitriles in chloroform, have been found to have bactericidal and fungicidal activity.
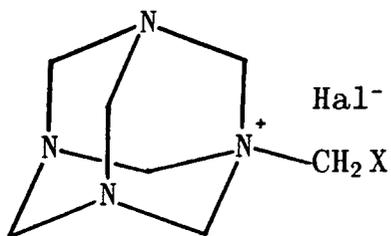
$X = COOCH_3,\ COOC_2H_5,$
$\qquad COOCH_2-C\equiv CH,\ CH_2-CBr=CH_2,$
$\qquad CH_2-CH=CH_2,\ CH_2-CCl=CH_2,\ CN$

$Hal = Cl,\ Br,\ I$

(<u>1.23</u>)

There is evidence[78-80] to suggest that reaction of hexamine with acid chlorides, such as acetyl and benzoyl chloride, in chloroform yields the N-acylhexaminium salts of structure (<u>1.24</u>).
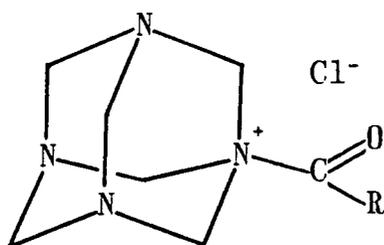
e.g. R = Me, Ph

(<u>1.24</u>)

The N-acetylium salt and similar structures formed by the interaction of hexamine with other electrophilic species are investigated by ^1H NMR in Chapter 7.

1.6.3 Complex formation with phenols

The action of hexamine on phenols leads primarily to the formation of hydrogen-bonded complexes, a large number having been reported in which 1,2 and 3 molecules of the phenol are combined with 1 molecule of hexamine.[81,82] In the case of phenol itself, hexamine triphenol, $C_6H_{12}N_4 \cdot 3C_6H_5OH$, is formed as a crystalline precipitate when concentrated aqueous solutions of hexamine and phenol are mixed at room temperature. A product of composition $C_6H_{12}N_4 \cdot C_6H_5OH$ has also been reported.[83] Hexamine forms 1:1 complexes with 1,3-dihydroxybenzene and 1,3-dihydroxy-5-methylbenzene, and a 1:2 complex with 1,3-dihydroxy-2,5-dimethylbenzene.

1.6.4 Complex formation with inorganic salts

Hexamine forms complexes with many inorganic salts, including those of alkali metals, alkaline earths, rare earths and transition metals. These compounds tend to conform to the type formula $MX_n \cdot nC_6H_{12}N_4$, where M is a metal ion of valence n, but the number of molecules of combined hexamine per molecule of salt is often lower, with several different complexes being formed with the same salt.[1] The salts of lithium, sodium, potassium, copper, silver, gold, magnesium, calcium, strontium, barium, zinc, cadmium, mercury, aluminium, titanium, lanthanum, cerium, neodymium, yttrium, erbium, thorium, tin, antimony, bismuth, chromium, molybdenum, tungsten, uranium, manganese, iron, cobalt, nickel, platinum and palladium are all reported to form complexes with hexamine. For example, when a moderately concentrated solution of hexamine in water is added to aqueous silver nitrate, the white crystalline compound

$AgNO_3 \cdot C_6H_{12}N_4$ is precipitated[84]

1.7 Reactions with halogens and inorganic halides

Hexamine in aqueous solution is decomposed by chlorine with formation of the explosive nitrogen trichloride[85] Reaction with sodium hypochlorite yields chloro- derivatives of hexamine, which are reported to be unstable and prone to explode on storage. Delepine reports[86] the formation of N-dichloropentamethylenetetramine (3,7-dichloro-1,3,5,7-tetraazabicyclo[3.3.1]nonane) (1.25) as thin leaflets on addition of a dilute sodium hypochlorite solution to a solution of hexamine.
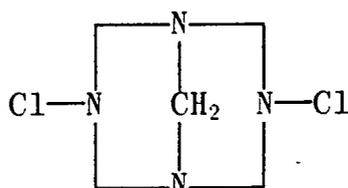
(1.25)

This compound explodes on heating to 78-82°C. Formation of a dichlorohexamine has been described by Leulier[87] and a tetrachloro- derivative by Buratti[88]

When bromine reacts with hexamine in chloroform solution a crystalline orange-red tetrabromide, $C_6H_{12}N_4Br_4$, is formed[89,90] On standing in air bromine is lost to form the yellow dibromide, $C_6H_{12}N_4Br_2$, which can also be prepared by the action of bromine water on aqueous hexamine solutions[17]

Reaction of hexamine with iodine yields the di- and tetraiodo- derivatives. The diiodide, $C_6H_{12}N_4I_2$, is

described[85] as red-brown needles or a yellow-green
crystalline powder, the tetraiodide occurring as dark brown
crystals or an orange-red powder. The production of mixed
halides, such as $C_6H_{12}N_4ICl$ has also been reported.[85]

Inorganic halides, such as phosgene, form addition
compounds with hexamine,[1] the phosgene derivative having
composition $2C_6H_{12}N_4 \cdot COCl_2$. Similarly, with sulphur
chloride, hexamine forms a substance of formula
$2C_6H_{12}N_4 \cdot S_2Cl_2$. This type of compound may also be formed
with PCl_3, $POCl_3$, and SO_2Cl_2.

A complex formed with silicon tetrafluoride has been
reported,[35] in which one SiF_4 unit is co-ordinated to two
hexamine molecules.

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CHAPTER 2

Experimental

2.1 Materials

2.1.1 Solvents

Water: Distilled water was used. Where necessary it was boiled for 20 minutes to expel dissolved carbon dioxide, and a soda-lime guard tube was used for protection from the atmosphere.

Acetone: General purpose reagent was used for washing apparatus. For synthetic work analytical grade was used without further treatment.

Acetonitrile: HPLC grade, used without further treatment.

Acetic anhydride: Analytical reagent. Dried with molecular sieve.

Acetic acid: Glacial, analytical grade. Dried with molecular sieve

Acetyl chloride: Commercial grade, used as supplied.

Chloroform: Analytical grade, used without further treatment.

Diethyl ether: Commercial grade, used as supplied.

Methanol: Analytical reagent, used without further treatment.

Ethanol: Analytical reagent, used as supplied.

Petroleum spirit (40-60°C): Commercial grade, used as supplied.

Petroleum ether (30-40°C): Commercial grade, used as supplied.

Triethylamine: Commercial grade, used as supplied.

Methylamine: 40% aqueous solution, commercial grade, used without further treatment.

Ammonia: 32% aqueous solution, commercial grade, used as supplied.

Formaldehyde: 37% aqueous solution, analytical grade, used as supplied.

Hydrochloric acid: 37% aqueous solution, analytical grade, used as supplied.

Nitric acid: 95% fuming, commercial sample, used without further treatment.

Sulphuric acid: 98%, analytical reagent, used as supplied.

Perchloric acid: 60-62% aqueous solution, analytical grade, used as supplied.

Isopropyl nitrate: Commercial grade, used as supplied.

Deuterium oxide: 99.8% D-atom, commercial sample, used without further treatment.

Deuterium chloride: 99%+ D-atom 20% solution in D_2O, commercial grade, used as supplied.

Chloroform-d: 99.8% D-atom, commercial grade, used as supplied.

DMSO-d$_6$: 99.9% D-atom, commercial sample, used without further treatment.

2.1.2 Gases

Ammonia: Dry, used as supplied from commercial cylinder.

Nitrogen dioxide: Dry, used as supplied from commercial cylinder.

Oxygen: Commercial cylinder, dried by passing through P_2O_5 before use.

2.1.3 Salts

Sodium nitrite: Analytical grade, used as supplied.

Sodium hydroxide: Analytical grade pellets, used as supplied.

Potassium hydroxide: Analytical grade pellets, used as supplied.

Sodium carbonate: Analytical grade, used as supplied.

Sodium bicarbonate: Commercial grade, used as supplied.

Sodium acetate: Anhydrous, commercial grade, used as supplied.

Di-sodium tetraborate (borax): Commercial grade, used as supplied.

Citric acid: Analytical grade, used without further treatment.

Tris(hydroxymethyl)aminomethane: Commercial grade, used as supplied.

Phosphorus pentoxide: Commercial grade, used as supplied.

Ammonium nitrate: Commercial grade, used as supplied.

Nitronium tetrafluoroborate: Commercial grade, used as supplied.

Nitrosonium hydrogen sulphate: Commercial grade, used as supplied.

Barium chloride: Commercial grade, used as supplied.

2.1.4 Substrates

[14]N-hexamine: Commercial sample, used as supplied.

p-chloroaniline: Commercial grade, used as supplied.

p-ethoxycarbonylaniline: Commercial grade, used as supplied.

Urethane: Commercial grade, used as supplied.

Acetamide: Commercial sample, used without further treatment.

Paraformaldehyde: Commercial sample, used as supplied.

[15]N-hexamine: Gaseous $^{15}NH_3$ (99% isotopic abundance) obtained from Amersham International was reacted under reduced pressure with a 37% aqueous formaldehyde solution[1]. The solution was evaporated to dryness and the ^{15}N-hexamine was extracted into chloroform. After removal of chloroform, the crude product was recrystallised from acetone to yield a white solid, m.p. 232-233°C (sub.), (lit., various). The 1% abundance of $^{14}NH_3$ gave a product with an isotopic composition of ca. 96% [$^{15}N_4$]-hexamine and 4% [$^{15}N_3$][$^{14}N_1$]-hexamine.

DPT: DPT was prepared using a method[8] adapted from the Hale[2] process. This involves gradually adding hexamine (10 g), with mixing, to 95% fuming nitric acid (35 g) at 0-10°C. On dilution of the solution with iced water (100 cm^3), RDX separated out, was filtered off and destroyed in aqueous sulphuric acid. The filtrate was neutralised with ammonia solution at 0°C. DPT precipiated out as a white solid, which was washed with water and recrystallised from acetone, m.p. 212°C (lit[3], 213°C). The ^{1}H NMR spectrum[4,5]

at 298 K in acetonitrile-d_3 showed a peak at $\delta 4.15$ (s, CH_2 bridge) and an AB quartet (J = 8 Hz) due to methylene protons with shifts of $\delta 4.85$ and $\delta 5.59$ respectively.

DAPT: was prepared by the method of Siele, Warman and Gilbert.[6,7] This involves addition of acetic anhydride (10 g), with stirring at 10°C, to a slurry of hexamine (5 g) in water (2.5 cm^3). The mixture was stirred for 30 minutes and then evaporated to dryness. The crude DAPT obtained was recrystallised from acetone to yield a white crystalline solid, m.p. 192-194°C (lit[7], 192°C). The ^{1}H NMR spectrum in chloroform-d showed bands at $\delta 2.08$ (s, acetyl) and $\delta 4.14$ (s, CH_2 bridge), with two AB quartets (J = 13 Hz) due to methylene protons with shifts of $\delta 4.22$ and $\delta 5.72$, and $\delta 4.76$ and $\delta 4.9$.

DADN: was prepared by the method of Yoshida and co-workers.[3] This involves adding hexamine (1.1 g), with stirring at 25-30°C, to a mixture of 95% nitric acid (4 cm^3) and 98% sulphuric acid (12 cm^3) over a period of 15 minutes. The reaction mixture was stirred for 1 hour, after which time ice-water (200 cm^3) was added, and the solution made alkaline by addition of solid sodium carbonate. The precipitated DADN that formed was filtered off, washed thoroughly with water, and dried. Recrystallisation from nitromethane yielded white crystals, m.p. 265-266°C (lit[3], 265°C). The ^{1}H NMR spectrum[3] in DMSO-d_6 showed a band at $\delta 2.2$ (acetyl) and a band at $\delta 5.44$ (ring CH_2).

DNPT: DNPT was prepared by the method of Bachmann and Deno.[9] In this process an ice-cold solution of sodium nitrite (3 g) in water (7.2 cm^3) was added to an ice-cold solution of hexamine (1 g) in acetic acid (2.8 cm^3)/water (64 cm^3). The solution was stirred for 1 hour, after which time the precipitated DNPT was filtered off, washed with distilled water, and dried to give a pale yellow powder, m.p. 207^oC (lit.[9], $206-209^oC$). In DMSO-d_6 the product gave a ^{1}H NMR spectrum consistent with that of Stefaniak,[4] which appears complex due to DNPT existing as two isomeric forms where the nitroso groups are either *syn* or *anti* to each other.

MDNA: was prepared by the method of Brian and Lamberton.[10] Acetamide (59 g) and paraformaldehyde (16.5 g) were heated together in an oil-bath at 155^oC for 17 hours with, and for a further 90 minutes without, a reflux condenser. The melt, on recrystallisation from ethanol, yielded methylenebis-N-acetamide (16.5 g), m.p. $196-197^oC$ (lit.[10] 196^oC). 95% nitric acid (30 cm^3) was added at $0-10^oC$ over 2.5 hours to a vigorously stirred suspension of methylenebis-N-acetamide (10 g) in acetic anhydride (30 cm^3). After a further hour at 10^oC the solution was run into a vigorously stirred ice (125 g)/ water (150 cm^3) mixture over a period of 1 hour. The crude nitration product, methylenebis-N-(N-nitroacetamide), was collected and dried. 2.5 g of this product was added portionwise to well-stirred aqueous ammonia (4.2 cm^3) kept at $10-20^oC$. After a further 5 minutes, barium chloride solution (3.6 g of dihydrate in 10.5 cm^3 water) was run in and the suspension was cooled, over a 20 minute period, to 4^oC. The barium salt of MDNA

was collected, washed with cold 50% aqueous acetone, and added to hydrochloric acid (2.5 cm^3 37% in 6.5 cm^3 water). After the solution had been filtered, extracted with ether (7 x 5 cm^3 portions), dried over sodium sulphate and concentrated to about 5 cm^3, MDNA was precipitated by addition of two volumes of warm petroleum spirit (b.p. 40-60°C). The crude MDNA was filtered off and recrystallised from diethyl ether to give a white crystalline solid, m.p. 98-100°C (lit[10] 98-101°C). The ^{1}H NMR spectrum in DMSO-d$_6$ showed singlets at δ4.8 due to methylene protons and at δ13 due to amino protons[11]

Nitramide: A sample of nitramide was prepared by the method of Marlies, La Mer and Greenspan[12] with the exception that isopropyl nitrate was used in place of ethyl nitrate. This preparation involves N-nitration of urethane (ethyl carbamate) with isopropyl nitrate and concentrated sulphuric acid, as shown in Equation (2.1), a dry-ice/methanol bath being used to cool the reaction mixture to -10°C.

$$NH_2COOC_2H_5 \quad + \quad {}^iPrONO_2 \quad (+ \quad H_2SO_4) \quad \longrightarrow \quad NO_2.NHCOOC_2H_5$$

urethane   nitrourethane

$$+ \quad {}^iPrOH$$

Eqn. (2.1)

The nitrourethane is extracted into ether, and partially purified by subsequent extraction into ammonia solution, so forming an aqueous solution of ammonium nitrourethane salt. The nitrourethane is liberated by acidification with

sulphuric acid, is extracted into ether, and is precipitated as its ammonium salt by bubbling dry ammonia through the ethereal solution. The nitrourethane is converted to potassium nitrocarbamate, as shown in Equation (2.2), by addition of the ammonium salt to cold alcoholic potassium hydroxide.

$$NO_2.NHCOOC_2H_5 + 2 KOH \longrightarrow NO_2.NKCOOK + C_2H_5OH + H_2O$$

Eqn. (2.2)
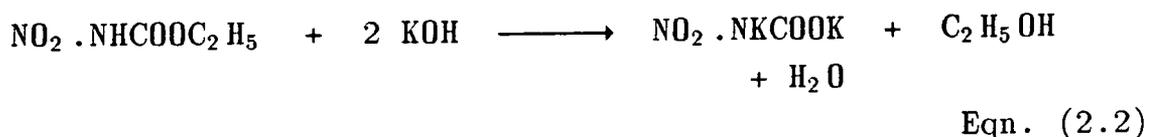
Nitramide is liberated by adding the potassium nitro-carbamate to sulphuric acid, as shown in Equation (2.3), and extraction into ether followed by solvent evaporation yielded shiny white leaflets.

$$NO_2.NKCOOK + 2 H_2SO_4 \longrightarrow NO_2.NH_2 + CO_2 + 2 KHSO_4$$

Eqn. (2.3)
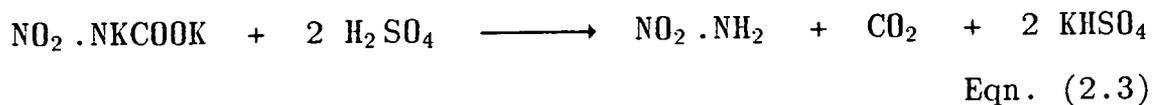
The product is extremely unstable, and should be stored in a desiccator in the refrigerator. It should be handled on an inert spatula, such as glass or teflon, as it attacks base metals to form products which catalyse its decomposition[12].

Dinitrogen pentoxide: This was prepared by the method of Reference 13, which involves oxidation of nitrogen dioxide/ dinitrogen tetraoxide by ozone. Ozone, formed by passage of dry oxygen through an ozoniser, was passed through a 20mm x 150mm glass tube packed for half its length with glass beads. Nitrogen dioxide was introduced perpendicular to the ozone stream, producing sufficient turbulence for mixing. The flow rate of the nitrogen dioxide was adjusted so that the orange-brown colour of the gas mixture was discharged just as it entered the glass-bead area of the reaction tube. The dinitrogen pentoxide collector, a removable trap, 50mm x 140mm, having an exit tube of 15mm, was attached to the reaction tube and cooled in a chloroform/dry ice slush bath. The exit tube from the trap led to a phosphorus pentoxide drying tower exhausting into a fume hood. Approximately 1 g of white crystalline dinitrogen pentoxide was collected and stored in a freezer to prevent decomposition.

3,7-bis(p-chloroaniline)-1,3,5,7-tetraazabicyclo[3.3.1]-nonane and **3,7-bis(p-ethoxycarbonylaniline)-1,3,5,7-tetraazabicyclo[3.3.1]nonane:** These compounds were synthesised by the procedure of Vaughan and co-workers[14] The appropriate arylamine (0.02 mol) in 2M-hydrochloric acid (30 cm^3, 0.022 mol) was diazotised at 0^{0}C with a solution of sodium nitrite (1.52 g, 0.022 mol) in water over a period of 1 hour. The diazonium salt solution was filtered, if necessary, and a mixture of concentrated ammonia (2.43 g, 0.04 mol) and 37% aqueous formaldehyde (16.17 g, 0.2 mol) was added slowly to the solution. After stirring for 15 minutes the mixture was neutralised with saturated sodium

bicarbonate solution, whereupon the product precipitated immediately. After filtering, drying, and recrystallisation from ethanol, pale yellow, needle-like crystals were obtained in the case of the p-ethoxycarbonyl derivative, m.p. 184-185°C (lit.[14] 185-186°C), and yellow-orange needle-like crystals in the case of the p-chloro derivative, m.p. 169-170°C (lit.[14] 166-168°C). A ^{1}H NMR spectrum[14] of the p-ethoxycarbonyl compound gave peaks at δ1.36 (t, J = 7 Hz, ethoxy CH_3), δ4.31 (q, J = 7.0 Hz, ethoxy CH_2), δ4.43 (s, CH_2 bridge), δ4.45 (br, ring CH_2), δ5.11 (br, ring CH_2), δ6.01 (br, ring CH_2) and δ7.44 (AB q, aromatic H). The p-chloro compound gave a similar spectrum with resonances at δ4.45 (s, CH_2 bridge), δ4.51 (br, ring CH_2), δ5.11 (br, ring CH_2), δ5.76 (br, ring CH_2) and δ7.08 (AB q, aromatic H).

1-(p-ethoxycarbonylphenyl)-3-methyltriazene: This compound was prepared by the method of Reference 15. In this process a solution of the p-ethoxycarbonylaniline (1 g) in 2M-hydrochloric acid (9 cm^3), diluted with water (20 cm^3) was diazotised at 0°C with sodium nitrite (0.45 g) in water, and the resulting solution was stirred for 1 hour or until clear. The diazonium salt solution was treated with 40% aqueous methylamine (2.8 cm^3) whereupon a precipitate of the triazene appeared immediately. Washing and drying yielded a yellow-green powder, m.p. 84-86°C (lit.[15] 83-85°C). The ^{1}H NMR spectrum[15] showed signals at δ1.33 (t, J = 7Hz, ethoxy CH_3), δ3.33 (br, s, N—CH_3), δ4.27 (q, J = 7 Hz, ethoxy CH_2), δ7.20 (br, d, J = 8 Hz, aromatic H) and δ7.80 (d, J = 8 Hz, aromatic H). Elemental analysis gave C 58.0%, H 6.3%, N 20.3%, O 15.5% (theoretical C 58.3%, H 6.0%, N 18.6%,

O 17.1%).

2.1.5 Buffer solutions

Buffer solutions were prepared by standard methods[16] and the pH tested using a PTI-6 Universal Digital pH Meter (accurate to ± 0.02 pH units).

Citrate buffer, pH 3.5-6.5: 0.2M-citric acid + 0.6M-sodium hydroxide solution.

Tris buffer, pH 7.0-7.7: 0.02M aqueous tris(hydroxymethyl) aminomethane + 0.02M-hydrochloric acid.

Borax buffer, pH 8.0-9.0: 0.025M aqueous borax + 0.1M-hydrochloric acid.

2.2 Measurement techniques

2.2.1 UV/visible measurements

All UV/visible spectra were recorded using fresh solutions in 1 cm quartz cells on either a Perkin-Elmer Lambda 3 or Philips PU8725 instrument. These same instruments were also used for kinetic and equilibrium measurements (made at $25^{\circ}C$), except in the case of fast reactions, in which case the stopped-flow spectrophotometer was used (see later).

All kinetic measurements were made under first-order conditions, and observed rate coefficients were determined by following the change in absorbance at an appropriate wavelength. Measured absorbance values were entered into a suitable program running on an Apple IIe microcomputer, which calculated the observed rate coefficient based on the following derivation.

For a decrease in absorbance:

$$-\frac{d[A]}{dt} = k_{obs}[A] \qquad [A] = \text{absorbance}$$

$$\int_{[A]_0}^{[A]_t} \frac{d[A]}{[A]} = -k_{obs}\int_0^t dt \qquad \begin{array}{l} \text{where} \\ [A]_0 = \begin{array}{l}\text{initial}\\\text{absorbance}\end{array} \\ [A]_t = \begin{array}{l}\text{absorbance}\\\text{at time } t\end{array} \end{array}$$

$$\ln\left[\frac{[A]_t}{[A]_0}\right] = -k_{obs}t$$

Assuming absorbance does not fade to zero:

$[A]_0$ becomes $[A]_0 - [A]_\infty$ and $[A]_t$ becomes $[A]_t - [A]_\infty$

where $[A]_\infty$ is the absorbance at 'infinite' time

Therefore
$$\ln\left[\frac{[A]_t - [A]_\infty}{[A]_0 - [A]_\infty}\right] = -k_{obs}t$$

$$[A]_0 - [A]_\infty = \text{constant}$$

Thus a plot of $\ln([A]_t - [A]_\infty)$ vs. time has a gradient of $-k_{obs}$. If following an increase in absorbance a plot of $\ln([A]_\infty - [A]_t)$ vs. time has a gradient of $-k_{obs}$.

For measurement of rate coefficients of reactions too fast for the conventional machines a Hi-Tech Scientific SF-3L stopped-flow spectrophotometer was used. This is shown schematically in Figure 2.1.

Figure 2.1 Schematic representation of stopped-flow
UV/visible spectrophotometer

T = three way taps

M = mixing point

O = observation point
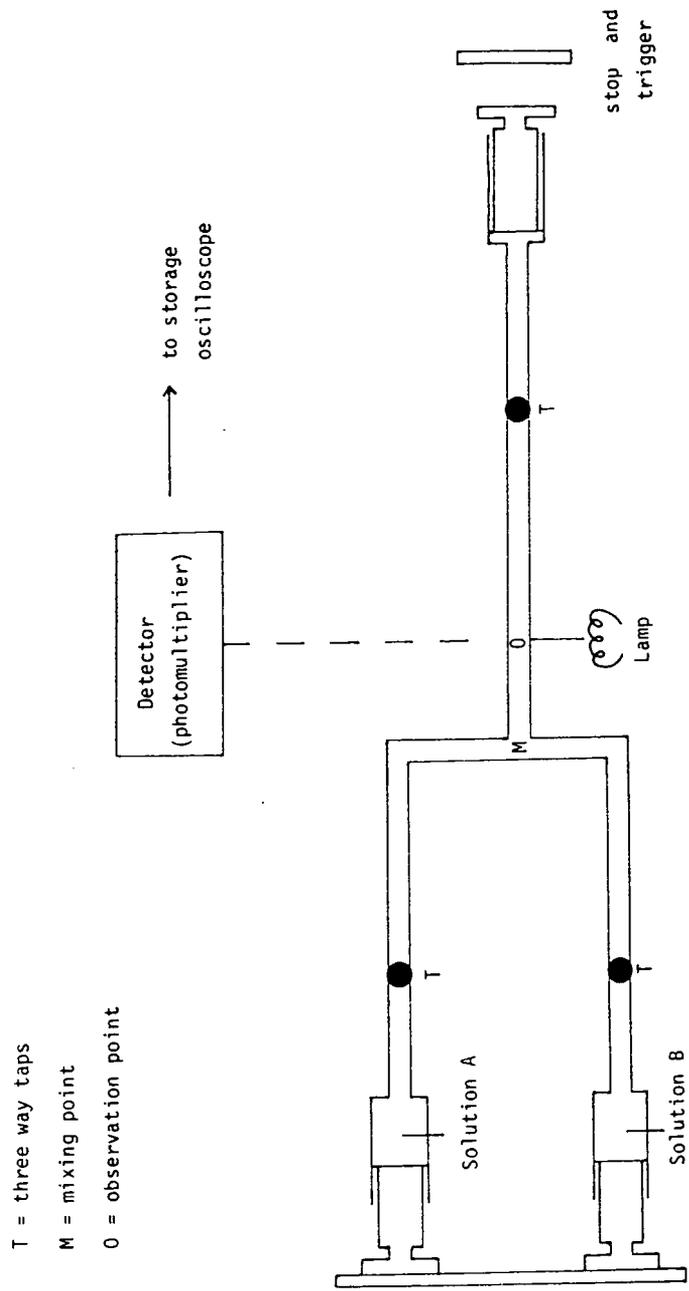
The two solutions, A and B, which undergo reaction are stored in glass reservoirs, and from there enter identical syringes. A single piston drives both syringes, so that equal volumes of each solution are mixed at point M (halving the concentration of each solution), before passing into a thermostatted 2 mm quartz cell at point O. When the plunger of the third syringe hits the stop the flow of reactants stops and the trigger causes monitoring of the reaction at O to begin. This is done by passing a beam of monochromatic light of the appropriate wavelength throught the cell. The reaction within the cell causes an increase or decrease in the transmitted light which is fed through a photomultiplier and displayed on an oscilloscope screen, from where voltage changes can be read off. Although the output of the photomultiplier is not linear, for a small percentage change the voltage can be assumed to be proportional to the absorbance of the reaction mixture, and therefore a plot of $\log|V_t - V_\infty|$ vs. time gives k_{obs} from $|gradient|$.

2.2.2 Mass spectrometry

Mass spectrometric measurements were made using a 7070E instrument supplied by V.G. Analytical Ltd. Both Electron Impact and Chemical Ionisation methods were employed, the most suitable for any particular compound being determined empirically.

2.2.3 NMR spectrometry

NMR spectra were recorded on either a Hitachi/Perkin-Elmer R-24B (60 MHz) or a Bruker AC250 (250 MHz) instrument. Shift measurements are quoted as 'δ' values relative to tetramethylsilane (TMS).

2.2.4 pH measurements

All pH measurements were carried out using a PTI-6 Universal Digital pH Meter (accurate to ± 0.02 pH units).

2.2.5 Quantitative determination of formaldehyde

As described by Vogel,[17] to an aqueous solution containing formaldehyde is added an excess of a saturated alcoholic dimedone (5,5-dimethylcyclohexane-1,3-dione) solution. This results in precipitation of the pale orange-brown 1:1 derivative, which after filtration, washing with water and recrystallisation from 50% aqueous ethanol yields a quantity of white crystals (m.p.[18] 189^{0}C) in molar proportion to the amount of formaldehyde in the solution under test.

2.3 Attempted preparation of DADN by acetylation of DPT

Since DPT is a by-product of the production of RDX, it would obviously be advantageous if it could be converted to DADN, and subsequently to HMX. Initial attempts to effect acetylation of DPT using acetic anhydride at 120°C, and acetyl chloride/acetic anhydride/glacial acetic acid mixtures at 120°C failed to produce any DADN. However, a procedure was found which yielded DADN from DPT which involved stirring 1 equivalent DPT and 2 equivalents acetyl chloride in an excess of acetic anhydride (solvent) at 20°C for 24hrs. Water was added to hydrolyse excess acetic anhydride, and the resulting acidic solution was rendered alkaline with sodium carbonate. The suspension formed was filtered off and washed with water before drying. The ^{1}H NMR spectrum showed the white solid to be a mixture of DPT and DADN, and so the mixture was washed very thoroughly with acetone and dried to give reasonably pure DADN in 50% yield. Increasing the reaction time to 48 hours did not increase the yield of DADN. Figures 2.2-2.5 show how the reaction has proceeded by use of ^{1}H NMR analysis.

