Unit 6: Acids and Bases
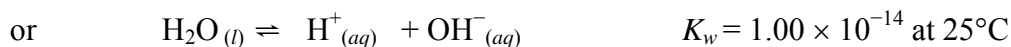
Chapter 16: Acids and Bases (pg. 568 to 571)

1. Brønsted acids are chemicals that can donate protons, whereas Brønsted bases are chemicals that can accept protons. In Brønsted definitions of acids and bases, water is an active participant during the dissociation process. When Brønsted acid dissociates, water acts like a base to accept a proton from the acid to form the hydronium ion (H_3O^+). When Brønsted base dissociates, water acts like an acid to donate a proton to the base to form hydroxide ion (OH^-). This is different than the simplistic Arrhenius definitions, in which acid is a proton producer and a base is a hydroxide producer. His definition did not utilize water in the dissociation process and hence cannot explain how weak bases (such as NH_3) that do not consist of OH^- ion have basic properties.

2. Most Brønsted bases consist of a lone pair because lone pair electrons would make that part of the molecule negative. Thereby, the base can better attract (or accept) a positive proton.

3. **(a)** both   **(b)** base   **(c)** acid   **(d)** base   **(e)** acid   **(f)** base   **(g)** base
    (h) base   **(i)** acid   **(j)** acid

5. **(a)** HCN (acid) / CN^- (base) ; CH_3COO^- (base) / CH_3COOH (acid)
    (b) HCO_3^- (acid) / CO_3^{2-} (base) ; HCO_3^- (base) / H_2CO_3 (acid)
    (c) $H_2PO_4^-$ (acid) / HPO_4^{2-} (base) ; NH_3 (base) / NH_4^+ (acid)
    (d) HClO (acid) / ClO^- (base) ; CH_3NH_2 (base) / $CH_3NH_3^+$ (acid)
    (e) H_2O (acid) / OH^- (base) ; CO_3^{2-} (base) / HCO_3^- (acid)
    (f) H_2O (acid) / OH^- (base) ; CH_3COO^- (base) / CH_3COOH (acid)

7. **(a)** CH_2ClCOO^- **(b)** IO_4^- **(c)** $H_2PO_4^-$ **(d)** HPO_4^{2-} **(e)** PO_4^{3-} **(f)** HSO_4^- **(g)** SO_4^{2-}
    (h) $HCOO^-$ **(i)** SO_3^{2-} **(j)** NH_3 **(k)** HS^- **(l)** S^{2-} **(m)** ClO^-

9. The ion-product of water is the dissociation constant of liquid water, K_w.

$$2\,H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + OH^-{}_{(aq)} \qquad\qquad K_w = 1.00 \times 10^{-14} \text{ at } 25°C$$

$$\text{or} \qquad H_2O_{(l)} \rightleftharpoons H^+{}_{(aq)} + OH^-{}_{(aq)} \qquad\qquad K_w = 1.00 \times 10^{-14} \text{ at } 25°C$$

10. $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ or $K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$

11. The equilibrium, $H_2O_{(l)} \rightleftharpoons H^+{}_{(aq)} + OH^-{}_{(aq)}$, is endothermic. This is because $K_w = 3.8 \times 10^{-14}$ at 40°C. Higher temperature causes more dissociation means that heat must be at the reactant side of the equation according to Le Châtelier's Principle.

12. pH is the negative logarithm of the $[H^+]$ or $[H_3O^+]$ and is a measurement of the acidity of a solution. Chemists normally choose to discuss acidity using pH because $[H^+]$ can vary widely (from very concentrated to very dilute). Therefore, it is much more convenient to change it into a logarithmic scale when comparing $[H^+]$ of different solutions.

13. Because the pH of the solution is 6.7 and it is less than 7, we can conclude that it is acidic. The pH of a solution can be 0. It would mean than the $[H^+] = 1$ mol/L. A negative pH is also possible and it would mean that $[H^+] > 1$ mol/L. A 16.4 mol/L of a strong acid solution (such as $HNO_{3\,(aq)}$) would have a pH of -1.21.

14. pOH is the negative logarithm of the $[OH^-]$ and is a measurement of the basicity of a solution. The equation that relates pH and pOH is: pH + pOH = 14.

