Chem101: General Chemistry
Chapter 9 - Acids and Bases

IV. ANSWERS AND SOLUTIONS TO ODD-NUMBERED PROBLEMS:

The Arrhenius Theory (Section 9.1)

9.1 Write the dissociation equations for the following that emphasize their behavior as Arrhenius acids:

a) HI　b) $HBrO$　c) HCN　d) $HClO_2$

SOLUTION:

a) $HI(aq) \rightarrow H^+(aq) + I^-(aq)$

b) $HBrO(aq) \rightarrow H^+(aq) + BrO^-(aq)$

c) $HCN(aq) \rightarrow H^+(aq) + CN^-(aq)$

d) $HClO_2(aq) \rightarrow H^+(aq) + ClO_2^-(aq)$

9.3 Each of the following produces a basic solution when dissolved in water. Identify those that behave as Arrhenius bases and write dissociation equations to illustrate that behavior.

a) $CsOH$　b) CH_3NH_2　c) NH_3　d) $Ca(OH)_2$

SOLUTION:

a) $CsOH(aq) \rightarrow Cs^+(aq) + OH^-(aq)$

b) not an Arrhenius base

c) not an Arrhenius base

d) $Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2\ OH^-(aq)$

The Brønsted Theory (Section 9.2)

9.5 Identify each Brønsted acid and base in the following equations. Note that the reactions are assumed to be reversible.

a) $HBr(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Br^-(aq)$

b) $H_2O(l) + N_3^-(aq) \rightleftharpoons HN_3(aq) + OH^-(aq)$

c) $H_2S(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HS^-(aq)$

d) $SO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HSO_3^-(aq) + OH^-(aq)$

e) $HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$

SOLUTION:

a) $HBr(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + Br^-(aq)$
acid base acid base

b) $H_2O(aq) + N_3^-(aq) \rightleftharpoons HN_3(aq) + OH^-(aq)$
acid base acid base

c) $H_2S(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + HS^-(aq)$
acid base acid base

d) $SO_3^{2-}(aq) + H_2O(aq) \rightleftharpoons HSO_3^-(aq) + OH^-(aq)$
base acid acid base

e) $HCN(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$
acid base acid base

9.7 Identify each conjugate acid-base pair in the equations you wrote for Exercise 9.5

SOLUTION:

	Acid	Conjugate Base	Base	Conjugate Acid
a)	HBr	Br	H_2O	H_3O^+
b)	H_2O	OH^-	N_3^-	HN_3
c)	H_2S	HS^-	H_2O	H_3O^+
d)	H_2O	OH^-	SO_3^{2-}	HSO_3^-
e)	HCN	CN^-	H_2O	H_3O^+

9.9 Write a formula for the conjugate base formed when each of the following behaves as a Brønsted acid:
a) HSO_3^- b) HPO_4^{2-} c) $HClO_3$ d) $CH_3NH_3^+$ e) $H_2C_2O_4$

SOLUTION:
a) SO_3^{2-} b) PO_4^{3-} c) ClO_3^- d) CH_3NH_2 e) $HC_2O_4^-$

9.11 Write a formula for the conjugate acid formed when each of the following behaves as a Brønsted base:
a) NH_2^- b) CO_3^{2-} c) OH^- d) $(CH_3)_2NH$ e) NO_2^-

SOLUTION:
a) NH_3 b) HCO_3^- c) H_2O d) $(CH_3)_2NH_2^+$ e) HNO_2

9.13 The following reactions illustrate Brønsted acid-base behavior. Complete each equation.
a) $HI(aq) + ? \rightarrow H_3O^+(aq) + I^-(aq)$
b) $NH_3(l) + ? \rightarrow NH_4^+ + NH_2^-$
c) $H_2C_2O_4(aq) + H_2O(l) \rightarrow ? + HC_2O_4^-(aq)$
d) $H_2N_2O_2(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ?$
e) $? + H_2O(l) \rightarrow H_3O^+(aq) + CO_3^{2-}(aq)$

SOLUTION:

a) H_2O b) NH_3 c) H_3O^+ d) $HN_2O_2^-$ e) HCO_3^-

9.15 Write equations to illustrate the acid-base reaction of each of the following pairs of Brønsted acids and bases:

Acid	Base
a) $HClO$	H_2O
b) $HClO_4$	NH_3
c) H_2O	NH_2^-
d) H_2O	ClO^-
e) $HC_2O_4^-$	H_2O

SOLUTION:

a) $HClO + H_2O \rightleftharpoons H_3O^+ + ClO^-$

b) $HClO_4 + NH_3 \rightleftharpoons NH_4^+ + ClO_4^-$

c) $H_2O + NH_2^- \rightleftharpoons NH_3 + OH^-$

d) $H_2O + ClO^- \rightleftharpoons HClO + OH^-$

e) $HC_2O_4^- + H_2O \rightleftharpoons H_3O^+ + C_2O_4^{2-}$

Naming Acids (Section 9.3)

9.17 A water solution of HF gas is used to etch glass. Name the water solution as an acid.

SOLUTION:

The gas is hydrogen fluoride; the acid is hydrofluoric acid.

9.19 Name the following acids. Refer to Table 4.6 as needed.

a) $H_2Se(aq)$ b) $HClO_3$ c) H_2SO_4 d) HNO_3

SOLUTION:

a) hydroselenic acid b) chloric acid c) sulfuric acid d) nitric acid

9.21 The acid $H_3C_6H_5O_7$ forms the citrate ion, $C_6H_5O_7^{3-}$, when all three hydrogens are removed. This acid is involved in an important energy-storing process in the body. Name the acid.

