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The Pyrolysis of Some Amine Salts of 2,4-Dichlorophenoxyacetic Acid

Shane S. Que Hee and Ronald G. Sutherland*

Amine salts of 2,4-dichlorophenoxyacetic acid (2,4-D) were pyrolyzed at various temperatures. Amide production was first observed at 80° by mass spectroscopy. The fastest rate of formation of the corresponding amide was between 140–150° for the methyl and *n*-butylamine salts, between 150–170° for the *n*-dodecyl and *n*-tetradecylamine salts, and above 160° for the dimethylamine salt. Beyond 160°, 2,4-dichlorophenol, imines, lactones, and other compounds were also formed.

The amides appeared to decompose completely above 200°. The activation energy for amide formation was calculated by differential scanning calorimetry (DSC), utilizing a new manipulative technique involving the Boltzmann distribution. The activation energies were found to be (44 ± 1) and (38 ± 3) kcal mol⁻¹ for the *n*-butyl and *n*-tetradecyl salts, respectively. The heats of fusion and stoichiometry of the salts were also determined.

Amine salts of 2,4-dichlorophenoxyacetic acid (2,4-D) are used widely for herbicidal applications. There is an extensive patent literature dealing with their synthesis (Dickson, 1952; Fischer, 1958; Harwood *et al.*, 1956). However, most of these patents often do not specify at what temperature the salt should be prepared or at what temperature amide production begins. In view of possible differences in biological effect between the salt and the corresponding amide, a detailed investigation of the rate of dehydration of the salts with respect to temperature was undertaken.

Reagents. Commercial solid *n*-tetradecyl- and *n*-dodecylamines (Aldrich) and liquid *n*-butylamine were purified by vacuum distillation. Reagent grade dimethylamine (Eastman) was used without further purification. A 20% aqueous solution of methylamine was used for the preparation of the methylamine salt. Commercial 2,4-D (Aldrich) was recrystallized from benzene until a constant melting point of 140.3 ± 0.5° was attained.

1:1 salts of 2,4-D were made by adding stoichiometric amounts of the amines dissolved in benzene-acetone solutions to solid 2,4-D at 10°. The solutions were shaken until all the 2,4-D had disappeared. The solvent was removed under vacuum at room temperature. The resultant salts were recrystallized from 1:1:1 ether-acetone-hexane until the melting points were constant.

EXPERIMENTAL SECTION

Pyrolysis Experiments. The solid salts were pyrolyzed in sealed Pyrex tubes covered with aluminum foil. Separate samples were pyrolyzed for 1 hr at temperatures from 30 to 200°. The tubes were then cooled before recording mass, infrared, and nuclear magnetic spin resonance spectra of the products.

Differential Scanning Calorimetry. Known amounts of salts were pyrolyzed in aluminum planchettes within

the heating chamber of a Perkin-Elmer differential scanning calorimeter DSC-1B. The pyrolysis was carried out at a constant heating rate of 10°/min using range, slope, average, and differential settings of 16, 210, 405, and 470. The heating chamber was purged by a slow flow of nitrogen. Pure benzoic acid was used as a standard to compute the heats of fusion of the salts.

To obtain the activation energy for salt decomposition, weighed samples of the butyl- and tetradecylamine salts in the planchettes were pyrolyzed past the melting point to a temperature, the "prepyrolysis temperature," where the baseline became stable again. The planchettes were cooled immediately to ambient temperature and then repyrolyzed at the same heating rate to a higher temperature, approximately 10° above the initial prepyrolysis temperature. This higher temperature became the new "prepyrolysis temperature." This cooling and heating procedure was repeated until the heat of fusion peak at the melting point became negligible. As the heat of fusion is an extensive property, the area under the curve representing the heat of fusion is proportional to the mass of salt melted, providing eutectic or solid solution does not occur. Thus, the mass of salt which has not decomposed up to the preceding "prepyrolysis temperature" can be found, as long as negligible decomposition occurs at the melting point.

The mass of salt decomposed up to the "prepyrolysis temperature" can thus be found by subtraction from the known initial amount. During the pyrolyses, no solid solution formation occurred as the melting point of each salt remained constant until amide formation was complete. As the heat of fusion decreased smoothly as pyrolysis proceeded, the likelihood of pyrolysis products having the same melting point as the salt was assumed to be negligible. Mass spectral examination also verified that very little salt was present when the peak at the melting point disappeared.

