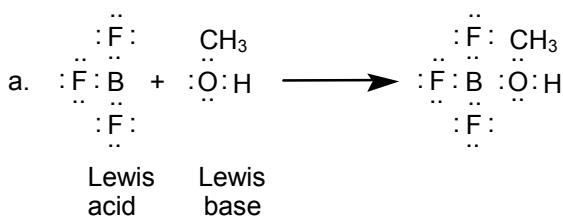# 16. ACIDS AND BASES

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.
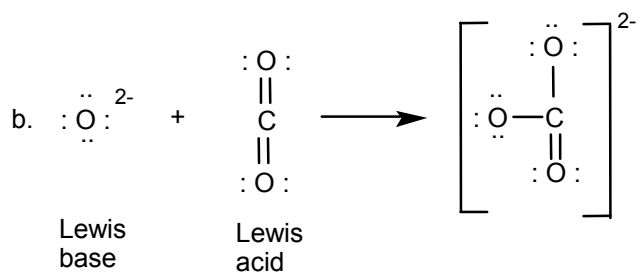
16.1 See labels below reaction:

$$H_2CO_3(aq) \quad + \quad CN^-(aq) \quad \rightleftharpoons \quad HCN(aq) \quad + \quad HCO_3^-(aq)$$

acid   base   acid   base

H$_2$CO$_3$ is the proton donor (Bronsted-Lowry acid) on the left, and HCN is the proton donor (Bronsted-Lowry acid) on the right. The CN$^-$ and HCO$_3^-$ ions are proton acceptors (Bronsted-Lowry bases). HCN is the conjugate acid of CN$^-$.

16.2 Part (a) involves molecules with all single bonds; part (b) does not, so bonds are drawn in.

a.

Lewis acid   Lewis base

(continued)

b. $\begin{array}{c} \cdot\cdot \\ :O: \\ \cdot\cdot \end{array}^{2-}$ $\;+\;$ $\begin{array}{c} :O: \\ \| \\ C \\ \| \\ :O: \end{array}$ $\;\longrightarrow\;$ $\left[\begin{array}{c} :\underline{O}: \\ | \\ :\underset{\cdot\cdot}{O}-C \\ \| \\ :\underline{O}: \end{array}\right]^{2-}$

Lewis   Lewis
base   acid

16.3   The $HC_2H_3O_2$ is a stronger acid than H_2S, and HS^- is a stronger base than the $C_2H_3O_2^-$ ion. The equilibrium favors the weaker acid and weaker base; therefore, the reactants are favored.

16.4   a. PH_3 b. HI c. H_2SO_3 d. H_3AsO_4 e. HSO_4^-

16.5   A 0.125 M solution of $Ba(OH)_2$, a strong base, ionizes completely to yield 0.125 M Ba^{2+} ion and 2 x 0.125 M, or 0.250 M, OH^- ion. Use the K_w equation to calculate the $[H_3O^+]$.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{(0.250)} = 4.\underline{00} \times 10^{-14} = 4.0 \times 10^{-14} \text{ M}$$

16.6   Use the K_w equation to calculate the $[H_3O^+]$.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.\underline{00} \times 10^{-9} = 1.0 \times 10^{-9} \text{ M}$$

Since the $[H_3O^+]$ concentration is less than 1.0×10^{-7}, the solution is basic.

16.7   Calculate the negative log of the $[H_3O^+]$:

pH $= -\log [H_3O^+] = -\log (0.045) = 1.3\underline{4}6 = 1.35$

16.8   Calculate the pOH of 0.025 M OH^-, and then subtract from 14.00 to find pH:

pOH $= -\log [OH^-] = -\log (0.025) = 1.6\underline{0}2$

pH $= 14.00 - 1.602 = 12.3\underline{9}7 = 12.40$

16.9   Because pH = 3.16, by definition the log $[H_3O^+]$ = -3.16. Enter this on the calculator and convert to the antilog (number) of -3.16.

$$[H_3O^+] = \text{antilog } (-3.16) = 10^{-3.16} = 6.\underline{9}1 \times 10^{-4} = 6.9 \times 10^{-4} \text{ M}$$

16.10   Find the pOH by subtracting the pOH from 14.00. Then enter -3.40 on the calculator to convert to the antilog (number) corresponding to -3.40.

$$\text{pOH} = 14.00 - 10.6 = 3.\underline{4}0$$

$$[H_3O^+] = \text{antilog } (-3.40) = 10^{-3.40} = \underline{3}.98 \times 10^{-4} = 4 \times 10^{-4} \text{ M}$$

■ Answers to Concept Checks

16.1   In any aqueous solution, you should consider the autoionization of water. And because we have a solution of a weak acid in water, you should also consider the equilibrium between this acid and water. Here are the two equilibria:
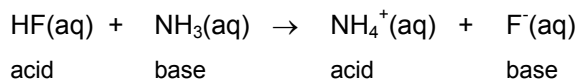
$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

$$HCHO_2(aq) + H_2O(l) \rightleftharpoons CHO_2^-(aq) + H_3O^+(aq)$$

The species present in these equilibria are: $H_2O(l)$, $H_3O^+(aq)$, $OH^-(aq)$, $HCHO_2(aq)$, and $CHO_2^-(aq)$.

16.2   The stronger acid gives up its proton more readily, and, therefore, its conjugate base ion holds onto a proton less strongly. In other words, the stronger acid has the weaker conjugate base. Because formic acid is the stronger acid, the formate ion is the weaker base. Acetate ion is the stronger base.

16.3   Look at each solution, and determine whether it is acidic, basic, or neutral. In solution A, the numbers of H_3O^+ and OH^- ions are equal, so the solution is neutral. For solution B, the number of H_3O^+ ions is greater than the number of OH^- ions, so the solution is acidic. In solution C, the number of H_3O^+ ions is less than the number of OH^- ions, so the solution is basic. Therefore, ranking from most acidic to least acidic (most basic) is B > A > C.

16.4   In order to qualitatively answer this problem, it is essential that all the solutions have the same solute concentrations. Bases produce solutions of pH greater than seven while acids produce solutions of pH less than seven. NH_3 and NaOH are bases, and HCl, and $HC_2H_3O_2$ are acids. NaOH is a stronger base than NH_3, so the NaOH solution would have the highest pH, followed by the NH_3 solution. $HC_2H_3O_2$ is a much weaker acid than HCl so the $HC_2H_3O_2$ solution would have a higher pH than the HCl solution. Therefore, the ranking from highest to lowest pH for solutions with the same solute concentrations is: NaOH > NH_3 > $HC_2H_3O_2$ > HCl.

■ Answers to Review Questions

16.1   You can classify these acids using the information in Section 16.1. Also recall that all diatomic acids of Group VIIA halides are strong except for HF.

    a. Weak   b. Weak   c. Strong

    d. Strong   e. Weak   f. Weak

16.2   In Section 16.1, we are told that all neutralizations involving strong acids and bases evolve 55.90 kJ of heat per one mol H_3O^+. Thus, the thermochemical evidence for the Arrhenius concept is based on the fact that when one mole of any strong acid (one mol H_3O^+) is neutralized by one mole of any strong base (one mol OH^-), the heat of neutralization is always the same ($\Delta H° = -55.90$ kJ/mol).

16.3   A Bronsted-Lowry acid is a molecule or ion that donates an H_3O^+ ion (proton donor) to a base in a proton-transfer reaction. A Bronsted-Lowry base is a molecule or ion that accepts an H_3O^+ ion (proton acceptor) from an acid in a proton-transfer reaction. An example of an acid-base equation:

$$HF(aq) \; + \; NH_3(aq) \; \rightarrow \; NH_4^+(aq) \; + \; F^-(aq)$$
$$\text{acid} \qquad\qquad \text{base} \qquad\qquad \text{acid} \qquad\qquad \text{base}$$

16.4   The conjugate acid of a base is a species that differs from the base by only one H_3O^+. Consider the base, HSO_3^-. Its conjugate acid would be H_2SO_3 but not H_2SO_4. H_2SO_4 differs from HSO_3^- by one H and one O.

