Picloram: Solubility and Acid–Base Equilibria Determined by Normal Pulse Polarography

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Solubility of picloram (4-amino-3,5,6-trichloropicolinic acid) in aqueous solution (1% 2-ethoxyethanol, μ = 0.3 M, Na_2SO_4) has been determined at 25 °C over the pH range 0.8–2.7. Solubilities are as follows: pH 0.84, 0.26 mM; 0.93, 0.34; 1.52, 0.35; 1.80, 0.58; 2.22, 1.02; 2.68, 1.87. Analysis of the solubility data as a function of pH yields the pK_a' value of 1.97 for the concentration acid dissociation constant of picloram under the above conditions.

The solubility and acid–base properties of picloram in aqueous solutions as a function of pH have an important bearing on the problems of its adsorption on soil and its potential as an environmental contaminant. These properties have been discussed in detail by Cheung and Biggar who have reviewed the literature on this subject and reported values of solubility for picloram as a function of temperature and pH and also values of the picloram acid dissociation constant as a function of temperature (Cheung and Biggar, 1974). The solubility and acid–base equilibrium constants appeared to us to be inconsistent with each other, and therefore it seemed prudent to redetermine these quantities by an independent method. Both the solubility and pK_a measurements are difficult due to the limited solubility of picloram in these solutions and the tendency of the neutral molecule, at equilibrium a solid at room temperature, to form oils which are difficult to remove from glassware and which tend to form emulsions with aqueous salt solutions. The limited solubility makes it difficult to determine pK_a values by potentiometric titration, the method of Cheung and Biggar, and also increases the likelihood of error in solubility measurements based on direct determination of picloram in solutions with the oil or microscopic solid particles. Because of the tediousness, difficulty, and potential for error of previously employed methods, we elected to determine solubility of picloram by normal pulse polarography (NPP).

The normal pulse polarographic method has advantages of sensitivity, simplicity, rigor, and freedom from interference due to presence of insoluble material. The use of pulse polarography for determination of picloram has been reported (Whittaker and Osteryoung, 1979). In the normal pulse polarographic mode the sensitivity for picloram is about 40 μA/mM and the limit of detection about 10^{-6} M. Calibration curves show linearity from the 10^{-6} M range to the solubility limit (ca. 10^{-4}–10^{-3} M in the present work). In equilibrated solutions calibration curves show no further increase in limiting current with increasing amounts of added picloram when the solubility is reached. By exploiting the quantitative theoretical relationship between current and concentration for diffusion controlled waves in normal pulse polarography (Osteryoung and Hasebe, 1976) and a unique feature of our instrumentation, the ability to change the position of the current sampling window, we were able to ensure that this effect was due to limited solubility rather than to a change in the sensitivity of the method.

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EXPERIMENTAL SECTION

Polarograms were obtained with a PARC Model 174 polarographic analyzer used with a PARC Model 174/70 drop timer and Houston omnigraphic Model 2000 X-Y recorder. The dropping mercury electrode had a flow rate of 2.20 mg/s at open circuit in distilled water. A Taccusel UAP 4 pulse polarograph used with a Taccusel PRT 30-0.1 potentiostat was used to obtain current-time profiles for individual pulses. These current-time profiles were obtained by varying the position of a current sampling window over the width of a constant width pulse. Standard Brinkmann cells were used with saturated calomel reference and platinum counter electrodes. Other experimental detail is reported elsewhere (Whittaker and Osteryoung, 1979).

Picloram standards (>99.9% pure) were obtained from Dow Chemical Co. All experiments were done in 0.1 F Na_2SO_4 (μ = 0.3) + H_2SO_4 + 1% 2-ethoxyethanol (2EE).