The above approach was modified so that the Boltzmann distribution could be used.

$$N_i = N_o e^{-\epsilon_i/RT} \quad (1)$$

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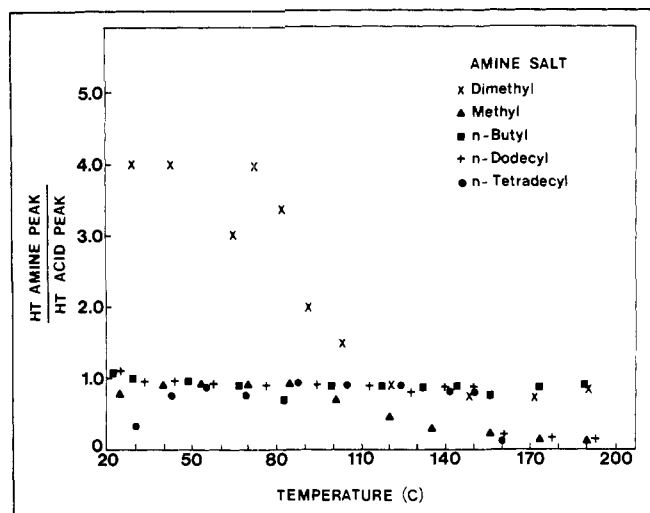


Figure 1. The stoichiometry of solid methyl-, *n*-butyl-, *n*-dodecyl-, *n*-tetradecyl-, and dimethylamine salts of 2,4-D at various temperatures.

where N_i is the number of molecules in state i whose potential energy is ϵ_i above that of the original state o . Equation 1 can be rewritten

$$N_A/N_B = e^{-\epsilon_i/RT} \quad (2)$$

Equation 2 gives the ratio of the number (N_A) of molecules at temperature T and time t which have passed over the energy barrier ϵ_i to the number of molecules of the same compound which have not (N_B). This energy barrier ϵ_i should be the activation energy. Using this approach, activation energies obtained using the Arrhenius relation can be double checked. Hitherto, this has not been possible. Equation 2 can be rewritten

$$m_A/m_B = e^{-\epsilon_i/RT} \quad (3)$$

$$X_A/X_B = e^{-\epsilon_i/RT} \quad (4)$$

$$X_A/X_B = k/A \quad (5)$$

where m and X signify mass and fraction of original mass, respectively, and k and A are the rate constant and the Arrhenius frequency factor, respectively, at temperature T .

Thus a plot of $\log(X_A/X_B)$ vs. $1/T$ should have a slope of $-\epsilon_i/2.303R$.

Accurate values of X_A and X_B require accurate values of the area corresponding to the original mass. For the *n*-butyl- and *n*-tetradecylamine salts, the area under the curve at the melting point was constant until 120°, thus allowing an accurate calculation of the area corresponding to the original mass.

The dimethyl- and methylamine salts could not be examined using this technique as decomposition occurred before and at their melting points.

RESULTS AND DISCUSSION

Though 1:1 amine-acid was reacted, this constitutes no proof that 1:1 salts were actually formed. The mass spectrum of an organic salt is usually characterized by the absence of any parent ion and, if both acid and basic parts are weak, by the spectrum's being a combination of the spectra of the two organic constituents. However, at low eV of the ion source (kept at 70° to prevent any amide formation), the only peaks observed were the parent ions of the amine and the acid, and the ratio of the peak heights of amine and acid parent ions should be related to the stoichiometry of the salt. That this in fact was the case is

