CHEMISTRY 204	Name KEY
Hour Exam II	
March 27, 2025	Signature
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	T.A
	Section

This exam contains 23 questions on 9 numbered pages. Check now to make sure you have a complete exam. You have two hours to complete the exam. Determine the **best** answer to the first 20 questions and enter these on the special answer sheet. Also, circle your responses in this exam booklet. Show all of your work and provide complete answers to questions 21, 22 and 23.

1-20	(60 pts.)	
21	(10 pts.)	
22	(20 pts)	
23	(30 pts)	
Total	(120 pts)	

Useful Information:

- Unless otherwise noted, all solutions referred to on this exam are aqueous solutions at 25°C.
- On this exam, H_3O^+ and H^+ are used interchangeably.

$$K_{\rm w} = [{\rm H}^+][{\rm O}{\rm H}^-] = 1.0 \text{ x } 10^{-14} \text{ at } 25^{\circ}{\rm C}.$$

For
$$ax^2 + bx + c = 0$$
, $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$pH = -log[H_3O^+] = -log[H^+]$$

$$K_{\rm a} = \frac{[{\rm H}^+]^2 - K_{\rm w}}{[{\rm HA}]_{\rm o} - \frac{[{\rm H}^+]^2 - K_{\rm w}}{[{\rm H}^+]}}$$

- 1. You go to the stockroom and find four bottles containing different aqueous weak acid solutions. The relative initial concentrations of each are such that the percent dissociation is the same in all of the samples. Which solution has the **lowest pH**?
 - a) HNO₂ b) HCN c) HF d) HC₂H₃O₂ e) All pH values are the same.
- 2. You go to the stockroom and find four bottles containing different aqueous weak acid solutions at the same concentration. You add concentrated HCl(*aq*) to each until the pH of each solution is 1.00. For which solution is the concentration of the conjugate base the **smallest**?
 - a) HNO₂ b) HCN c) HF d) HC₂H₃O₂ e) All concentrations are the same.
- 3. You dilute an aqueous solution by doubling the volume with water. The pH of the diluted solution increases by 0.3 compared to the original solution. How many of the following could correctly describe the original solution?
 - A weak acid.
 - A weak base
 - A strong acid
 - A strong base
 - A buffered solution
 - a) 1 b) 2 c) 3 d) 4 e) 5
- 4. Methylamine (CH₃NH₂) is an organic precursor to both pharmaceuticals and pesticides. It is water soluble and acts as a base in water ($K_b = 4.4 \times 10^{-4}$). Suppose you titrate a 2.00 x $10^{-3} M$ aqueous solution of methylamine to the equivalence point with a 2.00 x $10^{-3} M$ solution of HCl(*aq*). Determine the pH at equivalence.
 - a) 6.67 b) 6.74 c) 6.82 d) 7.18 e) 7.33
- 5. You dissolve 3.14 mg of NaOH in 314 L of aqueous solution. Determine the pH of this solution.
 - a) 7.40 b) 7.42 c) 7.45 d) 7.49 e) 7.54
- 6. Consider 0.010*M* solutions of four different aqueous acids. Which of the acid solutions has the **lowest pH**?
 - a) H₃PO₄ b) H₂SO₄ c) HCl d) HF e) Two or more are equally low.
- 7. Your lab partner is up to it again! Instead of diluting a 2.00*M* sulfuric acid solution with water, your lab partner used 2.00*M* hydrochloric acid. That is, 1.00 L of 2.00*M* H₂SO₄(*aq*) was mixed with 1.00 L of 2.00*M* HCl(*aq*). Determine [SO₄^{2–}] in this mixture at equilibrium.

- 8. Determine the pH of a 1.00*M* aqueous solution of NaHSO₃.
 - a) 3.50 b) 4.41 c) 6.09 d) 8.91 e) 10.50
- 9. You titrate 100.0 mL of a 0.500*M* HCN(*aq*) solution with 1.00*M* NaOH(*aq*). Determine the pH at the equivalence point.
 - a) 9.16 b) 9.25 c) 10.31 d) 11.36 e) 11.45
- 10. You discover an unlabeled solution of HCN(aq) and find that it has a pH of 6.79. What should the label read?
 - a) $1.6 \ge 10^{-5} M$ b) $2.6 \ge 10^{-5} M$ c) $4.2 \ge 10^{-5} M$ d) $6.8 \ge 10^{-5} M$ e) $8.3 \ge 10^{-5} M$
- 11. Approximately what mass of NaOH(s) must be added to 100.0mL of $1.00M \text{ HC}_2\text{H}_3\text{O}_2(aq)$ to result in a solution with pH = 3.14? Assume the solid sodium hydroxide does not appreciable change the volume.
 - a) 28 mg b) 63 mg c) 96 mg d) 117 mg e) 142 mg
- 12. About how much water **must be added** to 100.0mL of $1.00M \text{ HC}_2\text{H}_3\text{O}_2(\text{aq})$ to result in a solution with pH = 3.14?

a) 489 mL b) 589 mL c) 3.25 L d) 4.95 L e) 6.28 L

13. You are titrating 50.0 mL of a 0.100*M* aqueous solution of a weak diprotic acid with 0.100*M* NaOH(*aq*). Which of the following volumes of 0.100*M* NaOH(*aq*) does **not** constitute a "special point" at which we can determine the pH of the solution simply by using one or both of the K_a values of the acid?

a) 25.0 mL b) 50.0 mL c) 75.0 mL d) 100.0 mL e) They are all "special".

