CHAPTER 17 – Additional Aspects of Aqueous Equilibria

<u>Section 17.1 – The Common–Ion Effect</u>

(a) Solution #1 contains 16.4 g sodium acetate dissolved in a total volume of 1.00 L.

(i) How many moles of sodium acetate are present in this solution?

(ii) Sodium acetate is a soluble ionic compound, so it is classified as a (weak strong) electrolyte.

(iii) The concentration of $Na^+(aq)$ in Solution #1 is equal to _______________ *M*, and

the concentration of $C_2H_3O_2^-$ in Solution #1 is equal to _______________ *M*.

(b) Solution #2 is 0.20 *M* acetic acid. The K_a for acetic acid is equal to 1.8×10^{-5}.

(i) Acetic acid is classified as a (weak strong) acid, so it is a (weak strong) electrolyte.

(ii) Write the K_a expression for acetic acid.

(iii) Following Sample Exercise 16.12 as an example, calculate $[H^+]$, $[CH_3CO_2^-]$, and the pH of 0.20 *M* acetic acid.

(c) Solution #3 is prepared by mixing 500 mL of Solution #1 with 500 mL of solution #2 to make 1.00 L of solution.

(i) Assume that the equilibrium concentration of the hydronium ion is equal to *x*. Fill in the following RICE table.

R	acetic acid	+	water	$\rightleftharpoons$	hydronium	+	acetate
I	0.10 *M*		–		0 *M*		0.10 *M*
C			–				
E			–				

(ii) Because K_a is small, we can assume that *x* is small compared to the original concentrations of acetic acid and acetate. Therefore we can assume that

$(0.10 + x) =$ ____________ and $(0.10 - x) =$ __________

(iii) What is the concentration of the hydronium ion in Solution #3? Show your work below.

(iv) The pH of solution #3 is equal to ____________, and this value also

represents the (K_a K_b pK_a pK_b) of acetic acid.

(d) Write the value for $[H_3O^+]$ for Solutions #2 and #3 in the blanks below.

In Solution #2, $[H_3O^+]$ = ______________ and in Solution #3, $[H_3O^+]$ = ______________

(e) The difference in $[H_3O^+]$ in Solutions #2 and #3 can be explained in terms of the **common-ion effect**. Whenever a weak electrolyte (e.g., acetic acid) and a strong electrolyte (e.g., sodium acetate) containing a common ion (e.g., acetate) are together in solution, the weak electrolyte ionizes ___________ than it would if it were alone in solution.

(f) The common ion effect can be explained in terms of LeChâtelier's principle.

A weak acid like acetic acid only ionizes slightly in solution:

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

If additional acetate ions (from sodium acetate) are added to a solution of acetic acid, this change will cause the equilibrium above to shift toward the ______________________.

(g) A student is given a solution of 0.20 *M* hydrofluoric acid. The student then adds a few drops of 12 *M* HCl to the hydrofluoric acid solution.

(i) What will happen to the value of $[H_3O^+]$ after the HCl is added? Justify your answer.

(ii) What will happen to the value of $[F^-]$ after the HCl is added? Justify your answer.

The common-ion effect can be observed in both weak acids and in weak bases. The following example is for a weak base.

(h) A certain solution is 0.20 *M* ammonia. The K_b for ammonia is equal to 1.8×10^{-5}.

(i) Write the K_b expression for ammonia.

(ii) Calculate $[OH^-]$, $[NH_4^+]$, and the pH of 0.20 *M* ammonia.

(i) What would happen if ammonium chloride is added to a solution of 0.20 *M* ammonia?

(i) The concentration of the NH_4^+ ion would ________________

(ii) The concentration of the OH^- ion would ________________

(j) The common ion effect can be explained in terms of LeChâtelier's principle.

