

On the Association Constant of Calcium Oxalate

M V R RAO*, J S AGARWAL & NAYANTARA CHHOTRAY

Department of Chemistry, University of Delhi, Delhi 110007

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Calcium and oxalate ions form strong ion-pairs in dilute solutions ($\sim 10^{-4} M$). The thermodynamic association constant has been evaluated to be 2463 ± 220 at $37^\circ C$, using an ion-selective electrode to determine free calcium ion concentration. The ionic product (K_{ip}) at the inception of precipitation has also been evaluated and this is found to be a function of ionic strength as well as concentration of other complexing agents of oxalate, like magnesium ions.

Calcium oxalate occurs in two phases—as a monohydrate (whewellite) and as a dihydrate (weddelite)^{1a,b} in kidney and bladder stones. As part of our physicochemical studies on the genesis of the 'stones', we have presently investigated spontaneous formation of calcium oxalate crystallites in aqueous systems using an ion-selective electrode² (ISE) to follow changes in concentration of calcium ions.

A calcium ion-selective electrode (Orion, model 93-20) was used in conjunction with reference electrode (Orion, model 91-01). A Radiometer (PHM-26c) pH-meter was used to determine the e.m.f. of the cell.

The Nernstian response of the ISE was checked everyday and a calibration curve was obtained using a $0.1 M$ calcium ion standard solution (Orion 92-20-06) which was serially diluted to $0.0001 M$; the method was cross checked with an EDTA titration procedure using the ISE as an end point detector. The two sets of data agreed within $\pm 1\%$.

Calcium chloride solutions, prepared from AR grade sample, were taken in a double walled vessel through which water at $37.0 \pm 0.1^\circ C$ was circulated from a thermostat; the vessel had a close fitting perspex lid with suitable openings. A microburette was used to add a solution of sodium oxalate (AR). The solutions were continuously stirred with a magnetic stirrer. Precautions were taken to prevent dust particles from entering the systems; the solutions were filtered through millipore filters (Maxflow) before use.

In a typical 'titration', when sodium oxalate ($4.00 \times 10^{-3} M$) is added to $20 ml$ of calcium chloride solution ($3.45 \times 10^{-4} M$), initially there is a gradual but slow increase in e.m.f., till about $1.80 ml$ of the oxalate solution has been added; this is followed by a sharp increase in the e.m.f. The onset of second region coincides with the appearance of turbidity in the

solution. The curve ultimately has a near plateau region at low concentration of calcium ions. The solid phase that separated out has been identified as as monohydrate of calcium oxalate from X-ray powder diffraction pattern.

The initial part of the 'titration curve is of interest for the present discussion. On the addition of oxalate solution, the effective calcium ion concentration, after taking into account the dilution factor, decreases even before spontaneous precipitation sets in (from $3.45 \times 10^{-4} M$ in the original solution to $2.12 \times 10^{-4} M$ when $1.80 ml$ of sodium oxalate has been added). The fact that there is no formation of any precipitate before the visual observation of turbidity has also been verified by the determination of absorbance of the solution at $400 nm$ using a GILFORD UV-visible spectrophotometer. The absorbance, remains more or less unchanged till turbidity appears. The observed decrease in $[Ca^{2+}]$ on the addition of oxalate suggests that soluble calcium oxalate species are formed. The association constant, $\{[CaC_2O_4]/[Ca^{2+}][C_2O_4^{2-}]\}$, of calcium oxalate has been evaluated; the concentration of the ion pair has been calculated from the decrease in $[Ca^{2+}]$ on addition of oxalate. The ionic strength of the medium due to the presence of Ca^{2+} , Cl^- , Na^+ and $C_2O_4^{2-}$ ions, has been taken into account in deducing the mean activity coefficient ($\gamma_{\pm}$) from the extended form of the Debye-Hückel equation given by Davies³:

$$-\log \gamma_{\pm} = Az^2[I^{1/2}/(1 + I^{1/2}) - 0.3I]$$

where A is a constant, I is ionic strength and z is charge of the ion.

The equilibrium constant, K'_{asso} , and the thermodynamic equilibrium constant K_{asso} ($= K'_{asso} \times 1/\gamma_{\pm}^2$) have also been calculated; the activity coefficient of neutral calcium oxalate species has been assumed to be unity in the rather very dilute solutions. A typical set of data are given below; the deviation in the association constant has been calculated on the basis of data of four independent experiments.

$$[Ca^{2+}]_{initial} = 3.17 \times 10^{-4} M$$

$$[Ca^{2+}]_{free(IST)} = 2.12 \times 10^{-4} M$$

$$[Ca^{2+}]_{asso} \equiv [CaC_2O_4] = 1.05 \times 10^{-4} M$$

$$\left(\equiv [Ca^{2+}]_{initial} - [Ca^{2+}]_{free} \right)$$

$$[C_2O_4^{2-}]_{free} = 2.25 \times 10^{-4} M$$

$$\left(\equiv [\text{C}_2\text{O}_4^{2-}]_{\text{total}} - [\text{Ca}^{2+}]_{\text{asso}} \right)$$

$$K'_{\text{asso}} = 1754 \pm 159 \text{ M}^{-1}$$

$$I = 0.0015$$

$$\gamma_{\pm} = 0.8358$$

$$K_{\text{asso}} = 2463 \pm 220$$

It is interesting that the sparingly soluble calcium oxalate forms readily ion-pairs even in very dilute solutions and that the association constant is as high as 2463 at 37°C. Tomazic and Nancollas⁴ estimated the stability constant for this ion-pair to be 1869 at 37°C, using conductometric data on the kinetics of dissolution of whewellite. The present method, however, is based on a more direct technique of estimating free calcium ion concentration with an ISE.

The 'titration' experiment can also be used to determine the 'formation product', K'_{fp} , which is the ionic product of $[\text{Ca}^{2+}]$ and $[\text{C}_2\text{O}_4^{2-}]$ at the inception of spontaneous precipitation from a supersaturated solution⁵. The ionic product (K'_{fp}), has been calculated to be $5.30 \times 10^{-8} \text{ M}^2$ at 37°C, and this value is about

thirty-five times larger than the solubility product ($1.52 \times 10^{-9} \text{ M}^2$) at 37°C; the latter has also been determined using the calcium ion-selective electrode. This implies that calcium oxalate solution is supersaturated by a factor of 5.9 ($= [K'_{\text{fp}}/K_{\text{sp}}]^{1/2}$) before spontaneous precipitation of calcium oxalate monohydrate occurs under the parameters adopted in the present work. The formation product is also found to be a function of ionic strength as well as concentration of other complexing agents of oxalate, like magnesium ions etc. These data are of considerable relevance to discussion on the extent of supersaturation of urinary system which contains a host of ions, such as pyrophosphate and citrate, which readily form complexes with calcium ions.

References

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