SOLUTION:

citric acid

9.23 Refer to Table 4.6 and write the formula for chromic acid.

SOLUTION:

H_2CrO_4

The Self-Ionization of Water (Section 9.4)

9.25 Calculate the molar concentration of OH^- in water solutions with the following H_3O^+ molar concentrations:

a) 1.0×10^{-7} b) 3.2×10^{-3} c) 4.7×10^{-11} d) 1.2 e) 0.043

SOLUTION:

$$\text{a) } [OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7}\ M$$

$$\text{b) } [OH^-] = \frac{1.0 \times 10^{-14}}{3.2 \times 10^{-3}} = 3.1 \times 10^{-12}\ M$$

$$\text{c) } [OH^-] = \frac{1.0 \times 10^{-14}}{4.7 \times 10^{-11}} = 2.1 \times 10^{-4}\ M$$

$$\text{d) } [OH^-] = \frac{1.0 \times 10^{-14}}{1.2} = 8.3 \times 10^{-15}\ M$$

$$\text{e) } [OH^-] = \frac{1.0 \times 10^{-14}}{0.043} = 2.3 \times 10^{-13}\ M$$

9.27 Calculate the molar concentration of H_3O^+ in water solutions with the following OH^- molar concentrations:
a) 1.0×10^{-7} b) 5.2×10^{-4} c) 9.9×10^{-10} d) 0.092 e) 3.7

SOLUTION:

$$\text{a) } [H_3O^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7}\ M$$

$$\text{b) } [H_3O^+] = \frac{1.0 \times 10^{-14}}{5.2 \times 10^{-4}} = 1.9 \times 10^{-11}\ M$$

$$\text{c) } [H_3O^+] = \frac{1.0 \times 10^{-14}}{9.9 \times 10^{-10}} = 1.0 \times 10^{-5}\ M$$

$$\text{d) } [H_3O^+] = \frac{1.0 \times 10^{-14}}{9.2 \times 10^{-2}} = 1.1 \times 10^{-13}\ M$$

$$\text{e) } [H_3O^+] = \frac{1.0 \times 10^{-14}}{3.7} = 2.7 \times 10^{-15}\ M$$

9.29 Classify the solutions represented in Exercises 9.25 and 9.27 as *acidic, basic,* or *neutral.*

SOLUTION:

#9.25	a) neutral	b) acidic	c) basic	d) acidic	e) acidic
#9.27	a) neutral	b) basic	c) acidic	d) basic	e) basic

9.31 Explain what is wrong with each of the following statements. Water is the solvent.
a) a neutral solution is prepared in which $[H_3O^+] = 1.0 \times 10^{-4}$ and $[OH^-] = 1.0 \times 10^{-4}$
b) the $[H_3O^+]$ is measured for a basic solution and found to be 3.2×10^{-5} mol/L
c) an alkaline solution is prepared in which $[OH^-] = 1.4 \times 10^{-5}$ mol/L and $[H_3O^+] = 3.1 \times 10^{-10}$ mol/L

SOLUTION:
a) To be a neutral solution, $[H_3O^+] = [OH^-]$ and both must equal 1.0×10^{-7} M.
b) If the $[H_3O^+]$ is 3.2×10^{-5} M, the solution is acidic, not basic.
c) The product of the $[H_3O^+]$ times $[OH^-]$ does not equal 1.0×10^{-14}. One or both is wrong. If the $[OH^-]$ is accurate, the solution is basic as stated.

The pH Concept (Section 9.5)

9.33 Classify solutions with the following characteristics as *acidic, basic,* or *neutral.*
a) pH = 10 b) pH = 4 c) pH = 7.3 d) pH = 6

SOLUTION:
a) basic b) acidic c) basic d) acidic

9.35 Determine the pH of water solutions with the following characteristics. Classify each solution as *acidic, basic,* or *neutral.*
a) $[H^+] = 1.0 \times 10^{-5}$ M

b) $[OH^-] = 6.0 \times 10^{-3}$ M
c) $[H^+] = [OH^-]$
d) $[H^+] = 9.0 \times 10^{-4}$ M
e) $[OH^-] = 3.0 \times 10^{-9}$ M

SOLUTION:
Use the calculator to get pH. $pH = -\log[H^+]$
a) pH = 5.00; acidic
b) $[H^+] = \dfrac{1.0 \times 10^{-14}}{6.0 \times 10^{-3}} = 1.7 \times 10^{-12}$ M; pH = 11.78; basic
c) pH = 7.00; neutral
d) pH = 3.05; acidic
e) $[H^+] = \dfrac{1.0 \times 10^{-14}}{3.0 \times 10^{-9}} = 3.3 \times 10^{-6}$ M; pH = 5.48; acidic

9.37 Determine the pH of water solutions with the following characteristics. Classify each solution as *acidic, basic*, or *neutral.*
a) $[H^+] = 3.7 \times 10^{-4}$ M
b) $[H^+] = 7.4 \times 10^{-8}$ M
c) $[H^+] = 1.9 \times 10^{-10}$ M
d) $[OH^-] = 1.3 \times 10^{-1}$ M
e) $[OH^-] = 6.8 \times 10^{-7}$ M

SOLUTION:
Use the calculator to get pH. $pH = -\log[H^+]$
a) pH = 3.43; acidic
b) pH = 7.13; basic
c) pH = 9.72; basic

d) $[H^+] = \dfrac{1.0 \times 10^{-14}}{1.3 \times 10^{-1}} = 7.7 \times 10^{-14}$ M; pH = 13.11; basic

e) $[H^+] = \dfrac{1.0 \times 10^{-14}}{6.8 \times 10^{-7}} = 1.5 \times 10^{-8}$ M; pH = 7.83; basic

9.39 Determine the $[H^+]$ value for solutions with the following characteristics:
a) pH = 4.45 b) pH = 13.12 c) pH = 7.73

SOLUTION:
a) $[H^+] = 3.5 \times 10^{-5}$ M
b) $[H^+] = 7.6 \times 10^{-14}$ M
c) $[H^+] = 1.9 \times 10^{-8}$ M

