CHEMISTRY 1220
CHAPTER 16 PRACTICE EXAM

1. The pH of a 0.10 M solution of NH_3 containing 0.10 M NH_4Cl is 9.20. What is the **$[H_3O^+]$**?
a) 1.6×10^{-5} b) 1.0×10^{-1} c) 6.3×10^{-10} d) 1.7×10^{-10} e) 2.0×10^{-9}

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

$9.20 = -\log[H_3O^+]$

$[H_3O^+] = 10^{-9.20}$

$[H_3O^+] = 6.3 \times 10^{-10}$

2. Calculate the pH of an aqueous solution which is 0.0020 M $HClO_4$.
a) 1.30 b) 1.70 c) 2.30 d) 2.70 e) 2.00

$HClO_4(aq) \rightarrow H^+(aq) + ClO_4^-(aq)$

strong acid, fully dissociates

$[H^+] = 0.0020\ M$

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

$pH = -\log[0.0020]$

$pH = 2.70$

3. According to the Bronsted-Lowry Concept of acids and bases which of the following statements, a-d,
is FALSE?

a) A base is a species that accepts a proton. ✓

b) Acid-base reactions are restricted to aqueous solutions. ✗

c) Some species can act as either acids or bases, depending on what the other reactant is. ✓

d) NH_3 is a Bronsted base. ✓

e) All of the above, a-d, are part of this theory. ✗

4. Given that K_w for water is 2.40×10^{-14} (M^2) at 37°C, compute the pH of a neutral aqueous solution at 37°C (normal human body temperature). Answer the following TWO questions. What is the pH of a neutral solution at 37°C? AND If a solution has pH = 7.00 is it acidic, basic, or neutral at 37°C?

a) 6.82, acidic b) 6.82, basic c) 7.19, acidic d) 7.19, basic e) 7.00, neutral

$K_w = [H^+][OH^-] = 2.40 \times 10^{-14}$

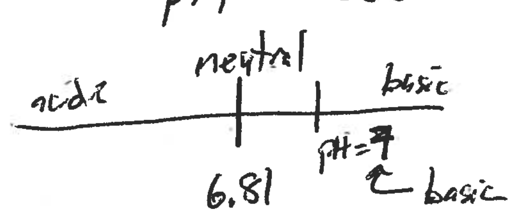
neutral soln $[H^+] = [OH^-] = x$

$x^2 = 2.40 \times 10^{-14}$

$x = 1.549 \times 10^{-7} = [H^+]$

$pH = -\log[1.549 \times 10^{-7}]$

$pH = 6.81$



5. The K_a values for HS^- and HPO_4^{2-} are 1.2×10^{-13} and 4.8×10^{-13} respectively. Therefore it follows the HS^- is a acid than HPO_4^{2-} and S_2^- is a base than PO_4^{3-}.
a) stronger, stronger b) stronger, weaker c) weaker, stronger d) weaker, weaker

$$HS^-_{(aq)} \rightleftharpoons H^+_{(aq)} + S^{2-}_{(aq)} \qquad K_a = 1.2 \times 10^{-13}$$

acid — conj. base

$$HPO_4^{2-}{}_{(aq)} \rightleftharpoons H^+_{(aq)} + PO_4^{3-}{}_{(aq)} \qquad K_a = 4.8 \times 10^{-13}$$

acid — conj. base

$K_a\ HPO_4^{2-} > K_a\ HS^-$ ∴ HPO_4^{2-} is a stronger acid than HS^- and PO_4^{3-} is a weaker conj. base than S^{2-}

6. What is the ionization constant of an acid if the hydronium ion concentration of a 0.500 M solution is 1.70×10^{-4} M?
a) 3.62×10^{-7} b) 2.89×10^{-8} c) 5.80×10^{-8} d) 1.16×10^{-7} e) 1.70×10^{-3}