Experimental Procedures. Solid picloram was placed in screw-top vials in which the cap was protected with metal foil. Ten milliliters of 0.1 F Na_2SO_4 (μ = 0.3 M) 1% in 2-ethoxyethanol (2EE) was added to each vial along with enough H_2SO_4 to adjust the pH to the approximate desired value [2EE was added to optimize the NPP method (Whittaker and Osteryoung, 1979)]. Amounts of picloram added at various nominal pH values were pH 1.0, 1.5, and 2.0, 5 mg/10 mL; pH 2.5, 10 mg/10 mL; and pH 3.0, 20 mg/10 mL. The test tubes were placed in an ultrasonic cleaner for 15 min and then equilibrated in a covered ethylene glycol bath which was maintained at 25.0 ± 0.2 °C with a Haack Model FJ thermoregulator. After equilibration, solution was removed from each vial with a pipet and placed in a thermostated (25.0 ± 0.2 °C) polarographic cell. Standards 1 × 10^{-4} F in picloram were run at the same temperature at each nominal pH value to obtain the true sensitivity of the measurement at that value. True pH values were determined by direct measurement after each polarographic run.

RESULTS AND DISCUSSION

If a polarographic limiting current obeys the Cottrell equation

$$i_l = nFAD^{1/2}C^b/(\pi t)^{1/2} \quad (1)$$

where i_l is the limiting current (μA), n the number of electrons transferred, F the Faraday (C), A the electrode area (cm^2), D the diffusion coefficient (cm^2 s^{-1}), C^b the bulk concentration (mM), and t the time (s), then the reaction giving rise to the faradaic current is said to be diffusion controlled. If it can be established through examination of the i_l-t behavior that a reaction is diffusion controlled over a range of concentrations, then variation in the apparent value of the constant $i_l/C^b = nFAD^{1/2}/(\pi t)^{1/2}$ at

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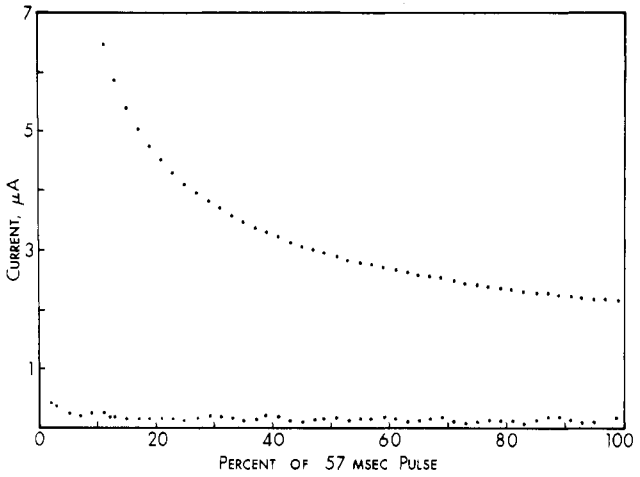

Figure 1. Current-time behavior during pulse application for picloram. Picloram concentration, 1.0×10^{-3} F; supporting electrolyte, 0.1 F Na_2SO_4, pH 1.9 (H_2SO_4), 1% 2EE; sampling time, 1.1 ms. Lower curve for supporting electrolyte only.

Table I. Picloram Solubility as a Function of pH (0.1 Na_2SO_4 + H_2SO_4 (μ = 0.3); 1% 2EE, 25 °C)

pH	solubility, mM
0.84	0.26
0.93	0.34
1.52	0.35
1.80	0.58
2.22	1.02
2.68	1.87

fixed A and t, if free from artifact, must be explained by concluding that the nominal value of C^b is not the true value. In the case of a slightly soluble substance a constant sensitivity (the ratio i_l/C^b) is obtained at sufficiently low concentrations but the sensitivity near the solubility limit is less because the true bulk concentration is less than the nominal concentration due to precipitation of material. The observed behavior is often complicated by kinetic phenomena involving the solid-solution equilibria. However, the polarographic measurement, under conditions of diffusion control, gives accurately the concentration of dissolved material and is not influenced by the presence of undissolved material. Therefore, NPP provides an ideal method for the determination of solubilities of slightly soluble electroactive molecules through the determination of the solution concentration of material in equilibrium with excess undissolved solute.

