14.1 19 (20) 21, 23, 25, (26) 29

CHAPTER 14

ACIDS AND BASES

Questions

19. Acids are proton (H^+) donors, and bases are proton acceptors.

HCO_3^- as an acid: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$

HCO_3^- as a base: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

$H_2PO_4^-$ as an acid: $H_2PO_4^- + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$

$H_2PO_4^-$ as a base: $H_2PO_4^- + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq)$

21. Basic solutions (at 25°C) have an $[OH^-] > 1.0 \times 10^{-7}$ M, which gives a pOH < 7.0. Because $[H^+][OH^-] = 1.0 \times 10^{-14}$ and pH + pOH = 14.00 for any aqueous solution at 25°C, a basic solution must also have $[H^+] < 1.0 \times 10^{-7}$ M and pH > 7.00. From these relationships, the solutions in parts b, c, and d are basic solutions. The solution in part a will have a pH < 7.0 (pH = 14.00 – 11.21 = 2.79) and is therefore not basic (solution is acidic).

23. 10.78 (4 S.F.); 6.78 (3 S.F.); 0.78 (2 S.F.); a pH value is a logarithm. The numbers to the left of the decimal point identify the power of 10 to which $[H^+]$ is expressed in scientific notation, for example, 10^{-11}, 10^{-7}, 10^{-1}. The number of decimal places in a pH value identifies the number of significant figures in $[H^+]$. In all three pH values, the $[H^+]$ should be expressed only to two significant figures because these pH values have only two decimal places.

25.

NH_3	+	NH_3	$\rightleftharpoons$	NH_2^-	+	NH_4^+
Acid		Base		Conjugate Base		Conjugate Acid

One of the NH_3 molecules acts as a base and accepts a proton to form NH_4^+. The other NH_3 molecule acts as an acid and donates a proton to form NH_2^-. NH_4^+ is the conjugate acid of the NH_3 base. In the reverse reaction, NH_4^+ donates a proton. NH_2^- is the conjugate base of the NH_3 acid. In the reverse reaction, NH_2^- accepts a proton. Conjugate acid-base pairs only differ by a H^+ in the formula.

27. a. These are solutions of strong acids like HCl, HBr, HI, HNO_3, H_2SO_4, and $HClO_4$. So 0.10 M solutions of any of the acids would be examples of a strong electrolyte solution that is very acidic.

b. These are solutions containing salts of the conjugate acids of the bases in Table 14.3. These conjugate acids are all weak acids, and they are cations with a 1+ charge. NH_4Cl, $CH_3NH_3NO_3$, and $C_2H_5NH_3Br$ are three examples of this type of slightly acidic salts.

Note that the anions used to form these salts are conjugate bases of strong acids; this is so because they have no acidic or basic properties in water (with the exception of HSO_4^-, which has weak acid properties).

c. These are solutions of strong bases like LiOH, NaOH, KOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$. All of these strong bases are strong electrolytes.

d. These are solutions containing salts of the conjugate bases of the neutrally charged weak acids in Table 14.2. These conjugate bases are all weak bases, and they are anions with a 1− charge. Three examples of this type of slightly basic salts are $NaClO_2$, $KC_2H_3O_2$, and CaF_2. The cations used to form these salts are Li^+, Na^+, K^+, Rb^+, Cs^+, Ca^{2+}, Sr^{2+}, and Ba^{2+} because these cations have no acidic or basic properties in water. Notice that these are the cations of the strong bases you should memorize.

e. There are two ways to make a neutral salt solutions. The easiest way is to combine a conjugate base of a strong acid (except for HSO_4^-) with one of the cations from a strong base. These ions have no acidic/basic properties in water, so salts of these ions are neutral. Three examples are NaCl, KNO_3, and SrI_2. Another type of strong electrolyte that can produce neutral solutions are salts that contain an ion with weak acid properties combined with an ion of opposite charge having weak base properties. If the K_a for the weak acid ion is equal to the K_b for the weak base ion, then the salt will produce a neutral solution. The most common example of this type of salt is ammonium acetate ($NH_4C_2H_3O_2$). For this salt, K_a for $NH_4^+ = K_b$ for $C_2H_3O_2^- = 5.6 \times 10^{-10}$. This salt at any concentration produces a neutral solution.

29. a. $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ or

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \quad K = K_w = [H^+][OH^-]$$

b. $HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$ or

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq) \quad K = K_a = \frac{[H^+][F^-]}{[HF]}$$

c. $$C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq) \quad K = K_b = \frac{[C_5H_5NH^+][OH^-]}{[C_5H_5N]}$$

31. a. This expression holds true for solutions of strong acids having a concentration greater than 1.0×10^{-6} *M*. 0.10 *M* HCl, 7.8 *M* HNO_3, and 3.6×10^{-4} *M* $HClO_4$ are examples where this expression holds true.

b. This expression holds true for solutions of weak acids where the two normal assumptions hold. The two assumptions are that water does not contribute enough H^+ to solution to make a difference, and that the acid is less than 5% dissociated in water (from the assumption that x is small compared to some number). This expression will generally hold true for solutions of weak acids having a K_a value less than 1×10^{-4}, as long as there is a significant amount of weak acid present. Three example solutions are 1.5 *M* $HC_2H_3O_2$, 0.10 *M* HOCl, and 0.72 *M* HCN.

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20a) Acidic

b) neutral

c) Basic

d) Acidic

26 a) A + B ⟷ A + B

A) B) B + A ⟷ A + B

B) Bronsted-Lowry possibly Arrhenius

Not Lewis