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[1.70 \times 10^{-4}][1.70 \times 10^{-4}]}{[0.500]}$$

$$K_a = 5.8 \times 10^{-8}$$

7. Consider the following salts. Which one(s) when dissolved in water will produce an acidic solution?

1) NH_4Cl 2) $KHSO_4$ 3) NaCN

a) only 1 b) only 2 c) only 3 d) 1 and 2 e) 2 and 3

$NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

conj. acid of w. base ↳ acidic

conj. base of strong acid → neutral ⇒ acidic

$KHSO_4 \rightarrow K^+(aq) + HSO_4^-(aq)$

conj. acid of strong base ↳ neutral

conj. base of strong acid

but it's diprotic

$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$

produces H^+ acidic

$NaCN \rightarrow Na^+(aq) + CN^-(aq)$

conj. acid of strong base

conj. base of weak acid ⇒ basic

8. A 0.010 M solution of HNO_2 is 19% ionized. What is the K_a?

a) 4.5×10^{-4} b) 3.9×10^{-4} c) 3.6×10^{-4} d) 5.0×10^{-4} e) 5.4×10^{-4}

$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$

$0.010(0.19) = 0.0019$

	HNO_2	H^+	NO_2^-
I	0.010	0	0
C	−0.0019	+0.0019	+0.0019
E	0.0081	0.0019	0.0019

$$K_a = \frac{(0.0019)(0.0019)}{0.0081}$$

$$K_a = 4.5 \times 10^{-4}$$

9. What is the ~~pH of a~~ 0.20 M NH_4Cl solution (K_b: $NH_3 = 1.8 \times 10^{-5}$)?
a) 2.72 b) 3.11 c) 4.98 d) 5.12 e) 7.61

$K_a \cdot K_b = K_w$

$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

$K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$

	NH_4^+	H_2O	NH_3	H_3O^+
I	0.20	–	0	0
C	−x	–	+x	+x
E	0.20−x	–	x	x

$\frac{x^2}{0.20} = 5.56 \times 10^{-10}$

$x = 1.05 \times 10^{-5} = [H_3O^+]$

$pH = -\log(1.05 \times 10^{-5})$

$pH = 4.98$

10. Ascorbic acid, $H_2C_6H_6O_2$, is a diprotic acid. The K_1 and K_2 values are 7.9×10^{-5} and 1.6×10^{-12} respectively. What is the $C_6H_6O_2^{2-}$ ion concentration in a 0.10 M solution of ascorbic acid?
a) 1.6×10^{-6} b) 1.6×10^{-12} c) 7.9×10^{-12} d) 2.8×10^{-3} e) 5.6×10^{-3}

$H_2C_6H_6O_2 \rightleftharpoons H^+ + HC_6H_6O_2^-$

	$H_2C_6H_6O_2$	H^+	$HC_6H_6O_2^-$
I	0.10	0	0
C	−x	+x	+x
E	0.10−x	x	x

$\frac{x^2}{0.10} = 7.9 \times 10^{-5}$

$x = 0.0028$

$HC_6H_6O_2^- \rightleftharpoons H^+ + C_6H_6O_2^-$

	$HC_6H_6O_2^-$	H^+	$C_6H_6O_2^-$
I	0.0028	0.0028	0
C	−x	+x	+x
E	0.0028−x	0.0028+x	x

$\frac{(0.0028+x)(x)}{0.0028-x} = 1.6 \times 10^{-12}$

$x = 1.6 \times 10^{-12}$

11. What is the pH of a solution of 0.31 M acid and 0.65 M of its conjugate base if the ionization constant, K_a, is 5.22×10^{-7}?

a) 6.60 b) 6.81 c) 7.00 d) 7.21 e) 7.42

$pK_a = 6.28$

$pH = pK_a + \log \left(\frac{[A^-]}{[HA]}\right)$

$pH = 6.28 + \log \left(\frac{0.65}{0.31}\right)$

$pH = 6.60$

12. Rubidium hydroxide is a strong base. Compute $[Rb^+]$ and $[OH^-]$ for a solution that is prepared by dissolving 2.0 g of RbOH in enough water to make 200.0 mL of solution. (atomic weights: Rb = 85.47, O = 16.00, H = 1.008)

a) 1.9×10^{-2}, 1.9×10^{-2}

b) 1.9×10^{-2}, 5.3×10^{-13}

c) 5.3×10^{-13}, ~~1.9×10^{-2}~~

d) 9.8×10^{-2}, 9.8×10^{-2}

e) 9.8×10^{-1}, 9.8×10^{-1}

$2.0\text{ g RbOH} \times \frac{1\text{ mol RbOH}}{102.47\text{ g}} \times \frac{1\text{ mol Rb}^+}{1\text{ mol RbOH}} = \frac{0.019518\text{ mol Rb}^+}{0.2000\text{ L}} = 9.8 \times 10^{-2}\text{ M}$

$2.0\text{ g RbOH} \times \frac{1\text{ mol RbOH}}{102.47\text{ g}} \times \frac{1\text{ mol OH}^-}{1\text{ mol RbOH}} = \frac{0.019518\text{ mol OH}^-}{0.2000\text{ L}} = 9.8 \times 10^{-2}\text{ M}$

13. You are given two solutions: 0.50 M HCl (aq) and 0.50 M $Ca(OH)_2$(aq). What is the [H+] in the HCl solution? What is the [OH-] in the $Ca(OH)_2$ solution? (The solutions are NOT mixed together).