9.41 Convert the following pH values into both $[H^+]$ and $[OH^-]$ values:
a) pH = 8.00 b) pH = 6.15 c) pH = 1.30

SOLUTION:

a) $[H^+] = 1.0 \times 10^{-8}$ M $[OH^-] = \dfrac{1.0 \times 10^{-14}}{1.0 \times 10^{-8}} = 1.0 \times 10^{-6}$ M

b) $[H^+] = 7.1 \times 10^{-7}$ M $[OH^-] = \dfrac{1.0 \times 10^{-14}}{7.1 \times 10^{-7}} = 1.4 \times 10^{-8}$ M

c) $[H^+] = 5.0 \times 10^{-2}$ M $[OH^-] = \dfrac{1.0 \times 10^{-14}}{5.0 \times 10^{-2}} = 2.0 \times 10^{-13}$ M

9.43 The pH values listed in Table 9.1 are generally the average values for the listed materials. Most natural materials such as body fluids and fruit juices have pH values that cover a range for different samples. Some measured pH values for specific body fluid samples are given below. Convert each one to $[H^+]$ and classify the fluid as *acidic, basic,* or *neutral*.

a) blood, pH = 7.41
b) gastric juice, pH = 1.60
c) urine, pH = 5.93
d) saliva, pH = 6.85
e) pancreatic juice, pH = 7.85

SOLUTION:

a) $[H^+] = 3.9 \times 10^{-8}$ M; basic
b) $[H^+] = 2.5 \times 10^{-2}$ M; acidic
c) $[H^+] = 1.2 \times 10^{-6}$ M; acidic
d) $[H^+] = 1.4 \times 10^{-7}$ M; acidic
e) $[H^+] = 1.4 \times 10^{-8}$ M; basic

9.45 The pH values of specific samples of food items are listed below. Convert each value to $[H^+]$ and classify the sample as *acidic, basic,* or *neutral*.

a) milk, pH = 6.39
b) coffee, pH = 5.10
c) orange juice, pH = 4.07
d) vinegar, pH = 2.65

SOLUTION:

a) $[H^+] = 4.1 \times 10^{-7}$ M; acidic
b) $[H^+] = 7.9 \times 10^{-6}$ M; acidic
c) $[H^+] = 8.5 \times 10^{-5}$ M; acidic
d) $[H^+] = 2.2 \times 10^{-3}$ M; acidic

Properties of Acids (Section 9.6)

9.47 Use the information in Table 9.4 and describe how you would prepare each of the following solutions.

a) about 750 mL of 0.5 M H_2SO_4 from dilute sulfuric acid
b) about 200 mL of 0.1 M NaOH from stock sodium hydroxide solution
c) about 1.0 L of 1.0 M acetic acid from glacial acetic acid

SOLUTION:

solve equation 7.9 for V_c. $V_c = V_d \times \dfrac{C_d}{C_c}$

a) $V_c = 750 \text{ mL} \times \dfrac{0.5 \cancel{M}}{3 \cancel{M}} = 125 \text{ mL}$

Measure 125 mL of the dilute (3 M) H_2SO_4 and put it in enough water (about 625 mL) to make 750 mL solution.

b) $V_c = 200 \text{ mL} \times \dfrac{0.1 \cancel{M}}{6 \cancel{M}} = 3 \text{ mL}$

Take 3 mL of stock NaOH and put it in enough water (about 197 mL) to make 200 mL solution.

c) $V_c = 1.0\,\cancel{L} \times \frac{1.0\,\cancel{M}}{18\,\cancel{M}} \times \frac{1000\text{ mL}}{1\,\cancel{L}} = 56\text{ mL}$
Measure 56 mL glacial acetic acid and put it in enough water (about 944 mL) to make 1.0 L solution.

9.49 Write balanced full equations to illustrate the following characteristics reactions of acids, using nitric acid (HNO_3).
a) reaction with water to form hydronium ions
b) reaction with the solid oxide CaO
c) reaction with the solid hydroxide, $Mg(OH)_2$
d) reaction with the solid carbonate, $CuCO_3$
e) reaction with the solid bicarbonate, $KHCO_3$
f) reaction with Mg metal

SOLUTION:
a) $HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$
b) $CaO(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$
c) $Mg(OH)_2(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$
d) $CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
e) $KHCO_3(s) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l) + CO_2(g)$
f) $Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$

9.51 Write each full equation of Exercise 9.49 in total ionic and net ionic form. Use Table 7.4 to decide which products will be soluble.

SOLUTION:
a) total ionic: $H^+(aq) + NO_3^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$
net ionic: $H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$
b) total ionic: $CaO(s) + 2H^+(aq) + 2NO_3^-(aq) \rightarrow Ca^{2+}(aq) + 2NO_3(aq) + H_2O(l)$
net ionic: $CaO(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l)$
c) total ionic: $Mg(OH)_2(s) + 2H^+(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + 2H_2O(l)$
net ionic: $Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
d) total ionic: $CuCO_3(s) + 2H^+(aq) + NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_3^-(aq) + H_2O(l) + CO_2(g)$
net ionic: $CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$
e) total ionic: $KHCO_3^-(s) + H^+(aq) + NO_3^-(aq) \rightarrow K^+(aq) + NO_3^-(aq) + H_2O(l) + CO_2(g)$
net ionic: $KHCO_3^-(s) + H^+(aq) \rightarrow H_2O(aq) + CO_2(g)$
f) total ionic: $Mg(s) + 2H^+(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + H_2(g)$
net ionic: $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

9.53 Write balanced full equations to illustrate five different reactions that could be used to prepare $BaCl_2$ from hydrochloric acid (HCl) and other appropriate substances.