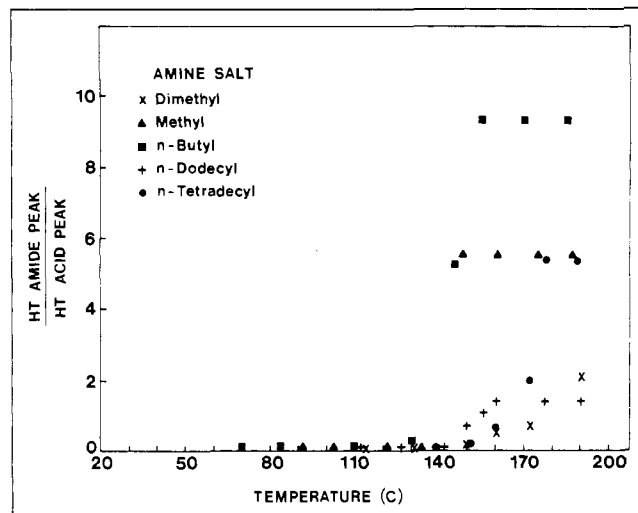


Figure 2. The extent of production of amide from corresponding amine salts of 2,4-D at various temperatures after 1 hr using the peak heights from low eV mass spectra.

shown in Figure 1. All the primary amine-acid salts are 1:1, but the dimethylamine salt appears to be 4:1, though it varies between 3:1, 2:1 and 1:1 beyond 80°. The reasons for this behavior are unknown. In most cases, the stoichiometry is constant to 140°.

The mass spectral data also indicated that amide production is detectable around 80 to 90° (Figure 2). The sharp N-H stretch of the amide is detectable by ir only at 120°, and the nmr first gives evidence of the amide at 140°. In addition, amide formation was not detected until above 120° by differential scanning calorimetry. This infers that mass spectrometry is the most sensitive technique by which amide production can be monitored. Furthermore, the reactant salt and the product amide can be clearly differentiated.

In view of the well known sensitivity of biological specimens to very small concentrations of impurities, amide production may be of some importance for those herbicidal formulations containing amine salts which have been subjected to temperatures above 80° during manufacture. The biological effects of 2,4-D amides have been studied (Krewson *et al.*, 1959; Marth *et al.*, 1959; Preston *et al.*, 1958), and it has been concluded that the "presence of an amide group on the side chain of chlorinated phenoxy plant regulators is perhaps almost as important as the position and degree of chlorination on the benzene nucleus" (Marth *et al.*, 1959).

The differing physical properties of the same salts prepared by different workers may also be explained by the fact that very different temperatures were often used in the syntheses.

Figure 2 compares the rate of formation of amide from the different salts. The ratio of the parent ion peak of the amide and the 220 peak of 2,4-D at 7 eV at a constant source temperature of 70° is taken to be the index of the rate of formation of the corresponding amides. The p-35 peak of the amide is also present, but at low eV this means that it has arisen from the parent peak of the amide as the molecular weight is even and not odd as expected of a compound with one nitrogen. Additionally, a metastable for Cl loss can also be seen in the low eV spectrum. That the monochloro compound is not present before fragmentation is also shown by the integrations in the aromatic region in the nmr spectra of the amides. Further, when the amide and p-35 peak heights are compared, they both increase at the same rate. As the C-Cl bond has a bond strength of approximately 80 kcal mol⁻¹ at 250° (Vedenev *et al.*, 1966), it is highly unlikely that

