Chapter 16 Acid-Base Equilibria and Solubility Equilibria

Student: ___________________________________________________________________________

NOTE: A table of ionization constants and $K_a$'s is required to work some of the problems in this chapter.

1. In which one of the following solutions will acetic acid have the greatest percent ionization?
   A. 0.1 M CH₃COOH
   B. 0.1 M CH₃COOH dissolved in 1.0 M HCl
   C. 0.1 M CH₃COOH plus 0.1 M CH₃COONa
   D. 0.1 M CH₃COOH plus 0.2 M CH₃COONa

2. Which one of the following is a buffer solution?
   A. 0.40 M HCN and 0.10 KCN
   B. 0.20 M CH₃COOH
   C. 1.0 M HNO₃ and 1.0 M NaNO₃
   D. 0.10 M KCN
   E. 0.50 M HCl and 0.10 NaCl

3. Which one of the following combinations cannot function as a buffer solution?
   A. HCN and KCN
   B. NH₃ and (NH₄)₂SO₄
   C. HNO₃ and NaNO₃
   D. HF and NaF
   E. HNO₂ and NaNO₂

4. Which of the following is the most acidic solution?
   A. 0.10 M CH₃COOH and 0.10 M CH₃COONa
   B. 0.10 M CH₃COOH
   C. 0.10 M HNO₂
   D. 0.10 M HNO₂ and 0.10 M NaNO₂
   E. 0.10 M CH₃COONa
5. Calculate the pH of a buffer solution that contains 0.25 M benzoic acid (C₆H₅CO₂H) and 0.15 M sodium benzoate (C₆H₅COONa). [\(K_a = 6.5 \times 10^{-5}\) for benzoic acid]
   
   A. 3.97  
   B. 4.83  
   C. 4.19  
   D. 3.40  
   E. 4.41 

6. A solution is prepared by mixing 500. mL of 0.10 M NaOCl and 500. mL of 0.20 M HOCl. What is the pH of this solution? [\(K_a(\text{HOCl}) = 3.2 \times 10^{-8}\)]
   
   A. 4.10  
   B. 7.00  
   C. 7.19  
   D. 7.49  
   E. 7.80 

7. Calculate the pH of a buffer solution prepared by dissolving 0.20 mole of cyanic acid (HCNO) and 0.80 mole of sodium cyanate (NaCNO) in enough water to make 1.0 liter of solution. [\(K_a(\text{HCNO}) = 2.0 \times 10^{-4}\)]
   
   A. 0.97  
   B. 3.10  
   C. 4.40  
   D. 3.70  
   E. 4.30 

8. Calculate the pH of a solution that is 0.410 M in HOCl and 0.050 M in NaOCl. [\(K_a(\text{HOCl}) = 3.2 \times 10^{-8}\)]
   
   A. 0.39  
   B. 3.94  
   C. 6.58  
   D. 7.49  
   E. 8.40 

9. Calculate the pH of a buffer solution prepared by dissolving 0.20 mole of sodium cyanate (NaCNO) and 1.0 mole of cyanic acid (HCNO) in enough water to make 1.0 liter of solution. [\(K_a(\text{HCNO}) = 2.0 \times 10^{-4}\)]
   
   A. 0.0  
   B. 3.0  
   C. 3.7  
   D. 4.4  
   E. 5.0
10. You are asked to go into the lab and prepare an acetic acid - sodium acetate buffer solution with a pH of 4.00 ± 0.02. What molar ratio of CH₃COOH to CH₃COONa should be used?
   A. 0.18  
   B. 0.84  
   C. 1.19  
   D. 5.50  
   E. 0.10

11. What is the net ionic equation for the reaction that occurs when small amounts of hydrochloric acid are added to a HOCl/NaOCl buffer solution?
   A. H⁺ + H₂O → H₃O⁺  
   B. H⁺ + OCl⁻ → HOCl  
   C. HOCl → H⁺ + OCl⁻  
   D. H⁺ + HOCl → H₂OCl⁺  
   E. HCl + HOCl → H₂O + Cl₂

12. Consider a buffer solution prepared from HOCl and NaOCl. Which is the net ionic equation for the reaction that occurs when NaOH is added to this buffer?
   A. OH⁻ + HOCl → H₂O + OCl⁻  
   B. OH⁻ + OCl⁻ → HOCl + O²⁻  
   C. Na⁺ + HOCl → NaCl + OH⁻  
   D. H⁺ + HOCl → H₂ + OCl⁻  
   E. NaOH + HOCl → H₂O + NaCl

13. Over what range of pH is a HOCl - NaOCl buffer effective?
   A. pH 2.0 - pH 4.0  
   B. pH 7.5 - pH 9.5  
   C. pH 6.5 - pH 8.5  
   D. pH 6.5 - pH 9.5  
   E. pH 1.0 - pH 14.0

14. Assuming equal concentrations of conjugate base and acid, which one of the following mixtures is suitable for making a buffer solution with an optimum pH of 9.2-9.3?
   A. CH₃COONa/CH₃COOH (Kₐ = 1.8 × 10⁻⁵)  
   B. NH₃/NH₄Cl (Kₐ = 5.6 × 10⁻¹⁰)  
   C. NaOCl/HOCl (Kₐ = 3.2 × 10⁻⁸)  
   D. NaN₂/HNO₂ (Kₐ = 4.5 × 10⁻⁵)  
   E. NaCl/HCl
15. Assuming equal concentrations of conjugate base and acid, which one of the following mixtures is suitable for making a buffer solution with an optimum pH of 4.6-4.8?