A weak base like ammonia only ionizes slightly in solution:

$$NH_3(aq) \quad + \quad H_2O(l) \quad \rightleftarrows \quad OH^-(aq) \quad + \quad NH_4^+(aq)$$

If additional ammonium ions (from ammonium chloride) are added to a solution of ammonia, this change will cause the equilibrium above to shift toward the _________________.

(k) A solution is prepared by mixing 500 mL of 1.0 *M* NH_3 with 500 mL of 1.0 *M* NH_4Cl to make 1.00 L of solution.

(i) Assume that the equilibrium concentration of the hydroxide ion is equal to *x*. Fill in the following RICE table.

R	ammonia	+	water	$\rightleftharpoons$	hydroxide	+	ammonium
I	0.50 *M*		–		0 *M*		0.50 *M*
C			–				
E			–				

(ii) Because K_b is small, we can assume that *x* is small compared to the original concentrations of ammonia and ammonium. Therefore we can assume that

$(0.50 + x) =$ ____________ and $(0.50 – x) =$ ___________

(iii) What is the concentration of the hydroxide ion in this solution? Show your work below.

(iv) The pOH of this solution is equal to ___________, and this value also

represents the (K_a K_b pK_a pK_b) of ammonia.

The pH of this solution is equal to ____________.

Section 17.2 – Buffered Solutions

(a) Solutions that contain a weak conjugate acid-base pair are called _______________________

solutions because they are able to _______________ drastic changes in pH when small

amounts of strong acid or strong base are added to them.

(b) Circle any of the following eight solutions that would represent a good buffer solution.

1 *M* HNO_3 and 1 *M* $NaNO_3$

1 *M* HF and 1 *M* NaF

1 *M* $HClO_3$ and 1 *M* $NaClO_3$

1 *M* CH_3CO_2H and 1 *M* $NaHCO_3$

1 *M* NH_3 and 1 *M* NH_4Cl

1 *M* HCN and 1 *M* NaCN

1 *M* CH_3NH_2 and 1 *M* CH_3CO_2H

1 *M* NaH_2PO_4 and 1 *M* Na_2HPO_4

(c) Justify the choices you made in part (b).

(d) The expression for the acid-dissociation constant for a weak acid HA is written as

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Rearrange the K_a expression to solve for [H$^+$].

(e) The equation you wrote in part (d) is Equation [17.5] on page 707.

From this relationship, we see that [H$^+$] and pH are determined by two factors:

the value of ____________ for the weak acid HA and

the ratio of the __________________________ of HA and A$^-$.

(f) Figure 17.2 on page 708 shows a diagram representing a buffer solution.

In this **initial** buffered solution (center of the diagram),

the concentration of HF is ___________________________ the concentration of F$^-$.

(g) Suppose that we add OH$^-$ ions to this buffered solution. Write the net-ionic equation for the reaction that would occur.

(h) Adding OH$^-$ ions to the HF/F$^-$ buffer will cause the concentration of HF to ________________

and the concentration of F$^-$ to _______________.

(i) Now suppose that we add H$^+$ ions to this buffered solution. Write the net-ionic equation for the reaction that would occur.

(j) Adding H$^+$ ions to the HF/F$^-$ buffer will cause the concentration of HF to ______________

and the concentration of F$^-$ to _______________.

(k) As long as the ratio $\frac{[HA]}{[A^-]}$ doesn't change very much, the change in pH

will be ______________ when either H^+ ions or OH^- ions are added to a buffer solution.

(l) If you need to calculate the pH of a buffer solution, the Henderson-Hasselbalch equation is very useful. Write Equation [17.9] on page 709.

(m) A certain buffer contains 1.0 M CH_3CO_2H and 1.0 M CH_3CO_2Na. Use the Henderson-Hasselbalch equation to calculate the pH of this solution. (K_a for acetic acid is 1.8×10^{-5})

(n) If a buffer solution contains equimolar quantities of a weak acid and its conjugate base, then

the pH of the buffer solution is equal to the ___________ of the weak acid.

