(CO)$_8$ can be prepared at low temperatures in an entirely analogous fashion to that used for Rh$_2$(CO)$_8$. Both complexes exist in the bridged form analogous to the bridged isomer of Co$_2$(CO)$_8$. By comparing the temperatures at which Rh$_2$(CO)$_8$ and Ir$_2$(CO)$_8$ disproportionate to the respective bridge forms of Rh$_4$(CO)$_{12}$ and Ir$_4$(CO)$_{12}$, one is led to believe that under comparable conditions Ir$_2$(CO)$_8$ is slightly less stable than Rh$_2$(CO)$_8$ although it is not possible to estimate the magnitude of this difference from the present data. The lower stability of Ir$_2$(CO)$_8$ compared to Rh$_2$(CO)$_8$ and their instability with respect to disproportionation to Ir$_4$(CO)$_{12}$ and Rh$_4$(CO)$_{12}$ above $-58°$ are certainly consistent with the nonobservation of the nonbridged forms of the complexes and with the difficulties previously experienced by various groups of workers in attempting to synthesize and characterize Rh$_2$(CO)$_8$ and Ir$_2$(CO)$_8$, particularly the latter, by room temperature techniques.

Reactions of Molecular Crystals with Gases. I. Reactions of Solid Aromatic Carboxylic Acids and Related Compounds with Ammonia and Amines[1,2]

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Abstract: Single crystals of carboxylic acids such as benzoic and 4-chlorobenzoic acid have been found to react quantitatively with ammonia gas to give microcrystalline aggregates preserving the shape of the original crystal. No accumulation of a 2:1 (acid–ammonia) ammonium salt was detected, and single crystals of this acid–ammonium salt were themselves found to react to form the 1:1 salt. *dl*-2-Butylamine vapor could be employed instead of ammonia in the reaction with *dl*-mandelic acid. The exchange of one base for another was found to be negligible under the reaction conditions as shown by the fact that ammonium *dl*-mandelate failed to react with *dl*-2-butylamine and *dl*-2-butylammonium mandelate was unreactive toward ammonia. Preliminary kinetic studies show no correlation of reaction rate with acidity (measured in aqueous solution), crystal density, or melting point. Microscopic observation of the reaction of single crystals suggests that such reactions may be highly anisotropic.

Although investigations of reactions believed to occur between molecular crystals and gases have been reported[3,4] for many years, there have been few studies of such reactions in depth, and it is sometimes[5] difficult to decide whether or not the reaction is complicated by the presence of melt produced in intermediate stages. The question of the generality of such reactions has not been answered; innumerable instances are known to the synthetic organic chemist in which a substance vulnerable to attack by air (oxygen and moisture) in solution or in the melt is indefinitely stable as the crystalline solid. Particularly striking examples of such nonreactivity are the behavior of the dimer of triphenylmethyl free radical, which in the solid state reacts with oxygen to form an outer coating of ditrityl peroxide that protects the crystal from further oxidation,[6] and the "molecular packaging" of unsaturated fatty acids which are protected from oxidation when incorporated into solid urea-inclusion complexes.[7] Of special interest have been attempts to induce a change in the course of a reaction by carrying it out in a way which should lead to direction of the reaction path by the rigid matrix of starting material. Thus, the orientation in nuclear bromination of aromatic compounds has been studied[8] as have the reactions of crystalline metal salts of phenol with methyl

(1) Taken from the Ph.D. Thesis of Rodger S. Miller, submitted to the University of Illinois, 1973, available from University Microfilms, Ann Arbor, Mich.

(2) We are indebted to the National Science Foundation (NSF-GH-33634) and to the Advanced Research Projects Agency of the Department of Defense (Contract ARPA-HC-15-67-0221) for support of this work.

(3) (a) For preliminary reports of the present work see R. S. Miller, D. Y. Curtin, and I. C. Paul, *J. Amer. Chem. Soc.*, 93, 2784 (1971); (b) *ibid.*, 94, 5117 (1972).

(4) In addition to the literature cited in ref 3a the following may be noted: (a) E. Hadjoudis, E. Kariv, and G. M. J. Schmidt, *J. Chem. Soc., Perkin Trans. 2*, 1056 (1972); (b) E. Hadjoudis and G. M. J. Schmidt, *ibid.*, 1060 (1972); (c) G. Friedman, C. Cohen, D. Wolf, and G. M. J. Schmidt, *Isr. J. Chem.*, 10, 559 (1972); (d) E. Hadjoudis, *ibid.*, 11, 63 (1973); (e) G. Adler, *ibid.*, 10, 563 (1972); (f) C. Ressler, *J. Org. Chem.*, 37, 2933 (1972); (g) J. R. Scheffer and M. D. Ouchi, *Tetrahedron Lett.*, 223 (1970); (h) J. P. Desvergne and J. M. Thomas, *Chem. Phys. Lett.*, 23, 343 (1973); (i) M. D. Cohen and B. S. Green, *Chem. Brit.*, 490 (1973); (j) H. W. Kaufman, D. Rabinovich, and G. M. J. Schmidt, *J. Chem. Soc., Perkin Trans. 2*, 433 (1974); (k) J. P. Desvergne, H. Bouas-Laurent, E. V. Blackburn, and R. Lapouyade, *Tetrahedron Lett.*, 947 (1974); (l) I. C. Paul and D. Y. Curtin, *Science*, in press.