16.5   You can write the equations by considering that $H_2PO_3^-$ is both a Bronsted-Lowry acid and a Bronsted-Lowry base. The $H_2PO_3^-$ acts as a Bronsted-Lowry acid when it reacts with a base such as OH^- :

$$H_2PO_3^-(aq) \quad + \quad OH^-(aq) \quad \rightarrow \quad HPO_3^{2-}(aq) \quad + \quad H_2O(l)$$
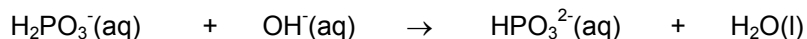
The $H_2PO_3^-$ acts as a Bronsted-Lowry base when it reacts with an acid such as HF:

$$H_2PO_3^-(aq) \quad + \quad HCl(aq) \quad \rightarrow \quad H_3PO_3(aq) \quad + \quad Cl^-(aq)$$
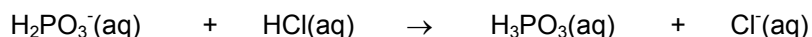
16.6   The Bronsted-Lowry concept enlarges on the Arrhenius concept in the following ways: (1) It expands the concept of a base to include any species that accepts protons, not just the OH^- ion or compounds containing the OH^- ion. (2) It enlarges the concepts of acids and bases to include ions as well as molecules. (3) It enables us to write acid-base reactions in nonaqueous solutions as well as in aqueous solutions, whereas the Arrhenius concept applies only to aqueous solutions. (4) It allows some species to be considered as acids or bases depending on the other reactant with which they are mixed.

16.7   According to the Lewis concept, an acid is an electron-pair acceptor and a base is an electron-pair donor. An example is

$$Ag^+(aq) \quad + \quad 2(:NH_3) \quad \rightarrow \quad Ag(NH_3)_2^+(aq)$$
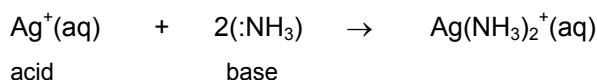
     acid   base

16.8   Recall that, the weaker the acid, the stronger it holds on to its proton(s). Thus, if a reaction mixture consists of a stronger acid and base and a weaker acid and base, the weaker acid side will always be favored because the proton(s) will bond more strongly to the weaker acid.

16.9   The two factors that determine the strength of an acid are (1) the polarity of the bond to which the H atom is attached, and (2) the strength of the bond, or how tightly the proton is held by the atom to which it is bonded. An increase in the polarity of the bond makes it easier to remove the proton, increasing the strength of the acid. An increase in the strength of the bond makes it more difficult to remove the proton, decreasing the strength of the acid. The strength of the bond depends in turn on the size of the atom, so larger atoms have weaker bonds, whereas smaller atoms have stronger bonds.

16.10   The self-ionization of water is the reaction of two water molecules in which a proton is transferred from one molecule to the other to form H_3O^+ and OH^- ions. At 25 °C, the K_w expression is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

16.11   The pH = -log $[H_3O^+]$ of an aqueous solution. Measure pH by using electrodes and a pH meter, or by interpolating the pH from the color changes of a series of acid-base indicators.

16.12   A solution of pH 4 has a $[H_3O^+] = 1 \times 10^{-4}$ M and is more acidic than a solution of pH 5, which has a $[H_3O^+] = 1 \times 10^{-5}$ M.

16.13   For a neutral solution, $[H_3O^+] = [OH^-]$; thus, the $[H_3O^+]$ of a neutral solution at 37 °C is the square root of K_w at 37 °C:

$$[H_3O^+] = \sqrt{2.5 \times 10^{-14}} = 1.\underline{5}8 \times 10^{-7} \text{ M}$$

$$pH = -\log(1.58 \times 10^{-7}) = 6.8\underline{0}1 = 6.80$$

16.14   Because pH + pOH = pK_w at any temperature,

pH + pOH = -log (2.5 x 10^{-14}) = 13.60

■ Answers to Conceptual Problems

16.15   It is not necessary to have the species NH_4OH in order to have OH^- in the solution. When ammonia reacts with water, hydroxide ion forms in the reaction.

$NH_3(aq)$ + $H_2O(l)$ ⇌ $NH_4^+(aq)$ + $OH^-(aq)$

16.16   A reaction where HPO_4^{2-} acts as an acid is

$HPO_4^{2-}(aq)$ + $OH^-(aq)$ ⇌ $PO_4^{3-}(aq)$ + $H_2O(l)$

A reaction where HPO_4^{2-} acts as a base is

$HPO_4^{2-}(aq)$ + $H_3O^+(aq)$ ⇌ $H_2PO_4^-(aq)$ + $H_2O(l)$

16.17   The hydroxide ion acts as a base and donates a pair of electrons on the O atom, forming a bond with CO_2 to give HCO_3^-.

16.18   Nitrogen has a greater electronegativity than carbon. You would expect the H—O bond in the H—O—N group to be more polar (with the H atom having a positive partial charge) than the H—O bond in the H—O—C group. Thus, based on their structure, you would expect HNO_2 to be the stronger acid.

16.19   When you lower the temperature of pure water, the value of K_w decreases. In pure water, the hydronium ion concentration equals the hydroxide ion concentration, so $K_w = [H_3O^+]^2$. When K_w decreases, the hydronium ion concentration decreases, and the corresponding pH increases.

16.20   Sodium hydroxide is a strong base, whereas ammonia is weak. As a strong base, NaOH exists in solution completely as ions, whereas NH_3 exists in solution as an equilibrium in which only part of the NH_3 has reacted to produce ions. Thus, a sodium hydroxide solution has a greater OH^- concentration than the same concentration solution of NH_3. At the same concentrations, the pH of the NaOH solution is greater (more basic) than that of the NH_3 solution.

16.21 A strong monoprotic acid will dissolve completely in water to form hydronium ions (H_3O^+) and anions in equal number. None of the original monoprotic acid molecules will remain. This is represented by the picture in the middle. The picture on the left represents an undissociated acid. No ions are present in the solution. The picture on the right represents a weak acid where only some of the acid molecules have dissociated to form ions.

16.22 A weak acid will dissolve to form a solution that has some hydronium ions present, some anions present, and some undissociated acid molecules present. This is represented by beaker A. The solution in beaker B is a weak base, and the solution in beaker C is a strong acid.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

16.23 The reaction with labels of "acid" or "base" written below is as follows:

$$OH^-(aq) \ + \ HF(l) \ \rightleftharpoons \ F^-(aq) \ + \ H_2O(l)$$

base   acid   base   acid

16.24 The reaction with labels of "acid" or "base" written below is as follows:

$$HOCl(aq) \ + \ H_2O(l) \ \rightleftharpoons \ H_3O^+(aq) \ + \ OCl^-(aq)$$

acid   base   acid   base

16.25 a. PO_4^{3-} b. HS^- c. NO_2^- d. $HAsO_4^{2-}$

16.26 a. HSO_4^- b. PH_3 c. Se^{2-} d. OBr^-

16.27 a. $HClO$ b. AsH_4^+ c. H_3PO_4 d. $HTeO_3^-$

16.28 a. H_2S b. NH_3 c. $HBrO_2$ d. $N_2H_5^+$

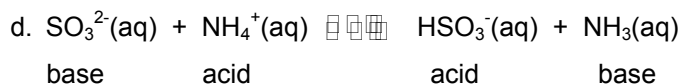16.29 Each equation is given below with the labels for acid or base:

a. $HSO_4^-(aq)$ + $NH_3(aq)$ ⇌ $SO_4^{2-}(aq)$ + $NH_4^+(aq)$
    acid   base   base   acid

The conjugate acid-base pairs are: HSO_4^-, SO_4^{2-}, and NH_4^+, NH_3.

b. $HPO_4^{2-}(aq)$ + $NH_4^+(aq)$ ⇌ $H_2PO_4^-(aq)$ + $NH_3(aq)$
    base   acid   acid   base

The conjugate acid-base pairs are: $H_2PO_4^-$, HPO_4^{2-}, and NH_4^+, NH_3.

c. $Al(H_2O)_6^{3+}(aq)$ + $H_2O(l)$ ⇌ $Al(H_2O)_5(OH)^{2+}(aq)$ + $H_3O^+(aq)$
    acid   base   base   acid

The conjugate acid-base pairs are: $Al(H_2O)_6^{3+}$, $Al(H_2O)_5(OH)^{2+}$, and H_3O^+, H_2O.

d. $SO_3^{2-}(aq)$ + $NH_4^+(aq)$ ⇌ $HSO_3^-(aq)$ + $NH_3(aq)$
    base   acid   acid   base

The conjugate acid-base pairs are: HSO_3^-, SO_3^{2-}, and NH_4^+, NH_3.