A. \( \text{CH}_3\text{COO}_2\text{Na}/\text{CH}_3\text{COOH} \) (\( K_a = 1.8 \times 10^{-5} \))
B. \( \text{NH}_3/\text{NH}_4\text{Cl} \) (\( K_a = 5.6 \times 10^{-10} \))
C. \( \text{NaOCl}/\text{HOCl} \) (\( K_a = 3.2 \times 10^{-8} \))
D. \( \text{NaNO}_2/\text{HNO}_2 \) (\( K_a = 4.5 \times 10^{-4} \))
E. \( \text{NaCl}/\text{HCl} \)

16. You have 500.0 mL of a buffer solution containing 0.20 M acetic acid (\( \text{CH}_3\text{COOH} \)) and 0.30 M sodium acetate (\( \text{CH}_3\text{COONa} \)). What will the pH of this solution be after the addition of 20.0 mL of 1.00 M NaOH solution? \( K_a = 1.8 \times 10^{-5} \)

A. 4.41
B. 4.74
C. 4.56
D. 4.92
E. 5.07

17. You have 500.0 mL of a buffer solution containing 0.30 M acetic acid (\( \text{CH}_3\text{COOH} \)) and 0.20 M sodium acetate (\( \text{CH}_3\text{COONa} \)). What will the pH of this solution be after the addition of 20.0 mL of 1.00 M NaOH solution? \( K_a = 1.8 \times 10^{-5} \)

A. 4.65
B. 4.71
C. 4.56
D. 4.84
E. 5.07

18. Calculate the percent ionization of cyanic acid, \( K_a = 2.0 \times 10^{-4} \), in a buffer solution that is 0.50 M HCNO and 0.10 M NaCNO.

A. 0.02%
B. 0.10%
C. 0.20%
D. 2.0%
E. 20%

19. In which one of the following solutions will acetic acid have the greatest percent ionization?

A. 0.1 M \( \text{CH}_3\text{COOH} \)
B. 0.1 M \( \text{CH}_3\text{COOH} \) dissolved in 0.1 M HCl
C. 0.1 M \( \text{CH}_3\text{COOH} \) dissolved in 0.2 M HCl
D. 0.1 M \( \text{CH}_3\text{COOH} \) plus 0.1 M \( \text{CH}_3\text{COONa} \)
E. 0.1 M \( \text{CH}_3\text{COOH} \) plus 0.2 M \( \text{CH}_3\text{COONa} \)
20. On the basis of the information below, a buffer with a pH = 9 can best be made by using

\[
\begin{array}{|c|c|}
\hline
\text{Acid} & \text{\(K_a\)} \\
\hline
\text{H}_3\text{PO}_4 & 7 \times 10^{-3} \\
\text{H}_2\text{PO}_4^- & 8 \times 10^{-8} \\
\text{HPO}_4^{2-} & 5 \times 10^{-13} \\
\hline
\end{array}
\]

A. pure NaH$_2$PO$_4$;
B. H$_2$PO$_4^-$ + PO$_4^{3-}$;
C. H$_2$PO$_4^-$ + HPO$_4^{2-}$;
D. HPO$_4^{2-}$ + PO$_4^{3-}$.

21. The pH at the equivalence point of a titration may differ from 7.0 due to

A. the initial concentration of the standard solution.
B. the indicator used.
C. the self-ionization of H$_2$O.
D. the initial pH of the unknown.
E. hydrolysis of the salt formed.

22. For which type of titration will the pH be basic at the equivalence point?

A. Strong acid vs. strong base.
B. Strong acid vs. weak base.
C. Weak acid vs. strong base.
D. all of the these
E. none of these

23. 50.00 mL of 0.10 M HNO$_2$ (nitrous acid, \(K_a = 4.5 \times 10^{-4}\)) is titrated with a 0.10 M KOH solution. After 25.00 mL of the KOH solution is added, the pH in the titration flask will be

A. 2.17
B. 3.35
C. 2.41
D. 1.48
E. 7.00

24. A titration of an acid and base to the equivalence point results in a noticeably acidic solution. It is likely this titration involves

A. a strong acid and a weak base.
B. a weak acid and a strong base.
C. a weak acid and a weak base (where \(K_a\) equals \(K_b\)).
D. a strong acid and a strong base.
25. Calculate the pH at the equivalence point for the titration of 0.20 M HCl with 0.20 M NH₃ (K_b = 1.8 \times 10^{-5}).

A. 2.87  
B. 4.98  
C. 5.12  
D. 7.00  
E. 11.12

26. What is the pH at the equivalence point in the titration of 100 mL of 0.10 M HCl with 0.10 M NaOH?

A. 1.0  
B. 6.0  
C. 7.0  
D. 8.0  
E. 13.0

27. What is the pH at the equivalence point in the titration of 100 mL of 0.10 M HCN (K_a = 4.9 \times 10^{-10}) with 0.10 M NaOH?

