CHAPTER 8 QUESTIONS

Multiple-Choice

Use the following information to answer questions 1-5.

A student titrates 20.0 mL of 1.0 *M* NaOH with 2.0 *M* formic acid, HCO_2H ($K_a = 1.8 \times 10^{-4}$). Formic acid is a monoprotic acid.

1. How much formic acid is necessary to reach the equivalence point?
   (A) 10.0 mL
   (B) 20.0 mL
   (C) 30.0 mL
   (D) 40.0 mL

2. At the equivalence point, is the solution acidic, basic, or neutral? Why?
   (A) Acidic; the strong acid dissociates more than the weak base
   (B) Basic; the only ion present at equilibrium is the conjugate base
   (C) Basic; the higher concentration of the base is the determining factor
   (D) Neutral; equal moles of both acid and base are present

3. If the formic acid were replaced with a strong acid such as HCl at the same concentration (2.0 *M*), how would that change the volume needed to reach the equivalence point?
   (A) The change would reduce the amount as the acid now fully dissociates.
   (B) The change would reduce the amount because the base will be more strongly attracted to the acid.
   (C) The change would increase the amount because the reaction will now go to completion instead of equilibrium.
   (D) Changing the strength of the acid will not change the volume needed to reach equivalance.

4. Which of the following would create a good buffer when dissolved in formic acid?
   (A) $NaCO_2H$
   (B) $HC_2H_3O_2$
   (C) NH_3
   (D) H_2O

5. $CH_3NH_2(aq) + H_2O(l) \leftrightarrow OH^-(aq) + CH_3NH_3^+(aq)$

   The above equation represents the reaction between the base methylamine ($K_b = 4.38 \times 10^{-4}$) and water. Which of the following best represents the concentrations of the various species at equilibrium?

   (A) $[OH^-] > [CH_3NH_2] = [CH_3NH_3^+]$
   (B) $[OH^-] = [CH_3NH_2] = [CH_3NH_3^+]$
   (C) $[CH_3NH_2] > [OH^-] > [CH_3NH_3^+]$
   (D) $[CH_3NH_2] > [OH^-] = [CH_3NH_3^+]$

Use the following information to answer questions 6-10.

The following reaction is found to be at equilibrium at 25°C:

$2SO_3(g) \leftrightarrow O_2(g) + 2SO_2(g)$ $\Delta H = -198$ kJ/mol

6. What is the expression for the equilibrium constant, K_c?

(A) $\dfrac{[SO_3]^2}{[O_2][SO_2]^2}$

(B) $\dfrac{2[SO_3]}{[O_2]2[SO_2]}$

(C) $\dfrac{[O_2][SO_2]^2}{[SO_3]^2}$

(D) $\dfrac{[O_2]2[SO_2]}{2[SO_3]}$

7. Which of the following would cause the reverse reaction to speed up?

(A) Adding more SO_3
(B) Raising the pressure
(C) Lowering the temperature
(D) Removing some SO_2

8. The value for K_c at 25°C is 8.1. What must happen in order for the reaction to reach equilibrium if the initial concentrations of all three species was 2.0 *M*?

(A) The rate of the forward reaction would increase, and $[SO_3]$ would decrease.
(B) The rate of the reverse reaction would increase, and $[SO_2]$ would decrease.
(C) Both the rate of the forward and reverse reactions would increase, and the value for the equilibrium constant would also increase.
(D) No change would occur in either the rate of reaction or the concentrations of any of the species.

9. Which of the following would cause a reduction in the value for the equilibrium constant?

(A) Increasing the amount of SO_3
(B) Reducing the amount of O_2
(C) Raising the temperature
(D) Lowering the temperature

10. The solubility product, K_{sp}, of AgCl is 1.8×10^{-10}. Which of the following expressions is equal to the solubility of AgCl?

(A) $\left(1.8\times10^{-10}\right)^2$ molar

(B) $\dfrac{1.8\times10^{-10}}{2}$ molar

(C) 1.8×10^{-10} molar

(D) $\sqrt{1.8\times10^{-10}}$ molar

11. A 0.1-molar solution of which of the following acids will be the best conductor of electricity?

(A) H_2CO_3
(B) H_2S
(C) HF
(D) HNO_3

12. Which of the following expressions is equal to the K_{sp} of Ag_2CO_3?

(A) $K_{sp} = [Ag^+][CO_3^{2-}]$
(B) $K_{sp} = [Ag^+][CO_3^{2-}]^2$
(C) $K_{sp} = [Ag^+]^2[CO_3^{2-}]$
(D) $K_{sp} = [Ag^+]^2[CO_3^{2-}]^2$

13. If the solubility of BaF_2 is equal to x, which of the following expressions is equal to the solubility product, K_{sp}, for BaF_2?

