

Chem 103

Lecture 3c
Acids and Bases2

Last time

- Finished Colligative properties
- Started Acids and Bases

Today

- pH, pOH calculations
- Buffer solutions

Recall K_a and K_b equilibria

- K_a equilibrium

$$\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-; K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$
- K_b equilibrium

$$\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^-; K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

KNOW how to write the K_a and K_b equilibria.

$$K_a K_b = K_w = 1.0 \times 10^{-14}$$

Be able to prove this:

Since $K_a = [\text{H}_3\text{O}^+][\text{A}^-]/[\text{HA}]$ for acid HA.

For its conjugate base, A^- , it is: $K_b = [\text{HA}][\text{OH}^-]/[\text{A}^-]$

So if multiply K_a of weak acid by K_b of its conjugate, then :

$$K_a K_b = ([\text{H}_3\text{O}^+][\text{A}^-]/[\text{HA}]) ([\text{HA}][\text{OH}^-]/[\text{A}^-])$$

$$K_a K_b = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1.0 \times 10^{-14}$$

This refers to conjugate pairs.

pH

Due to wide range of $[\text{H}_3\text{O}^+]$, log scale is convenient: "pH".

Define pH: $\text{pH} = -\log_{10}[\text{H}^+] = -\log[\text{H}_3\text{O}^+]$

pH of pure water: since $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{M}$

$$\text{pH} = -\log \{[\text{H}_3\text{O}^+]\} = -\log(1.0 \times 10^{-7} \text{M}) = 7.00$$

This is the *neutral* pH

(Rule of thumb: $\text{pH} > 7.0$ is *basic*. $\text{pH} < 7.0$ is *acidic*.)

pH of some common liquids

Stomach acid (gastric juice)	1.0
Lemon juice	2.0
Soda	3.0
Tomato	4.0
Black coffee	5.0
Cow's milk	6.0
Human saliva	7.0
Human blood	7.4
Sea water	8.4
Soap/detergent	10.0
Household ammonia	11.5
Oven cleaner	13.5

pH of acidic solutions

If a solution contains $2.0 \times 10^{-4} \text{ M H}_3\text{O}^+$, what's its pH?
Acidic or basic?

$$\text{pH} = -\log(2.0 \times 10^{-4} \text{M}) = 3.70 \text{ (use your calculators)}$$

It is acidic since $3.70 < 7.0$

Note: for calculating $[\text{H}^+]$ given the pH, use: $[\text{H}^+] = 10^{-\text{pH}}$

For a pH 4.70 solution, what is $[\text{H}_3\text{O}^+] = ?$

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-4.70} = 2.0 \times 10^{-5} \text{ (use your calculators)}$$

pOH; pH + pOH = 14.00

Definition: $\text{pOH} = -\log[\text{OH}^-]$

Since: $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

$$-\log\{[\text{H}_3\text{O}^+][\text{OH}^-]\} = -\log\{1.0 \times 10^{-14}\}$$

$$-\log[\text{H}_3\text{O}^+] - \log[\text{OH}^-] = 14.00$$

Or, $\text{pH} + \text{pOH} = 14.00$

(so if you know pH, you can know pOH)

Summary so far...

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{and} \quad \text{pOH} = -\log[\text{OH}^-]$$

$$\text{pH} + \text{pOH} = 14.00$$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \quad K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]}$$

$$K_a K_b = 1.0 \times 10^{-14}$$

$$[\text{H}^+] = 10^{-\text{pH}} \quad [\text{OH}^-] = 10^{-\text{pOH}}$$

Calculating pH of a solution

There are at least 5 scenarios we'll encounter involving pH calculations. You are expected to **MASTER** these calculations. (simple monoprotic acids/bases)

- 1) Strong acid solutions -use: $\text{pH} = -\log[\text{acid}]$
- 2) Strong base solutions -use: $\text{pH} = 14.00 - \log[\text{base}]$
- 3) Pure Weak acid solutions -use: K_a equilibrium, ICE calculation of $[\text{H}^+]$
- 4) Pure weak base solutions -use: K_b equilibrium, ICE calculation of $[\text{OH}^-]$
- 5) Buffer solutions: -use K_a equil. ICE calc'n of $[\text{H}^+]$

OR, use buffer equation: $\text{pH} = \text{p}K_a + \log \left\{ \frac{[\text{base}]}{[\text{acid}]} \right\}$

Calculations of salt solutions

A "salt" is a soluble ionic compound.

E.g. $\text{NaCH}_3\text{CO}_2 \rightarrow \text{Na}^+ + \text{CH}_3\text{CO}_2^-$; CH_3CO_2^- is acetate, a base.

$\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$; NH_4^+ is ammonium, an acid.

In above examples, Na^+ and Cl^- act like "spectator ions".

Remember: HCH_3CO_2 is acetic acid; NH_3 is ammonia.

Continuation...

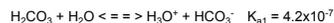
Previous 2 examples show need to determine what final solution contains. Treat the solution as being one of the various scenarios. Remember that if after neutralization, there is a mixture of strong and weak, the strong dominates the pH.

Polyprotic acids

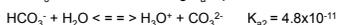
Polyprotic acids can donate more than 1 proton.

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:

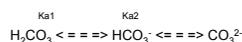


There are 2 K_a 's and deprotonation is *stepwise*

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?

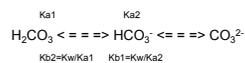
H_2CO_3 , a weak acid!

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?

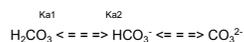
CO_3^{2-} , a weak base!

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

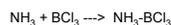
Lewis bases = electron pair donors

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?