A. 3.0  
B. 6.0  
C. 7.0  
D. 11.0  
E. 12.0

28. Calculate the pH of the solution resulting from the addition of 10.0 mL of 0.10 M NaOH to 50.0 mL of 0.10 M HCN (K_a = 4.9 \times 10^{-10}) solution.

A. 5.15  
B. 8.71  
C. 5.85  
D. 9.91  
E. 13.0

29. Methyl red is a common acid-base indicator. It has a K_a equal to 6.3 \times 10^{-6}. Its un-ionized form is red and its anionic form is yellow. What color would a methyl red solution have at pH = 7.8?

A. green  
B. red  
C. blue  
D. yellow  
E. violet
30. What mass of sodium fluoride must be added to 250. mL of a 0.100 M HF solution to give a buffer solution having a pH of 3.50? \( (K_a(HF) = 7.1 \times 10^{-4}) \)

A. 0.49 g  
B. 1.5g  
C. 3.4g  
D. 2.3g  
E. 0.75 g

31. What mass of ammonium nitrate must be added to 350. mL of a 0.150 M solution of ammonia to give a buffer having a pH of 9.00? \( (K_b(NH_3) = 1.8 \times 10^{-5}) \)

A. 7.6 g  
B. 2.4g  
C. 5.4g  
D. 11g  
E. 3.3g

32. 40.0 mL of an acetic acid of unknown concentration is titrated with 0.100 M NaOH. After 20.0 mL of the base solution has been added, the pH in the titration flask is 5.10. What was the concentration of the original acetic acid solution? \( (K_a(CH_3COOH) = 1.8 \times 10^{-5}) \)

A. 0.11 M  
B. 0.022 M  
C. 0.072 M  
D. 0.050 M  
E. 0.015 M

33. 25.0 mL of a hydrofluoric acid solution of unknown concentration is titrated with 0.200 M NaOH. After 20.0 mL of the base solution has been added, the pH in the titration flask is 3.00. What was the concentration of the original hydrofluoric acid solution. \( (K_a(HF) = 7.1 \times 10^{-4}) \)

A. 0.39 M  
B. 0.27 M  
C. 0.16 M  
D. 2.4M  
E. 0.23 M

34. For PbCl\(_2\) \( (K_{sp} = 2.4 \times 10^{-4}) \), will a precipitate of PbCl\(_2\) form when 0.10 L of 3.0 \( \times \) 10\(^{-2}\) M Pb(NO\(_3\))\(_2\) is added to 400 mL of 9.0 \( \times \) 10\(^{-2}\) M NaCl?

A. Yes, because \( Q > K_{sp} \).  
B. No, because \( Q < K_{sp} \).  
C. No, because \( Q = K_{sp} \).  
D. Yes, because \( Q < K_{sp} \).
35. The solubility of lead(II) iodide is 0.064 g/100 mL at 20°C. What is the solubility product for lead(II) iodide?

A. $1.1 \times 10^{-8}$  
B. $3.9 \times 10^{-6}$  
C. $1.1 \times 10^{-11}$  
D. $2.7 \times 10^{-12}$  
E. $1.4 \times 10^{-3}$

36. The molar solubility of magnesium carbonate is $1.8 \times 10^{-4}$ mol/L. What is $K_{sp}$ for this compound?

A. $1.8 \times 10^{-4}$  
B. $3.6 \times 10^{-4}$  
C. $1.3 \times 10^{-7}$  
D. $3.2 \times 10^{-8}$  
E. $2.8 \times 10^{-14}$

37. The molar solubility of manganese(II) carbonate is $4.2 \times 10^{-6}$ M. What is $K_{sp}$ for this compound?

A. $4.2 \times 10^{-6}$  
B. $8.4 \times 10^{-6}$  
C. $3.0 \times 10^{-16}$  
D. $1.8 \times 10^{-11}$  
E. $2.0 \times 10^{-3}$

38. The molar solubility of tin(II) iodide is $1.28 \times 10^{-2}$ mol/L. What is $K_{sp}$ for this compound?

A. $8.4 \times 10^{-6}$  
B. $1.28 \times 10^{-2}$  
C. $4.2 \times 10^{-6}$  
D. $1.6 \times 10^{-4}$  
E. $2.1 \times 10^{-6}$

39. The solubility of strontium carbonate is 0.0011 g/100 mL at 20°C. Calculate the $K_{sp}$ value for this compound.

A. $7.5 \times 10^{-6}$  
B. $1.5 \times 10^{-7}$  
C. $5.5 \times 10^{-9}$  
D. $7.5 \times 10^{-6}$  
E. $1.5 \times 10^{-3}$
40. The molar solubility of lead(II) iodate in water is $4.0 \times 10^{-5}$ mol/L. Calculate $K_{sp}$ for lead(II) iodate.

A. $1.6 \times 10^{-9}$
B. $6.4 \times 10^{-14}$
C. $2.6 \times 10^{-13}$
D. $4.0 \times 10^{-5}$
E. $4.0 \times 10^{-15}$

41. The solubility product for chromium(III) fluoride is $K_{sp} = 6.6 \times 10^{-11}$. What is the molar solubility of chromium(III) fluoride?

