

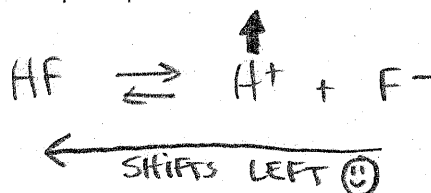
Practice Test Chapter 17

Buffers, Titrations, K_{sp}

Part I: Circle the letter of the best answer(s). No calculators are allowed on the multiple choice portion of the practice test!

Target 1: I can explain the Common Ion Effect and make associated calculations.

- The Common Ion Effect is defined as a(n) . . .
 - decrease in pH when an acid is added to a buffer solution which contains an acid with a greater K_a than the K_b of the conjugate base.
 - increase in pH when an acid is added to a buffer solution which contains an acid with a greater K_a than the K_b of the conjugate base.
 - decrease in the solubility of a strong acid when mixed with a weak base.
 - increase in the percent ionization of either a weak acid or a weak base upon the addition of a strong electrolyte containing an ion in common with the acid or base.
 - shift in equilibrium caused by the addition of a compound having an ion in common with the dissolved substance
- What change will be caused by addition of a small amount of HCl to a solution containing fluoride ions and hydrogen fluoride?
 - The concentration of hydronium ions will increase significantly.
 - The concentration of fluoride ions will increase as will the concentration of hydronium ions.
 - The concentration of hydrogen fluoride will decrease and the concentration of fluoride ions will increase.
 - The concentration of fluoride ion will decrease and the concentration of hydrogen fluoride will increase.
 - The fluoride ions will precipitate out of solution as its acid salt.



Target 2: I can describe how a buffer solution of a particular pH is made and how it operates to control pH. I can also calculate the change in pH of a simple buffer solution caused by adding a small amount of strong acid or base.

3. Which **one** of the following pairs cannot be mixed together to form a buffer solution?

- a. C_5H_5N, C_5H_5NHCl
- b. $HC_2H_3O_2, NaC_2H_3O_2$
- c. $KOH, HI \Rightarrow$ STRONG ACID / STRONG BASE
- d. NH_3, NH_4Cl
- e. $NaClO, HClO$

4. Which of the following solutions would make the best buffer?

- a. 0.10 M HCl and 0.10 M $NaCl$
- b. 0.10 M NH_4Cl and 0.10 M $NaCl$
- c. 0.10 M NH_4Cl and 0.10 M $HC_2H_3O_2$
- d. 0.10 M HF and 0.10 M $NaF \Rightarrow$ WEAK ACID & ITS CONJ. BASE
- e. None of the above would make very good buffers.

5. The net ionic equation describing the reaction of added $NaOH$ to a buffer solution containing HCO_3^- and CO_3^{2-} is _____.

- a. $OH^-(aq) + CO_3^{2-}(aq) \rightarrow HCO_4^{3-}(aq)$
- b. $OH^-(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_3^{2-}(aq)$
- c. $2 Na^+(aq) + CO_3^{2-}(aq) \rightarrow Na_2CO_3(aq)$
- d. $Na^+(aq) + HCO_3^-(aq) \rightarrow NaHCO_3(aq)$
- e. $Na^+(aq) + H_2O(l) \rightarrow H^+(aq) + NaOH(aq)$

6. The primary buffer system that controls the pH of blood is the _____ buffer system.

- a. carbon dioxide, carbonate
- b. carbonate, bicarbonate
- c. carbonic acid, oxygen
- d. carbonate, carbonic acid
- e. carbonic acid, bicarbonate

7. The K_a of benzoic acid is 6.30×10^{-5} . The pH of a buffer prepared by combining 50.0 mL of 1.00 M potassium benzoate and 50.0 mL of 1.00 M benzoic acid is _____.

- a. 1.709
- b. 0.851
- c. 3.406
- d. 4.201
- e. 2.383

Since you have same # of mols of acid & base component of the buffer, $pH = pK_a$!

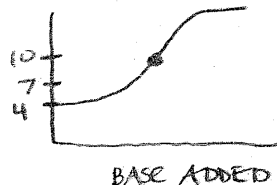
$$pH = pK_a + \log \frac{[BZ^-]}{[HBZ]}$$

No need for a calculator!
The pK_a (& pH) is a little less than the absolute value of the exponent 10^{-5} .

