**Chapter 16 exercise 13 المساق ك-103 الشعبة**

**Q1**. Practice exercise page 671

 Write the formula for the conjugate acid of the following, HSO3- , F- , PO43-  and CO.

Answer:

 HSO3-  + H+  → H2SO4

 F-  + H+  → HF

 PO43-  + H+  → HPO42-

 CO + H+ → HCO+

**Q2**. Practice exercise page 671

 When lithium oxide (Li2 O) is dissolved in water, the solution turns basic from the reaction of

 oxide ion (O2- ) with water. Write the reaction that occurs, and identify the conjugate acid –

 base pair.

Answer:

 Li2O (aq)  + H2O (l)  → LiOH + LiOH

 O2-(aq)  + H2O (l)  → OH- (aq)  + OH- (aq)

 base acid conjugate acid conjugate base

**Q3**. Practice exercise page 673

 For the following reactions, use figure 16.4 to predict whether the equilibrium lies predominantly

 to the left or to the right.

 a) HPO42- (aq)  + H2O (l) ­­ ­ ↔ H2PO4- (aq)  + OH- (aq)

 b) NH4+ (aq)  + OH- (aq) ↔ NH3 (aq)  + H2O (l)

Answer:

 a) OH‑ is in right column (strong base) than H2O , H2O is a conjugate weak acid . HPO42-

 HPO42-  is acid more than H2PO4-  , H2PO4-  is a weak base .

 OH-  more strong base than H2PO4-  , the reaction is shift to the left.

 b) OH-  is a strong base than H2O , H2O is conjugate weak acid. OH-  strong base than NH4+ ,

 the reaction shift to the right.

**Q4**. Practice exercise page 675

 Indicate whether solutions with each of the following in concentrations are neutral, acidic or basic.

 a)[H+] = 4 x 10-9 M, b) [OH-] = 1 x 10-7 M c) [OH-] = 7 x 10-13 M

 Answer:

 a) {H+] [OH-] = 1.0 x 10-14  [OH-] = (1.0 x 10-14) / (4 x 10-9) = 0.25 x 10-5 M

 [OH-] more concentrated than [H+], the solution is acidic.

 b) [OH-] = 1 x 10-7  [H+] = 1 x 10-7

 c) [H+] [OH-] = 1.0 x 10-14  [H+] = (1 x 10-14 ) / (7 x 10-13) = 0.143 x 10-1 M more acidic

**Q5**. Practice exercise page 675

 Calculate the concentration of OH-  in solution in which

 a) [H+] = 2 x 10-6 M b) [H+] = [OH+] c) [H+] = 100 x [OH+]

Answer:

 a) [OH+] = (1.0 x 10-14 ) / (2 x 10-6 ) = 0.5 x 10-8 M = 5 x 10-9 M

 b) [H+] [OH-] = 1.0 x 10-14

 (x)(x) = 1.0 x 10-14  x2  = 1.0 x 10-14  [OH-] = 1.0 x 10-7

 c) [OH-] = (1.0 x 10-14) (100 x [OH-]) , [OH-]2  = 1 x 10-6

 [OH-] = 1.0 x 10-8 M

**Q6**. Practice exercise page 677

 A solution formed by dissolving an anti-acid tablet has a pH of 9.18. Calculate [H+].

Answer:

 PH = - log [H+] = 9.18

 Log [H+] = - 9.18

 [H+] = antilog ( - 9.18) = 10- 9.18  = 6.6 x 10-10  M

**Q7**. Practice exercise 680

 An aqueous solution of HNO3 has a pH of 2.34. What is the concentration of the acid?

Answer:

 HNO3 → H+  + NO3-

 pH = - log [H+] = 2.34

 [H+] 10- 2.34  = 4.57 x 10-3 M concentration of HNO3 is 4.57 x 10-3 M

**Q8**. Practice exercise page 680

 What is the concentration of a solution of

 a) KOH for which pH is 11.89 b) Ca(OH)2 for which the pH is 11.68.