A. $1.6 \times 10^{-3}$ M
B. $1.2 \times 10^{-3}$ M
C. $6.6 \times 10^{-11}$ M
D. $2.2 \times 10^{-3}$ M
E. $1.6 \times 10^{-6}$ M

42. The solubility product for barium sulfate is $1.1 \times 10^{-10}$. Calculate the molar solubility of barium sulfate.

A. $5.5 \times 10^{-11}$ mol/L
B. $1.1 \times 10^{-5}$ mol/L
C. $2.1 \times 10^{-5}$ mol/L
D. $1.1 \times 10^{-10}$ mol/L
E. $2.2 \times 10^{-10}$ mol/L

43. The $K_{sp}$ for silver(I) phosphate is $1.8 \times 10^{-18}$. Calculate the molar solubility of silver(I) phosphate.

A. $1.6 \times 10^{-5}$ M
B. $2.1 \times 10^{-5}$ M
C. $3.7 \times 10^{-5}$ M
D. $7.2 \times 10^{-1}$ M
E. $1.8 \times 10^{-1}$ M

44. The solubility product for calcium phosphate is $K_{sp} = 1.3 \times 10^{-26}$. What is the molar solubility of calcium phosphate?

A. $1.3 \times 10^{-6}$ M
B. $1.5 \times 10^{-6}$ M
C. $2.6 \times 10^{-6}$ M
D. $4.6 \times 10^{-6}$ M
E. $6.6 \times 10^{-6}$ M
45. The $K_{sp}$ value for lead(II) chloride is $2.4 \times 10^{-4}$. What is the molar solubility of lead(II) chloride?

A. $2.4 \times 10^{-4}$ mol/L  
B. $6.2 \times 10^{-2}$ mol/L  
C. $7.7 \times 10^{-3}$ mol/L  
D. $3.9 \times 10^{-2}$ mol/L  
E. $6.0 \times 10^{-5}$ mol/L

46. Calculate the silver ion concentration in a saturated solution of silver(I) carbonate ($K_{sp} = 8.1 \times 10^{-12}$).

A. $5.0 \times 10^{-5}$ M  
B. $2.5 \times 10^{-4}$ M  
C. $1.3 \times 10^{-4}$ M  
D. $2.0 \times 10^{-4}$ M  
E. $8.1 \times 10^{-4}$ M

47. The $K_{sp}$ for silver(I) phosphate is $1.8 \times 10^{-18}$. Determine the silver ion concentration in a saturated solution of silver(I) phosphate.

A. $1.6 \times 10^{-5}$ M  
B. $2.1 \times 10^{-5}$ M  
C. $3.7 \times 10^{-5}$ M  
D. $1.1 \times 10^{-13}$ M  
E. $4.8 \times 10^{-5}$ M

48. Calculate the silver ion concentration in a saturated solution of silver(I) sulfate ($K_{sp} = 1.4 \times 10^{-5}$).

A. $1.5 \times 10^{-2}$ M  
B. $2.4 \times 10^{-2}$ M  
C. $3.5 \times 10^{-2}$ M  
D. $1.4 \times 10^{-5}$ M  
E. none of these

49. Calculate the concentration of chloride ions in a saturated lead(II) chloride ($K_{sp} = 2.4 \times 10^{-4}$) solution.

A. $2.4 \times 10^{-2}$ M  
B. $4.8 \times 10^{-1}$ M  
C. $3.9 \times 10^{-1}$ M  
D. $1.2 \times 10^{-1}$ M  
E. $7.8 \times 10^{-2}$ M
50. Calculate the concentration of fluoride ions in a saturated barium fluoride (K_{sp} = 1.7 \times 10^{-6}) solution.

A. 7.6 \times 10^{-3} M  
B. 1.5 \times 10^{-2} M  
C. 3.4 \times 10^{-5} M  
D. 1.7 \times 10^{-6} M  
E. 3.4 \times 10^{-6} M

51. Which of the following would decrease the K_{sp} for PbI_2?

A. Lowering the pH of the solution 
B. Adding a solution of Pb(NO_3)_2 
C. Adding a solution of KI 
D. None of these—the K_{sp} of a compound is constant at constant temperature.

52. Calculate the minimum concentration of Mg^{2+} that must be added to 0.10 M NaF in order to initiate a precipitate of magnesium fluoride. (For MgF_2, K_{sp} = 6.9 \times 10^{-9}).

A. 1.4 \times 10^{7} M  
B. 6.9 \times 10^{-2} M  
C. 6.9 \times 10^{-8} M  
D. 1.7 \times 10^{-7} M  
E. 6.9 \times 10^{-7} M

53. Calculate the minimum concentration of Cr^{3+} that must be added to 0.095 M NaF in order to initiate a precipitate of chromium(III) fluoride. (For CrF_3, K_{sp} = 6.6 \times 10^{-11}).

