In Class Problems
Precipitation–Dissolution Equilibrium & Kinetics
CENG 340–Introduction to Environmental Engineering
Instructor: Deborah Sills
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1. *Modified from Mihelcic and Zimmerman*
Ingesting cadmium may lead to kidney damage. According to the EPA, major sources of cadmium in drinking water include corrosion of galvanized pipes; erosion of natural deposits; discharge from metal refineries; and runoff from waste batteries and paints.

One method to remove metals, such as cadmium, from water is to raise the pH and cause them to precipitate as their metal hydroxides. The precipitation–dissolution equilibrium relationship for cadmium hydroxide is described by the following equation:

$$H_2O \underset{}{\overset{k_w}{\rightleftharpoons}} [H^+] + [OH^-]$$

$$[Cd(OH)_2] \xrightleftharpoons{K_{sp}} [Cd^{2+}] + 2\,[OH^-]$$

$$k_w = \frac{[H^+][OH^-]}{[H_2O]} = 10^{-14}$$

1 pure liquid

where $pK_{sp} = 13.85$

$K_{sp} = 10^{-13.85}$

$$K_{sp} = \frac{[Cd^{2+}][OH^-]^2}{[Cd(OH)_2]}$$

solid = 1

(a) In an attempt to remove cadmium from water by precipitating cadmium hydroxide, the pH of water was raised from 6.8 to 8.0. Was the dissolved cadmium concentration reduced to below 100 mg/L at the final pH?

pH = 8

$[H^+] = 10^{-8}$

$[OH^-] = \frac{10^{-14}}{10^{-8}} = 10^{-6}$

$MW_{Cd} = 112\, g/mole$

$$[Cd^{2+}] = \frac{K_{sp}}{[OH^-]^2} = \frac{10^{-13.85}}{(10^{-6})^2} = 1.4\times10^{-2}\ mole/L$$

$$[Cd^{2+}] = 1.4\times10^{-2}\ mole/L \times 112\, g/mole = 1.6\ g/L >> 100\, mg/L$$

$\boxed{\text{Dissolved Cadmium was not reduced to below 100 mg/L at pH = 8}}$

(b) What pH is required to reduce cadmium concentration to below the MCL of 5 ppb?

pH = ? If I find $[OH^-]$, I can find pH.

$$[Cd^{2+}] = 5\ ppb \times \frac{1\ ppm}{10^3\ ppb} = 5\times10^{-3}\ ppm = 5\times10^{-3}\ mg/L$$

$$[Cd^{2+}] = 5\times10^{-3}\ mg/L \times \frac{1\ mole}{112\ g} \times \frac{1\ g}{1000\ mg} = 4.46\times10^{-8}\ mole/L$$

$$[OH^-] = \sqrt{\frac{K_{sp}}{[Cd^{2+}]}} = \sqrt{\frac{10^{-13.85}}{4.46\times10^{-8}}} = 5.54\times10^{-4}$$

1

$$[H^+] = \frac{10^{-14}}{[OH^-]} = \frac{10^{-14}}{5.54\times10^{-4}} = 1.8\times10^{-11}\ mole/L$$

$$pH = -\log[H^+] = -\log(1.8\times10^{-11})$$

$\boxed{pH = 10.74}$