

1931

# The effect of substitution on the electron sharing ability of the two- and three- carbon radicals

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THE EFFECT OF SUBSTITUTION ON THE ELECTRON  
SHARING ABILITY OF THE TWO- AND THREE-  
CARBON RADICALS

BY

Otis D. Cole

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A Thesis Submitted to the Graduate Faculty  
for the Degree of

DOCTOR OF PHILOSOPHY

Major Subject Plant Chemistry

Approved:

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1931

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## INTRODUCTION

The fundamental method which is used by the organic chemist to make predictions of the properties and reactions of organic compounds is that of reasoning by analogy. If he can place a new substance in a class whose characteristics are known, its properties and reactions will be found to be similar to others of that type; but if he is making a study of a compound which falls in a class whose characteristics are not known he is at a loss in making predictions which will be of value in his work.

In no case is this more noticeable than in the study of negativity and valence, and this paper deals with a classification of radicals on the basis of their ability to share electrons. The question of negativity has occupied the minds of many organic chemists since the time of Ostwald<sup>1</sup>, whose studies have been generalized and enlarged by several later authors<sup>2</sup>. These generalizations may be interpreted to mean that the electron-sharing ability of any atom or group depends upon the strength of the atom or group to which it is linked.

It is generally considered that the union of two atoms,

- A and B, is accomplished in one of three ways: an electron is
1. Ostwald, J. prakt. Chem., 31, 433 (1885); Z. physik. Chem., 3, 170, 418 (1889).
  2. A general summary of theories of negativity and valence is given by Henrich, Johnson, and Hahn in Theories of Organic Chemistry, John Wiley & Sons, New York, p. 75-168, 1922.

transferred from A to B, it is transferred from B to A, or it is shared mutually between the two. Other factors which enter into the relative ease with which two atoms combine are the potential, the structure, and the mass of the atoms. These factors have all been grouped under one general heading called the electron-sharing ability of the atom. All data involving the electron-sharing ability of atoms or groups must be interpreted by considering the above factors.

From an examination of the above factors it follows that the polar properties of any series of compounds, such as the carboxylic acids,  $R-COOH$ , should be a function of the electron-sharing ability of the  $R$ - groups, if the interpretation of the electronic theories of valence is correct. If a mathematical relationship between the polar properties and the electron-sharing ability (the combined effect of the above factors) could be established, the organic chemist would have a better basis upon which to make his predictions.



### HISTORICAL

Hixon and Johns<sup>3</sup>, and Johns<sup>4</sup> pointed out on the above basis that the polar properties of several series of compounds containing only one polar group should be a function of the electron-sharing ability of the radicals or groups attached to the central group or atom. A test of this assumption was made by tabulating the dissociation constants of all the primary amines and carboxylic acids which permitted at least one comparison of a radical. In making this test the dissociation constants for all the amines were plotted against an empirical scale designated as the electron-sharing ability. The scale of electron-sharing ability was established by the use of an arbitrary curve:

$$y = \log K = -20 (e^{0.008x} - 0.75).$$

By first drawing the curve and then locating on it the points of  $\log K$  for the primary amines,  $R-NH_2$ , the positions for the various  $R$ - groups were fixed along the  $x$ -axis. It was found that, if the dissociation constants for any series of polar compounds were plotted against the electron-sharing ability of the  $R$ - groups present in the compounds, a smooth curve could be drawn through all the points. The curves were not found to be

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3. Hixon and Johns, J. Am. Chem. Soc., 49, 1786-95 (1927).

4. Johns, I. B., Unpublished Thesis, Iowa State College Library, (1930).

erratic at any point. The same orderly change in the dissociation constants of the amines ( $R-NH_2OH$ ) which may be said to be caused by the spacing of the radicals along the x-axis was observed in the acids ( $R-COOH$ ). It was predicted that the equilibrium constants for any given reversible reaction (that is, any measure of affinity) in any series of compounds would be a function of the electron-sharing ability of the radicals attached to the central group. Figure 1 shows this relationship for the acids  $R-COOH$ ,  $R-CH_2-COOH$ ,  $R-CH_2-CH_2-COOH$ , and the amines,  $R-NH_2$ . It will be observed that the slopes of the curves for the acids  $R-COOH$ ,  $R-CH_2-COOH$ , and  $R-CH_2-CH_2-COOH$  become progressively less.

These investigators implied that the difference between the slope of the curve for the amines ( $R-NH_2OH$ ) and that of the acids ( $R-COOH$ ) was due to the difference in size or mass of the central group; the smaller central group,  $-NH_2-$ , in the amines allowed more of the effect of the R- group to be transmitted through it than did the larger and heavier group,  $-C(=O)-$ , in the acids.

Johns<sup>5</sup> and others<sup>6</sup> extended the above concept of the electron-sharing ability to the ionization of organomercuric nitrates and chlorides and to the equilibrium reaction:



5. Johns, Peterson, and Hixon, J. Phys. Chem., 34, 2218-25 (1930).  
6. Johns with Hixon, J. Phys. Chem., 34, 2226-35 (1930).

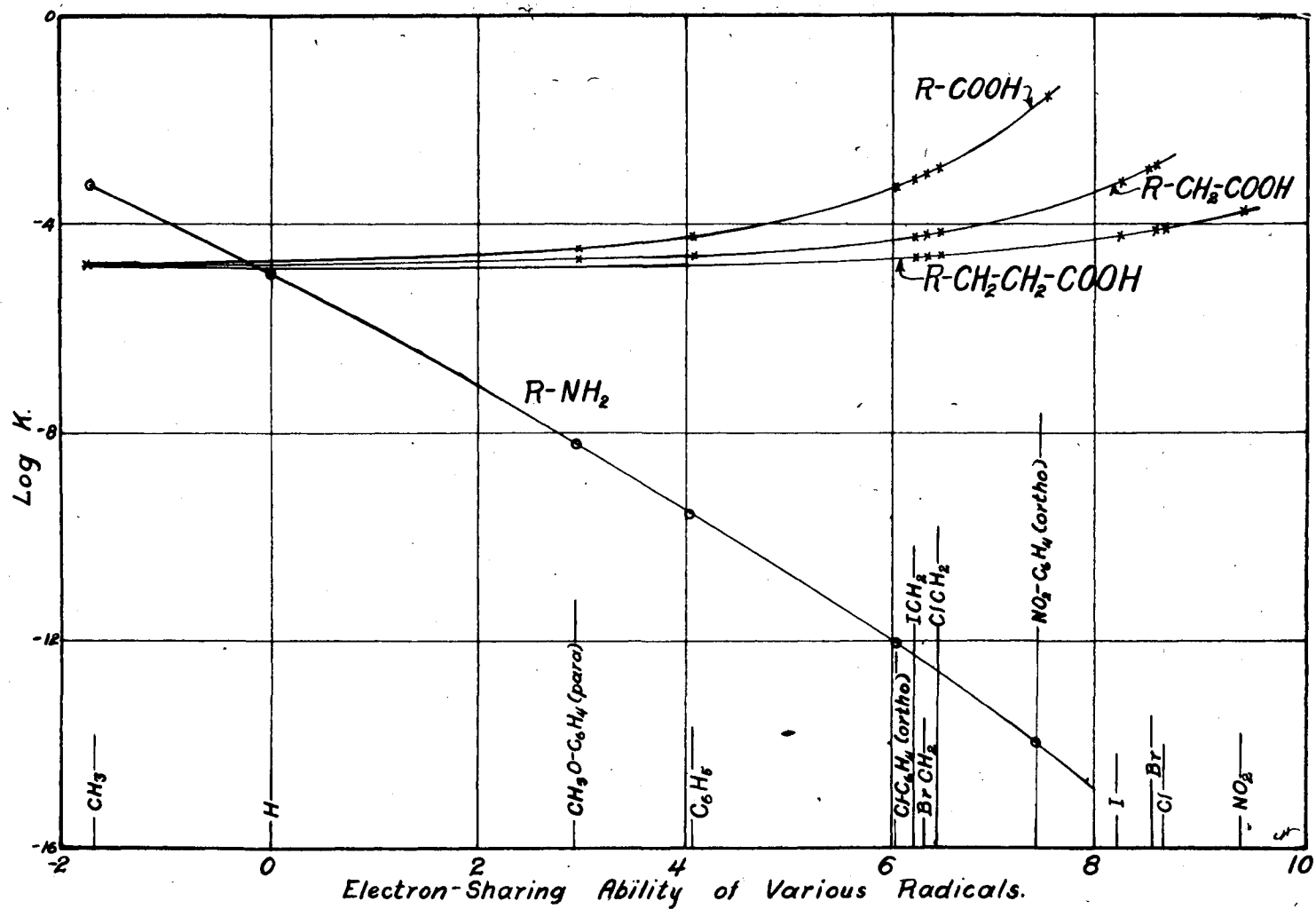


Fig. 1:  
Curves showing the relationship between the dissociation constants of the acids and amines.

Allison and Hixon<sup>7</sup> pointed out that strict interpretation of the data of Hixon and Johns shows that the electron-sharing ability of the carbon radicals varies from values more positive than hydrogen to values more negative than chlorine. They also presented evidence to show that the glucosyl radical (along with other polyalcoholic radicals) fell almost exactly half way between the methyl and phenyl radicals in electron-sharing ability.

Osburn<sup>8</sup> has shown by conductance measurements, and Ware<sup>9</sup> has shown by E. M. F. measurements that the dissociation constants of the arsonic acids are functions of the electron-sharing ability of the radicals attached to the  $-\text{As}(\text{OH})_2$  group. Osburn also pointed out the correlation of the toxicity of certain arseno- compounds with the electron-sharing ability of the radicals which were substituted in the compounds.

Craig<sup>10</sup>, and Hixon and Craig<sup>11</sup>, while investigating the insecticidal action of certain nitrogen heterocyclic compounds, found that the dissociation constants of the N-substituted and  $\alpha$ -substituted pyrrolidines were dependent on the negativity of the substituting R- group. The dissociation constants of the N-substituted pyrrolidines were almost identical

7. Allison with Hixon, J. Am. Chem. Soc., 50, 168-72 (1928).

8. Osburn, O.L. Unpublished Thesis, Iowa State College Library (1930).

9. Ware, F.E. Unpublished Thesis, Iowa State College Library (1930).

10. Craig, L.C. Unpublished Thesis, Iowa State College Library (1931).

11. Hixon and Craig, J. Am. Chem. Soc., (in the hands of the editor).

with those of the corresponding primary amines, while those of the  $\alpha$ -substituted pyrrolidines were very close to those of the corresponding R-CH<sub>2</sub>-NH<sub>2</sub> compounds. The pyridyl and pyrrol radicals were shown to be more negative than the phenyl radical.