(A) x^2
(B) $2x^2$
(C) $2x^3$
(D) $4x^3$

Use the following information to answer questions 14-16:

150 mL of saturated SrF_2 solution is present in a 250 mL beaker at room temperature. The molar solubility of SrF_2 at 298 K is 1.0×10^{-3} M.

14. What are the concentrations of Sr^{2+} and F^- in the beaker?

(A) $[Sr^{2+}] = 1.0 \times 10^{-3}\ M$ $[F^-] = 1.0 \times 10^{-3}\ M$
(B) $[Sr^{2+}] = 1.0 \times 10^{-3}\ M$ $[F^-] = 2.0 \times 10^{-3}\ M$
(C) $[Sr^{2+}] = 2.0 \times 10^{-3}\ M$ $[F^-] = 1.0 \times 10^{-3}\ M$
(D) $[Sr^{2+}] = 2.0 \times 10^{-3}\ M$ $[F^-] = 2.0 \times 10^{-3}\ M$

15. If some of the solution evaporates overnight, which of the following will occur?

(A) The mass of the solid and the concentration of the ions will stay the same.
(B) The mass of the solid and the concentration of the ions will increase.
(C) The mass of the solid will decrease, and the concentration of the ions will stay the same.
(D) The mass of the solid will increase, and the concentration of the ions will stay the same.

16. How could the concentration of Sr^{2+} ions in solution be decreased?

(A) Adding some NaF(*s*) to the beaker
(B) Adding some $Sr(NO_3)_2(s)$ to the beaker
(C) By heating the solution in the beaker
(D) By adding a small amount of water to the beaker, but not dissolving all the solid

17. A student added 1 liter of a 1.0 *M* KCl solution to 1 liter of a 1.0 *M* $Pb(NO_3)_2$ solution. A lead chloride precipitate formed, and nearly all of the lead ions disappeared from the solution. Which of the following lists the ions remaining in the solution in order of decreasing concentration?

(A) $[NO_3^-] > [K^+] > [Pb^{2+}]$
(B) $[NO_3^-] > [Pb^{2+}] > [K^+]$
(C) $[K^+] > [Pb^{2+}] > [NO_3^-]$
(D) $[K^+] > [NO_3^-] > [Pb^{2+}]$

18. $2HI(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + I_2(g) + \text{energy}$

A gaseous reaction occurs and comes to equilibrium, as shown above. Which of the following changes to the system will serve to increase the number of moles of I_2 present at equilibrium?

(A) Increasing the volume at constant temperature
(B) Decreasing the volume at constant temperature
(C) Increasing the temperature at constant volume
(D) Decreasing the temperature at constant volume

19.

$$2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)$$

The reaction above came to equilibrium at a temperature of 100°C. At equilibrium the partial pressure due to NOBr was 4 atmospheres, the partial pressure due to NO was 4 atmospheres, and the partial pressure due to Br_2 was 2 atmospheres. What is the equilibrium constant, K_p, for this reaction at 100°C?

(A) $\frac{1}{4}$

(B) $\frac{1}{2}$

(C) 1

(D) 2

20.

$$Br_2(g) + I_2(g) \leftrightarrow 2IBr(g)$$

At 150°C, the equilibrium constant, K_c, for the reaction shown above has a value of 300. This reaction was allowed to reach equilibrium in a sealed container and the partial pressure due to IBr(*g*) was found to be 3 atm. Which of the following could be the partial pressures due to $Br_2(g)$ and $I_2(g)$ in the container?

	$Br_2(g)$	$I_2(g)$
(A)	0.1 atm	0.3 atm
(B)	0.3 atm	1 atm
(C)	1 atm	1 atm
(D)	1 atm	3 atm

21. A laboratory technician wishes to create a buffered solution with a pH of 5. Which of the following acids would be the best choice for the buffer?

(A) $H_2C_2O_4$ $K_a = 5.9 \times 10^{-2}$
(B) H_3AsO_4 $K_a = 5.6 \times 10^{-3}$
(C) $H_2C_2H_3O_2$ $K_a = 1.8 \times 10^{-5}$
(D) HOCl $K_a = 3.0 \times 10^{-8}$

22. Which of the following species is amphoteric?

(A) H^+
(B) CO_3^{2-}
(C) HCO_3^-
(D) H_2CO_3

23. How many liters of distilled water must be added to 1 liter of an aqueous solution of HCl with a pH of 1 to create a solution with a pH of 2?

