

Chemistry 212
EXAM 2
February 17, 2004

100 (of 100)

KEY

Name _____

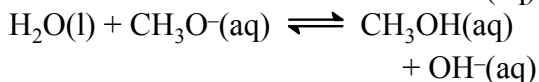
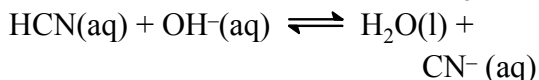
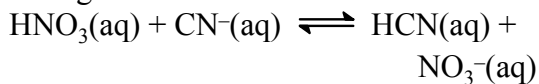
Part 1: Multiple Choice. (1 point each, circle only one answer,

1. For the equilibrium that exists in an aqueous solution of nitrous acid (HNO_2 , a weak acid), the equilibrium constant expression is:
 - a. $K = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$
 - b. $K = \frac{[\text{H}^+][\text{N}][\text{O}]^2}{[\text{HNO}_2]}$
 - c. $K = [\text{H}^+][\text{NO}_2^-]$
 - d. $K = \frac{[\text{H}^+]^2[\text{NO}_2^-]}{[\text{HNO}_2]}$
 - e. none of these
2. Which of the following is a conjugate acid/base pair?
 - a. HCl/OCl^-
 - b. $\text{H}_2\text{SO}_4/\text{SO}_4^{2-}$
 - c. $\text{NH}_4^+/\text{NH}_3$
 - d. $\text{H}_3\text{O}^+/\text{OH}^-$
 - e. none of these
3. At 0°C , the ion-product constant of water, K_w , is 1.2×10^{-15} . The pH of pure water at 0°C is:
 - a. 7.00
 - b. 6.88
 - c. 7.56
 - d. **7.46**
 - e. none of these
4. The equilibrium constant for the reaction $\text{A}^- + \text{H}^+ \rightleftharpoons \text{HA}$ is called:
 - a. K_a
 - b. K_b
 - c. $\frac{1}{K_a}$
 - d. $\frac{K_w}{K_b}$
 - e. $K_w K_a$
5. What is the equilibrium constant for the following reaction?

$$\text{N}_3^- + \text{H}_3\text{O}^+ \rightleftharpoons \text{HN}_3 + \text{H}_2\text{O}$$
 The K_a value for $\text{HN}_3 = 1.9 \cdot 10^{-5}$.
 - a. $5.3 \cdot 10^{-10}$
 - b. $1.9 \cdot 10^{-9}$
 - c. $1.9 \cdot 10^{-5}$
 - d. **$5.3 \cdot 10^4$**
 - e. $1.9 \cdot 10^9$
6. The hydrogen sulfate or bisulfate ion HSO_4^- can act as either an acid or a base in water solution. In which of the following equations does HSO_4^- act as an acid?
 - a. $\text{HSO}_4^- + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{SO}_4 + \text{OH}^-$
 - b. $\text{HSO}_4^- + \text{H}_3\text{O}^+ \leftrightarrow \text{SO}_3 + 2\text{H}_2\text{O}$
 - c. $\text{HSO}_4^- + \text{OH}^- \leftrightarrow \text{H}_2\text{SO}_4 + \text{O}^{2-}$
 - d. **$\text{HSO}_4^- + \text{H}_2\text{O} \leftrightarrow \text{SO}_4^{2-} + \text{H}_3\text{O}^+$**
 - e. none of these

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7–8. The following three equations represent equilibria that lie far to the right.



7. Identify the strongest acid.

- a. HCN
- b. HNO₃**
- c. H₂O
- d. OH⁻
- e. CH₃OH

8. Identify the strongest base.

- a. CH₃O⁻**
- b. CH₃OH
- c. CN⁻
- d. H₂O
- e. NO₃⁻

9. Calculate the [H⁺] in a solution that has a pH of 11.70.

- a. 2.3 M
- b. 11.7 M
- c. 5.0 × 10⁻³ M
- d. 2.0 × 10⁻¹² M**
- e. none of these

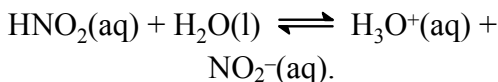
10. The pH of a solution at 25°C in which [OH⁻] = 3.4 • 10⁻⁵ M is:

- a. 4.5
- b. 10.5
- c. 9.5**
- d. 6.3
- e. none of these

11. In deciding which of two acids is the stronger, one must know:

- a. the concentration of each acid solution
- b. the pH of each acid solution
- c. the equilibrium constant of each acid**
- d. all of these
- e. both a and c must be known

12. Consider the reaction



Which species is the conjugate base?