16.30 Each equation is given below with the labels for acid or base:

a. $H_2PO_4^-(aq)$ + $HCO_3^-(aq)$ ⇌ $HPO_4^{2-}(aq)$ + $H_2CO_3(aq)$
    acid   base   base   acid

The conjugate acid-base pairs are: $H_2PO_4^-$, HPO_4^{2-}, and H_2CO_3, HCO_3^-.

b. $F^-(aq)$ + $HSO_4^-(aq)$ ⇌ $HF(aq)$ + $SO_4^{2-}(aq)$
    base   acid   acid   base
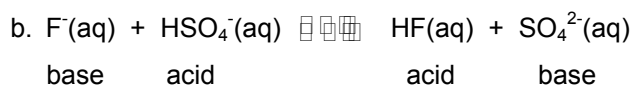
The conjugate acid-base pairs are: HF, F^-, and HSO_4^-, SO_4^{2-}.

c. $HSO_4^-(aq)$ + $H_2O(l)$ ⇌ $SO_4^{2-}(aq)$ + $H_3O^+(aq)$
    acid   base   base   acid

The conjugate acid-base pairs are: HSO_4^-, SO_4^{2-}, and H_3O^+, H_2O.

d. $H_2S(aq)$ + $CN^-(aq)$ ⇌ $HS^-(aq)$ + $HCN(aq)$

    acid   base   base   acid
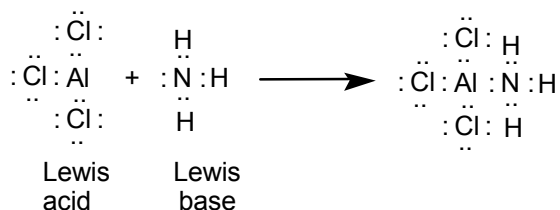
The conjugate acid-base pairs are: H_2S, HS^-, and HCN, CN^-.

16.31   The reaction is
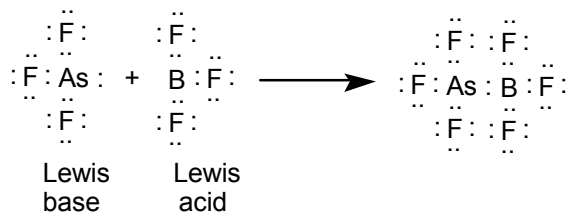
16.32   The reaction is

16.33 a. The completed equation is

$AlCl_3$ + Cl^- ⇌ $AlCl_4^-$
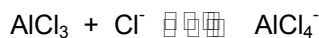
The Lewis formula representation is
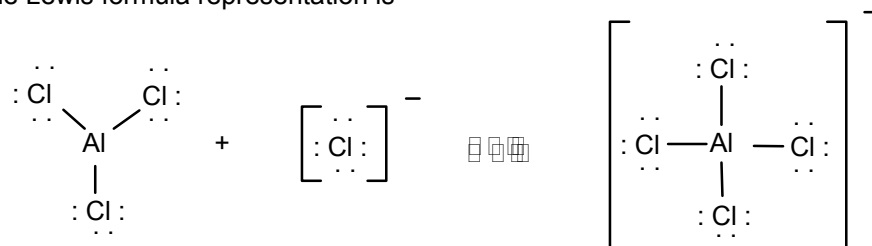
$AlCl_3$ is the electron-pair acceptor and is the acid. Cl^- is the electron pair donor and is the base.
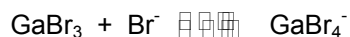
(continued)

b. The completed equation is

$I^- + I_2 \rightleftharpoons I_3^-$

The Lewis formula representation is

$$\left[:\ddot{I}: \right]^- + :\ddot{I}-\ddot{I}: \rightleftharpoons \left[:\ddot{I}-\ddot{I}-\ddot{I}: \right]^-$$

I_2 is the electron-pair acceptor and is the acid. The I^- ion is the electron pair donor and is the base.

16.34 a. The completed equation is

$GaBr_3 + Br^- \rightleftharpoons GaBr_4^-$

The Lewis formula representation is

$$\underset{:\ddot{B}r:}{\overset{:\ddot{B}r \quad \ddot{B}r:}{Ga}} \quad + \quad \left[:\ddot{B}r: \right]^- \rightleftharpoons \left[\underset{:\ddot{B}r:}{\overset{:\ddot{B}r:}{:\ddot{B}r-Ga-\ddot{B}r:}} \right]^-$$
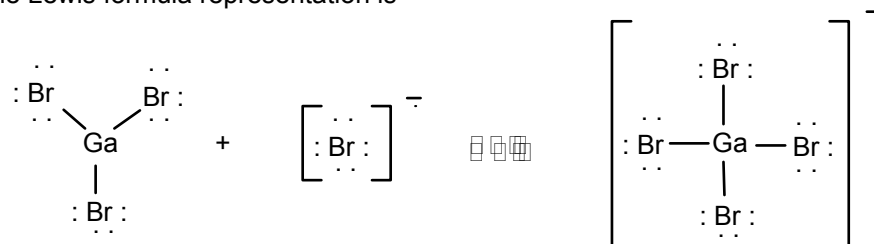
$GaBr_3$ is the electron-pair acceptor and is the acid. The Br^- ion is the electron pair donor and is the base.

b. The completed equation is

$BF_3 + F^- \rightleftharpoons BF_4^-$

The Lewis formula representation is

$$\underset{:\ddot{F}:}{\overset{:\ddot{F} \quad \ddot{F}:}{B}} \quad + \quad \left[:\ddot{F}: \right]^- \rightleftharpoons \left[\underset{:\ddot{F}:}{\overset{:\ddot{F}:}{:\ddot{F}-B-\ddot{F}:}} \right]^-$$
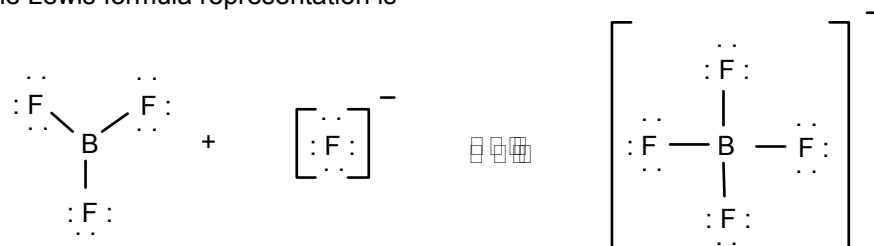
BF_3 is the electron-pair acceptor and is the acid. The F^- ion is the electron pair donor and is the base.

16.35 a. Each water molecule donates a pair of electrons to copper(II), making the water molecule a Lewis base and the Cu^{2+} ion a Lewis acid.

b. The AsH_3 donates a pair of electrons to the boron atom in BBr_3, making AsH_3 a Lewis base and the BBr_3 molecule a Lewis acid.

16.36 a. Each OH^- ion donates a pair of electrons to Al in $Al(OH)_3$, making OH^- a Lewis base and $Al(OH)_3$ a Lewis acid.

b. Each CN^- ion donates a pair of electrons to silver ion, making CN^- ion a Lewis base and the Ag^+ ion a Lewis acid.