SOLUTION:
a) $BaO(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l)$
b) $Ba(OH)_2(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$
c) $BaCO_3(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$
d) $Ba(HCO_3)_2(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l) + 2CO_2(g)$
e) $Ba(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$

9.55 Write balanced full, total ionic, and net ionic equations to illustrate each of the following reactions. All of the metals form 2+ ions.
a) zinc with H_2SO_4
b) magnesium with HCl
c) calcium with $HC_2H_3O_2$

SOLUTION:

a) full: $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
total: $Zn(s) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + H_2(g)$
net: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
b) full: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
total: $Mg(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq) + H_2(g)$
net: $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
c) full: $Ca(s) + 2HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + H_2(g)$
total: $Ca(s) + 2HC_2H_3O_2(aq) \rightarrow Ca^{2+}(aq) + 2C_2H_3O_2^-(aq) + H_2(g)$
net: $Ca(s) + 2HC_2H_3O_2(aq) \rightarrow Ca^{2+}(aq) + 2C_2H_3O_2^-(aq) + H_2(g)$

Properties of Bases (Section 9.7)

9.57 Write balanced full, total ionic, and net ionic equations to represent neutralization reactions between LiOH and the following acids. Use all H's possible for each acid.
a) HCl b) HNO_3 c) H_2SO_4

SOLUTION:

a) full: $HCl(aq) + LiOH(aq) \rightarrow LiCl(aq) + H_2O(l)$
total: $H^+(aq) + Cl^-(aq) + Li^+(aq) + OH^-(aq) \rightarrow Li^+(aq) + Cl^-(aq) + H_2O(l)$
net: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
b) full: $HNO_3(aq) + LiOH(aq) \rightarrow LiNO_3(aq) + H_2O(l)$
total: $H^+(aq) + NO_3^-(aq) + Li^+(aq) + OH^-(aq) \rightarrow Li^+(aq) + NO_3^-(aq) + H_2O(l)$
net: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
c) full: $H_2SO_4(aq) + 2LiOH(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(l)$
total: $2H^+(aq) + SO_4^{2-}(aq) + 2Li^+(aq) + 2OH^-(aq) \rightarrow 2Li^+(aq) + SO_4^{2-}(aq) + H_2O(l)$
net: $2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$ or $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

9.59 Some polyprotic acids can form more than one salt depending on the number of H's that react with base. Write balanced full, total ionic, and net ionic equations to represent the following neutralization reactions between KOH and:
a) H_2SO_4 (only react one H) b) H_2SO_4 (react both H's) c) H_3PO_4 (only react one H)

SOLUTION:

a) full: $H_2SO_4(aq) + KOH(aq) \rightarrow KHSO_4(aq) + H_2O(l)$
total: $H^+(aq) + HSO_4^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + HSO_4^-(aq) + H_2O(l)$
net: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
b) full: $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
total: $2H^+(aq) + SO_4^{2-}(aq) + 2K^+(aq) + 2OH^-(aq) \rightarrow 2K^+(aq) + SO_4^{2-}(aq) + H_2O(l)$
net: $2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$ or $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
c) full: $H_3PO_4(aq) + KOH(aq) \rightarrow KH_2PO_4(aq) + H_2O(l)$
total: $H^+(aq) + H_2PO_4^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + H_2PO_4^-(aq) + H_2O(l)$
net: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Salts (Section 9.8)

9.61 Identify with ionic formulas the cations and anions of the following salts:
a) LiCl b) $Cu(NO_3)_2$ c) $SrSO_4$
d) K_3PO_4 e) K_2HPO_4 f) $CaCO_3$

SOLUTION:

a) Li^+, Cl^- b) Cu^{2+}, $2NO_3^-$ c) Sr^{2+}, SO_4^{2-}

d) $3K^+$, PO_4^{3-} e) $2K^+$, HPO_4^{2-} f) Ca^{2+}, CO_3^{2-}

9.63 Identify with formulas the acid and base from which the anion and cation of each salt in Exercise 9.61 was derived. Pay special attention to salts derived from polyprotic acids and be sure to list the acid formula with all H's.

SOLUTION:

a) HCl, $LiOH$ b) HNO_3, $Cu(OH)_2$ c) H_2SO_4, $Sr(OH)_2$

d) H_3PO_4, KOH c) H_3PO_4, KOH f) H_2CO_3, $Ca(OH)_2$

9.65 Calculate the mass of water that would be released if the water of hydration were completely driven off 1.0 mol of: (a) plaster of paris and (b) gypsum (see Table 9.6). How would the products of these reactions compare?

SOLUTION:

From Table 9.6, plaster of paris is $CaSO_4 \cdot H_2O$; and gypsum is $CaSO_4 \cdot 2H_2O$

a) $1\ \cancel{mol\ CaSO_4 \cdot H_2O} \times \dfrac{1\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ CaSO_4 \cdot H_2O}} \times \dfrac{18\ g\ H_2O}{\cancel{mol\ H_2O}} = 18\ g\ H_2O$

b) $1\ \cancel{mol\ CaSO_4 \cdot 2H_2O} \times \dfrac{2\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ CaSO_4 \cdot 2H_2O}} \times \dfrac{18\ g\ H_2O}{\cancel{mol\ H_2O}} = 36\ g\ H_2O$

The products are water (H_2O) and $CaSO_4$ in both cases. The amount of water is greater from gypsum.