Answer:

 a) KOH → K+  + OH-

 pH = - log [H+] = 11.89

 [H+] = 10- 11.89  = 1.29 x 10-12 M

 [OH- ] = (1.0 x 10-14 ) / (1.29 x 10-12 ) = 0.775 x 10-2 M = 7.8 x 10-3 M

 b) Ca(OH)2  → Ca2+  + 2OH-

 pH = - log [H+] = 11.68 [H+] = 10- 11.68  = 2.089 x 10-12 M

 [H+] [OH-] = 1.0 x 10-14

 [OH-] = (1.0 x 10-14 ) / (2.089 x 10-12 ) = 0.4786 x 10-2 M

 for [OH-]2  = 2 x 0.4786 x 10-2  = 2.4 x 10-3 M

**Q9**. Practice exercise page 683

 Naicin, one of the B- vitamins, a 0.020 M solution of niacin has a pH of 3.26. What is the acid –

 constant -dissociation constant , Ka for niacin?

Answer:

 pH = - log [H+] = 3.26

 [H+] = 10- 3.26  = 5.495 x 10-4 M

 Niacin → H+  + niacin-

 initial 0.020 M 0 M 0 M

 change - 5.5 x 10-4 M + 5.5 x 10-4 M + 5.5 x 10-4 M

 equilibrium (0.02 – 5.5 x 10-4 ) 5.5 x 10-4 M 5.5 x 10-4 M

 Ka = [H+] [niacin- ] / [niacin] = (5.5 x 10-4 )2 / (0.02 – 5.5 x 10-4 ) = 1.5 x 10-5

**Q10**. Practice exercise page 684

 A o.020 solution of niacin has pH of 3.26. Calculate the percent ionization of the niacin/

Answer:

 pH = - log [H+] = 3.26

 [H+] = 10- 3.26  = 5.4954 x 10-4 M

 Percent ionization = [H+] equilibrium  / [niacin] = (5.4954 x 10-4 )(100) / 0.02 = 2.7%

**Q11**. Practice exercise page 686

 The Ka for niacin is 1.5 x 10-5. What is the ph of 0.010 M solution of niacin?

Answer:

 Niacin ↔ H+ + niacin-

 Niacin → H+  + niacin-

 initial 0.010 M 0 m 0 M

 change - X M + X M + X M

 equilibrium 0.01 – X X X

 Ka = (X) (X) / 0.01 – X 0.01 – X = 0.01

 X2  / 0.01 = 1.5 x 10-5  X = [H+] = 1.22 x 10-3 m

 pH = 3.41

**Q12**. Practice exercise page 690

 a) Calculate the pH of a 0.020 M solution of oxalic acid (H2C2O4 ) Ka1 = 5.0 x 10-5, Ka2 = 6.4 x 10-5

 b) Calculate the concentration of oxalic ion, C2O42-  in the solution.

Answer:

 a) H2C2O4 ↔ HC2O4-  + H+

 initial 0.020 M 0 M 0 M

 change - X M + X M + X M

 equilibrium 0.020 – X X M X M

 (X)(X) / 0.02 – X = 5.9 x 10-2

 X2  + 5.9 x 10-2  X - 0.118 x 10-2  = 0

 X = (- 5.9 x 10-2 ) ± [√ (5.9 x 10-2)2  - 4 (- 0.118 x 10-2 )] / 2

 X = 0.0158 pH = - log (0.0158)

 pH = 1.8

 b) HC2O4-  ↔ H+  + C­­2O4 -2

 initial 0.0158 M 0.0158 M 0 M

 change - Y + Y M + Y M

 equilibrium 0.0158 – Y 0.0158 + Y Y M

 (Y) (0.0158 + Y) / (0.0158 – Y) = 6.4 x 10-5  ( Y is very small can be neglected )

 Y \* 0.0158 / 0.0158 = 6.4 x 10-5  Y = [C2O42- ] = 6.4 xc 10-5

**Q13**. Practice exercise page 693

 A solution of NH3 in water has a pH = 11.17. What is the molarity of the solution?

Answer:

 NH3  + H2O ↔ NH4+  + OH-

 POH  = 14 - pH 14.00 – 11.17 = 2.83

 [OH- ] = 10-2.83  = 1.48 x 10-3 M

 NH3  + H2O ↔ NH4+  + OH-

 initial X 0 M 0 M

 change - 1.48 x 10-3 + 1.48 x 10-3  + 1.48 x 10-3 M

 equilibrium X – 1.48 x 10-3  1.48 x 10-3 1.48 x 10-3 M

 Kb = [NH4] [OH- ] / [NH3 ] = 1.8 x 10-5 = (1.48 x 10-3)2 / (X – 1.48 x 10-3 )

 X = [ (2.19 x 10-6 ) + (2.664 x 10-8)] / (1.8 x 10-5 ) = 0.123 M

**Q14**. Practice exercise page 695

 a) Which of the following anions has the largest base dissociation constant NO2- , PO43-  , N3-

 b) The base quinolone, its conjugate is pKa = 4.9. What is the base-dissociation constant for quinolone.