Figure 2.2 ^{1}NMR spectrum of DPT

Solvent: DMSO-d_6

H$_2$O in solvent

δ (ppm)
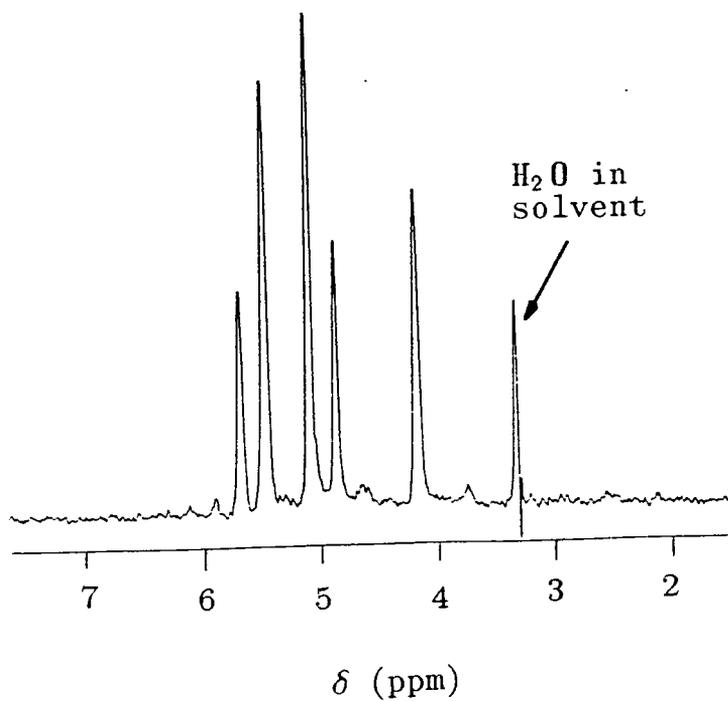
Figure 2.3 ^{1}NMR of mixture recovered from attempted acetolysis of DPT

Solvent: DMSO-d_6

H$_2$O in solvent

δ (ppm)
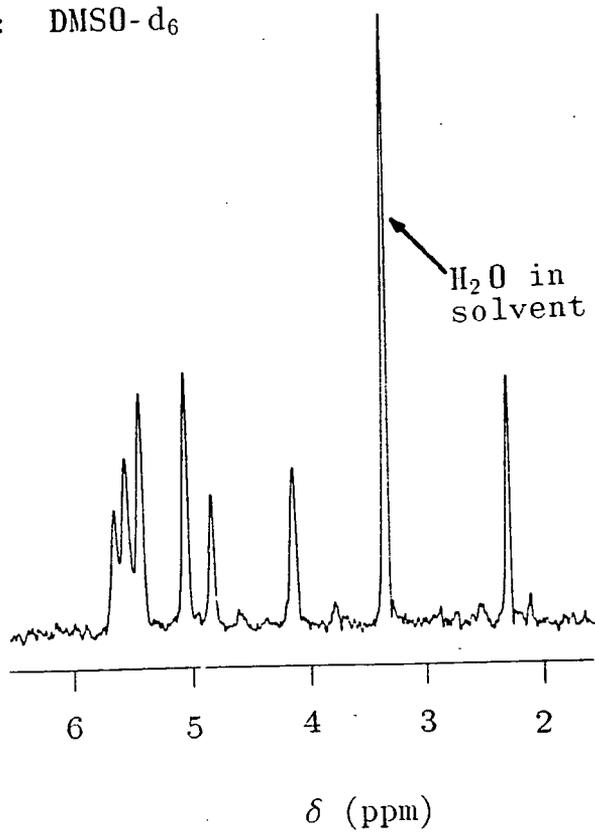
Figure 2.4 ^{1}H NMR of product after washing with acetone

Solvent: DMSO-d_6

H_2O in solvent

δ (ppm)
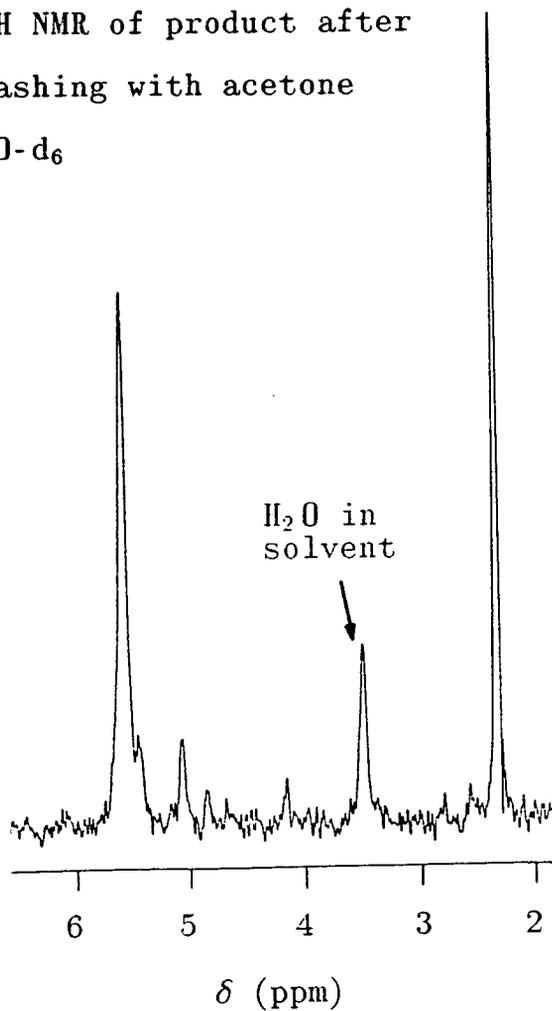
Figure 2.5 ^{1}H NMR of DADN

Solvent: DMSO-d_6

H_2O in solvent

δ (ppm)
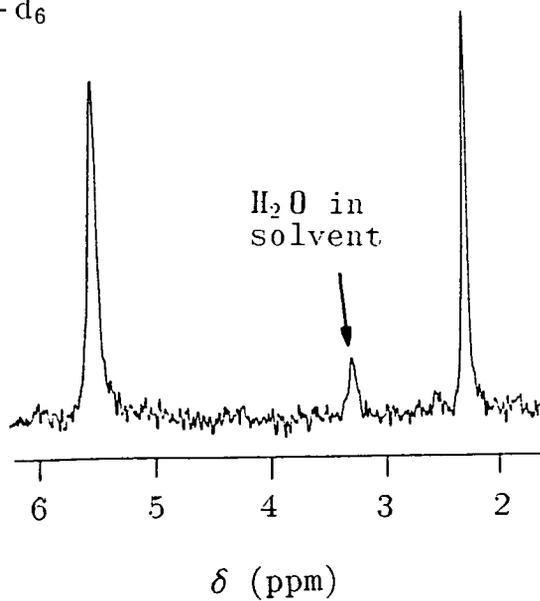
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CHAPTER 3

Mechanistic studies of the formation of DPT and DADN using ^{15}N-compounds

3.1 Introduction

In Chapter 1 it was seen how hexamine (1.1) can undergo reactions with a variety of electrophilic species to form N-substituted products, which either retain the eight-membered tetraaza ring structure of hexamine, or are of a six-membered triaza structure. From the viewpoint of the military, it is important that the factors which determine whether the nitration process in particular produces the six-membered RDX (1.2), or the eight-membered HMX (1.3), are understood. The tetranitro-substituted high explosive, HMX, is the desired product, but at present can only be produced in lower yield and at higher cost than the trinitro-substituted high explosive, RDX. The mechanisms for conversion of hexamine to RDX and HMX so far proposed[1-5] have generally suggested either the selective cleavage of bonds within the hexamine molecule during nitration, or total cleavage to smaller fragments, followed by nitration and recombination. There is evidence[6,7] for the latter pathway in the Bachmann process,[1,3,4] which involves nitration of hexamine using a mixture of ammonium nitrate, nitric acid and acetic anhydride.

Earlier work[8] on the mechanism of acetylation of hexamine to DAPT (1.6), using mixtures of pure ^{14}N- and ^{15}N-compounds as starting materials, demonstrated the reaction to occur via selective cleavage of the methylene bridge. It was decided to employ this particularly useful technique in re-investigating the mechanisms of formation of DPT (1.5) via nitration of hexamine, and formation of DADN (1.9) via nitration of DAPT.

3.2 Experimental

Preparation of ^{15}N-DAPT

^{15}N-DAPT was prepared from ^{15}N-hexamine by reaction with acetic anhydride in the presence of sodium hydroxide, as detailed by Siele, Warman and Gilbert[9,14] for the preparation of ^{14}N-DAPT, and described in Chapter 2. The product obtained gave a melting point and ^{1}H NMR consistent with that of the ^{14}N-compound.

A V.G. Analytical Ltd. 7070E instrument was employed for mass spectrometric measurements. Chemical Ionisation (CI) using iso-butane as a reagent gas was found to be the best method of mass spectrometric analysis for many of the compounds. This gives rise to the protonated species $(M+1)^+$. Results obtained using this technique are given in terms of the parent species whose mass, M, is one unit smaller than that observed. Electron Impact (EI) was used for other compounds.

Table 3.1 shows the results of mass spectral analysis of compounds prepared with nitrogen in natural abundance (^{14}N 99.63%, ^{15}N 0.37%). The peaks with masses M+1 and M+2 are due to naturally occurring ^{13}C and ^{15}N. Data are also given for ^{15}N-hexamine and ^{15}N-DAPT prepared from ^{15}NH$_3$. Peaks from the protonated species have been corrected, where necessary, for the presence of peaks from the parent species, which remain due to incomplete protonation.

Table 3.1 Mass spectrometric data for starting materials

Compound	Observed peak intensities			Theoretical peak intensities		
	M	M+1	M+2	M	M+1	M+2
^{14}N-hexamine (CI)	100	10.1	0.5	100	8.4	0.3
^{14}N-DAPT (CI)	100	13.0	1.5	100	11.9	1.1
^{14}N-DPT (CI)	100	8.5	1.1	100	8.1	1.1
^{15}N-hexamine (CI)	100	6.9	–	100	6.9	0.2
^{15}N-DAPT (CI)	100	9.7	1.6	100	10.4	0.9

The good agreement, in all cases, of the intensities of the M+1 and M+2 peaks with those calculated theoretically using known isotopic abundances is good evidence that the analytical technique is sound.

3.3 Investigation of the mechanism of formation of DPT via nitration of hexamine

In the method of preparation of DPT as detailed in Reference 10 (described in Chapter 2), ammonium nitrate is not present in the nitrating medium, thus eliminating the possibility of N-atom exchange prior to nitration. DPT is precipitated from the mother liquor by neutralisation with ammonia solution. Two experiments were conducted, one using $^{14}NH_3$ solution to neutralise, and one using $^{14}NEt_3$. This was done in order to establish whether ammonia was being incorporated in the final product, since triethylamine obviously could not be.

Experiment (a): DPT from a ^{14}N-hexamine/^{15}N-hexamine mixture using $^{14}NH_3$ solution to neutralise

DPT was prepared, using the method of Reference 10, from a mixture of 0.25 g (1.79 mmol) ^{14}N-hexamine and 0.25 g (1.74 mmol) ^{15}N-hexamine as starting material, (overall atomic ratio $^{14}N:^{15}N$ = 100:96). After removal of RDX, DPT was precipitated from the mother liquor by neutralisation with $^{14}NH_3$ solution. Mass spectroscopic data for the DPT obtained are given in Table 3.2.

Experiment (b): DPT from a ^{14}N-hexamine/^{15}N-hexamine using $^{14}NEt_3$ to neutralise

Experiment (a) was repeated, the procedure varying only in that $^{14}NEt_3$ was used to effect precipitation of DPT in place of $^{14}NH_3$ solution. Mass spectroscopic data for the DPT obtained are found in Table 3.2.

The M+1, M+2, M+3 and M+4 peaks have been corrected for the natural occurrence of ^{15}N and ^{13}C, as well as for the effects of incomplete protonation.

Table 3.2 Relative isotopic composition of DPT prepared in experiments (a) and (b)

	M $[^{14}N_4]$	M+1 $[^{14}N_3{}^{15}N_1]$	M+2 $[^{14}N_2{}^{15}N_2]$	M+3 $[^{14}N_1{}^{15}N_3]$	M+4 $[^{15}N_4]$
Expt.					
(a) (CI)	100	183.7	102.9	12.7	0.1
(b) (CI)	100	381.9	543.8	362.7	66.6
random[a]	100	396.0	576.0	364.0	85.0
selective[b]	100	-	-	4.0	96.0

a. this is the calculated relative distribution of isotopes for a random distribution with the starting composition used in experiments (a) and (b).

b. this is the calculated relative distribution of isotopes for totally selective ring cleavage with the starting composition used in experiments (a) and (b).

In order to eliminate the possibility of exchange between ^{14}N-hexamine and ^{15}N-hexamine prior to nitration, a third experiment, (c), was conducted.

Experiment (c): Investigation of possible N-atom exchange
between ^{14}N-hexamine and ^{15}N-hexamine

0.1 g (0.71 mmol) ^{14}N-hexamine and 0.1 g (0.69 mmol) ^{15}N-hexamine were thoroughly mixed and added to 10mls. 2M-nitric acid at 20°C. The hexamine mixture dissolved and the resulting solution was allowed to stand for 5 minutes, after which time the solution was cooled in an ice bath to effect precipitation of the mononitrate salt of hexamine, of the structure shown below.
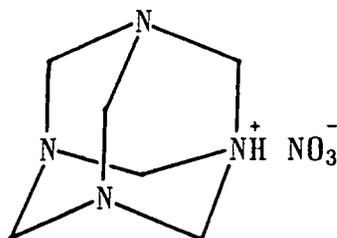
(<u>3.1</u>) hexamine mononitrate

The crystals were filtered off and quickly washed with 50% ethanol, then ethanol, then ether, followed by thorough drying.

The initial attempt at the above preparation did not result in precipitation of the mononitrate on cooling, and in an attempt to recover the hexamine or its salt, the solution was heated at 70°C on a rotary evaporator. The mononitrate salt obtained from this procedure gave interesting results, which are designated as being those of experiment (d).

Mass spectroscopic data obtained using both Chemical Ionisation and Electron Impact methods are given in Table 3.3. Data are also given for an authentic sample of

^{14}N-hexamine mononitrate prepared by this method from hexamine with natural isotopic abundance. The peak observed in the mass spectrum is that of the hexaminium cation of mass 141 (^{14}N-sample), and not that of the salt itself.

Table 3.3 Relative isotopic composition of hexamine mononitrate prepared in experiments (c) and (d)

Compound	Observed peak intensities			Theoretical peak intensities		
	M	M+1	M+2	M	M+1	M+2
^{14}N-hexamine mononitrate (CI)[a]	100	7.5	1.0	100	8.8	1.0
^{14}N-hexamine mononitrate (EI)	100	7.6	-	100	8.8	1.0

	M $[^{14}N_4]$	M+1 $[^{14}N_3{}^{15}N_1]$	M+2 $[^{14}N_2{}^{15}N_2]$	M+3 $[^{14}N_1{}^{15}N_3]$	M+4 $[^{15}N_4]$
Expt.					
(c) (CI)	100	9.6	17.3	17.1	104.8
(c) (EI)	100	5.7	2.7	9.3	99.6
(d) (CI)	100	298.7	355.3	245.2	61.8
total scrambling	100	400.0	588.0	376.0	88.0
no scrambling	100	-	-	4.0	96.0

a. extensive deprotonation of the hexaminium cation occurred, detected by the presence of a large M-1 peak corresponding to mass 140.

The results for the nitration of hexamine to DPT in Table 3.2 show conclusively that extensive non-selective ring cleavage is occurring, to give fragments containing single, chemically equivalent N-atoms. The results of experiment (c) demonstrate the lack of N-atom exchange between ^{14}N-hexamine and ^{15}N-hexamine in an acid medium at a temperature similar to that of the nitration process. However, experiment (d) shows that extensive scrambling of N-atoms occurs at elevated temperatures due to exchange, or fragmentation and recombination of the hexamine molecules.

Experiment (b) gave results in particularly good agreement with those calculated for total scrambling during the formation of DPT, the most abundant species being those containing both ^{14}N and ^{15}N. This is a great contrast to the results found for acetylation[8] and serves to confirm the work of Castorina and co-workers.[6,7] The results of experiment (a) also show extensive scrambling to occur during nitration, but there is a much larger relative abundance of ^{14}N-containing species than ^{15}N. This would indicate that ^{14}NH$_3$ used for neutralisation is being incorporated into the DPT formed by one of three mechanisms:

    i)   via reaction or exchange of ^{14}NH$_3$ with precursors to DPT.

    ii)   via exchange of ^{14}NH$_3$ with the formed DPT itself.

    iii)   via reaction of ^{14}NH$_3$ with liberated formaldehyde to form more ^{14}N-hexamine, which undergoes nitration to form ^{14}N-DPT exclusively.

The first of these three possibilities is the most likely on the basis of the observations made in this work. It is reasonable to disregard the second mechanism on the basis

that DPT immediately precipitates out on formation and thus cannot undergo exchange with species in solution, and the third option may be eliminated because of the fact that at the time when $^{14}NH_3$ is added the reaction mixture is diluted with water and unable to act as a nitrating medium.

3.4 Investigation of the mechanism of formation of DADN via nitration of DAPT

The nitration of DAPT to DADN was examined via a similar procedure to that used for investigation of the nitration of hexamine to DPT.

DADN was prepared by the method of Yoshida and co-workers[11] for the preparation of ^{14}N-DADN, as detailed in Chapter 2, from a mixture of 0.2 g (0.94 mmol) ^{14}N-DAPT and 0.2 g (0.93 mmol) ^{15}N-DAPT (overall atomic ratio ^{14}N:^{15}N = 100:97). Mass spectrometric data for the DADN obtained are found in Table 3.4. Data are also given for an authentic sample of ^{14}N-DADN.

Table 3.4 Relative isotopic composition of DADN prepared from ^{14}N-DAPT and ^{15}N-DAPT

Compound	Observed peak intensities			Theoretical peak intensities		
	M	M+1	M+2	M	M+1	M+2
^{14}N-DADN (CI)	100	12.6	1.5	100	11.6	1.8

	M	M+1	M+2	M+3	M+4
	$[^{14}N_4]$	$[^{14}N_3{}^{15}N_1]$	$[^{14}N_2{}^{15}N_2]$	$[^{14}N_1{}^{15}N_3]$	$[^{15}N_4]$
DADN from expt. (CI)	100	-	-	-	65.1
random	100	400.0	588.0	376.0	88.0
selective	100	-	-	4.0	96.0

The results show conclusively that nitration of DAPT to DADN proceeds exclusively via selective cleavage of the methylene bridge of the DAPT molecule, this being demonstrated by the absence of ^{14}N/^{15}N scrambling. This is in marked contrast to the nitration of hexamine to DPT, but shows a great similarity with the acetylation of hexamine to DAPT. In order to explain the results, it is necessary to look at the factors controlling cleavage of these ring systems.

3.5 The selective cleavage of 1,3,5,7-tetraazabicyclo-[3.3.1]nonanes

A study by Yoshida and co-workers[11] considered the general 1,3,5,7-tetraazabicyclo[3.3.1]nonane structure (3.2).

$$
\begin{array}{ccccc}
8 & \text{—} & \overset{1}{N} & \text{—} & 2 \\
| & & | & & | \\
R\text{—}\overset{7}{N} & \overset{9}{} & CH_2 & & \overset{3}{N}\text{—}R \\
| & & | & & | \\
6 & \text{—} & \underset{5}{N} & \text{—} & 4 \\
\end{array}
$$

(3.2)

$$
\begin{array}{c}
X \\
| \\
\overset{+}{N} \\
R\text{—}N \quad CH_2 \quad N\text{—}R \\
\overset{+}{N} \\
| \\
X
\end{array}
$$

(3.3)

$$
\begin{array}{c}
X \\
| \\
N \\
R\text{—}N \qquad N\text{—}R \\
N \\
| \\
X
\end{array}
$$
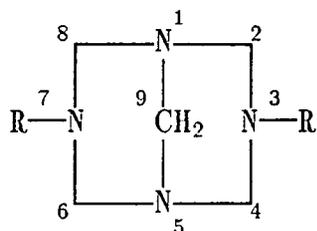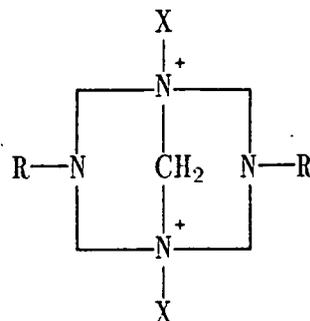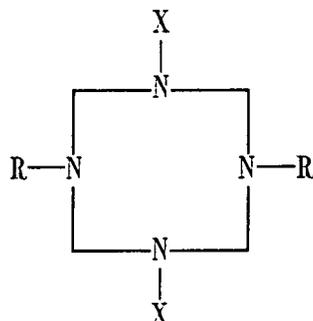
(3.4)

Selective cleavage of the methylene bridge in formation of (3.4) requires exclusive attack at C-9. Yoshida suggests that this is promoted by a reduction in the basicity (or nucleophilicity) of the 3,7-nitrogens, achieved by R- being electron-withdrawing, and the formation of an intermediate diquaternary salt of structure similar to (3.3). This is followed by hydrolytic cleavage of the methylene bridge via specific attack on C-9. This theory would therefore account

for the selective cleavage of DAPT (($\underline{3.2}$), R = Ac) during nitration to DADN (($\underline{3.4}$), R = Ac, X = NO_2). However, by analogy one might expect selective cleavage of DPT (($\underline{3.2}$), R = NO_2) during nitration to HMX (($\underline{3.4}$), R = NO_2, X = NO_2), but this has been shown to be clearly not the case.[6,7] If one applies this theory to the hexamine structure ($\underline{1.1}$) with regard to conversion to structures similar to ($\underline{3.2}$), since all four nitrogen atoms are equivalent then non-selective cleavage would be predicted. This has observed to be true for the case of nitration to DPT, but not for acetylation to DAPT.

Considering processes represented by the conversion of ($\underline{3.2}$) to ($\underline{3.4}$), since acetyl and nitro substituents clearly reduce the nucleophilicity of the 3,7-nitrogens, the flaw in the theory must arise from the postulated formation of the intermediate diquaternary salt ($\underline{3.3}$). It seems more probable that the molecule would undergo aminal-type cleavage[12,13] after formation of the monoquaternary salt ($\underline{3.5}$), to form a monocyclic iminium ion of type ($\underline{3.6}$).

($\underline{3.5}$)   ($\underline{3.6}$)
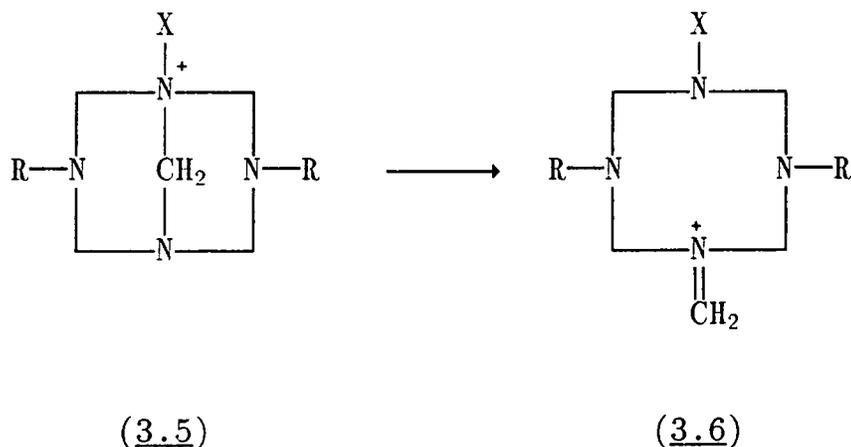
Eqn. (3.1)

The subsequent mechanism of the reaction depends on whether (3.6) undergoes ring cleavage or hydrolysis of the carbon-nitrogen double bond of the iminium group.

3.6 Discussion

Results for the nitration of hexamine with nitric acid indicate complete cleavage to species containing no more than one nitrogen atom. This is in marked contrast to the acetylation reaction[8] where the ring structure remains basically intact, but is similar to results obtained in the Bachmann reaction[6,7] The first step in both nitration and acetylation reactions is likely to be attack by an electrophile, E^+, at nitrogen, to give a cation (3.7) which can exist in a ring-opened form (3.8).

$$(1.1) + E^+ \rightleftharpoons (3.7) \rightleftharpoons (3.8)$$

(3.7)   (3.8)
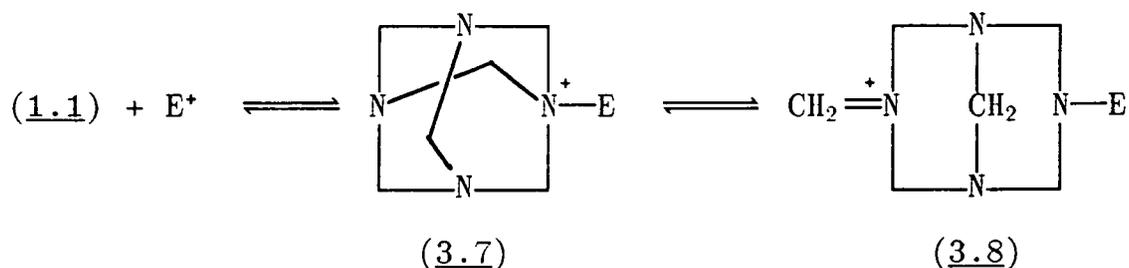
Scheme 3.1

The subsequent course of the reaction depends on the rate of hydrolysis of the $CH_2=N^+$ group of (3.8) compared to cleavage of other C–N bonds. In the acetylation reaction, carried out in the presence of water, and in a medium which is not highly acidic, hydrolysis will eliminate formaldehyde and yield a species which is readily acetylated to form DAPT. However, in nitric acid, hydrolysis of the $CH_2=N^+$

group, involving attack by water and proton elimination, will be more difficult, so that further nitration of the molecule accompanied by C-N bond cleavage and fragmentation occurs preferentially. This interpretation is also compatible with the mechanism of diazotisation of hexamine involving selective cleavage discussed in Chapter 1.

The selective cleavage observed in the conversion of DAPT to DADN shows that nitration does not necessarily result in extensive decomposition. Reaction is likely to involve the intermediate (3.9).

$$
\begin{array}{c}
\text{Ac} \\
|\\
\text{N}\\
\end{array}
$$

$$
\text{H}_2\text{C} \overset{+}{=\!\!=} \text{N} \qquad\qquad \text{N}\text{---}\text{NO}_2
$$

$$
\begin{array}{c}
\text{N}\\
|\\
\text{Ac}
\end{array}
$$
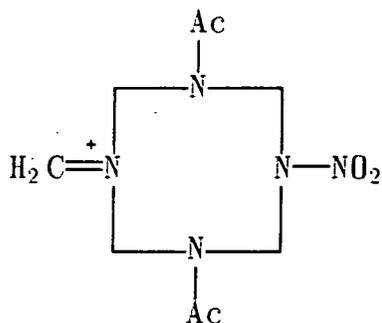
(3.9)

The presence of the electron-withdrawing nitro and acetyl groups will reduce the electron density at ring nitrogen atoms, reducing the possibility of further cleavage of the molecule. Nevertheless, there is evidence[6,7] that under the more severe conditions used in the Bachmann conversion of DPT to HMX extensive breakdown occurs.

A general conclusion of this, and previous work,[8] is that high yields of products containing eight-membered rings are only produced from hexamine derivatives by selective cleavage, in which the main ring structure remains intact. Extensive cleavage followed by recombination of fragments

usually results in the preferential formation of products with six-membered rings. A possible mechanism for the nitration of hexamine is discussed in Chapter 4.

3.7 References

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CHAPTER 4

The decomposition of DPT in acidic and basic aqueous solutions

4.1 Introduction

DPT (1.5) is one of the products of the nitration of hexamine (1.1) with nitric acid[1-5] and is important since it has previously[6,7] been postulated as an intermediate in the Bachmann[8-10] method of synthesis of RDX (1.2) and HMX (1.3) from hexamine (1.1). DPT is formed by neutralisation of the acidic aqueous filtrate produced during RDX synthesis,[1-4] and it therefore seems appropriate that the behaviour of DPT in acidic and basic aqueous media should be studied, as it may then be possible to speculate on the nature of the species formed in solution during the nitration process.

4.2 UV spectra of DPT

A sample of DPT was prepared as described in Chapter 2.

DPT was previously[11] found to be stable as a solution in acetonitrile, but decomposed in two stages in various aqueous media. The UV spectrum of DPT in acetonitrile gave λ_{max}= 241 nm, ϵ_{max}= 9760 1 mol^{-1} cm^{-1}. Preliminary investigations of the UV spectra of DPT in hydrochloric acid and aqueous sodium hydroxide solutions, prepared by addition of 0.1 cm^3 of a stock solution of 10^{-3} M DPT in acetonitrile to 2 cm^3 of the appropriate aqueous media, yielded the results given in Table 4.1 and Table 4.2 respectively. These spectra represent the product of the fast initial stage of the decomposition.

Table 4.1 UV data for DPT (4.8×10^{-5} M) in hydrochloric acid (5% acetonitrile) at 25°C

$[HCl]\backslash mol\ l^{-1}$	$\lambda_{max}\backslash nm$	$\epsilon_{max}\backslash l\ mol^{-1}\ cm^{-1}$
1.0	214	13760
0.5	213	13990
0.1	213	13970
0.05	213	13790
0.01	209	12340
0.005	205	12800

Table 4.2 UV data for DPT (4.8×10^{-5} M) in sodium hydroxide solution (5% acetonitrile) at 25°C

$[NaOH]\backslash mol\ l^{-1}$	$\lambda_{max}\backslash nm$	$\epsilon_{max}\backslash l\ mol^{-1}\ cm^{-1}$
1.0	229	16400
0.5	229	15490
0.1	230	15850
0.05	230	15760
0.01	232	13020
0.005	232	12760
0.001	236	11690

The results given in Table 4.1 show the hypsochromic shift from 240 nm to ca. 210 nm of the absorption maximum of DPT upon acidification, a process which has not been observed for any other secondary nitroamine.[12] This shift might be associated with protonation of one of the tertiary amino nitrogen atoms of DPT, but since this basic N-atom is separated from the N-nitro chromophores by methylene groups

this seems unlikely to have such an effect. The conclusion is therefore reached that the DPT must undergo decomposition in acidic solution. The results in Table 4.2 show the less pronounced hypsochromic shift to <u>ca</u>. 230 nm in basic media.

The work of Jones and Thorn[12] resulted in the following guidelines for use when interpreting the UV spectra of aliphatic nitroamines.

i) **Secondary nitroamines**

For compounds containing one secondary nitroamine group per molecule, in neutral solution ϵ_{max} approximates to 5500 l mol^{-1} cm^{-1}. Observation has indicated that for a molecule containing n secondary nitroamine functions ϵ_{max} approximates to 5500 × n l mol^{-1}cm^{-1}.

ii) **Primary nitroamines**

In neutral solution ϵ_{max} for the primary nitroamine functions is observed as approximately 7000 l mol^{-1} cm^{-1}.

iii) **Nitroamines in alkaline solution**

The absorption maximum of primary nitroamines in alkaline solution shows a shift of λ_{max} to longer wavelength and an increase in ϵ_{max}. For a molecule containing n primary nitroamine functions, ϵ_{max} approximates to 8500 × n l mol^{-1} cm^{-1}. Secondary nitromamines in alkaline solution, where observable, show no change in λ_{max} from that observed in neutral solution.

The value of ϵ_{max} of 9760 l mol^{-1} cm^{-1} for DPT in acetonitrile agrees quite well with the 5500 × n formula for secondary nitroamines. The increase in ϵ_{max} in acidic media to approximately 14000 l mol^{-1} cm^{-1}, and in basic media to approximately 16000 l mol^{-1} cm^{-1}, agrees quite well with the 7000 × n and 8500 × n formulae proposed for primary

nitroamines in acidic/neutral and alkaline solution respectively. It therefore appears that in acidic or basic media the secondary nitroamine functions of DPT are converted via some decomposition process to primary nitroamine functions.

4.3 Kinetic data for the decomposition of DPT

Work by Cooney[11,13] prior to that of the author, showed DPT to decompose in both acidic and basic aqueous media in two observable stages. The kinetics of the initial decomposition process were studied over a wide pH range, and so rate coefficients for this step were measured at a single acid and base concentration, for comparison purposes only. Cooney studied the decomposition of the intermediate formed in basic solution over a wide pH range, and a similar analysis of the intermediate formed in acid solution will be considered here.

The initial decomposition of DPT in 0.01M-hydrochloric acid and 0.01M-sodium hydroxide solution to form the intermediate species at <u>ca</u>. 210 nm and <u>ca</u>. 230 nm respectively was followed as a decrease in absorbance with time at 240 nm. Kinetic analysis gave:

$$k_{obs} = \underline{4.1 \times 10^{-3} \text{ sec}^{-1}} \text{ in the acid medium}$$

$$\text{and } k_{obs} = \underline{8.1 \times 10^{-3} \text{ sec}^{-1}} \text{ in the basic medium}$$

Once the initial product in acid or base had been formed, a second, very slow decomposition was seen to occur, resulting in fading of the UV spectrum to zero absorbance.

Kinetic analysis of this second decomposition in

0.01M-hydrochloric acid gave

k_{obs}= $\underline{5.0 \times 10^{-5} sec^{-1}}$, measured as a decrease in

absorbance at 210 nm,

and in 0.01M-sodium hydroxide solution gave

k_{obs}= $\underline{4.7 \times 10^{-5} sec^{-1}}$, measured as a decrease in

absorbance at 235 nm.

The observed rate coefficients for the decomposition of

the intermediate formed in acid solution over a broad pH

range are given in Table 4.3. The intermediate was prepared

by allowing a freshly prepared solution of DPT in

0.01M-hydrochloric acid to stand for 15 minutes, after which

time it was transferred to the appropriate medium,

compensating for the presence of acid in the solution where

necessary. In some cases catalysis by buffer components was

observed, and the values for the observed rate coefficient

have been extrapolated back to zero buffer concentration.