STATEMENT OF THE PROBLEM

In the table published by Hixon and Johns<sup>12</sup> no primary amines having dissociation constants between  $10^{-8}$  and  $10^{-9}$  are found. A search of later literature shows only three primary amines with dissociation constants in this intermediate range: glucosyl amine<sup>13</sup>,  $K_b=5.5 \times 10^{-7}$ , orthochlorobenzyl amine<sup>14</sup>,  $K_b=6.3 \times 10^{-8}$ , and indenyl amine<sup>15</sup>,  $K_b=5.75 \times 10^{-7}$ . As will be shown later, the last of these is probably in error\*. Unpublished data by Moore<sup>16</sup> shows that the dissociation constant of camphor amine also falls in this intermediate range ( $K_b = 1.1 \times 10^{-7}$ ). The lack of data on amines in this range would seem to indicate that these amines are unstable and either decompose quickly after forming or do not form readily.

From a consideration of structure and reactions, some of the unsubstituted aliphatic radicals should have electron-sharing abilities as great as that of the phenyl radical or greater. For example, qualitatively it is indicated that acetylene ( $\text{HC}\equiv\text{CH}$ ) and allylene ( $\text{CH}_2\text{C}\equiv\text{CH}$ ) on account of their ability to form salts with various metals, should contain very

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12. Hixon and Johns, Op. cit., 1788.

13. Vellinger, Compt. rend., 182, 1625 (1926); Allison with Hixon, Op. cit., 171.

14. Ware, F. E., Op. cit. p. 21.

15. Bourgeaud and Dondelinger, Bul. soc. chim., (4) 37, 277-87 (1925).

\* See Appendix.

16. Moore, P.A., Iowa State College. Private Communication. (1931).

negative groupings. If such is the case, the corresponding radicals have high electron-sharing abilities, although they are essentially aliphatic in structure. Figure 1 shows that the halogen substituted methyl radicals have electron-sharing abilities far in excess of the electron-sharing ability of the phenyl radical. From this it may be assumed that, by the introduction of various negative substituents into simple aliphatic radicals, the electron-sharing abilities of the radicals so formed will vary in value throughout the whole range-- from the completely saturated methyl radical on the one hand to the highly unsaturated radicals containing the triply-linked carbons on the other.

The purpose of this investigation is three-fold:

1. To determine the values of the electron-sharing abilities of some of the substituted two- and three- carbon aliphatic radicals.
2. To study the relative effects of different substituents.
3. To attempt to find some radicals which fall in the intermediate range between the aliphatic and aromatic radicals.

Simpler radicals would permit the study of properties of compounds in this range without involving the complexity of specific reactions characteristic of the biological compounds (carbohydrates and terpenes) falling within these limits.

## EXPERIMENTAL

### Plan of Procedure

The primary amines were chosen as compounds for study because the slope of the curve given by Hixon and Johns<sup>17</sup> is great enough to allow for more accurate differentiation of the values of the electron-sharing ability. The arsonic acids were desired since they would be of interest chemically.

In order to make predictions concerning radicals which might fall in the intermediate range, and in order to give a check upon the results obtained in this study, the graph shown in Figure 2 was constructed. This consists of two curves:  $\log K_b$  for the primary amines,  $R-NH_2$ , plotted against the electron-sharing ability of the radicals,  $R$ , and  $\log K_b$  for the primary amines  $R-CH_2-NH_2$  plotted against the same abscissa. The method of using the graph is as follows: The value of the abscissa for a certain radical,  $R$ , is located on the  $R-NH_2$  curve. This value is then used to locate a point on the  $R-CH_2-NH_2$  curve and the value of  $\log K_b$  for  $R-CH_2-NH_2$  is determined from this point. This procedure may be reversed and the value of  $\log K_b$  for  $R-NH_2$  may be determined from the value of  $\log K_b$  for  $R-CH_2-NH_2$ . The chief obstacle to the use of this graphical method is the fact that no points on the abscissa

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17. Hixon and Johns, Op. cit., 1789.



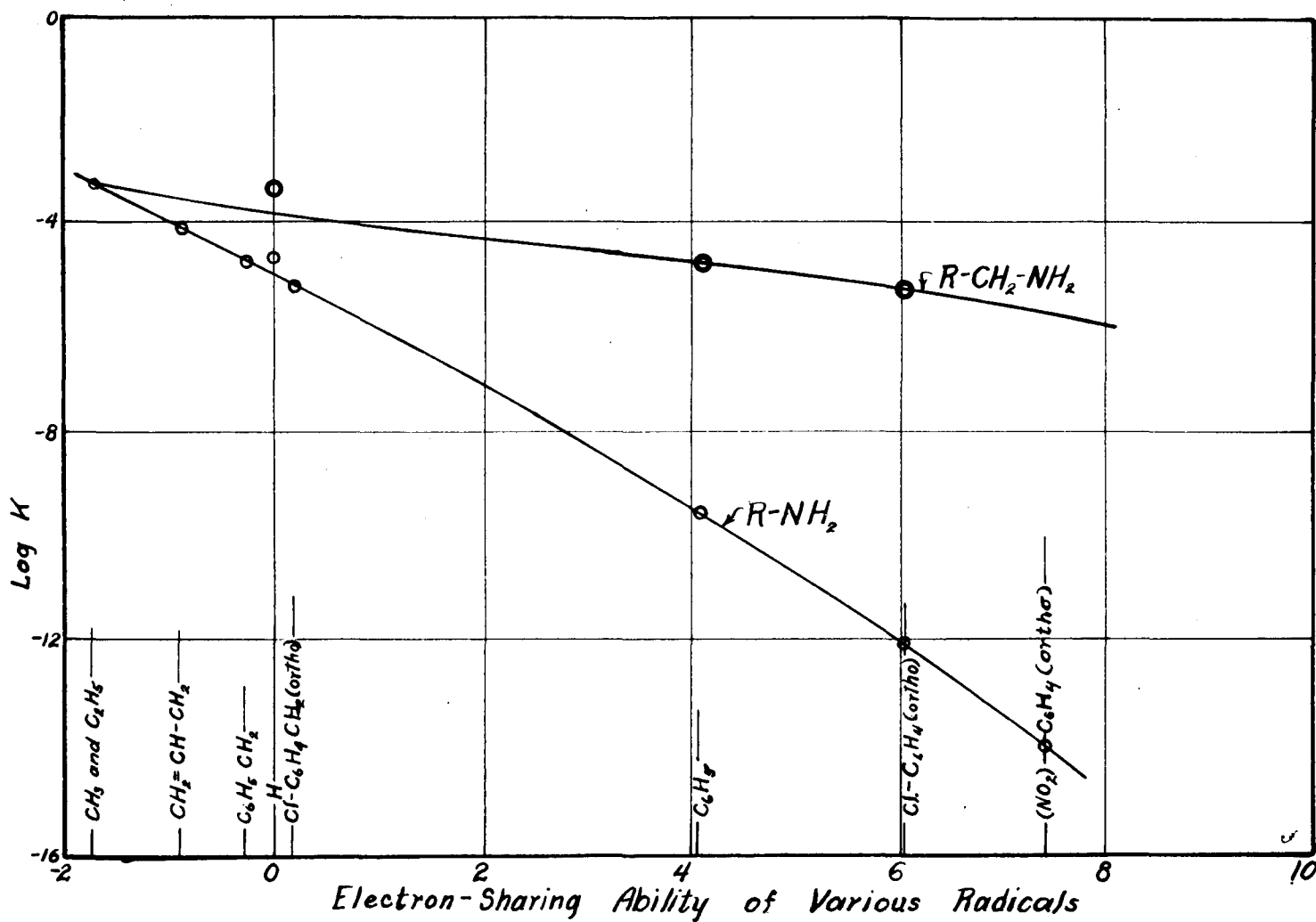


Fig. 2:  
Curves showing the effect of one CH<sub>2</sub>-group on the dissociation constants of the amines.

beyond that located by orthochlorobenzyl amine have been determined for the  $R-CH_2-NH_2$  curve.

The experimental work of this study consists of the following:

1. Preparation of the compounds for measurement.
  - a. Preparation of primary amines.
  - b. Preparation of arsonic acids.
2. Measurement of the dissociation constants and the determination of the electron-sharing ability of the radicals.

### Preparation of Compounds

#### Preparation of Primary Amines

The varying chemical nature of the primary amines used in this study made it impossible to use one general method for the preparation of all of them. For this reason each preparation will be discussed separately.

2,3-dibromopropyl amine hydrochloride. This compound was prepared from allyl amine according to Henry<sup>18</sup>. Fifteen and one-half grams of allyl amine were dissolved in slightly over one equivalent of dilute (1-1) hydrochloric acid. The reaction was very vigorous and the mixture was cooled in an ice-salt bath. When the temperature of the mixture had been 18. Henry, Ber., 8, 399 (1875).

lowered to about 0° - 5° C., bromine was added to the solution until a faint yellow color persisted. During this addition of bromine white crystals separated. When all the bromine had been added the mixture was cooled again to 0° C., and the crystals were filtered off by suction. The yield was 34.8 g. or 51%. The mother liquors were not concentrated, but were used to obtain the free base, which decomposed upon standing at room temperature for about two hours. It was impossible to dry the free base without decomposition. The hydrochloride salt was recrystallized once from water and twice from alcohol. Its melting point was 155.5° - 156.5°.

Analysis: Calculated for  $C_3H_5NBr_2Cl$ : N, 5.53%.

Found: N, 5.39%; 5.57%.

Propargyl amine hydrochloride was prepared according to the directions given by Paal and Herman for the preparation of the oxalate salt<sup>19</sup>. Three and seven-tenths grams of sodium were dissolved in about 75 cc. of absolute alcohol in a 100 cc. pressure bottle, and 10 g. of 2,3-dibromopropyl amine hydrochloride were added. The bottle was sealed and heated for 1 1/4 hours in a boiling water bath. After heating, the bottle and its contents were cooled to room temperature and the mixture was placed in a distilling flask and distilled. The distillate was made acid with dilute hydrochloric acid and evaporated in 19. Paal and Herman, Ber., 22, 3076-85 (1889).

vacuo. The residue was taken up in absolute alcohol and the propargyl amine hydrochloride was precipitated with dry ether in the form of a brownish-red powder. The yield of the crude material was 2 g. The hydrochloride decomposed before a recrystallization could be made. The melting point was not determined. The instability of this compound prevented its inclusion in the series of dissociation constant measurements.

2,3-dibromobutyl amine hydrochloride was prepared from crotonyl amine in a manner analogous to the preparation of 2,3-dibromopropyl amine hydrochloride from allyl amine. The crotonyl amine was prepared from crotonaldehyde by the formation of the oxime and subsequent reduction<sup>20</sup>.