16.37 The equation is $H_2S + HOCH_2CH_2NH_2 \rightarrow HOCH_2CH_2NH_3^+ + HS^-$. The H_2S is a Lewis acid, and $HOCH_2CH_2NH_2$ is a Lewis base. The hydrogen ion from H_2S accepts a pair of electrons from the N atom in $HOCH_2CH_2NH_2$.

16.38 The equation is $CaO + SO_2 \rightarrow CaSO_3$. The CaO is a Lewis base, and SO_2 is a Lewis acid. The oxide ion from CaO donates a pair of electrons to the sulfur of SO_2, forming the SO_3^{2-} ion in $CaSO_3$.

16.39 The reaction is $HSO_4^- + ClO^- \rightarrow HClO + SO_4^{2-}$. According to Table 16.2, HClO is a weaker acid than HSO_4^-. Because the equilibrium for this type of reaction favors formation of the weaker acid (or weaker base), the reaction occurs to a significant extent.

16.40 The reaction is $HCN + SO_4^{2-} \rightarrow CN^- + HSO_4^-$. According to Table 16.2, HCN is a weaker acid than HSO_4^-. Because the equilibrium for this type of reaction favors formation of the weaker acid (or weaker base), the reaction occurs in the opposite direction.

16.41 a. NH_4^+ is a weaker acid than H_3PO_4, so the left-hand species are favored at equilibrium.

b. HCN is a weaker acid than H_2S, so the left-hand species are favored at equilibrium.

c. H_2O is a weaker acid than HCO_3^-, so the right-hand species are favored at equilibrium.

d. H_2O is a weaker acid than $Al(H_2O)_6^{3+}$, so the right-hand species are favored at equilibrium.

16.42 a. HCO_3^- is a weaker acid than NH_4^+, so the right-hand species are favored at equilibrium.

b. H_2S is a weaker acid than H_2CO_3, so the left-hand species are favored at equilibrium.

c. H_2O is a weaker acid than HCN, so the left-hand species are favored at equilibrium.

d. HCN is a weaker acid than H_2CO_3, so the right-hand species are favored at equilibrium.

16.43   Trichloroacetic acid is the stronger acid because, in general, the equilibrium favors the formation of the weaker acid, which is formic acid in this case.

16.44   The BF_4^- ion is the weaker base because the equilibrium favors the formation of the weaker base.

16.45 a. H_2S is stronger because acid strength decreases with increasing anion charge for polyprotic acid species.

b. H_2SO_3 is stronger because, for a series of oxoacids, acid strength increases with increasing electronegativity.

c. HBr is stronger because Br is more electronegative than Se. Within a period, acid strength increases as electronegativity increases.

d. HIO_4 is stronger because acid strength increases with the number of oxygen atoms bonded to the central atom.

e. H_2S is stronger because, within a group, acid strength increases with the increasing size of the central atom in binary acids.

16.46 a. HNO_2 < HNO_3: Acid strength increases with the number of oxygen atoms bonded to the central atom.

b. HCO_3^- < H_2CO_3: Acid strength decreases with increasing anion charge for polyprotic acid species.

c. H_2S < H_2Te: Within a group, acid strength increases with increasing size of the central atom for binary acids.

d. H_2S < HCl: Within a period, acid strength increases with increasing electronegativity of the central atom for binary acids.

e. H_3AsO_4 < H_3PO_4: For a series of oxoacids, acid strength increases with increasing electronegativity.

16.47 a. $[H_3O^+]$ = 1.2 M

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{1.2} = 8.\underline{3}3 \times 10^{-15} = 8.3 \times 10^{-15} \text{ M}$$

b. $[OH^-]$ = 0.32 M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.32} = 3.\underline{1}2 \times 10^{-14} = 3.1 \times 10^{-14} \text{ M}$$

c. $[OH^-]$ = 2 x (0.085 M) = 0.170 M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.170} = 5.\underline{8}8 \times 10^{-14} = 5.9 \times 10^{-14} \text{ M}$$

d. $[H_3O^+]$ = 0.38 M

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{0.38} = 2.\underline{6}3 \times 10^{-14} = 2.6 \times 10^{-14} \text{ M}$$

16.48 a. $[OH^-]$ = 1.65 M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{1.65} = 6.0\underline{6}0 \times 10^{-15} = 6.06 \times 10^{-15} \text{ M}$$

b. $[OH^-]$ = 2 x (0.35 M) = 0.70 M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.70} = 1.\underline{4}2 \times 10^{-14} = 1.4 \times 10^{-14} \text{ M}$$

c. $[H_3O^+]$ = 0.045 M

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{0.045} = 2.\underline{2}2 \times 10^{-13} = 2.2 \times 10^{-13} \text{ M}$$

d. $[H_3O^+]$ = 0.58 M

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.00 \times 10^{-14}}{0.58} = 1.\underline{7}2 \times 10^{-14} = 1.7 \times 10^{-14} \text{ M}$$

16.49   The $[H_3O^+]$ = 0.050 M (HCl is a strong acid); using K_w, the $[OH^-]$ = 2.0 x 10^{-13} M.

16.50   The $[H_3O^+]$ = 0.020 M (HNO_3 is a strong acid); using K_w, the $[OH^-]$ = 5.0 x 10^{-13} M.

16.51 Because the $Ba(OH)_2$ forms two OH^- per formula unit, the $[OH^-] = 2 \times 0.0085 = 0.017$ M.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.017} = 5.\underline{8}8 \times 10^{-13} = 5.9 \times 10^{-13} \text{ M}$$

16.52 Because $Mg(OH)_2$ forms two OH^- per formula unit, the $[OH^-] = 2 \times 3.2 \times 10^{-4} = 6.4 \times 10^{-4}$ M.

$$[H_3O^+] = K_w \div [OH^-] = (1.0 \times 10^{-14}) \div (6.4 \times 10^{-4}) = 1.\underline{5}6 \times 10^{-11} = 1.6 \times 10^{-11} \text{ M}$$

16.53 a. 5×10^{-6} M $H_3O^+ > 1.0 \times 10^{-7}$, so the solution is acidic.

   b. Use K_w to determine $[H_3O^+]$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{5 \times 10^{-9}} = \underline{2}.0 \times 10^{-6} = 2 \times 10^{-6} \text{ M}$$

   Since 2×10^{-6} M $> 1.0 \times 10^{-7}$, the solution is acidic.

   c. When $[OH^-] = 1.0 \times 10^{-7}$ M, $[H_3O^+] = 1.0 \times 10^{-7}$ M, and the solution is neutral.

   d. 2×10^{-9} M $H_3O^+ < 1.0 \times 10^{-7}$, so the solution is basic.

16.54 a. Use K_w to determine $[H_3O^+]$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{2 \times 10^{-11}} = \underline{5}.0 \times 10^{-4} = 5 \times 10^{-4} \text{ M}$$

   Since 5×10^{-4} M $> 1.0 \times 10^{-7}$, the solution is acidic.

   b. 2×10^{-6} M $> 1.0 \times 10^{-7}$, so the solution is acidic.

   c. Use K_w to determine $[H_3O^+]$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{6 \times 10^{-10}} = \underline{1}.6 \times 10^{-5} = 2 \times 10^{-5} \text{ M}$$

   Since 2×10^{-5} M $> 1.0 \times 10^{-7}$, the solution is acidic.

   d. 6×10^{-3} M $> 1.0 \times 10^{-7}$ M, so the solution is acidic.