Table 4.3 Observed rate coefficients for the secondary decomposition of DPT (4.8×10^{-5} M) in aqueous solution at 25°C

Medium	pH	wavelength decomposition measured at \nm	k_{obs}\sec^{-1}
0.01M HCl	2.0	210	5.0×10^{-5}
0.001M HCl	3.0	210	4.3×10^{-5}
acetate buffer	5.0	220	extrapolates to zero
tris buffer	7.0	225	extrapolates to zero
tris buffer	9.0	225	extrapolates to zero
0.001M NaOH	11.0	225	4.0×10^{-3}
0.01M NaOH	12.0	225	5.0×10^{-2}
0.1M NaOH	13.0	225	0.53

A plot of $\log_{10}k_{obs}$ vs. pH is shown in Figure 4.1

Figure 4.1 A plot of $\log_{10}k_{obs}$ vs. pH for the secondary decomposition of DPT at 25°C
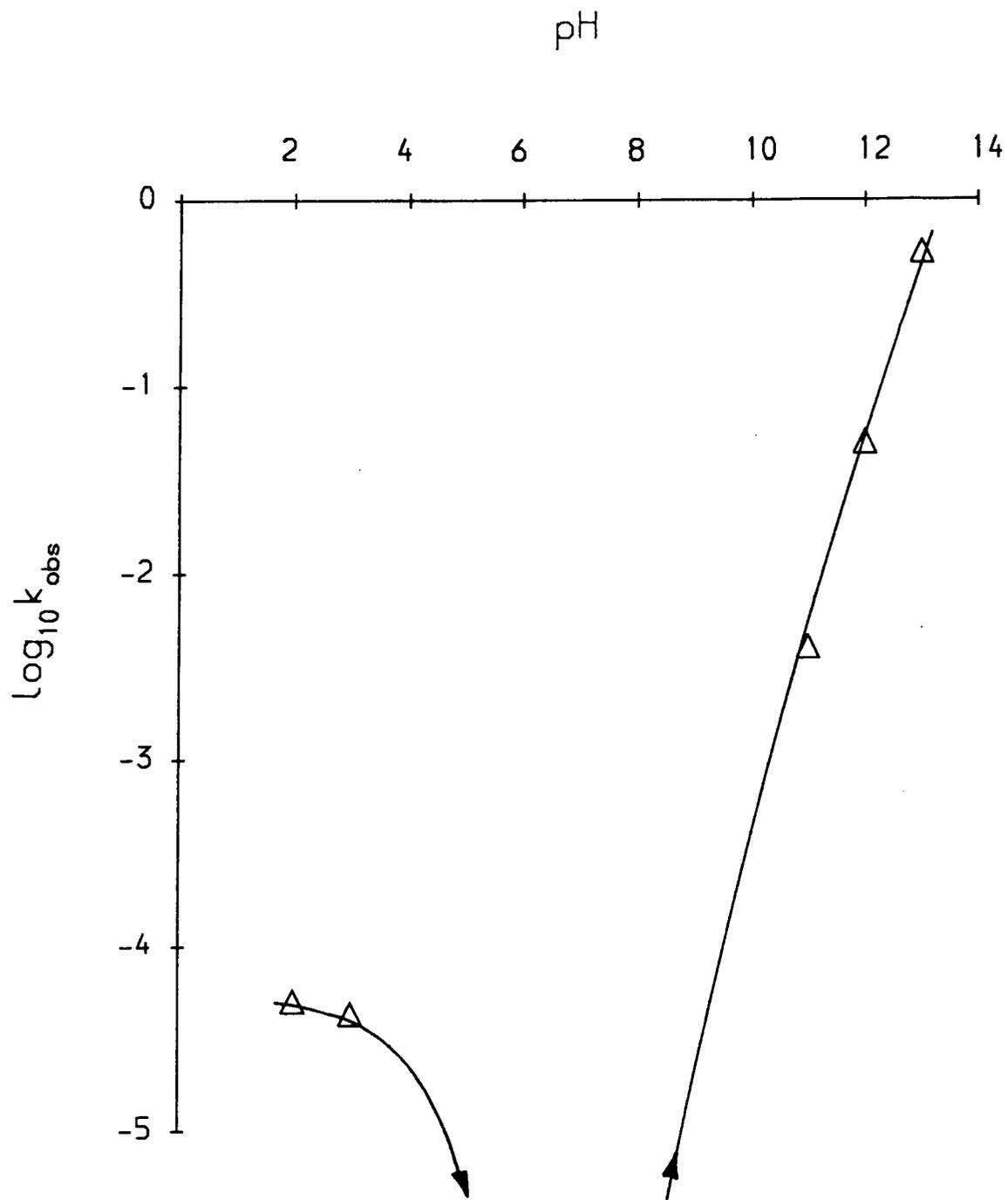
4.4 The role of methylenedinitroamine

It has been postulated that methylenedinitroamine (4.1), from hereon referred to as MDNA, may exist as a precursor to DPT in the nitration of hexamine.

$$\text{H}_2\text{C} \Big\langle \begin{array}{l} \text{NHNO}_2 \\ \text{NHNO}_2 \end{array}$$
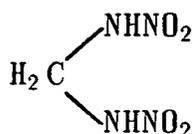
(4.1) MDNA

The strongest evidence for this is the work of Lamberton and co-workers,[14] who found that DPT could be precipitated from an aqueous solution of MDNA, partly neutralised with ammonia solution, by addition of 38% aqueous formaldehyde. After 24 hours a precipitate of DPT was recovered. The same reaction attempted in a non-aqueous solvent of ether[15] yielded a gummy product, assigned structure (4.3), and most probably formed by a series of condensation reactions, as outlined in Scheme 4.1.

(4.2) (4.3)
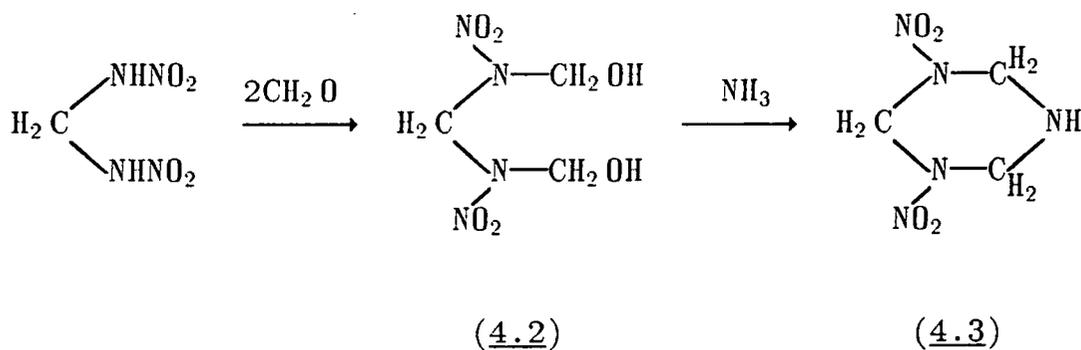
Scheme 4.1

The ability of MDNA to form the product of dimethyl-olation, ($\underline{4.2}$), was demonstrated.[15,28] These observations suggest that the formation of DPT from MDNA is dependent upon it undergoing some form of decomposition in aqueous solution, since in DPT the $N-NO_2$ functions are separated by CH_2-N-CH_2 groups. Chapman and co-workers[15] formed ($\underline{4.4}$) from MDNA, dry gaseous formaldehyde and alcoholic methylamine.
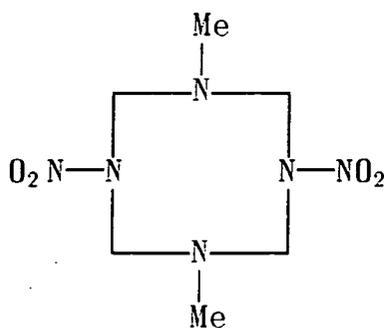
($\underline{4.4}$)

This product is identical with that formed by Wright and co-workers[2] from nitramide (NH_2NO_2), 37% aqueous formaldehyde and 10% aqueous methylamine. No MDNA was regenerated on hydrolysis. Thus the ability of MDNA to act as a precursor of nitramide is demonstrated. However, when Chapman attempted to form ($\underline{4.4}$) from MDNA, 40% aqueous formaldehyde, and methylamine the product ($\underline{4.5}$) was obtained which regenerated MDNA on hydrolysis.

$$O_2N-N \begin{array}{c} CH_2 \\ | \\ N-Me \\ | \\ CH_2 \end{array} N-NO_2,\ H_2C$$
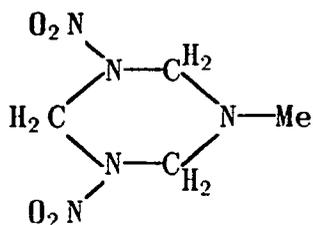
(4.5)

This result appears to contradict the earlier observation that DPT was formed when ammonia was used in place of methylamine. The explanation for this seems to lie in the fact that Lamberton and co-workers formed DPT by addition of 38% aqueous formaldehyde to an aqueous solution of MDNA already partially neutralised with ammonia, whereas Chapman and co-workers added methylamine to an aqueous solution of MDNA in 40% aqueous formaldehyde. Thus in the former case MDNA decomposes before addition of formaldehyde, but in the latter case it undergoes dimethylolation to form (4.2) before addition of methylamine. Work has already been done on the decomposition of MDNA in aqueous solution,[14] and this is re-investigated in Section 4.4.1.

It has been observed[16] that addition of MDNA to the Schiessler-Ross reaction, which uses paraformaldehyde and ammonium nitrate in acetic anhydride to form RDX, increases the yield. It is also shown that MDNA can be formed in small quantities in the reaction mixture employed, but not under the conditions suitable for RDX formation, indicating that MDNA is not necessarily an intermediate in the normal nitration reaction.

4.4.1 The decomposition of methylenedinitroamine in aqueous solution

It has been shown[14] that on heating in water MDNA ultimately yields formaldehyde, nitrous oxide and water, thus:

$$CH_2(NHNO_2)_2 \longrightarrow CH_2O + 2N_2O + H_2O \qquad \text{Eqn. (4.1)}$$
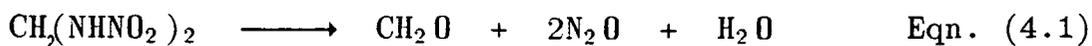
Equation (4.1) shows only the overall decomposition, and does not necessarily provide a true measure of the primary stage of decomposition, since intermediates between MDNA and the products shown are conceivable. Lamberton and co-workers[14] found MDNA to decompose in 11M-mineral acid and in 2M-sodium hydroxide solution, and at pH 3-8. Considerable stability was exhibited by MDNA at pH 1 and pH 10. The decomposition at pH 3-8 was found to lead to nitrous oxide and formaldehyde, as shown above, the primary stage of this decomposition appearing to result in the liberation of nitramide and conversion of the methylene group of MDNA to a form oxidisable by alkaline hypoiodite. The decomposition was auto-catalytic, and it was suggested that it proceeds via the mono-anion (4.6) thus:

(4.1)   (4.6)
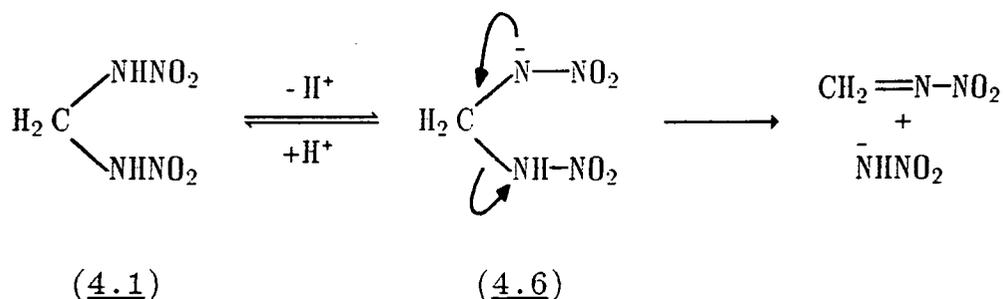
Scheme 4.2

The unionised species (4.1) present in pH 1 media, and the di-anion (4.7) present in pH 10 media would be stable, since one nitroamine group cannot facilitate the ejection of the other. The di-anion is likely to undergo tautomerisation in order to increase the charge separation between the N-atoms nearest the methylene group, as shown in Scheme 4.3.

(4.7)
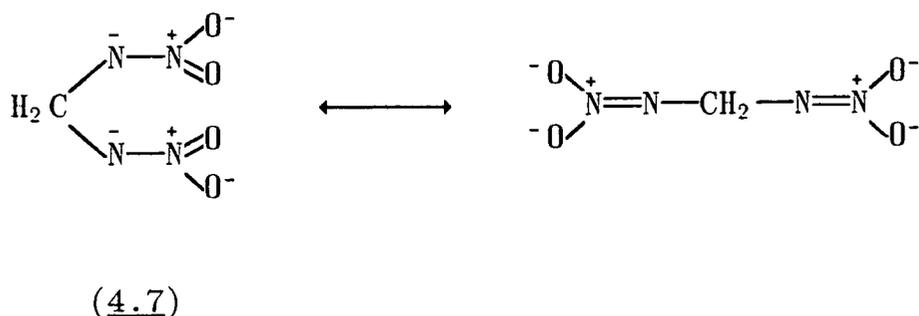
Scheme 4.3

The decomposition was found to be fastest at pH 5.4, this result being consistent with the calculated maximum concentration of the mono-anion species occurring at pH 5.8, based on the determined pK_a values of 5.0 and 6.6.[17]

4.4.2 UV spectra of methylenedinitroamine

A sample of MDNA was prepared by the method of Brian and Lamberton,[18] as described in Chapter 2.

The UV spectrum of MDNA in acetonitrile gave λ_{max} = 218 nm, ϵ_{max} = 12850 1 mol^{-1} cm^{-1}. Initial investigations of the UV spectra of MDNA in hydrochloric acid and aqueous sodium hydroxide solutions yielded the results given in Table 4.4 and Table 4.5 respectively.

Table 4.4 UV data for MDNA $(2.0 \times 10^{-5}$ M) in hydrochloric acid at 25°C

[HCl]\mol 1^{-1}	λ_{max}\nm	ϵ_{max}\1 mol^{-1} cm^{-1}
0.12	225	11900
1.15	226	10750
2.87	225	10000
5.73	225	10750
11.46	225	10250

Table 4.5 UV data for MDNA $(2.0 \times 10^{-5}$ M) in aqueous sodium hydroxide solution at 25°C

[NaOH]\mol 1^{-1}	λ_{max}\nm	ϵ_{max}\1 mol^{-1} cm^{-1}
0.1	233	13250
0.5	233	11850
1.0	233	13550
2.5	234	12500
5.0	226	27500

The results show MDNA to be stable in a range of acid solutions and basic solutions, although some change in the spectrum is seen in 5M-sodium hydroxide solution. Initial investigations of the stability of MDNA in aqueous solutions covering a wide pH range gave the results shown in Table 4.6.

Table 4.6 UV data for MDNA (2.0×10^{-5} M) in aqueous
solutions covering pH 0-14 at 25°C

| Medium | pH | Fresh solution | | After 48 hours | |
		λ_{max} \nm	ϵ_{max} \l mol^{-1}cm^{-1}	λ_{max} \nm	ϵ_{max} \l mol^{-1}cm^{-1}
1M HCl	0	225	11400	228	11750
10^{-2}M HCl	2	226	10000	229	12100
acetate buffer	3.7	227	13000	229	12000
acetate buffer	5.2	231	11250	-	-
phospate buffer	6.0	233	13500	-	-
tris buffer	7.4	231	15000	228	2650
borax buffer	8.0	233	16500	232	14350
borax buffer	9.7	233	16400	233	15800
10^{-2}M NaOH	12	233	16150	233	16350
1M NaOH	14	233	15150	-	-

The results agree with those of Lamberton and co-workers,[14] the most extensive decomposition occurring in the region pH 5-7, and in 1M-sodium hydroxide solution.

4.4.3 Quantitative determination of formaldehyde liberation during the decomposition of methylenedinitroamine

In order to ascertain that MDNA, upon decomposition, produces formaldehyde in a 1:1 molar ratio, in accordance with the stoichiometry of Equation (4.1), 0.25 g (1.84 mmol) MDNA was put in 10 cm^3 of pH 6.0 citrate buffer, and was left at 20^oC for 24 hours. The pH of the solution after this time was measured as 4.3, and bubbles of gas were seen to have formed in the solution. Addition of an excess of saturated alcoholic dimedone solution, as outlined by Vogel,[19] and described in Chapter 2, precipitated 0.51 g (1.75 mmol) of the pale orange-brown derivative. Within experimental error this is in a 1:1 molar ratio with the MDNA added, thus confirming that the methylene group of MDNA is wholly converted to formaldehyde during decomposition. The other product of this primary stage of decomposition is therefore a potential 2 equivalents of nitramide, which is known[20-27] to undergo decomposition to form nitrous oxide and water only.

4.4.4 Determination of a pH profile for the decomposition of methylenedinitroamine

Preliminary kinetic analysis of the decomposition of MDNA in various aqueous buffers revealed that there is no catalysis by components of the buffer, as can be seen from the specimen results given in Table 4.7. The observed rate coefficients for the decomposition in a range of buffers are given in Table 4.8.

Table 4.7 Specimen results to show the absence of catalysis by buffer components on the decomposition of MDNA

Buffer	Concentration	pH	$k_{obs} \backslash sec^{-1}$
citrate	0.06 M	6.1	3.9×10^{-5}
citrate	0.6 M	6.1	4.2×10^{-5}
tris/trisHCl	0.01 M	7.0	1.5×10^{-5}
tris/trisHCl	0.02 M	7.0	1.3×10^{-5}
tris/trisHCl	0.01 M	7.33	7.0×10^{-5}
tris/trisHCl	0.02 M	7.33	7.9×10^{-5}
tris/trisHCl	0.01 M	7.66	3.7×10^{-6}
tris/trisHCl	0.02 M	7.66	5.2×10^{-6}

Table 4.8 Observed rate coefficients for the decomposition of MDNA (4.0×10^{-5} M) in aqueous buffer solution at 25°C

Buffer medium	pH	$k_{obs} \backslash sec^{-1}$
0.6M citrate	3.5	2.5×10^{-5}
0.6M citrate	4.5	3.2×10^{-5}
0.6M citrate	5.0	5.0×10^{-5}
0.6M citrate	5.5	6.2×10^{-5}
0.6M citrate	6.0	4.4×10^{-5}
0.6M citrate	6.5	3.6×10^{-5}
0.02M tris	7.0	1.3×10^{-5}
0.02M tris	7.3	7.9×10^{-6}
0.02M tris	7.7	5.2×10^{-6}
0.025M borax	8.0	2.0×10^{-6}
0.025M borax	8.3	7.2×10^{-7}
0.025M borax	8.7	2.9×10^{-7}
0.025M borax	9.0	$< 2 \times 10^{-7}$
0.0001M NaOH	10.0	3.1×10^{-7}
0.001M NaOH	11.0	4.5×10^{-7}
0.01M NaOH	12.0	3.8×10^{-7}
0.1M NaOH	13.0	1.2×10^{-6}
1M NaOH	14.0	3.7×10^{-5}

On the basis of the results, it would seem that the maximum concentration of the monoanion (<u>4.6</u>) occurs at approximately pH 5.5, which is in agreement with the previously reported[14] figure of 5.4, both being slightly lower than the calculated value of 5.8, based on the quoted[14] pK_a values. The pK_a values for MDNA were calculated independently, as described in Section 4.4.5, as 4.77 and 6.39, giving a maximum concentration of the mono-anion (<u>4.6</u>) at pH 5.58.

Figure 4.2 shows the pH profile for the decomposition of MDNA based on the results in Table 4.8, superimposed on the theoretical profile calculated using Equation (4.2), derived in Scheme 4.4, substituting $k = 8 \times 10 \, \text{sec}^{-1}$, the appropriate value of $[H^+]$, and the new pK_a values of 4.77 and 6.39.

$$O_2N.HNCH_2NH.NO_2 \xrightleftharpoons{K_{a,1}} O_2N.HNCH_2\bar{N}.NO_2 \xrightleftharpoons{K_{a,2}} O_2N.\bar{N}CH_2\bar{N}.NO_2$$
$$+ H^+ \qquad\qquad + 2H^+$$
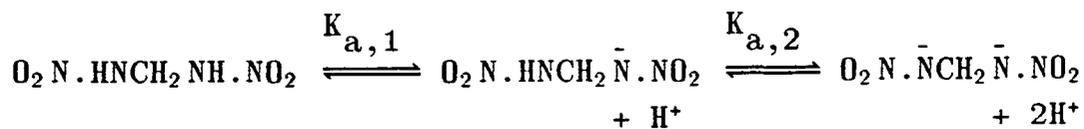
$$\underline{NH_2} \qquad\qquad\qquad \underline{NH^-} \qquad\qquad\qquad \underline{N^=}$$

$$K_{a,1} = \frac{[NH^-][H^+]}{[NH_2]}$$

$$K_{a,2} = \frac{[N^=][H^+]}{[NH^-]}$$

$$[N]_{stoich} = [NH_2] + [NH^-] + [N^=]$$

$$= [NH^-]\left[1 + \frac{[H^+]}{K_{a,1}} + \frac{K_{a.2}}{[H^+]}\right]$$

Assuming $\quad$ rate $= k[NH^-]$

$$= k_{obs}[N]_{stoich}$$

Where $\quad k_{obs} = k\,\frac{[NH^-]}{[N]_{stoich}}$

$$= \frac{k}{\left[1 + \frac{[H^+]}{K_{a,1}} + \frac{K_{a.2}}{[H^+]}\right]} \qquad \text{Eqn. (4.2)}$$

$$\underline{\text{Scheme 4.4}}$$

Figure 4.2 pH dependence of k_{obs} for decomposition of MDNA in aqueous solution at $25^{\circ}C$. The experimental points are shown as (0), and the solid line is calculated from Equation (4.2) with $k = 8 \times 10^{-5}$ sec^{-1}, $pK_{a,1} = 4.77$, $pK_{a,2} = 6.39$
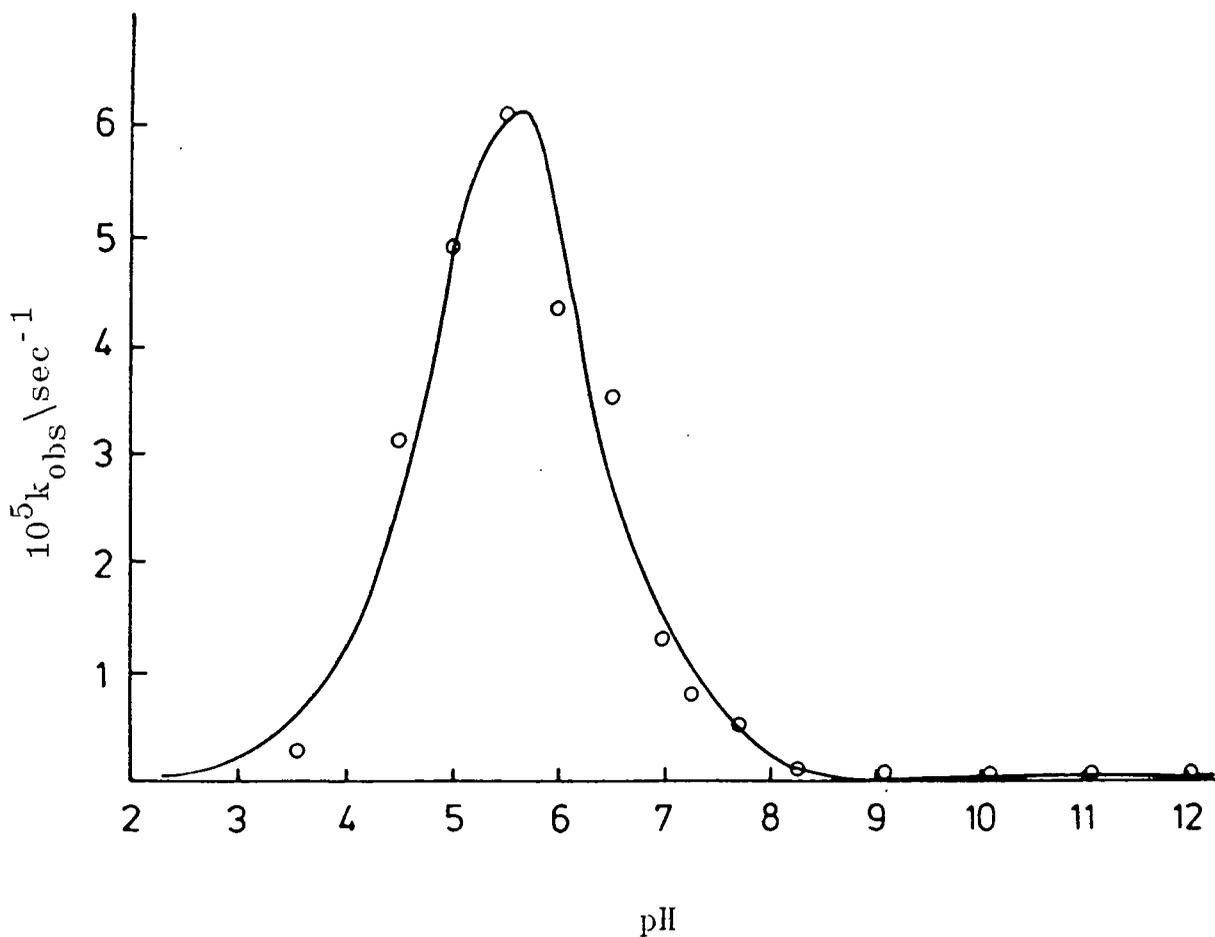
The pH profile is in agreement with the results of Lamberton and co-workers.[14] The experimental pH of maximum decomposition of 5.5 is in good agreement with the calculated value of 5.58. The profile is plotted as $\log_{10}k_{obs}$ vs. pH in Figure 4.3, and is not the same as the previously[11,13] determined profile for the decomposition of the intermediate formed by the primary decomposition of DPT in base, the maximum rate of decomposition in this case occurring at pH 9.5. Neither does the profile show any similarity with that measured for the decomposition of the intermediate species formed by the primary decomposition of DPT in acid solution (Figure 4.1). Also, the catalysis by buffer components seen by the author for the measurements in acid media, and by Cooney[11,13] in basic media, were not observed for the decomposition of MDNA. Thus MDNA is not the primary decomposition product of DPT in either acid or basic solution.

Figure 4.3   A plot of $\log_{10} k_{obs}$ vs. pH for the decomposition of MDNA at $25^{\circ}C$
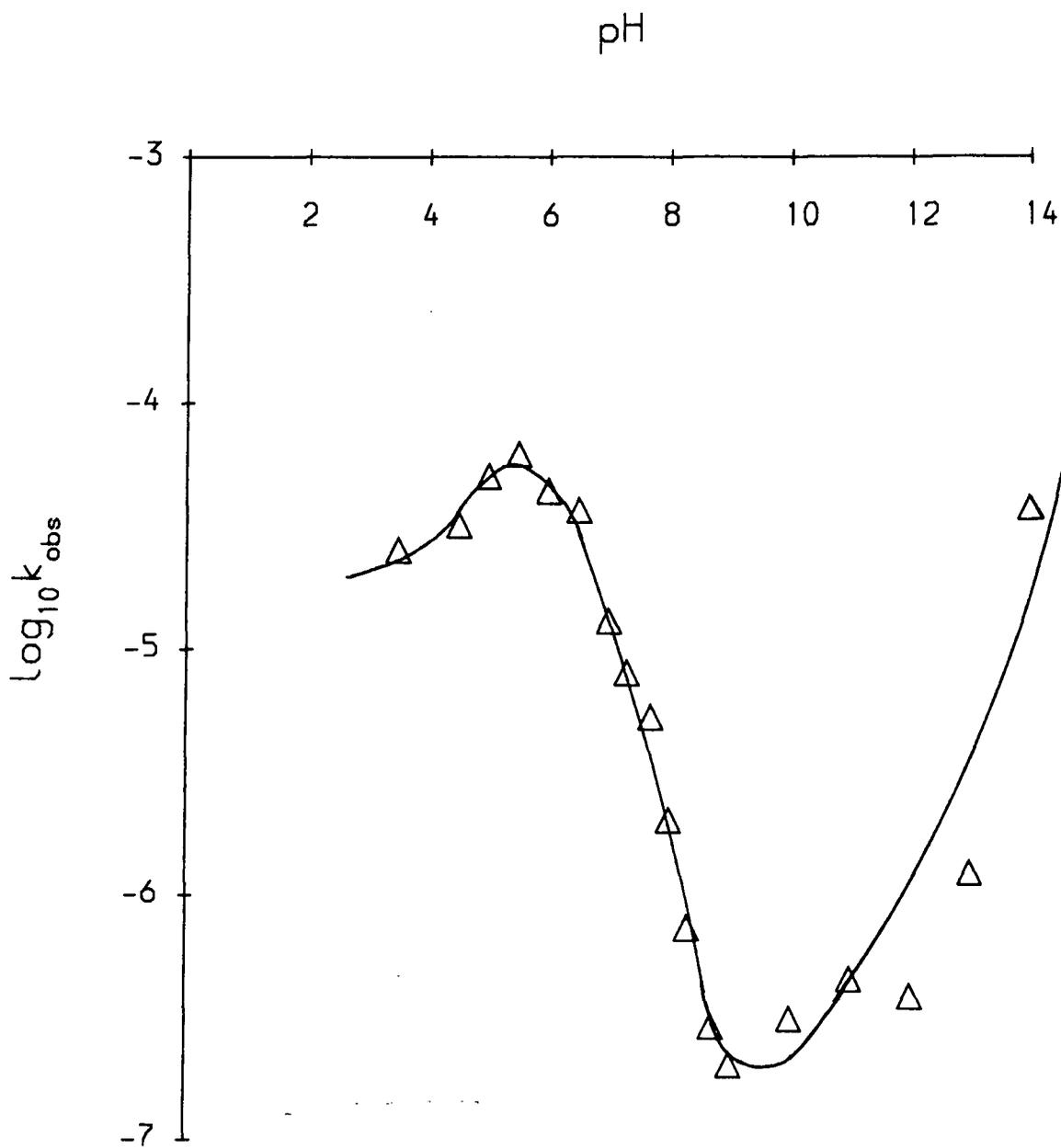
4.4.5 Determination of the pK$_a$ values of methylenedinitroamine

0.05M-sodium hydroxide solution was added to 50 cm^3 freshly prepared 0.005M aqueous MDNA solution in 0.5 cm^3 portions. The pH of the solution was measured after each addition. A total of 12 cm^3 of sodium hydroxide solution was added, this being in excess of the 10 cm^3 required for complete ionisation of the MDNA. Details of the variation of pH with base added are shown in Table 4.9.

Table 4.9 Variation of pH on addition of 0.05M-sodium hydroxide solution to 50 cm^3 0.005M aqueous methylenedinitroamine at 20°C

0.05M NaOH added \cm^3	Run 1 pH	Run 2 pH	Run 3 pH
0.0	3.52	3.55	3.57
0.5	3.85	3.94	3.93
1.0	4.15	4.19	4.21
1.5	4.37	4.40	4.41
2.0	4.56	4.58	4.60
2.5	4.73	4.74	4.76
3.0	4.90	4.89	4.91
3.5	5.06	5.06	5.07
4.0	5.23	5.23	5.23
4.5	5.40	5.40	5.42
5.0	5.59	5.58	5.60
5.5	5.78	5.76	5.78
6.0	5.95	5.93	5.94
6.5	6.13	6.10	6.11
7.0	6.25	6.25	6.26
7.5	6.43	6.40	6.41
8.0	6.59	6.57	6.58
8.5	6.78	6.74	6.74
9.0	6.97	6.95	6.96
9.5	7.24	7.27	7.28
10.0	7.86	8.02	7.94
10.5	10.09	10.27	10.25
11.0	10.68	10.75	10.76
11.5	10.93	10.98	10.97
12.0	11.08	11.12	11.13

$$H_2C \overset{\displaystyle NHNO_2}{\underset{\displaystyle NHNO_2}{<}} \quad \overset{K_{a,1}}{\rightleftharpoons} \quad H_2C \overset{\displaystyle \overset{-}{N}-NO_2}{\underset{\displaystyle NHNO_2}{<}} \quad \overset{K_{a,2}}{\rightleftharpoons} \quad H_2C \overset{\displaystyle \overset{-}{N}-NO_2}{\underset{\displaystyle \underset{-}{N}-NO_2}{<}}$$

$$(\underline{4.1}) \qquad\qquad\qquad (\underline{4.6}) \qquad\qquad\qquad (\underline{4.7})$$
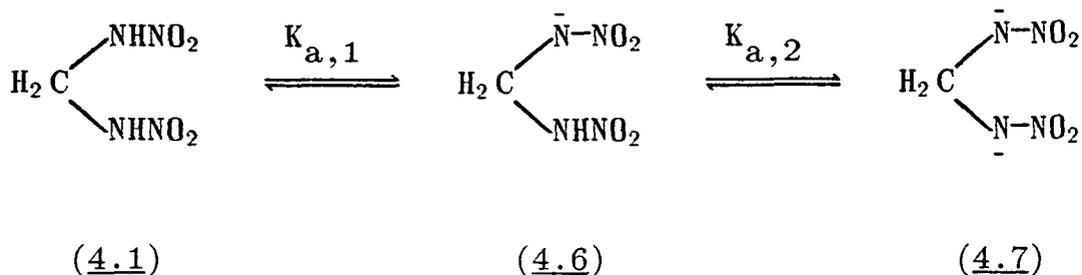
<u>Scheme 4.5</u>

It was found that the ionisation processes, shown in Scheme 4.5, overlap, since there is no sharp end-point for the addition of the first equivalent of base. This is due to titration of the second ionising group beginning before the first is complete as a consequence of ΔpK being small. This can be seen on the titration curve shown in Figure 4.4. The pK_a values were separated using the calculation method of Noyes[29] as detailed in Reference 30. This gave the results shown in Table 4.10.

