

1. Refer to the data you obtained earlier (Acids, Bases and pH). Compare the pH and $[H^+]$ of H_2SO_4 with that of HCl and HNO_3 . How do you account for any differences?

Solution	pH	Equilibrium $[H^+]$ or $[OH^-]$
0.100 M HCl	1.00	$[H^+] = 1.00 \times 10^{-1} M$
0.100 M HNO_3	1.00	$[H^+] = 1.00 \times 10^{-1} M$
0.100 M H_2SO_4	0.88	$[H^+] = 0.131 M$

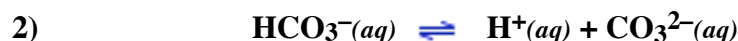
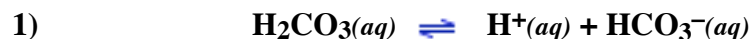
The difference in pH for the H_2SO_4 solution is due to the ability of H_2SO_4 to dissociate two protons.

2. H_2SO_4 is a polyprotic acid. What does this term mean? Write the dissociation equations which describe the stepwise behavior of H_2SO_4 .



3. Carbonic acid, H_2CO_3 , is a diprotic acid.

a) Write the two dissociation reactions showing the diprotic behavior.



b) If the initial concentration of H_2CO_3 is 0.100 M, calculate $[H^+]$. In your calculation, assume only the first dissociation occurs. (Note: The equilibrium constant for the first dissociation, K_{a1} , is 4.3×10^{-7} .)

	$H_2CO_3(aq)$	\rightleftharpoons	$H^+(aq)$	$+$	$HCO_3^-(aq)$	
initial	.1		1×10^{-7}		0	
change	-x		+x		+x	$x = [H_2CO_3]_{diss}$
equilibrium	.1 - x		$1 \times 10^{-7} + x$		0+x	

$$K_a = \frac{[H^+][HCO_3^-]}{[H_2CO_3]}$$

$$4.3 \times 10^{-7} = \frac{x^2}{.1 - x}$$

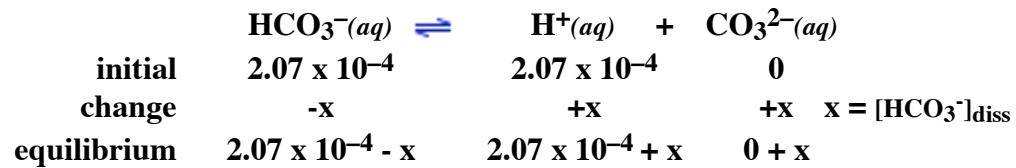
assume $x \ll .1$ and $1 \times 10^{-7} + x = x$

$$4.3 \times 10^{-8} = x^2$$

$$2.07 \times 10^{-4} = x = [\text{H}^+]$$

c) Now consider the second dissociation equation for which $K_{a2} = 5.6 \times 10^{-11}$. What is the initial concentration of $[\text{HCO}_3^-]$? What is the initial concentration of $[\text{H}^+]$? Calculate the final $[\text{H}^+]$ assuming the second dissociation occurs.

$$[\text{HCO}_3^-]_0 = [\text{H}^+]_0 = 2.07 \times 10^{-4} \text{ M}$$



$$K_a = \frac{[\text{H}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$$

$$5.6 \times 10^{-11} = \frac{(2.07 \times 10^{-4} + x)(x)}{2.07 \times 10^{-4} - x}$$

assume $x \ll 2.07 \times 10^{-4}$

$$5.6 \times 10^{-11} = \frac{(2.07 \times 10^{-4})(x)}{2.07 \times 10^{-4}}$$

$$5.6 \times 10^{-11} \text{ M} = x = [\text{H}^+]$$

$$[\text{H}^+]_{\text{total}} = [\text{H}^+]_{\text{H}_2\text{CO}_3 \text{ diss}} + [\text{H}^+]_{\text{HCO}_3^- \text{ diss}}$$

$$[\text{H}^+]_{\text{total}} = 2.07 \times 10^{-4} \text{ M} + 5.6 \times 10^{-11} \text{ M}$$

$$[\text{H}^+]_{\text{total}} = 2.07 \times 10^{-4} \text{ M}$$

Clearly, it is only the first dissociation which controls the pH of the solution.