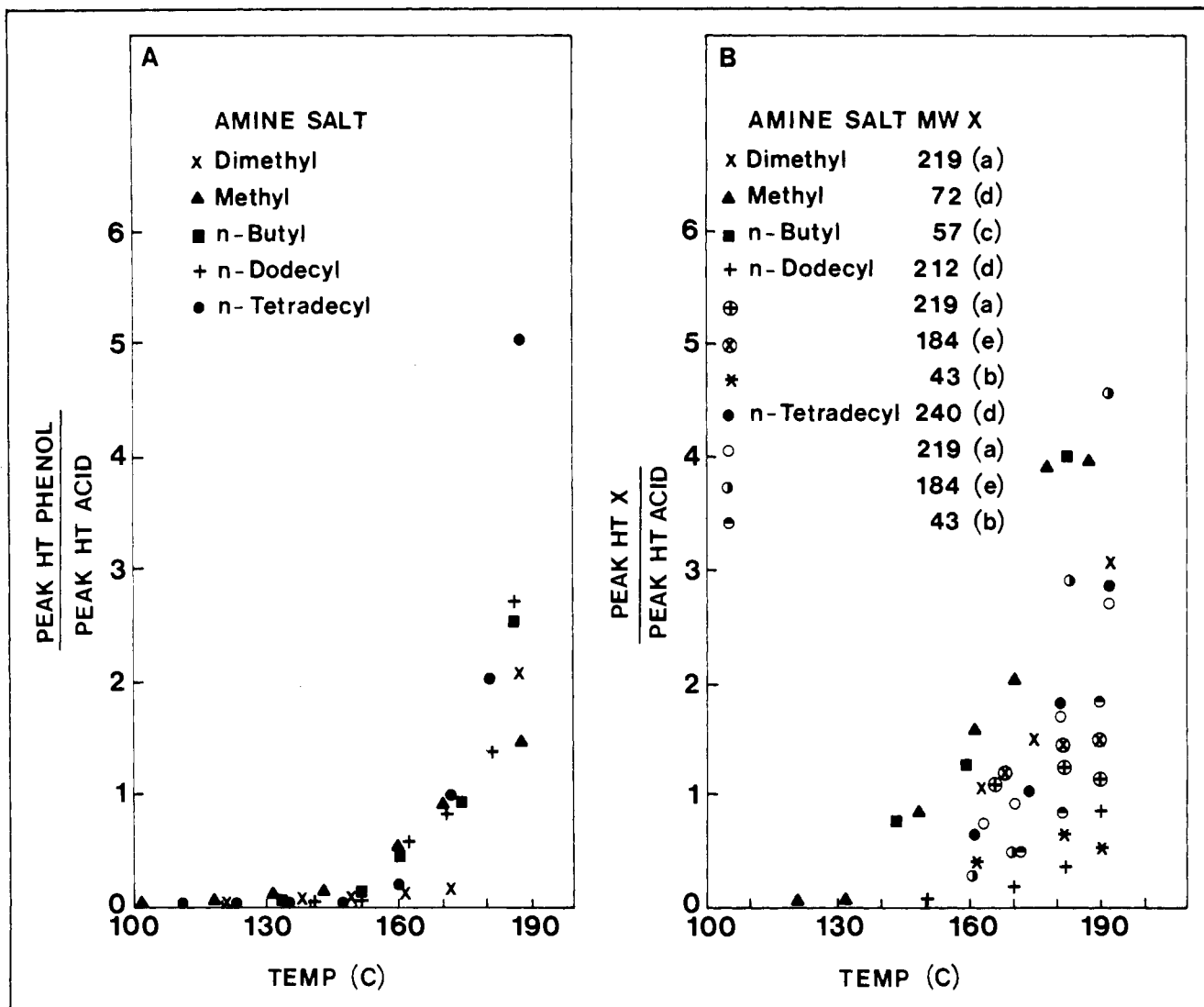


Figure 3. The extent of side product formation from some amine salts of 2,4-D at various temperatures after 1 hr using the peak heights from low eV mass spectra. A gives the extent of production of 2,4-dichlorophenol. B gives the extent of production of 2,4-dichlorophenoxyacetamide (a), methylimine (b), ethylimine (c), lactone-type compound (d), and unidentified compound (e).

Table I. Heats of Fusion and Melting Points of Several Amine Salts of 2,4-D Using Pure Benzoic Acid as External Standard

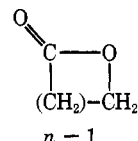
Compound	Observed mp, °C	ΔH_f , kcal/mol	Mw
2,4-D acid	138	5.60 ± 0.30	221
Dimethylamine salt	157	11.20 ± 0.52	266
Methylamine salt	140	6.53 ± 0.23	252
<i>n</i> -Butylamine salt	98	14.97 ± 0.74	294
<i>n</i> -Dodecylamine salt	61	31.15 ± 1.12	406
<i>n</i> -Tetradecylamine salt	57	39.20 ± 1.84	434

one chlorine atom has dissociated from the benzene nucleus under the thermal conditions applied above.

Figure 2 also shows that dehydration for the primary amine salts becomes slower as the chain length of the alkyl moiety increases. The greatest rate of amide production occurs between 140 and 150° for the methyl and *n*-butyl salts, and from 150 to 170° for the *n*-dodecyl and *n*-tetradecyl salts. The dimethylamine salt is much more stable than the primary amine salts, with amide production being appreciable only after 160°.