(5) R. E. Buckles, E. A. Hausman, and N. G. Wheeler, *J. Amer. Chem. Soc.*, 72, 2494 (1950).

(6) A. E. Tschitschibabin, *Ber.*, 40, 3056 (1907).

(7) L. C. Fetterly, "Non-stoichiometric Compounds," L. Mandelcorn, Ed., Academic Press, New York, N. Y., 1964, p 556 ff.
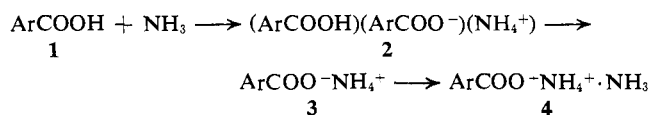
(8) M. M. Labes, H. W. Blakeslee, and J. E. Bloor, *J. Amer. Chem. Soc.*, 87, 4251 (1965); M. M. Labes and H. W. Blakeslee, *J. Org. Chem.*, 32, 1277 (1967).

iodide[9] and carbon dioxide.[10] Thus far, however, the search for highly specific orientational effects brought about by the solid state has not been extensively pursued.

A second attraction of reactions of solids arises from the fact that most organic crystals have all of the molecules in a single conformation. This conformational homogeneity should lead to greater stereospecificity when such a crystal reacts at the surface with its molecules so constituted that asymmetric induction could result. It was suggested by Morawetz[11] that achiral molecules which form chiral crystals, such as benzil, could, if single crystals were separated and caused to react appropriately, produce optically active products by asymmetric induction. An attempt to induce optical activity in this way by carrying out the reaction of diazomethane with a crushed single crystal of methanesulfinic acid (space group $P2_12_12_1$) suspended in pentane failed to produce optical activity in the product[12] and no optical activity could be detected when crushed benzil single crystals were suspended in water and treated with sodium borohydride or in pentane with lithium aluminum hydride at low temperatures.[13] Somewhat more promising have been reports of 6% of induced optical activity in the bromination of a dimethylchalcone[14] and a seeming preference for (S)-(+)-2-butylamine when crystalline (S)-(+)-mandelic acid was allowed to react with dl-butylamine vapor[3a] but all of these studies have been hampered by a general lack of knowledge of the fundamentals of reactions of molecular crystals with gases.

The present investigation was undertaken with the hope of clarifying some relatively simple solid–gas reactions, and the reactions of crystalline substituted benzoic acids were chosen as the point of departure.

The sequence of reactions which must be considered is the following.

$$\text{ArCOOH} + \text{NH}_3 \longrightarrow \underset{\textbf{2}}{(\text{ArCOOH})(\text{ArCOO}^-)(\text{NH}_4{}^+)} \longrightarrow$$

$$\underset{\textbf{3}}{\text{ArCOO}^-\text{NH}_4{}^+} \longrightarrow \underset{\textbf{4}}{\text{ArCOO}^-\text{NH}_4{}^+\cdot\text{NH}_3}$$

The earliest report of the reactions of solid aromatic carboxylic acids with ammonia gas seems to have been that of Pellizzari[15] who pulverized a number of solid acids and phenols and exposed them to dry ammonia vapor. No evidence was presented that the reaction mixture remained solid but it was noted that, whereas benzoic acid with excess ammonia was converted to a 1:1 salt **3**, 4-nitrobenzoic acid reacted with approximately 2 mol of ammonia to give a 1:2 salt, **4**, which reverted to the 1:1 salt, **3**, on standing exposed to the atmosphere. An independent study of Korczynski[16] confirmed this difference in behavior and showed in

addition that 4-chloro- and 4-bromobenzoic acid gave the 1:1 salt **3** but could not be induced to react with further ammonia to form **4**. The first attempt at a quantitative study of the reaction was by Hantszch[17] who studied the reaction of ammonia gas with powdered acid, often ground with powdered glass. Although the intermediate **2** was not considered in this early work, such 2:1 or acid salts are commonly observed in other situations and their structures have been extensively studied.[18]

Experimental Section

Melting points were determined with a Thomas-Hoover capillary apparatus and were corrected. Microanalyses were carried out by Mr. J. Nemeth and his associates. Infrared spectra were of Fluorolube mulls and were determined with a Perkin-Elmer Model 237b or 521 spectrometer calibrated with the 1601-cm^{-1} absorption of polystyrene. Nuclear magnetic resonance spectra were determined with Varian Associates Model A-60, A 60-A, A-56/60A, or T-60 spectrometers using 10% solutions in CCl_4 or $CDCl_3$ with tetramethylsilane as an internal standard. Optical rotations were measured with a Zeiss Circle Polarimeter 0.01° at 589 nm. X-Ray powder photographs were made with a Debye-Scherrer powder camera employing nickel-filtered Cu Kα radiation. Photographs were taken with a Beseler Topcon Super D 35-mm camera mounted on a Bausch and Lomb Model LS dynoptic polarizing microscope, using High Speed Ektachrome Type B.