 Answer:

 a) NO2-  is a conjugate base for the acid HNO2 Ka = 4.5 x 10-4

 PO43-  is a conjugate base for H3PO4 has three Ka 7.5 x 10-3 , 6.2 x 10-8  and 4.2 x 10-13

 N3- is a conjugate base for the acid Ka = 1.9 x 10-5

 Kb = (1.0 x 10-14 ) (4.5 x 10-4 ) = 0.22 x 10-10  for NO2-

 Kb = (1.0 x 10-14 ) (4.2 x 10-13 ) = 0.24 x 10-1  for PO43-

 Kb = (1.0 x 10-14 ) (1.9 x 10-5 ) = 0.53 x 10-9 for N3-

 Largest base dissociation constant is PO43-

 b) pKa + pKb = pKw

 4.90 + pKb = 14.00

 pKb = - log Kb = 9.1

 Kb = 10- 9.1  = 7.9 x 10-10

**Q15**. Practice exercise page 698

 In each of the following, indicate which salt in each of the following pair will form the more acidic (or more

 basic).

 (a) NaNO3  or Fe(NO3)3  ( b) KBr or KBrO (c) CH3NH3 Cl or BaCl2 (d) NH4NO2 or NH4NO3

 Answer:

 a) NaNO3 ↔ Na+  + NO3-

 Na+  ion from group 1A has no effect on pH

 NO3-  ion is the conjugate base of strong acid HNO3  has no effect on pH

 The solution is neutral

 Fe3+  is not from group 1A or 2A , decrease the pH

 NO3-  ion is the conjugate base of strong acid HNO3  has no effect on pH

 The solution is acidic

 Fe(NO3)3  more acidic than NaNO3

 b) K+  ion from group 1A has no effect on pH

 Br-  is a conjugate base of strong acid HBr it has no influence on pH

 KBr form a neutral solution

 KBrO, BrO-  is a conjugate base for a weak acid HBrO

 BrO- + H2O ↔ HBrO + OH-  the solution is basic

 KBr is more acidic than KBrO

**Exercises page 710**

**16.15**

 a) what is the difference between the Arrhenius and Bronsted-Lowry definition of an acid?

 b) NH3 (g) and HCl (g) react to form NH4Cl (s) (figure 16.3) which substance is the Bronsted-Lowry

 acid in this reaction? Which is the Bronsted-Lowry base?

Answer:

1. Arrhenius base is added to water leads to an increase in the concentration of OH- , while

Arrhenius acid in water an increase in the concentration of H+ .

Bronsted – Lowry base is accept a proton from H2O and the acid it donate a proton from H2O.

1. NH3  + HCl → NH4Cl

 HCl is the Bronsted – Lowry acid and NH3  is the Bronsted – Lowry base

**16.17**

 a) Give the conjugate base of the following Bronsted – Lowry acids ( i) HIO3  (ii) NH4+

 b) Give the conjugate acid of the following base (i) O2-  (ii) H2PO4-

Answer

 a) i) HIO3  → conjugate base is IO3-

 ii) NH4+ → conjugate base is NH3

b) i) O2-  → conjugate acid is OH-

 ii) H2PO4-  → conjugate acid is H3PO4

**16.19**

 Designate the Bronsted – Lowry acid and Bronsted – Lowry base on the left side of each

 of the following equations, and also designate the conjugate acid and conjugate base on

 the right side.

 a) NH4+ (aq)  + CN- (aq) ↔ HCN (aq)  + NH3 (aq)

 b) (CH3)3N (aq)  + H2O (l)  ↔ (CH3)NH+ )aq)  + OH- (aq)

 c) HCHO2 (aq)  + PO43-  ↔ CHO2- (aq)  + HPO42-  (aq)

 Answer:

 a) NH4+ (aq)  + CN- (aq) ↔ HCN (aq)  + NH3 (aq)

 acid base conjugate acid conjugate base

 b) (CH3)3N (aq)  + H2O (l)  ↔ (CH3)NH+ )aq)  + OH- (aq)

 base acid conjugate acid conjugate base

 c) HCHO2 (aq)  + PO43-  ↔ CHO2- (aq)  + HPO42-  (aq)

 acid base conjugate base conjugate acid

**16.21**

 a) The hydrogen oxalate ion (HC2O4- ) is amphiprotic. Write a balance chemical equation showing

 how it acts as an acid toward water and another equation showing how it acts as a base

 towered water.

 b) What is the conjugate acid of HC2O4- ? what is the conjugate base.