15. **(a)** $[H^+] = 3.8 \times 10^{-3}$ M   **(b)** $[H^+] = 6.2 \times 10^{-12}$ M
    (c) $[H^+] = 1.1 \times 10^{-7}$ M   **(d)** $[H^+] = 1.0 \times 10^{-15}$ M

17. **(a)** pH = 3.00   **(b)** pH = 13.89   **(c)** pH = 10.74   **(d)** pH = 3.28

19.

pH	$[H^+]$	Solution is
< 7	$> 1.0 \times 10^{-7}$ M	Acid
> 7	$< 1.0 \times 10^{-7}$ M	Basic
= 7	$= 1.0 \times 10^{-7}$ M	Neutral

21. $[H^+] = 2.5 \times 10^-$ M

23. pH = 0.118

25. The strength of an acid is defined by their percentage dissociation or K_a. A strong acid is an acid that dissociates 100% or has a $K_a > 1$. A weak acid is an acid that dissociates less than 100% or has a $K_a < 1$. One cannot label an acid as strong or weak just by the pH. The identity or the initial concentration of the acid needs to be given as well as the pH before the strength of the acid can be distinguished.

26. Strong Acids: $HClO_{4\,(aq)}$, $HClO_{3\,(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_{4\,(aq)}$, $HNO_{3\,(aq)}$
Some examples of Weak Acids: $HF_{(aq)}$, $HNO_{2\,(aq)}$, $H_2CO_{3\,(aq)}$, $H_2SO_{3\,(aq)}$, $CH_3COOH_{(aq)}$

27. Any acid that dissociates completely in water is a strong acid. From the list in the textbook, $HClO_{4\,(aq)}$ is the strongest acids. A base that dissociates completely in water is a strong base. Any alkali metal hydroxides (NaOH and KOH and others) and alkaline earth hydroxide ($Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, except $Mg(OH)_2$ and $Be(OH)_2$) are considered strong bases.

28. H_2SO_4 is a strong acid because the first proton completely dissociates from the parent molecule. The proton in HSO_4^-, however, does not dissociate completely. Hence, it makes HSO_4^- a weak acid.

29. Strong Acid (**b**), Weak Acid (**c**), Very Weak Acid (**d**)

31. (**a**) strong acid   (**b**) weak acid   (**c**) strong acid (first stage of ionization)   (**d**) weak acid
    (**e**) weak acid   (**f**) weak acid   (**g**) strong acid   (**h**) weak acid   (**i**) weak acid

33. (**a**) false   (**b**) false   (**c**) true   (**d**) false

35. The direction should favor formation of $F^-_{(aq)}$ and $H_2O_{(l)}$. Hydroxide ion is a stronger base than fluoride ion, and hydrofluoric acid is a stronger acid than water.

37. The bigger the K_a, the stronger is the acid because it can dissociate (or donate protons) more readily than an acid with a smaller K_a. When $K_a > 1$, we label the acid as a strong acid.

38. Since K_a is the equilibrium constant of acid dissociation and all equilibrium constant is dependent on temperature, K_a is also dependent on temperature.

39. Since the K_a values of strong acids are very large, the dissociations go to completion. Hence, there is no need to use the ICE box to calculate $[H_3O^+]$. The original concentration of the strong acid is the $[H_3O^+]$. Since equilibrium calculation is not needed for the determination of $[H_3O^+]$ of a strong acid, we do not usually quote the K_a values of strong acids.

40. Since all the acids have the same initial concentration of 0.40 mol/L, the acid that has the highest K_a would have the lowest pH (dissociates more H^+ and therefore lower pH). Conversely, the acid that has the lowest K_a would have the highest pH. Between HCOOH, $HClO_4$ and CH_3COOH, acetic acid ($CH_3COOH_{(aq)}$) has the smallest K_a. Hence, it would have the highest pH.

41. $[H^+] = [CN^-] = 8.6 \times 10^{-6}$ M; [HCN] = 0.15 M

43. $[F^-] / [HF] = 1.1 \times 10^3$

45. pH = 2.21

47. $K_a = 9.2 \times 10^{-4}$

49. Malonic acid, $[CH_2(COOH)_2]$, is diprotic because it can donate two protons (but one at a time).

$$HOOC-CH_2-COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HOOC-CH_2-COO^-_{(aq)}$$

$$HOOC-CH_2-COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + {}^-OOC-CH_2-COO^{2-}_{(aq)}$$