For the preparation of the crotonaldoxime, 87.5 g. (1.25 moles) of hydroxylamine hydrochloride were dissolved in 300 cc. of water in a 1-liter three-necked flask fitted with a stirrer. The solution was stirred and cooled to 0° - 5°C. in an ice-salt bath and 70 g. (0.66 mole) of anhydrous sodium carbonate were added, a little at a time to prevent loss by frothing. After all the sodium carbonate had been added, the stirring was continued while 70 g. (1 mole) of Eastman's crotonaldehyde (B.P. 102-3°) were added drop by drop. After a few minutes the mixture became cloudy and soon an oily precipitate of the oxime could be seen in the mixture. The tempera-

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20. Goldschmidt, Ber., 20, 728 (1887); Schindler, Monatsh. 12, 415-6 (1891).

ture of the reaction was not allowed to go above  $10^{\circ}\text{C}$ . at any time. After all the aldehyde had been added, the mixture was stirred for one-half hour, and then was extracted three times with 100 cc. portions of ether. The ether extract was dried overnight with anhydrous sodium sulfate and concentrated in a wide dish placed within a large vacuum desiccator. When all the ether had evaporated a mass of long needles remained. The yield was 26 g. or 35%. The melting point was  $120^{\circ}\text{C}$ . after one recrystallization from alcohol. The oil which remained after the separation of the crystals boiled at  $71^{\circ} - 76^{\circ}\text{C}$ . under a pressure of 8-10 mm. The distillate gave more crystals upon cooling. The oil and crystals gave the same amine upon reduction. This indicates that they are the two geometric isomers of the same oxime. Goldschmidt has stated that the solid form is the anti- form of the oxime. If this is the case, the oil is the syn- form, or a mixture of the two isomers.

Catalytic reduction of crotonaldoxime. (Dicrotonyl amine.) Adams, Cohen, and Rees<sup>21</sup> made a study of the catalytic reduction of compounds containing both the ethylenic linkage and nitro- groups, and found that, by using only absolute alcohol as a solvent without the addition of acid, the reduction took place only in the nitro- group. They assumed that the amine which was formed, being basic, poisoned the catalyst for

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21. Adams, Cohen, and Rees, J. Am. Chem. Soc., 49, 1093-9 (1927).

the reduction of the ethylenic linkage. The results of the above study indicated that catalytic reduction could be used for the reduction of crotonaldoxime to crotonyl amine without the formation of butyl amine from the reduction of the ethylenic linkage. The following experiment was performed: Four and one-fourth grams (0.05 mole) of crotonaldoxime were dissolved in 130 cc. of absolute alcohol and placed in the bottle of the shaking machine<sup>#</sup>. One-tenth gram of platinum black catalyst, prepared according to the directions given by Adams and Shriner<sup>22</sup>, was added to the solution and the bottle was put into the machine and shaken with an initial hydrogen pressure of 90 pounds. When the pressure had dropped as far as it would go the machine was stopped, the bottle was removed and the catalyst was filtered off from the solution of the amine. This solution was made acid with dilute hydrochloric acid and evaporated in vacuo. The residue was taken up in water, extracted with ether to remove any non-basic impurities, and evaporated again in vacuo. The residue was taken up in a small amount of water and decomposed with solid potassium

hydroxide to free the base which rose to the top of the solution

<sup>#</sup> This machine was a modification of the usual machine described in the literature, in that it had copper tubing connections instead of rubber tubing. The only rubber coming in contact with the reaction was the rubber gasket at the top of the bottle.

22. Adams and Shriner, J. Am. Chem. Soc., 45, 2171 (1923).

as an oil. This oil was separated and dried over solid potassium hydroxide. It had a boiling point of  $156-7^{\circ}$ . The compound had an herb-like odor. Three runs were made. The combined yield was 5.6 g. Dicrotonyl amine has a boiling point above  $130^{\circ}$  as reported by Mailhe<sup>23</sup>. A picrate of the base could not be prepared in a crystalline form as it formed an oil which could not be crystallized. A gold chloride double salt of the base was prepared. It had a melting point of  $156-7^{\circ}$  C.

Analysis: Calculated for  $C_8H_{16}NAuCl_4$ ; Au, 42.42%.

Found: Au, 42.51%, 42.42%.

The neutralization equivalent of the base was run with the following results:

Calculated for  $C_8H_{15}N$ : 125.0

Found: 125.2, 124.8, 124.6.

Reduction of crotonaldoxime by sodium amalgam. This procedure was carried out according to the directions given by Schindler<sup>24</sup>. Ten grams of crotonaldoxime were dissolved in 400 cc. of alcohol placed in a three-necked flask which was fitted with a stirrer. Thirty cc. of glacial acetic acid were added and the solution was stirred while 270 g. of 4% sodium amalgam were added over a period of about three hours. The reaction flask was cooled in a water bath so that the reaction never rose

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23. Mailhe, Bul. soc. chim., (4), 27, 226-29; 23, 235.  
24. Schindler, Op. cit., 415-16.

above room temperature. When all the amalgam had been added and the reaction had stopped, the solution was decanted from the mercury, made alkaline with solid sodium hydroxide, and steam distilled until the vapors gave no test for base with litmus. The distillate was made acid with hydrochloric acid and evaporated in vacuo to dryness. The residue was taken up in absolute alcohol and filtered to remove ammonium chloride. The alcoholic filtrate was evaporated in vacuo to dryness and the residue was taken up in a small amount of water. The base was freed from this solution by the addition of solid potassium hydroxide with cooling in an ice bath. The base separated in the form of an oil which rose to the top of the solution. It was separated, dried over solid potassium hydroxide, and distilled. The crotonyl amine boiled at 83-7° C. A higher boiling fraction was also obtained. It boiled up to 160° C. and probably contained some dicrotonyl amine. The crotonyl amine fraction, after standing overnight over solid potassium hydroxide, boiled at 84-5° C. The chloroplatinate double salt was prepared. It melted at 193° C., with decomposition, as reported by Schindler.

The bromination of the crotonyl amine was accomplished in the following manner: Five grams of crotonyl amine were placed in a beaker and enough hydrochloric acid was added to make it acid. The solution was cooled in an ice-salt bath and



bromine was added drop by drop, with stirring, until a faint yellow color persisted in the solution. A precipitate of dibromobutyl amine hydrochloride began to separate out before all the bromine was added. The mixture in the beaker was placed in a large vacuum desiccator and evaporated in the cold, to cut down loss by decomposition, until a crystalline mush was obtained. This was filtered with suction and dried. The yield was 17.2 g. or 91.5%. The melting point after three recrystallizations from alcohol was 199° C., with decomposition (darkened at 190° C.).

Analysis: Calculated for  $C_4H_{10}NBr_2Cl$ ; N, 5.23%,

Halogen as Cl, 48.01%.

Found: N, 4.79%, 4.68%; Halogen as Cl,

47.95%, 48.20%.

2-Butinyl amine. An attempt to prepare this compound was made, but the yield was so poor that further work on it was discontinued on account of the lack of materials. Three and five-tenths grams, (4 moles) of sodium were dissolved in 50 cc. of absolute alcohol in a pressure bottle. Ten grams (1 mole) of 2,3-dibromobutyl amine hydrochloride were added and the bottle was sealed and placed in a boiling water-bath and heated for 1 1/4 hours. At the end of this time the bottle and its contents were cooled to room temperature. The contents of the bottle were placed in a distilling flask and the alcohol and

base were distilled off from the residue of sodium bromide and sodium chloride which had separated during the reaction. The distillate was acidified with dilute hydrochloric acid and was evaporated to dryness in vacuo. The solid was extracted with 50 cc. of absolute alcohol and filtered to remove ammonium chloride, and the alcoholic solution was evaporated to dryness in vacuo. The residue was taken up in a small amount of water and solid potassium hydroxide was added to free the base. The yield was almost negligible; only a few droplets of an oil which had a strong odor resembling that of the ragweed were obtained. No physical constants or derivatives were obtained. The compound could probably be prepared if sufficient amounts of crotonyl amine were available.

2-bromoethyl amine hydrobromide. This compound was prepared by the hydrolysis of 2-bromoethylphthalimide with hydrobromic acid.

The 2-bromoethylphthalimide was prepared by the interaction of potassium phthalimide and ethylene bromide as described by Salzberg and Supniewski<sup>25</sup>. The material was used after one purification and recrystallization from alcohol in the presence of decolorizing charcoal. The melting point of this material was 79-80° C.; the melting point of the pure material is 83.5° C.

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25. Salzberg and Supniewski, *Organic Syntheses*, VII, 9, John Wiley and Sons, New York (1927).

The hydrolysis of the 2-bromoethylphthalimide was carried out as described by Gabriel and Stelzner<sup>26</sup>. Fifty grams of 2-bromoethylphthalimide, 50 cc. of 48% hydrobromic acid, and 25 cc. of glacial acetic acid were placed in a bomb tube and were heated for three hours at 170° C. The tube was cooled and opened, the contents were diluted with 150 cc. of water and the precipitated phthalic acid was filtered off with suction. The solution was evaporated in vacuo until a very thick brown syrup remained. Upon cooling, this syrup crystallized. This brown residue was taken up and recrystallized from hot alcohol. The yield of crude material was 25 g. or 62%. The melting point after one recrystallization from alcohol and ether and two recrystallizations from alcohol was 168-9° C. Gabriel<sup>27</sup> reports a melting point of 155-160° C.

Analysis: Calculated for C<sub>8</sub>H<sub>7</sub>NBr<sub>2</sub>; Br, 78.05%.

Found: Br, 78.15%, 78.25%.

The picrate was prepared by treating a water solution of the hydrobromide salt with a water solution of sodium picrate and concentrating the solution. The picrate was recrystallized from water and dried at 100° C. It melted at 130-131° C., as reported by Gabriel.

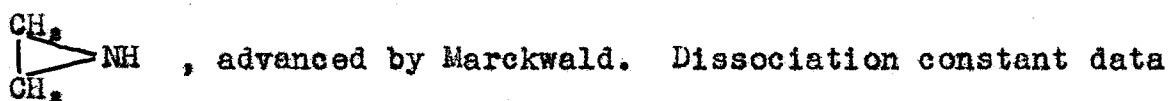
Vinyl amine. This amine was prepared from 2-bromoethyl amine hydrobromide by the removal of HBr with aqueous potassium

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26. Gabriel and Stelzner, Ber., 28, 2929 (1895).

27. Gabriel, Ber., 21, 567 (1888).

hydroxide solution according to Gabriel and Stelzner<sup>28</sup>. The base may also be prepared by heating the hydrobromide salt with water and silver oxide<sup>29</sup>. Formerly there was a controversy concerning the structure of this compound<sup>30</sup>, but in view of the evidence, the unsaturated amine structure,  $\text{CH}_2=\text{CH-NH}_2$ , seems more logical than the cyclic imine structure,



found by the author substantiate the unsaturated amine structure of Gabriel.

The compound was prepared as follows: The syrupy residue from the first evaporation, after the hydrolysis of 2-bromoethylphthalimide by hydrobromic acid was left in the distilling flask, which was fitted with a water-jacketed condenser, and 100 cc. of 33% potassium hydroxide solution was added. The mixture was warmed slightly and in a few minutes 2-bromoethyl amine separated as an oil at the bottom of the flask. Upon further heating this oil dissolved in the caustic solution with a slight crackling noise. After this reaction, the temperature was increased and the solution was distilled

28. Gabriel and Stelzner, Op. cit., 2929.

29. Gabriel, Op. cit., 567.

30. Gabriel, Ber., 21, 1049-57 (1888).

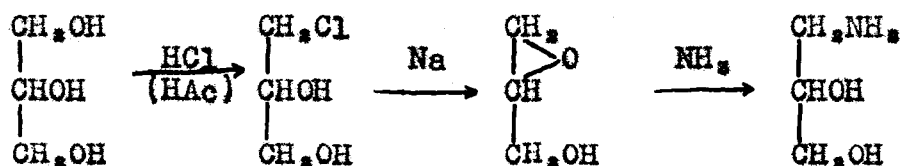
Howard and Marckwald, ibid., 32, 2036-38 (1899).