 Answer:

 a) HC2O4-  + H2O ↔ C2O42-  + H3O+ (behave as acid in water)

 HC2O4-  + H2O ↔ H2C2O4  + OH-  (behave as base in water)

 b) H2C2O4  is a conjugate acid of HC2O4-

 C2O42-  is a conjugate base of HC2O4-

**16.27**

 Predict the products of the following acid – base reactions, and predict whether the the

 equilibrium lies to the left or to the right of the equations.

 a) O2- (aq)  + H2O (l)  ↔

 b) CH3COOH (aq)  + HS-  ↔

 c) NO2-  + H2O (l)  ↔

 Answer:

 a) O2-  + H2O ↔ OH-  + OH-

 base acid conjugate acid conjugate base

 b) CH3COOH (aq)  + HS-  ↔ H2S + CH3COO-

 acid base conjugate acid conjugate base

1. NO2-  + H2O (l)  ↔ HNO2  + OH-

The equilibrium to the left

**16.31**

 Calculate [H+] for such of the following solutions, and indicate whether the solution is acidic,

 basic or neutral.

 A) [OH-] = 0.00045 M b) [OH-] 8.8 x 10-9 M

 c) a solution which [OH-] is 100 times greater then [H+]

Answer:

 a) pOH = - log [OH-] = - log (0.00045) = 3.35

pOH + pH = 14

 pH = 10.65

 b) pOH = - log [OH-] = - log (8.8 x 10-9 ) = 7.06

pOH = 14.00 – 7.06 = 6.94 the solution is acidic

 c) [H+] x 100 [H+] = 10-14

[H+] = 10-14 / 100 = 10-16

 [H+] = 10-8  pH = 8 the solution is basic

**16.33**

 At the freezing point of water 0˚C , Kw  = 1.2 x 10-15 . Calculate [H+] and [OH- ] for neutral solution

 of this temperature.

Answer:

 [OH- ] [H+ ] = 1.2 x 10-15  X2  = 1.2 x 10-15

 [OH- ] = [H+ ] = 3.5 x 10-6 M

**16.35**

 By what factor does [H+ ] change for pH change of

 a) 2.00 units b) 0.50 units

Answer:

 pH = - log [H+ ] [H+ ] = 10- 2.00  = 0.01 = 1 / 100

 [H+ ] = 10-0.50  = 0.316 = 1 / 0.316 = 3.2

**16.39**

 Complete the following table by calculating the missing entries and indicating whether the

 solution is acidic or basic

 H+  OH-  pH pOH acidic or basic

 7.5 x 10-3 M

 3.6 x 10-10 M

 8.25

 5.70

Answer:

 [H+ ] [OH- ] 1.0 x 10-14

 [OH- ] = (1.0 x 10-14 ) / (7.5 x 10-3 ) = 1.3 x 10-12 M

 pOH = - log (1.3 x 10-12 ) = 11.87

 pH = 14 – 11.87 = 2.13 the solution is acidic

 H+  OH-  pH pOH acidic or basic

 7.5 x 10-3 M 1.3 x 10-12 M 2.13 11.87 acidic

 2.8 x 10-5 M 3.6 x 10-10 M 4.56 9.44 acidic

 5.6 x 10-9 M 1.8 x 10-6 M 8.25 5.75 basic

 5.0 x 10-9 M 2.0 x 10-6 M 8.30 5.70 basic

**16.41**

 The average pH of normal arterial blood is 7.40. At normal body temperature is 37˚C, Kw  =

 2.4 x 10-14 . Calculate [H+ ], and pOH for the blood at this temperature.

