Problems, Chapter 17 (without solutions)

NOTE: Unless otherwise stated, assume T = 25.°C in all problems)

1) In which of these solutions will HNO_2 ionize less than it does in pure water?

a) 0.10 M NaCl
b) 0.10 M NaOH
c) 0.10 M KNO_3
d) 0.10 M $NaNO_2$

2) (17.6) Determine the pH of the following solutions:

a) A solution that is 0.20 M in NH_3 ($K_b(NH_3) = 1.8 \times 10^{-5}$)
b) A solution that is 0.20 M in NH_3 and 0.30 in NH_4Cl.

3) Solve an equilibrium problem (using an ICE table) to calculate the pH of a solution that is 0.195 M in $HC_2H_3O_2$ and 0.125 M in $KC_2H_3O_2$. Note that $K_a = 1.8 \times 10^{-5}$ for $HC_2H_3O_2$.

4) A buffer contains significant amounts of bromous acid ($HBrO_2$) and sodium bromite ($NaBrO_2$). Write equations showing how this buffer neutralizes added acid and added base.

5) Use the Henderson-Hasselbalch equation to calculate the pH of a solution that is 0.135 M in HClO and 0.155 M in KClO. $K_a = 2.9 \times 10^{-8}$ for HClO.

6) A student prepares 100.0 mL of a 0.120 M aqueous solution of hypochlorous acid (HClO, MW = 52.46 g/mol).

a) What is the pH of the above solution? Note that for chlorous acid at T = 25. °C, $K_a = 3.5 \times 10^{-8}$.

b) The student adds 1.00 g of potassium hypochlorite (KClO, MW = 75.56 g/mol) to the above solution of hypochlorous acid. What is the new value for the pH of the solution? You may assume there is no change in volume when the potassium hypochlorite is added.

c) Are either of the above solutions buffer solutions? Justify your answer.

7) (17.18) Which of the following solutions can act as a buffer?

a) KCN/HCN
c) NH_3/NH_4NO_3
b) $Na_2SO_4/NaHSO_4$
d) NaI/HI

8) What mass of ammonium chloride (NH_4Cl) should you add to 2.55 L of a 0.155 M NH_3 solution to obtain a buffer with a pH of 9.55?

9) (17.27) A 0.2688 g sample of a monoprotic acid neutralizes 16.4 mL of a 0.08133 M solution of KOH. Find the molar mass of the acid.

10) Potassium hydrogen phthalate ($KHC_8H_4O_2$, MW = 204.22 g/mol) is a solid ionic compound that forms a weak monoprotic acid, hydrogen phthalate ($HC_8H_4O_2^-$), when added to water. Because it can be obtained at high purity, potassium hydrogen phthalate is often used to standardize stock solutions of strong bases.

A 0.7184 g sample of potassium hydrogen phthalate is dissolved in water, and then titrated with a stock solution of potassium hydroxide (KOH, MW = 56.11 g/mol). After 24.82 mL of stock KOH solution is added the end point of the titration is reached.

a) Write the balanced equation for the reaction of hydrogen phthalate ion with potassium hydroxide.

b) What is the molarity of the stock solution of potassium hydroxide?

11) (17.30) In a titration experiment, 20.4 mL of a 0.833 M solution of HCOOH neutralizes 19.3 mL of a solution of $Ba(OH)_2$. What is the concentration of the $Ba(OH)_2$ solution?

12) For the following titrations indicate whether the pH at the equivalence point of the titration will be <u>much larger than 7.0</u>, <u>approximately 7.0</u>, or <u>much smaller than 7.0</u>.

a) Titration of CH_3COOH (a weak acid) with KOH (a strong soluble base).

b) Titration of HCl (a strong acid) with KOH (a strong soluble base).

13) Consider the titration of a 35.0 mL sample of 0.175 M HBr with a 0.200 M solution of KOH. Determine each quantity:

a) The initial pH

b) The volume of added base required to reach the equivalence point

c) The pH after the addition of 10.0 mL of base

d) The pH at the equivalence point

e) A suitable indicator for the titration (see Table 17.3)

f) The pH after adding 5.0 mL of base beyond the equivalence point

14) (17.45 a,c,e) Write balanced equations and solubility product expressions for each of the following compounds: a) CuBr(s) ; c) Ag_2CrO_4(s) ; e) $AuCl_3$(s)

15) Use the given <u>molar solubilities</u> in pure water to calculate K_{sp} for each compound.

a) $BaCrO_4$; molar solubility = 1.08 x 10^{-5} M

b) Ag_2SO_3; molar solubiliy = 1.55 x 10^{-5} M

c) $Pd(SCN)_2$; molar solubility = 2.22 x 10^{-8} M

16) (17.56) The pH of a saturated solution of a metal hydroxide with formula MOH is 9.68. Find K_{sp} for this compound.

17) Calculate the molar solubility of copper II sulfide (CuS) in each liquid or solution. Note that $K_{sp}(CuS)$ = 1.27 x 10^{-36}.

a) pure water

b) 0.25 M $CuCl_2$, which gives 0.25 M Cu^{2+}

c) 0.20 M K_2S , which gives 0.20 M S^{2-}

18) Find the number of moles of lead II bromide ($PbBr_2$, $K_{sp} = 6.6 \times 10^{-6}$) that will dissolve in 1.000 L of each of the above solutions.

a) Pure water.

b) A 0.200 M aqueous solution of potassium bromide (KBr), a soluble ionic compound.

19) (17.78) Which of the following ionic compounds will be more soluble in acid solution than in water: a) $BaSO_4$ b) $PbCl_2$ c) $Fe(OH)_3$ d) $CaCO_3$

20) Calculate the molar solubility of <u>copper II hydroxide</u> ($Cu(OH)_2$, $K_{sp} = 2.2 \times 10^{-20}$) in a solution buffered at each of the following values for pH.

a) pH = 4.0 b) pH = 7.0 c) pH = 9.0

21) (17.99) Calculate the solubility (in g/L) of Ag_2CO_3. $K_{sp} = 8.1 \times 10^{-12}$.

22) (17.106) Cacodylic acid ($(CH_3)_2AsO_2H$) has an ionization constant $K_a = 6.4 \times 10^{-7}$ at T = 25. °C.

a) Find the pH of 50.0 mL of a 0.100 M solution of cacodylic acid.

b) Find the pH of 25.0 mL of a 0.150 M solution of $(CH_3)_2AsO_2Na$, the sodium salt of the conjugate base of cacodylic acid.

c) The solutions in a and b are mixed. Find the pH of the resulting solution.