A. 0.023 M  
B. 0.032 M  
C. 7.7 \times 10^{-8} M  
D. 2.9 \times 10^{-9} M  
E. 6.9 \times 10^{-10} M

54. Will a precipitate form (yes or no) when 50.0 mL of 1.2 \times 10^{-3} M Pb(NO_3)_2 are added to 50.0 mL of 2.0 \times 10^{-4} M Na_2S? If so, identify the precipitate.

A. Yes, the precipitate is PbS. 
B. Yes, the precipitate is NaNO_3. 
C. Yes, the precipitate is Na_2S. 
D. Yes, the precipitate is Pb(NO_3)_2. 
E. No, a precipitate will not form.
55. Will a precipitate of magnesium fluoride form when 300. mL of $1.1 \times 10^{-3}$ M MgCl$_2$ are added to 500. mL of $1.2 \times 10^{-3}$ M NaF? ($K_{sp}$ (MgF$_2$) = $6.9 \times 10^{-9}$)

A. Yes, $Q > K_{sp}$
B. No, $Q < K_{sp}$
C. No, $Q = K_{sp}$
D. Yes, $Q < K_{sp}$

56. Will a precipitate of magnesium fluoride form when 200. mL of $1.9 \times 10^{-3}$ M MgCl$_2$ are added to 300. mL of $1.4 \times 10^{-2}$ M NaF? ($K_{sp}$ (MgF$_2$) = $6.9 \times 10^{-9}$)

A. Yes, $Q > K_{sp}$
B. No, $Q < K_{sp}$
C. No, $Q = K_{sp}$
D. Yes, $Q < K_{sp}$

57. Will a precipitate (ppt) form when 300. mL of $5.0 \times 10^{-5}$ M AgNO$_3$ are added to 200. mL of $2.5 \times 10^{-7}$ M NaBr? Answer yes or no, and identify the precipitate if there is one.

A. Yes, the ppt is AgNO$_3$(s).
B. Yes, the ppt is AgBr(s).
C. Yes, the ppt is NaBr(s).
D. Yes, the ppt is NaNO$_3$(s).
E. No, a precipitate will not form.

58. Will a precipitate (ppt) form when 20.0 mL of $1.1 \times 10^{-3}$ M Ba(NO$_3$)$_2$ are added to 80.0 mL of $8.4 \times 10^{-4}$ M Na$_2$CO$_3$?

A. Yes, the ppt is Ba(NO$_3$)$_2$.
B. Yes, the ppt is NaNO$_3$.
C. Yes, the ppt is BaCO$_3$.
D. Yes, the ppt is Na$_2$CO$_3$.
E. No, a precipitate will not form.

59. Will a precipitate (ppt) form when 300. mL of $2.0 \times 10^{-5}$ M AgNO$_3$ are added to 200. mL of $2.5 \times 10^{-9}$ M NaI? Answer yes or no, and identify the precipitate if there is one.

A. Yes, the ppt is AgNO$_3$(s).
B. Yes, the ppt is NaNO$_3$(s).
C. Yes, the ppt is NaI(s).
D. Yes, the ppt is AgI(s).
E. No, a precipitate will not form.
60. Which response has both answers correct? Will a precipitate form when 250 mL of 0.33 M Na₂CrO₄ are added to 250 mL of 0.12 M AgNO₃? (K_{sp}(Ag₂CrO₄) = 1.1 \times 10^{-12}) What is the concentration of the silver ion remaining in solution?

A. Yes, [Ag⁺] = 2.9 \times 10^{-6} M.
B. Yes, [Ag⁺] = 0.060 M.
C. Yes, [Ag⁺] = 1.3 \times 10^{-4} M.
D. No, [Ag⁺] = 0.060 M.
E. No, [Ag⁺] = 0.105 M.

61. To 1.00 L of a 0.100 M aqueous solution of benzoic acid (C₆H₅COOH) is added 1.00 mL of 12.0 M HCl. What is the percentage ionization of the benzoic acid in the resulting solution? [Given: K_a(C₆H₅COOH) = 6.5 \times 10^{-5}]

A. 3.3%
B. 12%
C. 1.3%
D. 0.52%
E. 0.065%

62. To 1.00 L of a 0.100 M aqueous solution of the weak base pyridine (C₅H₅N) is added 1.00 mL of 14.0 M NaOH. What fraction of the pyridine molecules in the resulting solution react with water to release hydroxide ions? [Given: K_b(C₅H₅N) = 1.7 \times 10^{-9}]

A. 1.2 \times 10^{-7}
B. 4.1 \times 10^{-5}
C. 0.14
D. 2.4 \times 10^{-10}
E. 1.3 \times 10^{-5}

63. Find the concentration of Pb²⁺ ions in a solution made by adding 5.00 g of lead(II) iodide to 500. mL of 0.150 M KI. [For PbI₂, K_{sp} = 1.39 \times 10^{-8}.]

A. 3.04 \times 10^{-4} M
B. 1.54 \times 10^{-5} M
C. 6.18 \times 10^{-6} M
D. 1.52 \times 10^{-4} M
E. 9.27 \times 10^{-8} M
64. Find the concentration of calcium ions in a solution made by adding 3.50 g of calcium fluoride to 750. mL of 0.125 M NaF. [For CaF₂, K_{sp} = 3.95 \times 10^{-11}.]