Answer:

 PH = - log [H+ ] = 7.40

 [H+ ] = 3.98 x 10-8 M = 4.00 x 10-8 M

 [H+ ] [OH- ] = 2.4 10-14  [OH- ] = 0.6 x 10-6 M

 pOH = - log (0.6 x 10-6 ) = 6.22

**16.45**

 Calculate the pH of the following strong acid solutions.

 a) 8.5 x 10-3 M HBr.

 b) 1.52 g of HNO3  in 575 mL of solution.

 c) 5.00 mL of 0.250 M HClO4 ­­diluted to 50.0 mL.

 d) a solution formed by mixing 10.0 mL of 0.100 M HBr with 20.0 mL of 0.2 M HCl.

Answer:

 a) HBr → H+  + Br-

 8.5 x 10-3 M 8.5 x 10-3 M 8.5 x 10-3 M

 pH = - log (8.5 x 10-3 ) = 2.07

 b) mole of HNO3  = (1.52 g) /(63 g/mol) = 0.024 mol

 molarity of HNO3  = (0.024 mol HNO3 )( 1000 mL / 575 mL) = 0.042 M [H+ ]

 pH = – log [H+ ] = - log (0.042) = 1.38

 c) N1 x V1 = N2 x V2

 0.250 M x 5.00 mL = N2 x 50.0 mL

 N2 = 0.0250 M molarity of HClO4  = molarity of [H+ ]

 pH = - log [H+ ] = - log (0.0250) = 1.60

 d) 10 mL + 20 mL = 30 mL volume of solution

 N1 x V1 = N2 x V2

 0.100 M x 10 mL = N2 x 30 mL N2 = 0.033 M of [HBr ] = [ H+ ]

 N1 x V1 = N2 x V2

 0.2 M x 20 mL = N2 x 30 mL N2 = 0.133 M of HCl = [H+ ]

 0.033 M [H+ ] from HBr + 0.133 M [H+] from HCl = 0.166 M

 pH = - log (0.166) = 0.778

**16.47**

 Calculate [OH- ] and pH for

 a) 1.5 x 10-3 M Sr(OH)2

 b) 2.250 g of LiOH in 250.0 mL of solution

 c) 1.00 mL of 0.175 M NaOH diluted to 2.000 L

 d) a solution formed by adding 5.00 mL of 0.105 M KOH to 15.0 mL of 9.5 x 10-2 M Ca(OH)2

Answer:

 a) Sr(OH)2  → Sr2+  + 2OH-

 [OH- ] = 2 x 1.5 x 10-3  = 3.0 x 10-3 M

 pOH = - log (3.0 x 10-3 ) = 2.523

 pH = 14 – 2.523 = 11.48

 b) mole of LiOH = 2.250 / 24 = 0.094

 [ LiOH] M = (0.094 mol LiOH) ( 1000 mL /250 mL) = 0.375 M of LiOH

 LiOH → Li+  + OH-

 0.375 0.375

 pOH = - log [OH- ] = - log (0.375) = 0.426

 pH = 14 – 0.426 = 13.57

 c) N1 x V1 = N2 x V2

 0.175 x 1.0 mL = N2 x 2000 mL N2 = 8.75 x 10-5 M NaOH

 pOH = - log [OH- ] = - log (8.75 x 10-5 ) = 4.058

 pH = 14 – 4.058 = 9.942

 d) total volume of solution 5.0 mL + 15 mL = 20 mL

 for KOH N1 x V1 = N2 x V2

 0.105 M x 5.0 mL = N2 x 20.0 mL N2 = 0.02625 M

 KOH for Ca(OH)2  9.5 x 10-2 M x 15 mL = N2 x 20 mL

 N2 = 0.07125 M of Ca(OH)2  = 0.07125 M [OH- ]

 total concentrations of [OH- ] = 0.02625 + 2x 0.07125 = 0.16875 M = 0.17 M

 pOH = - log [OH- ] = - log (0.17) = 0.773

 pH = 14 – 0.773 = 13.23

**16.53**

 Lactic acid (CH3CH(OH)COOH) has one acidic hydrogen. A 0.10 M solution lactic acid has pH = 2.44.

 Calculate Ka.

Answer:

 CH3CH(OH)COOH ↔ H+  + CH3CH(OH)COO-

 pH = - log [H+ ] = 2.44

 [H+ ] = 0.00363 M

 Ka = [CH3CH(OH)COO-  ] [ H+ ] / [CH3CH(OH)COO] = (0.000363)(0.00363) / 0.1

 Ka = 1.32 x 10-4  = 1.4 x 10-4

**16.55**

 A 0.10 M solution of chloroacetic acid (ClCH2COOH) is 11.0% ionized. Using this information,

 calculate [ClCH2COO-], [H+ ], and [ClCH2COOH] and Ka for chloroacetic acid.