16.55   The [H_3O^+] calculated below is > 1.0×10^{-7} M, so the solution is acidic.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{1.5 \times 10^{-9}} = 6.\underline{6}6 \times 10^{-6} = 6.7 \times 10^{-6} \text{ M}$$

16.56   The [H_3O^+] calculated below is < 1.0×10^{-7} M, so the solution is basic.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{8.4 \times 10^{-5}} = 1.\underline{1}9 \times 10^{-10} = 1.2 \times 10^{-10} \text{ M}$$

16.57 a. pH 4.6, acidic solution   b. pH 7.0, neutral solution

c. pH 1.6, acidic solution   d. pH 10.5, basic solution

16.58 a. pH 12.9, basic solution   b. pH 8.9, basic solution

c. pH 5.1, acidic solution   d. pH 1.6, acidic solution

16.59 a. Acidic (3.5 < 7.0)   b. Neutral (7.0 = 7.0)

c. Basic (9.0 > 7.0)   d. Acidic (5.5 < 7.0)

16.60 a. Basic (12.0 > 7.0)   b. Neutral (7.0 = 7.0)

c. Acidic (4.0 < 7.0)   d. Acidic (5.7 < 7.0)

16.61   Record the same number of places after the decimal point in the pH as the number of significant figures in the [H_3O^+].

a. -log (1.0×10^{-8}) = 8.0$\underline{0}0$ = 8.00   b. -log (5.0×10^{-12}) = 11.3$\underline{0}1$ = 11.30

c. -log (7.5×10^{-3}) = 2.1$\underline{2}4$ = 2.12   d. -log (6.35×10^{-9}) = 8.19$\underline{7}2$ = 8.197

16.62   Record the same number of places after the decimal point in the pH as the number of significant figures in the [H_3O^+].

a. -log (1.0×10^{-4}) = 4.0$\underline{0}0$ = 4.00   b. -log (3.2×10^{-10}) = 9.4$\underline{9}4$ = 9.49

c. -log (2.3×10^{-5}) = 4.6$\underline{3}8$ = 4.64   d. -log (2.91×10^{-11}) = 10.53$\underline{6}1$ = 10.536

16.63   Record the same number of places after the decimal point in the pH as the number of significant figures in the [H_3O^+].

-log (7.5×10^{-3}) = 2.1$\underline{2}49$ = 2.12

16.64 Record the same number of places after the decimal point in the pH as the number of significant figures in the $[H_3O^+]$.

$$-\log (5.0 \times 10^{-3}) = 2.3\underline{0}1 = 2.30$$

16.65 a. pOH = $-\log (5.25 \times 10^{-9})$ = 8.27$\underline{9}$8; pH = 14.00 - 8.2798 = 5.7$\underline{2}$02 = 5.72

b. pOH = $-\log (8.3 \times 10^{-3})$ = 2.0$\underline{8}$09; pH = 14.00 - 2.0809 = 11.91$\underline{9}$08 = 11.92

c. pOH = $-\log (3.6 \times 10^{-12})$ = 11.44$\underline{3}$6; pH = 14.00 - 11.4436 = 2.5$\underline{5}$63 = 2.56

d. pOH = $-\log (2.1 \times 10^{-8})$ = 7.67$\underline{7}$7; pH = 14.00 - 7.6777 = 6.3$\underline{2}$2 = 6.32

16.66 a. pOH = $-\log (4.83 \times 10^{-11})$ = 10.31$\underline{6}$05; pH = 14.00 - 10.31605 = 3.6$\underline{8}$39 = 3.68

b. pOH = $-\log (3.2 \times 10^{-5})$ = 4.4$\underline{9}$4; pH = 14.00 - 4.494 = 9.5$\underline{0}$51 = 9.51

c. pOH = $-\log (2.7 \times 10^{-10})$ = 9.5$\underline{6}$8; pH = 14.00 - 9.568 = 4.4$\underline{3}$1 = 4.43

d. pOH = $-\log (5.0 \times 10^{-4})$ = 3.3$\underline{0}$1; pH = 14.00 - 3.301 = 10.6$\underline{9}$8 = 10.70

16.67 First, convert the $[OH^-]$ to $[H_3O^+]$ using the K_w equation. Then, find the pH, recording the same number of places after the decimal point in the pH as the number of significant figures in the $[H_3O^+]$.

$$[H_3O^+] = K_w \div [OH^-] = (1.0 \times 10^{-14}) \div (0.0040) = 2.\underline{5}0 \times 10^{-12} \text{ M}$$

$$pH = -\log (2.50 \times 10^{-12}) = 11.6\underline{0}2 = 11.60$$

16.68 First, convert the $[OH^-]$ to $[H_3O^+]$ using the K_w equation. Then, find the pH, recording the same number of places after the decimal point in the pH as the number of significant figures in the $[H_3O^+]$.

$$[H_3O^+] = K_w \div [OH^-] = (1.0 \times 10^{-14}) \div (0.050) = 2.\underline{0}0 \times 10^{-13} \text{ M}$$

$$pH = -\log (2.00 \times 10^{-13}) = 12.6\underline{9}8 = 12.70$$

16.69 From the definition, pH = $-\log [H_3O^+]$ and $-pH = \log [H_3O^+]$, so enter the negative value of the pH on the calculator, and use the inverse and log keys (or 10^x) key to find the antilog of -pH:

$$\log [H_3O^+] = -pH = -5.12$$

$$[H_3O^+] = \text{antilog} (-5.12) = 10^{-5.12} = 7.\underline{5}8 \times 10^{-6} = 7.6 \times 10^{-6} \text{ M}$$

16.70   From the definition, pH = -log [H$_3$O$^+$] and -pH = log [H$_3$O$^+$], so enter the negative value of the pH on the calculator, and use the inverse and log keys (or 10^x) key to find the antilog of -pH:

log [H$_3$O$^+$] = -pH = -3.85

[H$_3$O$^+$] = antilog (-3.85) = $10^{-3.85}$ = 1.$\underline{4}$1 x 10^{-4} = 1.4 x 10^{-4} M

16.71   From the definition, pH = -log [H$_3$O$^+$] and -pH = log [H$_3$O$^+$], so enter the negative value of the pH on the calculator, and use the inverse and log keys (or 10^x) key to find the antilog of -pH. Then, use the K$_w$ equation to calculate [OH$^-$] from [H$_3$O$^+$].

log [H$_3$O$^+$] = -pH = -11.63

[H$_3$O$^+$] = antilog (-11.63) = $10^{-11.63}$ = 2.$\underline{3}$4 x 10^{-12} M

[OH$^-$] = K$_w$ ÷ [H$_3$O$^+$] = (1.0 x 10^{-14}) ÷ (2.34 x 10^{-12}) = 4.$\underline{2}$7 x 10^{-3}

= 4.3 x 10^{-3} M

16.72   From the definition, pH = -log [H$_3$O$^+$] and -pH = log [H$_3$O$^+$], so enter the negative value of the pH on the calculator, and use the inverse and log keys (or 10^x) key to find the antilog of -pH. Then, use the K$_w$ equation to calculate [OH$^-$] from [H$_3$O$^+$].

log [H$_3$O$^+$] = -pH = -9.61

[H$_3$O$^+$] = antilog (-9.61) = $10^{-9.61}$ = 2.$\underline{4}$5 x 10^{-10} M

[OH$^-$] = K$_w$ ÷ [H$_3$O$^+$] = (1.0 x 10^{-14}) ÷ (2.45 x 10^{-10}) = 4.$\underline{0}$8 x 10^{-5}

= 4.1 x 10^{-5} M

16.73   First, calculate the molarity of the OH$^-$ ion from the mass of NaOH. Then, convert the [OH$^-$] to [H$_3$O$^+$] using the K$_w$ equation. Then, find the pH, recording the same number of places after the decimal point in the pH as the number of significant figures in the [H$_3$O$^+$].

$$\frac{5.80 \text{ g NaOH}}{1.00 \text{ L}} \quad \text{x} \quad \frac{1 \text{ mol NaOH}}{40.01 \text{ g NaOH}} \quad = \quad \frac{0.1450 \text{ mol NaOH}}{1.00 \text{ L}} \quad = \quad 0.14\underline{5}0 \text{ M OH}^-$$

