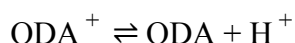


Electronic Supplementary Information

The detailed calculation processes of the ODA^+ and ODA^{2+} concentrations via the potentiometric titration:

For the ODA solutions in the presence of different H^+ concentrations, there were two ionization equations:



Because there was no precipitate in all solutions ($r = 1:1.5, 1:2$ and $1:4$, $[ODA] = 250$ mg/L), the first step should be dominant, whose equilibrium constant should be:

$$K_a = \frac{[ODA^+][H^+]}{[ODA^{2+}]}$$

(1)

The potentiometric titration was employed to determine the equilibrium constant of ODA. When the NaOH solution (its volume and concentration are V_0 and C_0 , respectively) was added to the ODA^{n+} solution (its volume and concentration are V_1 and C_1 , respectively), the ODA^+ and ODA^{2+} concentrations could be obtained by the following equation:

$$[ODA^+] = c_1 \frac{V_1}{V_1 + V_0}$$

(2)

$$[ODA^{2+}] = c_0 \frac{V_0}{V_0 + V_1} - c_1 \frac{V_1}{V_1 + V_0}$$

(3)

The solution pH was measured to calculate the H^+ concentration:

$$[H^+] = 10^{-pH}$$

(4)

The obtained $[ODA^+]$, $[ODA^{2+}]$, and $[H^+]$ were substitute into equation (1),

$$K_a = 2.84464623 \times 10^{-5}$$

Based on the equation (1),

$$[ODA^+] = K_a \frac{[ODA^{2+}]}{[H^+]} \quad (5)$$

And another equation,

$$[ODA^+] + [ODA^{2+}] = 1.2485 \text{ mmol/L (250 mg/L)} \quad (6)$$

Finally, we could calculate the $[ODA^+]$ and $[ODA^{2+}]$ at different pH values (Table S1).

Table S1 The ODA^+ and ODA^{2+} concentrations at different ODA:HCl molar ratios

ODA:HCl molar ratio	ODA^+ concentration (mmol/L)	ODA^{2+} concentration (mmol/L)
1:1.5	0.8655	0.3830
1:4	1.398×10^{-2}	1.2345