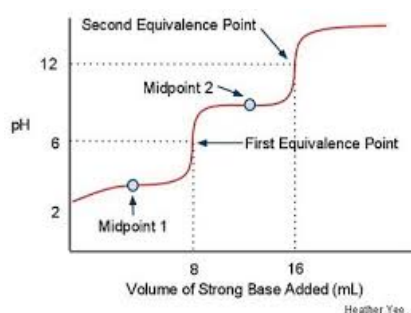


14.6 - 14.8 - Weak Bases, Polyprotic Acids, Salts



Bases

According to the Arrhenius model, a base is a species that releases hydroxide ions in solution. According to Bronsted-Lowry, a base is a proton acceptor.

The list of strong bases is the hydroxides of the following: Li, Na, K, Rb, Cs, Ca, Ba, and Sr. Resources Page 6.

Note: the alkaline earth hydroxides, while readily dissociated, are not very soluble.

1. pH of a Strong Base

Calculate the pH of a 5.0×10^{-2} M NaOH solution.

Since NaOH is a strong base, and dissociates completely, $[OH^-] = 5.0 \times 10^{-2}$ M.

Using the pOH approach:

$$pOH = -\log[OH^-] = -\log[5.0 \times 10^{-2}] = 1.30$$

Finally:

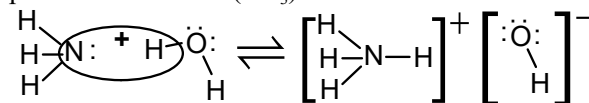
$$pH + pOH = 14.00$$

$$pH = 14.00 - 1.30 = \boxed{12.70}$$

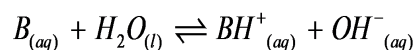
Proton Acceptors

Many bases do not contain the hydroxide ion.

These molecules strip a hydrogen ion from water, often furnishing a lone pair of electrons to which the H^+ is attracted, and release a hydroxide ion in the process. Ammonia (NH_3) is one:



The equilibrium expression for such an interaction is as follows, based on the following reaction:



$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

2. pH of a Weak Base

Calculations of weak bases are similar to that of weak acids.

Calculate the pH of a 1.0 M solution of methylamine (CH_3NH_2), $K_b = 4.38 \times 10^{-4}$.

$$K_b = 4.38 \times 10^{-4} = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} \Rightarrow \frac{(x)(x)}{1.0 - x} \approx \frac{(x)(x)}{1.0}$$

$$4.38 \times 10^{-4} \approx x^2$$

$$x \approx \sqrt{4.38 \times 10^{-4}} = 2.1 \times 10^{-2} M$$

Getting pOH:

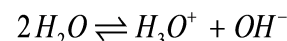
$$pOH = -\log[OH^-] = -\log[2.1 \times 10^{-2}] = 1.68$$

Finally pH: $pH + pOH = 14.00$

$$pH = 14.00 - 1.68 = \boxed{12.32}$$

Relations of K_b , K_a , and K_w

The relation between K_a and K_b of any aqueous compound are unified through the self-ionization of water:



In any aqueous solution, since:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

it turns out, for a species of interest, that:

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

Therefore, if a data table only has K_a or K_b information, you can find the missing K .

Polyprotic Acids

Polyprotic acids such as sulfuric and phosphoric acids can furnish more than one proton during a reaction. They do so stepwise: one proton at a time.

For each successive dissociation, there is a corresponding K_a , which becomes smaller with each hydrogen removed: $K_{a1} > K_{a2} > K_{a3}$. See Table 14.3.

The increasingly negative charge on the ion makes it less likely for the remaining dissociable hydrogen ions to detach.

3. pH of Sulfuric Acid

In cases where a concentrated solution of a polyprotic acid is to be calculated, the first ionization plays the largest contribution. In a dilute solution, however, that's not the case. Calculate the pH of a 1.00×10^{-2} M sulfuric acid solution.

Since sulfuric acid is a strong one, the first ionization is complete, contributing 0.0100 M of H^+ ions.

pH of Sulfuric Acid Solution

Setting up the equilibrium expression for the second dissociation vs. K_{a2} (from table):

$$K_{a2} = 1.2 \times 10^{-2} = \frac{[H^+][SO_4^{2-}]}{[HSO_4^-]} \Rightarrow \frac{(0.0100 + x)(x)}{0.0100 - x}$$

If we make the usual assumption that $0.0100 + x \sim 0.0100$, and $0.0100 - x \sim 0.0100$, then the value we get for x becomes larger than the amount of hydrogen released for the first ionization:

$$K_{a2} = 1.2 \times 10^{-2} = \frac{(0.0100)(x)}{0.0100}$$

$$x \approx 0.012 \text{ M}$$

Therefore, we have to use the full expression, and the quadratic formula (or graphing calculator).

pH of Sulfuric Acid Solution

Steering you towards using a graphing calculator (the quadratic process is on page 685 of your book):

$$K_{a2} = 1.2 \times 10^{-2} = \frac{(0.0100 + x)(x)}{0.0100 - x}$$

$$0 = \frac{0.0100x + x^2}{0.0100 - x} - 1.2 \times 10^{-2}$$

The two roots are -0.0263 , which cannot be correct, and 4.5×10^{-3} (which is correct).

Therefore, the amount of H^+ contributed is 0.0045 M, and the total becomes: $0.0100 \text{ M} + 0.0045 \text{ M} = 0.0145 \text{ M } H^+$, so:

$$pH = -\log[H^+] = -\log[0.0145] = \boxed{1.84}$$

Acid-Base Properties of Salts

Ionic compounds (called salts), regardless of their source, can be considered the offspring of a parent acid and a parent base.

Depending on the salt's parents, it will have acidic, basic, or neutral properties.

Consult your tables of acids and bases to remember which are strong or weak.

Salt Hydrolysis.

Weak acids and strong bases make basic salts,

Strong acids and weak bases make acidic salts.

Neutral salts: equally strong (or weak) acids and bases.

As they dissolve, they hydrolyze water (break it apart), releasing or combining with H^+ or OH^- ions.

Ex: Sodium fluoride (NaF) is a basic salt because it's made from NaOH (strong base) and HF (weak acid).

4. Parent Chemical Example

What are the parent chemicals of $ZnSO_4$?
Will it be acidic, basic or neutral in water?

$Zn(OH)_2$ (weak base), and H_2SO_4 – (strong acid).
It forms an acidic solution.

5. Salts as Weak Acids

Calculate the pH of a 0.10 M NH_4Cl solution.
First, determine what the parent chemicals are so you can find the appropriate K_a or K_b to use.

The parent chemicals are NH_3 (weak base) and HCl (strong acid). Note, however, it is the NH_4^+ ion which will contribute to the acidity of the solution, so use K_a for NH_4^+ for calculations ($= 5.6 E -10$).

Salts as Weak Acids Solution

Setting up the equilibrium expression in the usual way:

$$K_a = 5.6E-10 = \frac{[H^+][NH_3]}{[NH_4^+]} \Rightarrow \frac{(x)(x)}{0.10-x} \approx \frac{x^2}{0.10}$$

$$5.6E-11 \approx x^2$$

$$x \approx \sqrt{5.6E-11} = 7.5E-6M$$

pH:

$$pH = -\log[H^+] = -\log[7.5E-6] = \boxed{5.13}$$

Homework

Preview 15.1 - 15.3

14.6 - 14.8 Problems in your Booklet
Due: Next Class