Answer:

 Ionizes chloroacetic acid = 0.10 M x 11/100 = 0.011 M

 Remain chloroacetic acid = 0.10 – 0.011 = 0.089 M

 ClCH2COOH (aq) ↔ ClCH2COO- (aq)  + H+  (aq)

 initial 0.011 M 0 M 0 M

 change - 0.011 M + 0.011 M + 0.011 M

 equilibrium (0.10 – 0.011) 0.011 M 0.011 M

 Ka = (0.011) (0.011) / (0.089) = 1.4 x 10-3

**16.57**

 A particular sample of vinegar has a pH = 2.9. If acetic acid is the only acid that vinegar contains

 ( Ka = 1.8 x 10-5 ), calculate the concentration of acetic acid in the vinegar.

Answer:

 pH - log [H+ ] = 2.90

 [H+ ] = 0.00126 = the concentration of vinegar

 CH3COOH ↔ CH3 COO-  + H+

 Ka = [CH3COO-/] [H- ] / [CH3COOH] = 1.8 x 10-5

 (0.00126) (0.00126) / [CH3COOH] = 1.8 x 10-5

 [CH3COOH] = 0.0882 M

**16.63**

 Saccharin, a sugar substitute, is a weak acid with pKa = 2.32 at 25˚C. It ionizes in aqueous

 solution as follows:

 HNC7H4SO3 (aq) ↔ H+ (aq)  + NC7H4SO3- (aq)

 What is the pH of a 0.10 M solution of this substance?

Answer:

 pKa = - log Ka = 2.32

 Ka = 4.79 x 10-3  = [H+ ] [NC7H4SO3- ] / [HNC7H4SO3 ]

 HNC7H4SO3 (aq) ↔ H+ (aq)  + NC7H4SO3- (aq)

 initial 0.10 M 0 M 0 M

 change - X M + X M + X M

 equilibrium (0.10 – X) M X M X M

 X2  / (0.10 – X) = 4.79 x 10-4

 X = [H+ ] = 2.23 x 10-2 M

 pH = - log (2.23 x 10-2 ) = 1.652 = 1.7

**16.65**

 Calculate the percent ionization of hydrazoic acid (HN3) in solutions each of the following

 concentrations (Ka = 1.9 x 10-5 )

 a) 0.400 M b) 0.100 M c) 0.0400 M

Answer:

 a) percent ionization = [H+ ]equ  / [HA]initial  x 100%

 HN3  ↔ H+  + N3-

 Ka = [H+ ] [ N3- ] / [HN3] = 1.9 x 10-5  X = [H+ ] = [N3- ]

 X2  / (0.4 – X ) = 1.9 x 10-5  X = 2.77 x 10-3 M = [H+ ]

 Percent ionization = (2.77 x 10-3 )/ (0.4) x 100 = 0.69%

 b) [H+ ] [N3- ] / [HN3 ] = 1.9 x 10-3 M

 X2  / (0.1 – X) = 1.9 x 10-3  X = 1.38 x 10-3 M = [H+ ]

 Percent ionization = [H+ ] equ  / [HA] initial  x 100 = 1.4%

 c) X2  / (0.04 – X ) = 1.9 x 10-5  X = 0.872 x 10-3 M = [H+ ]

 percent ionization = ( 0.872 x 10-3 ) / 0.04 x 100 = 2.2%

**16.73**

 Write the chemical equations and Kb expression for the ionization of each of the following

 bases in aqueous solutions.

 a) dimethyl amine (CH3)2NH

 b) carbonate ion CO32-

 c) formate ion CHO2-

Answer:

 a) (CH3)2NH (aq)  + H2O (l)  ↔ (CH3)2NH2+ + OH (aq)

 Kb = [(CH3)2NH2+ ] [ OH- ] / [(CH3)2NH ]

 b) CO32-  (aq)  + H2O (l)  ↔ HCO3- (aq)  + OH+ (aq)

 Kb = [HCO3- ] [OH- ] / [CO32- ]

 c) HCOO-  (aq) + H2O (l)  ↔ HCOOH (aq)  + OH- (aq)

 Kb = [HCOOH] [OH- ] / [HCOO- ]