<u>Table 4.10</u>   The pK_a values of methylenedinitroamine at 20°C

Run no.	$pK_{a,1}$	$pK_{a,2}$
1	4.75	6.40
2	4.77	6.37
3	4.78	6.38
	Mean = 4.77 ± 0.02	Mean = 6.39 ± 0.02

On the basis of these results, the maximum concentration of the mono-anion ($\underline{4.6}$) occurs at pH 5.58.

Figure 4.4 A plot of pH vs. molar ratio base:acid for aqueous MDNA titrated aqainst aqueous sodium hydroxide
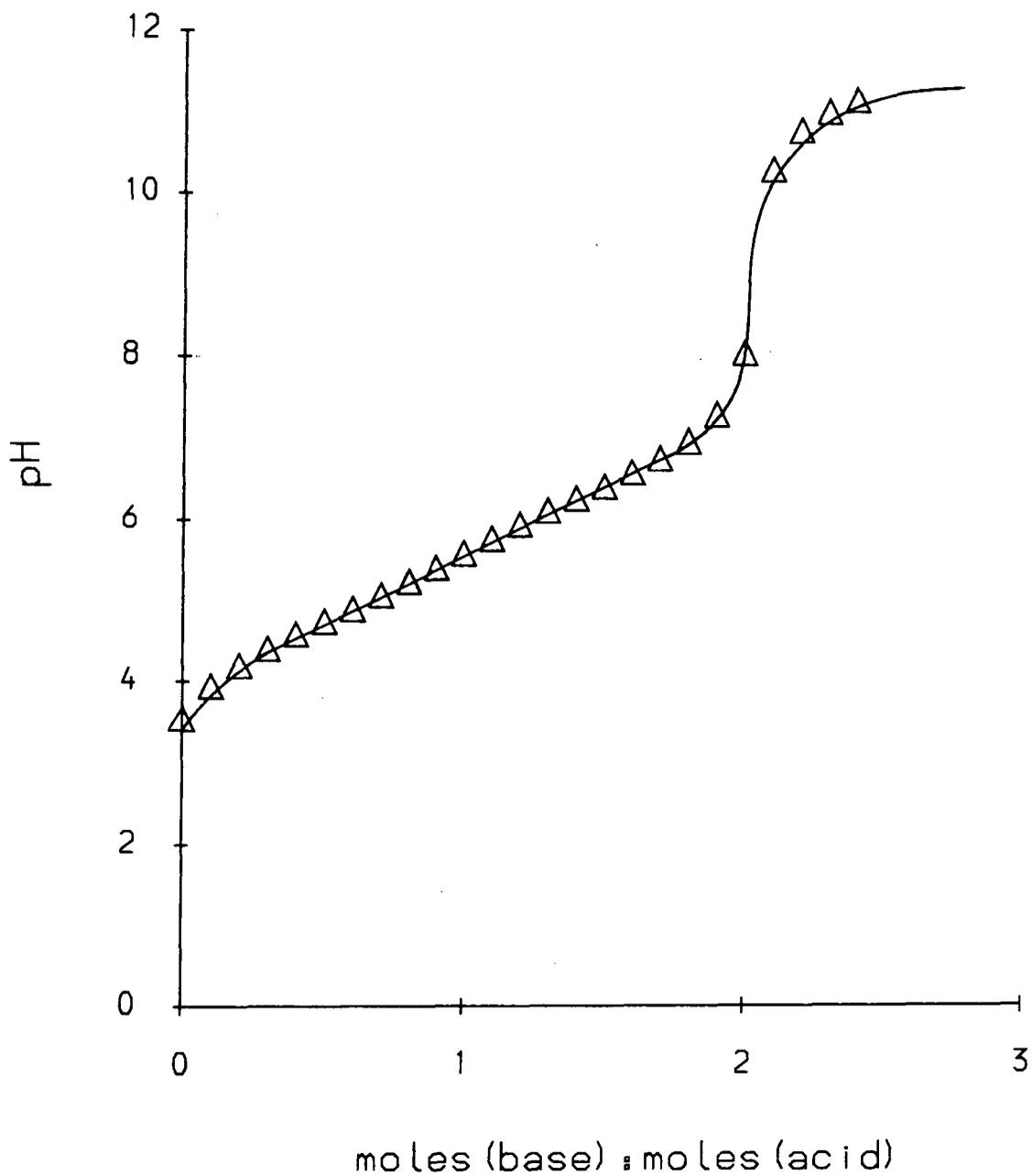
4.5 The role of nitramide

Wright and co-workers[31] found that if a solution of nitramide (NH_2NO_2) in water was treated with formaldehyde and neutralised with ammonia, a precipitate of DPT appeared over a five minute period. In light of this, and the fact that MDNA has been observed to act as a precursor of nitramide in the formation of DPT and related structures, it was decided to investigate this compound further with respect to its UV spectra and decomposition in aqueous solution.

A sample of nitramide was prepared by the method of Marlies, La Mer and Greenspan[32] as described in Chapter 2.

4.5.1 UV spectra of nitramide

An investigation of the UV spectra of nitramide (kept as a stock solution of 10^{-3}M nitramide in 0.01M-hydrochloric acid at 0°C, in which it was stable for several hours) in various aqueous media yielded the results shown in Table 4.11.

Table 4.11 UV data for nitramide (4.8×10^{-5} M) in aqueous solution at 25°C

Medium	pH	$\lambda_{max} \backslash nm$	$\epsilon_{max} \backslash l \ mol^{-1} \ cm^{-1}$
0.01M HCl	2.0	210	7550
0.001M HCl	3.0	210	7630
acetate buffer	5.0	220	6860
tris buffer	7.0	225	7480
tris buffer	9.0	225	6860
0.001M NaOH	11.0	225	7280
0.01M NaOH	12.0	225	7480
0.1M NaOH	13.0	225	7690

Note that λ_{max} undergoes a hypsochromic shift in acid solution from its value in neutral solution, to a value very close to that observed for the primary decomposition product of DPT in acidic solution. The values of ϵ_{max} are close to the value of 7000 $l \ mol^{-1} \ cm^{-1}$ predicted earlier for primary nitroamine groups, and at approximately 7500 $l \ mol^{-1} \ cm^{-1}$ are roughly half those of an acidic solution of DPT of the same concentration, presenting the possibility that in acid solution two equivalents of nitramide are formed from one equivalent of DPT.

4.5.2 Kinetic analysis of the decomposition of nitramide in aqueous solution

It has been stated earlier that nitramide decomposes to nitrous oxide and water in aqueous solution. The observed rate coefficients for this process over a range of pH are given in Table 4.12. In some cases catalysis by buffer components was observed and k_{obs} has been extrapolated back to zero buffer concentration.

Table 4.12   Observed rate coefficients for the decomposition of nitramide (4.8×10^{-5} M) in aqueous solution at 25°C

Medium	pH	wavelength decomposition measured at \nm	k_{obs}\sec^{-1}
0.01M HCl	2.0	210	6.0×10^{-5}
0.001M HCl	3.0	210	4.6×10^{-5}
acetate buffer	5.0	220	extrapolates to zero
tris buffer	7.0	225	extrapolates to zero
tris buffer	9.0	225	extrapolates to zero
0.001M NaOH	11.0	225	4.0×10^{-3}
0.01M NaOH	12.0	225	5.0×10^{-2}
0.1M NaOH	13.0	225	0.55

A plot of $\log_{10}k_{obs}$ vs. pH is shown in Figure 4.5. This shows no similarity with the comparable plot for the decomposition of MDNA (Figure 4.3).

Figure 4.5 A plot of $\log_{10}k_{obs}$ vs. pH for the decomposition of nitramide at 25°C
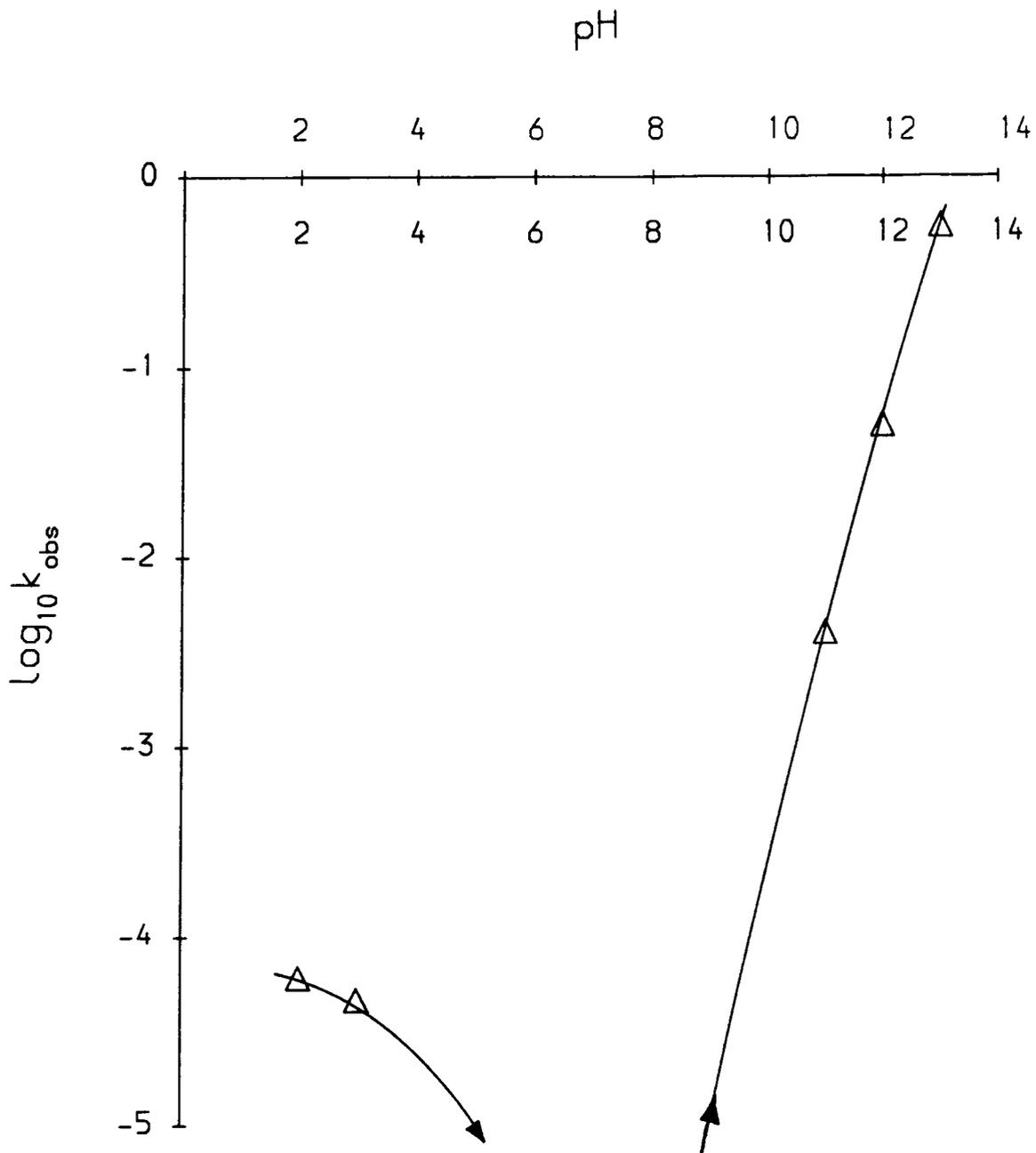
The increase in the rate of decomposition with increasing base concentration is particularly interesting. The generally accepted mechanism for nitramide decomposition involves isomerisation of structure (4.8) to its *aci*-nitro form (4.9), followed by elimination of H$^+$ and HO$^-$ from this species by attack of base on the N—H proton, as shown in Equations (4.3) and (4.4).

Eqn. (4.3)

(4.8)   (4.9)

$$B H^+ \; + \; NNO \; + \; HO^-$$

Eqn. (4.4)
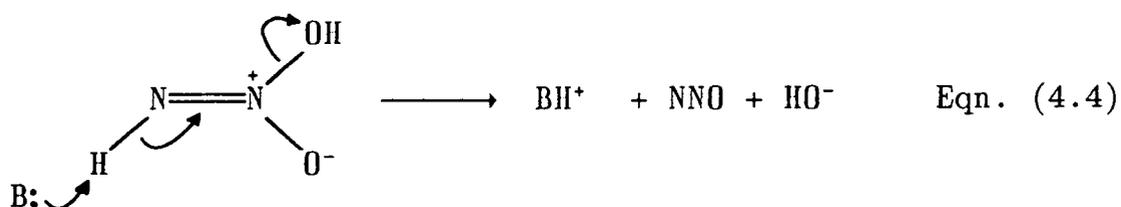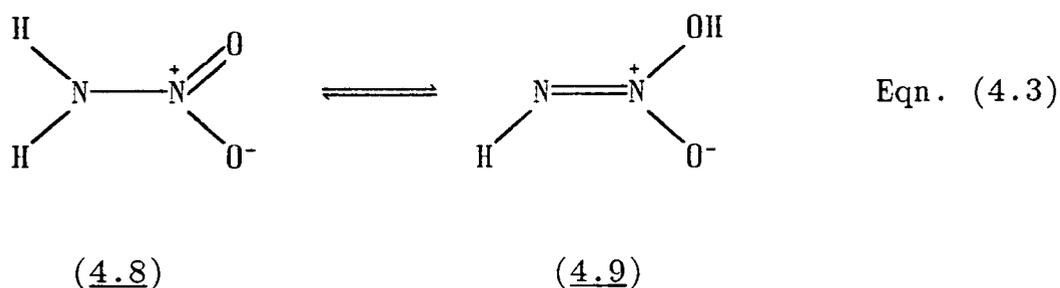
This mechanism requires the rate of reaction to be independent of [HO$^-$] in solution where nitramide, for which $pK_a = 6.55$,[20] is essentially completely ionised. The observation that k_{obs} increases in direct proportion with [HO$^-$] is in agreement with the work of Kresge and Tang.[20] They also found that in strongly basic buffers the system shows simple general base catalysis, and not the specific hydrogen ion-general base catalysis predicted by Equations (4.3) and (4.4). These results suggest two parallel decomposition pathways as outlined in Scheme 4.6.

Namely: (i) the traditional general base catalysed reaction of *aci*-nitramide, and (ii) the general base catalysed decomposition of the nitramide anion.

$$NH_2NO_2 \rightleftharpoons NHNO_2H \qquad\qquad \rightleftharpoons \qquad H^+ \;+\; NHNO_2^-$$

$$\text{(i)} \quad \underset{\text{Base}}{\big|\!\longrightarrow} \quad N_2O \;+\; H_2O \quad \underset{\text{Base}}{\longleftarrow\!\big|} \quad \text{(ii)}$$
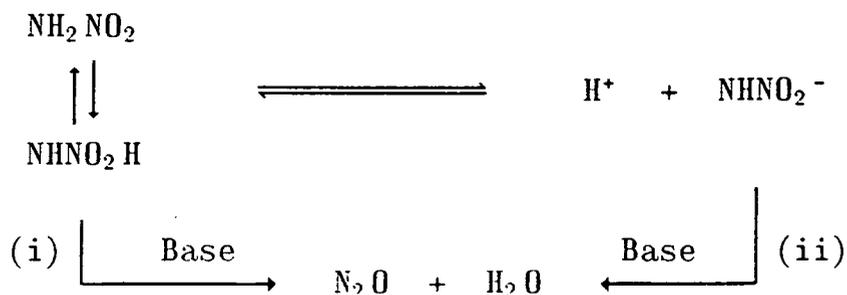
<u>Scheme 4.6</u>

It was noticed that the decomposition of nitramide in 0.01M-hydrochloric acid gave $k_{obs} = 6.0 \times 10^{-5}$ sec^{-1}, which is in agreement with an earlier[22] value of 5.0×10^{-5} sec^{-1}, and is very similar to the rate of the secondary decomposition of DPT in 0.01M-hydrochloric acid, giving $k_{obs} = 5.0 \times 10^{-5}$ sec^{-1}. Indeed, if the results (UV spectra and rate coefficients) for the decomposition of the intermediate formed from the primary decomposition of DPT in acid solution are compared with those for nitramide over the same pH range, as shown in Table 4.13, it is seen that they correlate remarkably well. This provides good evidence that the intermediate formed in acid solution is nitramide itself.

Table 4.13 Comparison of the UV spectra and rate coefficients for decomposition of nitramide and of intermediate in acid solution at $25^{\circ}C$

Medium	Nitramide		Intermediate[a]	
	λ_{max}\nm k_{obs}\sec^{-1}		λ_{max}\nm k_{obs}\sec^{-1}	
HCl (0.01 M)	210	6.0×10^{-5}	210	5.0×10^{-5}
HCl (0.001 M)	210	4.6×10^{-5}	210	4.3×10^{-5}
acetate[b] (pH 5)	220	1.8×10^{-3}	220	1.5×10^{-3}
tris[b] (pH 7)	225	9.0×10^{-3}	225	8.6×10^{-3}
tris[b] (pH 9)	225	1.4×10^{-2}	225	1.2×10^{-2}
NaOH (0.001 M)	225	4.0×10^{-3}	225	4.0×10^{-3}
NaOH (0.01 M)	225	5.0×10^{-2}	225	5.0×10^{-2}
NaOH (0.1 M)	225	0.55	225	0.53

a.   Prepared from reaction of DPT with 0.01M-HCl.

b.   Strong buffer catalysis is observed, values quoted for nitramide and for intermediate refer to solutions with precisely the same buffer composition. If extrapolated back to zero buffer concentration the values of k_{obs} are found to be zero.

4.6 Discussion

Wright and co-workers[2] proposed that DPT was produced in the Hale reaction dilution liquors from dimethylolnitramide (4.10) (the formaldehyde-addition product of nitramide).

$$O_2N \underset{\underset{CH_2OH}{|}}{\overset{\overset{CH_2OH}{|}}{N}}$$
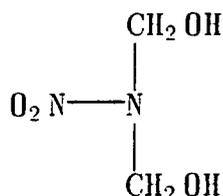
(4.10) dimethylolnitramide

They state that the precursor of DPT cannot be nitramide itself, based on a belief that nitramide is 'notoriously' unstable in acidic solution, but as the results in Section 4.5.2 show, the decomposition of nitramide in acid solution is very slow, having a half-life of ca. 3 hours. In addition to producing DPT itself from nitramide, formaldehyde and ammonia, as stated earlier, they also produced DPT from nitramide, formaldehyde and methylene-diamine, a compound of structure (4.11) from nitramide, formaldehyde and ethylenediamine, and the compound (4.4), seen earlier, from nitramide, formaldehyde and aqueous methylamine.

$$O_2N-N \underset{\underset{N}{\underset{|}{CH_2}}}{\overset{\overset{N}{\underset{|}{CH_2}}}{}} N-NO_2$$
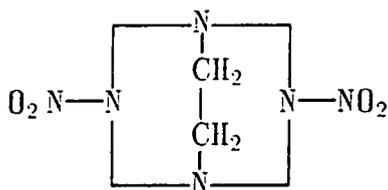
(4.11)

They gave evidence for the possible existence of dimethylolnitramide in an aqueous solution of paraformaldehyde and nitramide.

Lamberton and co-workers[33] suggest that compounds of the type ($\underline{4.12}$) behave predominantly as mixtures of nitramide (($\underline{4.13}$), R = H) and formaldehyde, through the mechanism shown in Equation (4.6), but have greater stability in mineral acids.

$$NO_2 NRCH_2 OH \rightleftharpoons NHRNO_2 + CH_2 O \qquad Eqn. \ (4.5)$$

$$\qquad (\underline{4.12}) \qquad\qquad (\underline{4.13})$$

$$\qquad Eqn. \ (4.6)$$
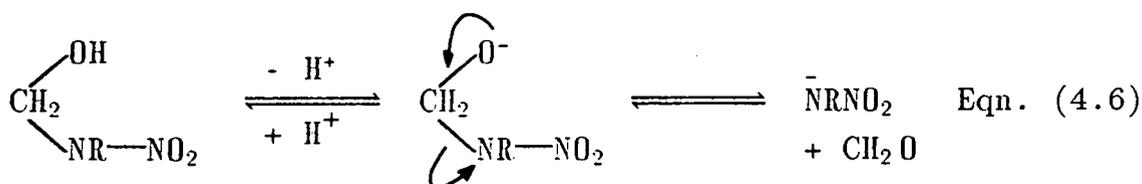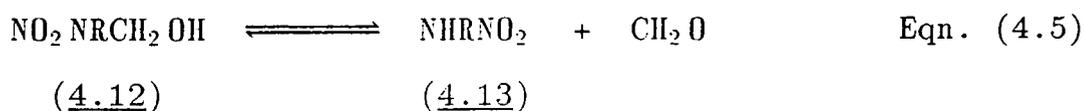
They propose that compounds of type ($\underline{4.12}$) are not precursors of type ($\underline{4.16}$), but that the reaction probably proceeds via decomposition of ($\underline{4.12}$) as shown in Equation (4.5), condensation of liberated formaldehyde with an amine ($\underline{4.14}$), Equation (4.7), to form ($\underline{4.15}$), and union thereof with nitramide, Equation (4.8).

$$CH_2 O + NHR_2 \rightleftharpoons NR_2 CH_2 OH \qquad Eqn. \ (4.7)$$

$$\qquad (\underline{4.14}) \qquad\qquad (\underline{4.15})$$

$$NR_2 CH_2 OH + NHRNO_2 \rightleftharpoons NO_2 NRCH_2 NR_2 + H_2 O \qquad Eqn. \ (4.8)$$

$$\qquad\qquad\qquad\qquad (\underline{4.16})$$
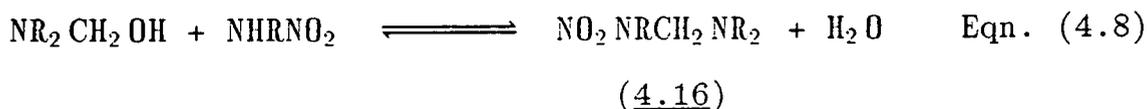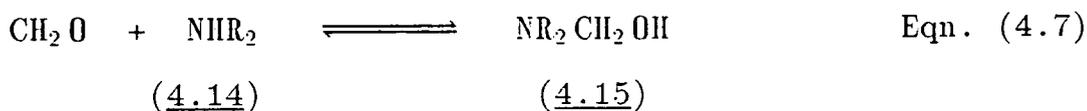
Myers and Wright,[34] however, favour direct condensation of
(4.12) and (4.14) to form (4.16). Lamberton proposes the
decomposition of (4.16) in water to yield formaldehyde and
salts of the nitramide anion with the amino-cation. In
alkali (4.16) is suggested to yield nitramide anion and an
equilibrium mixture of amine, formaldehyde and
hydroxymethylamine, and in acid to be stabilised, possibly
by formation of $NO_2NRCH_2\overset{+}{N}HR_2$ or $\overset{+}{N}HR_2CH_2OH$.

This information, coupled with the results obtained,
raises the possibility that DPT is formed, upon
neutralisation of the diluted nitration mixture, from
condensation of free nitramide and formaldehyde to form
dimethylolnitramide, and then the subsequent condensation of
2 equivalents of dimethylolnitramide with 1 equivalent of
methylenediamine formed from formaldehyde and ammonia, or
condensation of 2 equivalents of dimethylolnitramide with 2
equivalents of ammonia to form the 8-membered unbridged
structure (4.17), which then condenses with 1 equivalent of
formaldehyde to form DPT. This is summarised in Scheme 4.7.

$$NH_2NO_2 \quad + \quad 2\ CH_2O \quad \rightleftharpoons \quad O_2N-N(CH_2OH)_2$$

$$O_2N-N(CH_2OH)_2 \qquad (HOCH_2)_2N-NO_2$$

$$\Big\downarrow 2NH_3 \qquad CH_2(NH_2)_2 \rightleftharpoons CH_2O + 2NH_3 \Big\downarrow$$

(4.17) ring structure O_2N-N ... $N-NO_2$ with two NH bridges

$$\xrightarrow{\quad CH_2O \quad}$$

ring structure with O_2N-N ... CH_2 ... $N-NO_2$ and two N bridges
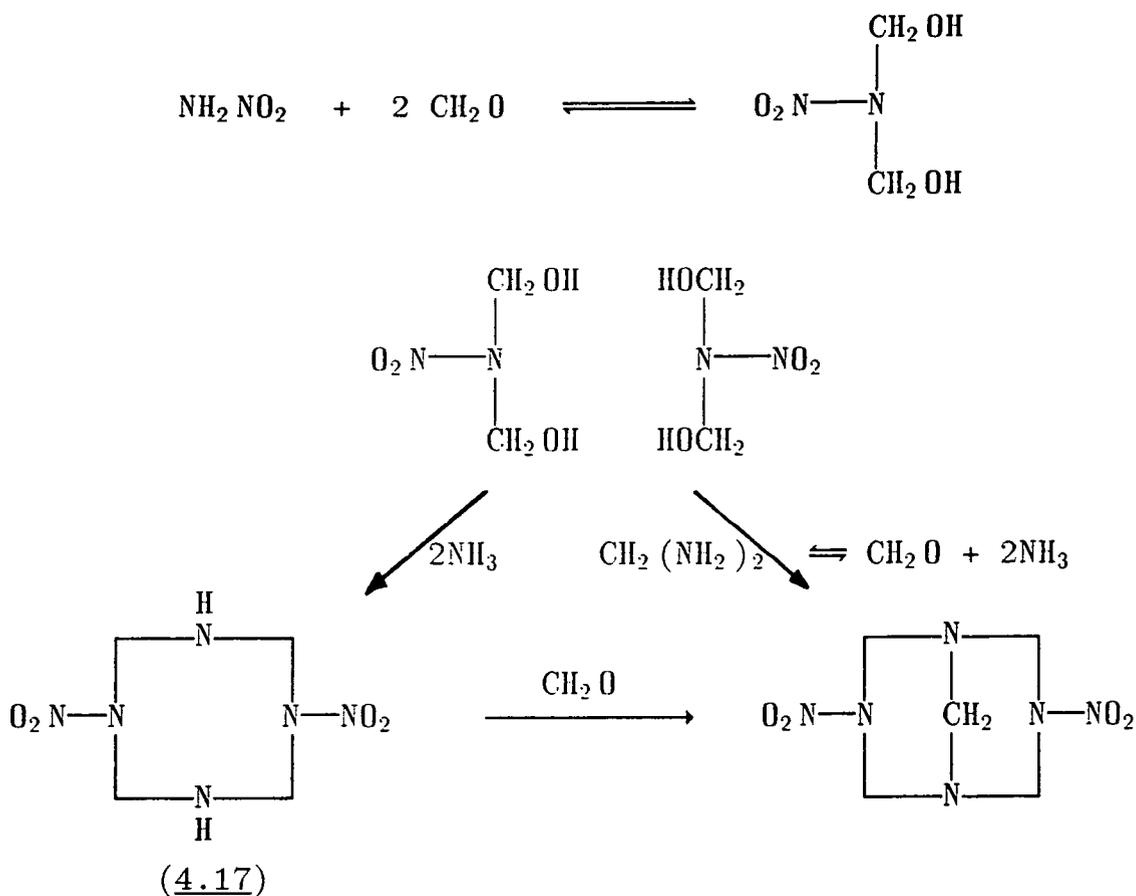
Scheme 4.7

The fact that DPT could still be precipitated when triethylamine was used to neutralise the diluted nitration mixture (see Chapter 3. Section 3.3) suggests that ammonia, or a formaldehyde condensation product thereof, must be formed from the initial degradation of the hexamine, and the molecule does not undergo total N-nitration. The mechanism of the initial degradation is not yet known.

Cooney[11] suggested that initial cleavage of C—N bonds in DPT could lead to a primary product which exists as (4.18) in alkaline media, (4.20) in acidic media, and (4.19) at intermediate pH.

$$O_2N\!-\!\overset{-}{N}\!-\!CH_2\!-\!\overset{\overset{\displaystyle R}{\displaystyle |}}{N}\!-\!CH_2\!-\!\overset{-}{N}\!-\!NO_2 \qquad (\underline{4.18})$$

$$O_2N\!-\!\overset{-}{N}\!-\!CH_2\!-\!\overset{\overset{\displaystyle R}{\displaystyle |}}{\underset{\displaystyle |}{\overset{+}{N}}}\!-\!CH_2\!-\!\overset{-}{N}\!-\!NO_2 \qquad (\underline{4.19})$$
$$\qquad\qquad\qquad\qquad H$$

$$O_2N\!-\!HN\!-\!CH_2\!-\!\overset{\overset{\displaystyle R}{\displaystyle |}}{\underset{\displaystyle |}{\overset{+}{N}}}\!-\!CH_2\!-\!NH\!-\!NO_2 \qquad (\underline{4.20})$$
$$\qquad\qquad\qquad\qquad H$$
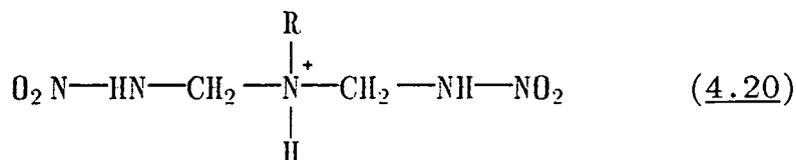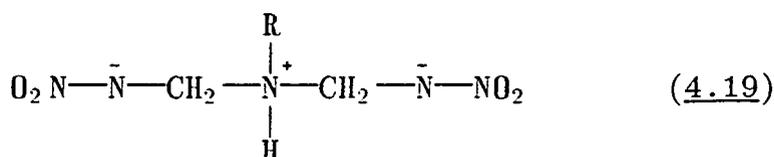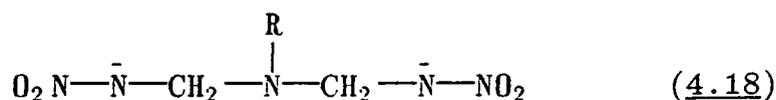
The subsequent decomposition of this intermediate might
lead to nitramide. However, since nitramide is unstable in
alkaline media it would not be detected. It seems probable
that if (4.18) exists in alkaline solution, it decomposes
rapidly on transfer to acidic solutions to give nitramide.
Thus, addition of dilute hydrochloric acid to a solution of
DPT in dilute sodium hydroxide gives the same UV spectrum
normally seen for DPT in acidic media. The results given in
this section suggest that (4.20) is not the species
ultimately formed from DPT in acid, although it may exist as
an intermediate. The possible mode of breakdown of DPT in
acid and base is summarised in Scheme 4.8, where (B) is a
low concentration ring-opened structure, conceivably similar
to (4.18), (4.19) or (4.20). (depending upon the pH), which
is in equilibrium with DPT. Intermediate (A) has been shown
in this chapter to be nitramide.

$$\text{DPT} \xrightarrow{k_1} (\underline{B}) \xrightarrow[\text{fast}]{\text{acid}} (\underline{A}),\ NH_2NO_2 \xrightarrow[\text{slow}]{k_2} \text{products}$$

$$\downarrow{\scriptstyle k_2 \atop \text{base}}$$

$$NH_2NO_2 \xrightarrow{\text{fast}} \text{products}\quad (N_2O\ +\ H_2O)$$

Scheme 4.8

The nature of the rate-determining step of the second stage of the reaction changes with pH. In basic solution, decomposition of (<u>B</u>) to give nitramide, which has been shown to decompose rapidly at high pH (see Table 4.12), will be rate-determining. In acid solution, nitramide decomposition becomes rate-determining. The work of Cooney[11,13] has shown that the change from one mechanism to the other probably occurs at <u>ca</u>. pH 8-9.