Duden and Macintyre, ibid., 33, 481-83 (1900).

Marckwald, ibid., 33, 764-66 (1900).

with a free flame until about 20 cc. of distillate had been obtained\*. The distillate was cooled in an ice-water bath and solid potassium hydroxide was added to salt out the base; this rose to the top of the solution in the form of a colorless oil. This oil was pipetted off; it was dried over solid potassium hydroxide, and was distilled. The fraction boiling at 55-69°C. was dried overnight over solid potassium hydroxide. After drying, it boiled at 56-56.5°C. The yield from 282 g. of 2-bromoethylphthalimide was 14 g. or 29.2%. The picrate was formed by adding an ether solution of picric acid to an ether solution of the base. It was filtered off by suction and recrystallized twice from alcohol to which a little ether had been added. The melting point was 142°C., with sintering beforehand, as reported by Gabriel.

Glyceryl amine (2,3-dihydroxypropyl amine). This compound was prepared from glycerol as indicated in the following steps:



The glycerol 1-monochlorohydrin was prepared by treating 90% glycerol with gaseous HCl at 105-110°C. using glacial

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\* Tests showed that 20 cc. of distillate contained most of the volatile base as none could be isolated from the next three 10 cc. portions.

acetic acid as a catalyst according to the directions given by Conant and Quayle<sup>31</sup>. The purified product distilled at 114-120° C. at a pressure of 14 mm.

The preparation of glycidol from glycerol 1-monochlorohydrin may be accomplished two ways, which are both modifications of the same general method. Aberhalden and Eichwald<sup>32</sup> treated an alcoholic solution of the bromohydrin with an alcoholic solution of sodium hydroxide. Bigot<sup>33</sup> prepared glycidol by treating an ether solution of the monochlorohydrin with 0.86 the calculated amount of sodium. Rider and Hill<sup>34</sup> modified this method by using a more dilute solution of the monochlorohydrin and only 0.7 of the calculated amount of sodium. Their investigations showed that the maximum yield was obtained by using these proportions. The method of Rider and Hill was used to prepare the glycidol for this synthesis.

Two hundred and twenty grams (2 moles) of glycerol 1-monochlorohydrin were dissolved in 1200 cc. of absolute ether in a 2-liter three-necked flask fitted with a mercury sealed stirrer and a reflux condenser. Thirty two and two-tenths grams (1.4 moles) of sodium were added in the form of thin flakes, and the stirrer was started. Sodium chloride began to separate from the reaction immediately. The mixture was allow-

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31. Conant and Quayle, *Organic Synthesis*, II, 29. John Wiley and Sons, New York (1922).  
32. Aberhalden and Eichwald, *Ber.*, 47, 2887 (1914).  
33. Bigot, *Ann. Chim. and Phys.*, (6) 22, 482-6 (1891).  
34. Rider and Hill, *J. Am. Chem. Soc.*, 52, 1525 (1930).

ed to react until all the sodium was used up. The sodium chloride was filtered off by suction and was washed with a little dry ether. The ether was distilled off on a water bath and the oil remaining was fractionated. The glycidol distilled at 70-75° C. at a pressure of 5 mm. The unused monochlorohydrin remained in the distilling flask. The yield of glycidol was 82 g., or 55%.

The preparation of glyceryl amine from glycidol was done according to the method of Knorr and Knorr<sup>35</sup>. Two hundred and fifty grams of 28% ammonium hydroxide solution were placed in a 1-liter balloon flask and 74 g. (1 mole) of glycidol were added. The contents of the flask were shaken to insure complete mixing and the mixture was allowed to stand overnight. The excess ammonia and the water were removed as completely as possible by evaporation in a wide evaporating dish on a steam hot plate. The residue was fractionated under reduced pressure. The glyceryl amine distilled at 132-136° C. at 7 mm. pressure. The yield from the first distillation was 39 g., or 43%. After drying for several months over solid potassium hydroxide, the boiling point of a sample of glyceryl amine was 137° C. at 7 mm. pressure. Repeated distillation tended to decompose the compound. Titration of samples, on the basis that only glyceryl amine and water were present, using sodium alizarin sulfonate

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35. Knorr and Knorr, Ber., 32, 752 (1899).

as an indicator, gave a purity of 69.7%. The 69.7% solution had a slight yellow color; its odor resembled that of ethanol amine; it was very hygroscopic. The oxalate, recrystallized from alcohol, had a melting point of 72-73° C. The chloroplatinate, recrystallized from water, had a melting point of 185° C.

Ethanol amine. The ethanol amine used in this study was obtained from Eastman Kodak Company. It was dried over solid potassium hydroxide and was purified by distilling under reduced pressure. It boiled at 61-62° C. at 5-6 mm.

#### Preparation of Arsonic Acids.

Allyl arsonic acid. This compound was prepared according to the general method given by Quick and Adams<sup>36</sup>. One hundred and ninety-eight grams (1 mole) of arsenous oxide were dissolved in 600 cc. of 10 N. sodium hydroxide (6 moles) in a 1-liter three-necked flask, fitted with a reflux condenser, mercury sealed stirrer, and a dropping funnel. The solution was heated nearly to boiling. When this temperature had been reached, 250 g. (2.06 moles) of allyl bromide were added dropwise over a period of about five hours, while the mixture was refluxed gently, with vigorous stirring. At the end of this time the reaction mixture was poured into a large beaker and

36. Quick and Adams, J. Am. Chem. Soc., 44, 810-11 (1922).



made just neutral to phenolphthalein with concentrated hydrochloric acid. The solution was then concentrated to about one-half its volume on the steam plate and the precipitated sodium chloride was filtered off. The hot filtrate was made acid with hydrochloric acid until it just turned Congo red to a blue color, and was cooled. The allyl arsonic acid separated in small needle-like crystals along with some sodium chloride. The mixture was heated to boiling and filtered hot. This removed most of the salt. The solution was cooled and allowed to stand overnight in the ice-box. The crystals which formed were filtered off. The yield was 201 g., or 59%. The melting point was 125° C. After two recrystallizations from alcohol and ethyl acetate, the melting point was 132° C. Adams and Quick report a melting point of 129° C. The allyl arsonic acid was in the form of very fine white needles.

2,3-dibromopropyl arsonic acid. This compound was desired as an intermediate in the preparation of glyceryl arsonic acid, but later developments showed that it was not satisfactory. It was prepared in the following manner: Thirty grams of allyl arsonic acid was dissolved in 150 cc. of water. The solution was cooled and 30 g. of bromine was added dropwise while the solution was stirred. The mixture became warm and the bromine was absorbed. Before all the bromine was added, the dibromopropyl arsonic acid began to separate from

the solution. The mixture was cooled to 0° C., and the product was filtered off by suction. The yield was 31 g., or 53%. The melting point, after two recrystallizations from alcohol, was 141.5° C. The crystalline form was small lustrous leaves.

Analysis: Calculated for  $C_2H_7AsO_3Br_2$ : Br, 49.05%.

Found: Br, 49.38%, 49.27%.

The mother liquor was concentrated in vacuo and allowed to stand in the ice-box for six days; feathery crystals appeared in the solution, but disappeared before they could be filtered.

The sodium salt was formed by adding sodium carbonate to a solution of the acid in water until it was just basic to litmus. Alcohol was added to the solution until a faint cloudiness showed. Then the solution was stirred and a few more drops of alcohol were added. The sodium salt separated in the form of fine white needles, which were filtered off and dried at 110° C.

The barium salt was prepared by adding equivalent amounts of the acid and barium chloride to hot water and stirring till they were dissolved. Ammonium hydroxide was added drop by drop; the solution turned cloudy at first, but with stirring a fine precipitate of small glistening platelets appeared. The addition of ammonia was continued until the solution was basic

to litmus, and the barium salt was filtered off with suction and washed with alcohol and ether, as it was insoluble in both these solvents.

Glyceryl arsonic acid. Attempts to prepare this compound from allyl arsonic acid by oxidation with alkaline potassium permanganate were not successful, as it was impossible to remove the finely divided manganese dioxide, which was formed in the reaction, from the solution in order to purify the product. An attempt was made to prepare the compound by the oxidation of 2,3-dibromopropyl arsonic acid with moist silver oxide, but no satisfactory results were obtained.

The glyceryl arsonic acid was finally prepared by using a modification of the method described by Eddee<sup>37</sup> for the preparation of 2-hydroxyethyl arsonic acid. This method consisted of the action of sodium hydroxide and arsenous oxide upon ethylene chlorohydrin. Eddee had great difficulty in obtaining the 2-hydroxyethyl arsonic acid in a crystalline form. The prediction was made that glyceryl arsonic would be more difficult to obtain in a crystalline form. Experiment proved this to be true.

Fifty grams (0.25 mole) of arsenous oxide and 60 g. (1.5 moles) of sodium hydroxide were dissolved in 150 cc. of water in a 1-liter three-necked flask, which was fitted with an efficient stirrer and a dropping funnel. The solution was

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37. Eddee, J. Am. Chem. Soc., 50, 1396 (1928).

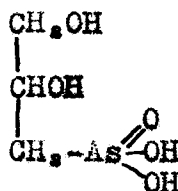
stirred and cooled to room temperature by placing the flask in a water-bath. Fifty-five grams (0.5 mole) of glycerol 1-monochlorohydrin were added dropwise during a period of about one-half hour, while the mixture was vigorously stirred. The reaction developed heat during the addition of the monochlorohydrin. When all the monochlorohydrin had been added and the temperature of the mixture had lowered to that of the room, the stirrer was stopped and the mixture was allowed to stand overnight. The reaction mixture was poured out into a beaker and was made acid to litmus with acetic acid. (No arsenous oxide precipitated from the acid solution.) The solution was then made strongly basic with ammonium hydroxide solution and an excess (as detected by starch-KI paper) of 30% hydrogen peroxide solution was added to oxidize the unreacted arsenous oxide to the pentavalent state. The pentavalent arsenic, so formed, was removed by the addition of magnesia mixture until no more precipitate was formed. The precipitate was removed by filtration, after the mixture had stood overnight.

After the removal of the pentavalent arsenic the filtrate was concentrated on the steam hot plate to about one-half its volume, and the mineral salts were filtered off. Hydrochloric acid was then added until the solution was just acid to Congo red. The solution was concentrated again and the metallic salts were filtered off. Absolute alcohol (200 cc.) was added

to precipitate more metallic salts, the solution was evaporated in vacuo and the process of addition of alcohol, filtration, and evaporation was repeated until no more metallic salts were thrown down by the addition of absolute alcohol. The final concentration was done in a vacuum desiccator.

No crystals could be obtained by cooling the syrup which was formed in the final concentration. The syrup, after standing for two years in a desiccator over concentrated sulfuric acid, solidified into a hard, transparent, glassy solid, which had no definite melting point. It was very hygroscopic; after it had stood in the air for only one hour, it became gummy.