Amides are not the only pyrolysis products above 160°. 2,4-Dichlorophenol is produced by all the salts above 120° (confirmed by low eV mass spectra [Figure 3A], glc, and a

positive ferric chloride test) but does not become a major product until above 160°. Similarly, imines (ethyl- by *n*-butylamine, and methyl- by the *n*-dodecyl- and *n*-tetradecylamine salts), and 2,4-dichlorophenoxyacetamide (except for the methyl- and *n*-butylamine salts) are also formed (Figure 3B). In addition, another product was tentatively identified as a compound of type



where *n* is the number of carbon atoms in the alkyl moiety. However, the butylamine salt did not produce this compound, and for the methylamine salt *n* appears to be equal to 1. Thus, between 140 and 160° when amide formation is greatest, few side products are formed (Figures 2, 3A, and 3B). For the salts requiring higher temperatures for amide formation, viz the dodecyl-, tetradecyl-, and dimethylamine salts, side product formation becomes as important as amide formation.

Table I, derived from DSC data, shows that the heats of fusion of the amine salts increase as the chain lengths of the alkyl moiety increase (Figure 4). However, the melting

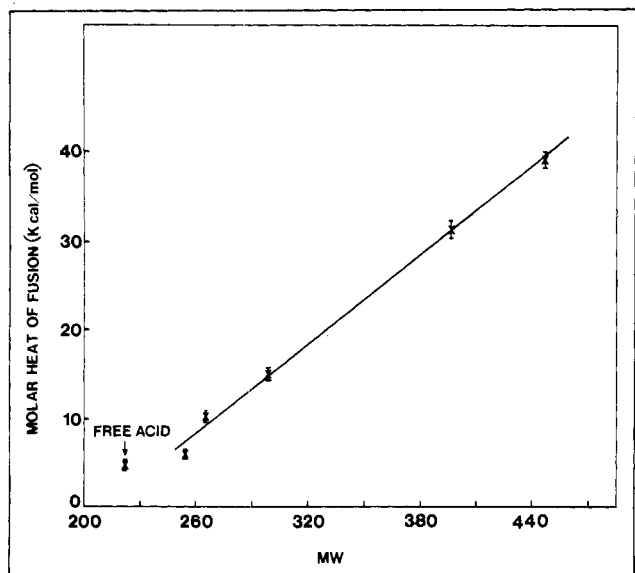


Figure 4. The variation of the molar heat of fusion of various amine salts of 2,4-D with molecular weight (mw), as determined from DSC data using benzoic acid as standard.

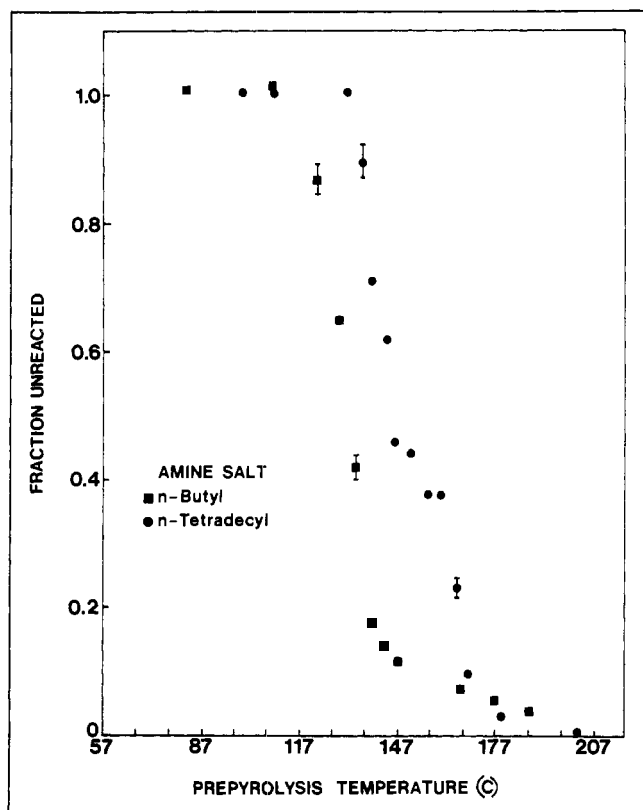


Figure 5. The extent of decomposition of the *n*-butyl- and *n*-tetradecylamine salts of 2,4-D using DSC and the prepyrolysis technique.

points decrease with increasing molecular weight. The most likely reason is that as the chain length becomes shorter, the ionic character of the salt increases.