Use H^2 equation.
☺

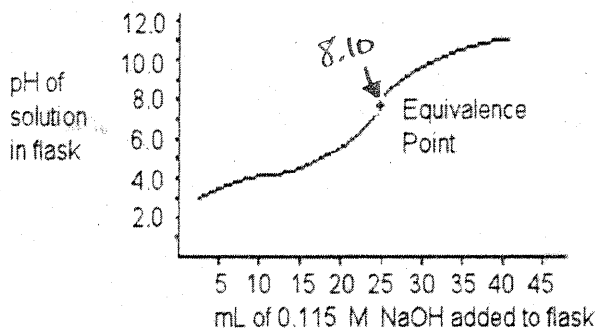
Target 3: I can describe the shape and key features of titration curves involving both strong and weak acids and bases.

8. An initial pH of 4.00, an equivalence point at pH 9.35, and a moderately short, nearly vertical middle section correspond to a titration curve for _____.
- strong acid titrated by a strong base
 - strong base titrated by a strong acid
 - weak acid titrated by a strong base
 - weak base titrated by a strong acid
 - weak base titrated by a weak acid



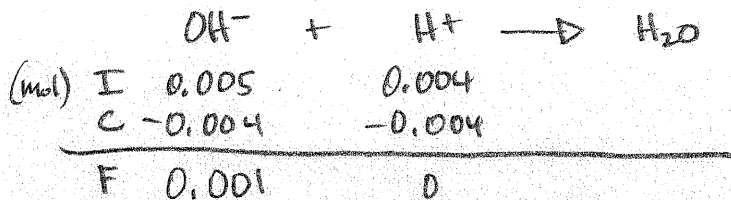
9. A 25.0 mL sample of a solution of an unknown compound is titrated with a 0.115 M NaOH solution. The titration curve below was obtained. The unknown compound is _____. (The pH = 8.10 at equivalence point.)
- a strong acid
 - a strong base
 - a weak acid
 - a weak base
 - neither an acid nor a base

↳ due to the hydrolysis of the basic salt!



Target 4: I can calculate the pH at any point, including the equivalence point, in acid base titrations.

10. The pH of a solution prepared by mixing 50.0 mL of 0.100 M NaOH and 40.0 mL of 0.100 M HNO₃ is approximately _____.
- 1.95
 - 4.87
 - 7.00
 - 9.13
 - 12.05



$$[\text{OH}^-] = \frac{0.001 \text{ mol}}{0.09 \text{ L}} = \frac{1 \times 10^{-3}}{9 \times 10^{-2}} \approx \frac{1}{10} \times 10^{-1} \text{ M OH}^- = 1 \times 10^{-2} \text{ M OH}^-$$

↑ 50 mL + 40 mL

$\approx \text{pOH} = 2$
 $\approx \text{pH} = 12$

11. Assume you added enough 0.30 M HCl to neutralize a certain volume of 0.10 M KOH. What is the approximate pH at equivalence point?
- a. 1.00
 b. 3.15
 c. 7.00
 d. 10.15
 e. There is not enough information to solve this problem.
- For a strong acid / strong base titration, a neutral salt is produced! No calculation needed!!*

12. The pH of a solution prepared by mixing 8.00 liters of 0.100 M NaOH and 2.00 liters of 0.200 M HC₂H₃O₂ is approximately _____.
- (The K_a of HC₂H₃O₂ is 1.8 × 10⁻⁵)

a. 1.39
 b. 5.28
 c. 7.00
 d. 9.03
 e. 12.61

OH⁻ + HC₂H₃O₂ → C₂H₃O₂⁻ + H₂O

	<i>(mol) I</i>	<i>0.8</i>	<i>0.4</i>	<i>0</i>	
	<i>C</i>	<i>-0.4</i>	<i>-0.4</i>	<i>+0.4</i>	
<i>F</i>	<i>0.4</i>	<i>0</i>	<i>0.4</i>	<i>0</i>	

STRONG BASE 0 WEAK BASE 0.4

[OH⁻] = 0.4 mol / 10 L
[OH⁻] = 0.04 M
[OH⁻] = 4 × 10⁻² M
pOH ≈ 2
pH ≈ 12

Target 5: I can calculate K_{sp} from solubility data and solubility from the value of K_{sp}.