Reactions of Carboxylic Acids with Ammonia and 2-Butylamine Vapor. A weighed sample of solid acid (previously dried under vacuum for at least 24 hr) was placed in an apparatus[19] which had been flushed with dry nitrogen for 12 hr. The system was evacuated and filled with anhydrous ammonia gas several times and then the absorption of ammonia was followed by the use of a manometer until the pressure was constant. The sample was removed, weighed, and submitted for microanalysis after treatment with a stream of dry nitrogen to remove residual ammonia (Table I).[20–23] Alternatively, single crystals with dimensions of 0.5 mm and larger were placed in a glass compartment equipped with inlet and outlet tubes to permit a continuous slow stream of ammonia gas to flow past the crystal while it was observed on a microscope stage. As reaction progressed, the initially clear crystal became opaque and extinction between crossed polarizing filters was no longer observed. The crystal could be seen to expand. In general, reaction of single crystals gave product "crystals" with greater cohesion than the initial crystals. Single crystals of benzoic acid and 4-chlorobenzoic acid were photographed with a Debye-Scherrer X-ray camera, removed, exposed to ammonia gas for 24 hr, and rephotographed (without grinding). The resulting photographs appeared to be identical with powder photographs of the ground salt and showed no evidence of crystallographic order.[24] Similarly Weissenberg photographs of a single crystal of the acid ammonium salt of 4-chlorobenzoic acid after exposure to ammonia gas for 48 hr showed only the presence of unoriented microcrystallites of the 1:1 salt.

The reactions with dl-2-butylamine were carried out by placing crystals (which had been dried at least 24 hr under vacuum) in a small open jar in a larger sealed jar containing a few drops (excess) of dl-2-butylamine which had been dried over 4 Å molecular sieves for at least 2 weeks prior to use. After reaction was completed the contents of the small jar were flushed with nitrogen, evacuated for 24 hr, and submitted for microanalysis without further purification. Results are summarized in Table I.

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(13) P. E. Bender, unpublished work (see footnote 7 in ref 9).

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(18) See J. C. Speakman, Struct. Bonding (Berlin), 12, 141 (1972), and references cited therein.

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(24) Copies of the X-ray photographs are available in the Ph.D. Thesis of R. S. M.[1]

Table I. Reaction of Crystalline Acids with Ammonia and dl-2-Butylamine Vapor

Acid	Mp of acid (°C)	Salt	Reaction time (days)	Anal. Found (calcd)			Rate of reaction with ammonia, 23 ± 1° (742 ± 8 mm)			
				C	H	N	Single crystals $10^4(k_{1/2} \pm$ stand. dev.) (sec⁻¹)		Powder $10^4(k_{1/2} \pm$ stand. dev.) (sec⁻¹)	
Benzoic[a]	121–122	$C_7H_6O_2 \cdot NH_3$[b]	13	60.5(60.4)	6.4(6.5)	10.0(10.1)	1.5	0.5[c]	2.1	0.5[d]
4-Chlorobenzoic[e]	241–242	$C_7H_5ClO_2 \cdot NH_3$[b]	3	48.2(48.4)	4.6(4.7)	8.2(8.1)	2.9	0.5[f]	20.	4.[d]
4-Bromobenzoic[g]	251–253	$C_7H_5BrO_2 \cdot NH_3$[b]					1.1	0.1[d]		
4-Nitrobenzoic[h]	240–241	$C_7H_5NO_4 \cdot NH_3$[i]	4	45.6(45.7)	4.4(4.4)	15.2(15.2)	0.31 (0.40)	0.03 (0.1)[d,i]	(0.5)	(0.3)[d]
2-Naphthoic[j]	183–184	$C_{11}H_8O_2 \cdot NH_3$	4	70.2(69.8)	5.9(5.9)	7.4(7.4)	0.31	0.08[d]	0.8	0.5[d]
α-trans-Cinnamic[k]	131–132	$C_9H_8O_2 \cdot NH_3$	3	65.4(65.4)	6.7(6.7)	8.5(8.5)	1.3	0.3[d]	2.4	0.2[d]
dl-Mandelic[l]	117–118	$C_8H_8O_3 \cdot NH_3$	3	56.9(56.8)	6.5(6.6)	8.3(8.3)				
		$C_8H_8O_3 \cdot C_4H_{11}N$[m]	1	64.3(64.0)	8.5(8.5)	6.1(6.2)				
		$C_8H_8O_3 . C_4H_{11}N$[n]	4	63.9(64.0)	8.4(8.5)	6.3(6.2)				
Dibenzylacetic[20,21,o]	83–86	$C_{16}H_{16}O_2 \cdot C_4H_{11}N$	10	76.4(76.6)	8.8(8.7)	4.9(4.5)				
2-Methoxy-1-naphthoic[22,p]	174–176	$C_{16}H_{16}O_2 \cdot 2C_4H_{11}N$[q]								
Phthalic[r]	191–192	$C_8H_6O_4 \cdot 2NH_3$					0.8	0.1[f]	1.62	0.08[d]
Ammonium hydrogen 4-chlorobenzoate[23,s]	235–237	$C_7H_5ClO_2 \cdot NH_3$		48.5(48.4)	4.6(4.7)	8.1(8.1)[t]	5.4	1.3[d]	6.2	2.9[d]

