Answer on Question #79395 Chemistry/ General Chemistry

Calculate the concentrations of all species present (H3O+, F−, HF, Cl−, and OH−) in a solution that contains 0.17 M HF (Ka=3.5×10−4) and 0.17 M HCl.
My answers are:0.17,0.17,0.17,0.17,5.89*10^-14 but it says the're wrong but i dont understand how.

Solution

HCl is a strong acid. It dissociates completely:

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

As mole ratio of $n(HCl):n(H^+):n(Cl^-)$ = 1:1:1 then initial concentrations of Cl^- and H^+ are $c(Cl^-)$=0.17M, $c(H^+)$ = 0.17M

HF is a weak acid. It does not dissociate completely:

$HF \leftrightarrow H^+ + F^-$

We should use ICE table to find equilibrium concentrations of all species:

	HF	H^+	F^-
Initial	0.17 M	0	0
Add	0	0.17 M	0
Change	-x	+x	+x
Equilibrium	0.17 – x	0.17 +x	x

$$K_a = \frac{(0.17 + x)x}{(0.17 - x)}$$

We make an assumption that x is very small (we'll check this assumption in the end of calculations- 5% rule), then

$0.17 + x \cong 0.17$

$0.17 - x \cong 0.17$

$$K_a = \frac{0.17x}{0.17}$$

$$K_a = x$$

$$x = 3.5 \times 10^{-4}$$

Select the smallest concentration for the 5% rule.

$$\frac{3.5 \times 10^{-4}}{0.17} \times 100\% = 0.2\%$$

This value is much less than 5%, so the assumptions are valid.

The concentrations at equilibrium are:

[HF] = 0.17-x = 0.17 – 3.5×10^{-4} = 0.1697 M

[H^+] = 0.17 + x = 0.17 + 3.5×10^{-4} = 0.1704 M

[F^-] = x=3.5×10^{-4} M

As Cl^- do not take part at any reaction its equilibrium concentration is the same as its initial concentration:

[Cl^-] = 0.17 M

Find equilibrium concentration of OH^-:

pH = -log[H^+]

pH = -log (0.1704)

pH = 0.7687

pOH = 14-pH

pOH = 14-0.7687

pOH=13.23

pOH= -log [OH^-]

13.23 = -log [OH^-]

[OH^-] = $10^{-13.23}$

[OH^-]=5.87×10^{-14} M

Answer:

[H_3O^+]	**[F^-]**	**[HF]**	**[Cl^-]**	**[OH^-]**
0.1704 M	**3.5×10^{-4}M**	**0.1697 M**	**0.17**	**5.87×10^{-14} M**