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THE DISSOCIATION OF EDTA AND EDTA-SODIUM SALTS

by

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In another paper the results of a study concerning the action of EDTA (ethylenediamine tetra-acetic acid), and of the disodium, trisodium and tetrasodium salts of EDTA on dental enamel surfaces are reported(1). The present study was undertaken in order to get a foundation for the preparation of the sodium salts. In order to prepare the salts it is necessary to know the dissociation constants and the dissociation curve of EDTA.

The EDTA has four carboxyl groups. Therefore, the dissociation takes place in four steps, each with its own dissociation constant. If we designate the EDTA as H_4Y, as proposed by *Schwarzenbach*(2), the dissociation takes place according to the following scheme,

$$H_4Y \rightleftarrows H^+ + H_3Y^-, \quad k_1 = \frac{[H^+]\cdot[H_3Y^-]}{[H_4Y]}$$

$$H_3Y^- \rightleftarrows H^+ + H_2Y^=, \quad k_2 = \frac{[H^+]\cdot[H_2Y^=]}{[H_3Y^-]}$$

$$H_2Y^= \rightleftarrows H^+ + HY^{\equiv}, \quad k_3 = \frac{[H^+]\cdot[HY^{\equiv}]}{[H_2Y^=]}$$

$$HY^{\equiv} \rightleftarrows H^+ + Y^{\equiv\equiv}, \quad k_4 = \frac{[H^+]\cdot[Y^{\equiv\equiv}]}{[HY^{\equiv}]}$$

Schwarzenbach & Ackermann(3) give the following values for the four dissociation constants: $pk_1 = 2.0$, $pk_2 = 2.6$, $pk_3 = 6.16$, and $pk_4 = 10.26$ (pk is the negative logarithm of the dissociation constant). Their determinations were carried out in solutions containing 0.1 N KCl. The concentration of EDTA was about 0.001 molar. In our experiments considerably higher concentrations were wanted, otherwise the exposure time would be too long. *Carini & Martell*(4) give the values $pk_3 = 6.320$ and $pk_4 = 11.014$. These are thermodynamic constants and only valid for zero ionic strength. When used for solutions of finite strength the activity of the ions must be known. Especially for the tervalent and the quadrivalent ions of the trisodium and tetrasodium salts respectively, the calculation of the activity would be too complicated. Therefore, it was decided to determine the stoichiometric constants for different concentrations. This was done by titrating EDTA solutions of different concentrations. One difficulty was encountered when preparing the solutions. The pure acid is only slightly soluble. The solubility is about 0.001 molar. The solubility of the sodium salts is considerably higher. In order to get higher concentrations the acid was dissolved in NaOH of different concentrations. In most cases a mixture of two sodium salts was obtained, with a preponderance of either disodium, trisodium or tetrasodium salt, depending on the strength of the NaOH solution. The solutions were then titrated with NaOH or HCl, in some instances with both. During the titration the pH was recorded, using a pH-meter from "Radiometer", Copenhagen, type 22, with a glass electrode type G 200 B. The acid used was Titriplex II from Merck. In this way the values of pk_3 and pk_4 could be determined fairly accurately. The value of pk_2 could not be determined directly, but in solutions containing a mixture of monosodium and disodium salts an approximate value for pk_2 could be calculated by means of the *Henderson—Hasselbalch* formula,

$$pH = pk_2 + \log \frac{[\text{disodium salt}]}{[\text{monosodium salt}]}.$$

No attempt was made to determine the first dissociation constant. Mostly saturated solutions were prepared. In some cases