a The commercial acid was sublimed and crystallized by slow evaporation of a solution in diethyl ether–ethanol (9:1), small tablets. b As previously reported, only a 1:1 salt was obtained.[15,16] c Average of three runs: 1.75 ± 1.2, 1.06 ± 0.2, 1.75 ± 0.11. d Average of two runs. e The commercial product was recrystallized from ethyl acetate (charcoal treatment) and then by slow evaporation from carbon tetrachloride containing a few drops of ethyl acetate, good quality plates. f Average of three runs. g Commercial acid was recrystallized from ethanol, then acetone, and finally by slow evaporation from an ethyl acetate solution after decolorization with activated charcoal, well-formed small plates. h After reprecipitation from aqueous base and recrystallization from ethyl acetate the commercial acid was recrystallized by slow cooling of an ethyl acetate solution, well-formed plates. i Previous reports[15,16,17a] suggest that the (NH₃)–acid salt first formed was converted to the 1:1 salt by the work-up procedure, as had been observed by Hantzsch.[17a] The first rate constant is that of the first half of the reaction and that in parentheses is for the overall reaction. j The commercial acid after crystallization from acetone (charcoal) and sublimation was crystallized by cooling a hot solution in ethyl acetate, tablets and prisms (not studied). k The commercial acid, after recrystallization from ethanol and sublimation, was crystallized from diethyl ether by slow evaporation, well-formed tablets. l The commercial acid was recrystallized from ethanol–diethyl ether (10:1) (charcoal) and then from diethyl ether by slow evaporation, well-formed rhombohedra. m Powder. n Large crystals. o Crystallized by slow evaporation of a diethyl ether solution, large crystals. p Crystallized by slow evaporation from diethyl ether solution; large crystals. q A 0.0095-g sample of acid absorbed 0.0016 g of NH₃ (theoretical for 1:2 salt, 4, 0.0016 g). r Commercial acid was recrystallized twice from 95% ethanol and then from hot water by slow cooling, colorless plates. s Prepared from the acid in hot ethanol containing 28% aqueous ammonium hydroxide by slow cooling, large plates. t Cl, 20.3; calcd, 20.5.

Since the 2:1 or acid salt 2 might reasonably be supposed to be an intermediate, reaction of a sample of crystalline 4-chlorobenzoic acid with ammonia gas was interrupted prior to completion and the resulting solid (Fluorlube mull) was submitted to analysis by infrared spectroscopy. The relatively intense peak at 1435 cm⁻¹ characteristic of the 2:1 salt was not evident.

Search for Amine–Ammonia Exchange. The irreversibility of the ammonia reaction with crystalline acids was shown by the fact that samples of the ammonium salt were stable under vacuum (1 mm pressure) at temperatures up to 56°. The further question of whether one amine could replace another already present in a crystalline salt was investigated. A sample of ammonium dl-mandelate which had been prepared by reaction of the acid with ammonia for 5 days (see Table I) was treated with dl-2-butylamine vapor for 24 hr (a length of time sufficient for the complete formation of the butylammonium salt from crystalline mandelic acid). The solid was dried under vacuum for 24 hr and the pmr spectrum was determined in D_2O: 5.15 (s, 1 H), 6.40 (s, 6 H). No alkyl proton signal attributable to a butyl group could be detected.

In a second experiment dl-2-butylammonium mandelate prepared from the crystalline acid with dl-butylamine vapor (analysis in Table I) was treated with ammonia gas for 24 hr and the product submitted for microanalysis.

Anal. Calcd for $C_8H_8O_3 \cdot C_4H_{11}N$: C, 64.0; H, 8.5; N, 6.2. Found: C, 64.2; H, 8.3; N, 6.2.

Reaction of (S)-(+)-Mandelic Acid with dl-2-Butylamine. Crystalline (S)-(+)-mandelic acid[25] with $[\alpha]_D$ +150° (lit.[26] +178°), mp 131–134°, was allowed to react incompletely with dl-2-butylamine vapor as described above for the dl-acid. After reaction the mandelate salt was decomposed with benzoyl chloride and 20% aqueous sodium hydroxide to give the benzamide of 2-butylamine. Unreacted liquid 2-butylamine was similarly converted to the amide,

and optical rotations were measured. A reaction for 4 hr at 25° with an initial amine/acid ratio of 0.80 gave benz-2-butylamide from the mandelate salt with $[\alpha]_D^{EtOH}$ +0.86 ± 0.17°. The benzamide from the amine (which had never had direct contact with the reacting crystalline phase) had $[\alpha]_D$ −1.38°. A second similar experiment for 19 hr at 25° with an initial amine/acid ratio of 0.89 gave values of +0.37 ± 0.09° and −0.36. If the fully resolved (R)-(−)-benz-2-butylamide is assumed to have $[\alpha]_D^{EtOH}$ −31°,[27] then the amide from these reactions had 1-4% of the maximum possible optical activity. The (S)-(+)-acid shows a slight preference for (R)-(−)-amine. The reaction was not studied further, however, because there appeared to be softening of the crystals during reaction which introduced the possibility of a liquid phase, a complication to be avoided.