The structure of glyceryl arsonic acid,



and the physical behavior of the compound indicated that perhaps intermolecular polymerization had taken place. A molecular weight determination was made to test this theory. The ebullition method, with absolute alcohol as the solvent, was used. The results did not check among themselves, but they all fell below the calculated molecular weight, and indicated that intermolecular polymerization had not taken place.

Barium glyceryl arsonate was prepared from the glassy

glyceryl arsonic acid. Two grams of the glassy glyceryl arsonic acid were dissolved in 20 cc. of water and 1.87 g. of barium carbonate were added to this solution. The solution was allowed to stand for an hour and then was filtered to remove the unreacted barium carbonate. The filtrate was made alkaline with 3 drops of ammonium hydroxide, and 100 cc. of alcohol was added. A heavy precipitate of glistening platelets formed. This precipitate was filtered off, washed with alcohol and ether, and sucked dry on a filter. The yield was 1.1 g.

Analysis: Calculated for  $\text{BaC}_2\text{H}_7\text{O}_5\text{As}\cdot 3\text{H}_2\text{O}$ : Ba, 35.27%;

$\text{H}_2\text{O}$ , 13.88%.

Found: Ba, 35.52%, 35.35%;  $\text{H}_2\text{O}$ , 13.91%.

Water was determined by drying at  $110^\circ\text{C}$ .

### Measurement of Dissociation Constants

Method for amines. It was evident, after reviewing the various methods given in the literature for the determination of dissociation constants of bases, that the method described by Carothers, Bickford, and Hurwitz<sup>38</sup> was the most accurate; it was also the most convenient to run. This method consists of half neutralizing a solution of base and then determining the hydrogen ion concentration of the half neutralized

38. Carothers, Bickford, and Hurwitz, J. Am. Chem. Soc., 49, 2908 (1927).

solution. The hydrogen ion concentration is equal to the hydrolysis constant of the neutral salt in the half neutralized solution. From this the dissociation constant of the base may be calculated by the relation

$$K_{\text{hyd}} = \frac{K_w}{K_b}$$

The best method for determining the hydrogen ion concentration of a series of solutions, which may vary from acidic to basic in nature, is by the use of the hydrogen electrode. However, some of the solutions, which were measured, poisoned the hydrogen electrode and made it impossible to use it for the measurement of the hydrogen ion concentration. In these cases, the pH was determined by the use of a quinhydrone electrode, or a suitable indicator was put in 10 cc. of the solution and a buffer solution was adjusted to give exactly the same color in 10 cc., with the same indicator. The hydrogen ion concentration of the buffer solution was then determined by the use of the hydrogen electrode.

The quinhydrone electrode can only be used when the solution to be measured is acidic in nature; in the case of most of the bases which were measured, it did not give absolutely constant readings. It was found that the initial readings were almost constant and these were used in all cases where the quinhydrone electrode was used for measuring the hydrogen ion concentrations of the solutions.

In the cases where the free bases could not be prepared, as 2,3-dibromopropyl amine, 2,3-dibromobutyl amine, and 2-bromoethyl amine, the method described by Moore<sup>39</sup> was attempted. This method consists of exactly half-neutralizing the hydrohalide salt of the base with standard sodium hydroxide solution and determining the hydrogen ion concentration of the resulting solution. Moore obtained exactly the same results by using aniline one-half neutralized with hydrochloric acid, and aniline hydrochloride one-half neutralized with sodium hydroxide. This method was not found to be satisfactory with 2,3-dibromopropyl amine hydrochloride and 2,3-dibromobutyl amine hydrochloride as the hydrogen ion concentration of the half neutralized changed upon standing and the readings which were obtained depended upon the length of time the solutions had been made up. The method seemed more satisfactory in the case of 2-bromoethyl amine hydrobromide, however, as it checked fairly well the results obtained by the use of the pure salt in solution. This is discussed later.

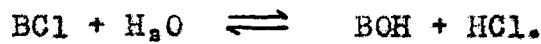
The dissociation constants of the above amines were calculated from the hydrolysis constants of the solutions of their hydrohalide salts in water. If the weight of salt used and the volume of solution are known, the total concentration of base existing as free base and as salt can be calculated. From

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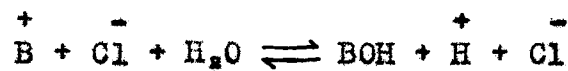
39. Moore, P. A., Iowa State College. Private communication. (1931).



these data the dissociation constant can be calculated as follows: If the equilibrium equation for the hydrolysis of the salt is written, we obtain



If we assume the salt and acid to be completely ionized in dilute solution, the equation may be written



The expression for the hydrolysis constant is then formulated:

$$\frac{(C_{\text{BOH}})(C_{\text{H}^+})(C_{\text{Cl}^-})}{(C_{\text{B}^+})(C_{\text{Cl}^-})(C_{\text{H}_2\text{O}})} = K_{\text{hyd.}} \quad (1)$$

If we multiply the left hand member of the expression by  $\frac{C_{\text{OH}^-}}{C_{\text{OH}^-}}$  we obtain

$$\frac{(C_{\text{OH}^-})(C_{\text{BOH}})(C_{\text{H}^+})}{(C_{\text{OH}^-})(C_{\text{B}^+})(C_{\text{H}_2\text{O}})} = K_{\text{hyd.}} = \frac{K_{\text{W}}}{K_{\text{b}}} \quad (2)$$

since the concentration of chloride ions is constant. From the hydrolysis equation

$$C_{\text{BOH}} = C_{\text{H}^+} \quad (3)$$

$$C_{\text{B}^+} = (C_{\text{BCl}} - C_{\text{H}^+}) \quad (4)$$

while the concentration of water may be taken as one. By substituting these values in equation (2), we obtain

$$\frac{(C_{\text{H}^+})^2}{(C_{\text{BCl}} - C_{\text{H}^+})} = K_{\text{hyd.}} = \frac{K_{\text{W}}}{K_{\text{b}}} \quad (5)$$

By solving equation (5) for  $K_{\text{b}}$ , we obtain

$$K_b = \frac{K_w (C_{\text{BCl}} - C_{\text{H}}^+)}{(C_{\text{H}}^+)^2} \quad (6)$$

At the half neutralization point

$$C_{\text{BOH}} = C_{\text{B}}^+ \quad (7)$$

and equation (2) becomes

$$C_{\text{H}}^+ = \frac{K_w}{K_b} \quad (8)$$

from which we obtain, by solving for  $K_b$ ,

$$K_b = \frac{K_w}{C_{\text{H}}^+} \quad (9)$$

The free amines were dried before use over solid potassium hydroxide. They were then purified by distillation in an atmosphere of dry air in all glass vacuum distillation apparatus as described by Ware<sup>40</sup>. This method enabled the amines to be sealed in small weighed bulbs immediately after distilling. In the case of low-boiling amines, like vinyl amine, the distillation was not carried out in vacuo, but at atmospheric pressure after the air in the apparatus had been dried by passing it through sulfuric acid. After the distillation, the system was evacuated by means of an oil pump for a few seconds; then dry air was allowed to enter the apparatus to fill the bulbs.

Before the bulbs were sealed, they were immersed first in hot, and then in cold water. The consequent expansion and contraction emptied the stems of the bulbs so they could be

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40. Ware, F. E., Op. cit., 14.

sealed easily without decomposing any of the amines. After they were sealed, the bulbs were washed in alcohol and then in acetone.

Samples of each fraction of each amine were titrated for purity with 0.1031 N. HCl, using sodium alizarin sulfonate as an indicator. The purity entered into the calculations for the amount of acid required for half neutralization if it fell below 99.7%. The impurity was assumed to be water in all cases.

The solutions for the determination of the hydrogen ion concentration of the half neutralized free bases were made up by placing weighed bulbs containing the amines in clean, dry 100 cc. volumetric flasks; the calculated amounts of 0.0531 N. HCl were added and the bulbs were broken by means of a glass rod. This was washed with conductivity water while it was being removed from the flasks. The flasks were then filled to the mark with conductivity water, and they were placed in a water thermostat at 25° C. until the measurements of hydrogen ion concentration were made.

The salts of the bases were purified before being used by careful recrystallization and drying in a vacuum desiccator over sulfuric acid. The purity was checked by melting point and analysis for nitrogen or halogen. Samples were weighed in small tubes and these were placed directly in clean, dry 250 cc. volumetric flasks to avoid loss by brushing or spilling. The

flasks were then filled to the mark with conductivity water. The solutions were kept in a water thermostat at 25° C. until the measurements were made.

The measurements, both with the hydrogen electrode and the quinhydrone electrode, were made against a saturated calomel half cell at 25° ± 0.1° C. The electrodes were checked, from time to time, with a 0.05 molar potassium hydrogen phthalate buffer solution.

Method for arsonic acids. The dissociation constants of the arsonic acids were determined by a modification of the method used for those of the bases<sup>41</sup>. It was assumed that only one hydrogen of the arsonic acids was ionized. This assumption was based upon the fact that in the di- and tri-basic inorganic acids, only one hydrogen is ionized to any appreciable extent. Table I shows a comparison of the dissociation constants for the first, second, and third hydrogens in several inorganic acids.<sup>42</sup>

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41. Carothers, Bickford and Hurwitz, Op. cit., 2908.

42. VanNostrand's Chemical Annual, D. VanNostrand Company, New York. page 74 (1926).

TABLE I

A COMPARISON OF THE DISSOCIATION CONSTANTS  
FOR THE DIFFERENT HYDROGENS IN SEVERAL  
INORGANIC ACIDS

Acid	$K_a$ 1st H	$K_a$ 2nd H	$K_a$ 3rd H
$H_3PO_4$	$1.1 \times 10^{-2}$	$2 \times 10^{-7}$	$3.6 \times 10^{-13}$
$H_3PO_3$	$5. \times 10^{-2}$	$2 \times 10^{-5}$	
$H_2SO_3$	$1.7 \times 10^{-2}$	$5 \times 10^{-6}$	
$H_2CO_3$	$3.0 \times 10^{-7}$	$6.0 \times 10^{-11}$	

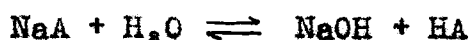
It was assumed from this data that the dissociation constant for the second hydrogen in the arsonic acids would be of the order  $1 \times 10^{-7}$  or  $1 \times 10^{-8}$  or less, and therefore would not interfere with the measurement of the dissociation constant for the first hydrogen.

As a check on the above point, a sample of allyl arsonic acid was titrated electrometrically with 0.1109 N. NaOH by the use of the quinhydrone electrode. Only one break appeared in the curve. The neutralization equivalent calculated from the end-point is 166.5. The molecular weight of allyl arsonic acid is 166. This shows that allyl arsonic acid acts

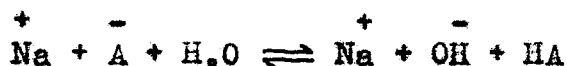
like a monobasic acid in its ionization behavior. The titration curve is shown in Figure III.