(A) 0.1 L
(B) 0.9 L
(C) 2 L
(D) 9 L

24. A 1-molar solution of a very weak monoprotic acid has a pH of 5. What is the value of K_a for the acid?

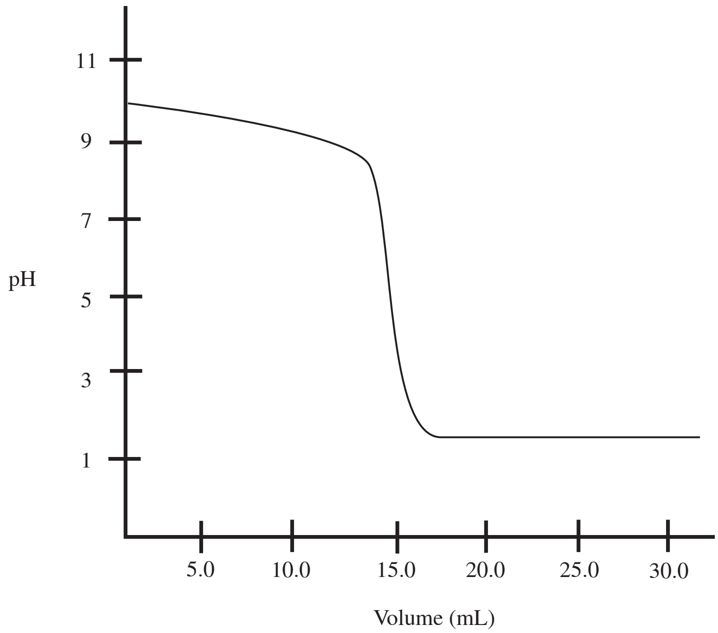
(A) $K_a = 1 \times 10^{-10}$
(B) $K_a = 1 \times 10^{-7}$
(C) $K_a = 1 \times 10^{-5}$
(D) $K_a = 1 \times 10^{-2}$

25. The value of K_a for HSO_4^- is 1×10^{-2}. What is the value of K_b for SO_4^{2-} ?

(A) $K_b = 1 \times 10^{-12}$
(B) $K_b = 1 \times 10^{-8}$
(C) $K_b = 1 \times 10^{-2}$
(D) $K_b = 1 \times 10^{2}$

Use the following information to answer questions 26-28.

The following curve is obtained during the titration of 30.0 mL of 1.0 *M* NH_3, a weak base, with a strong acid:



26. Why is the solution acidic at equilibrium?

    (A) The strong acid dissociates fully, leaving excess [H^+] in solution.
    (B) The conjugate acid of NH_3 is the only ion present at equilibrium.
    (C) The water which is being created during the titration acts as an acid.
    (D) The acid is diprotic, donating two protons for every unit dissociated.

27. What is the concentration of the acid?

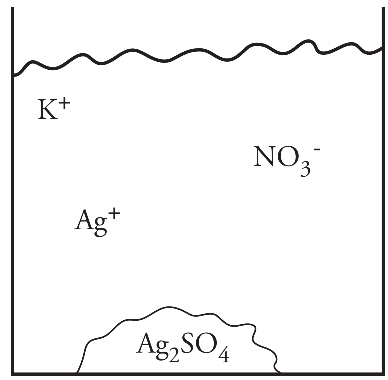
    (A) 0.5 *M*
    (B) 1.0 *M*
    (C) 1.5 *M*
    (D) 2.0 *M*

28. What ions are present in significant amounts during the first buffer region?

    (A) NH_3 and NH_4^+
    (B) NH_3 and H^+
    (C) NH_4^+ and OH^-
    (D) H_3O^+ and NH_3

Use the information below to answer questions 29-31.

Silver sulfate, Ag_2SO_4, has a solubility product constant of 1.0×10^{-5}. The below diagram shows the products of a precipitation reaction in which some silver sulfate was formed.



29. What is the identity of the excess reactant?

(A) $AgNO_3$
(B) Ag_2SO_4
(C) $NaNO_3$
(D) Na_2SO_4

30. If the beaker above were left uncovered for several hours:

(A) Some of the Ag_2SO_4 would dissolve.
(B) Some of the spectator ions would evaporate into the atmosphere.
(C) The solution would become electrically imbalanced.
(D) Additional Ag_2SO_4 would precipitate.

31. Which ion concentrations below would have led the precipitate to form?

(A) $[Ag^+] = 0.01\ M$ $[SO_4^{2-}] = 0.01\ M$
(B) $[Ag^+] = 0.10\ M$ $[SO_4^{2-}] = 0.01\ M$
(C) $[Ag^+] = 0.01\ M$ $[SO_4^{2-}] = 0.10\ M$
(D) This is impossible to determine without knowing the total volume of the solution.