[H$_3$O$^+$] = K$_w$ ÷ [OH$^-$] = (1.0 x 10^{-14}) ÷ (0.1450) = 6.$\underline{8}$96 x 10^{-14} M

pH = -log [H$_3$O$^+$] = -log (6.896 x 10^{-14}) = 13.1$\underline{6}$14 = 13.16

16.74 First, calculate the molarity of the $Ba(OH)_2$ from the mass of $Ba(OH)_2$, and multiply the molarity of $Ba(OH)_2$ by two to obtain $[OH^-]$. Then, convert the $[OH^-]$ to $[H_3O^+]$ using the K_w equation. Then, find the pH, recording the same number of places after the decimal point in the pH as the number of significant figures in the $[H_3O^+]$.

$$\frac{6.78 \text{ g Ba(OH)}_2}{1.00 \text{ L}} \times \frac{1 \text{ mol Ba(OH)}_2}{171.4 \text{ g Ba(OH)}_2} = \frac{0.039556 \text{ mol Ba(OH)}_2}{1.00 \text{ L}}$$
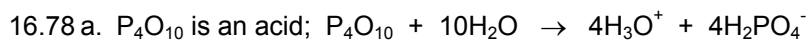
$[OH^-]$ = 2 mol OH^-/1 mol $Ba(OH)_2$ = 2 x 0.039556 M $Ba(OH)_2$ = 0.079$\underline{1}$1 M

$[H_3O^+]$ = K_w ÷ $[OH^-]$ = (1.0×10^{-14}) ÷ (0.07912) = 1.$\underline{2}$64 x 10^{-13} M

pH = -log $[H_3O^+]$ = -log (1.264×10^{-13}) = 12.8$\underline{9}$82 = 12.90

16.75 Figure 16.10 shows that the methyl-red indicator is yellow at pH values above about 5.5 (slightly past the midpoint of the range for methyl red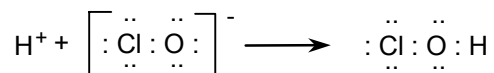). Bromthymol blue is yellow at pH values up to about 6.5 (slightly below the midpoint of the range for bromthymol blue). Therefore, the pH of the solution is between 5.5 and 6.5, and the solution is acidic.

16.76 Thymol blue is yellow at pH values above about 2.5. Bromphenol blue is yellow at pH values below about 3.5. Therefore, the pH of the aspirin solution must be in the range of 2.5 to 3.5. The solution is, of course, acidic.

◼ Solutions to General Problems

16.77 a. BaO is a base; $BaO + H_2O \rightarrow Ba^{2+} + 2OH^-$

b. H_2S is an acid; $H_2S + H_2O \rightarrow H_3O^+ + HS^-$

c. CH_3NH_2 is a base; $CH_3NH_2 + H_2O \rightarrow CH_3NH_3^+ + OH^-$

d. SO_2 is an acid; $SO_2 + 2H_2O \rightarrow H_3O^+ + HSO_3^-$

16.78 a. P_4O_{10} is an acid; $P_4O_{10} + 10H_2O \rightarrow 4H_3O^+ + 4H_2PO_4^-$

b. K_2O is a base; $K_2O + H_2O \rightarrow 2K^+ + 2OH^-$

c. N_2H_4 is a base; $N_2H_4 + H_2O \rightarrow N_2H_5^+ + OH^-$

d. H_2Se is an acid; $H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$

16.79 a. $H_2O_2(aq) + S^{2-}(aq) \rightarrow HO_2^-(aq) + HS^-(aq)$

b. $HCO_3^-(aq) + OH^-(aq) \rightarrow CO_3^{2-}(aq) + H_2O(l)$

c. $NH_4^+(aq) + CN^-(aq) \rightarrow NH_3(aq) + HCN(aq)$

d. $H_2PO_4^-(aq) + OH^-(aq) \rightarrow HPO_4^{2-}(aq) + H_2O(l)$

16.80 a. $H_2O(l) + HCl(aq) \rightarrow H_3O^+(aq) + Cl^-(aq)$

b. $HCO_3^-(aq) + HF(aq) \rightarrow H_2CO_3(aq) + F^-(aq)$

c. $NH_3(aq) + HBrO(aq) \rightarrow NH_4^+(aq) + BrO^-(aq)$
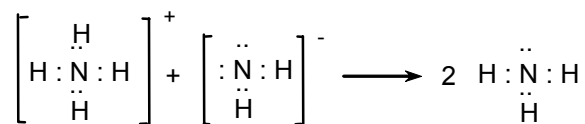
d. $H_2PO_4^-(aq) + H_2SO_3(aq) \rightarrow H_3PO_4(aq) + HSO_3^-(aq)$

16.81 a. The ClO^- ion is a Bronsted base, and water is a Bronsted acid. The complete chemical equation is $ClO^-(aq) + H_2O(l) \rightleftharpoons HClO(aq) + OH^-(aq)$. The equilibrium does not favor the products because ClO^- is a weaker base than OH^-. In Lewis language, a proton from H_2O acts as a Lewis acid by accepting a pair of electrons on the oxygen of ClO^-.

$$H^+ + \left[\ :\overset{..}{\underset{..}{Cl}}:\overset{..}{\underset{..}{O}}:\ \right]^- \longrightarrow :\overset{..}{\underset{..}{Cl}}:\overset{..}{\underset{..}{O}}:H$$

b. The NH_2^- ion is a Bronsted base, and NH_4^+ is a Bronsted acid. The complete chemical equation is $NH_4^+ + NH_2^- \rightleftharpoons 2NH_3$. The equilibrium favors the products because the reactants form the solvent, a weakly ionized molecule. In Lewis language, the proton from NH_4^+ acts as a Lewis acid by accepting a pair of electrons on the nitrogen of NH_2^-.

$$\left[\ H:\overset{\overset{\displaystyle H}{\displaystyle .}}{\underset{\underset{\displaystyle H}{\displaystyle ..}}{N}}:H\ \right]^+ + \left[\ :\overset{..}{\underset{\underset{\displaystyle H}{\displaystyle ..}}{N}}:H\ \right]^- \longrightarrow 2\ H:\overset{..}{\underset{\underset{\displaystyle H}{\displaystyle ..}}{N}}:H$$

16.82 a. The HS^- ion is a Bronsted base, and water is a Bronsted acid. The complete chemical equation is $HS^-(aq) + H_2O(l) \rightleftharpoons H_2S(aq) + OH^-(aq)$. The equilibrium does not favor the products because the HS^- ion is a weaker base than OH^-. In Lewis language, a proton from H_2O acts as a Lewis acid by accepting a pair of electrons on the sulfur of HS^-.

$$H:\overset{..}{\underset{..}{O}}:H + \left[\ :\overset{..}{\underset{..}{S}}:H\ \right]^- \longrightarrow H:\overset{..}{\underset{..}{S}}:H + OH^-$$
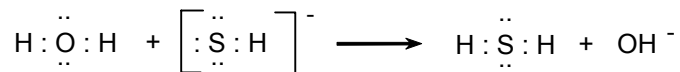
(continued)

b. The complete chemical equation is $Fe^{3+}(aq) + CN^-(aq) \rightarrow Fe(CN)^{2+}$. The reaction cannot be described in Bronsted language because no proton transfer occurs. In Lewis language, Fe^{3+} acts as a Lewis acid by sharing a pair of electrons on the carbon of CN^-.

$$Fe^{3+} + [:C \equiv N:]^- \longrightarrow [:Fe:C \equiv N:]^{2+}$$
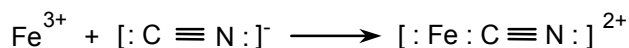
16.83 Table 16.2 shows that HNO_2 is a stronger acid than HF. Because an acid-base reaction normally goes in the direction of the weaker acid, the reaction is more likely to go in the direction written:

$$HNO_2 + F^- \rightleftharpoons HF + NO_2^-$$

16.84 The HS^- ion is acting as a base and will form its conjugate acid, H_2S. Table 16.2 shows that H_2S is a stronger acid than HCN. Because an acid-base reaction normally goes in the direction of the weaker acid, the reaction is more likely to go in the opposite direction:

$$H_2S + CN^- \rightleftharpoons HS^- + HCN$$

16.85 The order is $H_2S < H_2Se < HBr$. H_2Se is stronger than H_2S because, within a group, acid strength increases with increasing size of the central atom in binary acids. HBr is a strong acid, whereas the others are weak acids.