9.67 Write formulas for the acid and indicated solid that could be used to prepare each of the following salts:

a) $CuCl_2$ (solid is an oxide) b) $MgSO_4$ (solid is a carbonate) c) $LiBr$ (solid is a hydroxide)

SOLUTION:

a) HCl and CuO b) H_2SO_4 and $MgCO_3$ c) HBr and $LiOH$

9.69 Write balanced full equations to illustrate each salt preparation described in Exercise 9.67.

SOLUTION:

a) $CuO + 2HCl \rightarrow CuCl_2 + H_2O$

b) $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$

c) $LiOH + HBr \rightarrow LiBr + H_2O$

9.71 Determine the number of moles of each of the following salts that would equal 1 eq of salt:

a) KNO_3 b) Li_2CO_3 c) $SrCl_2$

SOLUTION:

a) Since 1 mol salt gives 1 mol (+) charge, 1 mol KNO_3 = 1 eq KNO_3

b) Since 1 mol salt gives 2 mol (+) charge, ½ mol Li_2CO_3 = 1 eq Li_2CO_3

c) Since 1 mol salt gives 2 mol (+) charge, ½ mol $SrCl_2$ = 1 eq $SrCl_2$

9.73 Determine the number of equivalents and milliequivalents in each of the following:

a) 0.10 mol KI b) 0.25 mol $MgCl_2$ c) 4.73×10^{-2} mol $AgNO_3$

SOLUTION:

a) $0.10\ \cancel{\text{mol KI}} \times \dfrac{1\text{ eq}}{1\ \cancel{\text{mol KI}}} = 0.10\text{ eq}$

$0.10\ \cancel{\text{eq}} \times 1000\text{ meq}/\cancel{\text{eq}} = 1.0 \times 10^2\text{ meq}$

b) $0.25\ \cancel{\text{mol MgCl}_2} \times \dfrac{2\text{ eq}}{\cancel{\text{mol MgCl}_2}} = 0.50\text{ eq}$

$0.50\ \cancel{\text{eq}} \times 1000\text{ meq}/\cancel{\text{eq}} = 5.0 \times 10^2\text{ meq}$

c) $4.73 \times 10^{-2}\ \cancel{\text{mol AgNO}_3} \times \dfrac{1\text{ eq}}{\cancel{\text{mol AgNO}_3}} = 4.73 \times 10^{-2}\text{ eq}$

$4.73 \times 10^{-2}\ \cancel{\text{eq}} \times 1000\text{ meq}/\cancel{\text{eq}} = 47.3\text{ meq}$

9.75 Determine the number of equivalents and milliequivalents in 5.00 g of each of the following salts. Include any waters of hydration given in the salt formula when you calculate salt formula weights.

a) $NaCl$ b) $NaNO_3$ c) Na_3PO_4 d) $MgSO_4 \cdot 7H_2O$

SOLUTION:

a) $5.00\ \cancel{\text{g NaCl}} \times \dfrac{1\ \cancel{\text{mole NaCl}}}{58.5\ \cancel{\text{g NaCl}}} \times \dfrac{1\text{ eq}}{\cancel{\text{mole NaCl}}} = 8.55 \times 10^{-2}\text{ eq} = 85.5\text{ meq}$

b) $5.00\ \cancel{\text{g NaNO}_3} \times \dfrac{1\ \cancel{\text{mole NaNO}_3}}{85\ \cancel{\text{g NaNO}_3}} \times \dfrac{1\text{ eq}}{\cancel{\text{mole NaNO}_3}} = 5.88 \times 10^{-2}\text{ eq} = 58.8\text{ meq}$

c) $5.00\ \cancel{\text{g Na}_3\text{PO}_4} \times \dfrac{1\ \cancel{\text{mole Na}_3\text{PO}_4}}{164\ \cancel{\text{g Na}_3\text{PO}_4}} \times \dfrac{3\text{ eq}}{\cancel{\text{mole Na}_3\text{PO}_4}} = 9.15 \times 10^{-2}\text{ eq} = 91.5\text{ meq}$

d) $5.00\text{ g MgSO}_4 \cdot 7\text{H}_2\text{O}$ (MW $= 24.3 + 32.1 + 64.0 + 7(18) = 246.4$)

$5.00\ \cancel{\text{g MgSO}_4} \cdot 7\cancel{\text{H}_2\text{O}} \times \dfrac{1\ \cancel{\text{mole MgSO}_4}}{246.4\ \cancel{\text{g MgSO}_4} \cdot 7\cancel{\text{H}_2\text{O}}} \times \dfrac{2\text{ eq}}{\cancel{\text{mole MgSO}_4}} = 4.06 \times 10^{-2}\text{ eq} = 40.6\text{ meq}$

9.77 A sample of intracellular fluid contains 45.1 meq/L of Mg^{2+} ion. Assume the Mg^{2+} comes from dissolved $MgCl_2$ and calculate the number of moles and number of grams of $MgCl_2$ that would be found in 250 mL of the intracellular fluid.

SOLUTION:

$\dfrac{45.1\ \cancel{\text{meq}}}{\cancel{\text{L soln}}} \times 0.250\ \cancel{\text{L soln}} \times \dfrac{1\ \cancel{\text{eq}}}{1000\ \cancel{\text{meq}}} \times \dfrac{1\text{ mol MgCl}_2}{2\ \cancel{\text{eq}}} = 5.64 \times 10^{-3}\text{ mol MgCl}_2$

$5.64 \times 10^{-3}\ \cancel{\text{mol MgCl}_2} \times \dfrac{95.2\text{ g MgCl}_2}{1\ \cancel{\text{mol MgCl}_2}} = 0.537\text{ g MgCl}_2$

Strengths of Acids and Bases (Section 9.9)

9.79 Illustrate the difference between weak, moderately strong, and strong acids by writing dissociation reactions for the hypothetical acid HB using arrows of various lengths.

SOLUTION:

Strong: $HB + H_2O \xrightarrow{\hspace{2em}} H_3O^+ + B^-$

Mod. Strong: $HB + H_2O \longrightarrow H_3O^+ + B^-$

Weak: $HB + H_2O \rightarrow H_3O^+ + B^-$

9.81 Arrange the four acids classified as weak in Table 9.7 in order of increasing strength (weakest first, strongest last).