Rates of Reactions of Crystalline Carboxylic Acids with Ammonia Gas. Rates of reaction of a number of the crystalline acids in Table I were followed at constant pressure with the apparatus[28] equipped with a servo-controlled syringe which feeds ammonia gas into the system to replace that which has reacted. The motor driving the syringe follows a signal from a pressure transducer monitoring the pressure in the system and an output signal proportional to the volume of ammonia added to the system to replace the ammonia used in the reaction is plotted against time with a strip-chart recorder. The single crystals studied were selected so that their long axis was 1.5-2 mm. and they were dried under vacuum for at least 24 hr prior to use. Powders were prepared by hand-grinding with a standard number of turns under conditions reproduced as exactly as possible from one run to another. During the kinetic studies the

(25) We are indebted to Professor John Wood for this sample which had been prepared by enzymatic destruction of the (R)-(−)-isomer.
(26) C. Jarowski and W. H. Hartung, J. Org. Chem., 8, 566 (1943).

(27) N. J. Leonard and E. W. Nommensen, J. Amer. Chem. Soc., 71, 2808 (1949); L. Verbit and P. J. Heffron, J. Org. Chem., 32, 3199 (1967).
(28) We are indebted to Mr. Ronald Anderson and his associates in the Electronics Shop of the School of Chemical Sciences for the design and construction of this apparatus. It is described more fully in the Ph.D. Thesis of R. S. M.[1]
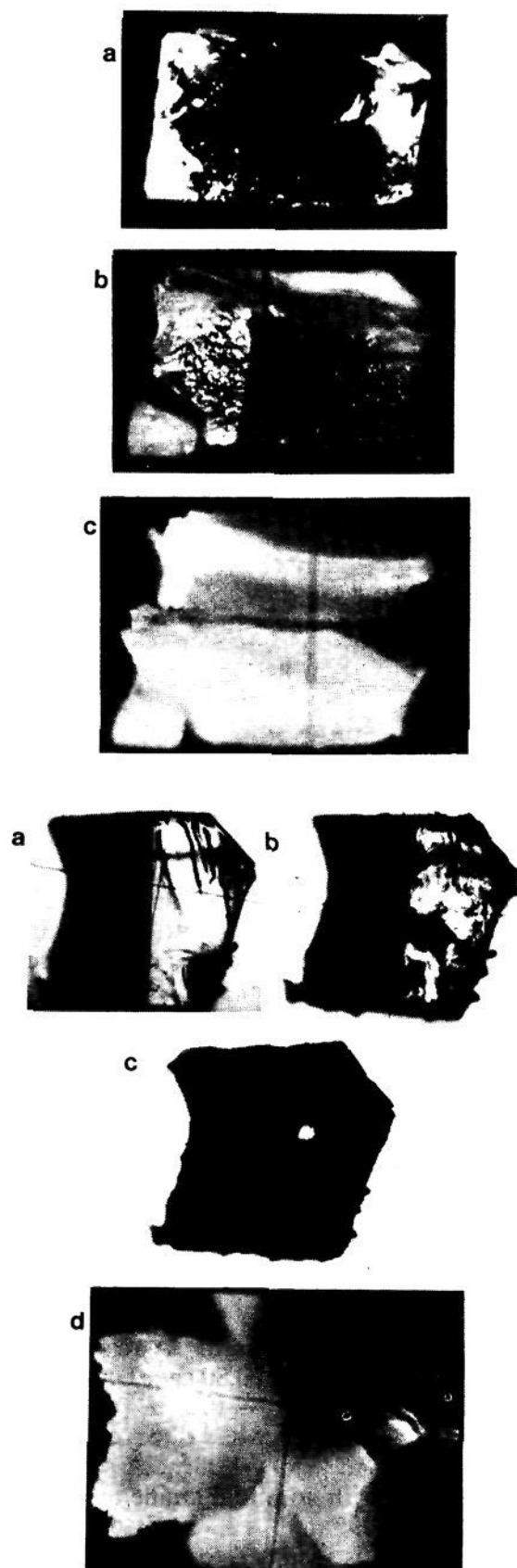
Figure 1. Upper sequence, reaction of dibenzylacetic acid with *dl*-2-butylamine: (a) before reaction; (b) after 3 min (note the preferential reaction at the inclined faces); (c) after 64 min showing the rupture of the crystal due to the great expansion required to accommodate the amine. Lower sequence, reaction of 2-methoxy-1-naphthoic acid: (a) starting crystal; (b) after 6 min; (c) after 12 min; (d) after 110 min.

room temperature was $23 \pm 1°$ and the pressure was 735 mm. If it is assumed that the rate is given by the expression

$$-\frac{dw}{dt} = kw^{2/3}$$

(as would be the case for a single particle if the instantaneous rate is proportional to the surface area of the reactant remaining as reaction begins at the surface and moves inward) then the integrated rate expression is

$$(M)^{1/3} = 1 - k_{2/3}t$$

were (M) is the mole fraction of reactant remaining and t is the time in seconds.[29] As will be discussed in the following papers,[30] how-

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(29) The kinetic situation in this case is like that discussed previously for the thermal decomposition of sulfonium salts. (a) W. D. Burrows, *J. Org. Chem.*, **33**, 3507 (1968); (b) W. D. Burrows and J. H. Cornell, *ibid.*, **32**, 3840 (1967).

ever, microscopic examination of single crystals suggested that reaction occurred primarily at the sides of the crystal, while the top and bottom faces reacted much more slowly. If this observation is incorporated into the model, the rate equation becomes

$$-\frac{dw}{dt} = kw^{1/2}$$

and the integrated expression

$$(M)^{1/2} = 1 - k_{1/2}t$$

Finally, it was pointed out by Burrows[29] that if a mixture of particles of varying size is considered the kinetic order depends on the distribution law assumed and with certain (reasonable) distribution laws a first-order equation is expected.

$$(M) = e^{-k_1 t}$$

In the present work, plots of $(M)^{1/3}$ or of $(M)^{1/2}$ *vs.* t gave reasonably straight lines, the plots of $(M)^{1/2}$ showing somewhat better standard deviations. Both of these plots showed less curvature than plots of $\ln(M)$ *vs.* t. Values of the rate constants $k_{1/2}$ are included in Table I.

