# The Chemistry of Acids & Bases

**STUDY LIST From Paul Groves**

## Brønsted-Lowry Acids

I can…

☐ recognize common acids and bases.

☐ write balanced equations for their ionization in water.

☐ demonstrate that $H_2O$ can sometimes be included as a reactant in ionization equations.

☐ recognize when a substance can be amphoteric.

☐ explain that a Brønsted base (proton acceptor) needs a lone pair while a Brønsted acid (proton donor) requires a proton (H atom) to donate.

☐ state the definition of Brønsted acid and base.

☐ recognize a Brønsted acid and base in a reaction and identify its conjugate.

## Strengths of Acids & Bases

☐ use Tables 17.3 and 17.4 to decide on the relative strengths of acids and bases.

☐ explain the relative strengths of oxoacids such as HOCl, HOBr, and HOI
as HOCl, $HClO_2$, $HClO_3$, and $HClO_4$

☐ explain the relative strengths of acids such as HF, HCl, HBr, and HI or
$H_2O$, $H_2S$, $H_2Se$, and $H_2Te$

☐ recognize that a **very** strong acid has a weak conjugate base and a **very** weak acid has a strong conjugate base.

☐ recognize that a weak acid may have a weak conjugate base.

☐ compare the conjugate bases of two acids with known relative strengths.

☐ state that an acid-base equilibrium favors the weak acid and weak base:

$$NH_3 + H_2O \leftrightarrows NH_4^+ + OH^-$$

☐ explain the above situation in terms of the strong donor having donated and the strong acceptor having accepted.

☐ know the strong acids and bases.

## pH, pOH, [H⁺], and [OH⁻]

☐ write the equation for the autoionization of $H_2O$.

☐ recognize that every aqueous solution contains both $[H^+]$ ions and $[OH^-]$ ions.

☐ write the $K_w$ expression for $H_2O$ and know the value of $K_w = [H^+][OH^-] = 1 \times 10^{-14}$ @ 25°C.

☐ use $K_w$ to calculate $[H^+]$ and $[OH^-]$.

☐ state that $pH = -\log[H^+]$ and $pOH = -\log[OH^-]$

☐ recognize that the $[H^+]$ of a strong acid solution comes from the concentration of the acid but the $[H^+]$ of a weak acid solution requires an equilibrium calculation.

☐ calculate $[H^+]$ as $10^{-pH}$

☐ determine the number of significant figures in a pH calculation.

## Equilibrium Calculations

☐ write the $K_a$ and $K_b$ expressions for weak acids and bases.

☐ solve "ice box" problems for acid and base equilibria.

☐ correctly use the "x is small enough to ignore" approximation.

☐ know when to use the short cut for $K_a$ problems
$x = \sqrt{K_a \times [HA]_i}$ where HA = generalized acid
(this also applies to $K_b$ problems)

## Hydrolysis

☐ state that salts (ionic compounds) can make a solution acidic, basic, or have no effect on the pH.

☐ write the ionization equations for the conjugate base of a weak acid, $A^- + H_2O \leftrightarrows HA + OH^-$

☐ determine the $K_b$ of a conjugate base from the $K_a$ of its acid knowing $K_a \cdot K_b = K_w$

☐ calculate the pH of a salt solution that contains the conjugate base of a weak acid or the conjugate acid of a weak base.

## What Can Be Acids?

☐ There are a few neutral molecules that are strong acids, HCl, HBr, HI, $HNO_3$, $H_2SO_4$, $HClO_3$ & $HClO_4$ that dissociate 100% in water.

☐ Neutral molecules that are weak acids include molecules such as HF, $H_2S$, and oxoacids that are not listed above ($H_2SO_3$, $HNO_2$, $H_2CO_3$, etc.), and many organic acids, $HC_2H_3O_2$, HCOOH, etc.