The arsonic acids were exactly half-neutralized with a standard sodium hydroxide solution, and the hydrogen ion concentrations of the resulting solutions were determined by the use of a quinhydrone electrode. The electrode gave constant readings and was satisfactory in every way.

The dissociation constants were calculated in the following manner: If the equilibrium reaction for the hydrolysis of the sodium salt of an acid is written, we obtain:



If we assume the salt and base to be completely ionized in dilute solution, we may write



The expression for the hydrolysis constant is then formulated:

$$\frac{(C_{\text{Na}^+})(C_{\text{OH}^-})(C_{\text{HA}})}{(C_{\text{Na}^+})(C_{\text{A}^-})(C_{\text{H}_2\text{O}})} = K_{\text{hyd.}} \quad (1)$$

If we multiply the left hand member of equation (1) by  $\frac{C_{\text{H}^+}}{C_{\text{H}^+}}$  we obtain

$$\frac{(C_{\text{H}^+})(C_{\text{OH}^-})(C_{\text{HA}})}{(C_{\text{H}^+})(C_{\text{A}^-})(C_{\text{H}_2\text{O}})} = K_{\text{hyd.}} = \frac{K_w}{K_a} \quad (2)$$

since the concentration of sodium ions is constant. By dividing the left hand member of equation (2) by  $\frac{C_{\text{H}^+}}{C_{\text{H}^+}}$  and setting

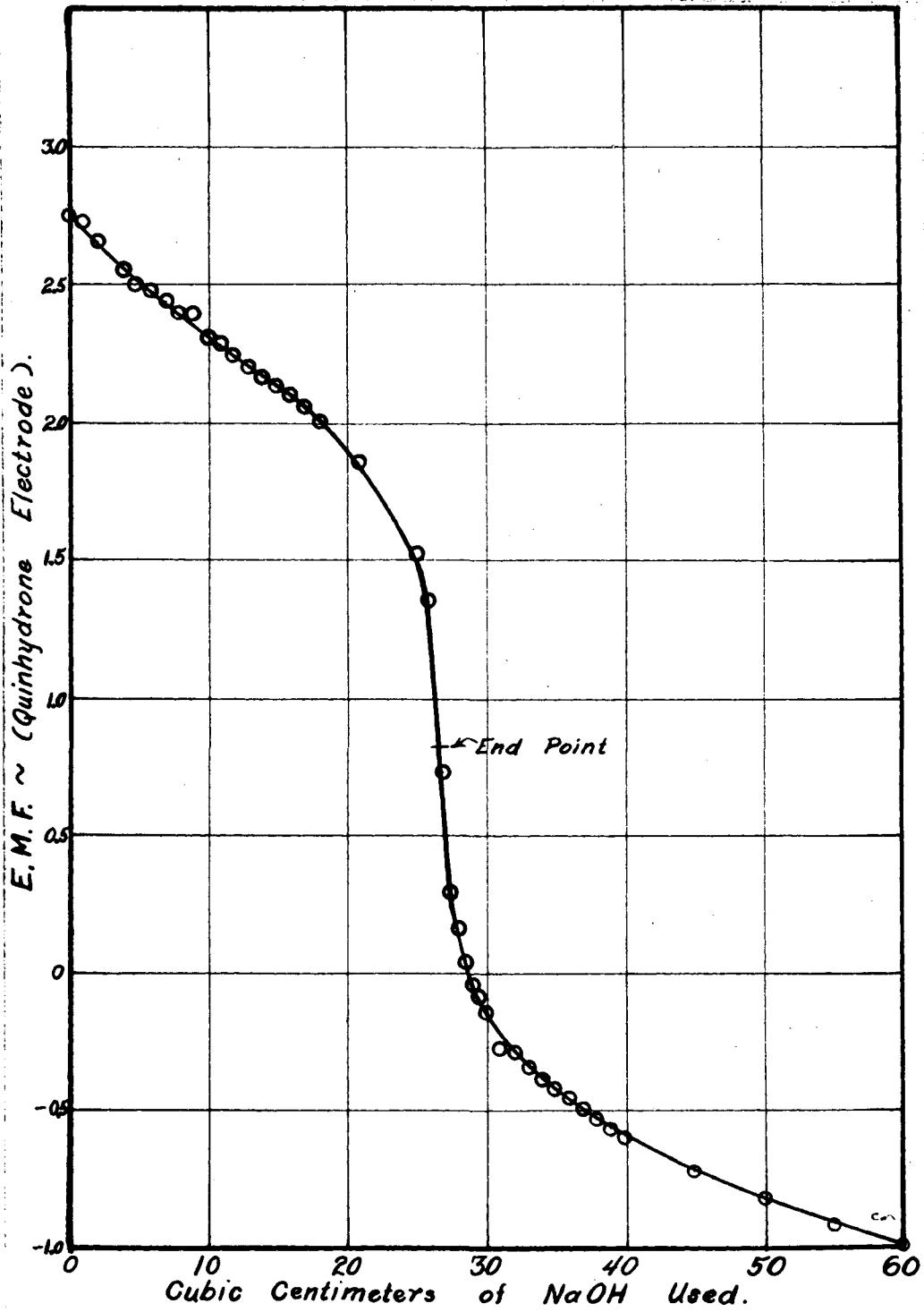


Fig. 3:  
Titration curve for allyl arsonic acid.

$C_{H_2O}$  equal to one, we obtain

$$\frac{(C_{OH^-})(C_{HA})}{C_A^+} = \frac{K_W}{K_a} \quad (3)$$

At the half-neutralization point,

$$C_{HA} = C_A^- \quad (4)$$

By substituting equation (4) in equation (3), we obtain

$$C_{OH^-} = \frac{K_W}{K_a} \quad (5)$$

from which we obtain, since  $K_W = (C_H^+)(C_{OH^-})$ ,

$$K_a = \frac{K_W}{C_{OH^-}} = C_H^+ \quad (6)$$

Before the solutions were made up for the determination of the hydrogen ion concentration, the arsonic acids were carefully recrystallized and dried in a vacuum desiccator over sulfuric acid. Samples were then weighed out in small tubes and placed directly in clean, dry 250 cc. volumetric flasks to avoid loss by brushing and spilling. A little conductivity water was then added to dissolve the samples. The calculated amount of 0.0571 N. NaOH was then added to half-neutralize exactly the acid, and the flasks were filled to the mark with conductivity water. The flasks were placed in a water thermostat at 25° C. until the hydrogen ion concentrations were determined.

In the case of glyceryl arsonic acid, the barium salt



was weighed and the free acid was obtained in solution by adding exactly the calculated equivalent amount of 0.0560 N.  $H_2SO_4$ . The precipitated barium sulfate was filtered off in the cold. The acid was then half-neutralized and the usual procedure was followed.

All measurements, as with the amines, were made against a saturated calomel half cell at  $25 \pm 0.1^\circ C$ . The quinhydrone electrode was checked, from time to time, with a 0.05 molar potassium acid phthalate buffer solution.

Results of the dissociation constant measurements. The results of the dissociation constant measurements on the amines are given in Table II. It will be noted that two values are given for the dissociation constant for 2-bromoethyl amine. The value for glyceryl amine may or may not be correct as the compound could only be obtained 69.7% pure. The half neutralization was made under the assumption that the impurity was water. The results of the dissociation constant measurements on the arsonic acids are given in Table III.

The results given in the table are shown graphically in Figure 4. The values for the electron-sharing ability of the radicals were determined from the values of the measured dissociation constants and the  $R-NH_2$  curve. In all the cases these values, when plotted on the  $R-CH_2-NH_2$  curve against the electron-sharing ability for R, gave very good agreement. The

values for the electron-sharing ability of the radicals, when no amine dissociation constants were given in the literature, were calculated from the equation

$$\log K_a = 20 (e^{9.5x - 5.6} - 0.24)$$

given by Hixon and Johns for the carboxylic acids, R-COOH.

The curve for the arsonic acids was obtained by plotting the values given by Ware<sup>43</sup> against the values of the electron-sharing ability of R obtained from the amine curve in the regular manner. Good agreement was obtained in all cases.

As will be noted from the graph, the values for the electron-sharing ability of the radicals measured in this study fall very close together, but the difference in the negative effects of the substituents is marked by a regular shift in position.

TABLE II

DISSOCIATION CONSTANTS OF SEVERAL NEGATIVELY  
SUBSTITUTED ALIPHATIC AMINES

Amine	E.M.F.* Hydrogen Electrode	E.M.F.* Quinhyd. Electrode	Salt or 1/2 neut. Base	$K_b$ **
2,3-dibromo- propyl amine	0.5707 <sup>(a)</sup>	---	Salt	$5.29 \times 10^{-6}$
2,3-dibromo- butyl amine	---	0.1291	Salt	$5.67 \times 10^{-6}$
2-bromoethyl amine	0.4937	0.1915	Salt	(b) $2.27 \times 10^{-8}$
2-bromoethyl amine	0.6438	---	1/2 neut.	(b) (c) $5.52 \times 10^{-8}$
Glyceryl amine	0.7806	---	1/2 neut.	$1.14 \times 10^{-5}$
Ethanol amine	0.7963	---	1/2 neut.	$2.05 \times 10^{-5}$
Vinyl amine	0.7104	---	1/2 neut.	$7.30 \times 10^{-7}$

\* The E.M.F. values given are average values against a saturated calomel half cell.

\*\* The values of  $K_b$  are average values.

(a)  $KH_2PO_4$  buffer solution and brom cresol purple were used.

(b) For a discussion of this value see page 52.

(c) Moore's method.

TABLE III

DISSOCIATION CONSTANTS OF SEVERAL NEGATIVELY  
SUBSTITUTED ARSONIC ACIDS

Arsonic Acid	E.M.F.* Quinhydrone Electrode.	Equivalents of Base Added	$K_a$ **
2,3-dibromo- propyl	0.2490	1/2	$3.6 \times 10^{-4}$
Glyceryl (a)	0.2160	1/2	$9.88 \times 10^{-5}$
Allyl	0.2175	1/2	$1.06 \times 10^{-4}$

\* The E.M.F. values given are average values against a saturated calomel half cell.

\*\* The values of  $K_a$  are average values.

(a) The acid was liberated from the barium salt.