13. The K_{sp} expression for silver phosphate is
- a. K_{sp} = [Ag⁺][PO₄³⁻]
 b. K_{sp} = [Ag⁺]²[PO₄³⁻]
 c. K_{sp} = [Ag⁺]³[PO₄³⁻]
 d. K_{sp} = [Ag⁺][PO₄³⁻]³
 e. K_{sp} = [Ag⁺][PO₄³⁻] / [Ag₃PO₄]
- Ag₃PO₄(s) ⇌ 3Ag⁺ + PO₄³⁻*

14. The concentration of lead ions in a saturated solution of lead (II) sulfide is _____ M. The solubility product constant of PbS is 3.0 × 10⁻²⁸.

a. 3.0 × 10⁻⁹
 b. 2.8 × 10⁻²
 c. 9.0 × 10⁻⁵⁶
 d. 1.7 × 10⁻¹⁴
 e. 3.0 × 10⁻²⁸

PbS(s) ⇌ Pb²⁺ + S²⁻
3.0 × 10⁻²⁸ = (x)(x) = x²
x = √(3 × 10⁻²⁸) = [Pb²⁺] = [S²⁻]

15. The solubility of manganese (II) hydroxide is 2.2 × 10⁻⁵ M. What is the K_{sp} of Mn(OH)₂?

a. 1.1 × 10⁻⁷
 b. 4.3 × 10⁻¹⁴
 c. 2.1 × 10⁻²¹
 d. 4.8 × 10⁻¹⁰
 e. 2.2 × 10⁻⁵

Mn(OH)₂(s) ⇌ Mn²⁺ + 2OH⁻; K_{sp} = [Mn²⁺][OH⁻]²
K_{sp} = (2.2 × 10⁻⁵)(4.4 × 10⁻⁵)²
K_{sp} = (2.2 × 10⁻⁵)(4.4 × 10⁻⁵)(4.4 × 10⁻⁵)
K_{sp} ≈ 40 × 10⁻¹⁵
K_{sp} = 4 × 10⁻¹⁴

16. Of the compounds listed in the table below, _____ is the least soluble.

compound	K _{sp}
CdS	8.0 × 10 ⁻²⁷
CuS	6.3 × 10 ⁻³⁶
PbS	8.0 × 10 ⁻²⁸
MnCO ₃	1.8 × 10 ⁻¹¹

The smaller the K_{sp}, the less soluble the salt!

- a. CdS
- b. CuS
- c. PbS
- d. MnCO₃
- e. The three sulfides are equally soluble, and all less than MnCO₃.

Target 6: I can solve K_{sp} problems associated with the common ion effect and changes in pH.

17. What is the solubility of PbCl₂ in a 0.15 M solution of HCl? The K_{sp} of PbCl₂ is 1.6 × 10⁻⁵.

- a. 2.0 × 10⁻⁹ M
- b. 1.1 × 10⁻¹⁴ M
- c. 1.8 × 10⁰ M
- d. 7.1 × 10⁻⁴ M
- e. 1.6 × 10⁵ M



$$1.6 \times 10^{-5} = (x)(0.15)^2 =$$

$$x = \text{sol. PbCl}_2 = \frac{1.6 \times 10^{-5}}{(1.5 \times 10^{-1})(1.5 \times 10^{-1})} \approx \frac{1.6 \times 10^{-5}}{2 \times 10^{-2}} = 0.8 \times 10^{-3} = 8 \times 10^{-4}$$

18. In which aqueous system is PbI₂ least soluble?

- a. H₂O
- b. 0.50 M HI
- c. 0.2 M NaCl
- d. 1.0 M NaCl
- e. 0.8 M KI



↳ greatest concn. of

19. The solubility of which one of the following will be least affected by the pH of the solution?

- a. Na₃PO₄
- b. NaF
- c. MnS
- d. CaCO₃
- e. KNO₃

↳ shifts furthest to left! (Le Chat's prin.)

It is the only neutral salt. ←

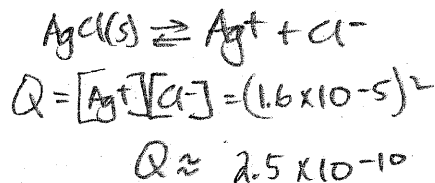
20. Which of the following statement(s) is/are TRUE? (If you think there is more than one true statement, darken them all in!)

- a. The solubility of AgCl can be increased by the addition of a solution of NaCl. *decreased*
- b. A solution of 0.1 M NaC₂H₃O₂ and 0.001 M HC₂H₃O₂ should act as a very effective buffer toward added base. *↳ a really low conc. of acid*
- c. The Henderson-Hasselbalch equation shows for a buffer solution that changing the total volume of the solution does not change the pH. ☹️
- d. The solubility of BaCO₃ will increase by the addition of a strong acid.