Figure 5 shows the progress of the pyrolysis of the butylamine and tetradecylamine salts using the differential scanning calorimetric technique described above. These two salts, representative of the long- and short-chain species, have melting points low enough to be amenable to the Boltzmann distribution technique for finding the activation energy (ϵ_i). Figure 6 shows plots of $\log(X_A/X_B)$ vs. $1/T$ for these representative salts using the data from Fig-

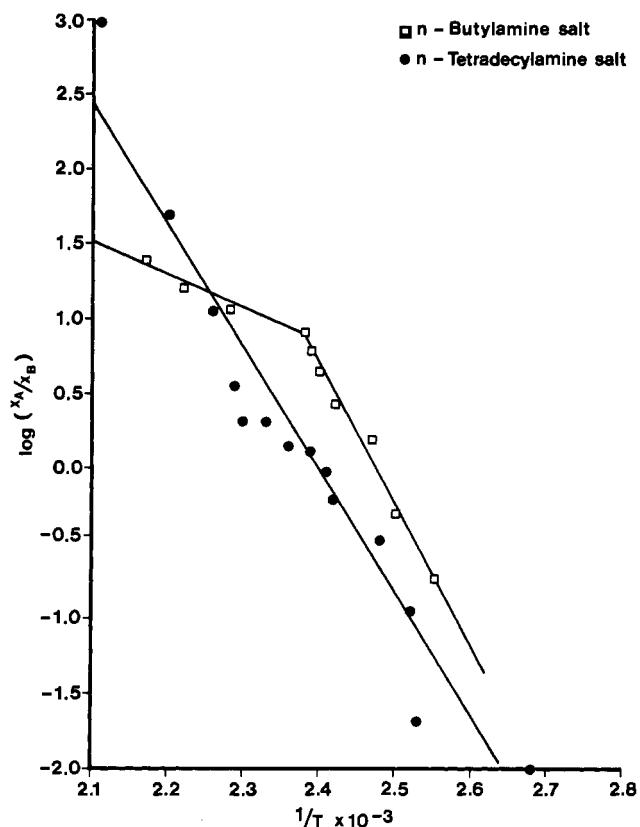
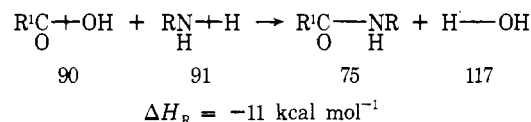


Figure 6. Plot of $\log(X_A/X_B)$ vs. $1/T$ for the calculation of the activation energy for the decomposition of the amine salts of Figure 5.

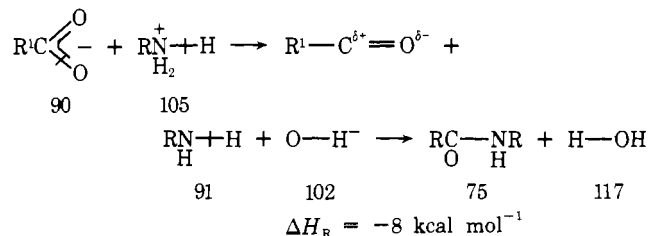
ure 5. ϵ_i was calculated to be (44 ± 1) kcal mol $^{-1}$ and (38 ± 3) kcal mol $^{-1}$ for the *n*-butyl- and the *n*-tetradecylamine salts, respectively. The figure for the *n*-butylamine salt was calculated from the low temperature part of the plot. The significance of the other straight line relation at higher temperatures is not known. No such phenomenon occurred for the *n*-tetradecylamine salt.

Mechanistically, there appear to be three possible schemes.

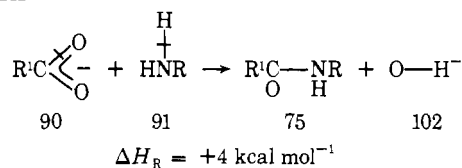
Scheme I



Scheme II



Scheme III



The bond energies (Vedenev *et al.*, 1966) in kcal mol⁻¹ are given beneath the bonds affected, the slash indicating the bond about to be broken. The enthalpy of reaction (ΔH_R) is also given. It is noted that the only endothermic process is that shown by Scheme III. However, when bond energy errors are considered, the energy input could be (4 ± 14) kcal mol⁻¹. The observed activation energies obtained in this paper are larger than the enthalpies of reaction of the three possible schemes. This is the expected result as the enthalpy of reaction is the minimum activation energy.