(o) Make a prediction about the pH values of each of the following buffer solutions.
Then use the Henderson-Hasselbalch equation to calculate the pH of each solution.

Composition of the Buffer Solution	Will the pH be higher or lower than 4.74?	Calculated pH value
1.0 M CH_3CO_2H and 0.80 M CH_3CO_2Na		
0.80 M CH_3CO_2H and 1.0 M CH_3CO_2Na		

(p) A certain buffer solution contains a mixture of NH_3 and NH_4Cl. The pH of this solution is 9.43. (K_a for $NH_4^+ = 5.6 \times 10^{-10}$)

(i) Calculate the pK_a of the ammonium ion. ________________

(ii) The pH of this solution is 9.43, which is (less greater) than the pK_a for NH_4^+.

Therefore, this solution must contain…

_______ more base (NH_3) than acid (NH_4^+)

_______ more acid (NH_4^+) than base (NH_3).

(iii) The concentration of NH_4^+ in this solution is 0.80 M. Use the Henderson-Hasselbalch equation to calculate the concentration of ammonia in this buffer solution.

(q) **Buffer capacity** is the amount of acid or base the buffer can ______________________

before the ________ of the buffer begins to change to an appreciable degree.

The **pH range** of any buffer is the pH range over which the buffer acts effectively.

Buffers most effectively resist a change in pH in either direction when the concentrations of

weak acid and conjugate base are ___________________________.

Buffers usually have a usable range defined as pH = ________________.

Buffer A	Buffer B
1.0 *M* CH_3CO_2H and 1.0 *M* CH_3CO_2Na	0.10 *M* CH_3CO_2H and 0.10 *M* CH_3CO_2Na

(r) The pH of Buffer A is (less than greater than equal to) the pH of Buffer B.

(s) Buffer A has a (smaller greater) buffer capacity than Buffer B because ___________

____

(t) Which one of these acids would be more suitable to use for making a buffer solution at a pH of 9.00?

HCN ($K_a = 4.9 \times 10^{-10}$) HF ($K_a = 6.8 \times 10^{-4}$)

(u) What other substance is needed to make the buffer described in part (t)? _____________

(v) A certain buffer contains 1.0 *M* CH_3CO_2H and 1.0 *M* CH_3CO_2Na at a pH of 4.74.
The volume of the buffer is 1.00 L.

(i) Suppose that 1 mL of 10.0 *M* NaOH are added to this buffer.
How many moles of OH^- ions does this quantity represent? _________________

(ii) Assume that the volume of the solution is still 1.00 L. After the NaOH has been added to

the buffer, $[CH_3CO_2H]$ = _____________ and $[CH_3CO_2^-]$ = ______________

(iii) Use the Henderson-Hasselbalch equation to calculate the pH of the buffer after the NaOH is added to it.

(iv) As a comparison, calculate the pH if 1 mL of 10.0 *M* NaOH is added to 1.00 L of pure water. Assume that the final volume of solution is still 1.00 L.

(w) A certain buffer contains 1.0 *M* CH_3CO_2H and 1.0 *M* CH_3CO_2Na at a pH of 4.74.
The volume of the buffer is 1.00 L.

(i) Suppose that 1 mL of 10.0 *M* HCl are added to this buffer.
How many moles of H^+ ions does this quantity represent? ____________________

(ii) Assume that the volume of the solution is still 1.00 L. After the HCl has been added to

the buffer, $[CH_3CO_2H]$ = _____________ and $[CH_3CO_2^-]$ = _______________

(iii) Use the Henderson-Hasselbalch equation to calculate the pH of the buffer after the HCl is added to it.

(iv) As a comparison, calculate the pH if 1 mL of 10.0 *M* HCl is added to 1.00 L of pure water. Assume that the final volume of solution is still 1.00 L.

Section 17.3 – Acid-Base Titrations

In an acid-base titration, a solution containing a known concentration of base is slowly added to a sample of acid in which the concentration is unknown. Alternatively, a known concentration of acid is slowly added a sample of base in which the concentration is unknown.