↳ since it is basic salt, it will dissolve well in an acid!!

Target 7: I can predict whether a precipitate will form when two solutions are mixed, given appropriate K_{sp} values.

21. Equal volumes of $1.6 \times 10^{-5} \text{ M KCl}$ and $1.6 \times 10^{-5} \text{ M AgNO}_3$ are mixed. The K_{sp} for silver chloride is 1.6×10^{-10} . As these two solutions are mixed . . .
- a precipitate of AgCl forms.
 - there is no precipitate formed.
 - NaCl will precipitate.
 - AgNO₃ will precipitate.
 - the $[\text{Na}^+]$ will become 0.020 M .



$Q > K \dots$ you get ppt!!

2011

Part 2: CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures.

1. Each of three beakers contains 25.0 mL of a 0.100 M solution of HCl, NH₃, or NH₄Cl, as shown below. Each solution is at 25°C .

Beaker 1: 25.0 mL of 0.100 M HCl

Beaker 2: 25.0 mL of 0.100 M NH_3

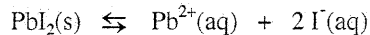
Beaker 3: 25.0 mL of $0.100 \text{ M NH}_4\text{Cl}$

- Determine the pH of the solution in beaker 1. Justify your answer.
- In beaker 2, the reaction $\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$ occurs. The value of K_b for $\text{NH}_3(aq)$ is 1.8×10^{-5} at 25°C .
 - Write the K_b expression for the reaction of $\text{NH}_3(aq)$ with $\text{H}_2\text{O}(l)$.
 - Calculate the $[\text{OH}^-]$ in the solution in beaker 2.
- In beaker 3, the reaction $\text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq)$ occurs.
 - Calculate the value of K_a for $\text{NH}_4^+(aq)$ at 25°C .
 - The contents of beaker 2 are poured into beaker 3 and the resulting solution is stirred. Assume that volumes are additive. Calculate the pH of the resulting solution.
- The contents of beaker 1 are poured into the solution made in part (c)(ii). The resulting solution is stirred. Assume that volumes are additive.
 - Is the resulting solution an effective buffer? Justify your answer.
 - Calculate the final $[\text{NH}_4^+]$ in the resulting solution at 25°C .

2006

2. Answer the following questions that relate to the solubility of salts lead and barium.

a) A saturated solution is prepared by adding excess $\text{PbI}_2(\text{s})$ to distilled water to form 1.0 liter of solution at 25°C . The concentration of the $\text{Pb}^{2+}(\text{aq})$ in the saturated solution is found to be $1.3 \times 10^{-3} \text{ M}$. The chemical equation for the dissolution of $\text{PbI}_2(\text{s})$ is shown below:

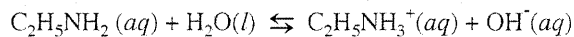


- Write the equilibrium constant expression.
 - Calculate the molar concentration of the $\text{I}^{-}(\text{aq})$.
 - Calculate the value of K_{sp} .
- b) A saturated solution is prepared by adding $\text{PbI}_2(\text{s})$ to distilled water to form 2.00 liters of solution at 25°C . What are the molar concentrations of the $\text{Pb}^{2+}(\text{aq})$ and $\text{I}^{-}(\text{aq})$? Justify your answer.
- c) Solid NaI is added to a saturated solution of PbI_2 at 25°C . Assume that the volume of the solution does not change, does the molar concentration of the $\text{Pb}^{2+}(\text{aq})$ increase, decrease, or remain the same? Justify your answer.
- d) The value the K_{sp} of BaCrO_4 is 1.2×10^{-10} . When a 500-mL sample of $8.2 \times 10^{-6} \text{ M Ba}(\text{NO}_3)_2$ is added to 500. mL of $8.2 \times 10^{-6} \text{ M Na}_2\text{CrO}_4$, no precipitate is observed.
- Assume that the volumes are additive, calculate the molar concentrations of $\text{Ba}^{2+}(\text{aq})$ and $\text{CrO}_4^{2-}(\text{aq})$ in the 1.00 liter of solution.
 - Using the molar concentrations of the $\text{Ba}^{2+}(\text{aq})$ and $\text{CrO}_4^{2-}(\text{aq})$ above, show why a precipitate does not form. You must include a calculation as part of your answer.