Morawetz and Otaki (1963) observed that Scheme III best explained the mechanism for the formation of amides in aqueous solution. In addition, they showed that the activation energy decreased with increasing chain length, as observed here for the solid state. However, the activation energies they obtained were approximately half those of the present study. The differences may be caused by the different acids employed for the amidification.

As long as adequate temperature control is used, quantitative yields of amides can be obtained. These results agree with those of Hunter (1941), who prepared long-chain amides of many acids, including phenylacetic acid, by pyrolysis in sealed tubes.

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Estimation of Thiabendazole in the Milligram and Submilligram Ranges

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A procedure is described for the estimation of thiabendazole suspensions of ~ 0.1 -2.0 mg/ml concentration. The procedure depends on the extraction of a copper-thiabendazole-tetramethylguanidine complex by chloroform from an aqueous suspension at pH 12.0-12.5. Absorbance of

the complex is measured at 350-400 nm, or the color may be compared with artificial permanent standards. In a modification that permits estimation of ~ 2 μ g/ml, absorbance of the copper-thiabendazole-tetramethylguanidine complex in chloroform is measured at 311 nm.

Aqueous thiabendazole [2-(4'-thiazolyl)benzimidazole] (TBZ) suspensions of up to 1000 ppm are widely used as a soak for agricultural products to control diseases and spoilage caused by *Fusarium*, *Penicillium*, and other fungi. Methods available to determine the concentration of TBZ in the soak tanks were summarized by Holmes (1972). These included measurement of the absorbance of a dilute hydrochloric acid solution of TBZ at 302 nm, a titration using silver nitrate, and estimation by refractive index. Miller *et al.* (1971) proposed an indirect estimation by determining excess methylmercuric chloride from the reaction of a known amount of methylmercuric chloride with TBZ. The color of the methylmercuric dithizonate formed from the excess methylmercuric chloride and a known amount of dithizone was compared with artificial permanent color standards. Silver nitrate may be substituted for methylmercuric chloride in this reaction (Miller and Csonka, 1973). A test kit for TBZ from Lamotte Chemical Co. is now available using the official AOAC procedure (1965). However, a rapid tankside test is needed in some industries, such as the flower bulb industry in Washington State.

Miller *et al.* (1971) observed that the copper ion reacts with TBZ in alkaline suspension to form a deep green insoluble compound. Under proper conditions, a similar compound may be extracted rapidly by chloroform. The

resulting chloroform solution may be measured spectrophotometrically or compared with artificial permanent color standards to estimate TBZ in the soaking tanks.

MATERIALS AND METHODS

For the estimation of TBZ in bulb treating tanks, up to 20 ml of a carefully mixed sample of the tank suspension containing 2-12 mg of TBZ is placed in a separatory funnel. If the sample volume is <10 ml, water is added to make the volume to 10 ml. Two-tenths of a milliliter of practical grade tetramethylguanidine (TMG) is added, followed by 10.0 ml of chloroform. After the addition of 0.5 ml of 1% cupric chloride dihydrate, the funnel is shaken vigorously at once for 1 min. The layers are allowed to separate and the chloroform phase is passed through a small plug of cotton in the stem of the separatory funnel to remove suspended water. The absorbance of the chloroform phase may be measured spectrophotometrically in a cuvette of 1.00 cm light path at any selected wavelength from 350 to 400 nm and compared with the absorbance of 2.0-12.0 mg of TBZ carried through the procedure and measured at the same wavelength.

Comparison of the chloroform phase may be made with permanent color standards in similar tubes. The composition of the color standards representing 2.0-12.0 mg of TBZ as copper-TBZ-TMG complex in 10 ml of chloroform is given in Table I.

A modified procedure may be used to determine ~ 2 μ g/ml of TBZ. The solution or suspension of TBZ is made alkaline with 0.2 ml of TMG, and 5.00 ml of chloroform is added, followed by 0.1 ml of 1% cupric chloride dihydrate. After extraction and filtration of the chloroform phase

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