SOLUTION:
Use the % dissociation (smallest = weakest) or K_a (smallest = weakest).
H_3BO_3, H_2CO_3, $HC_2H_3O_2$, and HNO_2

9.83 Write dissociation reactions and K_a expressions for the following weak acids:
a) hypobromous acid, HBrO
b) sulfurous acid, H_2SO_3 (1st H only)
c) hydrogen sulfite ion, HSO_3^-
d) hydroselenic acid, H_2Se (1st H only)
e) arsenic acid, H_3AsO_4 (1st H only)

SOLUTION:

a) $HBrO \rightleftharpoons H^+ + BrO^- \qquad K_a = \dfrac{[H^+][BrO^-]}{[HBrO]}$

b) $H_2SO_3 \rightleftharpoons H^+ + HSO_3^- \qquad K_a = \dfrac{[H^+][HSO_3^-]}{[H_2SO_3]}$

c) $HSO_3^- \rightleftharpoons H^+ + SO_3^{2-} \qquad K_a = \dfrac{[H^+][SO_3^{2-}]}{[HSO_3^-]}$

d) $H_2Se \rightleftharpoons H^+ + HSe^- \qquad K_a = \dfrac{[H^+][HSe^-]}{[H_2Se]}$

e) $H_3AsO_4 \rightleftharpoons H^+ + H_2AsO_4^- \qquad K_a = \dfrac{[H^+][H_2AsO_4^-]}{[H_3AsO_4]}$

9.85 Equal molar solutions are made of three monoprotic acids: HA, HB, and HC. The pHs of the solution are respectively 4.82, 3.16, and 5.47. Rank the acids in order of increasing acid strength and explain your reasoning.

SOLUTION:
The strongest acid will have the highest $[H_3O^+]$ and the smallest pH. HC, HA, HB

9.87 Arsenic acid (H_3AsO_4) is a moderately weak, triprotic acid. Write equations showing its stepwise dissociation. Which of the three anions formed in these reactions will be the strongest Brønsted base? Which will be the weakest Brønsted base? Explain your answers.

SOLUTION:
(1) $H_3AsO_4 \rightleftharpoons H^+ + H_2AsO_4^-$
(2) $H_2AsO_4 \rightleftharpoons H^+ + HAsO_4^{2-}$
(3) $HAsO_4 \rightleftharpoons H^+ + AsO_4^{3-}$
H_3AsO_4 is the strongest acid; its conjugate base, $H_2AsO_4^-$, is the weakest base.
$HAsO_4^{2}$ is the weakest acid; its conjugate base, AsO_4^{3-}, is the strongest base.

Analysis of Acids and Bases (Section 9.10)

9.89 Describe the difference between the information obtained by measuring the pH of an acid solution and by titrating the solution with base.

SOLUTION:
The pH of a solution measures the $[H^+]$ actually present in the solution at that instant. The titration determines the total amount of $[H^+]$ ultimately available from the acid for reaction with the base.

9.91 Determine the number of moles of NaOH that could be neutralized by each of the following:
a) 1.00 L of 0.200 M HCl b) 500 mL of 0.150 M HNO_3

SOLUTION:

$$\text{a) } 1.00\ \cancel{\text{L soln}} \times \frac{0.200\ \cancel{\text{mol HCl}}}{\cancel{\text{L soln}}} \times \frac{1\ \text{mol NaOH}}{1\ \cancel{\text{mol HCl}}} = 0.200\ \text{mol NaOH}$$

$$\text{b) } 0.500\ \cancel{\text{L soln}} \times \frac{0.150\ \cancel{\text{mol HNO}_3}}{\cancel{\text{L soln}}} \times \frac{1\ \text{mol NaOH}}{1\ \cancel{\text{mol HNO}_3}} = 0.0750\ \text{mol NaOH}$$

Titration Calculations (Section 9.11)

9.93 Write a balanced full equation to represent the neutralization reaction between NaOH and each of the following acids:
a) molybdic acid, H_2MoO_4 b) permanganic acid, $HMnO_4$ c) phosphoric acid, H_3PO_4

SOLUTION:
a) $H_2MoO_4 + 2NaOH \rightarrow Na_2MoO_4 + 2H_2O$
b) $HMnO_4 + NaOH \rightarrow NaMnO_4 + H_2O$
c) $H_3PO_4 + NaOH \rightarrow Na_3PO_4 + 3H_2O$

9.95 Write a balanced full equation to represent the neutralization reaction between HCl and each of the following bases:
a) $Cd(OH)_2$ b) $Cr(OH)_3$ c) $Fe(OH)_2$

SOLUTION:
a) $Cd(OH)_2 + 2HCl \rightarrow CdCl_2 + 2H_2O$
b) $Cr(OH)_3 + 3HCl \rightarrow CrCl_3 + 3H_2O$
c) $Fe(OH)_2 + 2HCl \rightarrow FeCl_2 + 2H_2O$

9.97 A 25.0 mL sample of gastric juice is titrated with 0.0210 M NaOH solution. The titration to the equivalence point requires 29.8 mL of NaOH solution. If the equation for the reaction is
$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
what is the molarity of HCl in the gastric juice?