2009 B

3. A pure 14.85 g sample of the weak base ethylamine, $\text{C}_2\text{H}_5\text{NH}_2$, is dissolved in enough distilled water to make 500. mL of solution.

(a) Calculate the molar concentration of the $\text{C}_2\text{H}_5\text{NH}_2$ in the solution. The aqueous ethylamine reacts with water according to the equation below.



- (b) Write the equilibrium-constant expression for the reaction between $\text{C}_2\text{H}_5\text{NH}_2(\text{aq})$ and water.
- (c) Of $\text{C}_2\text{H}_5\text{NH}_2(\text{aq})$ and $\text{C}_2\text{H}_5\text{NH}_3^{+}(\text{aq})$, which is present in the solution at the higher concentration at equilibrium? Justify your answer.

(d) A different solution is made by mixing 500. mL of 0.500 M $\text{C}_2\text{H}_5\text{NH}_2$ with 500. mL of 0.200 M HCl . Assume that volumes are additive. The pH of the resulting solution is found to be 10.93.

- Calculate the concentration of $\text{OH}^{-}(\text{aq})$ in the solution.
- Write the net-ionic equation that represents the reaction that occurs when the $\text{C}_2\text{H}_5\text{NH}_2$ solution is mixed with the HCl solution.
- Calculate the molar concentration of the $\text{C}_2\text{H}_5\text{NH}_3^{+}(\text{aq})$ that is formed in the reaction.
- Calculate the value of K_b for $\text{C}_2\text{H}_5\text{NH}_2$.

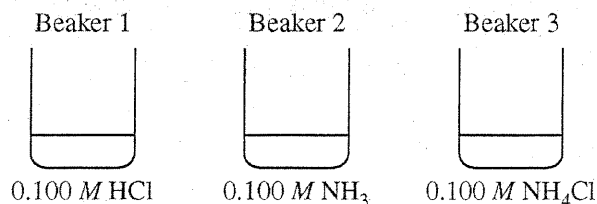
4. A sample of 0.1276 g of an unknown monoprotic acid was dissolved in 25.0 mL of water and titrated with 0.0633 M NaOH. The volume of base required to reach equivalence point was 18.4 mL.

a) Calculate the molar mass of the acid.

b) After 10.0 mL of the base has been added in the titration, the pH was determined to be 5.87. What is the K_a of the unknown acid?

AP[®] CHEMISTRY
2011 SCORING GUIDELINES

Question 1



1. Each of three beakers contains 25.0 mL of a 0.100 M solution of HCl, NH₃, or NH₄Cl, as shown above. Each solution is at 25°C.

(a) Determine the pH of the solution in beaker 1. Justify your answer.

$\text{pH} = -\log[\text{H}^+] = -\log(0.100) = 1.000$	1 point is earned for the correct pH.
--	---------------------------------------

- (b) In beaker 2, the reaction $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$ occurs. The value of K_b for $\text{NH}_3(\text{aq})$ is 1.8×10^{-5} at 25°C.

(i) Write the K_b expression for the reaction of $\text{NH}_3(\text{aq})$ with $\text{H}_2\text{O}(\text{l})$.

$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$	1 point is earned for the correct expression.
--	---

(ii) Calculate the $[\text{OH}^-]$ in the solution in beaker 2.

<p>Let $[\text{OH}^-] = x$, then $K_b = \frac{(x)(x)}{(0.100 - x)}$</p> <p>Assume that $x \ll 0.100 \text{ M}$, then</p> $1.8 \times 10^{-5} = \frac{x^2}{0.100} \Rightarrow x = [\text{OH}^-] = 1.3 \times 10^{-3} \text{ M}$	<p>1 point is earned for the correct setup.</p> <p>1 point is earned for the correct answer.</p>
---	--

- (c) In beaker 3, the reaction $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$ occurs.

(i) Calculate the value of K_a for $\text{NH}_4^+(\text{aq})$ at 25°C.

$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$	1 point is earned for the correct answer.
--	---

**AP[®] CHEMISTRY
2011 SCORING GUIDELINES**

Question 1 (continued)

- (ii) The contents of beaker 2 are poured into beaker 3 and the resulting solution is stirred. Assume that volumes are additive. Calculate the pH of the resulting solution.