- a. HNO₂(aq)
- b. H₂O(l)
- c. H₃O⁺(aq)
- d. NO₂⁻(aq)**
- e. Two of these

13. Calculate the pH of 0.250 M HNO₃(aq).

- a. 0.60**
- b. 2.50
- c. 12.0
- d. 1.20
- e. 13.4

14. Calculate the pH of a 0.10 M solution of Ca(OH)₂.

- a. 13.30**
- b. 13.00
- c. 0.20
- d. 0.10
- e. none of these

15. A 0.400-g sample of NaOH(s) is added to enough water to make 250.0 mL of solution. The pH of this solution is:

- a. 12.000
- b. 1.398
- c. 2.000
- d. 12.602**
- e. none of these

16. Calculate the pH of a 5.0 M solution of aniline ($C_6H_5NH_2$; $K_b = 3.8 \cdot 10^{-10}$):
- 4.36
 - 9.64**
 - 0.070
 - 9.30
 - none of these
17. Calculate the percentage of pyridine (C_5H_5N) that forms pyridinium ion, $C_5H_5NH^+$, in a 0.10 M aqueous solution of pyridine ($K_b = 1.7 \cdot 10^{-9}$):
- 0.0060%
 - 1.6%
 - 0.77%
 - 0.060%
 - 0.013%**
18. The dihydrogenphosphate ion, $H_2PO_4^-$, has both a conjugate acid and a conjugate base. These are, respectively:
- H_3PO_4 , PO_4^{3-}
 - H_3PO_4 , HPO_4^{2-}**
 - $H_2PO_4^-$, HPO_4^{2-}
 - HPO_4^{2-} , PO_4^{3-}
 - HPO_4^{2-} , H_3PO_4
19. We dilute 1.00 mL of 1.00 M HCl solution to 100.0 mL. What is $[OH^-]$ in this solution at 25 °C?
- 1.00×10^{12} M
 - 1×10^2 M
 - 0.010 M
 - 7.00×10^{-4} M
 - 1.00×10^{-12} M**
20. Which of the following acid-base reactions will lie predominantly toward the products? ($K_a(CH_3CO_2H) = 1.8 \cdot 10^{-5}$, $K_b(NH_3) = 1.8 \cdot 10^{-5}$)
- Reaction 1: $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
- Reaction 2: $CH_3CO_2H(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CH_3CO_2^-(aq)$
- Reaction 3: $CH_3CO_2H(aq) + NH_3(aq) \leftrightarrow NH_4^+(aq) + CH_3CO_2^-(aq)$
- 1 only
 - 2 only
 - 3 only**
 - 1 and 2 only
 - 1 and 3 only
21. At the neutralization point of the titration of an acid with base, what condition is met?
- Volume of base added from buret equals volume acid in reaction flask.
 - Molarity of base from the buret equals molarity of acid in reaction flask.
 - Moles of base added from the buret equals moles of acid in the reaction flask.**
 - % ionization of base added from the buret equals % ionization of the acid in flask.
 - All of the above conditions are met.
22. If you mix equal molar quantities of NaOH and CH_3CO_2H , what are the principal species present in the resulting solution?
- Na^+ , $CH_3CO_2^-$, OH^- , and H_2O
 - Na^+ , $CH_3CO_2^-$, CH_3CO_2H , OH^- , and H_2O**
 - Na^+ , CH_3CO_2H , OH^- , and H_2O
 - Na^+ , $CH_3CO_2^-$, H_3O^+ , and H_2O
 - Na^+ , CH_3CO_2H , H_3O^+ , and H_2O