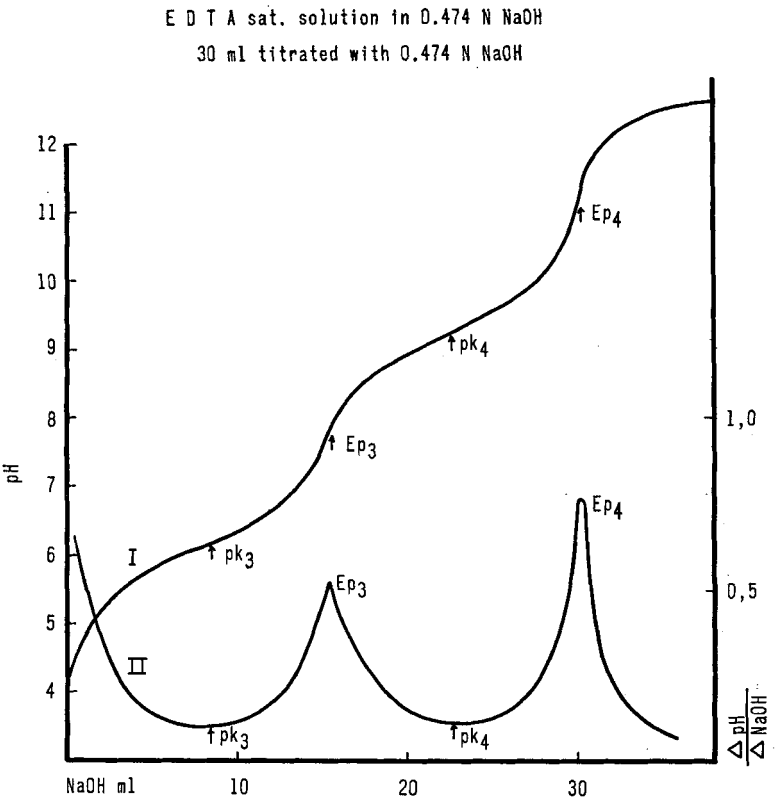

Fig. 1.

an almost pure solution of only one sodium salt was obtained. Thus, it was possible to determine approximately the solubility of the sodium salts of EDTA. At the equivalence points the solution contained almost exclusively one sodium salt. Therefore, the pH at the equivalence points corresponds to the pH of a solution of one single sodium salt.

RESULTS

Fig. 1 shows a typical titration curve, curve I. In this case the solution contained chiefly disodium salt and a little mono-

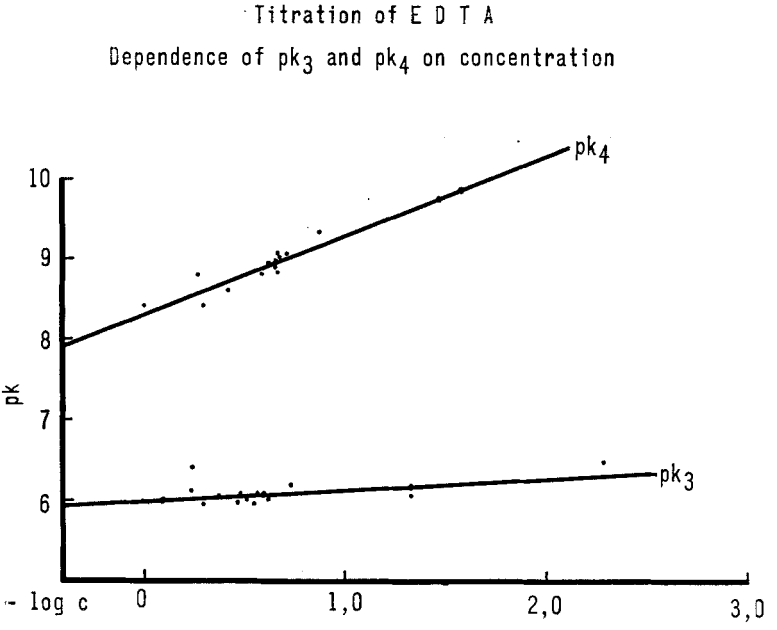

Fig. 2.

sodium salt. It is seen that the equivalence points are fairly well recognizable. The values of pk_3 and pk_4 cannot be read directly from the titration curve. In curve II are plotted the reciprocal values of the buffer capacity,

$$\frac{\triangle pH}{\triangle V}$$

i.e. the change in pH with each addition of NaOH divided by the added volume. In this curve the minima correspond to the values of pk_3 and pk_4. It is seen that these minima are fairly well defined. The equivalence points which correspond to the maxima can be determined with accuracy.

In Fig. 2 are shown the values obtained for pk_3 and pk_4. It is seen that both pk_3 and pk_4 vary proportionally with the negative logarithm of the concentration. (The concentrations given are the initial concentrations corrected for the increase in volume due to addition of titration fluid). The curve that will best fit the data