<p>In the resulting solution, $[\text{NH}_3] = [\text{NH}_4^+]$;</p> $K_a = 5.6 \times 10^{-10} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$ <p>Thus $[\text{H}_3\text{O}^+] = 5.6 \times 10^{-10}$; $\text{pH} = -\log(5.6 \times 10^{-10}) = 9.25$</p>	<p>1 point is earned for noting that the solution is a buffer with $[\text{NH}_3] = [\text{NH}_4^+]$.</p> <p>1 point is earned for the correct pH.</p>
---	---

- (d) The contents of beaker 1 are poured into the solution made in part (c)(ii). The resulting solution is stirred. Assume that volumes are additive.

- (i) Is the resulting solution an effective buffer? Justify your answer.

<p>The resulting solution is not an effective buffer. Virtually all of the NH_3 in the solution formed in (c)(ii) will react with the H_3O^+ from solution 1:</p> $\text{NH}_3 + \text{H}_3\text{O}^+ \rightarrow \text{NH}_4^+ + \text{H}_2\text{O}$ <p>leaving mostly NH_4^+ in the final solution. Since only one member of the $\text{NH}_4^+/\text{NH}_3$ conjugate acid-base pair is left, the solution cannot buffer both base and acid.</p>	<p>1 point is earned for the correct response with an acceptable justification.</p>
---	---

- (ii) Calculate the final $[\text{NH}_4^+]$ in the resulting solution at 25°C.

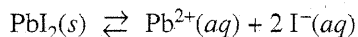
<p>moles = (volume)(molarity)</p> <p>moles H_3O^+ in sol. 1 = $(0.0250)(0.100) = 0.00250$ mol</p> <p>moles NH_3 in sol. 2 = $(0.0250)(0.100) = 0.00250$ mol</p> <p>moles NH_4^+ in sol. 3 = $(0.0250)(0.100) = 0.00250$ mol</p> <p>When the solutions are mixed, the H_3O^+ and NH_3 react to form NH_4^+, resulting in a total of 0.00500 mol NH_4^+. The final volume is the sum $(25.0 + 25.0 + 25.0) = 75.0$ mL.</p> <p>The final concentration of NH_4^+ = $(0.00500 \text{ mol}/0.0750 \text{ L}) = 0.0667 \text{ M}$.</p>	<p>1 point is earned for the correct calculation of moles of NH_4^+.</p> <p>1 point is earned for the correct calculation of the final volume <u>and</u> concentration.</p>
---	---

AP[®] CHEMISTRY
2006 SCORING GUIDELINES

Question 1

2 Answer the following questions that relate to solubility of salts of lead and barium.

- (a) A saturated solution is prepared by adding excess $\text{PbI}_2(s)$ to distilled water to form 1.0 L of solution at 25°C. The concentration of $\text{Pb}^{2+}(aq)$ in the saturated solution is found to be $1.3 \times 10^{-3} M$. The chemical equation for the dissolution of $\text{PbI}_2(s)$ in water is shown below.



- (i) Write the equilibrium-constant expression for the equation.

$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$	One point is earned for the correct expression.
---	---

- (ii) Calculate the molar concentration of $\text{I}^{-}(aq)$ in the solution.

By stoichiometry, $[\text{I}^{-}] = 2 \times [\text{Pb}^{2+}]$, thus $[\text{I}^{-}] = 2 \times (1.3 \times 10^{-3}) = 2.6 \times 10^{-3} M$	One point is earned for the correct concentration.
--	--

- (iii) Calculate the value of the equilibrium constant, K_{sp} .

$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2 = (1.3 \times 10^{-3})(2.6 \times 10^{-3})^2$ $= 8.8 \times 10^{-9}$	One point is earned for a value of K_{sp} that is consistent with the answers in parts (a)(i) and (a)(ii).
--	--

- (b) A saturated solution is prepared by adding $\text{PbI}_2(s)$ to distilled water to form 2.0 L of solution at 25°C. What are the molar concentrations of $\text{Pb}^{2+}(aq)$ and $\text{I}^{-}(aq)$ in the solution? Justify your answer.

The molar concentrations of $\text{Pb}^{2+}(aq)$ and $\text{I}^{-}(aq)$ would be the same as in the 1.0 L solution in part (a) (i.e., $1.3 \times 10^{-3} M$ and $2.6 \times 10^{-3} M$, respectively). The concentrations of solute particles in a saturated solution are a function of the constant, K_{sp} , which is independent of volume.	One point is earned for the concentrations (or stating they are the same as in the solution described in part (a)) and justification.
--	---

AP[®] CHEMISTRY
2006 SCORING GUIDELINES

Question 1 (continued)

- (c) Solid NaI is added to a saturated solution of PbI_2 at 25°C . Assuming that the volume of the solution does not change, does the molar concentration of $\text{Pb}^{2+}(\text{aq})$ in the solution increase, decrease, or remain the same? Justify your answer.