A. \(3.16 \times 10^{-10}\) M
B. \(2.53 \times 10^{-9}\) M
C. \(4.29 \times 10^{-4}\) M
D. \(6.32 \times 10^{-10}\) M
E. \(2.15 \times 10^{-4}\) M

65. A saturated sodium carbonate solution at 0°C contains 7.1 g of dissolved sodium carbonate per 100 mL of solution. The solubility product constant for sodium carbonate at this temperature is

A. 1.2
B. 0.30
C. \(3.0 \times 10^{-4}\)
D. 0.90
E. \(1.2 \times 10^{-3}\)

66. A saturated sodium carbonate solution at 100°C contains 45.5 g of dissolved sodium carbonate per 100 mL of solution. The solubility product constant for sodium carbonate at this temperature is

A. 79.0
B. 0.316
C. 0.0790
D. 36.8
E. 316

67. What volume of 0.0500 M sodium hydroxide should be added to 250 mL of 0.100 M HCOOH to obtain a solution with a pH of 4.50? [K_a(HCOOH) = 1.7 \times 10^{-4}]

A. 540 mL
B. 420 mL
C. 80 mL
D. 340 mL
E. 500 mL

68. What volume of 0.200 M potassium hydroxide should be added to 300 mL of 0.150 M propanoic acid (C_2H_5COOH) to obtain a solution with a pH of 5.25? [K_a(C_2H_5COOH) = 1.34 \times 10^{-5}]

A. 32 mL
B. 210 mL
C. 160 mL
D. 65 mL
E. 13 mL
69. What is the effective pH range for a sodium acetate/acetic acid buffer? (For CH₃COOH, $K_a = 1.8 \times 10^{-5}$)

70. Calculate the pH of a solution that is 0.15 M CH₃COOH and 0.75 M CH₃COONa.

71. What is the optimum pH of a sodium formate/formic acid buffer? (For formic acid, $K_a = 1.7 \times 10^{-4}$)

72. Describe how to make a sodium formate (HCOONa)/formic acid (HCOOH) buffer that has a pH of 4.77.

73. Write an equation showing the net reaction that occurs when a strong acid is added to a CO₃²⁻/HCO₃⁻ buffer solution (for carbonic acid, $K_{a1} = 4.2 \times 10^{-7}$, $K_{a2} = 2.4 \times 10^{-8}$).

74. Write an equation showing the net reaction that occurs when a strong base is added to a CO₃²⁻/HCO₃⁻ buffer solution (for carbonic acid, $K_{a1} = 4.2 \times 10^{-7}$, $K_{a2} = 2.4 \times 10^{-8}$).
75. Calculate the percent ionization of formic acid in a 0.010 M HCOOH solution. \( K_a = 1.7 \times 10^{-4} \)

76. Calculate the percent ionization of formic acid in a solution that is 0.010 M HCOOH and 0.005 M HCOONa and compare your answer to the percent ionization you would calculate if the sodium formate were not present. Explain the difference, if any. \( K_a = 1.7 \times 10^{-4} \)

77. Calculate the percent ionization of formic acid in a solution that is 0.010 M HCOOH and 0.050 M HCOONa. \( K_a = 1.7 \times 10^{-4} \)

78. What molar ratio of benzoate ion to benzoic acid would be required to prepare a buffer with a pH of 5.20? \( K_a(C_6H_5COOH) = 6.5 \times 10^{-5} \)

79. Write a net ionic equation for the reaction that occurs when a small amount of hydrochloric acid is added to a buffer solution containing \( \text{NH}_4\text{Cl} \) and \( \text{NH}_3 \).

80. Write a net ionic equation for the reaction occurring when a small amount of sodium hydroxide solution is added to a buffer solution containing \( \text{NH}_4\text{Cl} \) and \( \text{NH}_3 \).
81. Write a net ionic equation for the reaction that occurs when a small amount of nitric acid is added to a NaNO₂/HNO₂ buffer.

82. Write a net ionic equation for the reaction occurring when a small amount of sodium hydroxide is added to a NaNO₂/HNO₂ buffer.

83. Calculate the pH at the equivalence point for the titration of 0.25 M CH₃COOH with 0.25 M NaOH. (For CH₃COOH, Kₐ = 1.8 × 10⁻⁵)

84. Calculate the pH at the equivalence point for the titration of 0.22 M HCN with 0.22 M NaOH. (Kₐ = 4.9 × 10⁻¹⁰ for HCN.)

85. Bromothymol blue is a common acid-base indicator. It has a Kₐ equal to 1.6 × 10⁻⁷. Its un-ionized form is yellow and its conjugate base is blue. What color would a solution have at pH = 5.8?

86. Describe how to prepare 500 mL of a cyanic acid (HCNO)/sodium cyanate (NaCNO) buffer having a pH of 4.80. (Kₐ(HCNO) = 2.0 × 10⁻⁴)
87. Solid sodium iodide is slowly added to a solution that is 0.0050 M Pb\(^{2+}\) and 0.0050 M Ag\(^+\).
(K\(_{sp}\) (PbI\(_2\)) = 1.4 \times 10^{-8}; K\(_{sp}\) (AgI) = 8.3 \times 10^{-17})

What compound will precipitate first?