(a) The point at which stoichiometrically equivalent quantities of acid and base have been brought

together is called the _____________________________________. This point can be signaled

by a color change in an acid-base __________________________ such as phenolphthalein.

(b) Alternatively, a pH meter can be used to monitor the progress of the reaction, producing a

pH ____, a graph of the pH as a function of the

volume of titrant added. The shape of this graph makes it possible to determine the

equivalence point.

See Figure 17.6 on page 714 for a set-up of an acid-base titration. The solution in the buret is normally of known concentration, and it is called the **titrant**. The titrant is slowly added from the buret to a solution in a beaker or a flask, which is called the **analyte**. If an indicator is used, a few drops of indicator are added to the flask at the beginning of the titration. When the color change occurs, it is called the **end point**. The titration is stopped at the end point. If a pH meter is used to monitor the pH, no indicator is needed. You can add titrant past the equivalence point because the titration curve can be analyzed after the experiment to locate the equivalence point.

Strong Acid – Strong Base Titration

(c) Consider the following titration experiment.
The buret contains 0.10 *M* NaOH. The flask contains 50.0 mL of 0.10 *M* HCl.

The initial pH of 0.10 *M* HCl is ___________

(d) After 49.0 mL of 0.10 *M* NaOH have been added to the flask,
we have the following stoichiometry.

Initial moles of HCl = _______________ Moles of NaOH added = ___________________

Moles of excess HCl = ______________ Total volume of solution = __________________

$[H^+]$ = _______________ pH of solution = ________________

(e) After 50.0 mL of 0.10 *M* NaOH have been added, the equivalence point is reached because

moles of NaOH added are ______________ the initial moles of HCl in the flask. At the

equivalence point of a strong acid – strong base titration, the pH is equal to ___________

because both the cation and the anion of the salt that is formed are _____________________
and do not affect the pH of the solution.

(f) After 51.0 mL of 0.10 *M* NaOH have been added to the flask,
we have the following stoichiometry.

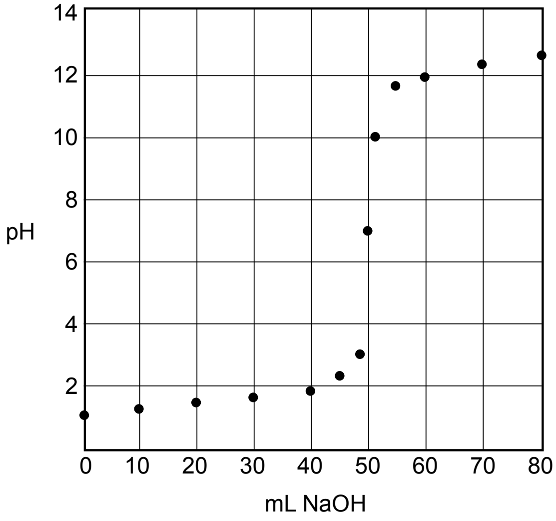
Initial moles of HCl = _______________ Moles of NaOH added = ___________________

Moles of excess NaOH = ____________ Total volume of solution = __________________

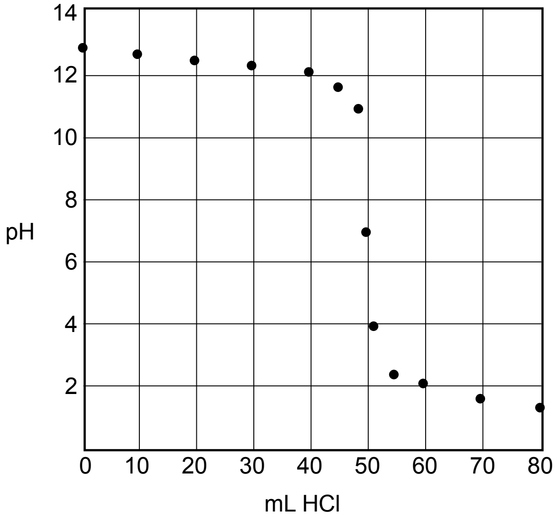
$[OH^-]$ = ______________ pH of solution = ________________

(g) Sketch the shape of the pH titration curve for each of the following titrations.