23. What is the concentration of CrO_4^{2-} in a saturated solution of PbCrO_4 if $K_{\text{sp}} = 1.8 \times 10^{-14}$?
- $1.3 \times 10^{-7} \text{ M}$**
 - $7.5 \times 10^{-6} \text{ M}$
 - $1.8 \times 10^{-4} \text{ M}$
 - $1.3 \times 10^{-4} \text{ M}$
 - $5.1 \times 10^{-3} \text{ M}$
24. For thallium bromide, TlBr , $K_{\text{sp}} = 3.4 \times 10^{-6}$. How many grams of KBr ($\text{MM} = 119.0 \text{ g/mol}$) must be added to 100. mL of $5.5 \times 10^{-4} \text{ M TlNO}_3$ to just initiate precipitation?
- 0.74 g
 - 0.074 g**
 - 0.065 g
 - 0.0065 g
 - 0.0033 g
25. Silver chromate, Ag_2CrO_4 , has a K_{sp} of $9.0 \cdot 10^{-12}$. Calculate the solubility in mol/L of silver chromate.
- $1.3 \cdot 10^{-4} \text{ M}$**
 - $7.8 \cdot 10^{-5} \text{ M}$
 - $9.5 \cdot 10^{-7} \text{ M}$
 - $1.9 \cdot 10^{-12} \text{ M}$
 - $9.8 \cdot 10^{-5} \text{ M}$
26. The solubility in mol/L of $\text{M}(\text{OH})_2$ in 0.010 M KOH is $1.0 \cdot 10^{-5} \text{ mol/L}$. What is the K_{sp} for $\text{M}(\text{OH})_2$?
- $1.0 \cdot 10^{-9}$**
 - $1.0 \cdot 10^{-10}$
 - $1.0 \cdot 10^{-12}$
 - $4.0 \cdot 10^{-15}$
 - $1.0 \cdot 10^{-2}$
27. Chromate ion is added to a saturated solution of Ag_2CrO_4 to reach 0.10 M CrO_4^{2-} . Calculate the final concentration of silver ion at equilibrium (K_{sp} for Ag_2CrO_4 is $9.0 \cdot 10^{-12}$)
- $1.7 \cdot 10^{-6} \text{ M}$
 - $9.5 \cdot 10^{-6} \text{ M}$**
 - $6.6 \cdot 10^{-6} \text{ M}$
 - $5.5 \cdot 10^{-11} \text{ M}$
 - $1.1 \cdot 10^{-13} \text{ M}$
28. The solubility of $\text{Mg}(\text{OH})_2$ ($K_{\text{sp}} = 8.9 \cdot 10^{-12}$) in 1.0 L of a solution buffered (with large capacity) at pH 10.0 is:
- $8.9 \cdot 10^9$ moles
 - $8.9 \cdot 10^{-4}$ moles**
 - $8.9 \cdot 10^{-1}$ moles
 - $8.9 \cdot 10^{-7}$ moles
 - none of these
29. Sodium chloride is added slowly to a solution that is 0.010 M in Cu^+ , Ag^+ , and Au^+ . The K_{sp} values for the chloride salts are $1.9 \cdot 10^{-7}$, $1.6 \cdot 10^{-10}$, and $2.0 \cdot 10^{-13}$, respectively. Which compound will precipitate first?
- $\text{CuCl}(\text{s})$
 - $\text{AgCl}(\text{s})$
 - $\text{AuCl}(\text{s})$**
 - All will precipitate at the same time.
 - Cannot be determined.
30. For which process is ΔS negative?
- evaporation of 1 mol of $\text{CCl}_4(\text{l})$
 - mixing 5 mL ethanol with 25 mL water
 - compressing 1 mol Ne at constant temperature from 1.5 atm to 0.5 atm**
 - raising the temperature of 100 g Cu from 275 K to 295 K
 - grinding a large crystal of KCl to powder