88. Solid sodium iodide is slowly added to a solution that is 0.0050 M Pb\(^{2+}\) and 0.0050 M Ag\(^+\).
(K\(_{sp}\) (PbI\(_2\)) = 1.4 \times 10^{-8}; K\(_{sp}\) (AgI) = 8.3 \times 10^{-17})

Calculate the Ag\(^+\) concentration when PbI\(_2\) just begins to precipitate.

89. Solid sodium iodide is slowly added to a solution that is 0.0050 M Pb\(^{2+}\) and 0.0050 M Ag\(^+\).
(K\(_{sp}\) (PbI\(_2\)) = 1.4 \times 10^{-8}; K\(_{sp}\) (AgI) = 8.3 \times 10^{-17})

What percent of Ag\(^+\) remains in solution at this point?

90. The K\(_{sp}\) of CaF\(_2\) is 4 \times 10^{-11}. What is the maximum concentration of Ca\(^{2+}\) possible in a 0.10 M NaF solution?

91. 5.0 mL of 12 M NH\(_3\) is added to 500. mL of 0.050 M AgNO\(_3\). What concentration of silver ion will exist after equilibrium is established? (K\(_f\) for Ag(NH\(_3\))\(_2\)\(^+\) = 1.5 \times 10^7)
92. Will a precipitate of AgCl form when 0.050 mol NaCl(s) and 0.050 mol AgNO₃(s) are dissolved in 500 mL of 3.0 M NH₃? (Kf for Ag(NH₃)₂⁺ = 1.5 × 10⁷; Ksp(AgCl) = 1.6 × 10⁻¹⁰)

93. Will Fe(OH)₃ precipitate from a buffer solution that is 0.60 M CH₃COOH and 0.10 M CH₃COONa, if the solution is also made to be 0.001 M in Fe³⁺? (For Fe(OH)₃, Ksp = 6.8 × 10⁻₃₆.)

94. The concentration of Mg²⁺ in seawater is 5.0 × 10⁻² M. What hydroxide concentration is needed to remove 90% of the Mg²⁺ by precipitation? (For Mg(OH)₂, Ksp = 1.2 × 10⁻¹₁)

95. Ammonium chloride solutions are slightly acidic, so they are better solvents than water for insoluble substances such as Ca(OH)₂. Identify two reactions that, added together, give the overall reaction below. Then determine Kc for the overall reaction. (Ksp for Ca(OH)₂ is 4.0 × 10⁻⁹; K₅ for NH₃ is 1.8 × 10⁻⁵.)

\[
\text{Ca(OH)}_2(aq) + 2\text{NH}_4^+(aq) \rightleftharpoons \text{Ca}^{2+}(aq) + 2\text{NH}_3(aq) + 2\text{H}_2\text{O}
\]

96. Calculate the equilibrium constant Kc for the following overall reaction. (For AgCl, Ksp = 1.6 × 10⁻¹⁰; for Ag(CN)₂⁻, Kf = 1.0 × 10⁻²¹)

\[
\text{AgCl(s)} + 2\text{CN}^-(aq) \rightleftharpoons \text{Ag(CN)}_2^-(aq) + \text{Cl}^-(aq)
\]
97. Calculate the equilibrium constant $K_c$ for the net reaction shown below.

(For AgI, $K_{sp} = 8.3 \times 10^{-17}$; for Ag(NH$_3$)$_2^+$, $K_f = 1.5 \times 10^7$.)

$$\text{AgI(s) + 2NH}_3\text{(aq)} \rightleftharpoons \text{Ag(NH}_3)_2^+(aq) + I^-(aq)$$

98. The solubility of Ba(NO$_3$)$_2$ is 130.5 g/L at 0°C. How many moles of dissolved salt are present in 4.0 L of a saturated solution of Ba(NO$_3$)$_2$ at 0°C?

99. A sample of rainwater collected near a lead smelter is analyzed for acid content. Experiments show that a 100. mL sample of the rainwater is neutralized by 22.4 milliters of 0.0122 M NaOH. Assuming that the acid present is sulfuric acid, which resulted from the reaction of SO$_2$ with water, what is the molarity of acid in the rainwater?

100. A wildlife biologist is interested in testing the pH of the water in a lake. He obtains a 200. mL sample of the water and titrates this sample with a .050 M NaOH solution. Neutralization of the lake water requires 40. mL of the NaOH solution.

Estimate the pH of the lake.
101. A wildlife biologist is interested in testing the pH of the water in a lake. He obtains a 200. mL sample of the water and titrates this sample with a .050 M NaOH solution. Neutralization of the lake water requires 40. mL of the NaOH solution.

Is the lake acidic, basic, or neutral?

102. A wildlife biologist is interested in testing the pH of the water in a lake. He obtains a 200. mL sample of the water and titrates this sample with a .050 M NaOH solution. Neutralization of the lake water requires 40. mL of the NaOH solution?

If the size of the lake can be described as 1.1 km long by 2.3 km wide, and has a depth of 10. m, estimate how many moles of acid are present in the lake water.