SOLUTION:

$$\text{moles HCl} = 29.8\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.0210\ \cancel{\text{mol NaOH}}}{\cancel{\text{L}}} \times \frac{1\ \text{mol HCl}}{1\ \cancel{\text{mol NaOH}}} = 6.26 \times 10^{-4}\ \text{mol HCl}$$

$$\text{molarity of HCl} = \frac{6.26 \times 10^{-4}\ \text{mol HCl}}{25.0\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \text{L}} = 0.0250\ \text{M HCl}$$

9.99 A 20.00 mL sample of each of the following acid solution is to be titrated to the equivalence point using 0.120 M NaOH solution. Determine the milliliters of NaOH solution that will be needed for each acid sample.
a) 0.180 M HCl
b) 0.180 M H_2SO_4
c) 0.100 M HCl
d) 10.00 g of H_3PO_4 in 250 mL of solution

e) 0.150 mol H_2MoO_4 in 500 mL of solution
f) 0.215 mol H_2MoO_4 in 700 mL of solution

SOLUTION:
At the equivalence point, the mol H^+ = mol OH^- or meq acid = meq base.
$V_aM_a = V_bM_b$, where the molarity of the acid and base must be the molarity of H^+ or OH^-.
For the 0.120 M NaOH, the solution is 0.120 M in OH^-.

a) For 0.180 M HCl, the solution is 0.180 M in H^+.

$$V_b = V_a \times \frac{M_a}{M_b} = 20.00 \text{ mL} \times \frac{0.180 \cancel{M}}{0.120 \cancel{M}} = 30.0 \text{ mL NaOH}$$

b) For 0.180 M H_2SO_4, there are 2 H^+ per H_2SO_4; so it is 0.360 M in H^+.

$$V_b = V_a \times \frac{M_a}{M_b} = 20.00 \text{ mL} \times \frac{0.360 \cancel{M}}{0.120 \cancel{M}} = 60.0 \text{ mL NaOH}$$

c) For 0.100 M HCl, the solution is 0.100 M in H^+.

$$V_b = V_a \times \frac{M_a}{M_b} = 20.00 \text{ mL} \times \frac{0.100 \cancel{M}}{0.120 \cancel{M}} = 16.7 \text{ mL NaOH}$$

d) $$\text{mol } H_3PO_4 = 10.00 \cancel{\text{g } H_3PO_4} \times \frac{1 \text{ mol } H_3PO_4}{97.99 \cancel{\text{g } H_3PO_4}} = 0.1021 \text{ mol } H_3PO_4$$

$$\text{molarity} = \frac{0.1021 \text{ mol } H_3PO_4}{0.250 \text{ L soln}} = 0.4084 \text{ M}$$

For 0.4084 M H_3PO_4, there are 3 H^+ per H_3PO_4; so it is 1.225 M in H^+.

$$V_b = V_a \times \frac{M_a}{M_b} = 20.00 \text{ mL} \times \frac{1.225 \cancel{M}}{0.120 \cancel{M}} = 204 \text{ mL NaOH}$$

e) $$\text{molarity of } H_2MoO_4 = \frac{0.150 \text{ mol } H_2MoO_4}{0.500 \text{ L}} = 0.300 \text{ M}$$

Since there are $2H^+$ per H_2MoO_4, the solution is 0.600 M in H^+.

$$V_b = V_a \times \frac{M_a}{M_b} = 20.00 \text{ mL} \times \frac{0.600 \cancel{M}}{0.120 \cancel{M}} = 100 \text{ mL NaOH}$$

f) $$\text{molarity of } H_2MoO_4 = \frac{0.215 \text{ mol } H_2MoO_4}{0.700 \text{ mL}} = 0.307 \text{ M}$$

Since there are $2H^+$ per H_2MoO_4, the solution is 0.614 M in H^+.

$$V_b = V_a \times \frac{M_a}{M_b} = 20.00 \text{ mL} \times \frac{0.614 \cancel{M}}{0.120 \cancel{M}} = 102 \text{ mL NaOH}$$

9.101 The following acid solutions were titrated to the equivalence point with the base listed. Use the titration data to calculate the molarity of each acid solution.
a) 25.00 mL of HI solution required 27.15 mL of 0.250 M NaOH solution
b) 20.00 mL of H_2SO_4 solution required 11.12 mL of 0.109 M KOH solution
c) 25.00 mL of gastric juice (HCl) required 18.40 mL of 0.0250 M NaOH solution

SOLUTION:

a) $$\text{moles HI} = 27.15\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.250\ \cancel{\text{mol NaOH}}}{\cancel{\text{L}}} \times \frac{1\text{ mol HI}}{1\ \cancel{\text{mol NaOH}}} = 6.79 \times 10^{-3}\text{ mol HI}$$

$$\text{molarity HI} = \frac{6.79 \times 10^{-3}\text{ mol}}{25.00\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{1\text{ L}} = 0.272\text{ M HI}$$

b) $$\text{moles } H_2SO_4 = 11.12\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.109\ \cancel{\text{mol KOH}}}{\cancel{\text{L}}} \times \frac{1\text{ mol } H_2SO_4}{2\ \cancel{\text{mol KOH}}} = 6.06 \times 10^{-4}\text{ mol } H_2SO_4$$

$$\text{molarity } H_2SO_4 = \frac{6.06 \times 10^{-4}\text{ mol}}{20.00\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{1\text{ L}} = 0.0302\text{ M } H_2SO_4$$

c) $$\text{moles HCl} = 18.40\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.0250\ \cancel{\text{mol NaOH}}}{\cancel{\text{L}}} \times \frac{1\text{ mol HCl}}{1\ \cancel{\text{mol NaOH}}} = 4.6 \times 10^{-4}\text{ mol HCl}$$

$$\text{molarity HCl} = \frac{4.6 \times 10^{-4}\text{ mol}}{25.00\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{1\text{ L}} = 0.0184\text{ M HCl}$$

9.103 A 20.00 mL sample of diprotic oxalic acid ($H_2C_2O_4$) solution is titrated with 0.250 M NaOH solution. A total of 27.86 mL of NaOH is required. Calculate:

a) the number of moles of oxalic acid in the 20.00 mL sample
b) the molarity of the oxalic acid solution
c) the number of grams of oxalic acid in the 20.00 mL sample