31. A 100-mL sample of water is placed in a coffee cup calorimeter. When 1.0 g of an ionic solid is added, the temperature decreases from 21.5°C to 20.8°C as the solid dissolves. For the dissolving of the solid
- $\Delta H < 0$
 - $\Delta S_{\text{univ}} > 0$
 - $\Delta S_{\text{sys}} < 0$
 - $\Delta S_{\text{surr}} > 0$
 - none of these
32. ΔS_{surr} is _____ for exothermic reactions and _____ for endothermic reactions.
- favorable, unfavorable**
 - unfavorable, favorable
 - favorable, favorable
 - unfavorable, unfavorable
 - cannot tell
33. Calculate the standard entropy change for the following reaction,
 $\text{Cu(s)} + 1/2 \text{O}_2(\text{g}) \rightarrow \text{CuO(s)}$
 given that $S^\circ[\text{Cu(s)}] = 33.15 \text{ J/K}\cdot\text{mol}$, $S^\circ[\text{O}_2(\text{g})] = 205.14 \text{ J/K}\cdot\text{mol}$, and $S^\circ[\text{CuO(s)}] = 42.63 \text{ J/K}\cdot\text{mol}$
- 195.66 J/K
 - 93.09 J/K
 - 45.28 J/K
 - 93.09 J/K**
 - 195.66 J/K
34. At what temperature would a given reaction become spontaneous if $\Delta H = +119 \text{ kJ}$ and $\Delta S = +263 \text{ J/K}$?
- 452 K**
 - 2210 K
 - 382 K
 - 2.21 K
 - 363 K
35. All of the following have $\Delta G^\circ_f = 0$ **EXCEPT**
- $\text{O}_2(\text{g})$
 - $\text{Br}_2(\text{g})$**
 - $\text{H}_2(\text{g})$
 - Ca(s)
 - Hg(l)
36. Consider the freezing of liquid water at -10°C . For this process what are the signs for ΔH , ΔS , and ΔG ?
- | | ΔH | ΔS | ΔG |
|----|------------|------------|------------|
| a. | + | - | 0 |
| b. | - | + | 0 |
| c. | - | + | - |
| d. | + | - | - |
| e. | - | - | - |
37. Given the following data, calculate the normal boiling point for formic acid (HCOOH).
- | | ΔH_f° (kJ/mol) | S° (J/mol K) |
|-------------------|-----------------------------|---------------------|
| HCOOH(l) | -410 | 130 |
| HCOOH(g) | -363 | 251 |
- 2.57 K
 - 1730°C
 - 388°C
 - 82°C
 - 115°C**
38. The standard free energy of formation of nitric oxide, NO , at 1000. K (roughly the temperature in an automobile engine during ignition) is 78 kJ/mol. Calculate the equilibrium constant for the reaction
 $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$
 at 1000. K. ($R = 8.3145 \text{ J/K}\cdot\text{mol}$)
- $1.8 \cdot 10^{-19}$
 - 0.99
 - $4.3 \cdot 10^{-10}$
 - $7.1 \cdot 10^{-9}$**
 - $8.4 \cdot 10^{-5}$

39. Water gas, a commercial fuel, is made by the reaction of hot coke carbon with steam.
 $C(s) + H_2O(g) \leftrightarrow CO(g) + H_2(g)$
 When equilibrium is established at $800^\circ C$ the concentrations of CO , H_2 , and H_2O are $4.00 \cdot 10^{-2}$, $4.00 \cdot 10^{-2}$, and $1.00 \cdot 10^{-2}$ mole/liter, respectively. Calculate the value of ΔG° for this reaction at $800^\circ C$. ($R = 8.3145 \text{ J/K}\cdot\text{mol}$)
- a. 109 kJ
 b. -43.5 kJ
 c. 193 kJ
d. 16.3 kJ
 e. none of these
40. Given the reaction:
 $CH_3CO_2H(aq) \rightleftharpoons H^+(aq) + CH_3CO_2^-(aq)$
 at $25^\circ C$, $K_a = 1.8 \cdot 10^{-5}$
 What is ΔG° at $25^\circ C$? ($R = 8.3145 \text{ J/K}\cdot\text{mol}$)
- a. -27,000 J
b. +27,000 J
 c. -2300 J
 d. +2300 J
 e. +270 J

Show your work Problems (i.e. a number is not an acceptable answer, I must see your calculations for full credit!)