Rates of reaction of powders of benzoic acid were not significantly affected by stirring with a magnetic stirrer at 40 rpm nor by drying at 56° under vacuum as compared with a sample allowed to stand exposed to the atmosphere for 8 hr. Although the cell containing the sample was not evacuated to remove air before admission of ammonia gas, the procedure employed was carefully reproduced from one run to the next. Reactions carried out with an interval of 6 months between them gave results in reasonable agreement.

Discussion

The results show that a number of aromatic carboxylic acids, not only as finely ground powders as had earlier been investigated[15-17] but also as single crystals with dimensions of a millimeter or more, are converted quantitatively to ammonium salts by reaction with gaseous ammonia. In many cases when single crystals are employed they can be observed under a microscope to undergo highly anisotropic reaction.[3a,30] The product "crystal" is opaque, and crystals of benzoic or 4-chlorobenzoic acid mounted in an X-ray camera after reaction *but without grinding* gave good quality powder photographs of the 1:1 ammonium salt. This suggests that the product is an aggregate of microcrystallites with external dimensions of approximately 10^3 Å.[31] No evidence for topotactic behavior[32] has been observed. Examination of partially reacted crystals of 4-chlorobenzoic acid suggests that there is no significant accumulation of 2:1 (acid:ammonia) salt, **2**, as a reaction intermediate and, in fact, crystals of the salt **2** also react readily with ammonia gas to give the 1:1 salt **3**. The increase in molecular volume in going from the acid to the 2:1 salt **2** is approximately 6–8% with an overall increase of 15–20% in the change from acid to 1:1 salt.[3a,17] However, formation of diammonium phthalate has been estimated to lead to a 44% increase, and even larger changes are involved in the reactions in which 2-butylamine serves as the gas. In Figure 1 are shown the results of two such reactions. The reaction of ammonia to form 1:1 salts has been shown to be irreversible for the compounds studied and it has further

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(30) R. S. Miller, I. C. Paul, and D. Y. Curtin, *J. Amer. Chem. Soc.*, **96**, 6334 (1974); R. S. Miller, D. Y. Curtin, and I. C. Paul, **96**, 6340 (1974).

(31) A. R. Stokes, "X-Ray Diffraction in Polycrystalline Materials," H. S. Peiser, H. P. Rookesby, and A. J. C. Wilson, Ed., Chapman and Hall, London, 1960, p 409; L. V. Azaroff and M. J. Buerger, "The Powder Method in X-Ray Crystallography," McGraw-Hill, New York, N. Y., 1958, p 254 ff.

(32) J. Z. Gougoutas, *Isr. J. Chem.*, **10**, 395 (1971).

been found that the ammonium cation in ammonium *dl*-mandelate was not replaced by *dl*-2-butylamine nor was 2-butylamine in its mandelate salt replaced appreciably by ammonia under conditions employed for the acid–base reactions.

A mechanism which fits the available facts involves attack of ammonia or amine vapor at the exposed carboxyl groups on the crystal surface to form a monolayer of salt. As the salt forms, disorientation of the layer occurs sufficiently to permit ammonia to diffuse to the next layer of acid molecules where the process is repeated. Certain difficulties in obtaining confirmation of the mechanism from the kinetic order should be noted. Three rate expressions considered in the Experimental Section are shown below in their expanded

$$(M)^{1/3} = 1 - k_{2/3}t \text{ or } (M) =$$
$$1 - 3k_{2/3}t + 3(k_{2/3}t)^2 - (k_{2/3}t)^3$$

$$(M)^{1/2} = 1 - k_{1/2}t \text{ or } (M) = 1 - 2k_{1/2}t + (k_{1/2}t)^2$$

$$(M) = e^{-k_1 t} = 1 - k_1 t/1! + (k_1 t)^2/2! -$$
$$(k_1 t)^3/3! + (k_1 t)^4/4! - \ldots$$

forms where (M) is mole fraction of starting material remaining at time t. Comparison of these three rate expressions suggests that distinction between them requires care, particularly if data are collected for only the first 50% of the reaction or less.[33]

A more serious problem is the difficulty due to the speed of the reaction and the rapid liberation of the heat of reaction[17,34] caused by the use of undiluted ammonia gas.

Gas–solid reactions (other than those studied by Hantzsch[17]) which have been treated as first order include the reaction of oxygen gas with bissalicylalethylenediaminecobalt[35] and its 3,3′-diethoxy derivative as well as the isomerization of crystalline substituted *cis*-cinnamic acids to the trans isomers with iodine vapor.[4b] A rate law which, when put into the form employed in the previous discussion is $(M) = 1 - (kt)^{1/2}$, was assumed for the reaction of nitric oxide or oxygen gas with a solid solution of amide radicals (produced by γ-irradiation) in crystalline amides.[36] It may be noted that in this case there is a very low concentration (*ca.* 10^{-7}

mol/g) of the reactive species so that there is little disruption of the crystal, and a rate law typical for a diffusion process applies.