32. In a voltaic cell with a Cu (s) | Cu^{2+} cathode and a Pb^{2+} | Pb (s) anode, increasing the concentration of Pb^{2+} causes the voltage to decrease. What is the reason for this?

(A) The value for Q will increase, causing the cell to come closer to equilibrium.
(B) The solution at the anode becomes more positively charged, leading to a reduced electron flow.
(C) The reaction will shift to the right, causing a decrease in favorability.
(D) Cell potential will always decrease anytime the concentration of any aqueous species present increases.

33. Which of the following could be added to an aqueous solution of weak acid HF to increase the percent dissociation?

(A) NaF(*s*)
(B) $H_2O(l)$
(C) NaOH(*s*)
(D) $NH_3(aq)$

34. A bottle of water is left outside early in the morning. The bottle warms gradually over the course of the day. What will happen to the pH of the water as the bottle warms?

(A) Nothing; pure water always has a pH of 7.00.
(B) Nothing; the volume would have to change in order for any ion concentration to change.
(C) It will increase because the concentration of $[H^+]$ is increasing.
(D) It will decrease because the auto-ionization of water is an endothermic process.

Free-Response Questions

1. The value of the solubility product, K_{sp}, for calcium hydroxide, $Ca(OH)_2$, is 5.5×10^{-6}, at 25°C.

   (a) Write the K_{sp} expression for calcium hydroxide.
   (b) What is the mass of $Ca(OH)_2$ in 500 mL of a saturated solution at 25°C?
   (c) What is the pH of the solution in (b)?
   (d) If 1.0 mole of OH^- is added to the solution in (b), what will be the resulting Ca^{2+} concentration? Assume that the volume of the solution does not change.

2. For sodium chloride, the solution process with water is endothermic.

   (a) Describe the change in entropy when sodium chloride dissociates into aqueous particles.
   (b) Two saturated aqueous NaCl solutions, one at 20°C and one at 50°C, are compared. Which one will have higher concentration? Justify your answer.
   (c) Which way will the solubility reaction shift if the temperature is increased?
   (d) If a saturated solution of NaCl is left out overnight and some of the solution evaporates, how will that affect the amount of solid NaCl present?

3. A student tests the conductivity of three different acid samples, each with a concentration of 0.10 *M* and a volume of 20.0 mL. The conductivity was recorded in microsiemens per centimeter in the table below:

Sample	Conductivity (μS/cm)
1	26,820
2	8655
3	35,120

   (a) The three acids are known to be HCl, H_2SO_4, and H_3PO_4. Identify which sample is which acid. Justify your answer.
   (b) The HCl solution is then titrated with a 0.150 *M* solution of the weak base methylamine, CH_3NH_2.
   ($K_b = 4.38 \times 10^{-4}$)
   (i) Write out the net ionic equation for this reaction.
   (ii) Determine the pH of the solution after 20.0 mL of methylamine has been added.

4. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92.4$ kJ

When the reaction above took place at a temperature of 570 K, the following equilibrium concentrations were measured:

$[NH_3] = 0.20$ mol/L
$[N_2] = 0.50$ mol/L
$[H_2] = 0.20$ mol/L

(a) Write the expression for K_c and calculate its value.
(b) Calculate ΔG for this reaction.
(c) Describe how the concentration of H_2 at equilibrium will be affected by each of the following changes to the system at equilibrium:
 (i) The temperature is increased.
 (ii) The volume of the reaction chamber is increased.
 (iii) N_2 gas is added to the reaction chamber.
 (iv) Helium gas is added to the reaction chamber.

5. In an acidic medium, iron (III) ions will react with thiocyanate (SCN^-) ions to create the following complex ion:

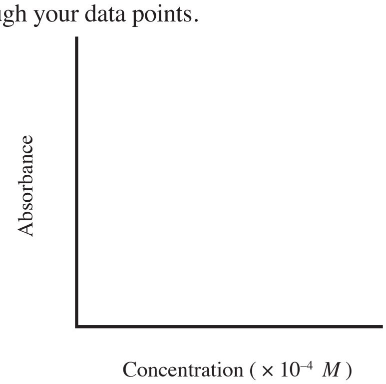
$Fe^{3+}(aq) + SCN^-(aq) \leftrightarrow FeSCN^{2+}(aq)$

Initially, the solution is a light yellow color due to the presence of the Fe^{3+} ions. As the $FeSCN^{2+}$ forms, the solution will gradually darken to a golden yellow. The reaction is not a fast one, and generally after mixing the ions the maximum concentration of $FeSCN^{2+}$ will occur between 2–4 minutes after mixing the solution.