41. An aqueous solution contains dissolved $C_6H_5NH_3Cl$ and $C_6H_5NH_2$ ($pK_b = 9.42$). The concentration of $C_6H_5NH_2$ is 0.50 M and the pH is 4.20.
- a. Calculate the concentration of $C_6H_5NH_3^+$ in this buffer solution (4 points)
 pK_b for $C_6H_5NH_2 = -\log(3.8 \cdot 10^{-10}) = 9.42$; pK_a for $C_6H_5NH_3^+ = 14.00 - 9.42 = 4.58$
 $pH = pK_a + \log([C_6H_5NH_2]/[C_6H_5NH_3^+])$, $4.20 = 4.58 + \log(0.50 \text{ M}/[C_6H_5NH_3^+])$
 $-0.38 = \log(0.50 \text{ M}/[C_6H_5NH_3^+])$, $[C_6H_5NH_3^+] = [C_6H_5NH_3Cl] = 1.2 \text{ M}$
- b. Calculate the pH after 4.0 g of $NaOH$ (s) is added to 1.0 L of this solution (Neglect any volume changes). (6 points)

$$4.0 \text{ g } NaOH \cdot 1 \text{ mol } NaOH/40.00 \text{ g} \cdot 1 \text{ mol } OH^-/\text{mol } NaOH = 0.10 \text{ mol } OH^-$$

$$[OH^-] = 0.10 \text{ mol}/1.0 \text{ L} = 0.10 \text{ M}$$

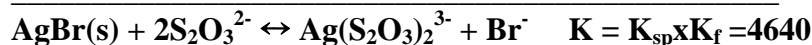
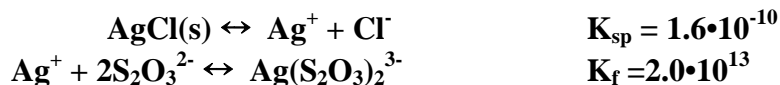
	$C_6H_5NH_3^+$	+	OH^-	\leftrightarrow	$C_6H_5NH_2$	+	H_2O
Before	1.2 M		0.10 M		0.50 M		
Change	-0.10		-0.10		+0.10		
After	1.1		0		0.60		

A buffer solution exists. $pH = 4.58 + \log(0.60/1.1) = 4.32$

42. A solution of sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) is used to dissolve silver chloride (AgCl , $K_{\text{sp}} = 1.6 \cdot 10^{-10}$). What mass of AgCl can dissolve in 1.00 L of 0.250 M $\text{Na}_2\text{S}_2\text{O}_3$? Ag^+ reacts with $\text{S}_2\text{O}_3^{2-}$ to form a complex ion:



(10 points)



	$\text{AgCl}(\text{s})$	+	$2\text{S}_2\text{O}_3^{2-}$	\leftrightarrow	$\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$	+	Cl^-
Initial			0.250 M		0		0
	S mol/L AgCl(s) dissolved to reach equilibrium						
Change	-s		-2s		+s		+s
Equil.			0.250 - 2s		s		s

$$K = s^2 / (0.250 - 2s)^2 = 4640; \text{ Taking the square root of both sides:}$$

$$s / (0.250 - 2s) = 68.12, s = 17.03 - 136.24 s, s = 0.124 \text{ mol/L}$$

$$1.00 \text{ L} \times 0.124 \text{ mol AgCl/L} \times 143.35 \text{ g AgCl/mol AgCl} = 17.8 \text{ g AgCl}$$

43. A 0.20 M sodium chlorobenzoate ($\text{NaC}_7\text{H}_4\text{ClO}_2$) solution has a pH of 8.65. Calculate the pH of a 0.20 M chlorobenzoic acid ($\text{HC}_7\text{H}_4\text{ClO}_2$) solution. (10 points)

From the pH, $\text{C}_7\text{H}_4\text{ClO}_4^-$ is a weak base. Use the weak base data to determine K_b for $\text{C}_7\text{H}_4\text{ClO}_4^-$ (which I abbreviate as CB^-)