50.0 mL of 0.10 *M* HCl
is titrated with 0.10 *M* NaOH



50.0 mL of 0.10 *M* NaOH
is titrated with 0.10 *M* HCl



(h) The features of a strong acid – strong base titration are as follows.

The initial pH of the strong acid is very ___________.

As the titrant (strong base) is added, the pH increases _______________ at first, and then

increases _______________ in the vicinity of the equivalence point.

At the equivalence point the pH of the solution is __________ because the salt that is formed in

the neutralization reaction is _________________.

After the equivalence point, the pH of the solution is very _____________.

Weak Acid – Strong Base Titration

(i) Consider the following titration experiment.
The buret contains 0.10 *M* NaOH. The flask contains 50.0 mL of 0.10 *M* $HC_2H_3O_2$.
The K_a for $HC_2H_3O_2$ is 1.8 x 10^{-5}.

Acetic acid is a _______ acid, so the pH of 0.10 *M* $HC_2H_3O_2$ should be ________ than 1.00.

(j) Calculate the pH of 0.10 *M* $HC_2H_3O_2$.

(k) After 25.0 mL of 0.10 *M* NaOH have been added to the flask,
we have the following stoichiometry.

Initial moles of $HC_2H_3O_2$ = _______________ Moles of NaOH added = _________________

Moles of $HC_2H_3O_2$ present in the solution = _______________

Moles of $C_2H_3O_2^-$ present in the solution = _______________

(l) If you use the Henderson-Hasselbalch equation at this point to calculate the pH of the solution,
you will discover something familiar from Section 17.2.

At the halfway-point in a weak acid – strong base titration, a buffer solution exists in which

the moles of the weak acid (HA) are _________________ the moles of the conjugate base (A^-).

At the halfway-point in any weak acid – strong base titration, the pH = ______ of the weak acid.

(m) After 50.0 mL of 0.10 *M* NaOH have been added to the flask, we have the following stoichiometry.

Initial moles of $HC_2H_3O_2$ = ______________ Moles of NaOH added = ________________

Moles of $HC_2H_3O_2$ present in the solution = ______________

Moles of $C_2H_3O_2^-$ present in the solution = ______________

Total volume of solution = ___________

$[C_2H_3O_2^-]$ = _____________

(n) In the strong acid – strong base titration discussed earlier, the pH at the equivalence point was equal to 7 because the salt formed at the equivalence point (NaCl) was neutral.

However, in a weak acid – strong base titration, the salt present in solution at the equivalence point is not neutral. In this example, we have a solution of sodium acetate.

Sodium acetate ($NaC_2H_3O_2$) behaves as a ___________ in solution.

Write a net-ionic equation that shows why the acetate ion behaves this way in water.

$$\rightleftharpoons$$

(o) The K_a for acetic acid is equal to 1.8×10^{-5}. Calculate the value of K_b for acetate.

(p) The concentration of acetate in solution at the equivalence point of this weak acid – strong base titration is equal to 0.050 *M*. Use the K_b expression to calculate the pH at this point.

(q) After 51.0 mL of 0.10 *M* NaOH have been added to the flask, we have the following stoichiometry.

Initial moles of $HC_2H_3O_2$ = ______________ Moles of NaOH added = ________________

Moles of excess NaOH = _______________ Total volume of solution = _______________

$[OH^-]$ = ________________ pH of solution = ________________

(r) It is important to compare and contrast the pH titration curves for a strong acid strong base titration with a weak acid – strong base titration.