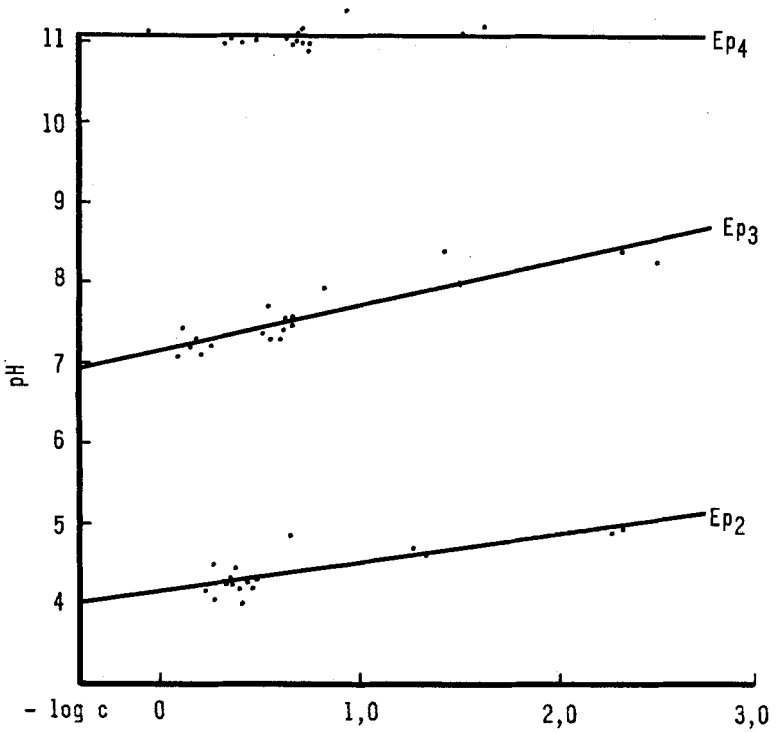

Fig. 3.

is a straight line. The equations for the two lines have been calculated by the method of least squares,

$$pk_3 = 6.0 + 0.13 \ (- \log c)$$
$$pk_4 = 8.31 + 0.98 \ (- \log c)$$

Thus, the values can either be taken directly from the diagram or be calculated from the equations.

Fig. 3 shows the results for the equivalence points. Also the values for disodium and trisodium salts vary proportionally with the negative logarithm of the concentration. The equivalence

point of the tetrasodium salt, however, was found to be independent of the concentration within the range studied.

The equations for the curves that fit the data are,

$Ep_2 = 4.15 + 0.356$ (— log c) for disodium salt
$Ep_3 = 7.15 + 0.56$ (— log c) for trisodium salt
$Ep_4 = 11.05$ for tetrasodium salt

As already mentioned, the EDTA (in the form of the free acid) is little soluble. When it is dissolved in NaOH with a large excess of OH^- ions, it is immediately converted to the tetrasodium salt, $H_4Y + 4NaOH \rightarrow Na_4Y + 4H_2O$. When more acid is added after all the NaOH has reacted, the acid combines with the tetra salt to form the trisodium salt,

$$3Na_4Y + H_4Y \rightarrow 4Na_3HY.$$

The trisodium salt is less soluble than the tetra salt. Therefore, if the NaOH is strong enough, there may be formed more trisodium salt than the amount which can be held in solution, and trisodium salt will be precipitated. If the NaOH solution is not strong enough, the concentration of trisodium salt will be too low for precipitation to occur. On further addition of EDTA the acid will react to form disodium salt,

$$2Na_3HY + H_4Y \rightarrow 3Na_2H_2Y.$$

The disodium salt is less soluble than the trisodium salt. Therefore, if the NaOH solution is not strong enough to make a supersaturated solution of trisodium salt, further addition of NaOH may cause precipitation of disodium salt or of monosodium salt, which is still less soluble. In many of the experiments it was, therefore, possible to determine the solubility of the sodium salts. The determinations were carried out at room temperature without any attempt to control the temperature. The values must, therefore, only be considered approximate. The following values were found:

For tetrasodium salt 1.7 molar, for trisodium salt 0.60 molar, for disodium salt 0.30 molar, and for monosodium salt 0.012 molar. The results are summarized in Table 1.

Table 1

Compound	Solubility moles/litre	pH in sat. sol.	Diss. const.	Equivalence points
EDTA	0.001	3.03	2.0[1])	—
Monosodium	0.012	—	2.6[1])	—
Disodium	0.30[2])	4.5	6.10+0.13(—log c)	4.15+0.356(—log c)
Trisodium	0.60	7.7	8.31+0.98(—log c)	7.15+0.56 (—log c)
Tetrasodium	1.7	11.05		11.05

[1]) The values are taken from *Schwarzenbach & Ackermann* (3).

[2]) For the solubility of disodium salt *Blaedel & Knight* (5) give the figure 11.1 g/100 ml solution at 21°C. This represents a concentration of almost exactly 0.3 molar.