103. NaCl is added slowly to a solution that is 0.010 M each in Cu⁺, Ag⁺, and Au⁺. The K_{sp}'s for CuCl, AgCl, and AuCl are $1.9 \times 10^{-7}$, $1.8 \times 10^{-10}$, and $2.0 \times 10^{-13}$, respectively. Which compound will precipitate first?

104. A 50.0 mL sample of $2.0 \times 10^{-4}$ M Cu(NO₃)₂ is added to 50.0 mL of 4.0 M NaCN. The formation constant of the complex ion Cu(CN)₃²⁻ is $1.0 \times 10^{9}$. What is the copper(I) ion concentration in this system at equilibrium?
105. How many moles of NaF must be dissolved in 1.00 liter of a saturated solution of PbF$_2$ at 25°C to reduce the [Pb$^{2+}$] to 1.0 × 10$^{-6}$ M? (K$_{sp}$ of PbF$_2$ at 25°C = 4.0 × 10$^{-8}$)

106. At 25°C, the base ionization constant for NH$_3$ is 1.8 × 10$^{-5}$. Determine the hydroxide ion concentration in a 0.150 M solution of ammonia at 25°C.

107. At 25°C, the base ionization constant for NH$_3$ is 1.8 × 10$^{-5}$. Determine the pH of a solution prepared by adding 0.0500 mol of solid ammonium chloride to 100. mL of 0.150 M ammonia.

108. At 25°C, the base ionization constant for NH$_3$ is 1.8 × 10$^{-5}$. Determine the percentage ionization of a 0.150 M solution of ammonia at 25°C.

109. At 25°C, the base ionization constant for NH$_3$ is 1.8 × 10$^{-5}$. If 0.0800 mole of solid magnesium chloride is dissolved in a solution prepared by adding 0.0500 mol of solid ammonium chloride to 100. mL of 0.150 M ammonia, will a precipitate of magnesium hydroxide form? [Assume the volume of the solution is unchanged. The solubility product constant for magnesium hydroxide is 1.5 × 10$^{-11}$.] 

110. The percent ionization of a weak acid HA is greater in a solution containing the salt NaA than it is in a solution of the weak acid only.

True    False
111. A mixture made from 10 mL of 1 M HCl and 20 mL of 1 M CH₃COONa would be classified as a buffer solution.

   True    False

112. All indicators are weak acids that are one color in acidic solution and another color in basic solution.

   True    False

113. For any conjugate acid-base pair, it is true that Kₘₐₜ = KₐKₗₘₚₜ.

   True    False

114. The pH of a solution that is 0.20 M CH₃COOH and 0.20 M CH₃COONa should be higher than the pH of a 0.20 M CH₃COOH solution.

   True    False
**Chapter 16 Acid-Base Equilibria and Solubility Equilibria Key**

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58. C
59. D
60. A
61. D
62. A
63. C
64. B
65. A
66. E
67. B
68. C
69. 3.7-5.7
70. 5.44
71. 3.77

72. Prepare a solution having a molar ratio of 10.0 mol HCOONa to 1.0 mol HCOOH.

73. $\text{CO}_3^{2-} + \text{H}^+ \rightarrow \text{HCO}_3^-$ and $\text{HCO}_3^- + \text{H}^+ \rightarrow \text{H}_2\text{CO}_3$

74. $\text{HCO}_3^- + \text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}$

75. 13%

76. 3.4%; the addition of formate suppresses the ionization of formic acid by shifting the ionization equilibrium towards the un-ionized acid.

77. 0.34%

78. 10.3

79. $\text{H}^+ + \text{NH}_3 \rightarrow \text{NH}_4^+$

80. $\text{OH}^- + \text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}_2\text{O}$

81. $\text{H}^+ + \text{NO}_2^- \rightarrow \text{HNO}_2$

82. $\text{OH}^- + \text{HNO}_2 \rightarrow \text{NO}_2^- + \text{H}_2\text{O}$

83. 8.92

84. 11.18

85. Yellow

86. Dissolve amounts of the two compounds equal to a molar ratio of 12.6 mol NaCNO to 1.0 mol HCNO in enough water to yield 500. mL of solution.

87. AgI

88. $5.0 \times 10^{-14}$ M

89. $1.0 \times 10^{-9}$ %

90. $4.0 \times 10^{-9}$ M
91. $8.4 \times 10^{-5}$ M
92. Yes
93. Yes
94. $4.9 \times 10^{-5}$ M

95. $\text{Ca(OH)}_2(s) \rightleftharpoons \text{Ca}^{2+}(aq) + 2\text{OH}^-(aq)$
$2\text{NH}_4^+(aq) + 2\text{OH}^-(aq) \rightleftharpoons 2\text{NH}_3(aq) + 2\text{H}_2\text{O}$
$K_c = 120$

96. $1.6 \times 10^{11}$
97. $1.2 \times 10^{-9}$
98. 2.0 moles
99. $1.37 \times 10^{-3}$ M
100. 2
101. acidic
102. $2.5 \times 10^9$ mol
103. AuCl
104. $1.3 \times 10^{-14}$ M
105. 0.20 mol
106. $1.6 \times 10^{-3}$ M
107. 8.73
108. 1.1%
109. Yes
110. FALSE
111. TRUE
112. FALSE
113. FALSE
114. TRUE