Marks 6

- The pH of the ocean before the industrial revolution was around 8.22. Show that this pH corresponds to a concentration of $[OH^-(aq)] = 1.7 \times 10^{-6}$ M.

By definition, $pH = -\log_{10}[H^+(aq)]$. Hence, a pH of 8.22 corresponds to:

$$[H^+] = 10^{-pH}\ M = 10^{-8.22}\ M$$

$[H^+(aq)]$ and $[OH^-(aq)]$ are linked through K_w. Assuming a temperature of 298 K,

$$K_w = [H^+(aq)][OH^-(aq)] = 1.00 \times 10^{-14}$$

Hence:

$$[OH^-] = K_w / [H^+(aq)] = 1.00 \times 10^{-14} / 10^{-8.22}\ M = 1.7 \times 10^{-6}\ M$$

All forms of life depend on iron and the concentration of iron in the oceans and elsewhere is one of the primary factors limiting the growth rates of the most basic life forms. One reason for the low availability of iron(III) is the insolubility of the hydroxide, $Fe(OH)_3$, which has a K_{sp} of only 1×10^{-39}. What was the maximum concentration of $Fe^{3+}(aq)$ at a pH of 8.22?

$Fe(OH)_3(s)$ dissolves according to the chemical equation:

$$Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$$

The solubility product is therefore given by:

$$K_{sp} = [Fe^{3+}(aq)][OH^-(aq)]^3$$

As $[OH^-(aq)] = 1.7 \times 10^{-6}$ M:

$$[Fe^{3+}(aq)] = K_{sp} / [OH^-(aq)]^3 = 1 \times 10^{-39} / (1.7 \times 10^{-6})^3\ M = 2 \times 10^{-22}\ M$$

Answer: **2×10^{-22} M**

Industrialisation has led to an increase in atmospheric CO_2. What effect has this had on the amount of $Fe^{3+}(aq)$ in sea water?

CO_2 dissolves in water to give acidic solution that reacts with OH^- ions.

$$2OH^-(aq) + CO_2(aq) \rightarrow CO_3^{2-}(aq) + H_2O$$

Increased amounts of CO_2 is thus likely to decrease the amount of OH^- in the sea water. From Le Chatelier's principle, the decrease in $[OH^-]$ will result in the equilibrium for the dissolution of $Fe(OH)_3$ above shifting to the right. This will lead to an increase in $[Fe^{3+}(aq)]$.