SOLUTION:

a) $$\text{mol } H_2C_2O_4 = 27.86\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.250\ \cancel{\text{mol NaOH}}}{\cancel{\text{L}}} \times \frac{1\text{ mol } H_2C_2O_4}{2\ \cancel{\text{mol NaOH}}} = 3.48 \times 10^{-3}\text{ mol } H_2C_2O_4$$

b) $$\text{molarity} = \frac{3.48 \times 10^{-3}\text{ mol } H_2C_2O_4}{20.00\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{1\text{ L}} = 0.174\text{ M } H_2C_2O_4$$

c) $$3.48 \times 10^{-3}\ \cancel{\text{mol } H_2C_2O_4} \times \frac{90.0\text{ g } H_2C_2O_4}{\cancel{\text{mol } H_2C_2O_4}} = 0.313\text{ g } H_2C_2O_4$$

Hydrolysis Reactions of Salts (Section 9.12)

9.105 A solution of solid NH_4Cl in pure water is acidic (the pH is less than 7). Explain.

SOLUTION:

The NH_4^+ is a Brønsted acid, establishing an equilibrium with H_2O:

$$NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$$

The H_3O^+ formed makes the solution acidic (pH is less than 7).

9.107 Predict the relative pH (*greater than 7, less than 7,* etc) for water solutions of the following salts. Table 9.9 may be useful. For each solution in which the pH is greater or less than 7, explain why and write a net ionic equation to justify your answer.

a) potassium sulfite, K_2SO_3
b) lithium nitrite, $LiNO_2$

c) sodium carbonate, Na_2CO_3
d) methylammonium chloride, CH_3NH_3Cl (CH_3NH_2 is a weak base)

SOLUTION:
a) Basic. SO_3^{2-} is a weak base. $SO_3^{2-} + H_2O \rightleftharpoons HSO_3^- + OH^-$
b) Basic. NO_2^- is a weak base. $NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$
c) Basic. CO_3^{2-} is a weak base. $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$
d) Acidic. $CH_3NH_3^+$ is a weak acid. $CH_3NH_3^+ + H_2O \rightleftharpoons H_3O^+ + CH_3NH_2$

9.109 A chemist has 20.00 mL samples of 0.100 M acid A and 0.100 M acid B in separate flasks. Both acids are monoprotic. Unfortunately, the flasks are not labeled, so the scientist does not know which sample is in which flask. But fortunately, it is known that acid A is strong and acid B is weak. Before thinking about the problem, the chemist adds 20.00 mL of 0.100 M NaOH to each flask. Explain how the chemist could use a pH meter (or pH paper) to determine which flask originally contained which acid.

SOLUTION:
The 20.00 mL of NaOH would react with all of the acid in both cases, giving essentially a solution of the sodium salt of the acid. The salt of the strong acid would not hydrolyze and would have a pH = 7. The salt of the weak acid would hydrolyze, giving a basic solution (pH greater than 7).

9.111 How would the pH of equal molar solutions of the following salts compare (*highest, lowest,* etc)? NaH_2PO_4, Na_2HPO_4, and Na_3PO_4

SOLUTION:
The Na_3PO_4 would have the highest pH, definitely basic, because the PO_4^{3-} is the strongest base of the three anions. The NaH_2PO_4 would be the lowest pH, definitely acidic, because $H_2PO_4^-$ is the strongest acid of the three anions. The pH of the Na_2HPO_4 would be higher than NaH_2PO_4 and lower than Na_3PO_4.

Buffers (Section 9.13)

9.113 Could a mixture of ammonia (NH_3), a weak base, and ammonium chloride (NH_4Cl) behave as a buffer when dissolved in water? Use equations to justify your answer.

SOLUTION:
Yes. The NH_3 could react with added acid, and the NH_4^+ could react with added base. The equations are:
$NH_3 + H^+ \rightarrow NH_4^+$ and $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$

9.115 Calculate the pH of a buffer made by dissolving 1 mol formic acid (HCOOH) and 1 mol sodium formate (HCOONa) in 1 L of solution (see Table 9.9).

SOLUTION:
From Table 9.9, $pK_a = 3.74$ Using Equation 9.54, $pH = pK_a + \log\frac{[B^-]}{[HB]}$

$[B^-] = [HB] = 1.0$ $pH = 3.74 + \log\frac{1.0}{1.0} = 3.74$

9.117 Which of the following acids and its conjugate base would you use to make a buffer with the highest buffer capacity and a pH of 3.00? Explain your reasons. formic acid, lactic acid, nitrous acid

SOLUTION:
The best buffer capacity is when the pH equals the pK_a. The closer the pH is to pK_a, the higher the buffer capacity. Since pK_a for nitrous acid is the closest to 3.00, a mixture of nitrous acid and nitrite ion would be best.

9.119 Calculate the pH of the buffers with the acid and conjugate base concentrations listed below.

a) $[HPO_4^{2-}] = 0.33$ M, $[PO_4^{3-}] = 0.52$ M
b) $[HNO_2] = 0.029$ M, $[NO_2^-] = 0.065$ M
c) $[HCO_3^-] = 0.50$ M, $[CO_3^{2-}] = 0.15$ M

SOLUTION:

a) pK_a for $HPO_4^{2-} = 12.66$

$$pH = 12.66 + \log\frac{0.52}{0.33} = 12.86$$

b) pK_a for $HNO_2 = 3.33$

$$pH = 3.33 + \log\frac{0.065}{0.029} = 3.68$$

c) pK_a for $HCO_3^- = 10.25$

$$pH = 10.25 + \log\frac{0.15}{0.50} = 9.73$$

9.121 A citric acid-citrate buffer has pH = 3.20. You want to increase the pH to a value of 3.35. Would you add critic acid or sodium citrate to the solution? Explain.

SOLUTION:
Add sodium citrate. There are at least two viable explanations: (1) to increase pH, add the Brønsted base, citrate; or (2) in Equation 9.54, if the $[B^-]$ increases, the value of the log term increases, which increases pH.