DISCUSSION

A clear picture of the dissociation of EDTA may be obtained from Fig. 4.

When a solution of pure EDTA is titrated with NaOH, the composition of the solution will gradually change with the amount of NaOH added. At the start there will be only pure EDTA, then on adding NaOH, monosodium salt will be formed in increasing amount until one equivalent of NaOH is added. On further addition, NaOH reacts with the monosodium salt to form disodium salt, $NaH_3Y + NaOH \rightarrow Na_2H_2Y + H_2O$.

The concentration of monosodium salt will, therefore, decrease while disodium salt increases until two equivalents of NaOH are added. On further addition disodium salt will decrease and trisodium salt will be formed. Finally, only tetrasodium salt will be present. Each of the salts will thus go through a maximum, corresponding to its equivalence point. Unlike the combination of EDTA with Ca and other bivalent and polyvalent metals which are complex bound and not ionized, the sodium salts of EDTA are dissociated. Thus the EDTA is present as an anion. The charge of the anion depends on the pH of the solution. When the dissociation constants are known, the relative amounts of the different forms can easily be calculated.

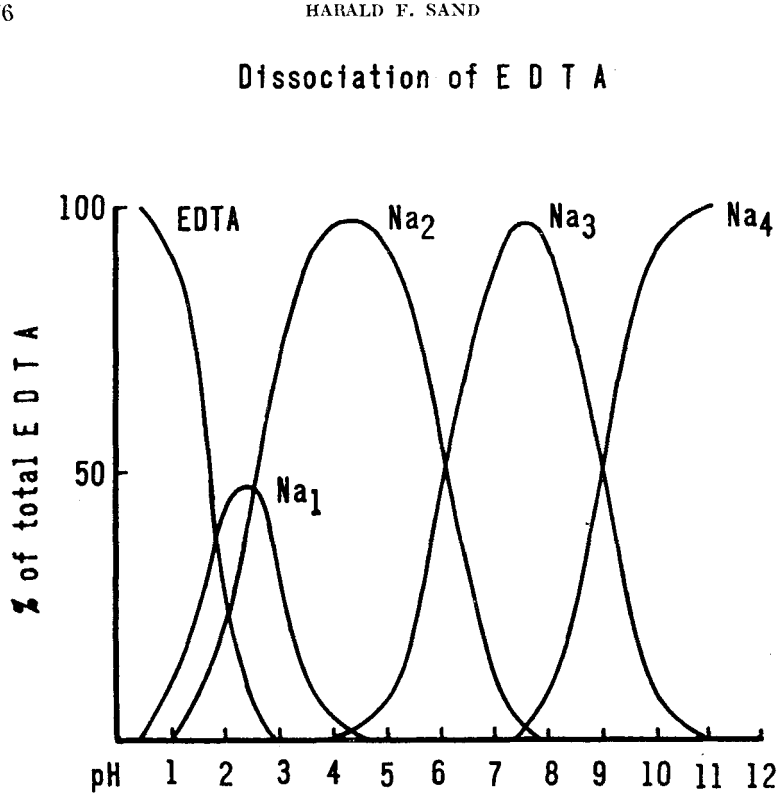

Fig. 4.

In a strongly acid solution, up to about pH 1.0, there is mostly free acid and a little monosodium salt present. Therefore, if c is the total amount of EDTA present we get,

$$c = H_4Y + H_3Y^-; \quad [H_3Y] = \frac{k_1 \cdot [H_4Y]}{[H]}$$

If we substitute and rearrange:

$$\frac{[H_4Y]}{c} = \frac{[H]}{[H] + k_1}$$

From this we can calculate the fraction of EDTA present as free acid. At a pH between 1.0 and 2.5 both free acid, monosodium salt and disodium salt will be present and we get,

$$c = H_4Y + H_3Y + H_2Y; \quad [H_4Y] = \frac{[H_3Y] \cdot [H]}{k_1};$$

$$[H_2Y] = \frac{[H_3Y] \cdot k_2}{[H]}; \quad \frac{[H_3Y]}{c} = \frac{[H] \cdot k_1}{[H]\ k_1 + [H]^2 + k_1\ k_2}.$$