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CHAPTER 5

Studies on dinitroso-pentamethylenetetramine

5.1 Introduction

It was seen in Chapter 1 how nitrosation of hexamine
(1.1) in aqueous solution, by addition of hydrochloric or
acetic acid and a solution of sodium nitrite[1] yields the
eight-membered bicyclic dinitroso species, DNPT (1.12), or
the six-membered trinitroso species, TMTN (1.11), depending
upon the pH. There is particular interest in the eight-
membered dinitroso compound as a possible intermediate in
HMX synthesis. It is likely that DNPT could undergo further
nitrosation/nitration and oxidation to give a method of
production of HMX economically competitive with the Bachmann
process. It is clear that formation of DNPT from hexamine
can occur by either selective cleavage of the hexamine
molecule during nitrosation, with the basic ring structure
remaining intact, or via cleavage to simple fragments,
followed by nitrosation and recombination, the nature of the
products presumably being determined by which particular
mechanistic route is followed. It was therefore decided to
study the mechanisms of formation and decomposition of DNPT.

5.2 Mechanistic studies using ^{15}N-compounds

Earlier work on the mechanisms of acetylation of hexamine to DAPT (1.6)[2] nitration of hexamine to DPT, and nitration of DAPT to DADN (1.9) (Chapter 3 and Reference 3), using mixtures of ^{14}N- and ^{15}N-compounds as starting materials gave valuable insights into the mechanisms of these reactions, and thus it was decided to use a similar technique to study the nitrosation of hexamine to DNPT.

^{15}N-hexamine was prepared, as described in Chapter 2, from ^{15}NH$_3$ (99% isotopic abundance). The experiment involving nitrosation of a mixture of ^{14}N- and ^{15}N-hexamine was conducted by following the method of Bachmann and Deno[1] described in Chapter 2. for the preparation of ^{14}N-DNPT. 0.25 g (1.79 mmol) ^{14}N-hexamine and 0.25 g (1.74 mmol) ^{15}N-hexamine (overall atomic ratio ^{14}N:^{15}N = 100:96) was used to give the mixed ^{14}N/^{15}N product.

The best method of mass spectrometric analysis was found to be Electron Impact. A sample of DNPT prepared from hexamine with nitrogen in natural abundance showed only a small parent peak (186), but a fragment of mass 170, corresponding to loss of a single oxygen atom, was prominent, and gave M+1 and M+2 peaks whose intensities were close to those predicted theoretically. as shown in Table 5.1.

Table 5.1   Mass spectrometric data for ^{14}N-DNPT

	Peak intensities		
	M	M+1	M+2
theoretical (M = 186)	100	10.4	2.0
observed (M = 170)	100	8.0	1.1

The data in Table 5.2 for DNPT prepared from the mixture of ^{14}N- and ^{15}N-hexamine correspond to masses 170-174, and have been corrected for M+1 and M+2 peaks due to naturally occurring ^{13}C and ^{15}N.

Table 5.2   Relative isotopic composition of DNPT prepared by nitrosation of a mixture of ^{14}N- and ^{15}N-hexamine

	M	M+1	M+2	M+3	M+4
	$[^{14}N_4]$	$[^{14}N_3{}^{15}N_1]$	$[^{14}N_2{}^{15}N_2]$	$[^{14}N_1{}^{15}N_3]$	$[^{15}N_4]$
experimental	100	40	14	43	69
selective[a]	100	0	0	4	93
random[b]	100	390	564	352	81

a.   calculated distribution for selective cleavage of methylene bridge.

b.   calculated for random isotopic distribution.

It can be seen that the most abundant species are those containing the isotopically pure $[^{15}N_4]$ and $[^{14}N_4]$ compositions. Thus the predominant mechanism of nitrosation involves cleavage of a methylene bridge of the hexamine molecule, with the ring structure remaining intact. The small quantities of DNPT containing a 3:1 isotopic ratio indicates the presence of a minor pathway involving cleavage to a species containing three ring nitrogens[2].

Comparing the present case of nitrosation of hexamine with that of acetylation[2], it appears that when reaction is carried out under mild conditions and in media which are not highly acidic, then little ring cleavage is observed, and the major products contain the eight-membered ring structure. In contrast, under more severe conditions and at higher acidity, as in the nitration process (Chapter 3 and Reference 3), extensive cleavage occurs followed by recombination of fragments which preferentially form six-membered rings. The observation of Bachmann and Deno[1] that the acidity of the nitrosating medium determines whether the major product is the eight-membered DNPT (1.12) or the six-membered TMTN (1.11) is significant, since this is compatible with a general mechanism, Scheme 3.1, proposed earlier, for reaction of hexamine with electrophiles, in which the course of reaction depends on the rate of hydrolysis of the $H_2C{=\!\!=}N^+$ group compared to cleavage of other C—N bonds.

During acetylation and nitrosation reactions carried out in the presence of water, and in mildly acidic conditions, hydrolysis will eliminate formaldehyde to give a species which readily reacts further with electrophiles yielding

eight-membered products. However, at the higher acidity and low water concentrations employed in the nitration process, the hydrolysis of the $H_2C=N^+$ group, involving attack by water and proton elimination[4,5] will be more difficult, so that cleavage of other C—N bonds becomes preferential.

5.3   The decomposition of DNPT in aqueous solutions

The decomposition of DNPT in aqueous solutions was examined by monitoring the fade in UV absorbance at 225 nm (λ_{max} = 228 nm, ϵ = 10,100 l mol^{-1} cm^{-1}). DNPT was found to be stable in neutral or alkaline media for prolonged periods of time, but in acid solutions an accurately first-order process was observed, resulting in complete fading of the UV absorbance. The variation of the observed rate coefficient, k_{obs}, with acidity was measured in aqueous hydrogen chloride (i.e. HCl in H_2O) and aqueous deuterium chloride (i.e. DCl in D_2O), and gave the results shown in Tables 5.3 and 5.4 respectively.

It is known[6] that in dilute solutions the acidity function for the DCl/D_2O system is identical to that for HCl/H_2O. At acid concentrations less than 0.2 M there is a curved dependence of $\log(k_{obs})$ on H_0 (see Figure 5.1) which is consistent with rate determining decomposition of the protonated substrate[8] as shown in Equation (5.1).

Table 5.3 Observed rate coefficients for the decomposition
of DNPT (4×10^{-5} M) in HCl/H$_2$O measured at 25°C

[HCl]\mol l^{-1}	k$_{obs}$\s^{-1}
0.001	6.0×10^{-4}
0.002	1.2×10^{-3}
0.003	2.0×10^{-3}
0.004	2.2×10^{-3}
0.005	2.9×10^{-3}
0.006	3.2×10^{-3}
0.007	3.6×10^{-3}
0.008	4.4×10^{-3}
0.01	5.1×10^{-3}
0.02	0.011
0.03	0.011
0.04	0.016
0.05	0.018
0.06	0.022
0.08	0.026
0.10	0.022
0.20	0.036
0.50	0.083
1.00	0.22
1.50	0.69
2.00	1.27

Table 5.4 Observed rate coefficients for the decomposition of DNPT (4×10^{-5} M) in DCl/D_2O measured at $25^{\circ}C$

$[DCl]\backslash mol\ l^{-1}$	$k_{obs}\backslash s^{-1}$
0.001	2.0×10^{-3}
0.002	3.5×10^{-3}
0.003	6.2×10^{-3}
0.004	7.9×10^{-3}
0.005	8.7×10^{-3}
0.006	8.9×10^{-3}
0.007	0.010
0.01	0.012
0.02	0.020
0.03	0.023
0.04	0.023
0.05	0.036
0.06	0.041
0.08	0.047
0.10	0.056
0.20	0.087

Figure 5.1 Acidity dependence of k_{obs} for decomposition of

DNPT at $25^{\circ}C$ in aqueous hydrochloric acid (x)

and in deuterium chloride in deuterium oxide

(o):

line A (————) is calculated from Equation

(5.2) with $k = 0.05$ sec^{-1}, $K = 11.4$ 1 mol^{-1};

line B (— — — —) is calculated from Equation

(5.2) with $k = 0.05$ sec^{-1}, $K = 40$ 1 mol^{-1}

$$\text{DNPT} \quad + \quad \text{H}^+ \quad \underset{K}{\rightleftharpoons} \quad \text{DNPTH}^+ \quad \xrightarrow{k} \quad \text{products} \qquad \text{Eqn. (5.1)}$$
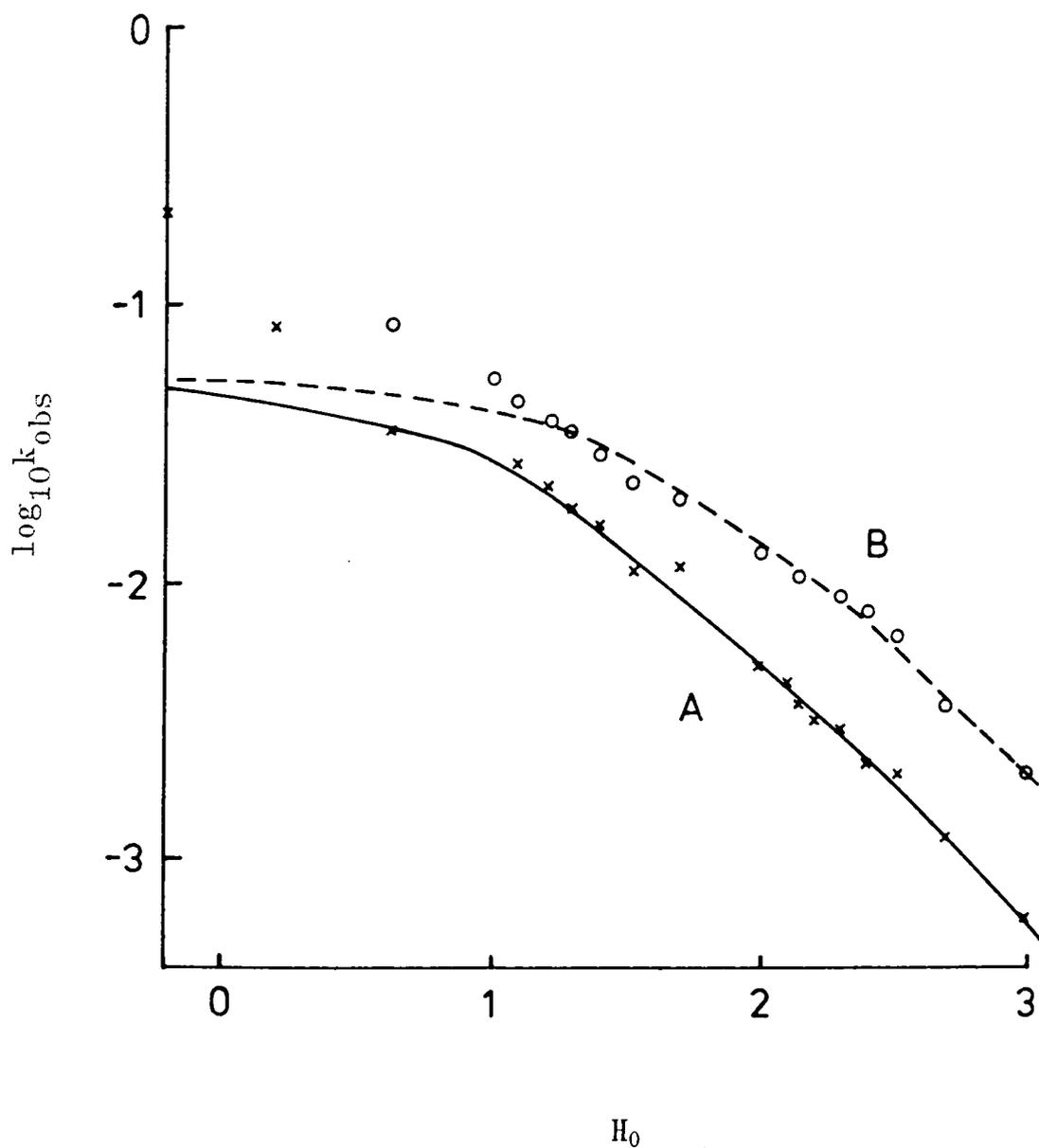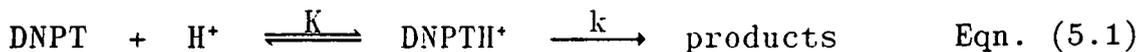
$$[\text{DNPT}] + [\text{DNPTH}^+] + [\text{products}] = \text{constant}$$

$$[\text{DNPTH}^+] = K[\text{DNPT}][\text{H}^+]$$

$$\text{therefore} \quad [\text{DNPT}] + K[\text{DNPT}][\text{H}^+] + [\text{products}] = \text{constant}$$

$$[\text{DNPT}]\left[1 + K[\text{H}^+]\right] + [\text{products}] = \text{constant}$$

$$\frac{d[\text{DNPT}]}{dt}\left[1 + K[\text{H}^+]\right] + \frac{d[\text{products}]}{dt} = 0$$

$$\text{rate} = k[\text{DNPTH}^+] = \frac{d[\text{products}]}{dt}$$

$$\frac{d[\text{products}]}{dt} = kK[\text{DNPT}][\text{H}^+]$$

$$\text{therefore} \quad \frac{d[\text{DNPT}]}{dt}\left[1 + K[\text{H}^+]\right] + kK[\text{DNPT}][\text{H}^+] = 0$$

$$\text{and} \quad -\frac{d[\text{DNPT}]}{dt} = k_{obs}[\text{DNPT}]$$

$$-k_{obs}[\text{DNPT}]\left[1 + K[\text{H}^+]\right] + kK[\text{DNPT}][\text{H}^+] = 0$$

$$k_{obs} = \frac{kK[\text{H}^+]}{1 + K[\text{H}^+]}$$

expressing in terms of an acidity function:

$$k_{obs} = \frac{kKh_0}{1 + Kh_0} \qquad \text{Eqn. (5.2)}$$

Scheme 5.1

Figure 5.1 shows that at acid concentrations less than 0.2 M there is a good fit with the theoretical equation (5.2) for HCl/H_2O when $k = 0.05$ s^{-1}, $K = 11.4$ 1 mol^{-1}, and for DCl/D_2O when $k = 0.05$ s^{-1}, $K = 40$ 1 mol^{-1}. The identical values for k obtained in water and deuterium oxide are consistent with this step involving C—N bond cleavage. However, the hydronation constant, K, is 3.5 times greater in deuterium oxide than in water, this higher basicity being in agreement with the literature values for other nitrogen bases.

A feature that is observed here with the dinitroso derivative, but is not observed with the corresponding diacetyl[8] or dinitro[9] derivative, is that at higher acidities values of k_{obs} are higher than predicted by Equation (5.2). This is interpretable in terms of partial di-hydronation of DNPT in acid media more concentrated than 0.2 M.

5.3.1 Determination of formaldehyde liberated during the decomposition of DNPT in acid media

It was thought that DNPT would produce formaldehyde when it undergoes decomposition in acid media, and this was confirmed when neutralisation of a reaction in 1M-hydrochloric acid precipitated 2.8 equivalents of formaldehyde-dimedone adduct[10] (see Chapter 2). In a control experiment, neutralisation of 5 equivalents of formaldehyde and 2 equivalents of ammonia in similar acidic medium yielded 3.6 equivalents of dimedone-adduct. Thus it is shown that some condensation of formaldehyde and ammonia occurs in neutral solution.

During the decomposition of DPT, nitramide, NH_2NO_2, has been shown[9] to be an intermediate in acid solution. No intermediates were observed during the decomposition of DNPT, but, however, it is known[11] that nitrosamine, NH_2NO, decomposes very rapidly in acid media to give nitrogen[12] and water. Nitrosamine cannot, therefore, be ruled out as an intermediate species during the decomposition of DNPT. Combined with the formaldehyde determination our results are consistent with Scheme 5.2, the rate determining step being C—N bond cleavage in the hydronated substrate.

$$\text{DNPT} \; \underset{}{\overset{H^+}{\rightleftharpoons}} \; \text{DNPTH}^+ \; \longrightarrow \; 2NH_2NO \; + \; 5CH_2O \; + \; 2NH_3$$

$$\downarrow$$

$$2N_2 \; + \; 2H_2O$$
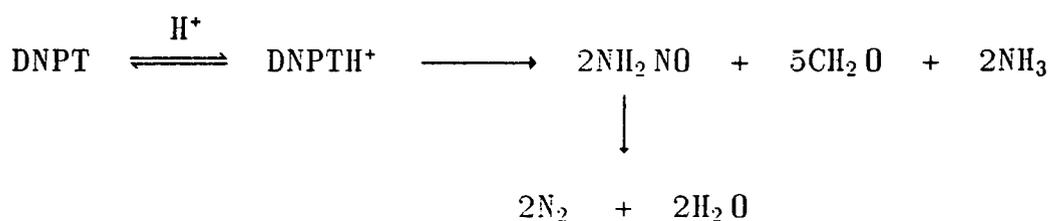
Scheme 5.2

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CHAPTER 6

Studies on the decomposition of

3,7-bis(arylazo)-1.3.5.7-tetraaza-

bicyclo[3.3.1]nonanes in acid solution

6.1 Introduction

The formation of 3,7-bis(arylazo)-1,3,5,7-tetraaza-bicyclo[3.3.1]nonanes (1.14) by diazonium coupling with hexamine (1.1)[1-3] was described in Chapter 1. Although the compounds themselves are of no direct military use, it was decided that it might be useful to compare their behaviour in various media with that of DPT (1.5), DNPT (1.12), and DAPT (1.6), since all are based on the same bicyclic ring structure with electron-withdrawing groups on the 3,7-nitrogens.

6.2 UV spectra of bis(arylazo) compounds

3,7-bis(p-chlorophenylazo)-1,3,5,7-tetraazabicyclo-[3.3.1]nonane (6.1) and 3,7-bis(p-ethoxycarbonylphenylazo)-1,3,5,7-tetraazabicyclo[3.3.1]nonane (6.2) were synthesised by the method of Reference 3, described in Chapter 2.
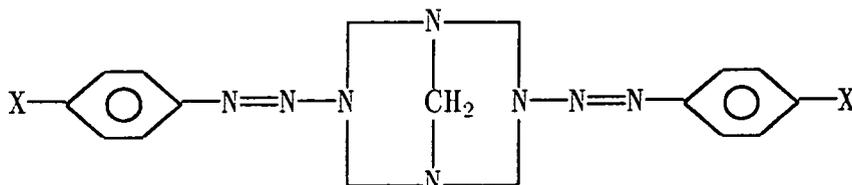
(6.1)   X = Cl

(6.2)   X = CO_2Et

The p-chloro derivative (6.1) was found to be stable as a 10^{-3} M stock solution in acetonitrile, as was the p-ethoxy-carbonyl derivative (6.2). Data for the UV absorbance maxima of these compounds and the corresponding anilines from which they were formed are given in Table 6.1.

Table 6.1 UV data for bis(arylazo) compounds and
corresponding anilines in acetonitrile at 25°C

Compound	$\lambda \backslash nm$	$\epsilon \backslash l \; mol^{-1} \; cm^{-1}$
(6.1)	224	15,700
	285	30,600
	310	shoulder
p-chloroaniline	247	13,100
	301	1,800
(6.2)	228	13,700
	305	35,300
	340	shoulder
p-ethoxycarbonyl-aniline	220	10,200
	285	22,250

Compounds (6.1) and (6.2) were seen to precipitate out of
solution when the acetonitrile stock was added to 100%
aqueous media, and thus all reactions had to be carried out
in a medium containing a high percentage of acetonitrile.
On addition of small quantities of water or aqueous sodium
hydroxide to acetonitrile solutions of (6.1) and (6.2) the
resulting UV spectrum was stable over a period of several
hours, and was comparable with that of the compounds in
acetonitrile only. However, addition of small quantities of
sulphuric acid, both aqueous and in acetonitrile only,
resulted in a shift of the absorbance at _ca_. 300 nm to

ca. 290 nm in the case of (6.2), and a fading of the absorbance for both compounds. An increase in the absorbance at ca. 225 nm was seen for (6.1) and (6.2), to ultimately give a spectrum very similar to that of the parent aniline in the same medium in each case.

Initial attempts to measure the rate of decomposition of (6.1) and (6.2) in sulphuric acid/acetonitrile solutions were complicated by the occurrence of an 'ageing' effect, in which the acid solutions appeared to change concentration over the period of a day. This was indicated by marked inconsistencies in the observed rate coefficients calculated for reaction in the same medium measured at differing times. This effect has been observed by other workers[4] and it has also been shown that sulphuric acid is incompletely dissociated in acetonitrile[4]. Because of these factors, it was decided to use perchloric acid/acetonitrile solutions for further work, since perchloric acid is completely dissociated in acetonitrile[4,5] No ageing effect was observed for these solutions.

The UV spectra of (6.1) and (6.2) in perchloric acid/acetonitrile were in agreement with those measured in sulphuric acid/acetonitrile, showing the compounds to decompose in the acid solutions to ultimately give a spectrum similar to that of the corresponding aniline in the same medium.

6.3 Kinetic data for the decomposition of 3,7-bis(p-ethoxy-carbonylphenylazo)-1,3,5,7-tetraazabicyclo[3.3.1]nonane in acid solution

UV spectra of the decomposition of (6.2) in perchloric acid/'pure' acetonitrile solutions, containing 0.05% water showed the process to occur in two distinct stages. Figure 6.1 shows the rapid decrease in absorbance at 290 nm, to form a species at ca. 270 nm, via no observable intermediates. Figure 6.2 illustrates how this species undergoes a much slower decomposition, with a decrease in absorbance at 270 nm, and a corresponding increase at ca. 220 nm.

It was found that addition of water to the reaction medium at a concentration of 1-5% resulted in the absorbance at 290 nm fading in a single process, as can be seen in Figure 6.3.

Figure 6.1 Successive UV spectra, taken at 15 second
            intervals, to illustrate the initial
            decomposition of (6.2) (4 x 10^{-5} M) in 0.001M-
            perchloric acid/acetonitrile solutions
            containing 0.05% water at 25°C
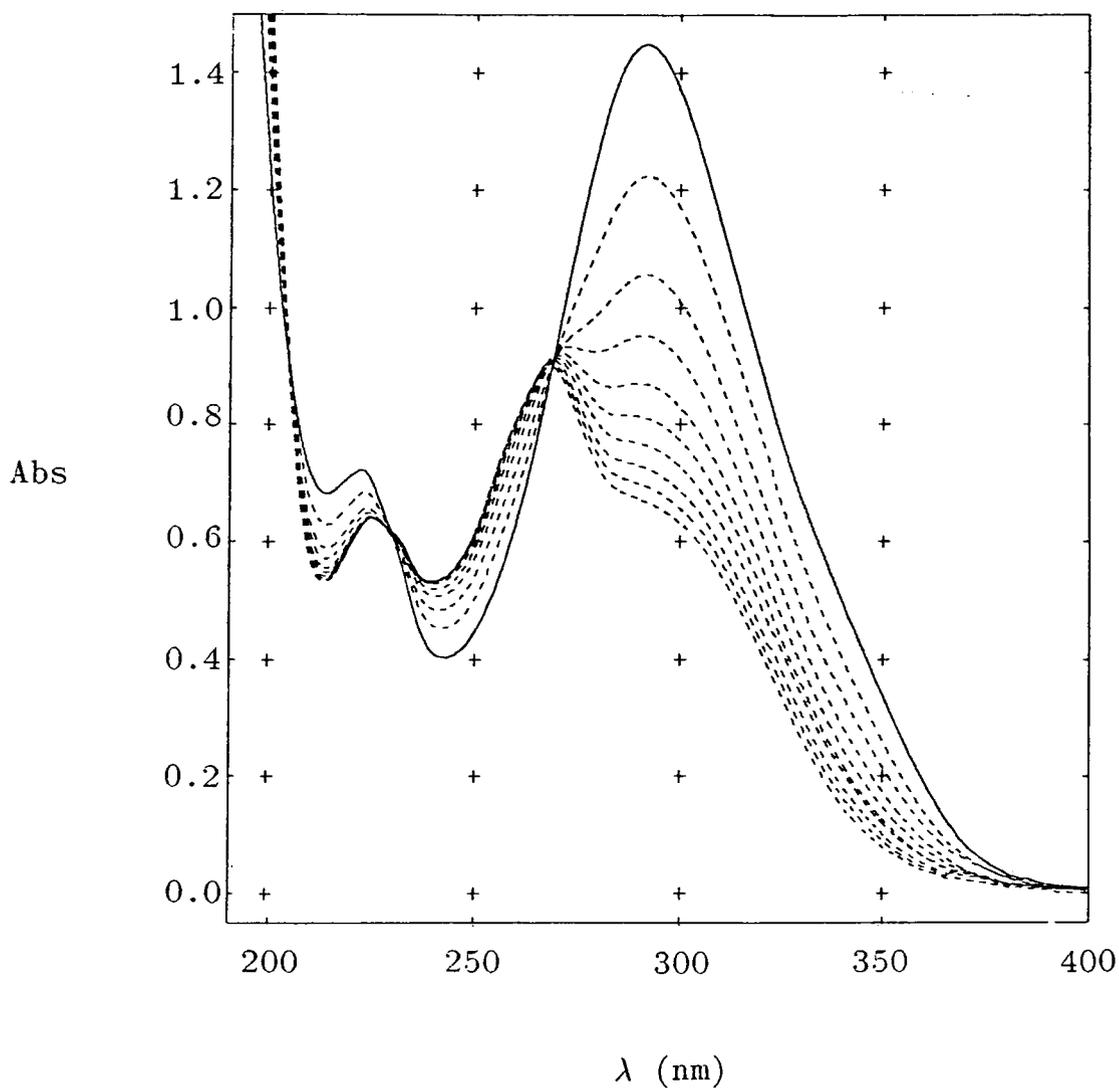
Figure 6.2 Successive UV spectra, taken at 2 minute intervals, to illustrate the initial decomposition of (6.2) (4 x 10^{-5} M) in 0.001M-perchloric acid/acetonitrile solutions containing 0.05% water at 25°C
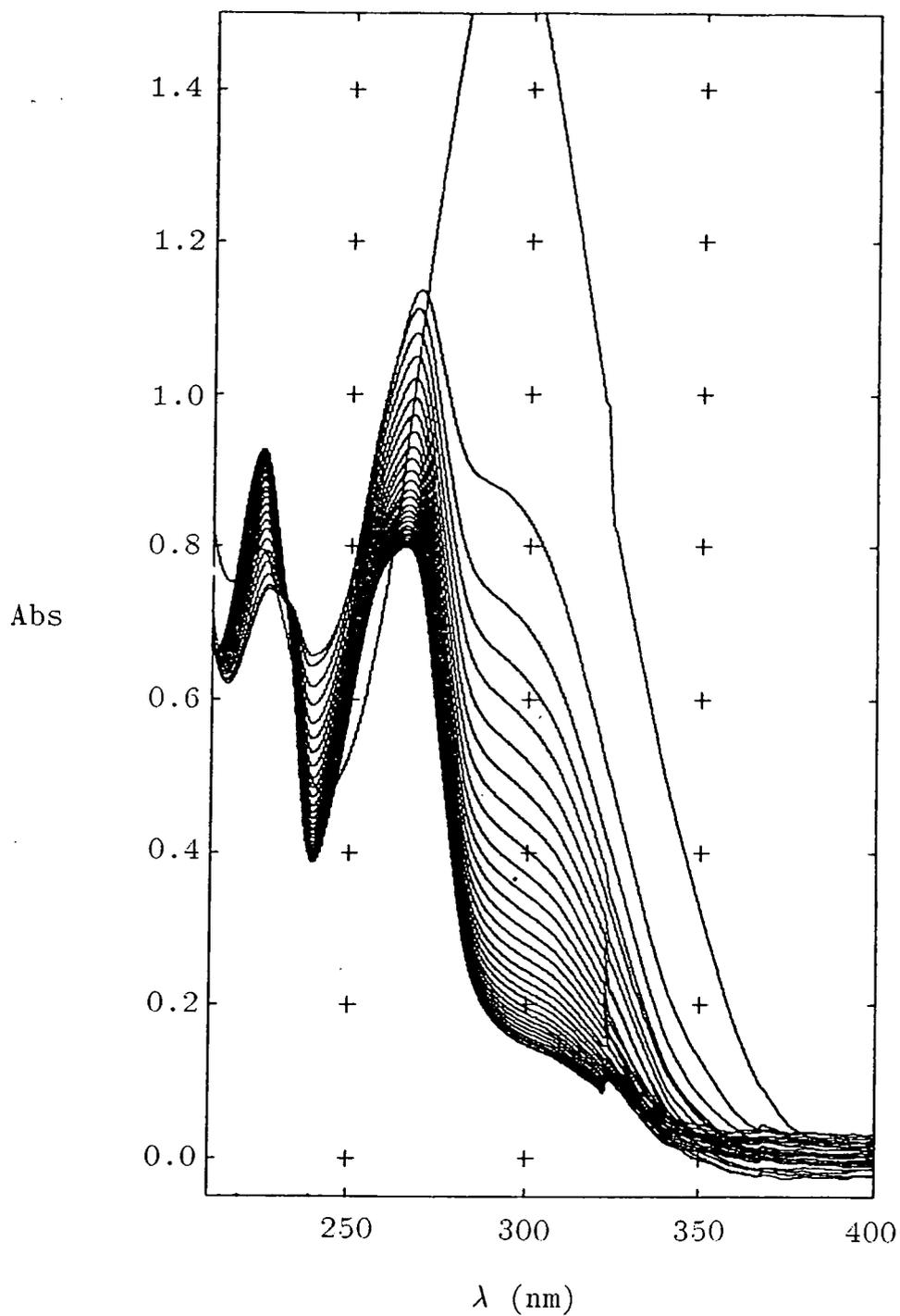
Figure 6.3   Successive UV spectra, taken at 30 second
intervals, to illustrate the single step
decomposition of (6.2) (4 x 10^{-5} M) in 0.001M-
perchloric acid/acetonitrile solutions
containing 1% water at 25°C
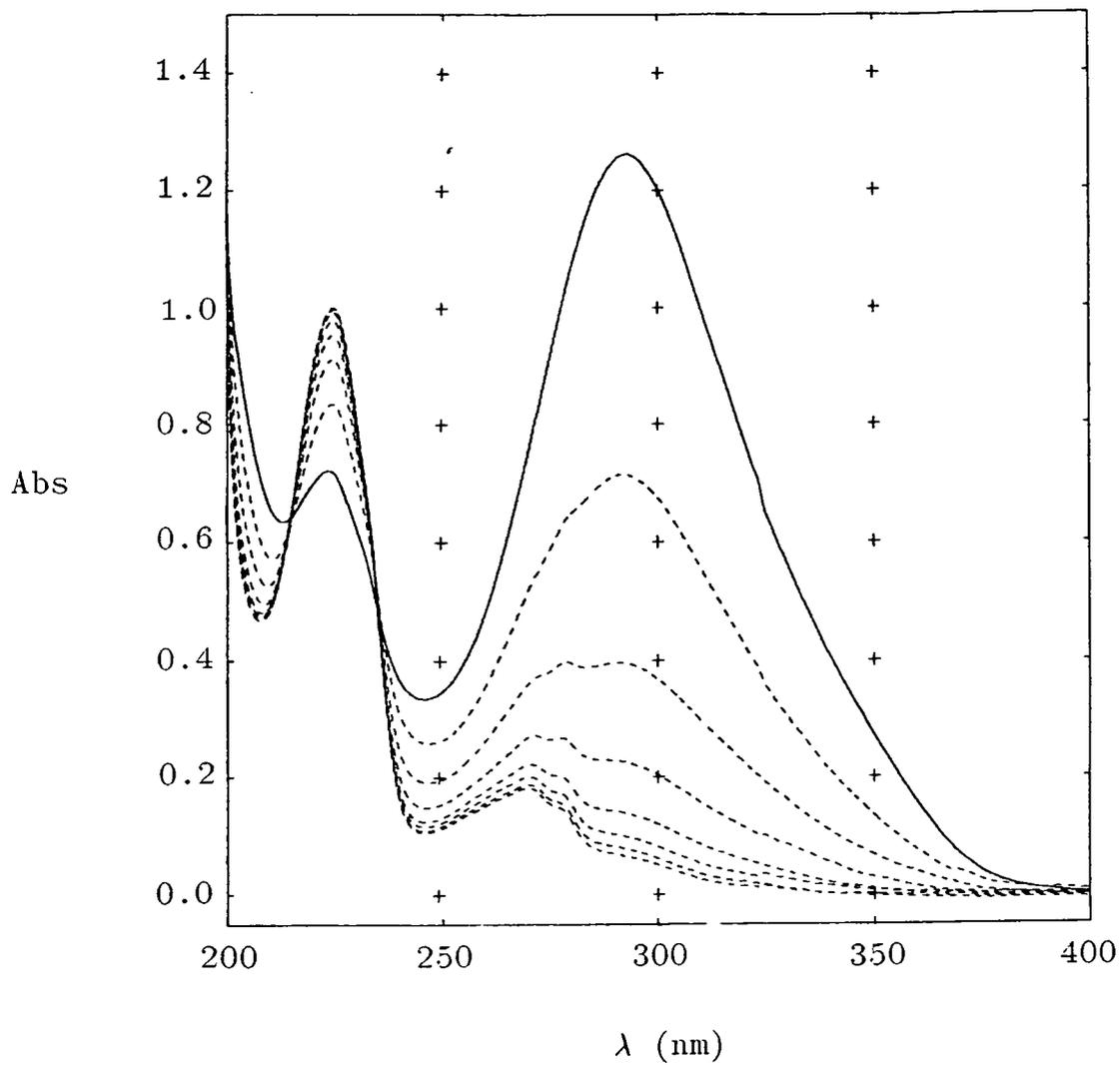
Table 6.2 shows the acidity dependence of the observed, separable, first order rate coefficients for the two step decomposition of (6.2) in perchloric acid/acetonitrile solutions containing 0.05% water, measured as fading UV absorbances at 300 nm and 270 nm. Tables 6.3-6.7 show the observed first order rate coefficients for the single step decomposition of (6.2) in perchloric acid/acetonitrile solutions containing 1-5% water, measured as a fading UV absorbance at 300 nm.