	CB^-	+	H_2O	\leftrightarrow	HCB	+	OH^-
Initial	0.20 M				0		~0
Equil.	0.20 - x				x		x

$$\text{Since pH} = 8.65, \text{pOH} = 5.35 \text{ and } [\text{OH}^-] = 10^{-5.35} = 4.5 \cdot 10^{-6} \text{ M} = x$$

$$K_b = [\text{HCB}][\text{OH}^-] / [\text{CB}^-] = x^2 / (0.20 - x) = (4.5 \cdot 10^{-6})^2 / (0.20 - 4.5 \cdot 10^{-6}) = 1.0 \cdot 10^{-10}$$

Since CB^- is a weak base, HCB, chlorobenzoic acid, is a weak acid. Solving the weak acid problem:

	HCB	\leftrightarrow	CB^-	+	H^+
Initial	0.20 M		0		~0
Equil.	0.20 - x		x		x

$$K_a = K_w / K_b = 1.0 \cdot 10^{-14} / 1.0 \cdot 10^{-10} = 1.0 \cdot 10^{-4} = x^2 / (0.20 - x) \approx x^2 / (0.20)$$

$$x = [\text{H}^+] = 4.5 \cdot 10^{-3} \text{ M}; \text{ pH} = 2.35$$

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44. Consider the titration of 100.0 mL of 0.200 M acetic acid ($K_a = 1.8 \cdot 10^{-5}$) by 0.100 M KOH. Calculate the pH of the resulting solution after the following volumes of KOH have been added.

a. 0.00 mL (4 points)

Only a weak acid is present. Solving the weak acid problem:

$$K_a = 1.8 \cdot 10^{-5} = x^2 / (0.20 - x) \approx x^2 / (0.20), x = [\text{H}_3\text{O}^+] = 1.9 \cdot 10^{-3} \text{ M}$$

$$\text{pH} = 2.72$$

b. 50.0 mL (4 points)

The added base will react completely with the best acid present, $\text{CH}_3\text{CO}_2\text{H}$.

$$\text{mmol CH}_3\text{CO}_2\text{H present} = 100.0 \text{ mL} \cdot (0.200 \text{ mmol/L}) = 20.0 \text{ mmol}$$

$$\text{mmol OH}^- \text{ added} = 50.0 \text{ mL} \cdot (0.100 \text{ mmol/L}) = 5.00 \text{ mmol}$$

$$\text{mmol CH}_3\text{CO}_2\text{H present after reaction: } 20 \text{ mmol} - 5 \text{ mmol} = 15 \text{ mmol}$$

After the reaction, both weak acid and its conjugate base are present. Using Henderson-Hasselbach equation to solve for pH:

$$\text{pH} = \text{p}K_a + \log[\text{CH}_3\text{CO}_2^-] / [\text{CH}_3\text{CO}_2\text{H}] = -\log(1.8 \cdot 10^{-5}) + \log(5/15)$$

$$\text{pH} = 4.74 + (-0.477) = 4.26$$

c. 100.0 mL (4 points)

$$\text{mmol CH}_3\text{CO}_2\text{H present} = 100.0 \text{ mL} \cdot (0.200 \text{ mmol/L}) = 20.0 \text{ mmol}$$

$$\text{mmol OH}^- \text{ added} = 100.0 \text{ mL} \cdot (0.100 \text{ mmol/L}) = 10.00 \text{ mmol}$$

$$\text{mmol CH}_3\text{CO}_2\text{H present after reaction: } 20 \text{ mmol} - 10 \text{ mmol} = 10 \text{ mmol}$$

$$\text{pH} = \text{p}K_a + \log[\text{CH}_3\text{CO}_2^-] / [\text{CH}_3\text{CO}_2\text{H}] = -\log(1.8 \cdot 10^{-5}) + \log(10/10)$$

$$\text{pH} = \text{p}K_a = 4.74$$

d. 150.0 mL (4 points)

$$\text{mmol CH}_3\text{CO}_2\text{H present} = 100.0 \text{ mL} \cdot (0.200 \text{ mmol/L}) = 20.0 \text{ mmol}$$

$$\text{mmol OH}^- \text{ added} = 150.0 \text{ mL} \cdot (0.100 \text{ mmol/L}) = 15.00 \text{ mmol}$$

