Because HF is a weak acid and HCl is a strong acid, the major species in solution are HF, H+ , and Cl–. The Cl–, which is the conjugate base of a strong acid, is merely a spectator ion in any acid–base chemistry.

	$HF_{(aq)}$ →	$H^+_{(aq)}$ +	$F^-_{(aq)}$
Initial	0.17M	0.17M	…
Change	-xM	+xM	+xM
Equilibrium	(0.17-x)M	(0.17+x)M	xM

$K_a = 3.5\cdot10^{-4} = \frac{[H^+][F^-]}{[HF]} = \frac{(0.17+x)*x}{0.17-x}$

If we assume that x is small relative to 0.17 M, this expression simplifies. $\frac{0.17x}{0.17}$ $= 3.5\cdot10^{-4}$

$x = 3.5\cdot10^{-4}$ M = $[F^-]$

This F– concentration is substantially smaller than it would be in a 0.17 M solution of HF with no added HCl. The common ion, H+ , suppresses the ionization of HF. The concentration of H+(aq) is

$[H_3O^+] = 0.17 + 3.5\cdot10^{-4} = 0.17035$ M

$[HF] = 0.17 - 3.5\cdot10^{-4} = 0.16965$ M

pH = 0.77

pOH = 14 – pH = 13.23

$[OH^-] = 5.89\cdot10^{-14}$ M

[H+] is due entirely to the HCl; the HF makes a negligible contribution by comparison.

$[Cl^-] = 0.17035$ M