(i) A solution of a weak acid has an initial pH that is ________________ than the pH of a strong acid of the same concentration.

(ii) At the halfway-point in a weak acid titration, the pH = __________ of the weak acid.

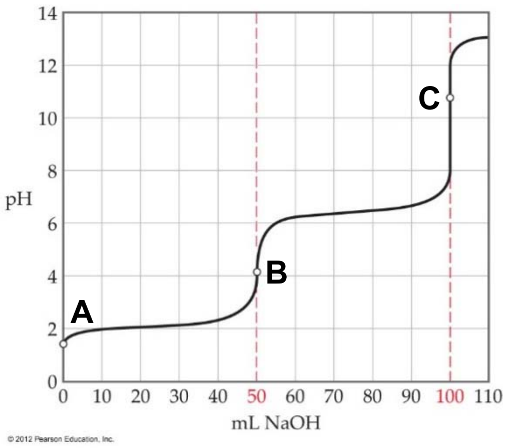
(iii) Near the equivalence point, the "rapid rise" portion of the curve is ____________________ for the weak acid titration than it is for the strong acid titration.

(iv) In a weak acid titration, the pH at the equivalence point is ________________ than 7.0 because the salt that is produced at the equivalence point behaves as a ______________.

After the equivalence point, the titration curves for a weak acid and a strong acid should appear similar to each other.

Polyprotic Acid – Strong Base Titration

(s) The diagram at right shows the titration curve for an experiment in which 50.0 mL of 0.10 *M* H_3PO_3 is titrated with 0.10 *M* NaOH.



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(i) What is the dominant species present in the solution at point A of the titration? __________

(ii) What is the dominant species present in the solution at point B of the titration? __________

(iii) What is the dominant species present in the solution at point C of the titration? __________

An acid-base indicator is often used in titrations instead of a pH meter.

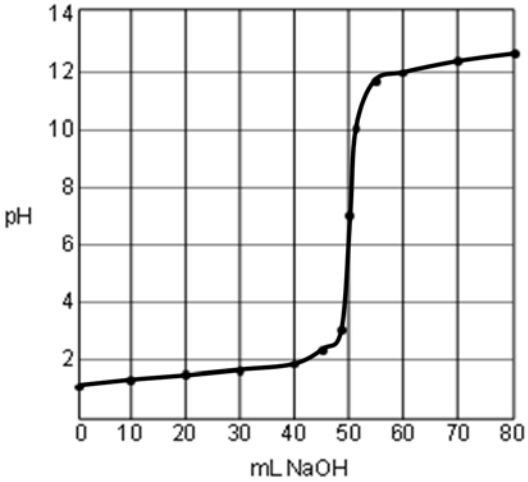
(t) The pH changes very _______________ as you get close to the equivalence point.

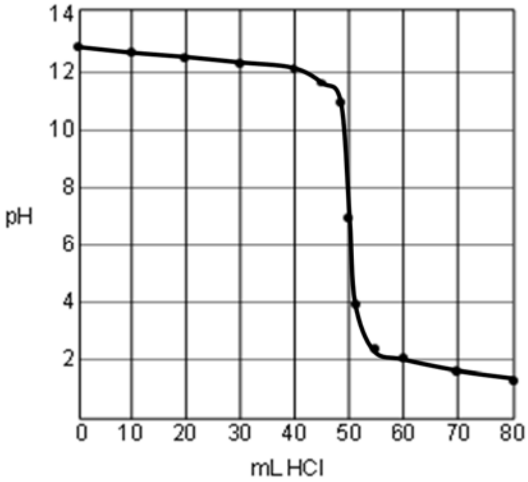
In this region, one ___________ of titrant can change the pH by several units.