$$\text{mmol CH}_3\text{CO}_2\text{H present after reaction: } 20 \text{ mmol} - 15 \text{ mmol} = 5 \text{ mmol}$$

$$\text{pH} = \text{p}K_a + \log[\text{CH}_3\text{CO}_2^-] / [\text{CH}_3\text{CO}_2\text{H}] = -\log(1.8 \cdot 10^{-5}) + \log(15/5)$$

$$\text{pH} = 4.74 + (0.48) = 5.22$$

e. 200.0 mL (4 points)

$$\text{mmol CH}_3\text{CO}_2\text{H present} = 100.0 \text{ mL} \cdot (0.200 \text{ mmol/L}) = 20.0 \text{ mmol}$$

$$\text{mmol OH}^- \text{ added} = 200.0 \text{ mL} \cdot (0.100 \text{ mmol/L}) = 20.00 \text{ mmol}$$

$$\text{mmol CH}_3\text{CO}_2\text{H present after reaction: } 20 \text{ mmol} - 20 \text{ mmol} = 0 \text{ mmol}$$

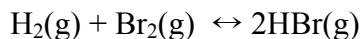
$$\text{Total volume} = 300 \text{ mL. } [\text{CH}_3\text{CO}_2^-] = 20.0 \text{ mmol} / 0.3 \text{ L} = 0.0667 \text{ M}$$

This is the equivalence point. Enough base has been added to convert all acid into its conjugate base. This is therefore a weak base problem:

$$K_b = K_w / K_a = 5.6 \cdot 10^{-10} = x^2 / (0.0667 - x) \approx x^2 / 0.0667, x = [\text{OH}^-] = 6.1 \cdot 10^{-6} \text{ M}$$

$$\text{pOH} = 5.21; \text{pH} = 8.79$$

45. Consider the reaction



Where $\Delta H^\circ = -103.8 \text{ kJ/mol}$. In a particular experiment, equal moles of $\text{H}_2(\text{g})$ at 1.00 atm and $\text{Br}_2(\text{g})$ at 1 atm were mixed in a 1.00 L flask at 25°C and allowed to reach equilibrium. Then the molecules of H_2 at equilibrium were counted using a very sensitive technique, and $1.10 \cdot 10^{13}$ molecules were found. For this reaction, calculate the values of K , ΔG° , and ΔS° . ($R = 8.3145 \text{ J/K}\cdot\text{mol} = 0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K}$) (10 points)

At equilibrium:

$$P_{\text{H}_2} = nRT/V = (1.10 \cdot 10^{13} / 6.022 \cdot 10^{23}) \text{ mol} \times (0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(298 \text{ K})/1\text{L} = 4.47 \cdot 10^{-10} \text{ atm}$$

The pressure of H_2 decrease from 1 atm to $4.47 \cdot 10^{-10}$ atm. Essentially, all of the H_2 and Br_2 has reacted. Therefore, $P_{\text{HBr}} = 2.00 \text{ atm}$ since there is a 2:1 ratio between HBr and H_2 in the balanced equation. Since we began with equal moles of H_2 and Br_2 , then we will have equal moles of H_2 and Br_2 at equilibrium. Therefore, $P_{\text{H}_2} = P_{\text{Br}_2} = 4.47 \cdot 10^{-10} \text{ atm}$.

$$K = P_{\text{HBr}}^2 / (P_{\text{H}_2} \cdot P_{\text{Br}_2}) = (2.00)^2 / (4.47 \cdot 10^{-10})^2 = 2.00 \cdot 10^{19}$$

$$\Delta G^\circ = -RT \ln K = -8.3145 \text{ J/K}\cdot\text{mol}(298 \text{ K}) \cdot \ln(2.00 \cdot 10^{19}) = -1.10 \cdot 10^5 \text{ J/mol}$$

$$\Delta S^\circ = (\Delta H^\circ - \Delta G^\circ)/T = -103,800 \text{ J/mol} - (-1.10 \cdot 10^5 \text{ J/mol})/298 \text{ K} = 20 \text{ J/K}\cdot\text{mol}$$