Acids and Bases

- Properties
- Defining Acids and Bases
- pH and pOH

Arrhenius Definition

Arrhenius (1884) said that acids and bases release specific ions in water:

Acids - dissociate to produce H^+ ions in water

Bases - dissociate to produce OH^- ions in water

Bronsted-Lowry Definition

- Bronsted and Lowery independently (1923) said that acids and bases can be thought of H^+ donors and acceptors:

  - Acids donate H^+ ions
  - Bases accept H^+ ions

Water can either accept or donate a H^+ ions. When water accepts a H^+ ion (H_3O^+), it is called hydronium.

Conjugate Acid-Base Pairs

In acid-base equilibria, both the forward and reverse reactions involve proton transfers. In the reaction:

NH_3 (g) + H_2O (l) <--> NH_4^+ (aq) + OH^- (aq)

Because it is a reversible reaction, NH_4^+ is involved in a reverse proton transfer, in which it gives up a proton.

In the reverse reactions, the products are called conjugate acid and conjugate base to identify them as reverse reactants.

Categories for acids-bases

1) Strong acids completely transfer their protons to water, the conjugate bases do not accept (or negligibly accept) protons

2) Weak acids partially dissociate or donate protons to solution. The weak conjugate base also partially accepts protons.

3) Substances with negligible acidity that contain hydrogen have strong conjugate bases.

Self-ionization of Water

Water can self ionize, which means that if conditions are right, two molecules of water can produce a hydronium ion and a hydroxide ion:

2 $H_2O_{(l)}$ <==> $H_3O^+_{(aq)}$ + $OH^-_{(aq)}$

When this happens, we can write the equilibrium constant expression, which is given a special symbol: K_w

All aqueous solutions have a $K_w = 1.0 \times 10^{-14}$

pH

The pH scale, designed by Sorensen, was a proposal that expresses acidity and basicity in a more compact form.

Since the molar concentration of hydronium is different in different substances, we use a scale to show this concentration.

Formula for pH:

$$pH = -\log [H_3O^+]$$

A pH of 0 is very acidic. A pH of 14 is very basic. A pH of 7 is neutral.

pOH

Similar to pH, except pOH is a scale to show the concentration of OH^- ions in solution.

Formula for pOH:

$$pOH = -\log [OH^-]$$

Would a substance with a pOH of 6 be an acid or base? How about a pOH of 10?

Strong Acids & Bases

Strong acids and bases dissociate completely in water. Therefore, the molarity of the $[H^+]$ will always be equal to the molarity of the monoprotic acid:

$$HCl\ (g) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

If the molarity of the HCl was 0.1 M, then the $[H^+]$ will also be 0.1M

What if the acid or base is a polyprotic, like H_2SO_4?

Weak Acids

Because weak acids only partially dissociate in water, the dissociation is in equilibrium, and we can write an equilibrium expression (K_a)

Ex: $HCHO_2$ (aq) <==> H^+ (aq) + CHO_2^- (aq)

$$K_a = \frac{[H^+][CHO_2^-]}{[HCHO_2^-]}$$

Weak Acids & pH

From the pH of a given concentration solution, it is possible to determine the K_a of the acid.

Ex. A prepared solution of 0.10 M formic acid, $HCHO_2$, has a measured pH of 2.38. What is the K_a for the acid?

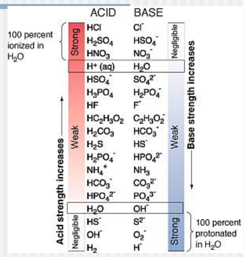
Weak Acids & pH

Using the K_a for an acid, it is possible to determine the pH of a solution.

Ex. What is the pH of a 0.30 M acetic acid solution ($K_a = 1.8 \times 10^{-5}$)

Relative Strengths of Acids & Bases

- The more readily a substance gives up a proton, the less readily the conjugate base accepts a proton.
- Or, the stronger the acid, the weaker the conjugate base.



Weak Bases

- Similar to weak acids, a weak base is a base that only partially dissociates in water, like NH_3.
- $NH_{3\ (aq)} + H_2O_{\ (l)} \Longleftrightarrow NH_4^+{}_{\ (aq)} + OH^-{}_{\ (aq)}$
- The base dissociation constant (K_b) is calculated in a similar way to K_a.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

- From the $K_{b,}$ it is possible to calculate the pOH.

K_b problems

1) A 0.5 M solution of methylamine, CH_3NH_2, has a pH of 12.2. What is the K_b of the base?
2) A 0.5 M solution of NH_3 is created. The K_b for ammonia is 1.8×10^{-5}.

a) What is the OH^- concentration for the solution at equilibrium?
b) What is the pH of the solution?

Weak Acid Lab Ka's

Weak Acid Formula	Ka
$KHSO_3$	6.4×10^{-8}
$HC_2H_4O_3$	1.6×10^{-4}
$KHC_8H_4O_4$	3.9×10^{-6}
$CH_3CO_2C_6H_4COOH$	3.2×10^{-4}
$KHSO_4$	1.0×10^{-2}
KH_2PO_4	6.2×10^{-8}
$KHC_4H_4O_6$	4.6×10^{-5}

Relationship of K_a and K_b

When two reactions are added together to give a third reaction, the equilibrium constant of the third reaction is the product of the first two.

With acid-conjugate base pairs:

$$K_a \cdot K_b = K_w$$

$$pK_a + pK_b = pK_w$$

Salt solutions

Salts that react with water in solution to create H^+ or OH^- ions undergo hydrolysis and create solutions with pHs that may be different than neutral.

The pH of a salt solution can be predicted by looking at the strength of the acids and bases that made it.

Rules for pH of salt solutions

1) Salts derived from a strong acid and a strong base is neutral.
2) Salts derived from a strong base and a weak acid will have pH > 7.
3) Salts derived from a weak base and a strong acid will have a pH < 7.
4) Salts derived from a weak acid and a weak base will have a pH dependent on which is greater between the K_a of the conjugate acid and the K_b of the conjugate base.

pH of a salt

- A solution of 0.15 M NaF is made. What is the pH of that solution?

Titration

Titration is a way to identify unknown concentrations of acids or bases.

In titration reactions, you neutralize an unknown acid (base) with a known concentration of base (acid). By knowing the amount of moles of base (acid) added, you can determine the moles of acid (base) neutralized.

Titration Curves

Titration curves are designed to graphically represent and determine the equivalence point during a titration.

Equivalence point - point at which the [H_3O^+] is equal to the [OH^-]

End point of an indicator - pH where the indicator changes color.