50. Amphoteric means that a chemical species can be an acid or a base. $H_3PO_{4\,(aq)}$ is a weak acid. Hence, at equilibrium, there will still be a lot of H_3PO_4 molecules around.
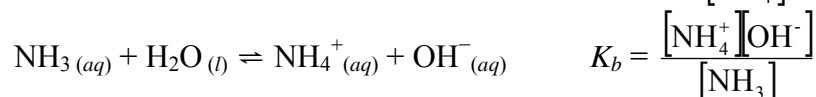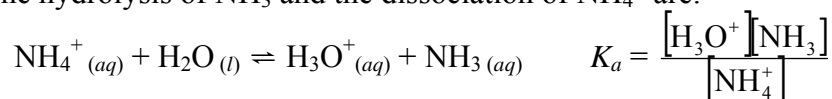
$$H_3PO_{4\,(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + H_2PO_4^-{}_{(aq)}$$

$$H_2PO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HPO_4^{2-}{}_{(aq)}$$

$$HPO_4^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + PO_4^{3-}{}_{(aq)}$$

The species at equilibrium would be H_3PO_4 (acid), $H_2PO_4^-$ (amphoteric), HPO_4^{2-} (amphoteric), PO_4^{3-} (base), H_3O^+ (acid), and H_2O (amphoteric).

51. $[H^+] = [SO_4^{2-}] = 0.045$ M; $[HSO_4^-] = 0.16$ M

53. The hydrolysis of NH_3 is $NH_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$. Ammonia, NH_3, is a weak base because it has a $K_b = 1.8 \times 10^{-5} < 1$ and it does not accept proton 100% of the time. At equilibrium, the major species would be $NH_{3\,(aq)}$ as only a small amount of OH^- is formed. $[NH_3]_{eq} > [OH^-]_{eq}$

54. $K_w = K_a \times K_b$. The hydrolysis of NH_3 and the dissociation of NH_4^+ are:

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + NH_{3\,(aq)} \qquad K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]}$$

$$NH_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)} \qquad K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$K_a \times K_b = \frac{[H_3O^+][\cancel{NH_3}]}{[\cancel{NH_4^+}]} \times \frac{[\cancel{NH_4^+}][OH^-]}{[\cancel{NH_3}]} = [H_3O^+][OH^-] = K_w$$

55. **(a)** pH = 11.1   **(b)** pH = 8.96

57. Original $[NH_3] = 0.15$ M

65. Salt hydrolysis is when an ionic compound with a weak base as its anion is dissolved in water. The weak base anion accepts a proton from water to produce its conjugate base and OH^-. As a result, the pH of the solution increases from 7.

68. $KF - F^-$ is a weak base and will undergo hydrolysis
$NH_4NO_2 - NO_2^-$ is a weak base and will undergo hydrolysis
$MgSO_4 - SO_4^{2-}$ is a weak base and will undergo hydrolysis
$KCN - CN^-$ is a weak base and will undergo hydrolysis
$NaC_6H_5COO - C_6H_5COO^-$ is a weak base and will undergo hydrolysis
$Na_2CO_3 - CO_3^{2-}$ is a weak base and will undergo hydrolysis
$KHCOO - HCOO^-$ is a weak base and will undergo hydrolysis

71. There are two possibilities: (i) MX is the salt of a strong acid and a strong base so that neither the cation nor the anion reacts with water to alter the pH and (ii) MX is the salt of a weak acid and a weak base with K_a for the acid equal to K_b for the base. The hydrolysis of one would be exactly offset by the hydrolysis of the other.

73. pH = 9.15

93. $K = 4.0 \times 10^{-2}$

Chapter 17: Acid-Base Equilibria (pg. 604)

1. A buffer solution can resist the change in pH when a small amount of strong acid or strong base is added to it. It is consisted of relatively equal concentration of a weak acid and its conjugate base.
(Example 0.1 mol/L of CH_3COOH / 0.1 mol/L of $NaCH_3COO$)

5. Buffer Solution Systems: **(b)** and **(c)**

17. Molar Mass = 202 g/mol

19. **(a)** pH = 5.28   **(b)** pH = 8.72

21. An acid-base indicator is itself a weak acid. Most acid-base indicators are organic molecules. The acid form of these molecules has a different colour than their conjugate base counterparts. Since they have different K_a, their colour change can be of different pH ranges. Hence, we can approximate the pH of a solution if we use a multiple number of different acid-base indicators.

22. The acid-base indicator should change colour at roughly the same pH as the equivalence point of the titration.

23. An indicator is typically a weak organic acid. The indicator will react with the base component of the titration. If a large amount of indicator is used, it will react with a significant amount of the base affecting the equivalence point of the titration.