Conditions providing diffusion-controlled well-formed NPP reduction waves for picloram have been reported (Whittaker and Osteryoung, 1979). We have verified diffusion control for the experimental conditions reported above according to the criterion that eq 1 is obeyed. The current decay within a pulse for 10^{-3} F picloram is shown in Figure 1. The experiment was carried out using the Taccusel pulse polarograph and potentiostat. A 5-s drop time was used to minimize the effect of change in area of the drop during the pulse time. With a 4.9-s delay before pulse application, the change in area during a 50-ms pulse is 0.7%, which is negligible. The least-squares linear regression analysis of log i vs. log t for the data of Figure 1 gave a slope of 0.503 (0.496 for 1×10^{-4} F picloram) which indicates diffusion control.

Standards 10^{-4} F in picloram gave sensitivities of 49, 42, 39, 41, and 48 μA/mM at nominal pH values of 1.0, 1.5, 2.0, 2.5, and 3.0, respectively. Solubilities calculated using these sensitivity data are given in Table I. Figure 2 illustrates the excellent agreement between these data and those of Cheung and Biggar and suggests the care with

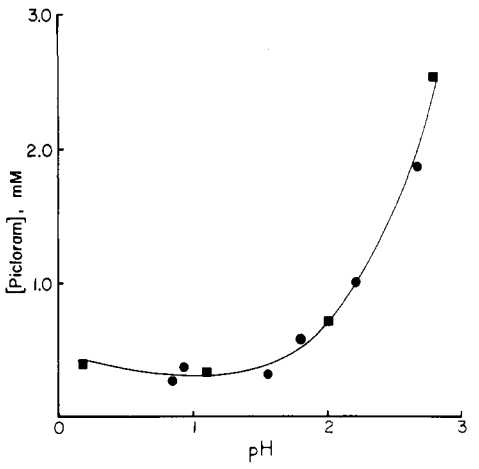

Figure 2. Solubility of picloram as a function of pH. Solution Na_2SO_4 + H_2SO_4 (μ = 0.3), 1% 2EE: 25.0 ± 0.02 °C. This work (●). Interpolated from data of Cheung and Biggar (1974) (■).

which their solubility study was carried out. Therefore their pK_a value, 3.4, should be reexamined. Casual examination of Figure 2 suggests that pK_a is about 2.

A more detailed analysis is based on the nomenclature of Cheung and Biggar for the convenience of the reader. The total solubility is given by

$$S = [HP] + [H^+P^-] + [HPH^+] + [P^-]$$

where HP is dissolved neutral picloram molecule, H^+P^- is the zwitterion, HPH^+ the conjugate acid of HP, and P^- the conjugate base. The equilibrium constants of interest are defined as follows:

$$HP = H^+ + P^- \qquad K_a$$

$$HPH^+ = HP + H^+ \qquad K_2$$

$$H^+P^- = H^+ + P^- \qquad K_3$$

In the pH range 0–2 the solubility is nearly pH independent and due to $[HP] + [H^+P^-]$. The suggestion of increasing solubility with decreasing pH is undoubtedly due to the increasing importance of HPH^+. Sargent and Blackmun (1970) point out that the methods of Clark and Perrin (1964) yield estimates of −2.0 and −1.3 for pK_a values for the base functionalities of picloram. Thus we estimate $K_2 = 20$ and at pH 0 $[HPH^+]$ is 5% of the total. At pH values above 2, the solubility increases sharply due to the increased fraction of P^-. Thus over the range of interest, the solubility is well approximated by

$$S = [HP] + [H^+P^-] + [P^-]$$

$$S = [HP](1 + K_a/K_3 + K_a/[H^+])$$

A plot of S vs. $1/[H^+]$ is a straight line with slope $K_a[HP]$ and intercept $[HP](1 + K_a/K_3)$. Experimental values obtained from least-squares analysis of the data of Table I are $(3.36 \pm 0.62) \times 10^{-3}$ and 0.313 ± 0.130 mM, respectively. The ratio of slope to intercept is

$$\frac{K_aK_3}{K_a + K_3} = (1.07 \pm 0.37) \times 10^{-2}\ \text{M}$$

We define this ratio as

$$K_a' = \frac{K_aK_3}{K_a + K_3} = \frac{[H^+][P^-]}{[HP] + [H^+P^-]}$$

the combined equilibrium constant for the dissociation of

both the neutral and zwitterion forms to form the conjugate base. Thus for picloram under the conditions of these experiments $K_a' = 1.07 \times 10^{-2}$ M; $pK_a' = 1.97$. It should be pointed out that this constant is a mixed constant whose value depends on activity effects. The hydrogen ion term is expressed in activity while $[P^-]$, [HP], and $[H^+P^-]$ are in concentration units.

