Dougherty Valley • AP Chemistry

**S-84**

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 H2O 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, HClO2, HClO3, and HClO4
* explain the relative strengths of acids such as
HF, HCl, HBr, and HI or
H2O, H2S, H2Se, and H2Te
* 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:

NH3 + H2O ⮀ NH4+ + 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 H2O.
* recognize that every aqueous solution contains both [H+] ions and [OH-] ions.
* write the Kw expression for H2O and know the value of Kw = [H+] [OH-] =1 x 10-14 @ 25°C.
* use Kw 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 Ka and Kb 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 Ka problems x =  where HA = generalized acid (this also applies to Kb 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 + H2O ⮀ HA + OH
* determine the Kb of a conjugate base from the Ka of its acid knowing Ka · Kb = Kw
* 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, HNO3, H2SO4, HClO3 & HClO4 that dissociate 100% in water.
* Neutral molecules that are weak acids include molecules such as HF, H2S, and oxoacids that are not listed above (H2SO3, HNO2, H2CO3, etc.), and many organic acids, HC2H3O2, HCOOH, etc.
* Cations (positive ions) can be acids, such as NH4+, but metal ions can also make solutions acidic especially if they have 2+ or 3+ charges. They are surrounded by H2O in solution and polarize the H-O bond enough that some of the H atoms leave. (Ex: Cu2+, Al3+, and Fe3+)
Fe(H2O)63+ + H2O ⮀ H3O+ + Fe(H2O)5OH2+ 
* Anions (negative ions) can also be acids. If an anion still has a H atom, it can donate it. Common examples are bicarbonate ion, HCO3 and dihydrogen phosphate ion, H2PO4, which acts as the acid in baking soda.

Diprotic Acids

* write the two acid ionization steps for any diprotic acid.
Example: H2S + H2O ⮀ H3O+ + HS
 HS + H2O ⮀ H3O+ + S2
Ka1 =  Ka2 = 
Ka1 = 1 x 10-7 Ka2 = 1 x 10-19
* state that the first ionization constant, Ka1, is always larger than the second constant, Ka2.
* state that the only strong diprotic acid, H2SO4, is only strong for the first ionization.
Ka2 = 1.2 x 10-2.
* explain that when calculating the pH of a diprotic acid, only the first ionization needs to be considered.
 H2S + H2O ⮀ H3O+ + HS
 HS + H2O ⮀ H3O+ + S2
The [H3O+] from the first ionization is small. The [H3O+] from the second ionization is *very* small and made even smaller by the presence of H3O+ from the first ionization (Le Châtelier’s Principle).
*Bottom Line: treat the acid as monoprotic and use Ka1 as the Ka*.
* state that [S2] = Ka2.
explain this fact by using information from the previous situation. Because of the first ionization:
 H2S + H2O ⮀ H3O+ + HS
 [H3O+] = [HS]
The changes caused to these two concentrations by the second ionization are small (and made negligible by Le Châtelier’s Principle.)
So, Ka2 =  becomes Ka2 = [S2]!
* combine the two ionization steps and combine the two Ka’s.
 H2S + 2 H2O ⮀ 2 H3O+ + S2-
 Ka = Ka1 · Ka2 = 1 x 10-26
* use this equation and Ka 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 S2- concentration in a saturated solution (0.10 M) of H2S, in which the pH has been adjusted to 6.00 by the addition of HCl? For H2S, Ka1 = 1.1 x 10-7 and Ka2 = 1.0 x 10-19.