[H+] [OH-]

a) [H+] = 0.50 M, [OH-] = 0.50 M
b) [H+] = 0.25 M, [OH-] = 1.0 M
c) [H+] = 0.50 M, [OH-] = 0.25 M
d) [H+] = 0.25 M, [OH-] = 0.25 M
e) [H+] = 0.50 M, [OH-] = 1.0 M

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$ strong acid ⇒ fully dissociates

0.50 M   0.50 M

$Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$ strong base ⇒ fully dissociates

0.50 M —×2→ 1.0 M

14. How many grams of phosphoric acid are there in 175 mL of a 3.5 M solution of phosphoric acid (MW 98.00 g/mol)?

a) 0.61 g
b) 60 g
c) 21 g
d) 4.9 g
e) 610 g

$\frac{3.5 \text{ mol } H_3PO_4}{L} \times 0.175 L \times \frac{98.00 g}{1 \text{ mol } H_3PO_4} = 60.03 g$

15. A solution is prepared by dissolving 516.5 mg of oxalic acid ($C_2H_2O_4$, 90.00 g/mol) to make 100.0 mL of solution. A 10.00 mL portion is then diluted to 250.0 mL. What is the molarity of the final solution?

a) 2.295×10^{-3} M
b) 6.341×10^{-2} M
c) 3.172×10^{-3} M
d) 4.685×10^{-2} M
e) 1.889×10^{-3} M

$$0.5165\, g\, C_2H_2O_4 \times \frac{1\, mol}{90.00\, g} = \frac{0.005739\, mol}{0.1000\, L} = 0.05739\, M$$

$$0.05739\, \frac{mol}{L} \times 0.010\, L = \frac{5.739 \times 10^{-4}\, mol}{0.250\, L} = 2.296 \times 10^{-3}\, M$$

16. What is the conjugate base of methylamine, CH_3NH_2?

a) CH_3NH^+ b) CH_3NH^- c) $CH_3NH_2^+$ d) $CH_3NH_2^-$ e) $CH_3NH_3^+$

17. What is the conjugate acid of methylamine, CH_3NH_2?

a) CH_3NH^+ b) CH_3NH^- c) $CH_3NH_2^+$ d) $CH_3NH_2^-$ e) $CH_3NH_3^+$

acid

$$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3NH^-(aq)$$

conj. base

base

$$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + CH_3NH_3^+(aq)$$

conj. acid

18. The K_a values for HCNO and HNO_2 are 2.2 x 10^{-4} and 4.5 x 10^{-4} respectively. Therefore it follows the HCNO is a __ acid than HNO_2 and CNO^- is a ___ base than NO_2^-.

a) stronger, stronger

b) stronger, weaker

c) weaker, stronger

d) weaker, weaker

$HCNO_{(aq)} \rightleftharpoons H^+_{(aq)} + CNO^-_{(aq)}$ $K_a = 2.2 \times 10^{-4}$

$HNO_{2\,(aq)} \rightleftharpoons H^+_{(aq)} + NO_2^-_{(aq)}$ $K_a = 4.5 \times 10^{-4}$

K_a for HCNO is less than HNO_2 ∴ HCNO is weaker

weaker the acid (HCNO) the stronger the conj. base (CNO^-)

19. What change will be observed for the following reaction if a few drops of NaOH are added?

$$HNO_2 + H_2O \rightleftharpoons NO_2^- + H_3O^+$$

a) a decrease in the fraction of acid dissociated

b) an increase in the fraction of acid dissociated

c) no change in the fraction of acid dissociated

OH^- will neutralize H_3O^+, which will ↓ $[H_3O^+]$ and the Eq. will shift (R)

This increases the fraction of acid dissociated.

