- Calculate the pH of a 0.10 mol L^{-1} solution of HF. (The pK_a of HF is 3.17.)

Marks 6

HF is a weak acid so the equilibrium concentrations need to be calculated using a reaction table:

	HF	$\rightleftharpoons$	$H^+(aq)$	$F^-(aq)$
Initial	**0.10**		**0**	**0**
Change	**-*x***		**+*x***	**+*x***
Equilibrium	**0.10 – *x***		***x***	***x***

As pK_a = -log$_{10}K_a$, at equilibrium,

$$K_a = \frac{[H^+(aq)][F^-(aq)]}{[HF(aq)]} = \frac{(x)(x)}{(0.10-x)} = \frac{x^2}{(0.10-x)} = 10^{-3.17}$$

As K_a is so small, x will be tiny and 0.10 – x ~ 0.10 and so

$x^2 = 10^{-3.17} \times 0.10$ **or** $x = [H^+(aq)] = 0.00822$ **M**

As pH = -log$_{10}$[H^+(aq)],

pH = -log$_{10}$(0.00822) = 2.09

Answer: **2.09**

What mass of NaF needs to be added to 100.0 mL of the above solution to make a buffer with a pH of 3.00?

Using the Henderson-Hasselbalch equation for the pH of the buffer:

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$

To make a buffer with pH = 3.00 and [acid] = [HF] = 0.10 M:

$$3.00 = 3.17 + \log\left(\frac{[F^-]}{0.10}\right) \text{ or } \log\left(\frac{[F^-]}{0.10}\right) = =0.17$$

Hence,

$$[F^-] = 0.10 \times 10^{-0.17} = 0.068 \text{ M}$$

The number of moles in 100.0 mL is thus 0.0068 mol. As NaF will dissolve to give one F^- per formula unit, this is also the number of moles of NaF required.

ANSWER CONTINUES ON THE NEXT PAGE

The formula mass of NaF is (22.99 (Na) + 19.00 (F)) g mol^{-1} = 41.99 g mol^{-1}. The mass of NaF required is thus:

mass = number of moles × formula mass
= (0.0068 mol) × (41.99 g mol^{-1}) = 0.28 g

Answer: **0.28 g**

Explain why HCl is a much stronger acid than HF.

Cl is a much larger atom than F and is less electronegative. The H–Cl bond is therefore much longer and weaker than the H–F bond. The H–Cl bond is therefore easier to break and it is the stronger acid.

HF is actually a weak acid. F is smaller and more electronegative than O, so the H–F bond is stronger than the O–H bond. There is consequently little dissociation of HF when it is dissolved in water.