Table 6.2   Observed rate coefficients for the decomposition of (6.2) (4×10^{-5} M) in perchloric acid/ acetonitrile solutions containing 0.05% water at 25°C

$[HClO_4] \backslash mol\ l^{-1}$	$k_{obs,1} \backslash sec^{-1}$ [measured at 300 nm]	$k_{obs,2} \backslash sec^{-1}$ [measured at 270 nm]
0.001	0.038	1.0×10^{-3}
0.002	0.057	1.2×10^{-3}
0.003	0.080	1.1×10^{-3}
0.004	0.11	1.1×10^{-3}
0.005	0.11	9.5×10^{-4}
0.006	0.10	1.1×10^{-3}
0.007	0.12	1.3×10^{-3}
0.008	0.10	1.3×10^{-3}
0.009	0.10	1.3×10^{-3}

It can be seen from these results that $k_{obs,2}$ is independent of acid concentration.

Table 6.3   Observed rate coefficients for the decomposition
of (6.2) (4 x 10^{-5} M) in perchloric acid/
acetonitrile solutions containing 1% water
at 25°C

$[HClO_4]\backslash mol\ l^{-1}$	$k_{obs}\backslash sec^{-1}$
0.001	0.013
0.002	0.016
0.003	0.016
0.004	0.019
0.005	0.018
0.006	0.019
0.007	0.019
0.008	0.022
0.009	0.022

Table 6.4 Observed rate coefficients for the decomposition
of (6.2) (4 x 10^{-5} M) in perchloric acid/
acetonitrile solutions containing 2% water
at 25°C

$[HClO_4]\backslash mol\ l^{-1}$	$k_{obs}\backslash sec^{-1}$
0.001	0.013
0.002	0.018
0.003	0.020
0.004	0.023
0.005	0.024
0.006	0.026
0.007	0.024
0.008	0.027
0.009	0.029

Table 6.5 Observed rate coefficients for the decomposition of ($\underline{6.2}$) (4×10^{-5} M) in perchloric acid/ acetonitrile solutions containing 3% water at 25°C

$[\text{HClO}_4] \backslash \text{mol } l^{-1}$	$k_{obs} \backslash \text{sec}^{-1}$
0.001	6.7×10^{-3}
0.002	0.011
0.003	0.013
0.004	0.018
0.005	0.017
0.006	0.018
0.007	0.020
0.008	0.021

Table 6.6 Observed rate coefficients for the decomposition of (6.2) (4×10^{-5} M) in perchloric acid/ acetonitrile solutions containing 4% water at 25°C

$[HClO_4]\backslash mol\ l^{-1}$	$k_{obs}\backslash sec^{-1}$
0.001	3.2×10^{-3}
0.002	6.8×10^{-3}
0.003	9.8×10^{-3}
0.004	0.013
0.005	0.013
0.006	0.015
0.007	0.017
0.008	0.018
0.009	0.020

Table 6.7   Observed rate coefficients for the decomposition
of ($\underline{6.2}$) (4×10^{-5} M) in perchloric acid/
acetonitrile solutions containing 5% water
at 25°C

$[HClO_4] \backslash mol \ l^{-1}$	$k_{obs} \backslash sec^{-1}$
0.001	2.4×10^{-3}
0.002	4.8×10^{-3}
0.003	6.4×10^{-3}
0.004	8.3×10^{-3}
0.005	8.8×10^{-3}
0.006	0.011
0.007	0.012
0.008	0.013
0.009	0.013

Table 6.8 shows the effect of varying the water concentration at a fixed acid concentration on the observed rate coefficients. Notice that two distinct processes are observable in the decomposition of (6.2) up to a water concentration of 0.3 %.

Table 6.8 Observed rate coefficients for the decomposition of (6.2) (4×10^{-5} M) in 0.005M-perchloric acid/ acetonitrile solutions of varying water concentration at 25°C

% H_2O	$k_{1,obs}\backslash sec^{-1}$	$k_{2,obs}\backslash sec^{-1}$
0.05	0.11	1.2×10^{-3}
0.1	0.021	2.1×10^{-3}
0.2	0.017	3.6×10^{-3}
0.3	0.017	5.4×10^{-3}
1	0.018	-
2	0.022	-
3	0.016	-
4	0.010	-
5	7.4×10^{-3}	-

By making a comparison with the decomposition processes previously observed for other 3,7-substituted bicyclo-nonanes, such as DPT ($\underline{1.5}$) and DNPT ($\underline{1.12}$), it is possible to interpret these results in terms of Scheme 6.1, where k_1 and k_2 are the respective rate coefficients for the fast and slow decomposition processes observed in perchloric acid/ acetonitrile solutions containing less than 1% water. The final decomposition product is shown as the aniline, since this is suggested from comparison of UV spectra, as described earlier.

$$ArN_2 - N \underset{\underline{P}}{\underbrace{\begin{array}{c} N \\ CH_2 \\ N \end{array}}} N - N_2Ar \ + \ H^+ \ \overset{K_{H^+}}{\rightleftharpoons} \ \underline{P}H^+ \ \overset{k_1}{\longrightarrow} \ \underline{R}H^+ \ \overset{k_2}{\underset{H_2O}{\big\downarrow}} \ ArNH_2$$

Scheme 6.1

In any acid medium $\underline{P}$ is likely to undergo a rapid equilibrium protonation on one of the the ring amino nitrogen atoms followed by ring cleavage, to form a protonated ring-opened intermediate, $\underline{R}H^+$. It is proposed that at low water concentration k_1 is larger than k_2, the second stage of decomposition being water dependent, and thus two processes are observed, with an initial fast ring-opening resulting in formation of an observable intermediate, which undergoes a subsequent slow decomposition to the aniline. When water is present at a concentration above 1% the second process becomes faster

than the first, i.e. $k_2 > k_1$, and thus the intermediate and associated second decomposition process are not observed.

If one treats the two decomposition processes as separate reactions, as is reasonable when there is a large difference in rates, it is possible to derive the equations given below.

For the first process:

$$k_{obs,1} = \frac{k_1 K_{H^+}[H^+]}{1 + K_{H^+}[H^+]} \qquad \text{Eqn. (6.1)}$$

(see Scheme 5.1 for derivation)

For the second process:

$$k_{obs,2} = k_2 \qquad \text{Eqn. (6.2)}$$

Tables 6.9-6.14 compare the observed rate coefficients for the initial decomposition of (6.2) in acid solutions of varying water concentration with those calculated from Equation (6.1) using the values of k_1 and K_{H^+} given with each table.

Table 6.9 Comparison of observed and calculated rate coefficients for the initial decomposition of (**6.2**) (4×10^{-5} M) in perchloric acid/ acetonitrile solutions containing 0.05% water

$[HClO_4] \backslash mol\ l^{-1}$	$k_{obs,1} \backslash sec^{-1}$	$k^a_{obs,calc} \backslash sec^{-1}$
0.001	0.038	0.041
0.002	0.057	0.060
0.003	0.080	0.071
0.004	0.11	0.078
0.005	0.11	0.083
0.006	0.10	0.086
0.007	0.12	0.089
0.008	0.10	0.091
0.009	0.10	0.093

a. calculated from Equation (6.1) using values of
   $k_1 = 0.11\ sec^{-1}$, $K_{H^+} = 600\ l\ mol^{-1}$

It can be seen from Table 6.9 that the values of $k_{obs,1}$ have reached a limiting value rather more quickly than is predicted by calculation. This phenomenon remains unexplained.

<u>Table 6.10</u> Comparison of observed and calculated rate
coefficients for the initial decomposition of
(<u>6.2</u>) (4×10^{-5} M) in perchloric acid/
acetonitrile solutions containing 1% water

$[HClO_4]\backslash mol \ l^{-1}$	$k_{obs,1}\backslash sec^{-1}$	$k_{obs,calc}^a\backslash sec^{-1}$
0.001	0.013	0.011
0.002	0.016	0.018
0.003	0.016	0.022
0.004	0.019	0.025
0.005	0.018	0.027
0.006	0.019	0.028
0.007	0.019	0.029
0.008	0.022	0.030
0.009	0.022	0.031

a. calculated from Equation (6.1) using values of
k_1 = 0.040 sec^{-1}, K_{H^+} = 400 l mol^{-1}

Once again, as can be seen from the table, the values of
$k_{obs,1}$ have reached a limiting value faster than is
predicted by calculation. However, this might be explained
by the fact that at the higher acid concentrations the rate
of the first decomposition process is approaching the
predicted rate of the second (shown later), and under these
circumstances Equation (6.1) is not applicable.

Table 6.11 Comparison of observed and calculated rate
coefficients for the initial decomposition of
($\underline{6.2}$) (4 x 10^{-5} M) in perchloric acid/
acetonitrile solutions containing 2% water

$[HClO_4]\backslash mol\ l^{-1}$	$k_{obs,1}\backslash sec^{-1}$	$k^a_{obs,calc}\backslash sec^{-1}$
0.001	0.013	0.010
0.002	0.018	0.015
0.003	0.020	0.019
0.004	0.023	0.022
0.005	0.024	0.024
0.006	0.026	0.026
0.007	0.024	0.027
0.008	0.027	0.028
0.009	0.029	0.029

a. calculated from Equation (6.1) using values of
k_1 = 0.040 sec^{-1}, K_{H^+} = 300 l mol^{-1}

Table 6.12 Comparison of observed and calculated rate
coefficients for the initial decomposition of
(6.2) (4×10^{-5} M) in perchloric acid/
acetonitrile solutions containing 3% water

$[HClO_4]\backslash mol\ l^{-1}$	$k_{obs,1}\backslash sec^{-1}$	$k^a_{obs,calc}\backslash sec^{-1}$
0.001	6.7×10^{-3}	6.6×10^{-3}
0.002	0.011	0.011
0.003	0.013	0.014
0.004	0.018	0.016
0.005	0.017	0.018
0.006	0.018	0.020
0.007	0.020	0.021
0.008	0.021	0.022

a. calculated from Equation (6.1) using values of
$k_1 = 0.033\ sec^{-1}$, $K_{H^+} = 250\ l\ mol^{-1}$

Table 6.13 Comparison of observed and calculated rate
coefficients for the initial decomposition of
(6.2) (4×10^{-5} M) in perchloric acid/
acetonitrile solutions containing 4% water

$[HClO_4] \backslash mol\ l^{-1}$	$k_{obs,1} \backslash sec^{-1}$	$k_{obs,calc}^{a} \backslash sec^{-1}$
0.001	3.2×10^{-3}	3.6×10^{-3}
0.002	6.8×10^{-3}	6.7×10^{-3}
0.003	9.8×10^{-3}	9.2×10^{-3}
0.004	0.013	0.011
0.005	0.013	0.013
0.006	0.015	0.015
0.007	0.017	0.016
0.008	0.018	0.018
0.009	0.020	0.019

a. calculated from Equation (6.1) using values of
$k_1 = 0.040\ sec^{-1}$, $K_{H^+} = 100\ l\ mol^{-1}$

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Table 6.14 Comparison of observed and calculated rate
coefficients for the initial decomposition of
(6.2) (4 x 10^{-5} M) in perchloric acid/
acetonitrile solutions containing 5% water

[HClO$_4$]\mol l^{-1}	$k_{obs,1}$\sec^{-1}	$k_{obs,calc}^{a}$\sec^{-1}
0.001	2.4 x 10^{-3}	2.7 x 10^{-3}
0.002	4.8 x 10^{-3}	4.9 x 10^{-3}
0.003	6.4 x 10^{-3}	6.7 x 10^{-3}
0.004	8.3 x 10^{-3}	8.3 x 10^{-3}
0.005	8.8 x 10^{-3}	9.6 x 10^{-3}
0.006	0.011	0.011
0.007	0.012	0.012
0.008	0.013	0.013
0.009	0.013	0.013

a. calculated from Equation (6.1) using values of
k_1 = 0.027 sec^{-1}, K_{H^+} = 110 l mol^{-1}

The values of k_1 and K_{H^+} obtained are summarised in Table 6.15. Also shown are the values of k_2 ($k_2 = k_{obs,2}$) measured at 0.05-0.3% water, and the calculated values for 1-5% water, based on the linear dependence on acid concentration apparent from the measured values.

Table 6.15 Values of k_1, k_2 and K_{H^+} for the decomposition of (6.2) in perchloric acid/acetonitrile solutions containing varying concentrations of water at $25^{\circ}C$

% H_2O	$k_1 \backslash sec^{-1}$	$K_{H^+} \backslash 1\ mol^{-1}$	$k_2 \backslash sec^{-1}$	
0.05	0.11	600	1.2×10^{-3}	
0.1	-	-	2.1×10^{-3}	
0.2	-	-	3.6×10^{-3}	
0.3	-	-	5.4×10^{-3}	
1	0.040	400	0.02	
2	0.040	300	0.04	
3	0.033	250	0.06	a
4	0.040	100	0.08	
5	0.027	110	0.10	

a. calculated from measured values at lower water concentration

That the calculated value of k_1 in solutions containing only 0.05% water is roughly one hundred times larger than the measured value of k_2 (i.e. the second step determines the overall rate of decomposition) is consistent with the observation of two consecutive decomposition processes in such media, as outlined in Scheme 6.1. However, at higher water concentrations the calculated values of k_2 become higher than k_1 (i.e. the first step is determines the overall rate of decomposition), and thus the second stage of decomposition should not be observed, as is the case. At water concentrations of 1% and 2% the calculated values of k_2 are similar to those of k_1, although the decomposition of (6.2) in these media apparently occurs by a single observable process. It is possible that under these particular conditions k_1 and the true value of k_2 are virtually identical, so that the two first order processes appear as one, but it is more likely that inaccuracies in the calculations have obscured the more probable explanation that k_2 is, in fact, larger than k_1.

The calculated values of k_1 are similar in the solutions of higher water concentration, which is consistent with the first decomposition step being independent of water concentration.

The decrease in the calculated values of K_{H^+} as one moves to higher water concentrations is indicative of the greater proton-solvating power that is expected of such media, thus reducing the concentration available for protonation of (6.2). Such an effect was demonstrated in its simplest form by an experiment in which water was added to a solution of p-ethoxycarbonylaniline in 0.005M-perchloric acid/

acetonitrile solution containing little water. This
resulted in an increase in the absorbance at ca. 285 nm to a
value nearer that measured in acetonitrile only, as shown in
Figure 6.4, and is attributable to decreasing protonation of
the aniline as the water concentration increases. This
effect was mimicked by the species formed in solution via
decomposition of (6.2).

Table 6.16 illustrates how the pK_a of the aniline
decreases with increasing water concentration, these values
being calculated using Equation (6.3) and data derived from
the spectra shown in Figure 6.4.
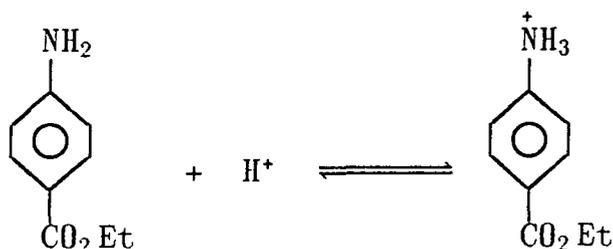
$$\text{gives} \quad pK_a = -\log_{10}\left[\frac{[NH_2 . C_6 H_4 . CO_2 Et][H^+]}{[^+NH_3 . C_6 H_4 . CO_2 Et]}\right] \qquad \text{Eqn. (6.3)}$$

Scheme 6.2

Table 6.16 pK_a values for p-ethoxycarbonylaniline in
acetonitrile containing varying concentrations
of water at $25^{\circ}C$

% H_2O	$[NH_2.C_6H_4.CO_2Et]$ \mol l^{-1}	$[^+NH_3.C_6H_4.CO_2Et]$ \mol l^{-1}	pK_a
1	9.00×10^{-7}	7.91×10^{-5}	4.24
2	6.30×10^{-6}	7.37×10^{-5}	3.37
3	1.44×10^{-5}	6.56×10^{-5}	2.96
4	2.52×10^{-5}	5.48×10^{-5}	2.64
5	3.19×10^{-5}	4.81×10^{-5}	2.48
6	3.91×10^{-5}	4.09×10^{-5}	2.32
7	4.18×10^{-5}	3.82×10^{-5}	2.26
8	4.72×10^{-5}	3.28×10^{-5}	2.14

Figure 6.4 UV spectra for p-ethoxycarbonylaniline
$(8 \times 10^{-5} \text{ M})$ in 0.005M-perchloric acid/
acetonitrile solutions containing varying
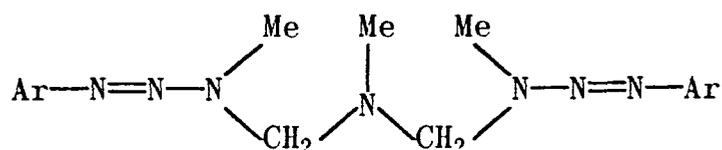concentrations of water at 25°C
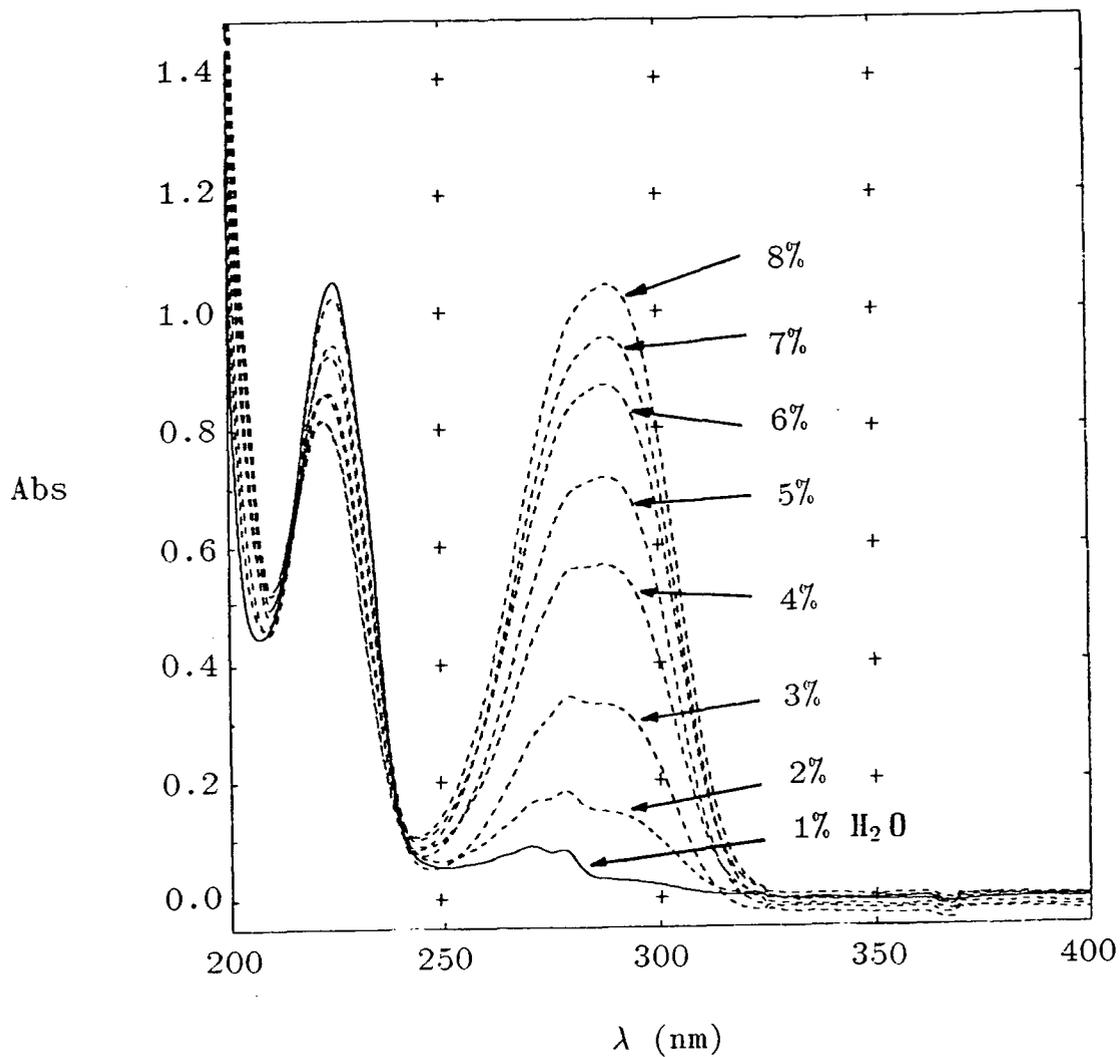
6.4 The role of triazenes

The results for decomposition of (6.2) in media rich in acetonitrile indicate the presence of an intermediate, R in Scheme 6.1. A similar intermediate is likely to exist during decomposition of (6.1). A major difference between the studies of the decomposition of (6.2) and of the related compounds DPT and DNPT is the solvent, the decomposition of the latter two compounds being studied in aqueous media. The results for the decomposition of (6.2) show that in media containing more than 1% water the intermediate R is no longer observable. It was shown in Chapter 4 how the observed intermediate in the decomposition of the 3,7-dinitro derivative, DPT (1.5), in acid solution was nitramide (NH_2NO_2) or its protonated cation, and it was postulated in Chapter 5 that the probable, though unobserved, intermediate in the decomposition of the 3,7-dinitroso derivative, DNPT (1.12), in acid solution was nitrosamine (NH_2NO) or its protonated cation. Thus, by analogy one might expect the intermediate in the decomposition of 3,7-bis(arylazo) structures such as (6.1) and (6.2) in acid solution to be a triazene of structure (6.3) and (6.4) respectively, or its protonated cation.

$$Ar-N{=}N-NH_2 \qquad (6.3) \quad Ar = p\text{-}Cl.C_6H_4$$

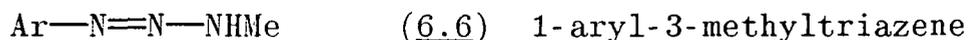$$(6.4) \quad Ar = p\text{-}EtO_2C.C_6H_4$$
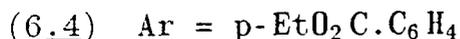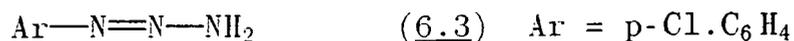
This was also suggested by Vaughan and co-workers[3] based on their finding that bis-triazenes (6.5), close structural analogues of the bis(arylazo)tetraazabicyclononanes, undergo acid-catalysed hydrolysis to the corresponding aniline via
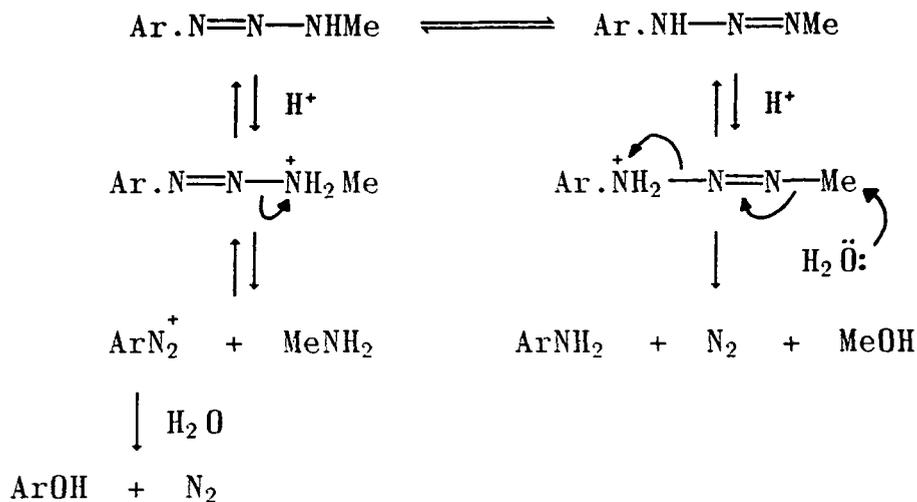
the intermediate 1-aryl-3-methyltriazene (6.6).

$$Ar—N{=}N—N\begin{smallmatrix}Me\\ \\CH_2\end{smallmatrix}\quad N\begin{smallmatrix}Me\\ \\ \end{smallmatrix}\quad CH_2\quad N\begin{smallmatrix}Me\\ \\ \end{smallmatrix}—N{=}N—Ar$$

(6.5)

Ar—N=N—NHMe   (6.6)   1-aryl-3-methyltriazene

Formation of the aniline when 1-aryltriazenes undergo hydrolysis has also been demonstrated by other workers,[6-10] and it is of particular relevance here, since this was identified as the final product in the decomposition of (6.1) and (6.2), and was also identified as such by Vaughan and co-workers in their studies of the decomposition of (6.2) in mixed-solvent buffer systems. It is interesting to note that formation of the aniline requires (6.6) to undergo tautomerisation[11-16] prior to decomposition via elimination of nitrogen, otherwise phenol and methylamine would be produced, as illustrated in Scheme 6.3.

$$\text{Ar.N} = \text{N} - \text{NHMe} \quad \rightleftharpoons \quad \text{Ar.NH} - \text{N} = \text{NMe}$$

$$\Big\updownarrow \text{H}^+ \qquad\qquad\qquad\qquad \Big\updownarrow \text{H}^+$$

$$\text{Ar.N} = \text{N} - \overset{+}{\text{N}}\text{H}_2\,\text{Me} \qquad \text{Ar.}\overset{+}{\text{N}}\text{H}_2 - \text{N} = \text{N} - \text{Me}$$

$$\Big\updownarrow \qquad\qquad\qquad\qquad \Big\downarrow \quad \text{H}_2\ddot{\text{O}}:$$

$$\text{Ar}\overset{+}{\text{N}_2} \;+\; \text{MeNH}_2 \qquad \text{ArNH}_2 \;+\; \text{N}_2 \;+\; \text{MeOH}$$

$$\Big\downarrow \text{H}_2\text{O}$$

$$\text{ArOH} \;+\; \text{N}_2$$

Scheme 6.3

In order to assess properly the possibility that the triazene (6.4) was the intermediate observed during the decomposition of (6.2) in perchloric acid/acetonitrile solutions containing only 0.05% water it would be necessary to measure its UV spectra and rates of decomposition in the appropriate media. However, isolation of (6.4) has so far proved impossible, and studies were therfore conducted on its closest analogue, the 3-methyl derivative (6.6).

6.4.1 UV spectra of 1-(p-ethoxycarbonylphenyl)-3-methyltriazene

A sample of (6.6) was prepared by the method of Reference 11, as described in Chapter 2.

The triazene was found to be stable as a 10^{-3} M stock solution in acetonitrile, the UV spectrum showing absorbances at λ = 305 nm, ϵ = 15,900 l mol^{-1} cm^{-1} and λ = 224 nm, ϵ = 8,500 l mol^{-1} cm^{-1}. In dilute perchloric acid/acetonitrile solutions containing 0.05% water the absorbance at 305 nm was seen to fade rapidly, with a weaker absorbance appearing at ca. 265 nm. The final spectrum resembled that of the corresponding aniline in the same medium. Indeed, in perchloric acid solutions containing up to 10% water a rapid decomposition was observed to ultimately form the aniline. Addition of aqueous sodium hydroxide solution or water to acetonitrile solutions of (6.6) did not alter the UV spectrum.

The effect discussed earlier, whereby addition of water to a solution of the aniline or the product formed from decomposition of (6.2) in perchloric acid/acetonitrile solutions resulted in a shift in the protonation equilibrium, was also seen for the species formed via decomposition of (6.6).

6.4.2 Kinetic data for the decomposition of 1-(p-ethoxy-carbonylphenyl)-3-methyltriazene in acid solutions

Kinetic measurements of the decomposition of (6.6) in perchloric acid/acetonitrile solutions containing varying concentrations of water, made by following the fade in UV absorbance at 300 nm, showed this to occur via a single first order process. This was found to be highly dependent on water concentration, as illustrated in Table 6.17, and independent of acid concentration, as can be seen from Table 6.18. Table 6.19 shows results obtained for a situation where water concentration increased as acid concentration increased, the change in observed rate coefficients being wholly attributable to the former.