20. Given K_a values of 1.0 x 10^{-10} and 6.8 x 10^{-8} for C_6H_5OH and $C_5H_5NH^+$ respectively, calculate the equilibrium constant for the following reaction.

$$C_6H_5O^- + C_5H_5NH^+ \rightleftharpoons C_6H_5OH + C_5H_5N$$

a) 6.8 x 10^2 b) 0.15 c) 1.5 x 10^{-3} d) 6.8 x 10^{-2} e) 6.8 x 10^{-8}

$C_6H_5OH \rightleftharpoons H^+ + C_6H_5O^-$ $K_a = 1.0 \times 10^{-10}$

$C_5H_5NH^+ \rightleftharpoons H^+ + C_5H_5N$ $K_a = 6.8 \times 10^{-8}$

$C_6H_5O^- + H^+ \rightleftharpoons C_6H_5OH$ $K = \frac{1}{1.0 \times 10^{-10}}$

$C_5H_5NH^+ \rightleftharpoons H^+ + C_5H_5N$ $K = 6.8 \times 10^{-8}$

$C_6H_5O^- + C_5H_5NH^+ \rightleftharpoons C_6H_5OH + C_5H_5N$ $K = 6.8 \times 10^2$

21. The value of K_a in water at 25°C for benzoic acid ($C_6H_5CO_2H$) is 6.46 x 10^{-5} M. Calculate the pH of an aqueous solution with a total concentration of benzoic acid equal to 0.025 M.

a) 1.29 b) 2.09 c) 2.90 d) 3.10 e) 3.90

	C_6H_5COOH	$\rightleftharpoons$ H^+ +	$C_6H_5COO^-$
I	0.025	0	0
C	$-x$	$+x$	$+x$
E	$0.025-x$	x	x

$$\frac{x^2}{0.025} = 6.46 \times 10^{-5}$$

$x = 0.00127$

pH $= -\log(0.00127)$

pH $= 2.90$

22. The value of K_a in water at 25°C for chloroacetic acid is 1.35×10^{-3} M. Calculate the pH of an aqueous solution with an initial concentration of chloroacetic acid equal to 0.10 M.

a) 1.35 b) 1.96 c) 2.14 d) 3.65 e) 3.35

$HA \rightleftharpoons H^+ + A^-$

	HA	H^+	A^-
I	0.10	0	0
C	$-x$	$+x$	$+x$
E	$0.10-x$	x	x

$$\frac{x^2}{0.10-x} = 1.35\times10^{-3}$$

$$x^2 + 1.35\times10^{-3}x - 1.35\times10^{-4} = 0$$

$$x = 0.010963$$

$pH = -\log(0.010963)$

$pH = 1.96$

23. Consider the following salts. Which one(s) when dissolved in water will produce an acidic solution?

1) NH_4Cl 2) $KHSO_4$ 3) NaCN

a) only 1 b) only 2 c) only 3 d) 1 and 2 e) 2 and 3

$NH_4Cl \rightarrow NH_4^+ + Cl^-$
acidic

$KHSO_4 \rightarrow K^+ + HSO_4^-$
acidic

$HSO_4^- + H_2O(l) \rightleftharpoons H^+ + SO_4^{2-}$

$NaCN \rightarrow Na^+ + CN^-$
basic

24. A 1.50 g sample of Vitamin C is dissolved in 100.0 mL of water and titrated with 0.250 M NaOH to the methyl orange equivalence point. The volume of the base used is 34.1 mL. What is the molecular weight of Vitamin C assuming one dissociable proton per molecule?

a) 176 b) 164 c) 152 d) 146 e) 139

MW = g/mol

$\frac{0.250 \text{ mol NaOH}}{L} \times 0.0341 L = 0.008525 \text{ mol NaOH} \times \frac{1 \text{ mol } OH^-}{1 \text{ mol NaOH}}$

At equiv pt. 1 mol H^+ = 1 mol OH^-

↳ 0.008525 mol H^+

$MW = 1.50 g / 0.008525 \text{ mol}$

$= 176$ g/mol

25. A 25.00 mL sample of 0.100 M HCl is titrated with 0.100 M NaOH. What is the pH of the solution at the points where 24.9 and 25.1 mL of NaOH have been added.

a) 3.00, 11.00
b) 3.30, 10.70
~~c) 3.30, 10.30~~
d) 3.70, 10.30
e) 3.70, 10.70

$\frac{0.100 \text{ mol HCl}}{L} \times 0.02500 L = 0.0025 \text{ mol } H^+$

$0.100 \text{ mol NaOH} \times 0.02490 L = 0.002490 \text{ mol } OH^-$

$[H^+] = \frac{1 \times 10^{-5} \text{ mol}}{0.0495 L} = 2.02 \times 10^{-4} M$

0.0025 mol H^+

pH = 3.70

$\frac{0.100 \text{ mol NaOH}}{L} \times 0.0251 L = 0.00251 \text{ mol NaOH} = 0.00251 \text{ mol } OH^-$