16.86 The order is $HBrO < HBrO_2 < HClO_2$. $HClO_2$ is stronger than $HBrO_2$ because, for oxoacids, acid strength increases with increasing electronegativity. $HBrO_2$ is stronger than HBrO because acid strength increases with increasing oxidation number of Br.

16.87 The KOH is a strong base and is fully ionized in solution, so you can use its formula and molar concentration to determine the $[OH^-]$ of the solution. Therefore, the 0.25 M KOH contains 0.25 M OH^-. The $[H_3O^+]$ is obtained from the K_w expression:

$$K_w = 1.0 \times 10^{-14} = [H_3O^+] \times 0.25 \text{ M } OH^-$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{0.25} = 4.\underline{0}0 \times 10^{-14} = 4.0 \times 10^{-14} \text{ M}$$

16.88   The $Sr(OH)_2$ is a strong base and is fully ionized in solution, so you can use its formula and molar concentration to determine the [OH⁻] of the solution. Therefore, the 0.35 M $Sr(OH)_2$ contains 2 x 0.35 = 0.70 M OH⁻. The [H_3O^+] is obtained from the K_w expression:

$$K_w = 1.0 \times 10^{-14} = [H_3O^+] \times 0.70 \text{ M OH}^-$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{0.70} = 1.\underline{4}2 \times 10^{-14} = 1.4 \times 10^{-14} \text{ M}$$

16.89   Enter the H_3O^+ concentration of 1.5 x 10⁻³ into the calculator, press the log key, and press the sign key to change the negative log to a positive log. This follows the negative log definition of pH. The number of decimal places of the pH should equal the significant figures in the H_3O^+.

$$pH = -\log [H_3O^+] = -\log (1.5 \times 10^{-3}) = 2.8\underline{2}3 = 2.82$$

16.90   Enter the H_3O^+ concentration of 2.5 x 10⁻² into the calculator, press the log key, and press the sign key to change the negative log to a positive log. This follows the negative log definition of pH. The number of decimal places of the pH should equal the significant figures in the H_3O^+.

$$pH = -\log [H_3O^+] = -\log (2.5 \times 10^{-2}) = 1.6\underline{0}2 = 1.60$$

16.91   Find the pOH from the pH using pH + pOH = 14.00. Then, calculate the [OH⁻] from the pOH by entering the pOH into the calculator, pressing the sign key to change the positive log to a negative log, and finding the antilog. On some calculators, the antilog is found by using the inverse of the log; on other calculators, the antilog is found using the 10^x key. The number of significant figures in the [OH⁻] should equal the number of decimal places in the pOH.
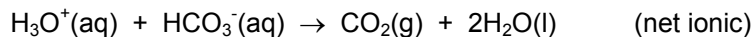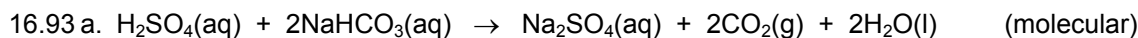
$$pOH = 14.00 - 3.15 = 10.85$$

$$[OH^-] = \text{antilog} (-10.85) = 10^{-10.85} = 1.\underline{4}2 \times 10^{-11} = 1.4 \times 10^{-11} \text{ M}$$

16.92   Find the pOH from the pH using pH + pOH = 14.00. Then, calculate the [OH⁻] from the pOH by entering the pOH into the calculator, pressing the sign key to change the positive log to a negative log, and finding the antilog. On some calculators, the antilog is found by using the inverse of the log; on other calculators, the antilog is found using the 10^x key. The number of significant figures in the [OH⁻] should equal the number of decimal places in the pOH.

$$pOH = 14.00 - 4.05 = 9.95$$

$$[OH^-] = \text{antilog} (-9.95) = 10^{-9.95} = 1.\underline{1}2 \times 10^{-10} = 1.1 \times 10^{-10} \text{ M}$$

16.93 a. $H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$ (molecular)

$H_3O^+(aq) + HCO_3^-(aq) \rightarrow CO_2(g) + 2H_2O(l)$ (net ionic)

b. The total moles of H_3O^+ from the H_2SO_4 is

$$\text{mol } H_3O^+ = \frac{0.437 \text{ mol } H_2SO_4}{1 \text{ L}} \times 0.02500 \text{ L} \times \frac{2 \text{ mol } H_3O^+}{1 \text{ mol } H_2SO_4}$$

$$= 0.02185 \text{ mol } H_3O^+$$

The moles of H_3O^+ that reacted with the NaOH are given by

$$\text{mol } H_3O^+ = \frac{0.108 \text{ mol NaOH}}{1 \text{ L}} \times 0.0287 \text{ L} = 0.003100 \text{ mol}$$

The moles of $NaHCO_3$ present in the original sample are equal to the moles of H_3O^+ that reacted with the HCO_3^-, which is given by

Total moles H_3O^+ - moles H_3O^+ reacted with the NaOH = moles HCO_3^-

0.02185 mol - 0.003823 mol = 0.01803 = 0.0180 mol $NaHCO_3$

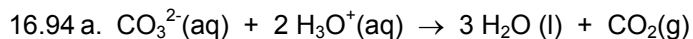c. The mass of $NaHCO_3$ (molar mass 84.01 g/mol) present in the original sample is

$$0.01803 \text{ mol } NaHCO_3 \times \frac{84.01 \text{ g } NaHCO_3}{1 \text{ mol } NaHCO_3} = 1.514 \text{ g}$$

Thus, the percent $NaHCO_3$ in the original sample is given by

$$\text{Percent } NaHCO_3 = \frac{1.514 \text{ g}}{2.500 \text{ g}} \times 100\% = 60.56 = 60.6 \text{ percent}$$

The percent KCl in the original sample is

Percent KCl = 100 - 60.56 = 39.44 = 39.4 percent

16.94 a. $CO_3^{2-}(aq) + 2 H_3O^+(aq) \rightarrow 3 H_2O(l) + CO_2(g)$

b. The total moles of H_3O^+ from the HCl is

$$\text{mol } H_3O^+ = \frac{0.798 \text{ mol HCl}}{1 \text{ L}} \times 0.02500 \text{ L} = 0.01995 \text{ mol}$$

The moles of H_3O^+ that reacted with the NaOH is given by

$$\text{mol } H_3O^+ = \frac{0.108 \text{ mol NaOH}}{1 \text{ L}} \times 0.0287 \text{ L} = 0.003100 \text{ mol}$$

The moles of Na_2CO_3 present in the original sample are equal to the moles of H_3O^+ that reacted with the CO_3^{2-}, which is given by

Total moles H_3O^+ - moles H_3O^+ reacted with the NaOH = moles CO_3^{2-}

$$0.01995 \text{ mol } - 0.003100 \text{ mol } = 0.01685 = 0.0169 \text{ mol } Na_2CO_3$$

c. The mass of Na_2CO_3 (molar mass 105.99 g/mol) present in the original sample is

$$0.01685 \text{ mol } Na_2CO_3 \times \frac{105.99 \text{ g } Na_2CO_3}{1 \text{ mol } Na_2CO_3} = 1.785 \text{ g}$$

Thus, the percent Na_2CO_3 in the original sample is given by
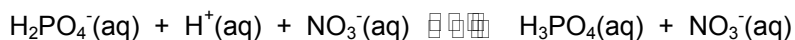
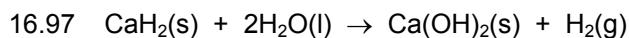$$\text{Percent } Na_2CO_3 = \frac{1.785 \text{ g}}{2.500 \text{ g}} \times 100\% = 71.40 = 71.4 \text{ percent}$$

The percent NaCl in the original sample is

Percent NaCl = 100 - 71.40 = 28.60 = 28.6 percent

16.95 $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$

$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

$HCO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightleftharpoons Na^+(aq) + CO_3^{2-}(aq) + H_2O(l)$

$HCO_3^-(aq) + H^+(aq) + Cl^-(aq) \rightleftharpoons H_2O(l) + CO_2(g) + Cl^-(aq)$

16.96 $H_2PO_4^-$ is amphiprotic or amphoteric.