14. How many of the following will **not** produce a buffered solution?

- 100.0 mL of 0.100*M* Na₂CO₃(*aq*) and 50.0 mL of 0.100*M* NaOH(*aq*)
- 100.0 mL of 0.100*M* Na₂CO₃(*aq*) and 50.0 mL of 0.100*M* HCl(*aq*)
- 100.0 mL of 0.100*M* Na₂CO₃(*aq*) and 25.0 mL of 0.200*M* HCl(*aq*)
- 100.0 mL of 0.100*M* Na₂CO₃(*aq*) and 75.0 mL of 0.200*M* HCl(*aq*)
- 50.0 mL of 0.200*M* Na₂CO₃(*aq*) and 5.00 mL of 1.00*M* HCl(*aq*)

a) 1 b) 2 c) 3 d) 4 e) 5

- 15, 16. Recall the demonstration from lecture in which we produced the solids silver chromate and silver chloride by adding $AgNO_3(aq)$ to a solution containing aqueous sodium chromate, $Na_2CrO_4(aq)$, and aqueous sodium chloride, NaCl(aq).
- 15. Determine the solubility of Ag₂CrO₄(*s*) ($K_{sp} = 9.0 \times 10^{-12}$) in pure water at 25°C.
 - a) $9.0 \ge 10^{-12} M$ b) $3.0 \ge 10^{-6} M$ c) $1.3 \ge 10^{-4} M$ d) $1.7 \ge 10^{-4} M$ e) $2.6 \ge 10^{-4} M$
- 16. Determine the solubility of Ag₂CrO₄(*s*) ($K_{sp} = 9.0 \ge 10^{-12}$) in 1.00*M* NaCl(*aq*) 25°C. The K_{sp} for AgCl = 1.6 $\ge 10^{-10}$.
 - a) $1.6 \ge 10^{-10} M$ b) $2.6 \ge 10^{-4} M$ c) 0.046 Md) 0.50 Me) 1.00 M

17-18. Consider the formation of the complex ion $Ag(NH_3)_2^+(aq)$ when $Ag^+(aq)$ and $NH_3(aq)$ react:

 $\begin{array}{ll} \operatorname{Ag}^+(aq) + \operatorname{NH}_3(aq) & \Longrightarrow & \operatorname{Ag}(\operatorname{NH}_3)^+(aq) & \operatorname{K}_1 = 2.1 \times 10^3 \\ \operatorname{Ag}(\operatorname{NH}_3)^+(aq) + \operatorname{NH}_3(aq) & \Longleftrightarrow & \operatorname{Ag}(\operatorname{NH}_3)_2^+(aq) & \operatorname{K}_2 = 8.2 \times 10^3 \end{array}$

50.0 mL of 1.00 x 10^{-3} M AgNO₃ is reacted with 50.0 mL of 15.00 M NH₃.

17. Determine the equilibrium concentration of Ag^+ .

a)	$5.2 \ge 10^{-13} M$
b)	$3.9 \ge 10^{-12} M$
c)	$2.6 \ge 10^{-11} M$
d)	$3.2 \ge 10^{-9} M$
e)	$3.2 \ge 10^{-8} M$

18. Assuming that none of the silver containing species have an effect on the pH, determine the pH of the solution. (K_b for NH₃ = 1.8 x 10⁻⁵).

a) 7.00	b) 9.26	c) 11.63	d) 12.06	e) 12.22
a) 7.00	0) 9.20	C) 11.03	u) 12.00	6) 12.2

19-20. Indicate which of the graphs below **best** represents each plot described. A graph may be used once, more than once, or not at all.



- 19. $[A^{2-}](y)$ vs. $[H_2A](x)$ for which H_2A is a typical weak aqueous diprotic acid. E
- 20. pH (y) vs. [NaHA] (x) for which H_2A is a typical weak aqueous diprotic acid. E

- 21. As we discussed in class, we make assumptions when solving acid-bases problems, which is fine as long as we know what these assumptions are and how to check them. For example, in solving problems with a **weak acid** (**HA**) **in water**, we often simplify our calculations and make the assumption that at equilibrium, $[H_3O^+] = [A^-]$. Let's evaluate this assumption under various conditions. Please address the following questions in your answer:
 - Why is the assumption $[H_3O^+] = [A^-]$ not exactly true? Why are we able to generally make this assumption?
 - As we **dilute** the acid, does this assumption become **more or less valid**?
 - Consider two solutions each with a **different weak acid** at the **same initial concentration**. Is the assumption **more or less valid** for the acid with the **higher** *K*_a **value**?