$- 0.0025 \text{ mol } H^+$

$1 \times 10^{-5} \text{ mol } OH^-$

$[OH^-] = \frac{1 \times 10^{-5} \text{ mol } OH^-}{0.0501 L} = 1.996 \times 10^{-4} M$

pOH = 3.70 pH = 10.3

26. What is the pH of a solution of 0.65 M acid and 0.51 M of its conjugate base if the pK_a is 5.30?

a) 5.19 b) 5.41 c) 5.62 d) 5.85 e) 6.05

$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$

$pH = 5.30 + \log\left(\frac{0.51}{0.65}\right)$

$pH = 5.19$

27. Hydrosulfuric acid (H_2S) has $K_1 = 1.1 \times 10^{-7}$ and $K_2 = 1.0 \times 10^{-13}$. What is the HS^- ion concentration of a 0.10 M solution of H_2S?

a) 1.0×10^{-4} b) 1.0×10^{-5} c) 3.3×10^{-4} d) 3.3×10^{-5} e) 1.1×10^{-7}

$H_2S \rightleftharpoons H^+ + HS^-$ $K_1 = 1.1 \times 10^{-7}$

	H_2S	H^+	HS^-
I	0.10	0	0
C	$-x$	$+x$	$+x$
E	$0.10 - x$	x	x

$\frac{x^2}{0.10} = 1.1 \times 10^{-7}$

$x = 1.0488 \times 10^{-4}$

28. A 0.0184 M solution of HCNO is 12.8% ionized. What is the K_a?

a) 1.1×10^{-3} b) 1.5×10^{-3} c) 1.9×10^{-3} d) 3.5×10^{-4} e) 2.9×10^{-4}

$HCNO \rightleftharpoons H^+ + CNO^-$

0.0184(0.128)

0.0023552

I 0.0184 0 0

C −x +x +x

E 0.016045 0.002355 0.002355

$$K_a = \frac{[0.002355][0.002355]}{[0.016045]}$$

$K_a = 3.5 \times 10^{-4}$

29. Given the following K_a values, determine which species is the strongest base.

HF 6.8×10^{-4} HNO_2 4.5×10^{-4} HCNO 2.2×10^{-4}

a) F^- b) NO_2^- c) CNO^- d) HF e) HCNO

↑ K_a : stronger the acid

↳ weaker conj. base

↓ K_a : weaker the acid

↳ stronger conj. base

HCNO has smallest K_a of the three

↳ weakest acid

↳ strongest base

30. A 25.00 mL sample of 0.100 M CH_3CO_2H is titrated with 0.100 M NaOH. What is the pH of the solution at the points where 25.0 and 25.5 mL of NaOH have been added? ($K_a = 1.8 \times 10^{-5}$)

a) 8.72, 11.00 b) 8.72, 9.85 c) 7.00, 10.00 d) 7.00, 9.85 e) 7.00, 8.00

$\frac{0.100 \text{ mol } CH_3COOH}{L} \times 0.02500 L = 0.0025 \text{ mol } CH_3COOH$

$\frac{0.100 \text{ mol NaOH}}{L} \times 0.02500 L = 0.0025 \text{ mol } OH^-$

$CH_3COOH + NaOH \rightarrow NaCH_3COO + H_2O$

	CH_3COOH	NaOH	$NaCH_3COO$
I	0.0025 mol	0	
add		0.0025 mol	
Neut	−0.0025	−0.0025	+0.0025 mol
End	0	0	0.0025 mol

$[C_2H_3O_2^-] = \frac{0.0025 \text{ mol}}{0.050 L}$

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

	$C_2H_3O_2^-$	H_2O	OH^-	$HC_2H_3O_2$
I	0.05	—	0	0
C	−x	—	+x	+x
E	0.05	—	x	x

$K_a \cdot K_b = K_w$

$K_b = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}}$

$K_b = \frac{x^2}{0.05} = 5.56 \times 10^{-10}$

$x = 5.27 \times 10^{-6}$ pOH = 5.27 pH = 8.72

At 25.5 mL (excess base)

$\frac{0.100 \text{ mol } OH^-}{L} \times 0.0255 L = 0.00255 \text{ mol } OH^-$

$-0.00250 \text{ mol } H^+$

$0.00005 \text{ mol } OH^- / 0.055 L$

pOH = 3

pH = 11