Ideally, the indicator should begin and end its color change anywhere on the

_______________________ portion of the titration curve.

(u) Methyl red changes color from red to yellow over a pH range of 4.8 – 6.0. Phenolphthalein changes color from colorless to pink over a pH range of 8.2 – 10.0. The following diagram represents the pH titration curves for a strong acid – strong base titration (left) and a strong base – strong acid titration (right). Explain why both methyl red and phenolphthalein are good choices for estimating the equivalence point in these titrations.





(v) Use the information in Figure 17.14 on page 721 to explain why methyl red is an unsatisfactory choice for an indicator in the titration of acetic acid with NaOH.

(w) Use the information in Figure 17.15 on page 722 to explain why phenolphthalein is an unsatisfactory choice for an indicator in the titration of ammonia with HCl.

Section 17.4 – Solubility Equilibria

(a) In a saturated solution, there is an equilibrium between the dissolved particles and the undissolved solid. For each of the following solids, write a balanced chemical equation for the equilibrium that would be established when the solid dissolves in water to form a saturated solution. Use Equation [17.15] on page 723 as an example.

(i) barium fluoride ____

(ii) lead(II) carbonate ____

(iii) silver chromate ____

(b) The equilibrium constant for the process in which a solid dissolves in water is known as the solubility-product constant, or K_{sp}. Write the K_{sp} expression for the reactions written in part (a).

(i) barium fluoride ____

(ii) lead(II) carbonate ____

(iii) silver chromate ____

(c) It is important to distinguish carefully between solubility and K_{sp}.

Solubility is defined as the quantity of a solute that dissolves (in water) to form a

____________________ solution. The units of solubility are normally g/L or mol/L.

The solubility-product constant (K_{sp}) is a unitless number that represents the equilibrium

constant for the equilibrium between an ionic solid and its ____________________ solution.

The magnitude of the K_{sp} can be used to compare the solubility of two solids that have similar chemical formulas.

(d) Based on the information given at the top of page 724 (“Give It Some Thought”), which of the following salts should have the highest molar solubility in water?

silver chloride　　silver bromide　　silver iodide

(e) Based on the information given in Table D.3 on page 1063, which of the following salts should have the highest molar solubility in water?

barium sulfate　　calcium sulfate　　lead(II) sulfate

If you are given the solubility of an ionic solid, you can calculate the value of K_{sp}.

(f) Read Sample Exercise 17.10 on page 724. Then fill in the missing information in the table below.

Name	Formula	Cation concentration in a saturated solution (*M*)	Anion concentration in a saturated solution (*M*)	Calculated value of K_{sp} (2 sig figs)
copper(II) carbonate		$[Cu^{2+}] = 1.5 \times 10^{-5}$ *M*	$[CO_3^{2-}] =$	
silver carbonate		$[Ag^+] = 2.5 \times 10^{-4}$ *M*	$[CO_3^{2-}] =$	
copper(II) hydroxide		$[Cu^{2+}] = 2.3 \times 10^{-7}$ *M*	$[OH^-] =$	

If you are given the K_{sp} for an ionic solid, you can calculate the molar solubility.

(g) Read Sample Exercise 17.11 on page 725. Then fill in the missing information in the table below.

Name	Formula	K_{sp}	Cation concentration in a saturated solution (*M*)	Anion concentration in a saturated solution (*M*)
lead(II) chromate		2.8×10^{-13}	$[Pb^{2+}] =$	$[CrO_4^{2-}] =$
silver chromate		1.2×10^{-12}	$[Ag^+] =$	$[CrO_4^{2-}] =$
lead(II) fluoride		3.6×10^{-8}	$[Pb^{2+}] =$	$[F^-] =$

Section 17.5 – Factors that Affect Solubility

(a) Three factors that can affect the solubility of ionic compounds are the presence of ____________________ ions, the _________ of the solution, and the presence of complexing agents such as NH_3, CN^-, or OH^-.

(b) In the last section, you calculated the molar solubility of lead(II) fluoride to be 2.1×10^{-3} *M*. According to the common-ion effect, the solubility of PbF_2 in a solution that already contains either Pb^{2+} ions or F^- ions should be (less greater) than the solubility of PbF_2 in pure water.