The data and analysis presented here bear only peripherally on the question of the importance of the zwitterion. We suspect that the equilibrium between HP and H^+P^- strongly favors H^+P^-, i.e., $K_a \gg K_3$ and $K_a' \approx K_3$. Two lines of reasoning support this conclusion. First, the highly polar solvent and rather large salt concentrations should favor the zwitterion. This general observation has been substantiated for pyridine monocarboxylic acids (Green and Tong, 1956). Second, although Cheung and Biggar do not specify the ionic strength of their experiments, from their experimental description, the ionic strength must be rather low. Yet our solubility data are in excellent agreement, even in the high pH range. At an ionic strength of 0.3 M (our conditions) the activity coefficient of P^- is approximately 0.71, while in the low ionic strength of Cheung and Biggar, the activity coefficient is not less than about 0.95. Thus about 25% difference in solubility would be expected in the basic pH range where the P^- concentration term dominates. In addition, Cheung and Biggar state that the solubility in this range is unaffected by ionic strength up to 0.5 M KCl. We conclude that in fact the zwitterion rather than the neutral molecule is the dominant form and the activity effects on the zwitterion and the conjugate base are similar.

The solubility data of Cheung and Biggar are quite correct, but the many values previously reported for the acid dissociation constant of picloram are too small (pK_a' too large). These pK_a values range from 2.8 to ca. 4.1. [Hamaker has stated that the value of 4.1 is inappropriate because it was determined in 50% acetone (Hamaker, 1975)]. They have been determined from the pH at the half-equivalence point in a potentiometric titration of picloram with strong base. Such titrations are fraught with difficulty, for they must be done at low concentration to ensure that all the material is in solution. With respect to the value of 3.4 obtained by Cheung and Biggar, titrations were done using 1 mM solutions of picloram (Cheung, 1976) yet over a considerable portion of the pH range of interest the measured solubility is less than 1 mM. This could easily account for the high values obtained. Similar experimental difficulties could reasonably account for the other reported values which are larger than that inferred here from the solubility measurements. Apparently the potentiometric titration, which has great merit at higher concentrations, is an unreliable technique for pK_a determination under these circumstances. On the other hand, the inference of the pK_a value from solubility data appears to be quite straightforward. The slight uncertainty in interpretation of the pK_a, that is, the judgment of the relative importance of HP and H^+P^- under the experimental conditions, should not be allowed to obscure the simplicity and reliability of the determination of the phenomenological constant, K_a'. Furthermore, normal pulse polarography has proven to be a rapid, convenient, and error-free technique for the determination of solubilities of slightly soluble compounds.

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LITERATURE CITED

Cheung, M. W., Biggar, J. W., *J. Agric. Food Chem.* **22**, 202 (1974).
Cheung, M. W., private communication. CIBA-GIGY Corporation, Greensboro, NC, 1976.
Clark, J., Perrin, D. D., *Q. Rev.* **18**, 295 (1964).
Green, R. W., Tong, H. K., *J. Am. Chem. Soc.* **78**, 4896 (1956).
Hamaker, J. W., private communication. DOW Chemical USA, Walnut Creek, CA, 1975.
Osteryoung, J., Hasebe, K., *Rev. Polarogr. Japan*, **21**, 1 (1976).
Sargent, J. A., Blackman, G. E., 1970, *J. Exp. Bot.* **21**, 219 (1970).
Whittaker, J. W., Osteryoung, J., *J. Agric. Food Chem.*, in press (1979).

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