

## **Acids and Bases**



A Lewis acid-base reaction occurs when a base donates a pair of electrons to an acid. A **Lewis acid-base adduct**, a compound that contains a coordinate covalent bond between the Lewis acid and the Lewis base, is formed.

The boron atom in boron trifluoride, BF<sub>3</sub>, has only six electrons in its valence shell. Being short of the preferred octet, BF<sub>3</sub> is a very good Lewis acid and reacts with many Lewis bases; a fluoride ion is the Lewis base donating one of its lone pairs:



## **Relative Strengths of Acids and Bases**

The reaction of an acid with waters:

$$HA_{(aq)} + H_2O_{(I)} \Leftrightarrow H_3O^+_{(aq)} + A^-_{(aq)}$$

Water is the base that reacts with the acid HA,  $A^-$  is the conjugate base of the acid HA, and the hydronium ion is the conjugate acid of water. A strong acid yield 100% (or very nearly so) of H<sub>3</sub>O<sup>+</sup> and A<sup>-</sup> when the acid ionizes in water. A weak acid gives small amounts of H<sub>3</sub>O<sup>+</sup> and A<sup>-</sup>.

The equilibrium constant for an acid is called the acid-ionization constant, Ka:

$$K_{a} = \frac{[H_{3}O^{+}][A^{-}]}{[HA]}$$

Where the concentrations are those at equilibrium. Although water is a reactant in the reaction, it is the solvent as well, so we do not include  $[H_2O]$  in the equation. The larger the  $K_a$  of an acid, the larger of the concentration of  $H_3O^+$  and  $A^-$  relative to the concentration of the nonionized acid, HA. Thus a stronger acid has a larger denization constant than does a weaker acid. The ionization constant increase.

We can rank the strengths of bases by their tender(), to form hydroxide ions in aqueous solution. The reaction a Bronsted-Leavy has with water is given by:  $E_{(aq)} + H_2O_{(l)} \Leftrightarrow HB^+_{(aq)} + OH^-_{(aq)}$ 

Water is the acid that reacts with the base,  $HB^+$  is the conjugate acid of the base B, and the hydroxide ion is the conjugate base of water. A strong base yields 100% (or very nearly so) of  $OH^-$  and  $HB^+$  when it reacts with water.

We can measure the relative strengths of bases by measuring their **base-ionization** constant  $(K_b)$  in aqueous solutions.

$$\mathsf{K}_{\mathsf{b}} = \frac{[HB^+][OH^-]}{[B]}$$

Consider the ionization reactions for a conjugate acid-base pair, HA-A<sup>-</sup>:

$$2H_2O_{(I)} \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$
$$K_a \times K_b = \frac{[H_3O^+][A^-]}{[HA]} \times \frac{[HA][OH]}{[A^-]} = [H_3O^+][OH^-] = K_w$$

The extent to which an acid, HA, donates protons to water molecules depend on the strength of the conjugate base,  $A^-$ , of the acid. If  $A^-$  is a strong base, any protons that are donated to water molecules are recaptured by  $A^-$ . Thus there is relatively little  $A^-$  and  $H_3O^+$  in solution, and the acid, HA, is weak. If  $A^-$  is a weak base, water binds the proton more strongly, and the solution contains primarily  $A^-$  and  $H_3O^+$ — the acid is strong. Strong acids form very weak conjugate bases, and weak acids form stronger conjugate bases.

Strong Acid	Weak Acid
<ul> <li>Completely ionized (α = 1)</li> <li>Example: HCl, HBr, HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, HClO<sub>4</sub></li> <li>The ionization reaction proceeds one way to the right (irreversible reaction) HCl<sub>(aq)</sub> → H<sup>+</sup><sub>(aq)</sub> + Cl<sup>-</sup><sub>(aq)</sub></li> </ul>	<ul> <li>Partially ionized (0 &lt; α &lt; 1)</li> <li>Example: CH<sub>3</sub>COOH, H<sub>3</sub>PO<sub>4</sub>, H<sub>2</sub>S, etc.</li> <li>The ionization reaction, des back and forth (recessible reaction)</li> <li>CECCOH<sub>(aq)</sub> ⇔ H<sup>+</sup><sub>(aq)</sub> + CH<sub>3</sub>COO<sup>-</sup><sub>(aq)</sub></li> </ul>
Strong Base	Weak Base
• Completely ionized ( $\alpha$ = 1)	• Partially ionized ( $0 < \alpha < 1$ )
<ul> <li>Example: LiOH, NaOH, KOH, Ca(OH)<sub>2</sub>, Sr(OH)<sub>2</sub>, Ba(OH)<sub>2</sub></li> <li>The ionization reaction proceeds one way to the right (irreversible reaction) NaOH<sub>(aq)</sub> → Na<sup>+</sup><sub>(aq)</sub> + OH<sup>-</sup><sub>(aq)</sub></li> </ul>	<ul> <li>Example: Fe(OH)<sub>2</sub>, Al(OH)<sub>3</sub>, NH<sub>3</sub>, etc.</li> <li>The ionization reaction goes back and forth (reversible reaction) Al(OH)<sub>3(s)</sub> ⇔ Al<sup>3+</sup>(aq) + 3OH<sup>-</sup>(aq) NH<sub>3(g)</sub> + H<sub>2</sub>O(I) ⇔ NH<sub>4</sub><sup>+</sup>(aq) + OH<sup>-</sup>(aq)</li> </ul>

## **Acid and Base Reaction (Neutralization Reaction)**

A solution is neutral when it contains equal concentrations of hydronium and hydroxide ions. The following four situations with illustrate how solutions with various pH values can arise following a neutralization reaction using stoichiometrically equivalent quantities: than 5,0 x  $10^{-9}$  M (ph > 8,3), it is red or pink. Substances such as phenolphthalein, which can be used to determine the pH of solution, are called **acid-base indicators.** 

Figure below present several indicators, their colors, and their color-change intervals.







 $NH_4^+_{(aq)} \Leftrightarrow NH_{3(aq)} + H^+_{(aq)}$ 

point, however, the neutralization is complete and the pH of the solution will depend on the extent of the hydrolysis of the salt formed, which is  $CH_3COONa$ . The number of moles of NaOH in 10.0 mL is

10.0 mL X 
$$\frac{0.100 \text{ mole NaOH}}{1 \text{ L NaOH soln}}$$
 X  $\frac{1 \text{ L}}{1000 \text{ mL}}$  = 1.00 x 10<sup>-3</sup> mole

The number of moles of  $CH_3COOH$  originally present in 25.0 mL of solution is

25.0 mL X 
$$\frac{0.100 \text{ mole } CH_3COOH}{1 L CH_3COOH \text{ soln}}$$
 X  $\frac{1 L}{1000 \text{ mL}}$  = 2.50 x 10<sup>-3</sup> mole

We work with moles at this point because when two solutions are mixed, the solution volume increases. The molarity will change but the number of moles will remain the same. The changes in number of moles:



- 15. What is the pH at each of the following points in the titration of 25.00 mL of 0.100M HCl with 0.100 M NaOH?
  - (a) Before the addition of any NaOH (initial pH)
  - (b) After the addition of 24.00 mL 0.100 M NaOH (before the equivalence point)
  - (c) After the addition of 25.00 mL 0.100 M NaOH (the equivalence point)