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

$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq)$

$H_2PO_4^-(aq) + K^+(aq) + OH^-(aq) \rightleftharpoons K^+(aq) + HPO_4^{2-}(aq) + H_2O(l)$

$H_2PO_4^-(aq) + H^+(aq) + NO_3^-(aq) \rightleftharpoons H_3PO_4(aq) + NO_3^-(aq)$

16.97 $CaH_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$

The hydride ion is a stronger base because it took an H^+ from water, leaving the OH^- ion. Every time a strong base is added to water, it will react with the water leaving the OH^- as the product, so a strong base cannot exist in water.

16.98 $Ca_3N_2(s) + 6H_2O(l) \rightarrow 3Ca(OH)_2(s) + 2NH_3(g)$

$NaNH_2(s) + H_2O(l) \rightarrow NaOH(aq) + NH_3(g)$
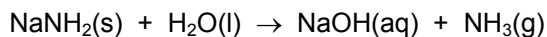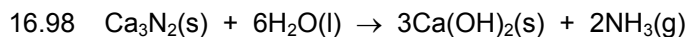
N^{3-} is a stronger base than NH_2^- because it has more negative charge, so it will have a greater attraction for H^+. Also, we could consider that NH_2^- is a N^{3-} that has already reacted with two H^+'s. Stronger bases than OH^- will produce OH^- in aqueous solution.
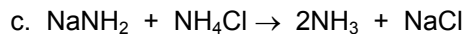
16.99 a. $2HF(l) \rightleftharpoons H_2F^+ + F^-$

b. NaF will be a base because F^- is a conjugate base of HF.

c. $HClO_4 + HF \rightarrow H_2F^+ + ClO_4^-$
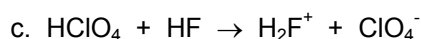
The conjugate acid is H_2F^+.

16.100 a. $2NH_3(l) \rightleftharpoons NH_4^+ + NH_2^-$

b. NH_2^- is the conjugate base of NH_3, so $NaNH_2$ will be a base.

c. $NaNH_2 + NH_4Cl \rightarrow 2NH_3 + NaCl$

16.101 The reaction of ammonia with water is given by

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

The initial concentration of NH_3 (molar mass 17.03 g/mol) is

$$\text{Molarity} = \frac{4.25 \text{ g } NH_3 \times \dfrac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3}}{0.2500 \text{ L}} = 0.9982 \text{ M}$$

Since the NH_3 is 0.42 percent reacted, the concentration of OH^- is

$[OH^-] = 0.9982 \text{ M} \times 0.0042 = 0.00419 \text{ M}$

$pOH = -\log [OH^-] = -\log (0.00419) = 2.378$

$pH = 14 - pOH = 14 - 2.378 = 11.622 = 11.62$

16.102 The reaction of $C_2H_5NH_2$ with water is given by

$$C_2H_5NH_2(aq) \ + \ H_2O(l) \ \rightleftharpoons \ C_2H_5NH_3^+ (aq) \ + \ OH^-(aq)$$

The initial concentration of $C_2H_5NH_2$ (molar mass 45.09 g/mol) is given by

$$\text{Molarity} \ = \ \frac{0.675 \text{ g } C_2H_5NH_2 \ \times \ \dfrac{1 \text{ mol } C_2H_5NH_2}{45.09 \text{ g } C_2H_5NH_2}}{0.1000 \text{ L}} \ = \ 0.14\underline{9}7 \text{ M}$$

Since the $C_2H_5NH_2$ is 0.98 percent reacted, the concentration of OH^- is

$[OH^-] = 0.1497 \text{ M} \times 0.0098 = 0.001\underline{4}7 \text{ M}$

$pOH = - \log [OH^-] \ = \ - \log (0.00147) \ = \ 2.8\underline{3}3$

$pH = 14 \ - \ pOH \ = \ 14 \ - \ 2.8\underline{3}3 \ = \ 11.1\underline{6}7 \ = \ 11.17$

■ Solutions to Cumulative-Skills Problems

16.103 For $(HO)_mYO_n$ acids, acid strength increases with n regardless of the number of OH's. The structure of H_3PO_4 is $(HO)_3PO$; because H_3PO_3 and H_3PO_4 have about the same acidity, H_3PO_3 must also have n = 1; thus, m = 2. This leaves one H, which must bond to phosphorus, giving a structure of $(HO)_2(O)PH$. Assuming that only two H's react with NaOH, the mass of NaOH that reacts with 1.00 g of H_3PO_3 (PA) is calculated as follows:

$$1.00 \text{ g PA} \ \times \ \frac{1 \text{ mol PA}}{81.994 \text{ g PA}} \ \times \ \frac{2 \text{ mol NaOH}}{1 \text{ mol PA}} \ \times \ \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}}$$

$$= \ 0.97\underline{5}6 \ = \ 0.976 \text{ g NaOH}$$

16.104 For $(HO)_mYO_n$ acids, acid strength increases with n regardless of the number of OH's. The structure of H_3PO_4 is $(HO)_3PO$; because H_3PO_2 and H_3PO_4 have the same acidity, H_3PO_2 must also have n = 1; thus, m = 1. This leaves two H's, which must bond to phosphorus, giving a structure of $(HO)(O)PH_2$. Assuming that only one H reacts with NaOH, the mass of NaOH that reacts with 1.00 g of H_3PO_2 (HA) is calculated as follows:

$$1.00 \text{ g HA} \ \times \ \frac{1 \text{ mol PA}}{65.994 \text{ g PA}} \ \times \ \frac{1 \text{ mol NaOH}}{1 \text{ mol HA}} \ \times \ \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}}$$

$$= \ 0.60\underline{6}1 \ = \ 0.606 \text{ g NaOH}$$

16.105 BF_3 acts as a Lewis acid, accepting an electron pair from NH_3:

$$BF_3 \ + \ :NH_3 \ \rightarrow F_3B:NH_3$$
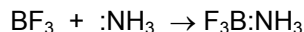
The NH_3 acts as a Lewis base in donating an electron pair to BF_3. When 10.0 g of each are mixed, the BF_3 is the limiting reagent because it has the higher formula weight. The mass of $BF_3:NH_3$ formed is

$$10.0 \text{ g } BF_3 \ \times \ \frac{1 \text{ mol } BF_3}{67.81 \text{ g } BF_3} \ \times \ \frac{1 \text{ mol } BF_3:NH_3}{1 \text{ mol } BF_3} \ \times \ \frac{84.84 \text{ g } BF_3:NH_3}{1 \text{ mol } BF_3:NH_3}$$

$$= \ 12.\underline{5}1 \ = \ 12.5 \text{ g } BF_3:NH_3$$

16.106 BF_3 is the Lewis acid and accepts an electron pair from ether:

$$BF_3 \ + \ :OR_2 \ \rightarrow F_3B:OR_2$$
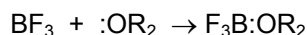
The ether ($:OR_2$) acts as a Lewis base in donating an electron pair to BF_3. When 10.0 g of BF_3 and 20.0 g of ether are mixed, the BF_3 is the limiting reagent because the formula weights are nearly equal. The mass of $BF_3:OR_2$ formed is

$$10.0 \text{ g } BF_3 \ \times \ \frac{1 \text{ mol } BF_3}{67.81 \text{ g } BF_3} \ \times \ \frac{1 \text{ mol } BF_3:OR_2}{1 \text{ mol } BF_3} \ \times \ \frac{141.93 \text{ g } BF_3:OR_2}{1 \text{ mol } BF_3:OR_2}$$

$$= \ 20.\underline{9}3 \ = \ 20.9 \text{ g } BF_3:OR_2$$