Table 6.17 Observed rate coefficients for the decomposition of (6.6) (5×10^{-5} M) in 0.001M-perchloric acid/ acetonitrile solutions containing varying concentrations of water at 25°C

% H_2O	$k_{obs} \backslash sec^{-1}$
0.05	0.055
0.1	0.11
0.2	0.32

Table 6.18   Observed rate coefficients for the decomposition
of (6.6) (5 x 10^{-5} M) in perchloric acid/
acetonitrile solutions containing ca. 0.5% water
at 25°C

$[HClO_4]$\mol 1^{-1}	% H_2O	k_{obs}\sec^{-1}
0.006	0.55	0.51
0.013	0.54	0.45
0.025	0.52	0.32
0.050	0.49	0.21

Table 6.19   Observed rate coefficients for the
decomposition of (6.6) (8 x 10^{-5} M) in
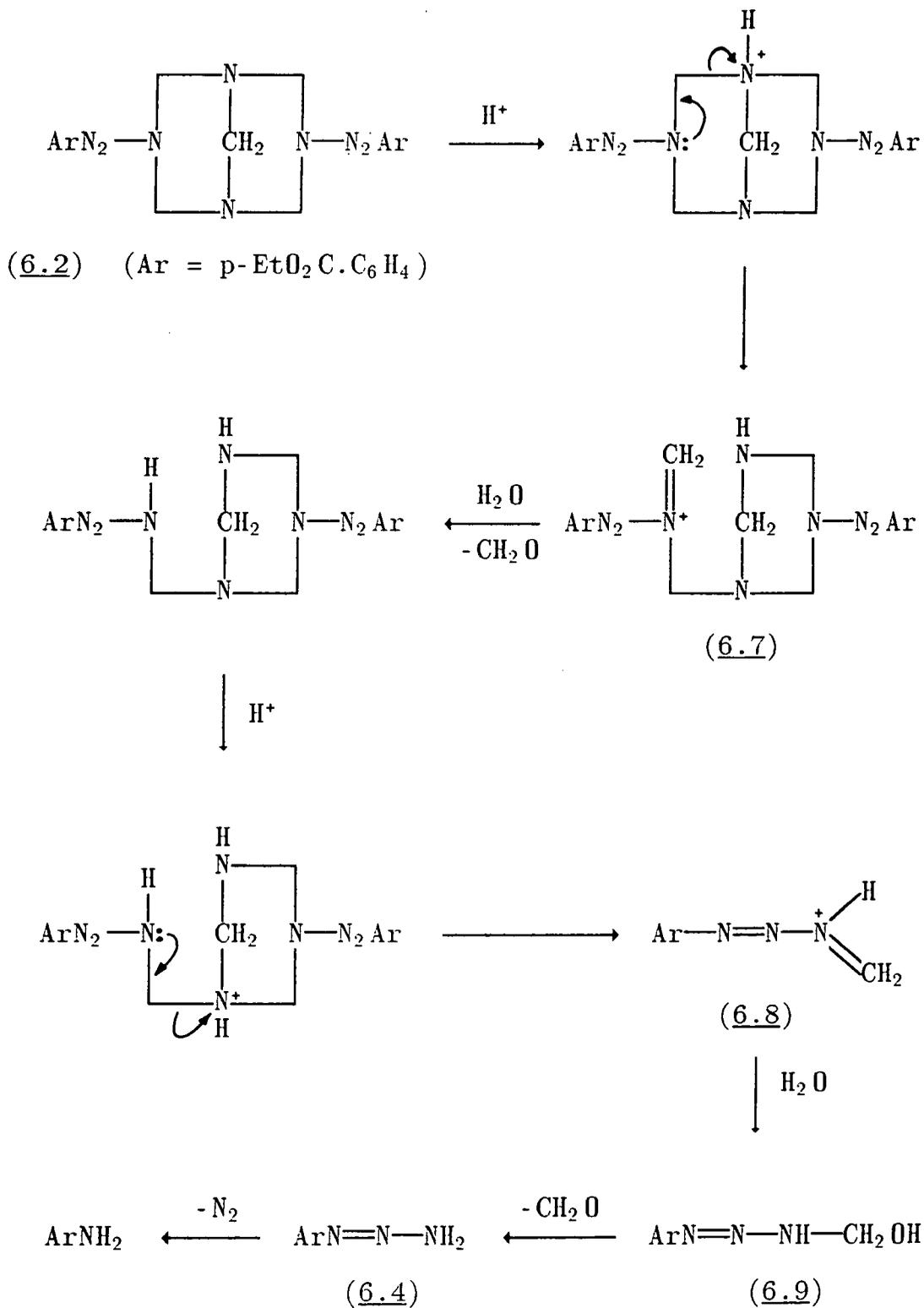perchloric acid/acetonitrile solutions at 25°C

$[HClO_4]$\mol 1^{-1}	% H_2O	k_{obs}\sec^{-1}
0.001	0.055	0.055
0.002	0.059	0.064
0.003	0.064	0.073
0.004	0.068	0.080
0.005	0.073	0.081
0.006	0.077	0.085
0.007	0.082	0.089
0.008	0.086	0.10
0.009	0.091	0.11

The observed rate coefficients for this decomposition
process are about one hundred times larger than those
measured for the secondary decomposition of (6.2) in the
same media, shown in Table 6.2. Whilst these results do not

represent the rates of decomposition of the proposed intermediate (<u>6.4</u>), it is expected that substitution of a hydrogen atom for a methyl group would lead to an increase in rate, and thus it is apparent that the intermediate observed in the decomposition of (<u>6.2</u>) is not the triazene (<u>6.4</u>).

6.5 Discussion

Vaughan and co-workers[3] proposed the mechanism shown in Scheme 6.4 for the decomposition of compounds such as (6.2) in acid solutions.

$(\underline{6.2})$ $(Ar = p\text{-}EtO_2C.C_6H_4)$

$$ArN_2\text{—}NH_2 \xleftarrow{\ -N_2\ } ArN{=}N\text{—}NH_2 \xleftarrow{\ -CH_2O\ } ArN{=}N\text{—}NH\text{—}CH_2OH$$

$(\underline{6.4})$ $(\underline{6.9})$

Scheme 6.4

- 157 -

The observations made by the author of the decomposition of (6.2) are in agreement with this scheme, which is of particular interest with regard to the proposed methyleniminium ion intermediates (6.7) and (6.8). These species undergo subsequent hydrolysis in the presence of water, liberating formaldehyde and the amine.[17,18] This process would normally occur rapidly, but it is possible that under the conditions when the water concentration in the acid solution was only 0.05% this process was slow and rate-limiting, thus enabling direct observation of the decomposition of the iminium intermediate. This is borne out by the fact that at higher water concentrations the intermediate was not observed, it presumably having undergone rapid hydrolysis upon formation. It is difficult to ascertain whether the intermediate observed is (6.7) or (6.8), but since the hydrolysis of both would occur under the same conditions, and the subsequent decomposition of the observed species resulted in direct formation of the aniline, it is reasonable to suggest that this was (6.8) and not (6.7). It is also possible for (6.8) to undergo stabilising tautomerisation because of the presence of the hydrogen atom attached to the iminium nitrogen, thus increasing the probability of this species being observed rather than the unstabilised structure, (6.7). There is other evidence for the existence of such an intermediate in the acid catalysed hydrolysis of triazenes,[7] and it has proved possible to observe methyleniminium ions by ^{1}H NMR at room temperature.[19]

The hydroxymethyltriazene (6.9) postulated as a precursor to the dihydrogen-substituted triazene (6.4) has eluded synthesis[3] but the N-methyl derivative of (6.9) has been synthesised, and was found to decompose with elimination of formaldehyde and nitrogen, to give the aniline and methanol, probably via the methyltriazene (6.6)[6].

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CHAPTER 7

^{1}H NMR studies on the interaction of electrophiles with hexamine

7.1 Introduction

The reactions of hexamine to form its acetylated, nitrated, nitrosated and diazotised derivatives is postulated to occur via interaction with electrophilic species formed in solution during synthesis. For instance, nitration is thought to involve attack of NO_2^+, formed in the nitrating medium, on the amino nitrogen atoms of the hexamine molecule. Analogously, nitrosation will occur via attack of NO^+, diazotisation via attack of ArN_2^+, and acetylation via attack of $CH_3C\overset{\pm}{=}O$. ^{1}H NMR studies of the reaction mixtures themselves have been conducted for the acetylation of hexamine and subsequently formed acetylated derivatives,[1-4] and NMR data have been reported[5] for several compounds which have been characterised as products or intermediates in the formation of HMX (1.3) from nitration of hexamine.

It was decided that a ^{1}H NMR study of the interaction between hexamine and the electrophilic species alone, NO_2^+ from a nitronium salt for example, might give an insight into how the hexamine molecule undergoes initial ring cleavage. However, it was generally found that ring-opened species such as (7.1) or (7.2) were not observed, but rather quaternary salts of the type (7.3) (see Chapter 1).


(7.1)   (7.2)
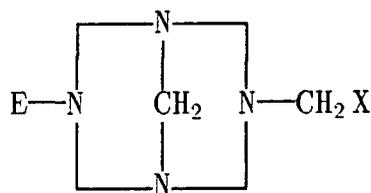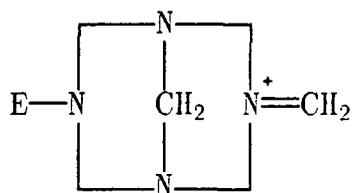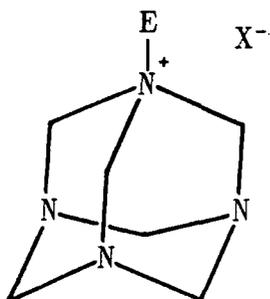
$$E = NO_2, \ NO, \ Ac \ etc$$
$$X^- = Cl^-, \ BF_4^- \ etc.$$

(7.3)

7.2 Interaction of hexamine with nitronium ion

When a commercial sample of nitronium tetrafluoroborate ($NO_2^+BF_4^-$) was dissolved in dimethyl sulphoxide-d_6 a reaction was seen to occur, involving effervescence and the formation of a clear deep red-brown solution. Addition of hexamine to this solution resulted in further effervescence and a fading of the coloration to the original colourless appearance of the DMSO-d_6. This phenomenon is explained in terms of formation of the O-nitro-sulphoxonium ion (7.4), which then undergoes reaction with hexamine, losing NO_2^+ to reform dimethyl sulphoxide.
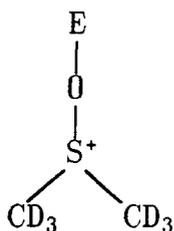

(7.4) E = NO_2
(7.5) E = NO

Formation of other nitro-substituted cations of suitable molecules, which also act as nitrating agents has been reported[6]. A ^1H NMR spectrum of nitronium tetrafluoroborate

in DMSO-d$_6$ showed only a singlet at $\delta 14.65$ typical of açidic protons.

The ^{1}H NMR spectrum of a 1:1 solution of hexamine and nitronium tetrafluoroborate in DMSO-d$_6$ at 25°C is shown in Figure 7.1. This can be explained by the formation of a species of structure (7.6).

(7.6) = NO$_2$
(7.7) = CH$_3$
(7.8) = NO
(7.9) = Ac
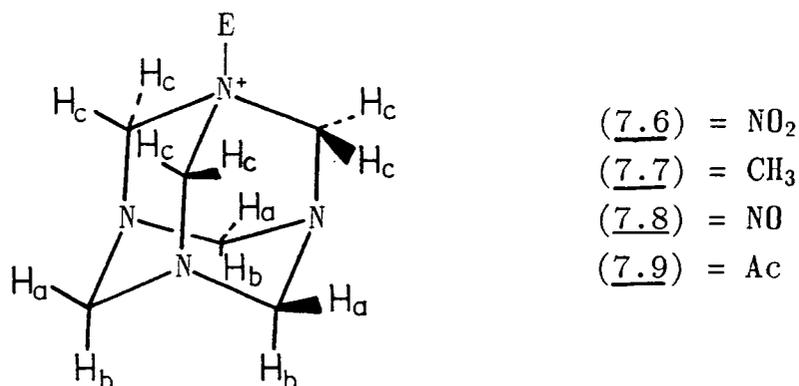
The protons H$_c$ are equivalent and give a singlet resonance at $\delta 4.92$. Protons H$_a$ and H$_b$ are not equivalent, giving rise to an AB quartet, with a doublet from protons H$_a$ centred at $\delta 4.56$, and a doublet from protons H$_b$ centred at $\delta 4.41$. The coupling constant between H$_a$ and H$_b$ is 12.5 Hz. These results compare well with those of Farminer and Webb[7], who studied the quaternary salt of structure (7.7). The ^{1}H NMR data for (7.6) is compared with that for (7.7) in Table 7.1.

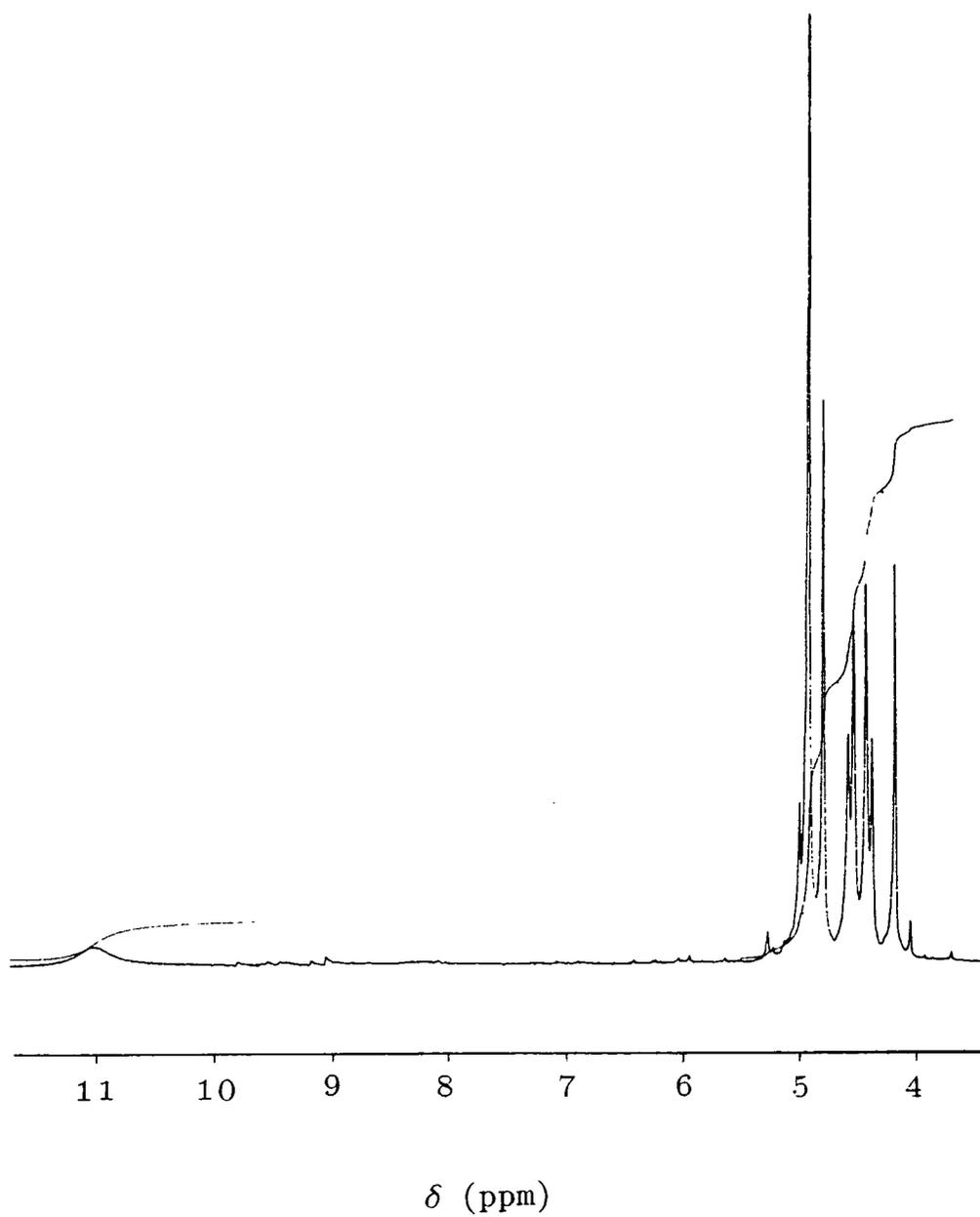Figure 7.1 ^{1}H NMR spectrum of 1:1 solution of hexamine and nitronium tetrafluoroborate in DMSO-d_6

δ (ppm)

Table 7.1 Comparison of ^{1}H NMR data for (7.6) (formed from nitronium tetrafluoroborate) and (7.7)

Compound	δ (ppm)		
	N—CH$_2$—N	N$^{\pm}$—CH$_2$—N	others
(7.6)	4.41 (d)	4.92 (s)	11.05 (br) (NH$^+$)
	4.56 (d) (J = 12.5 Hz)		5.00 (s) [protonated (7.6)]
			4.82 (s) [protonated unreacted hexamine]
			4.22 (s) [decompn. product of hexamine]
(7.7)	4.44 (d)	5.08 (s)	2.5 (s) (N$^{\pm}$—CH$_3$)
	4.64 (d) (J = 12 Hz)		

The other peaks in the spectrum have been assigned on the basis that nitronium tetrafluoroborate renders the solution acidic by reaction with water in the atmosphere and in the DMSO-d$_6$, so forming nitric acid. This accounts for the absence of a singlet water peak at $\delta 3.33$, as normally seen in DMSO-d$_6$ alone[8,9] although a signal arising from protonated water is not observed, possibly due to only a low concentration being present. A singlet is seen at $\delta 4.82$ due to mono-protonated unreacted hexamine, which is consistent with a literature value[7] for this species in the same solvent, the effect of protonation being to move the signal downfield of the value of $\delta 4.55$ measured for hexamine in DMSO-d$_6$ only. This resonance occurs as a singlet due to rapid intramolecular proton exchange. The sharp singlet at

$\delta 5.00$ is assigned to the mono-protonated form of (7.6), the other resonances of this species occurring at exactly the same shift as those of unprotonated (7.6). This is consistent with the data obtained by Farminer and Webb[7] for the mono-protonated and unprotonated forms of (7.7). Further evidence for this is the observation that the integral over the AB quartet is of exactly the same magnitude as the combined integrals over the peaks at $\delta 4.92$ and $\delta 5.00$. It is reasonable to suggest that both mono-protonated and unprotonated forms of (7.6) could exist under conditions where $[NO_2^+] \gg [H^+]$, and that rapid intermolecular proton exchange does not occur between these species in this medium, since the NH^+ resonance at $\delta 11.05$ (arising from N-protonated (7.6) and hexamine) is seen to be broad due to the exchange process being slow.

The singlet at $\delta 4.22$ has proved difficult to assign, the most reasonable conclusion reached so far being that this is due to a decomposition product of hexamine. The identity of this species remains unknown, but the concentration seems to be dependent upon the same factors as determine the concentration of (7.6), since the size of the peak at $\delta 4.22$ is observed to remain at the same ratio to those from (7.6) as the concentration of (7.6) increases.

7.3 Interaction of hexamine with nitrosonium ion

Dissolving a commercial sample of nitrosonium hydrogen sulphate ($NO^+HSO_4^-$) in DMSO-d_6 resulted in effervescence and the formation of a clear, bright green solution. This was interpreted as formation of the O-nitroso-sulphoxonium ion (7.5), in an analogous manner to the formation of (7.4). Addition of hexamine, to give a 1:1 solution, caused the green coloration to fade completely, presumbaly due to loss of the nitroso group from (7.5) to the hexamine molecule. The ^{1}H NMR spectrum of the solution is shown in Figure 7.2, and is assigned to the structure (7.8). The chemical shifts are given in Table 7.2.

Table 7.2 ^{1}H NMR data for (7.8)

	δ (ppm)	
N—CH$_2$—N	N$^{\pm}$—CH$_2$—N	others
4.47 (d)	5.02 (s)	8.1 (br) (NH$^+$)
4.61 (d) (J = 12.5 Hz)		4.85 (s) [protonated unreacted hexamine]
		4.29 (s) [decompn. product of hexamine]

Notice that no resonances are observed for a mono-protonated form of (7.8), even though the signal from unreacted hexamine is shifted downfield suggesting that the medium is acidic, as is also indicated by the broad NH$^+$ resonance at δ8.1. This might reflect the greater proton

affinity of the hexamine molecule in a solution where $[H^+]$ is low, or alternatively that protonation has occurred, but has no noticeable effect on the shifts of the H_c protons, as well as those of H_b and H_a. The singlet at $\delta 4.29$, previously assigned to a decomposition product of hexamine, is still present, indicating that the formation of this species is not dependent on the nature of the electrophile in solution, but more probably on the acidic nature of the medium.

Figure 7.2 ^{1}H NMR spectrum of a 1:1 solution of hexamine and nitrosonium hydrogen sulphate in DMSO-d_6

δ (ppm)
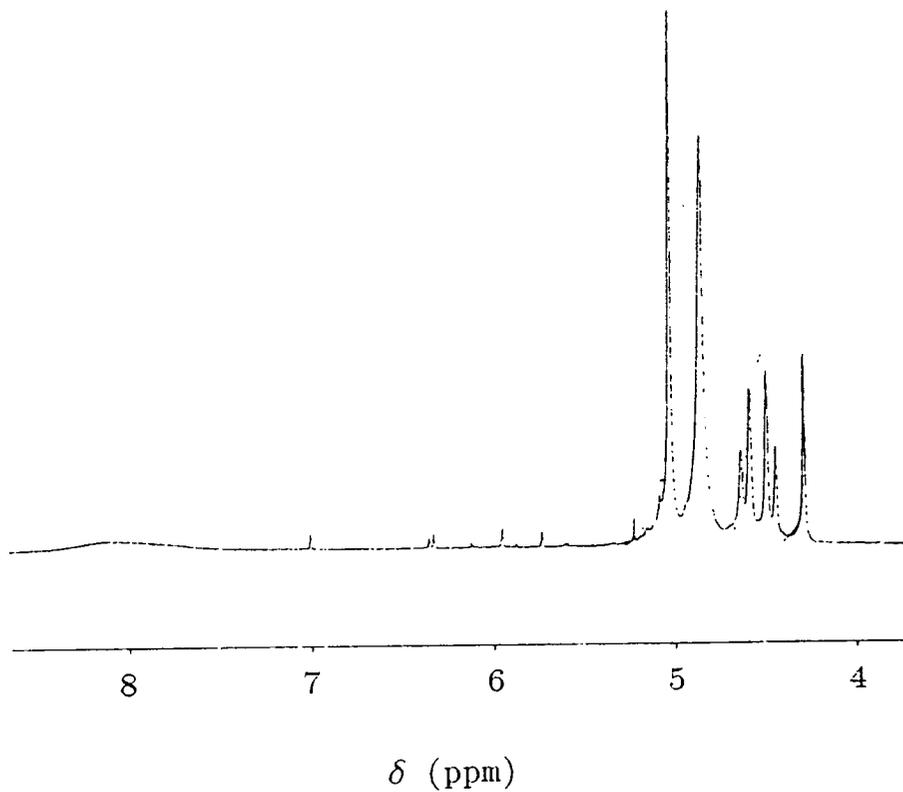
7.4 Interaction of hexamine with fuming nitric acid

Addition of several drops of fuming nitric acid to a solution of hexamine in DMSO-d_6 resulted in formation of a clear, colourless solution, the ^{1}H NMR spectrum of which (Figure 7.3) is consistent with formation of the species (7.6). This is presumably due to attack of NO_2^+ present in the nitric acid on the hexamine molecule. The peaks are much weaker than those for the same species formed from nitronium tetrafluoroborate reflecting the limited concentration of NO_2^+ available from the acid. The chemical shifts are shown in Table 7.3.

Table 7.3 ^{1}H NMR data for the species (7.6) formed with fuming nitric acid

	δ (ppm)	
N—CH$_2$—N	N$^\pm$—CH$_2$—N	others
4.40 (d)	4.94 (s)	8.71 (s) (H$_3$O$^+$)
4.54 (d)		8.0 (br) (NH$^+$)
(J = 11 Hz)		7.2 (t) (NH$^+$, J=50Hz)
		4.87 (s) [protonated unreacted hexamine]
		4.21 (s) [decompn. product of hexamine]

The spectrum is too weak to observe any formation of mono-protonated (7.6) which might be expected in an obviously acidic medium such as this. The shift of the hexamine resonance indicates protonation to have occurred, and a singlet characteristic of protonated water molecules

is seen at $\delta 8.71$. Signals from two different NH^+ functions are seen. One occurs as a broad peak at $\delta 8.0$, and the other as a 1:1:1 triplet (J_{NH} = 50 Hz) at $\delta 7.2$ due to the nitrogen atom splitting the hydrogen signal. Observation of such a triplet shows proton exchange to be non-existent at this site.

Figure 7.3

Figure 7.3 ^{1}H NMR spectrum of a solution of hexamine in DMSO-d_6 containing fuming nitric acid

δ (ppm)
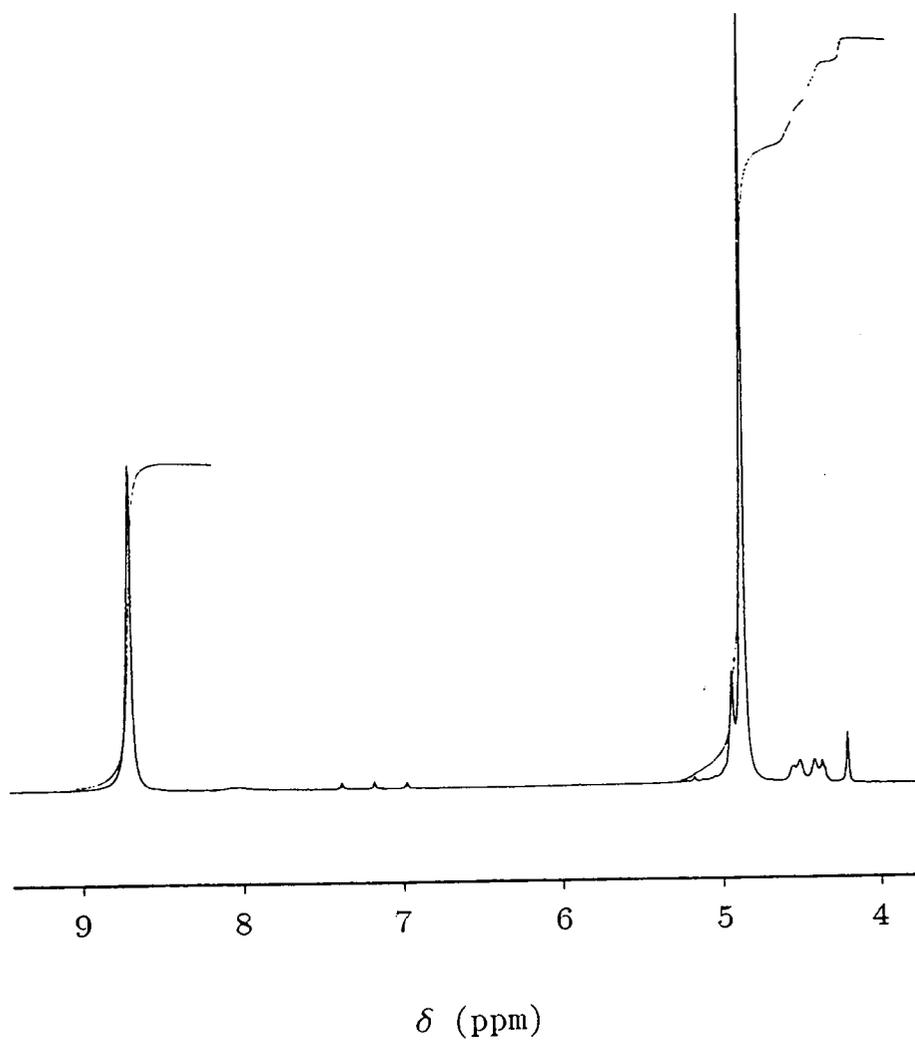
7.5 Interaction of hexamine with dinitrogen pentoxide

A sample of dinitrogen pentoxide (N_2O_5) was prepared, as described in Chapter 2. It can be a source of nitronium ion, undergoing heterolytic dissociation as shown in Equation (7.1)[10]

$$N_2O_5 \rightleftharpoons NO_2^+ + NO_3^- \qquad \text{Eqn. (7.1)}$$
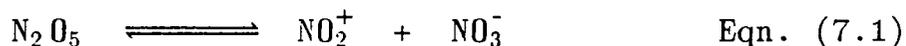
It can be used in appropriate solvents to nitrate various substrates, and is currently being investigated as a potential nitrating agent in the preparation of RDX and HMX[11]

When a sample of dinitrogen pentoxide was dissolved in DMSO-d_6 effervescence was seen to occur, resulting in formation of a clear red-brown solution, presumably due to formation of the species (7.4) proposed earlier. Addition of excess hexamine caused the solution to return to its original colourless state after further effervescence. The ^{1}H NMR spectrum of this solution, consistent with formation of (7.6), is shown in Figure 7.4, the chemical shifts being shown in Table 7.4.

Table 7.4 ^{1}H NMR data for the species (7.6) formed with dinitrogen pentoxide

δ (ppm)

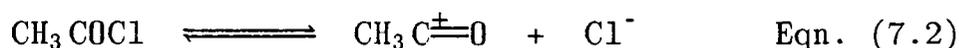| N—CH$_2$—N	N$^{\pm}$—CH$_2$—N	others	
4.46 (d)	5.00 (s)	4.76 (s)	[protonated unreacted hexamine]
4.61 (d) (J = 12.5 Hz)		4.26 (s)	[decompn. product of hexamine]

Again formation of a mono-protonated form of (7.6) is not observed in a medium which is indicated to be acidic by the shift of the hexamine resonance. The signal at δ4.26 is still present and of an intensity in the same proportion to that of the signals from (7.6) as seen in earlier spectra.

Figure 7.4 ^{1}H NMR spectrum of a solution of excess hexamine and dinitrogen pentoxide in DMSO-d_6

δ (ppm)

7.6 Interaction of hexamine with acetyl chloride

Addition of acetyl chloride to a solution of hexamine (1:1) in dry chloroform resulted in the immediate precipitation of a white crystalline solid, which when filtered off was found to be unstable in air. It was previously postulated[1] that this compound had structure (7.9), the anion of the salt being Cl^-, and the ^{1}H NMR spectrum in DMSO-d_6 (Figure 7.5) appears to confirm this, showing similar characteristics to the species (7.6)-(7.8). Such formation can only occur if acetyl chloride undergoes the dissociation shown in Equation (7.2).

$$CH_3COCl \rightleftharpoons CH_3C^{\pm}=O + Cl^- \qquad \text{Eqn. (7.2)}$$

The chemical shifts measured are shown in Table 7.5.

Table 7.5 ^{1}H NMR data for the species (7.9), formed as the salt, from 1:1 hexamine/acetyl chloride in chloroform, measured in DMSO-d_6

δ (ppm)

N—CH$_2$—N	N$^{\pm}$—CH$_2$—N	others
4.44 (d)	5.01 (d)	8.30 (s)
4.59 (d)		7.2 (t) (NH$^+$, J=50Hz)
(J = 12.5 Hz)		
		4.82 (s) [protonated hexamine]
		3.42 (s) [H$_2$O in solvent]
		2.00 (s) [acetyl of (7.9) ?]

