

Chem 103 Lecture 3b

Last time: pH calculations: 4 scenarios

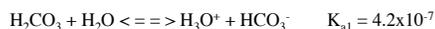
Today: polyprotic acids buffers

Polyprotic acids

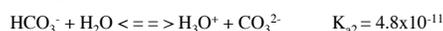
Polyprotic acids can donate more than 1 proton.

Examples: H_2CO_3 (carbonic acid); H_3PO_4 (phosphoric acid).

H_2CO_3 in water has following K_a equilibria:



NaHCO_3 in water has following K_a equilibria :



There are 2 K_a 's and deprotonation is *stepwise*

pH of pure polyprotic weak acid.

pH=? for 0.100 M H_2CO_3 ? ($K_{a1}=4.2 \times 10^{-7}$, $K_{a2}=4.7 \times 10^{-11}$)

Solution: there are 2 equilibria simultaneously occurring:



In this problem, what is the major species?

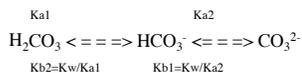
H_2CO_3 , a weak acid!

Treat it as a *monoprotic weak acid* HA, using only K_{a1}

pH of pure weak polyprotic base

pH=? for 0.100M Na_2CO_3 . ($K_{a1}=4.2 \times 10^{-7}$, $K_{a2}=4.7 \times 10^{-11}$)

We have the same 2 equilibria as in the previous example:



In this problem, what is the major species?

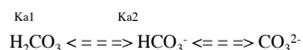
CO_3^{2-} , a weak base!

Treat it as a monoprotic weak base, using $K_{b1} = 10^{-14.00}/K_{a2}$

pH of pure intermediate form

pH=? for 0.100M NaHCO_3 . ($K_{a1}=4.2 \times 10^{-7}$, $K_{a2}=4.7 \times 10^{-11}$)

We have the same 2 equilibria as in the previous example:



Express K_a 's as $\text{p}K_a$'s: $\text{p}K_{a1} = -\log(4.2 \times 10^{-7}) = 6.38$; $\text{p}K_{a2} = 10.33$

What is the major species in this problem?

It's the intermediate species: HCO_3^- . (without proof)

$\text{pH} = (1/2)(\text{p}K_{a1} + \text{p}K_{a2}) = (1/2)(6.38 + 10.33) = 8.36$

Lewis Acids

Lewis acids = electron pair acceptors

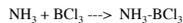
Lewis bases = electron pair donors

Consider ammonia, NH_3 and a boron trichloride, BCl_3

Do the Lewis structures of both of them.

NH_3 has a lone pair; BCl_3 has an empty orbital

NH_3 can "donate" its lone pair to BCl_3 to form a coordinate covalent bond or "dative" bond:



Which is the base? Which is the acid?