<p>$[\text{Pb}^{2+}]$ will decrease.</p> <p>The $\text{NaI}(\text{s})$ will dissolve, increasing $[\text{I}^-]$; more $\text{I}^-(\text{aq})$ then combines with $\text{Pb}^{2+}(\text{aq})$ to precipitate $\text{PbI}_2(\text{s})$ so that the ion product $[\text{Pb}^{2+}][\text{I}^-]^2$ will once again attain the value of 8.8×10^{-9} (K_{sp} at 25°C).</p>	<p>One point is earned for stating that $[\text{Pb}^{2+}]$ will decrease.</p> <p>One point is earned for justification (can involve a Le Chatelier argument).</p>
--	--

- (d) The value of K_{sp} for the salt BaCrO_4 is 1.2×10^{-10} . When a 500. mL sample of $8.2 \times 10^{-6} M$ $\text{Ba}(\text{NO}_3)_2$ is added to 500. mL of $8.2 \times 10^{-6} M$ Na_2CrO_4 , no precipitate is observed.

- (i) Assuming that volumes are additive, calculate the molar concentrations of $\text{Ba}^{2+}(\text{aq})$ and $\text{CrO}_4^{2-}(\text{aq})$ in the 1.00 L of solution.

<p>New volume = 500. mL + 500. mL = 1.000 L, therefore $[\text{Ba}^{2+}]$ in 1.000 L is one-half its initial value:</p> $[\text{Ba}^{2+}] = \frac{500. \text{ mL}}{1,000. \text{ mL}} \times (8.2 \times 10^{-6} M) = 4.1 \times 10^{-6} M$ $[\text{CrO}_4^{2-}] = \frac{500. \text{ mL}}{1,000. \text{ mL}} \times (8.2 \times 10^{-6} M) = 4.1 \times 10^{-6} M$	<p>One point is earned for the correct concentration.</p>
---	---

- (ii) Use the molar concentrations of $\text{Ba}^{2+}(\text{aq})$ ions and $\text{CrO}_4^{2-}(\text{aq})$ ions as determined above to show why a precipitate does not form. You must include a calculation as part of your answer.

<p>The product $Q = [\text{Ba}^{2+}][\text{CrO}_4^{2-}]$</p> $= (4.1 \times 10^{-6} M)(4.1 \times 10^{-6} M)$ $= 1.7 \times 10^{-11}$ <p>Because $Q = 1.7 \times 10^{-11} < 1.2 \times 10^{-10} = K_{sp}$, no precipitate forms.</p>	<p>One point is earned for calculating a value of Q that is consistent with the concentration values in part (d)(i).</p> <p>One point is earned for using Q to explain why no precipitate forms.</p>
---	--

AP[®] CHEMISTRY
2009 SCORING GUIDELINES (Form B)

#3

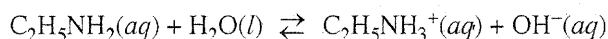
Question 1 (10 points)

A pure 14.85 g sample of the weak base ethylamine, $C_2H_5NH_2$, is dissolved in enough distilled water to make 500. mL of solution.

- (a) Calculate the molar concentration of the $C_2H_5NH_2$ in the solution.

$n_{C_2H_5NH_2} = 14.85 \text{ g } C_2H_5NH_2 \times \frac{1 \text{ mol } C_2H_5NH_2}{45.09 \text{ g } C_2H_5NH_2}$ $= 0.3293 \text{ mol } C_2H_5NH_2$ $M_{C_2H_5NH_2} = \frac{0.3293 \text{ mol } C_2H_5NH_2}{0.500 \text{ L}} = \mathbf{0.659 \text{ M}}$	<p>One point is earned for the correct number of moles.</p> <p>One point is earned for the correct concentration.</p>
---	---

The aqueous ethylamine reacts with water according to the equation below.



- (b) Write the equilibrium-constant expression for the reaction between $C_2H_5NH_2(aq)$ and water.

$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$	<p>One point is earned for the correct expression.</p>
---	--

- (c) Of $C_2H_5NH_2(aq)$ and $C_2H_5NH_3^+(aq)$, which is present in the solution at the higher concentration at equilibrium? Justify your answer.