It can certainly be questioned that our data define unambiguously the rate law for the reactions of ammonia gas with crystalline acids or that the absolute values for the rate constants are reliable. Nevertheless conclusions can be drawn about the significance of the relative rates. These are presented in Table II. They

Table II. Relative Rates of Reaction of Crystalline Acids with Ammonia Gas

Acid	Melting point (°C)[a]	Density[b] (g/cm³)	Relative rates Single crystal[c]	Powder[c]	Relative acid dissociation constant (H₂O solution)[a]
2-Naphthoic	186	1.32	1.0	1.0	1.9
4-Nitrobenzoic	242	1.60	1.0	0.6	10.0
Phthalic	210	1.58	2.6	2.0	35
1-Naphthol	96	1.29		2.0[d]	10⁻⁵
4-Bromobenzoic	255	1.86	3.5		2.9
α-*trans*-Cinnamic	136	1.24	4.2	3.0	1.0
Benzoic acid	122	1.32	4.8	2.6	1.7
4-Chlorobenzoic	243	1.54	9	25.0	2.8
Ammonium acid 4-chlorobenzoate		1.49	17	18	

[a] "Lange's Handbook of Chemistry" 10th ed, McGraw-Hill, New York, N. Y., 1961. [b] J. D. H. Donnay and H. M. Ondik, "Crystal Data," 3rd ed, U. S. Department of Commerce, National Bureau of Standards, and the Joint Committee on Powder Diffraction Standards. [c] Present work. [d] The value for 1-naphthol is calculated from Hantzsch's relative rates for naphthol and benzoic acid.[17a]

are somewhat arbitrarily divided into slow, medium, and fast groups, and it is reassuring that this division is the same whether the rates for single crystals or for powders are compared. Hantzsch's data[17a] for 1-naphthol at 0° are included since this substance has a layered crystal structure[37] similar to the carboxylic acids[3] and shows most dramatically the lack of correlation of rate with acidity over a rather large range of acidities. It is noteworthy that the group of slowest reacting acids contains both the most and the least acidic compounds listed in Table II. A similar lack of correlation of rate with melting point or crystal density (rough measures of relative crystal stability) is apparent.

In a number of other investigations of solid–gas reactions of molecular crystals, large changes in rate have been found to be produced by seemingly trivial changes in molecular structure. For example, crystalline tetramethylrubrene undergoes photochemical oxidation much more readily than the parent rubrene.[38] The reaction with oxygen gas of triphenylmethyl radical produced in a matrix of crystalline triphenylacetic acid is very much faster than is the reaction of triphenylmethyl radical in a matrix of triphenylmethyl bromide or silver triphenylacetate.[39] The related reaction[36] of oxygen gas with radicals imbedded in crystalline amides already mentioned shows similar great varia-

(33) For example recalculation of the data of Hantszch and Wiegner (ref 17b, data in the table on p 481) gives a rate constant and standard deviation for benzoic acid and ammonia from least-squares analysis of $(2.84 \pm 0.02) \times 10^{-3}$ sec⁻¹, whereas when data are plotted as a one-half order reaction there is curvature and a markedly higher standard deviation ($k_{1/2} = (8.72 \pm 0.36) \times 10^{-4}$ sec⁻¹). On the other hand data for the first 60% reaction give a one-half order constant with a much smaller standard deviation ($k_{1/2} = (10.8 \pm 0.03) \times 10^{-4}$). Although the first-order plot is more nearly linear, the authors state that data for other kinetic runs whose first-order plots showed curvature were discarded. It may be noted that of 12 kinetic runs with benzoic acid purified by sublimation and ground with varying amounts of ground glass this is stated to be the only one which gave a "good" (first order) "constant." Eleven points were reported covering the range of 16–94% reaction. The sample was ground with twice its weight of ground glass.

(34) Dr. J. H. Engelmann (unpublished work) has found that the reaction of 4-chlorobenzoic acid with undiluted ammonia is too rapid to follow kinetically even with the recording apparatus described in the Experimental Section.

(35) C. H. Barkelew and M. Calvin, *J. Amer. Chem. Soc.*, **68**, 2257 (1946); see L. H. Voght, Jr., H. M. Faigenbaum, and S. E. Wlberley, *Chem. Rev.*, **63**, 269 (1963); A. E. Martell and M. Calvin, "Chemistry of Metal Chelate Compounds," Prentice-Hall, Englewood Cliffs, N. J., 1952. The 3,3-difluoro derivative of this compound gave data which best fit a second-order plot.

(36) A. Faucitano, A. Perotti, and G. Adler, *Mol. Cryst. Liquid Cryst.*, **9**, 296 (1969).

(37) B. Robinson and A. Hargreaves, *Acta Crystallogr.*, **17**, 944 (1964).

(38) R. M. Hochstrasser and G. B. Porter, *Quart. Rev., Chem. Soc.*, **14**, 146 (1960).