(c) Read Sample Exercise 17.12 on page 727. Then calculate the molar solubility of PbF_2 (K_{sp} = 3.6 x 10^{-8}) in a solution that is

(i) 0.010 *M* in $Pb(NO_3)_2$

(ii) 0.010 *M* NaF

(d) The pH of a solution affects the solubility of any substance whose anion is basic.

For example, $Mg(OH)_2$ becomes more soluble in water as the pH of the solution is

(decreased increased).

(e) The solubility of PbF_2 will (decrease increase) as the solution becomes more acidic.

This observation can be explained in terms of LeChâtelier's principle. The additional H^+ ions

react with the ________ ions to form _________. This causes the concentration of the F^- ions

in the solution to (decrease increase). This causes a shift in the equilibrium of PbF_2

toward the (left right).

(f) The solubility of PbI_2 is <u>not</u> affected by lowering the pH of the solution. Explain.

(g) Which of the following substances is more soluble in an acidic solution than it is in pure water?

$CuCl_2$ $Ni(OH)_2$ $MgCO_3$ AgBr $Ca(NO_3)_2$

(h) Write a balanced equation for the dissolution of barium fluoride in water.
The K_{sp} for barium fluoride is equal to 1.7×10^{-6}.

(i) Write a balanced equation for the dissociation of hydrofluoric acid in water.
The K_a for hydrofluoric acid is equal to 6.8×10^{-4}.

(j) Write a balanced chemical equation that shows what happens when solid barium fluoride is added to a strong acid solution containing H^+ ions.

(k) Use the values of K_{sp} for BaF_2 and K_a for HF to calculate the value of K_c for the reaction written in part (j). If the value of K is greater than 1, it indicates that the reaction proceeds spontaneously to the right.

(l) An assembly of a metal ion and a substance known as a Lewis base (which is an electron pair donor) is called a ___________________ ion. Examples of these types of ions are listed in Table 17.1 on page 732.

Section 17.6 – Precipitation and Separation of Ions

(a) Comparing the reaction quotient Q with the solubility-product constant K_{sp} helps to predict whether or not a precipitate will form for a given set of conditions. There are three possibilities. Fill in the blanks with the appropriate conclusions regarding the formation of a precipitate.

If $Q > K_{sp}$, ____

If $Q = K_{sp}$, ____

If $Q < K_{sp}$, ____

(b) The K_{sp} for $BaSO_4$ is equal to 1.1×10^{-10}. Predict whether or not a precipitate will form in each situation.

(i) 25 mL of 2.0×10^{-4} *M* $Ba(NO_3)_2$ is mixed with 25 mL of 1.0×10^{-4} *M* Na_2SO_4.

(ii) 75 mL of 5.0×10^{-6} *M* $BaCl_2$ is mixed with 75 mL of 3.0×10^{-5} *M* K_2SO_4.

(c) A beaker contains 1 liter of a solution in which both calcium ions and barium ions are present. In this solution, $[Ca^{2+}] = [Ba^{2+}] = 2.0 \times 10^{-4}$ *M*.

A solution of 0.10 *M* NaF is slowly added from a buret into the beaker.

At 25°C, the value of K_{sp} for CaF_2 is 3.9×10^{-11}, and the value of K_{sp} for BaF_2 is 1.7×10^{-6}.

(i) Predict which salt, CaF_2 or BaF_2 will precipitate first. Justify your answer.

(ii) Calculate the minimum concentration of fluoride that is required in order for the salt you selected in part (c)(i) to form a precipitate.

(c) (iii) Calculate the minimum volume of 0.10 *M* NaF that is required in order for the salt you selected in part (c)(i) to form a precipitate.

Section 17.7 – Qualitative Analysis for Metallic Elements

This section presents information that is not included in the AP Chemistry curriculum and will not be covered on the AP Exam.