In the same manner the amount of di-, tri- and tetrasodium salts can be calculated. In Fig. 4 the calculation is based on the following values for the dissociation constants,

$$pk_1 = 2.0, \ pk_2 = 2.60, \ pk_3 = 6.15, \ pk_4 = 9.0.$$

As expected, the pk_3 and pk_4 values were found to vary with the concentration. Consequently, also the equivalence points of the salts should vary. In the equivalence points the pH of the solution is the same as in a solution of the pure salt. The pH of trisodium salt, for example, can be approximately calculated from the formula,

$pH = \frac{1}{2} \ (pk_3 + pk_4)$.
For a 0.1 molar solution we get,
$pH = \frac{1}{2} \ (6.13 + 9.29) = 7.71$.
From the formula for the equivalence points we get,
$Ep_3 = 7.15 + 0.56 \ (- \log c)$.
For a 0.1 molar solution we then have,
$Ep_3 = 7.15 + 0.56 = 7.71$.
Thus, there is complete agreement.

It was quite unexpected to find that the equivalence point of the tetra salt was independent of the concentration. The pH of the tetrasodium solution can be approximately calculated from the formula,

$pH = \frac{1}{2} \ (pk_w + pk_4 + \log c)$, ($pk_w$ is the negative logarithm of the dissociation constant for water). Since the value of pk_4 may vary from about 8.3 to about 11.0, one would expect that the pH of a tetrasodium solution should vary. As explained earlier it was found that the pk_4 in a solution of a certain concentration can be calculated from the formula,

$pk_4 = 8.31 + 0.98 \ (- \log c)$. If this value is substituted in the equation for pH above we have,

$$pH = \tfrac{1}{2}\ (pk_w + 8.31 - 0.98 \log c + \log c)$$
$$pH = \tfrac{1}{2}\ (14 + 8.31 + 0.02 \log c)$$
$$pH = 11.16 + 0.01 \log c.$$

From this we see, that a change in concentration from 1.0 to 0.01 molar will only change the pH by 0.02 units, which is about the error of measurement.

It is seen that the agreement between the calculated value and the value found in the experiments is not so good as was the case with trisodium salt. This might be expected since the formula for the pH of the tetrasodium salt is less accurate.

Fig. 2 shows that pk_4 varies considerably more than pk_3. This might be expected, since the function of ions are related to the ionic strength rather than to the concentration. The ionic strength is defined by the formula,

$$\tfrac{1}{2} \Sigma\ (c_1 \cdot z_1^2 + c_2 \cdot z_2^2 \ldots\ldots\ldots\ldots\ldots\ c_n \cdot z_n^2)$$

Since the ionic strength varies with the charge of an ion to the second power, it is clear that the ionic strength of the tetrasodium solution increases more with increasing concentration than does the ionic strength of trisodium salt.

If Table 1 is compared with the dissociation curve, Fig. 4, it may at first sight seem surprising that a solution of the free acid has a pH as high as 3.03. But it must be remembered that a saturated solution of EDTA has a concentration of only 0.001 molar. In the first step the acid splits off one H^+ ion. If the dissociation were complete, one mole of EDTA would give one mole of H^+ ions. In a 0.001 molar solution of EDTA the concentration of H ions cannot exceed 0.001 molar, i.e. the pH cannot be lower than 3.0. It must also be remembered that in a solution of 0.001 normal HCl the pH is above 3. Thus, the lower part of the dissociation curve for EDTA (the free acid) can only be obtained by adding a very strong acid (for example 1.0 N HCl) to the solution.

It is also seen from Fig. 4 that the maximum concentration of the monosodium salt is only about 47 % of the total amount of EDTA in the solution. This means that it will be impossible to prepare a solution which contains only the ions of monosodium salt. Even if a pure solution of monosodium salt were dissolved

in water, it would form also H_2Y ions and undissociated free acid. This is explained by the fact that there is very little difference between the first dissociation constant (pk_1 2.0) and the second one (pk_2 2.60). But even the disodium and the trisodium salts do not quite reach a concentration of 100 % of the total.

Wandelt(6) gives the pH-values of the different sodium salts of EDTA. For the free acid as well as for monosodium and disodium salts he gives the value pH 4.0. The pH of a solution of monosodium salt can be calculated approximately from the formula,

$$pH = \tfrac{1}{2} \ (pk_1 + pk_2)$$

and for disodium salt,

$$pH = \tfrac{1}{2} \ (pk_2 + pk_3).$$

It is seen from these equations that if the pH of the monosodium and the disodium salts should be the same, then the pk_1 and pk_3 must have the same values. That is, of course, impossible.