(39) E. G. Janzen, F. J. Johnston, and C. L. Ayers, *J. Amer. Chem. Soc.*, **89**, 1176 (1967).

tion with changes in the amide structure. The Co[II] complexes of bissalicylaldehyde imines exist in a variety of crystalline forms some of which have reacted rapidly and reversibly with oxygen gas while others are inert.[35,40] An example of biochemical interest is the reported[4f] ready dehydrogenation by oxygen of the crystalline hydrate of L-1,4-cyclohexadienyl-1-alanine whereas the unhydrated L, D, and DL forms of the substance are stable under the same conditions. It has been recognized in these cases that there was involved initially a surface reaction and that the subsequent rate depends on the ability

(40) W. P. Schaeffer and R. E. Marsh, *Acta Crystallogr., Sect. B*, **25**, 1575 (1969).

of the gas to gain access to the interior of the crystal. Unfortunately, there has been a paucity of crystallographic structural data in most of these cases. When direct microscopic examination of single crystals has been carried out no anisotropy has been reported. In contrast, the reactions of ammonia with acids almost always show much more rapid reaction at some crystal faces than at others (Figure 1).[3] An understanding of this phenomenon in terms of the crystal structure seems to be a prerequisite to a further understanding of the mechanisms of such reactions. This subject will be developed further in the following papers.[30]

Reactions of Molecular Crystals with Gases. II.
The X-Ray Structure of Crystalline 4-Chlorobenzoic Acid and the Anisotropy of Its Reaction with Ammonia Gas[1,2]

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Abstract: The crystal structure of 4-chlorobenzoic acid (1) has been determined from three-dimensional reflection data collected on a diffractometer. The crystals are triclinic with $a = 14.392 (2)$ Å, $b = 6.227 (1)$ Å, $c = 3.861 (1)$ Å, $\alpha = 88.68 (3)°$, $\beta = 100.12 (4)°$, and $\gamma = 93.31 (3)°$; the space group is $P\bar{1}$. The structure was refined to an R factor of 0.071 on 924 independent reflections. The crystal is comprised of stacks of hydrogen-bonded dimers as are many other substituted benzoic acid crystals. The dimers are so arranged that there is a nearly linear C-Cl $\cdots$ Cl-C interaction (C-Cl $\cdots$ Cl angle of 167°) with a Cl $\cdots$ Cl contact of 3.44 Å. The progress of the reaction of single crystals of 1 with ammonia gas was followed under the microscope. The crystals grow as plates with (100) being the developed face of the plate. A highly anisotropic direction of attack was observed, with the rate of reaction being quite rapid from the sides of the plate while the top face remained clear. This behavior can be rationalized in that the faces where reaction is rapid have carboxyl functional groups completely exposed to attack from the surface, whereas the {100} faces have a layer of closely packed 4-chlorophenyl groups (with the C-Cl vector pointing toward the surface of the {100} faces) shielding the carboxyl groups from attack. Should, nonetheless, a surface layer of molecules react, progress of the reaction toward the center of the crystal is hindered effectively by the large distances between neighboring carboxyl groups in a direction normal to the {100} faces.

Microscopic examination of the reactions of crystalline carboxylic acids with ammonia gas suggested[3] that such reactions are in general anisotropic. In order to examine possible correlation of reactivity with the structure of the crystal faces, it is necessary to know the crystal structures of the pertinent acids. Many of the desired crystal structures had been reported but that of 4-chlorobenzoic acid, one of the fastest reacting and highest melting acids, was not known in any detail. This paper describes the determination of the X-ray structure of 4-chlorobenzoic acid and discusses the correlation of chemical reactivity with crystal structure for this compound.

A morphological study on 4-chlorobenzoic acid (1) was carried out by Steinmetz[4] and is reported by Groth.[5] Single-crystal cell data were presented by Toussaint[6] with subsequent minor changes by Pollock and Woodward.[7] Toussaint also published a picture of a projection of the structure but without further details.[6] In the course of a study of some other aromatic acids, Schmidt[8] commented in general terms upon the probable packing arrangement in the unit cell although no additional experimental data were gathered.

Experimental Section

A plate-like crystal (0.10 × 0.20 × 0.045 mm) was obtained from a solution of the acid in a 2% mixture of ethyl acetate in carbon tetrachloride. Crystal data: for $C_7H_5O_2Cl$, $M = 156.6$; triclinic,

(1) Taken from the Ph.D. Thesis of Rodger S. Miller, submitted to the University of Illinois, 1973, available from University Microfilms, Ann Arbor, Mich.

(2) We are indebted to the National Science Foundation (NSF GH-33634) and to the Advanced Research Projects Agency of the Department of Defense (Contract ARPA-HC-15-67-0221) for support of this work.

(3) R. S. Miller, D. Y. Curtin, and I. C. Paul, *J. Amer. Chem. Soc.*, **96**, 6329 (1974).

(4) H. Steinmetz, *Z. Kristallogr., Kristallgeometrie, Kristallphys. Kristallchem.*, **53**, 473 (1914).

(5) P. Groth, "Chemische Krystallographie," Vol. 4, Engelmann, Leipzig, 1917, pp 466–467.

(6) J. Toussaint, *Acta Crystallogr.*, **4**, 71 (1951).

(7) J. M. Pollock and I. Woodward, *Acta Crystallogr.*, **7**, 605 (1954).

(8) G. M. J. Schmidt, *J. Chem. Soc.*, 2014 (1964).