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:

$$ODA^{2+} \rightleftharpoons ODA^{+} + H^{+}$$

$$ODA^{+} \rightleftharpoons ODA + H^{+}$$

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:

$$Ka = \frac{[ODA^+][H^+]}{[ODA^{2+}]} \tag{1}$$

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:

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

$$[ODA^{2+}] = c_0\frac{V_0}{V_0+V_1} - c_1\frac{V_1}{V_1+V_0} \tag{3}$$

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

$$[H^+] = 10^{-pH} \tag{4}$$

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

$$Ka = 2.84464623\times10^{-5}$$

Based on the equation (1),

$$[ODA^+] = Ka\frac{[ODA^{2+}]}{[H^+]} \tag{5}$$

And another equation,

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

Finally, we could calculated 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