**16.83**

 Calculate [OH- ] and pH for each of the following solutions

 Ka for HCN = 4.9 x 10-10  , Ka for H2CO3  = 5.6 x 10-11  , Ka for HNO2 = 4.5 x 10-4 )

 a) 0.10 M NaCN

 b) 0.080 M Na2 CO3

 c) a mixture that is 0.10 M in NaNO2  and 0.20 M in Ca(NO2)2

Answer:

 a) NaCN → Na+  + CN-

 CN-  + H2O → HCN + OH-

 Kb x Ka = Kw Kb = (1.0 x 10-14 ) / (4.9 x 10-10 ) = 0.204 x 10-4

 Kb = [HCN] [OH- ] / [CN- ] = 0.204 x 10-4

 CN-  + H2O → HCN + OH-

 initial 0.10 M 0 M 0 M

 change - X M + X M + X M

 equilibrium 0.10 – X X M X M

 (X)(X) / 0.1 – X = 0.204 x 10-4  X = 1.43 x 10-3  M = [OH- ]

 [OH- ] [H+ ] = 1.0 x 10-14  [H+ ] = (1.0 x 10-14 ) / (1.43 x 10-3 ) = 0.7 x 10-11

 pH = - log (0.7 x 10-11 ) = 11.15

 b) Na2CO3 → 2Na+  + CO32-

 CO32-  + H2O → HCO3- + OH-

 Kb = (1.0 x 10-14 ) / (5.6 x 10-11 ) = 0.18 x 10-3

 Kb = [HCO3- ] [ OH- ] / [CO32- ] = 0.18 x 10-3

 X2 /( 0.08 – X) = 0.18 x 10-3  X = 3.6 x 10-3  = [OH- ]

 [H+ ] [OH- ] = 10-14  [H+ ] = (1.0 x 10-14 ) / ( 3.8 x 10-3 ) = 0.263 x 10-11  M

 pH = - log [H+ ] = - log (0.263 x 10-11 ) = 11.58

 c) NaNO2 → Na+  + NO2-

 NO2-  + H2O → HNO2  + OH-

**16.85**

 Predict whether aqueous solutions of the following compounds are acidic, basic or neutral

 a) NH4Br b) FeCl3  c) Na2CO3  d) HClO4  e) NaHC2O4

 Answer:

 a) NH4Br → NH4 +  + Br-

 NH4+  is the conjugate acid of base NH3

 Br- is conjugate base of a strong acid HBr , it has no influence on pH.

 The solution of the salt is acidic

 b) FeCl3  → Fe3+  + 3Cl-

 Fe3+  is not from group 1A or group 2A, decrease the pH

 Cl-  is the conjugate base of strong HCl , has no influence on pH

 The solution of the salt is acidic

 c) Na2CO3 → 2Na+  + CO32-

 Na+  is from group 1A has no effect on pH

 CO32-  is a conjugate base of weak acid H2CO3

The solution of the salt is basic

 d) KClO4  → K+  + ClO4 -

 K+ is from group 1A has no effect on pH

 ClO4-  is a conjugate base of the acid HClO4 has no effect on pH

 The solution of the salt is neutral

 e) NaHC2O4 → Na+ + HC2O4-

 Na+  is from group 1A has no influence on pH

 HC2O4 –  is a conjugate acid of a base C2O42-

 The solution of the salt is acidic

**16.87**

 An unknown salt is either NaF, NaCl , or NaOCl. When 0.05 mole of salt is dissolved in water

 to form 0.500 L of solution. The pH of solution is 8.08. What is the identity of the salt?

 ( Kb for the F-  = 1.5 x 10-11 , Kb for ClO-  = 0.334 x 10-6 , Cl-  is from strong acid HCl)

Answer:

 pH + pOH = 14

 8.08 + pOH = 14 pOH = 5.92

 pOH = - log [OH- ] = 5.92

 [OH-] = 1.2 x 10-6  M

 Kb = (X)(X) / 0.1 = ( 1.20 x 10-6 )(1.2 x 10-6 ) / 0.1 = 1.45 x 10-11

 Kb x Ka = 10-14  Ka = (1.0 x 10-14 ) / (1.45 x 10-11 ) = 7.14 x 10-3

 Kb for the anion salt is 1.5 x 10-11 , Ka for the conjugate acid = 7.14 x 10-3 . The conjugate

 acid is F-  , the salt is NaF.