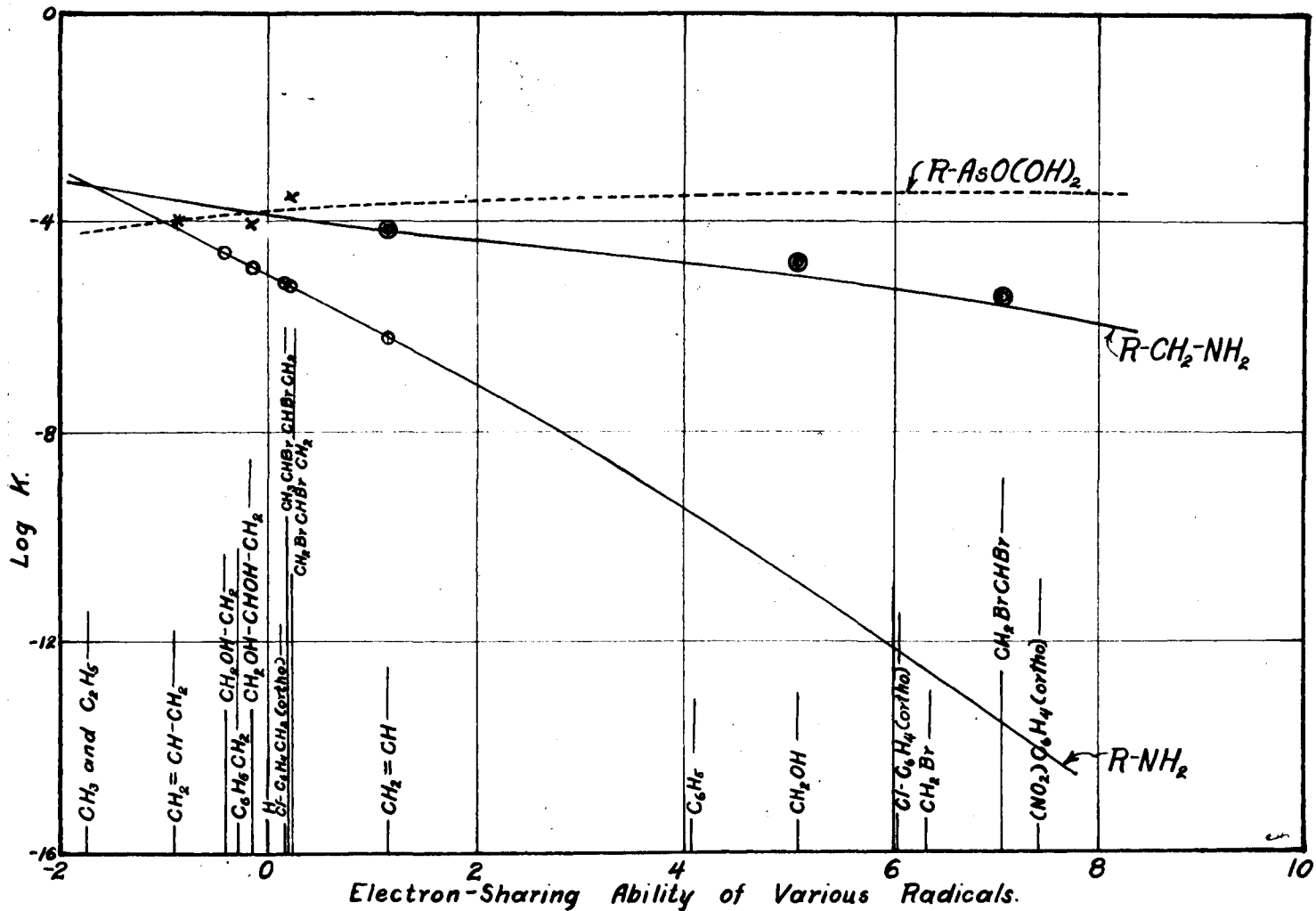
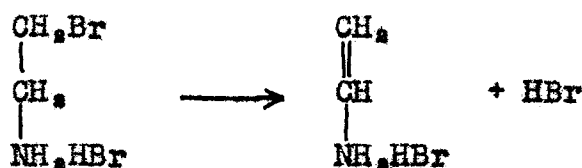


Fig. 4:  
Curves showing the results of the experimental work.

The values for the dissociation constant of 2-bromoethyl amine by measurements both of the half-neutralized salt and of the neutral salt were abnormally low when checked by the graphical method of plotting  $\log K_b$  for  $\text{CH}_2\text{Br}-\text{CH}_2-\text{NH}_2$  against the electron-sharing ability for  $\text{CH}_2\text{Br}$ -. In the half neutralized solution the values of the E.M.F. came to a maximum, as is customary with the hydrogen electrode, and then gradually fell off over a period of about two hours, which was the longest any one solution was tested. The values in the solutions of the neutral salt remained constant. This indicated that in the solution of the salt there might be a reaction of the following nature:



The presence of the HBr in the solution would increase the hydrogen ion concentration and would consequently lower the measured dissociation constant.

In order to test for the presence of the above reaction the following experiment was performed: A sample of 2-bromoethyl amine hydrobromide was weighed and dissolved in 200 cc. of conductivity water. This solution was arranged in an electrometric titration apparatus so that standard base could be added and E.M.F. measurements could be taken by means of a hydrogen electrode against a saturated calomel half cell. A

reading was taken on the solution of the salt and the calculated amount of 0.0571 N. NaOH for 0.3 equivalent of the salt was added. A reading of the E.M.F. was taken and more base was added from time to time in order to bring the E.M.F. up to the value for 0.3 equivalent of base which was taken as an arbitrary standard for comparison.

The results of the above experiment are shown in Table IV. This table shows the number of equivalents of base added, the drift in the E.M.F. values before the base was added, and the time that elapsed during the experiment. It will be noted that the E.M.F. drifts toward a higher value after the addition of 1.31 equivalents of base. This is probably due to the fact that 1.31 equivalents of base is a trifle more than should have been added at that time. This is also the case with the values after the addition of 1.42 equivalents of base. It is interesting to note that the E.M.F. reading for the solution after the addition of 1.50 equivalents of base checks the reading for the half neutralized solution of vinyl amine fairly closely (0.7130 compared to 0.7104 for vinyl amine). The value when 2.00 equivalents of base were added was 0.7860 v. This checks fairly well the value of 0.7780 v. for a solution of vinyl amine of approximately the same concentration when it is remembered that the above solution may have a mixture of  $\text{CH}_2\text{Br}-\text{CH}_2-\text{NH}_2\cdot\text{HBr}$ ,  $\text{CH}_2\text{Br}-\text{CH}_2-\text{NH}_2$ ,  $\text{CH}_2=\text{CH}-\text{NH}_2$ , NaBr, and NaOH present in it.

The graph of equivalents of NaOH required to keep a constant reading of E.M.F. against the time elapsed is shown in Figure 5.

The results of the above experiment show without a doubt that the measurement of the dissociation constant of 2-bromoethyl amine by any method, which depends upon the measurement of some property of an aqueous solution of the compound, is impossible. The value in the table is in error and should not be used for the determination of the electron-sharing ability of the 2-bromoethyl radical.



TABLE IV

TABLE SHOWING THE DRIFT IN E.M.F. VALUES WITH TIME IN A  
PARTIALLY NEUTRALIZED SOLUTION OF 2-BROMOETHYL  
AMINE HYDROBROMIDE

Time in Minutes	Equivalents of Base Added	E.M.F. by the Hydrogen Electrode
0	0	0.4960
0*	0.3	0.6880
30	0.3	0.6490
30*	0.5	0.6880
45	0.5	0.6740
45*	0.59	0.6880
60	0.59	0.6810
60*	0.65	0.6880
75	0.65	0.6760
75*	0.75	0.6880
125	0.75	0.6450
125*	0.95	0.6880
160	0.95	0.6650
160	1.00	0.6710
160*	1.13	0.6880
190	1.13	0.6750
190*	1.31	0.6880
205	1.31	0.6890
215	1.31	0.6960
225	1.31	0.6880
235	1.31	0.6780
235*	1.42	0.6880
240	1.42	0.6920
260	1.42	0.7030
260	1.50	0.7130
270	1.50	0.7090
280	1.50	0.7040
290	1.50	0.7010
325	1.50	0.6970
325	1.75	0.7350
340	1.75	0.7350
370	1.75	0.7270
370	2.00	0.7880
385	2.00	0.7830
390	2.00	0.7810
395	2.00	0.7760

\* These points are plotted in Figure 5.

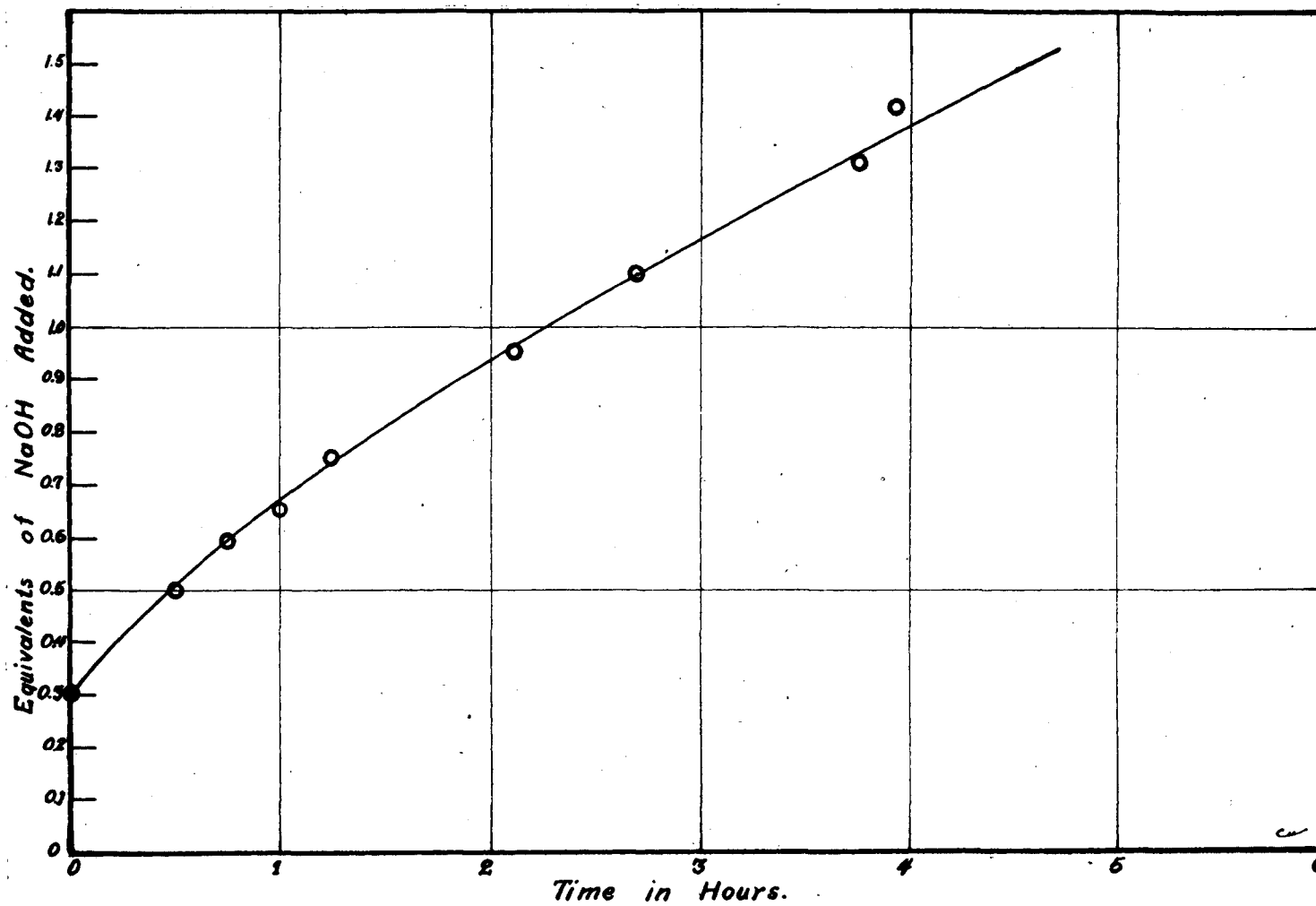


Fig. 5:  
 Curve showing the amounts of base necessary to keep the E.M.F. of 2-bromoethyl amine solution constant over a period of time.