Wandelt(6) wants to study the action of different EDTA-sodium salts. But he makes a mistake when he adjusts all solutions to the same pH (6.9). Because, as seen from Fig. 4, in doing so he ends up in all cases with a solution of practically the same composition, i.e. about 90 % trisodium salt and 10 % disodium salt. Thus, when he finds a difference in the action of the different solutions, this must be due to a difference in the concentration and not to a difference in composition.

SUMMARY

By titrating EDTA solutions of varying concentrations, information has ben obtained concerning the third and the fourth dissociation constant, the pH in solutions of EDTA-disodium, trisodium and tetrasodium salts, and the solubility of the sodium salt of EDTA. The following values have been found,

$$pk_3 = 6.10 + 0.13 \ (- \log c), \quad pk_4 = 8.31 + 0.98 \ (- \log c).$$

Disodium salt pH $= 4.15 + 0.356 \ (- \log c)$
Trisodium salt pH $= 7.15 + 0.56 \ (- \log c)$
Tetrasodium salt pH $= 11.05$,

and for the solubilities,

EDTA	0.001	moles/litre
Monosodium salt	0.012	"
Disodium salt	0.30	"
Trisodium salt	0.60	"
Tetrasodium salt	1.7	"

RÉSUMÉ

LA DISSOCIATION DE L'EDTA (ETHYLÈNE DIAMINE-ACIDE TÉTRA-ACETIQUE) ET DES SELS SODIQUES DE L'EDTA

Par le titrage de solutions d'EDTA de différentes concentrations, l'auteur a obtenu des renseignements sur la troisième et la quatrième constante de dissociation, le pH dans les solutions des sels *disodique*, *trisodique* et *tétrasodique* d'EDTA, et la solubilité du sel sodique d'EDTA. Les valeurs suivantes ont été trouvées:

$$pk_3 = 6{,}10 + 0{,}13 \ (- \log c), \quad pk_4 = 8{,}31 + 0{,}98 \ (- \log c).$$

pH du sel disodique $= 4{,}15 + 0{,}356 \ (- \log c)$
pH du sel trisodique $= 7{,}15 + 0{,}56 \ (- \log c)$
pH du sel tétrasodique $= 11{,}05$,

et pour les solubilités:

EDTA	0,001	moles p. litre
Sel monosodique	0,012	"
Sel disodique	0,30	"
Sel trisodique	0,60	"
Sel tétrasodique	0,7	"

ZUSAMMENFASSUNG

DIE DISSOZIATION VON EDTA UND EDTA-NATRIUMSALZ

Durch Titration von EDTA-Lösung verschiedener Konzentration wurden Informationen erhalten über die 3. und 4. Dissociationskonstante, den pH-Wert des Di-Natrium-, Tri-Natrium- und

Tetra-Natriumsalzes des EDTA und ihrer Löslichkeit. Die folgenden Werte wurden dabei gefunden:

$pk_3 = 6.10 + 0.13$ (— log c), $pk_4 = 8.31 + 0.98$ (— log c).

Di-Natriumsalz pH	$= 4.15 + 0.356$ (— log c)
Tri-Natriumsalz pH	$= 7.15 + 0.56$ (— log c)
Tetra-Natriumsalz pH	$= 11.05$

und für die Löslichkeit:

EDTA	0.001	mol/Liter
Mono-Natriumsalz	0.012	"
Di-Natriumsalz	0.30	"
Tri-Natriumsalz	0.60	"
Tetra-Natriumsalz	1.7	"

RESUMEN

LA DISOCIACIÓN DEL EDTA Y DE LAS SALES SÓDICAS DE EDTA

Titulando soluciones de EDTA en concentraciones variables se ha obtenido información acerca de la constante de disociación de tercero y cuarto orden, del pH en soluciones de sales de EDTA disódica, trisódica y tetrasódica, y la solubilidad de la sal sódica de EDTA. Se encontraron los siguientes valores:

$pk_3 = 6.10 + 0.13$ (— log c), $pk_4 = 8.31 + 0.98$ (— log c).

pH de la sal disódica	$= 4.15 + 0.356$ (— log c)
pH de la sal trisódica	$= 7.15 + 0.56$ (log c)
pH de la sal tetrasódica	$= 11.05$.

Respecto de las solubilidades se obtubieron los siguientes resultados:

EDTA	0.001	moles/litro
Sal monosódica	0.012	"
" disódica	0.30	"
" trisódica	0.60	"
" tetrasódica	1.7	"

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