Explain your answers without using sample calculations. **Full credit** is reserved for a **coherent**, **legible**, **and logical discussion** that is **limited to this page**. [10 points]

As we dilute less valid.

Higher K_a value, more valid.

- 22. In this problem you will compare aqueous solutions of the weak acids HF ($K_a = 7.2 \times 10^{-4}$), and HC₂H₃O₂ ($K_a = 1.8 \times 10^{-5}$). Consider two 1.00-L solutions, one of 1.00*M* HF, and the other of 1.00*M* HC₂H₃O₂.
 - a. We discussed how diluting a weak acid increases the percent dissociation. Consider that we dilute each 1.00-L sample of acid such that it is 99.0% dissociated.
 - Determine the **final volume** of each solution.
 - Determine the following **ratio** in each solution:

 $\frac{[H_3O^+]}{[A^-]}$ (where A⁻ is the conjugate base of each acid)

Show all work. Full credit is reserved for a coherent and logical approach that we can follow. [10 points]

For HF: V = 136,125L Ratio of [H₃O⁺]/[F⁻]= 1.000189 = 1

For HC₂H₃O₂: V = 7,807,571 L Ratio of [H₃O⁺]/[C₂H₃O₂⁻] = 1.43375 22. b. As we discussed, the concentration for any given species cannot go to zero at equilibrium, so an acid will not dissociate completely. Determine the **theoretical maximum percent dissociation** for **each** acid (HF, $K_a = 7.2 \times 10^{-4}$, and HC₂H₃O₂, $K_a = 1.8 \times 10^{-5}$) when we dilute them.

Show all work and briefly explain your reasoning. Full credit is reserved for a coherent and logical approach that we can follow. Please report your answer with five significant figures. [10 points]

For **HF**, percent dissociation = **99.986%**

For HC₂H₃O₂, percent dissociation = 99.448%

- 23. When sending samples away to Waste Management, we need to specify the identities and amounts of everything in the container. Let's consider the waste generated from the world famous "water to Kool-Aid to milk" demonstration that I performed in lecture. In cleaning up this demonstration, we pour everything back into the "pitcher" (Erlenmeyer flask). Forgetting about the few drops of phenolphthalein (which will not significantly change our calculations), we use the following for the demonstration:
 - 970.0 mL deionized water
 - 10.00 mL of 1.000*M* NaOH(*aq*)
 - 10.00 mL of $6.000M H_2SO_4(aq) (K_{a2} = 1.2 \times 10^{-2})$
 - 10.00 mL saturated (that is, maximum concentration) BaCl₂(*aq*)

Knowing that the K_{sp} for BaCl₂ = 22.45, the K_{sp} for Ba(OH)₂ = 5.0 x 10⁻³, and the K_{sp} for BaSO₄ = 1.5 x 10⁻⁹, **determine the identities and concentrations of all species** in the solution, the **pH of the solution**, and **the mass of any solid(s) present**, after all solutions have been returned to the flask and reach equilibrium.

Show all work and briefly explain your reasoning/approach. Full credit is reserved for a systematic, coherent solution that we can follow. Please use the next page if needed. [30 points].

Cations	Anions
$[Na^+] = 0.0100M$	$[HSO_4^-] = 0.036316M$
$[H_3O^+] = 0.073684M$	$[SO_4^{2-}] = 0.0059143M$
$[Ba^{2+}] = 2.5362 \times 10^{-7} M$	$[Cl^{-}] = 0.03554M$
	$[OH^{-}] = 1.35714 \times 10^{-13} M$

pH = 1.13

Mass BaSO₄(s) = (0.01777 mol)(233.38g/mol) = **4.14716g BaSO₄(s**)

Ba(**OH**)₂ **won't ppt** (Q = 4.67 x $10^{-33} \ll K_{sp}$ (5.0 x 10^{-3})) BaCl₂ won't ppt (Q = 3.20 x $10^{-10} \ll K_{sp}$ (22.45)); very diluted from original NaCl "very soluble"

Checks

(Positive Charge = 0.0836845) (Negative Charge = 0.0836846)

Original "SO4" (H₂SO₄) = 0.060 mol Final "SO4" (HSO₄⁻ + SO₄²⁻ + BaSO₄) = 0.036316 + 0.0059143 + 0.01777 = 0.060

 $K_{\rm w}$, $K_{\rm sp}$ (BaSO₄), $K_{\rm a2}$ (HSO₄⁻) all check

23. Please continue any work for #23 on this page.