

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:



- 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 K_w expression for H₂O and know the value of K_w = [H⁺] [OH⁻] = 1 × 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 $x = \sqrt{K_a \times [\text{HA}]_i}$ where HA = generalized acid (this also applies to K_b problems)

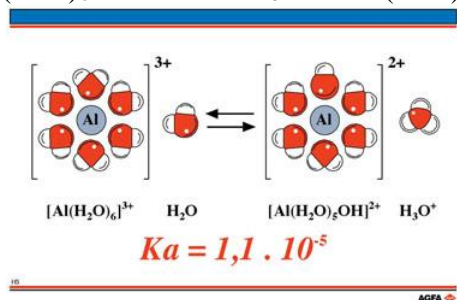
Hydrolysis

- 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, $\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^-$
- determine the K_b of a conjugate base from the K_a of its acid knowing K_a · 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, 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⁺)

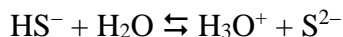
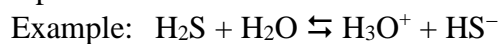
$$\text{Fe}(\text{H}_2\text{O})_6^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Fe}(\text{H}_2\text{O})_5\text{OH}^{2+}$$



- 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.

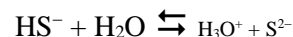


$$K_{a1} = \frac{[\text{H}_3\text{O}^+][\text{HS}^-]}{[\text{H}_2\text{S}]} \quad K_{a2} = \frac{[\text{H}_3\text{O}^+][\text{S}^{2-}]}{[\text{HS}^-]}$$

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

- state that the first ionization constant, K_{a1} , is always larger than the second constant, K_{a2} .
- state that the only strong diprotic acid, H₂SO₄, 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.



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 K_{a1} as the K_a .

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

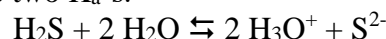


$$[\text{H}_3\text{O}^+] = [\text{HS}^-]$$

The changes caused to these two concentrations by the second ionization are small (and made negligible by Le Châtelier's Principle.)

So, $K_{a2} = \frac{[\text{H}_3\text{O}^+][\text{S}^{2-}]}{[\text{HS}^-]}$ becomes $K_{a2} = [\text{S}^{2-}]!$

- combine the two ionization steps and combine the two K_a 's.



$$K_a = K_{a1} \cdot K_{a2} = 1 \times 10^{-26}$$

- 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 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, $K_{a1} = 1.1 \times 10^{-7}$ and $K_{a2} = 1.0 \times 10^{-19}$.