☐ Cations (positive ions) can be acids, such as $NH_4^+$, but metal ions can also make solutions acidic especially if they have 2+ or 3+ charges. They are surrounded by $H_2O$ in solution and polarize the H-O bond enough that some of the H atoms leave. (Ex: $Cu^{2+}$, $Al^{3+}$, and $Fe^{3+}$)
$Fe(H_2O)_6^{3+} + H_2O \leftrightarrows H_3O^+ + Fe(H_2O)_5OH^{2+}$



$[Al(H_2O)_6]^{3+}$   $H_2O$   $[Al(H_2O)_5OH]^{2+}$   $H_3O^+$

$Ka = 1,1 . 10^{-5}$
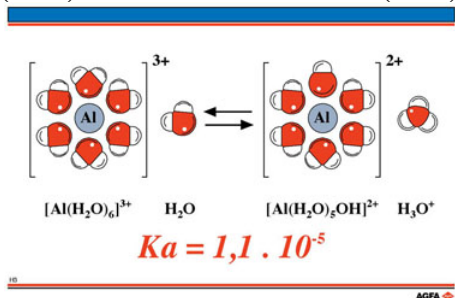
☐ Anions (negative ions) can also be acids. If an anion still has a H atom, it can donate it. Common examples are bicarbonate ion, $HCO_3^-$ and dihydrogen phosphate ion, $H_2PO_4^-$, which acts as the acid in baking soda.

## Diprotic Acids

☐ write the two acid ionization steps for any diprotic acid.
Example:  $H_2S + H_2O \leftrightarrows H_3O^+ + HS^-$
$HS^- + H_2O \leftrightarrows H_3O^+ + S^{2-}$

$K_{a1} = \dfrac{[H_3O^+][HS^-]}{[H_2S]}$   $K_{a2} = \dfrac{[H_3O^+][S^{2-}]}{[HS^-]}$

$K_{a1} = 1 \times 10^{-7}$   $K_{a2} = 1 \times 10^{-19}$

☐ state that the first ionization constant, $K_{a1}$, is always larger than the second constant, $K_{a2}$.

☐ state that the only strong diprotic acid, $H_2SO_4$, is only strong for the first ionization.
$K_{a2} = 1.2 \times 10^{-2}$.

☐ explain that when calculating the pH of a diprotic acid, only the first ionization needs to be considered.

$$H_2S + H_2O \leftrightarrows H_3O^+ + HS^-$$
$$HS^- + H_2O \leftrightarrows H_3O^+ + S^{2-}$$

The $[H_3O^+]$ from the first ionization is small. The $[H_3O^+]$ from the second ionization is *very* small and made even smaller by the presence of $H_3O^+$ from the first ionization (Le Châtelier's Principle).
*Bottom Line: treat the acid as monoprotic and use $K_{a1}$ as the $K_a$.*

☐ state that $[S^{2-}] = K_{a2}$.
explain this fact by using information from the previous situation. Because of the first ionization:

$$H_2S + H_2O \leftrightarrows H_3O^+ + HS^-$$
$$[H_3O^+] = [HS^-]$$
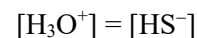
The changes caused to these two concentrations by the second ionization are small (and made negligible by Le Châtelier's Principle.)
So, $K_{a2} = \dfrac{[H_3O^+][S^{2-}]}{[HS^-]}$ becomes $K_{a2} = [S^{2-}]$!

☐ combine the two ionization steps and combine the two $K_a$'s.
$H_2S + 2 H_2O \leftrightarrows 2 H_3O^+ + S^{2-}$
$K_a = K_{a1} \cdot K_{a2} = 1 \times 10^{-26}$

☐ use this equation and $K_a$ when a problem is given in which the pH is adjusted to a known value by the addition of another acid.

### *Example from the Ch 17 Practice Test:*

13. What is the $S^{2-}$ concentration in a saturated solution (0.10 <u>M</u>) of $H_2S$, in which the pH has been adjusted to 6.00 by the addition of HCl? For $H_2S$, $K_{a1} = 1.1 \times 10^{-7}$ and $K_{a2} = 1.0 \times 10^{-19}$.