<p>$C_2H_5NH_2$ is present in the solution at the higher concentration at equilibrium. Ethylamine is a weak base, and thus it has a small K_b value. Therefore only partial dissociation of $C_2H_5NH_2$ occurs in water, and $[C_2H_5NH_3^+]$ is thus less than $[C_2H_5NH_2]$.</p>	<p>One point is earned for the correct answer with justification.</p>
---	---

AP[®] CHEMISTRY
2009 SCORING GUIDELINES (Form B)

Question 1 (continued)

- (d) A different solution is made by mixing 500. mL of 0.500 M C₂H₅NH₂ with 500. mL of 0.200 M HCl. Assume that volumes are additive. The pH of the resulting solution is found to be 10.93.

- (i) Calculate the concentration of OH⁻(aq) in the solution.

<p>pH = -log[H⁺] [H⁺] = 10^{-10.93} = 1.17 × 10⁻¹¹ $[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1.00 \times 10^{-14}}{1.17 \times 10^{-11}} = 8.5 \times 10^{-4} \text{ M}$</p> <p>OR</p> <p>pOH = 14 - pH = 14 - 10.93 = 3.07 pOH = -log[OH⁻] [OH⁻] = 10^{-3.07} = 8.5 × 10⁻⁴ M</p>	<p>One point is earned for the correct concentration.</p>
---	---

- (ii) Write the net-ionic equation that represents the reaction that occurs when the C₂H₅NH₂ solution is mixed with the HCl solution.

$\text{C}_2\text{H}_5\text{NH}_2 + \text{H}_3\text{O}^+ \rightarrow \text{C}_2\text{H}_5\text{NH}_3^+ + \text{H}_2\text{O}$	<p>One point is earned for the correct equation.</p>
---	--

- (iii) Calculate the molar concentration of the C₂H₅NH₃⁺(aq) that is formed in the reaction.

<p>moles of C₂H₅NH₂ = 0.500 L × $\frac{0.500 \text{ mol}}{1.00 \text{ L}}$ = 0.250 mol</p> <p>moles of H₃O⁺ = 0.500 L × $\frac{0.200 \text{ mol}}{1.00 \text{ L}}$ = 0.100 mol</p> <table border="1" style="margin: 10px auto; border-collapse: collapse; text-align: center;"> <thead> <tr> <th></th> <th>[C₂H₅NH₂]</th> <th>[H₃O⁺]</th> <th>[C₂H₅NH₃⁺]</th> </tr> </thead> <tbody> <tr> <td>initial value</td> <td>0.250</td> <td>0.100</td> <td>~ 0</td> </tr> <tr> <td>change</td> <td>-0.100</td> <td>-0.100</td> <td>+0.100</td> </tr> <tr> <td>final value</td> <td>0.150</td> <td>~ 0</td> <td>0.100</td> </tr> </tbody> </table> <p>$[\text{C}_2\text{H}_5\text{NH}_3^+] = \frac{0.100 \text{ mol C}_2\text{H}_5\text{NH}_3^+}{1.00 \text{ L}} = 0.100 \text{ M}$</p>		[C ₂ H ₅ NH ₂]	[H ₃ O ⁺]	[C ₂ H ₅ NH ₃ ⁺]	initial value	0.250	0.100	~ 0	change	-0.100	-0.100	+0.100	final value	0.150	~ 0	0.100	<p>One point is earned for the correct number of moles of C₂H₅NH₂ and H₃O⁺.</p> <p>One point is earned for the correct concentration.</p>
	[C ₂ H ₅ NH ₂]	[H ₃ O ⁺]	[C ₂ H ₅ NH ₃ ⁺]														
initial value	0.250	0.100	~ 0														
change	-0.100	-0.100	+0.100														
final value	0.150	~ 0	0.100														

AP[®] CHEMISTRY
2009 SCORING GUIDELINES (Form B)

Question 1 (continued)

(iv) Calculate the value of K_b for $C_2H_5NH_2$.

$$[C_2H_5NH_2] = \frac{0.150 \text{ mol } C_2H_5NH_2}{1.00 \text{ L}} = 0.150 \text{ M}$$

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} = \frac{(0.100)(8.5 \times 10^{-4})}{0.150} = 5.67 \times 10^{-4}$$

One point is earned for the correct calculation of the molarity of $C_2H_5NH_2$ after neutralization.

One point is earned for the correct value.