Figure 7.5 ^{1}H NMR spectrum of species (7.9), formed as the salt, from 1:1 hexamine/acetyl chloride in chloroform, measured in DMSO-d_6

δ (ppm)
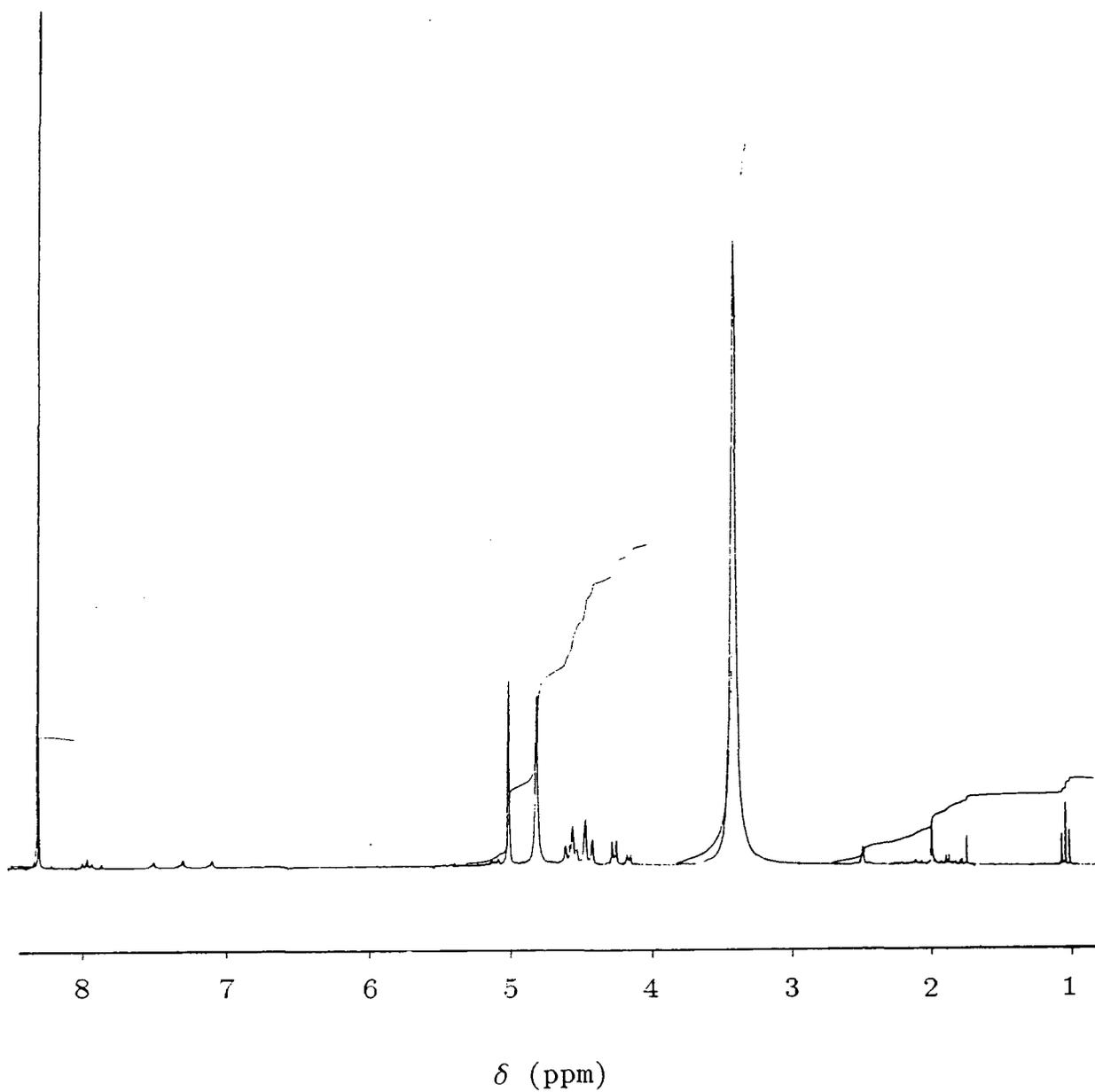
No evidence of mono-protonated ($\underline{7.9}$) is seen. The presence of a protonated hexamine peak is probably due to some formation of the acetate salt of hexamine by reaction of ($\underline{7.9}$) with water in the air and in the solvent to form acetic acid. An indication that this has only occurred to a limited extent is that the water peak at $\delta 3.42$ is at a shift close to that expected for unprotonated H_2O in DMSO-d_6. The sharp singlet at $\delta 8.30$ is assigned to an unknown decomposition product of ($\underline{7.9}$), possibly an aldehyde. Various other small peaks seen in the spectrum cannot be assigned, but are probably due to decomposition products.

Addition of excess acetyl chloride to a solution of hexamine in DMSO-d_6 (1:1) resulted in formation of a clear, colourless solution which gave the ^{1}H NMR spectrum shown in Figure 7.6. The chemical shifts are given in Table 7.6, and indicate formation of the species ($\underline{7.9}$).

Table 7.6 ^{1}H NMR data for (7.9) formed in DMSO-d_6 solution

from hexamine and excess acetyl chloride

δ (ppm)

N—CH$_2$—N	N$^{\pm}$CH$_2$—N	others	
4.41 (d)	5.04 (s)	5.31 (s)	[protonated (7.9) ?]
4.54 (d) (J = 12.5 Hz)		5.17 (s)	[protonated (7.9) ?]
		4.21 (s)	[decompn. product of hexamine]
		2.19 (s)	
		1.97 (s)	
		1.84 (s)	[acetyl of (7.9) ?]

The singlet peaks seen at δ5.31 and δ5.17 may arise from the H$_c$ protons of protonated forms of (7.9), and indeed minor peaks can be seen in the region of the AB quartet which also indicate this. The acidic medium required for this might arise from hydrolysis of acetyl chloride by water present in the solvent to form acetic acid. Also, it is noteworthy that the integral over the AB quartet region matches exactly in magnitude the integral over the three signals at δ5.04, δ5.17 and δ5.31, also providing evidence that this has occurred. Also, there appears to be more than one peak attributable to acetyl groups. The small singlets at δ2.19 and δ1.97 may be due to the acetyl groups of protonated forms of (7.9), although one may be due to acetic acid. Notice that the resonance from protonated excess hexamine is absent, indicating it to have all undergone reaction.

Figure 7.6 ^{1}H NMR spectrum of species (7.9) formed in
DMSO-d_6 solution from hexamine and excess acetyl
chloride

δ (ppm)
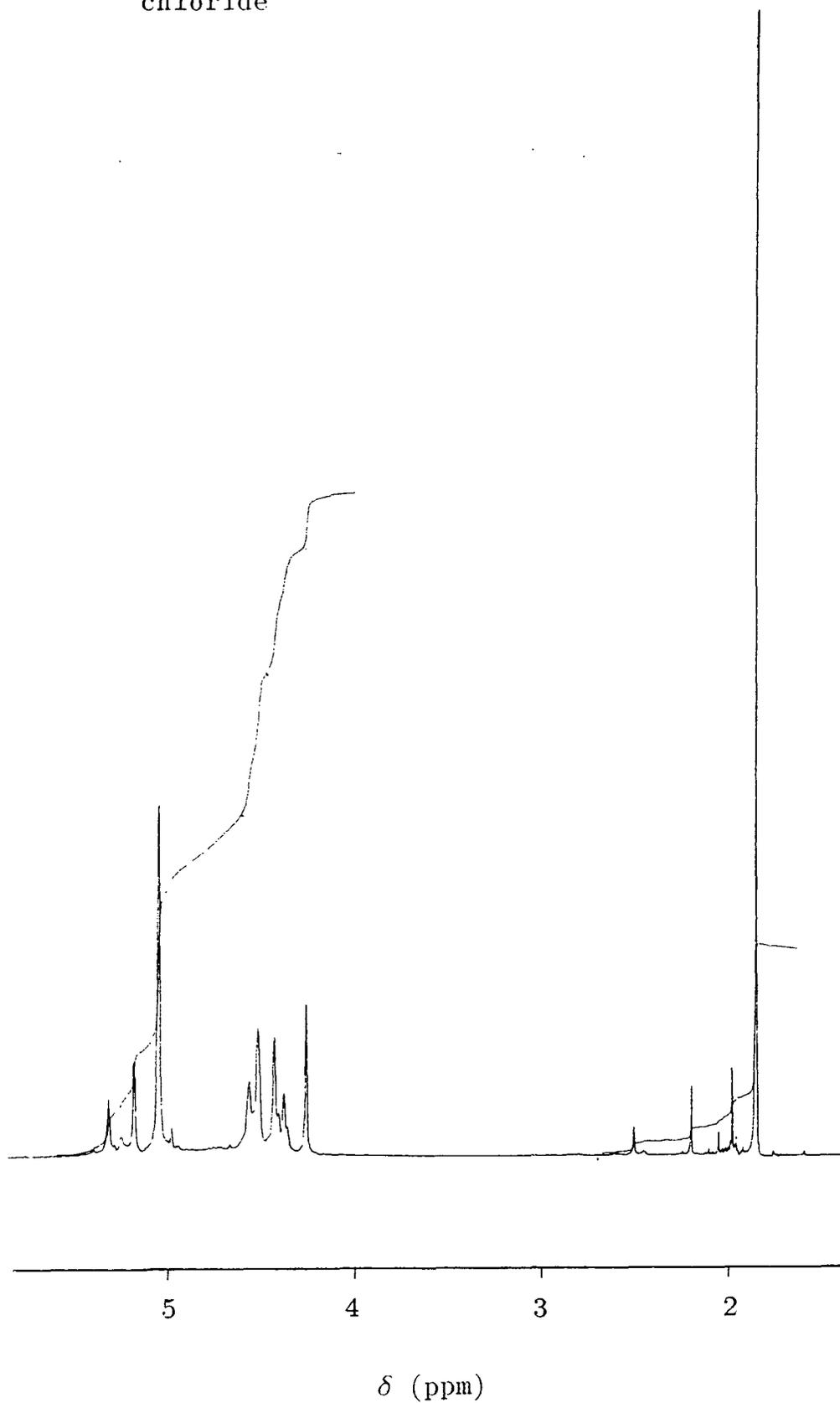
7.7 Interaction of DAPT with acetyl chloride

It was thought that a comparison of the interaction of a bicyclic hexamine derivative, such as DPT (1.5) or DAPT (1.6), with the same electrophiles used with hexamine might be useful, as this might yield further information as to the mechanism of formation of the unbridged eight-membered ring species, such as HMX (1.3), TAT (1.8) and DADN (1.9). However, ^{1}H NMR spectra obtained were generally very complex and uninterpretable, perhaps indicative of extensive decomposition having taken place. The one exception to this was the spectrum obtained when excess acetyl chloride was added to a solution of DAPT in chloroform-d_6, giving a clear colorless solution. This is shown in Figure 7.7, with shifts listed as singlet resonances in Table 7.7.

Table 7.7 ^{1}H NMR data for a CDCl$_3$ solution of DAPT and excess acetyl chloride

δ (ppm)

species formed	others
5.45 (s, 2H)	2.66 (s) $\begin{bmatrix} \text{excess} \\ \text{AcCl} \end{bmatrix}$
5.13 (s, 2H)	
5.06 (s, 2H)	
4.59 (s, 2H)	
4.55 (s, 2H)	
2.37 (s, 3H)	
2.18 (s, 3H)	
2.16 (s, 3H)	

Figure 7.7 ^{1}H NMR spectrum of a CDCl$_3$ solution of DAPT and excess acetyl chloride

δ (ppm)
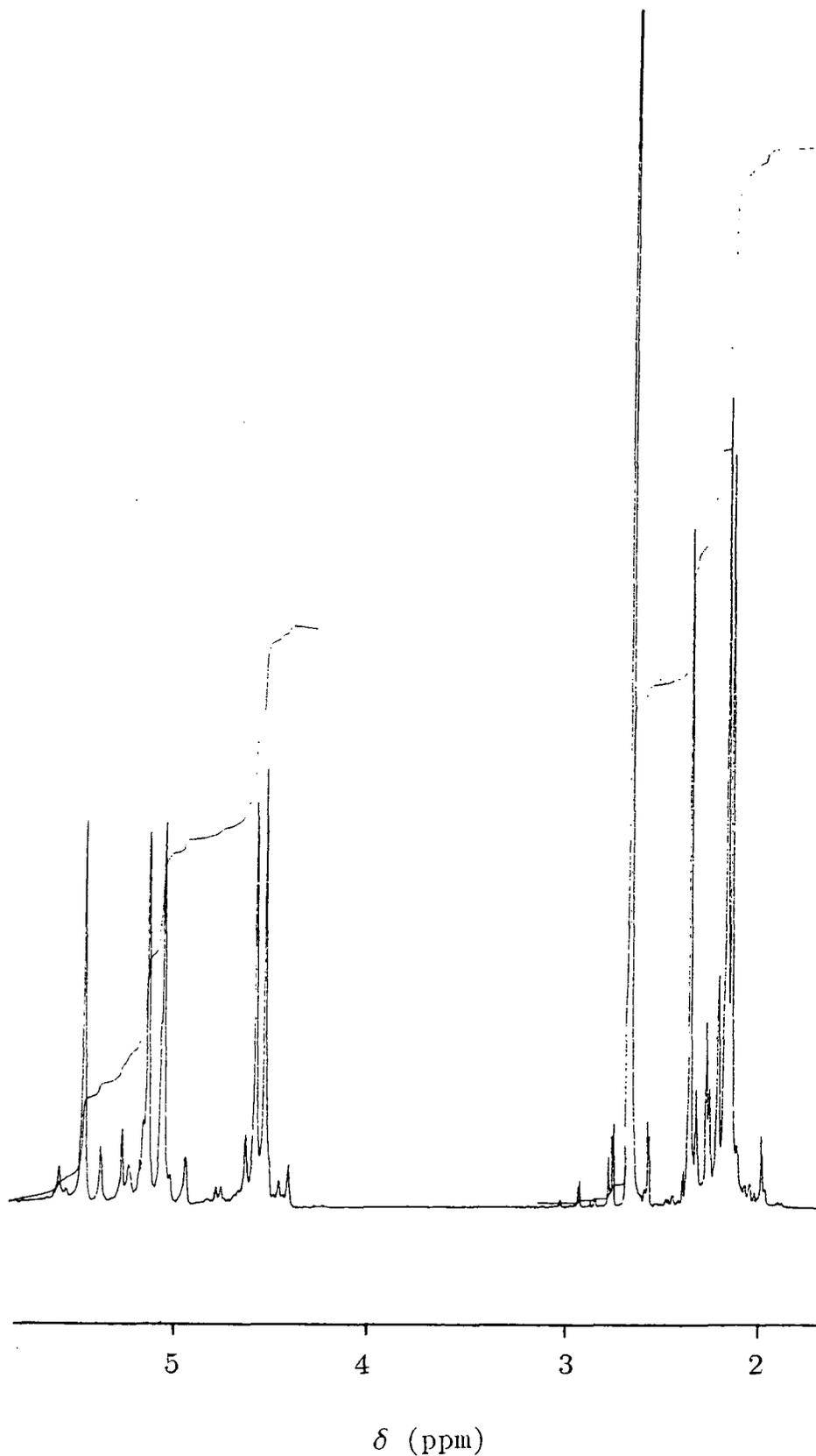
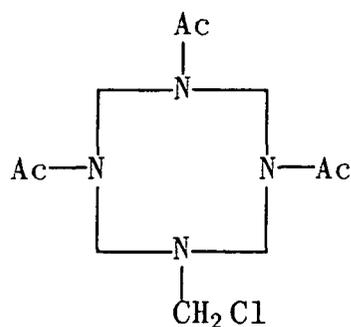Although the spectrum appears quite simple, it is not easily interpreted. At first it might appear that there is a doublet centred at $\delta 5.09$ and another centred at $\delta 4.57$. However, this cannot be the case, since no apparent coupling exists between any of the signals. The most likely species formed by reaction of DAPT and acetyl chloride are (7.10) and (7.11).
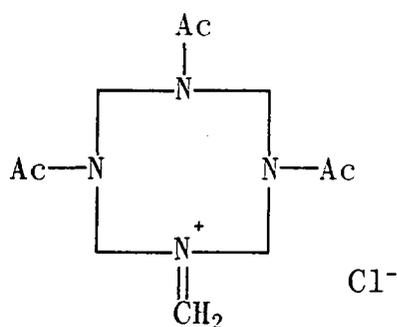
(7.10)

(7.11)

The spectrum is not that of (7.10), since for this structure a signal from the methyleniminium protons would be expected at $\delta 8-9$.[12] At first sight it might also seem unreasonable to suggest that Figure 7.7 is the spectrum of species (7.10), since one might expect there to be either two signals from ring methylene protons if molecular flexing occurs, or two AB quartets if the ring is more rigid, the former being most likely by comparison with other similar structures. However, it is possible, even if protons on any one methylene group are equivalent due to molecular flexing, that hindered rotation of the N-acetyl groups and/or steric hindrance about the Cl-atom leads to there being five different proton environments if one includes the non-ring methylene group, and hence the five singlets seen at $\delta 4.55$,

$\delta4.59$, $\delta5.06$, $\delta5.13$ and $\delta5.45$. It is reasonable to suggest that the 'lone' singlet at $\delta5.45$ is due to the methylene group adjacent to the Cl-atom, whilst the two pairs of singlets further upfield arise from very similar, but non-equivalent, methylene protons either on the same side of the ring as the Cl-atom, or opposite it. Such a scenario would also explain the presence of three different acetyl peaks at $\delta2.16$, $\delta2.18$ and $\delta2.37$. The singlet at $\delta2.66$ is attributable to excess acetyl chloride, since when excess DAPT is present this peak does not appear in the spectrum.

7.8   References

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7.   A.R. Farminer and G.A. Webb, *Organic Magnetic Resonance*, 1976, <u>8</u>, 102.

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9.   D.H. Williams and I. Fleming, 'Spectroscopic methods in organic chemistry', 3rd. edn., McGraw-Hill, London, 1980, p.143.

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11.   P. Golding, Ministry of Defence, private communication.

12.   J.B. Lambert, G. Wang, D.E. Huseland and L.C. Takiff, *J.Org.Chem.*, 1987, <u>52</u>, 68.

APPENDIX I

RESEARCH COLLOQUIA, SEMINARS, LECTURES

AND CONFERENCES

The Board of Studies in Chemistry requires that each postgraduate research thesis contains an appendix listing:

(A)   all research colloquia, seminars and lectures arranged by the Department of Chemistry during the period of the author's residence as a postgraduate student;

(B)   lectures organised by Durham University Chemical Society;

(C)   all research conferences attended and papers presented by the author during the period when research for the thesis was carried out;

(D)   details of the postgraduate induction course.

(A) <u>Research colloquia, seminars and lectures organised by</u>

<u>Durham University Chemistry Department, 1986-1989</u>

(those attended are marked *)

29.10.86   Prof. E.H. Wong (University of New Hampshire, U.S.A.), 'Coordination Chemistry of P-O-P Ligands'. (*)

5.11.86   Prof. D. Dopp (University of Duisburg), 'Cyclo-additions and Cyclo-reversions Involving Captodative Alkenes'. (*)

26.11.86   Dr. N.D.S. Canning (University of Durham), 'Surface Adsorption Studies of Relevance to Heterogeneous Ammonia Synthesis'. (*)

3.12.86   Dr. J. Miller (Dupont Central Research), 'Molecular Ferromagnets: Chemistry and Physical Properties'. (*)

8.12.86   Prof. T. Dorfmuller (University of Bielefeld), 'Rotational Dynamics in Liquids and Polymers'.

28.1.87   Dr. W. Clegg (University of Newcastle-upon-Tyne), 'Carboxylate Complexes of Zinc: Charting a Structural Jungle'.

4.2.87   Prof. A. Thomson (University of East Anglia), 'Metalloproteins and Magnetooptics'.

11.2.87   Dr. T. Shepherd (University of Durham), 'Pteridine Natural Products: Synthesis and Use in Chemotherapy'. (*)

17.2.87   Prof. E.H. Wong (University of New Hampshire, U.S.A.), 'Symmetrical Shapes from Molecules to Art and Nature'.

4.3.87   Dr. R. Newman (University of Oxford), 'Change and Decay: A Carbon-13 CP/MAS NMR Study of Humification and Coalification Processes'.

11.3.87   Dr. R.D. Cannon (University of East Anglia), 'Electron Transfer in Polynuclear Complexes'.

17.3.87   Prof R.F. Hudson (University of Kent), 'Aspects of Organophosphorus Chemistry'.

18.3.87   Prof. R.F. Hudson (University of Kent), 'Homolytic Rearrangements of Free Radical Stability'.

6.5.87   Dr. R. Bartsch (University of Sussex), 'Low Co-ordinated Phosphorus Compounds'. (*)

7.5.87   Dr. M. Harmer (I.C.I. Chemicals & Polymer Group), 'The Role of Organometallics in Advanced Materials'.

11.5.87   Prof. S. Pasynkiewicz (Technical University, Warsaw), 'Thermal Decomposition of Methyl Copper and its Reactions with Trialkylaluminium'.

27.5.87   Dr. R.M. Blackburn (University of Sheffield), 'Phosphonates as Analogues of Biological Phosphate Esters'.

24.6.87   Prof. S.M. Roberts (University of Exeter), 'Synthesis of Novel Antiviral Agents'. (*)

26.6.87   Dr. C. Krespan (E.I. Dupont de Nemours), 'Nickel (0) and Iron (0) as Reagents in Organofluorine Chemistry'.

4.11.87   Mrs. M. Mapletoft (Durham Chemistry Teachers' Centre), 'Salters' Chemistry'.

19.11.87   Dr. J. Davidson (Herriot-Watt University), 'Metal Promoted Oligemerisation Reactions of Alkynes'. (*)

10.12.87   Dr.C.J. Ludman (University of Durham), 'Explosives'. (*)

16.12.87   Mr. R.M. Swart (I.C.I.), 'The Interaction of Chemicals with Lipid Bilayers'.

16.3.88   Mr. L. Bossons (Durham Chemistry Teachers' Centre), 'GSCE Practical Assessment'.

7.4.88   Prof. M.P. Hartshorn (University of Canterbury, New Zealand), 'Aspects of Ipso-Nitration'. (*)

13.4.88   Mrs. E. Roberts (SATRO Officer for Sunderland), Talk - Durham Chemistry Teachers' Centre, 'Links Between Industry and Schools'.

18.4.88   Prof. C.A. Nieto de Castro (University of Lisbon and Imperial College), 'Transport Properties of Non-polar Fluids'. (*)

19.4.88   Graduate Chemists (Northeast Polytechnics and Universities), R.S.C. Graduate Symposium. (*)

25.4.88   Prof. D. Birchall (I.C.I Advanced Materials), 'Environmental Chemistry of Aluminium'.

27.4.88   Dr. J.A. Robinson (University of Southampton), 'Aspects of Antibiotic Biosynthesis'.

27.4.88   Dr. R. Richardson (University of Bristol), 'X-Ray Diffraction from Spread Monolayers'.

28.4.88   Prof. A. Pines (University of California, Berkeley, U.S.A.), 'Some Magnetic Moments'.

11.5.88   Dr. W.A. McDonald (I.C.I. Wilton), 'Liquid Crystal Polymers'. (*)

11.5.88   Dr. J. Sodeau (University of East Anglia), Durham Chemistry Teachers' Centre Lecture, 'Spray Cans, Smog and Society'.

8.6.88   Prof. J.-P. Majoral (Universite Paul Sabatier), 'Stabilisation by Complexation of Short-Lived Phosphorus Species'.

29.6.88   Prof. G.A. Olah (University of Southern California), 'New Aspects of Hydrocarbon Chemistry'. (*)

18.10.88   Dr. J. Dingwall (Ciba Geigy), 'Phosphorus-containing Amino Acids: Biologically Active Natural and Unnatural Products'.

18.10.88   Mr. F. Bollen (Durham Chemistry Teachers' Centre), 'The Use of SATIS in the classroom'.

18.10.88   Dr. C.J. Ludman (Durham University), 'The Energetics of Explosives' (*).

9.11.88   Dr. G. Singh (Teesside Polytechnic), 'Towards Third Generation Anti-Leukaemics'.

16.11.88   Dr. K.A. McLauchlan (University of Oxford), 'The Effect of Magnetic Fields on Chemical Reactions'.

2.12.88   Dr. G. Hardgrove (St. Olaf College, U.S.A.), 'Polymers in the Physical Chemistry Laboratory'.

9.12.88   Dr. C. Jaeger (Friedrich-Schiller University GDR), 'NMR investigations of Fast Ion Conductors of the NASICON Type'.

14.12.88   Dr. C. Mortimer (Durham University Teachers' Centre), 'The Hindenberg Disaster - An Excuse for Some Experiments'

25.1.89   Dr. L. Harwood (University of Oxford), 'Synthetic Approaches to Phorbols Via Intramolecular Furan Diels-Alder Reactions: Chemistry Under Pressure

1.2.89   Mr. T. Cressey and Mr. D. Waters (Durham Chemistry Teachers' Centre), 'GCSE Chemistry 1988: A Coroner's Report'.

13.2.89   Prof. R.R. Schrock (M.I.T.), 'Recent Advances in Living Metathesis' (*).

15.2.89   Dr. A.R. Butler (St. Andrews University), 'Cancer in Linxiam: The Chemical Dimension'.

22.2.89   Dr. G. MacDougall (Edinburgh University), 'Vibrational Spectroscopy of Model Catalytic Systems'.

1.3.89   Dr. R.J. Errington (University of Newcastle-upon-Tyne), 'Polymetalate Assembly in Organic Solvents'.

9.3.89   Dr. I. Marko (Sheffield University), 'Catalytic Asymmetric Osmylation of Olefins'.

14.3.89   Mr. P. Revell (Durham Chemistry Teachers' Centre), 'Implementing Broad and Balanced Science 11-16'.

15.3.89   Dr. R. Aveyard (University of Hull), 'Surfactants at your Surface'.

20.4.89   Dr. M. Casey (University of Salford), 'Sulphoxides in Stereoselective Synthesis'.

27.4.89   Dr. D. Crich (University College London), 'Some Novel Uses of Free Radicals in Organic Synthesis' (*).

3.5.89   Mr. A. Ashman (Durham Chemistry Teachers' Centre), 'The Chemical Aspects of the National Curriculum'.

3.5.89   Dr. P.C.B. Page (University of Liverpool), 'Stereocontrol of Organic Reactions Using 1,3-dithiane-1-oxides'.

10.5.89   Prof. P.B. Wells (Hull University), 'Catalyst Characterisation and Activity'.

11.5.89   Dr. J. Frey (Southampton University, 'Spectroscopy of the Reaction Path: Photodissociation Raman Spectra of NOCl'.

16.5.89   Dr. R. Stibr (Czechoslovak Academy of Sciences), 'Recent Developments in the Chemistry of Intermediate-Sited Carboranes'.

17.5.89   Dr. C.J. Moody (Imperial College), 'Reactive Intermediates in Heterocyclic Synthesis' (*).

23.5.89   Prof. P. Paetzold (Aachen), 'Iminoboranes $XB\equiv NR$: Inorganic Acetylenes ?'.

14.6.89   Dr. M.E. Jones (Durham Chemistry Teachers' Centre), 'GCSE and A-level Chemistry 1989'.

15.6.89   Prof. J. Pola (Czechslovak Academy of Sciences), 'Carbon Dioxide Laser Induced Chemical Reactions - New Pathways in Gas-Phase Chemistry'.

28.6.89   Dr. M.E. Jones (Durham Chemistry Teachers' Centre), 'GCSE and A-level Chemistry 1989'.

11.7.89   Dr. D. Nicholls (Durham Chemistry Teachers' Centre), 'Liquid Air Demonstration'.

(B) <u>Lectures organised by Durham University Chemical</u>

   <u>Society 1986-1989</u>

(those attended are marked *)

16.10.86 Prof. N.N. Greenwood (University of Leeds),
          'Glorious Gaffes in Chemistry'. (*)

23.10.86 Prof. H.W. Kroto (University of Sussex),
          'Chemistry in Stars, between Stars and in the
          Laboratory'.

30.10.86 Prof. D. Betteridge (B.P. Research), 'Can
          Molecules Talk Intelligently'.

 6.11.86 Dr. R.M. Scrowston (University of Hull), 'From
          Myth and Magic to Modern Medicine'. (*)

13.11.86 Prof. Sir G. Allen (Unilever Research),
          'Biotechnology and the Future of the Chemical
          Industry'.

20.11.86 Dr. A. Milne and Mr. S. Christie (International
          Paints), 'Chemical Serendipity - A Real Life Case
          Study'.

27.11.86 Prof. R.L. Williams (Metropolitan Police Forensic
          Science), 'Science and Crime'. (*)

22.1.87   Prof. R.H. Ottewill (University of Bristol),
          'Colloid Science: A Challenging Subject'.

 5.2.87   Dr. P. Hubberstey (University of Nottingham),
          'Demonstration Lecture on Various Aspects of
          Alkali Metal Chemistry'. (*)

12.2.87   Dr. D. Brown (I.C.I. Billingham), 'Industrial
          Polymers from Bacteria'.

19.2.87   Dr. M. Jarman (Institute of Cancer Research), 'The
          Design of Anti-Cancer Drugs'. (*)

 5.3.87   Prof. S.V. Ley (Imperial College), 'Fact and
          Fantasy in Organic Synthesis'.

 9.3.87   Prof. F.G. Bordwell (Northeastern University,
          U.S.A.), 'Carbon Anions, Radicals, Radical Anions
          and Radical Cations'. (*)

12.3.87   Dr. E.M. Goodger (Cranfield Institute of
          Technology), 'Alternative Fuels for Transport'.
          (*)

15.10.87 Dr. M.J. Winter (University of Sheffield),
          'Pyrotechnics (Demonstration Lecture)'. (*)

22.10.87 Prof. G.W. Gray (University of Hull), 'Liquid
          Crystals and their Applications'. (*)

29.10.87 Mrs. S. van Rose (Geological Museum), 'Chemistry of Volcanoes'. (*)

5.11.87 Dr. A.R. Butler (University of St. Andrews), 'Chinese Alchemy'. (*)

12.11.87 Prof. D. Seebach (E.T.H. Zurich), 'From Synthetic Methods to Mechanistic Insight'. (*)

19.11.87 Prof. P.G. Sammes (Smith, Kline and French), 'Chemical Aspects of Drug Development'. (*)

26.11.87 Dr. D.H. Williams (University of Cambridge), 'Molecular Recognition'. (*)

3.12.87 Dr. J. Howard (I.C.I. Wilton), 'Liquid Crystal Polymers'. (*)

21.1.88 Dr. F. Palmer (University of Nottingham), 'Luminescence (Demonstration Lecture)'. (*)

28.1.88 Dr. A. Cairns-Smith (University of Glasgow), 'Clay Minerals and the Origin of Life'.

11.2.88 Prof. J.J. Turner (University of Nottingham), 'Catching Organometallic Intermediates'. (*)

18.2.88 Dr. K. Borer (University of Durham Industrial Research Laboratories), 'The Brighton Bomb - A Forensic Science View'. (*)

25.2.88 Prof. A. Underhill, (University of Bangor), 'Molecular Electronics'. (*)

3.3.88 Prof. W.A.G. Graham (University of Alberta, Canada), 'Rhodium and Iridium Complexes in the Activation of Carbon-Hydrogen Bonds'.

6.10.88 Prof. R. Schmutzler (University of Braunschweig), 'Fluorophosphines Revisited - New Contributions to an Old Theme'.

13.10.88 Dr. D.R. Marshall (University of Bangor), 'The Best Bang Since the Big One'. (*)

21.10.88 Prof. P. von Rague Schleyer (University of Erlangen), 'The Fruitful Interplay Between Calculational and Experimental Chemistry'.

27.10.88 Prof. W.C. Rees (Imperial College), 'Some Very Heterocyclic Compounds'. (*)

10.11.88 Prof. J.I.G. Cadogan (B.P. Research), 'From Pure Science to Profit'. (*)

24.11.88 Dr. R.W. Walker and Dr. R.R. Baldwin (University of Hull), 'Combustion - Some Burning Problems'. (*)

1.12.88 Dr. R. Snaith (University of Cambridge), 'Egyptian Mummies - What, Where, Why and How ?' (*)

26.1.89 Prof. K.R. Jennings (University of Warwick), 'Chemistry of the Masses'.

2.2.89 Prof. L.D. Hall (Addenbrookes' Hospital), 'NMR - A Window to the Human Body'. (*)

9.2.89 Prof. J. Baldwin (University of Oxford), 'Recent Advances in the Bioorganic Chemistry of Penicillin Biosynthesis'. (*)

16.2.89 Prof. J.B. Aylett (Queen Mary College), 'Silicon-based Chips: The Chemists Contribution'. (*)

23.2.89 Dr. B.F.G. Johnson (University of Cambridge), 'The Binary Carbonyls'.

(C) Conferences attended

European Symposium on Organic Reactivity II, University of Padova, Italy, 27th August - 1st September 1989.
Paper presented: 'Reactions of hexamethylenetetramine and its derivatives with electrophiles'.

(D) First year induction course, October 1986

This course consists of a series of one hour lectures on the services available in the department.

1. Departmental organisation.

2. Safety matters.

3. Electrical appliances and infra-red spectroscopy.

4. Chromatography and microanalysis.

5. Atomic absorptiometry and inorganic analysis.

6. Library facilities.

7. Mass spectroscopy.

8. Nuclear magnetic resonance spectroscopy.

9. Glassblowing technique.