A student creates four solution with varying concentration of $FeSCN^{2+}$ and gathers the following data at 298 K using a spectrophotometer calibrated to 460 nm:

$[FeSCN^{2+}]$	Absorbance
1.1×10^{-4} *M*	0.076
1.6×10^{-4} *M*	0.112
2.2×10^{-4} *M*	0.167
2.5×10^{-4} *M*	0.199

(a) (i) On the axes below, create a Beer's Law calibration plot for $[FeSCN^{2+}]$. Draw a best-fit line through your data points.



 (ii) The slope of the best-fit line for the above set of data points is 879 and the *y*-intercept is –0.024. Write out the equation for this line.

To determine the equilibrium constant for the reaction, a solution is made up in which 5.00 mL of 0.0025 *M* $Fe(NO_3)_3$ and 5.00 mL of 0.0025 *M* KSCN are mixed. After 3 minutes, the absorbance of the solution is found to be 0.134.

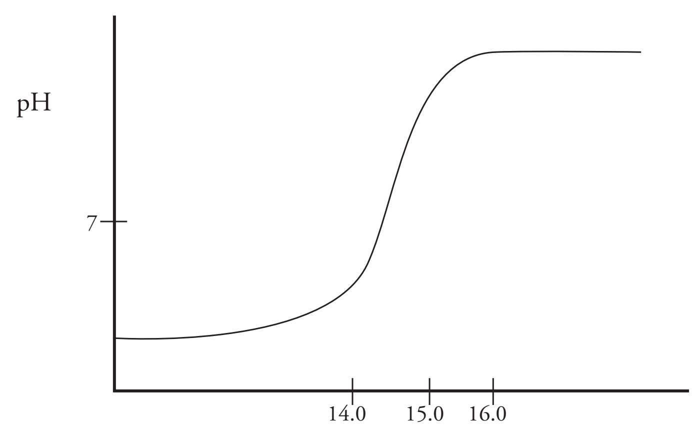
(b) (i) Using your Beer's Law best-fit line from (a), calculate $[FeSCN^{2+}]$ once equilibrium has been established.
(ii) Calculate $[Fe^{3+}]$ and $[SCN^-]$ at equilibrium.
(iii) Calculate K_{eq} for the reaction.

After equilibrium is established, the student heats the solution and observes that it becomes noticeably lighter.

(c) (i) Did heating the mixture increase the equilibrium constant, decrease it, or have no effect on it? Why?
(ii) Is the equilibrium reaction exothermic or endothermic? Justify your answer.

6. $$HA + OH \rightleftharpoons A^- + H_2O(l)$$

A student titrates a weak acid, HA, with some 1.0 *M* NaOH, yielding the following titration curve:



(a) Which chemical species present in solution dictates the pH of the solution in each of the volume ranges listed below?
(i) 1.0 mL–14.0 mL
(ii) 15.0 mL
(iii) 16.0 mL–30.0 mL

(b) At which volumes is:
(i) $[HA] > [A^-]$?
(ii) $[HA] = [A^-]$?
(iii) $[HA] < [A^-]$?

(c) At which point in the titration (if any) would the concentration of the following species be equal to zero? Justify your answers.
(i) HA
(ii) A^-

(d) If the titration were performed again, but this time with 2.0 *M* NaOH, name two things that would change about the titration curve, and explain the reasoning behind your identified changes.

7. A student performs an experiment to determine the concentration of a solution of hypochlorous acid, HOCl ($K_a = 3.5 \times 10^{-8}$). The student starts with 25.00 mL of the acid in a flask and titrates it against a standardized solution of sodium hydroxide with a concentration of 1.47 *M*. The equivalence point is reached after the addition of 34.23 mL of NaOH.

   (a) Write the net ionic equation for the reaction that occurs in the flask.
   (b) What is the concentration of the HOCl?
   (c) What would the pH of the solution in the flask be after the addition of 28.55 mL of NaOH?
   (d) The actual concentration of the HOCl is found to be 2.25 *M*. Quantitatively discuss whether or not each of the following errors could have caused the error in the student's results.
      (i) The student added additional NaOH past the equivalence point.
      (ii) The student rinsed the buret with distilled water but not with the NaOH solution before filling it with NaOH.
      (iii) The student measured the volume of acid incorrectly; instead of adding 25.00 mL of HOCl, only 24.00 mL was present in the flask prior to titration.