## DISCUSSION

The experimental work shows that the method for the determination of dissociation constants as ordinarily used will not give the correct results for 2-bromoethyl amine. The instability of the compound leads to erroneous results. A close approximation of the electron-sharing ability of the 2-bromoethyl radical might be obtained by the measurement of the dissociation constant of 2-chloroethyl amine. The chlorine should be more firmly bound to the carbon and therefore the measurement of the hydrogen ion concentration of the solution of 2-chloroethyl amine hydrohalide should give a measure of the dissociation constant that would not be in error like that of the bromo-compound.

The structure of vinyl amine is indicated to be  $\text{CH}_2=\text{CH}-\text{NH}_2$  from the dissociation constant. If it had the cyclic imine structure, its dissociation constant would be more in the order of the constants for pyrrolidine and piperidine, although less, due to the effect of the smaller ring.  $K_b$  for pyrrolidine is  $1.3 \times 10^{-3}$ , and  $K_b$  for piperidine is  $1.6 \times 10^{-3}$ . It is hardly to be expected that, with the removal of two  $\text{CH}_2$ - groups, the dissociation constant would drop from the order of  $10^{-3}$  to that of  $10^{-7}$ . Furthermore, the graphical agreement of the value of  $\log K_b$  for allyl amine, when plotted against the value of the electron-sharing ability found for the vinyl radical by

means of the R-NH<sub>2</sub> curve, is extraordinarily close to the established R-CH<sub>2</sub>-NH<sub>2</sub> curve.

### CONCLUSIONS

From the limited data presented in this study the following conclusions may be safely made:

1. The effect of substitution of negative groups on the electron-sharing ability of aliphatic radicals is most marked in the smaller radicals.

2. The introduction of a second negative group into an aliphatic radical, if it is attached to a carbon atom other than the first, has relatively less effect than the introduction of the first negative group.

3. The value of the electron-sharing ability of the vinyl radical is almost exactly half way between the values of the methyl and phenyl radicals.

4. It is impossible to measure the dissociation constant of 2-bromoethyl amine by any method which depends upon the aqueous solution of the compound.

SUMMARY

1. Diacrotonyl amine has been prepared by the catalytic reduction of crotonaldoxime.
2. 2,3-dibromobutyl amine hydrochloride has been reported for the first time.
3. An attempt to prepare 2-butinyl amine has been made. The lack of materials prevented further study of this compound.
4. 2,3-dibromopropyl arsonic acid has been reported for the first time. Its sodium and barium salts have been prepared.
5. Glyceryl arsonic acid has been reported for the first time. Its barium salt has been prepared.
6. Dissociation constants for several negatively substituted aliphatic amines have been determined.
7. Dissociation constants for three negatively substituted aliphatic arsonic acids have been determined.
8. The electron-sharing abilities of the radicals have been determined from the dissociation constants of the above compounds. They were found to vary regularly in accordance with the theory.
9. Substantiation for the structure of vinyl amine has been advanced.
10. It has been found impossible to measure the dissociation constant of 2-bromoethyl amine in aqueous solution.

APPENDIX

The value of  $5.75 \times 10^{-7}$  given by Bourgeaud and Dondelinger<sup>44</sup> for the dissociation constant of indanyl amine is probably in error. When the values of  $\log K_p$  for the  $R(\text{CH}_3)_2\text{N}$  and  $R(\text{C}_2\text{H}_5)_2\text{N}$  series are plotted (Figure 6) against the electron-sharing ability values for R, which were determined from the dissociation constants of the  $\text{RNH}_2$  compounds, a smooth curve results. The only eccentricity is found in the values for dimethyl indanyl amine and diethyl indanyl amine. This eccentricity could result in one of two ways: either the value of the primary indanyl amine dissociation constant is wrong, or the tertiary indanyl dimethyl amine and indanyl diethyl amine were not pure and an error was made in their measurement.

The only criterion of purity of the amine hydrochlorides used by Bourgeaud and Dondelinger was the titration of the hydrochlorides with NaOH in the presence of phenolphthalein.

The theory of electron-sharing ability of organic radicals is either invalid in the case of the indanyl radical or the constants reported by Bourgeaud and Dondelinger are in error. The theory holds for the amines, carboxy acids, arsonic acids, organo-mercuric nitrates, organo-mercuric iodides,  $\alpha$ -substituted pyrrolidines, and N-substituted pyrrolidines. Therefore it is 44. Bourgeaud and Dondelinger, Op. cit., pp. 277-87.

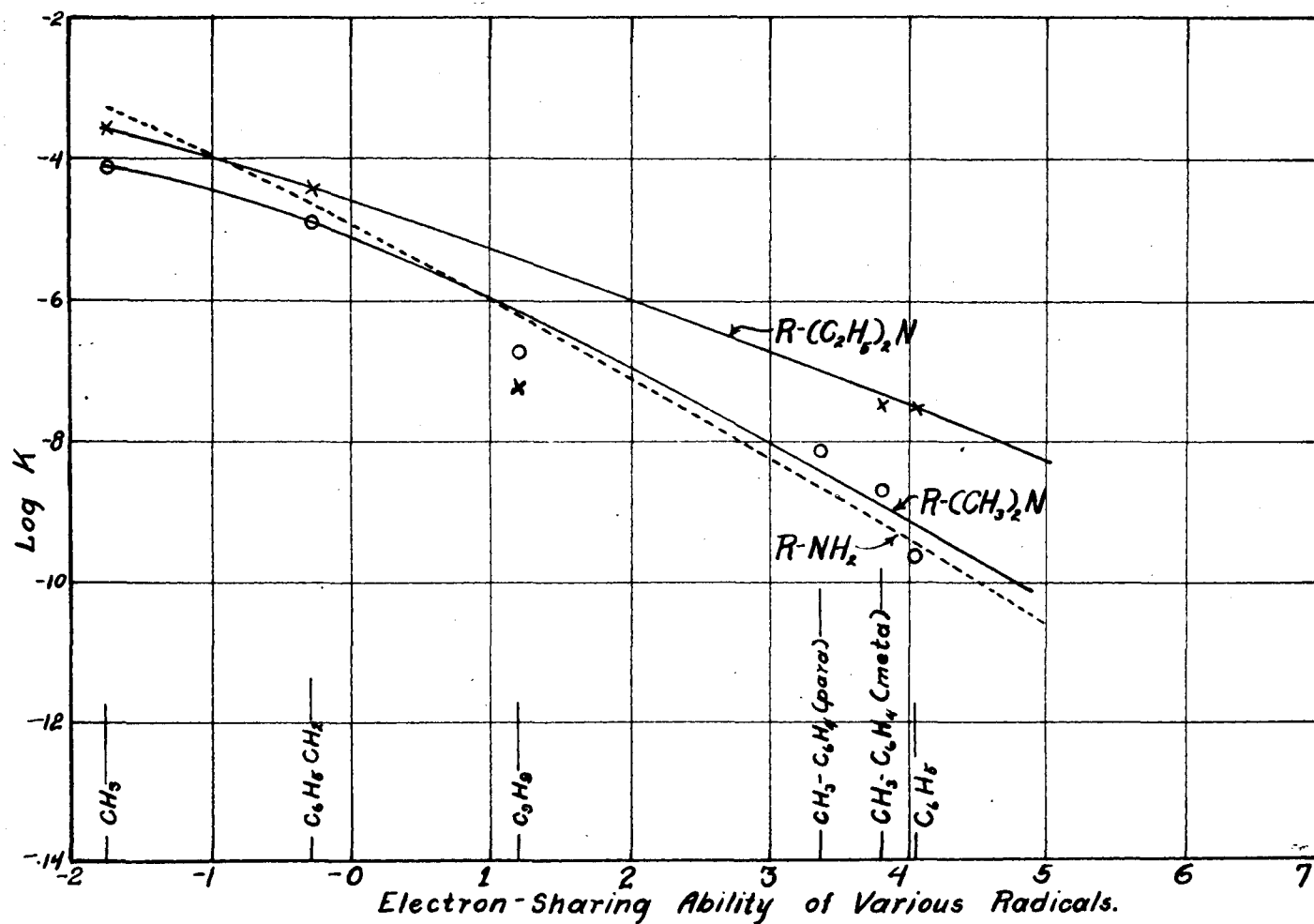
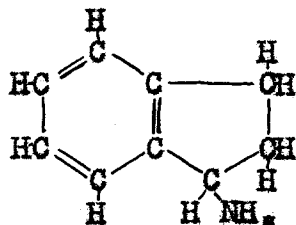


Fig. 6:  
Curves showing the discrepancy in the value for the indanyl radical by means of the dimethyl and diethyl tertiary amine series.

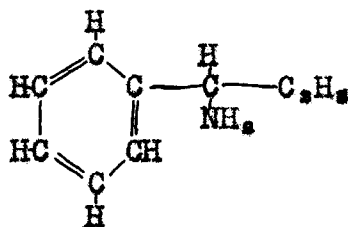


possible that the measurement is in error in some way.

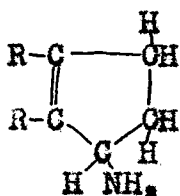
From a consideration of the structure of indanyl amine,



it can be considered to be closely related in structure to phenyl-ethyl-methyl-amine,



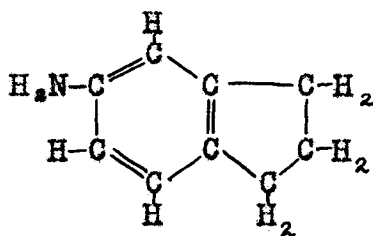
the dissociation constant of which should be near that of benzylamine,  $2 \times 10^{-5}$ . The same prediction is reached by considering the amine,



to be closely related to the aliphatic amine,  $\text{CH}_2\text{-CH=CH-CH}_2\text{-NH}_2$ , which has a dissociation constant of  $4 \times 10^{-4}$ . The R- groups are negative and would tend to make the constant lower but not as low as  $10^{-7}$  as reported by Bourgeaud and Dondelinger.

If the negativity in the phenyl ring were greatly shifted

toward the five membered ring, an explanation of the low dissociation constant could be made. This shift of negativity could be proved by the measurement of the dissociation constant of 5-amino-indene:



If the measurement shows this base to have a dissociation constant more positive than aniline, the negativity in the phenyl ring has been absorbed in the five membered ring and the value given by Bourgeaud and Dondelinger may be correct. If, however, the value is in the neighborhood of  $10^{-10}$ , which is the region where the xylidines' dissociation constants are found, the negativity has not shifted and the value given by Bourgeaud and Dondelinger should be incorrect.

In view of the above considerations the author is inclined to believe that the values given by the above investigators are rather seriously in error with respect to the indanyl radical and that the indanyl radical should not be in as negative a position in the series as their values would place it.