Exam 2 Test Prep

% Ionization	Buffer	pH	Common-Ion Effect	Irreversible Reaction
Cation	Oxyacid	Anion	pOH	Equivalence Point
Binary Acid	Lewis Acid	Titration	Reversible Reaction	Bronsted-Lowry Base
Lewis Base	Hydrolysis	Carboxylic Acid	Amphoteric	Bronsted-Lowry Acid

1. An acid that consists of an H, O, and one other element, which is a nonmetal, is an: oxyacid

2. An acid that consists of an H and one other element is a: binary acid

3. A proton donor: bronsted-lowry acid

4. The quantity of weak acid that ionizes in a solution, expressed as a percentage: % ionization

5. A negatively charged ion: anion

6. When the amount of acid is equal to the amount of base in a titration/reaction: equivalence point

7. A proton acceptor: bronsted-lowry base

8. A positively charged ion: cation

9. An organic acid that contains the -COOH group: carboxylic acid

10. Represents the hydrogen ion concentration: pH

11. If we have a solution containing several types of ions and equilibrium is achieved, when we add another species containing the same ion, to the existing solution, reduction in the degree of dissociation of the first species is observed. common-ion effect

12. A solution of a weak conjugate acid-base pair that resists drastic changes in pH: buffer

13. The chemical breakdown of a compound due to its reaction with water: hydrolysis

14. Represents the hydronium ion concentration: pOH

15. An e^- pair acceptor: lewis acid

16. A technique in which an acid or base of known concentration is added to an acid or base of unknown concentration: titration

17. The system must take another path to return to the original state: irreversible process

18. An e^- pair donor: lewis base

19. A compound that can act as either an acid or a base: amphoteric

20. The system can follow the same path in reverse to get back to the original state: reversible process

1. Calculate the pH of 0.00125 M HCl ($K_a = 3.5 \times 10^{-2}$)

R	HCl ←→	H^+ +	Cl^-
I	0.00125	0	0
C	− x	+ x	+ x
E	0.000125 − x	x	x

$K_a = \frac{x^2}{0.00125} = 3.5 \times 10^{-2}$

$[H^+] = x = 6.6 \times 10^{-6}$

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

$pH = 5.18$

2. What is the pH of a buffer that is 0.12 M in lactic acid ($HC_3H_5O_3$) and 0.10 M in sodium lactate? For lactic acid, $K_a = 1.4 \times 10^{-4}$

R	$HC_3H_5O_3$ (aq) ←→	H^+ (aq) +	$C_3H_5O_3^-$ (aq)
I	0.12M	0	0.10M
C	−x	+x	+x
E	(0.12 − x)	x	0.10 + x

$K_a = 1.4 \times 10^{-4} = \frac{x(0.10)}{0.12} =$

3. The K_a for niacin (C_5H_4NCOOH) is 1.6×10^{-5}. What is the pH of a 0.010 M solution of niacin?

R	C_5H_4NCOOH ←→	$C_5H_4NCOOH^-$ +	H^+
I	0.010	0	0
C	− x	+ x	+ x
E	0.010 − x	x	x

$K_a = \frac{x^2}{0.010} = 1.6 \times 10^{-5}$

$x = 3.9 \times 10^{-4}$

$pH = -\log(3.9 \times 10^{-4})$

$pH = 3.41$

4. Write the equilibrium equation for the base and solve for the pH: $K_b = 1.8 \times 10^{-5}$

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3][H_2O]}$

	NH_3 +	H_2O ←→	NH_4^+ +	OH^-
I	0.15	—	0	0
C	-x	—	+x	+x
E	0.15	—	x	x

$K_b = \frac{x^2}{0.15} = 1.8 \times 10^{-5}$

$x = [OH^-] = 1.6 \times 10^{-3}$

$pOH = -\log(1.6 \times 10^{-3})$
$= 2.80$

$pH = 14 - 2.8 = 11.2$

5. Calculate the pH of 0.119 M pyridine: $K_b = 1.7 \times 10^{-9}$

R	C_5H_5N +	H_2O ←→	$C_5H_5NH^+$ +	OH^-
I	0.119		0	0
C	-x		+x	+x
E	0.119		x	x

$K_b = 1.7 \times 10^{-9} = \frac{x^2}{0.119}$

$x = [OH^-] = 1.4 \times 10^{-5}\ M$

$pOH = -\log(1.42 \times 10^{-5})$
$= 4.848$

$pH = 14 - 4.848$
$= 9.15$

6. Find the molar solubility of Ag_2CrO_4 in pure water if the solubility product constant for silver chromate is 1.1 x 10^{-12}.

$K_{sp} = 1.1 \times 10^{-12}$

R	Ag_2CrO_4 →	$2Ag^+$ +	CrO_4^{2-}
I	—		
C	—	+2x	+x
E	—	2x	x

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

$K_{sp} = [2x]^2[x]$

$1.1 \times 10^{-12} = (4x^2)(x)$

$\sqrt[3]{2.75 \times 10^{-13}} = \sqrt[3]{x^3}$

$x = 6.5 \times 10^{-5}\ M$

7. What are ways that we can make a buffer?

- mix a weak acid and a salt of its conjugate base or a weak base and a salt of its conjugate acid

- add strong acid and partially neutralize a weak base or add strong base and partially neutralize a weak acid

8. What is the concentration of [Cl^-] in the final solution if you pour 10.0 mL of 0.10 M NaCl, 10.0 mL of 0.10 M KOH, and 5.0 mL of 0.20 M HCl solutions together to make a total volume 100.0 mL.

$(M_1V_1 + M_2V_2)/V$

$(0.10 \times 10 + 0.20 \times 5)/100$

$2/100$

$0.02M$

9. If the pH of a saturated solution of $Ba(OH)_2$ is 12. What is the value of solubility product (Ksp) of $Ba(OH)_2$?

$Ba(OH)_2 \rightleftharpoons Ba^{2+} + 2OH^-$

$pH = 12$

$pOH = 14 - 12 = 2$

$[OH^-] = 10^{-2} = 1 \times 10^{-2}$

law of cons. of ions $= 0.5 \times 10^{-2} = [Ba^{2+}]$

$Ksp = [Ba^{2+}][OH^-]^2$

$Ksp = (0.5 \times 10^{-2})(1.0 \times 10^{-2})$

$Ksp = 5.0 \times 10^{-7}$

10. How do we determine if a precipitate will form if we have the values of the reaction quotient, Q, and the solubility product constant, K_{sp}?

If $Q = Ksp$, the solution is saturated

If $Q > Ksp$, a precipitate will form

If $Q < Ksp$, a precipitate will not form

11. What is the pH range of an acid? Of a base? Neutral?

acid: 1 - 6

base: 8 - 14

neutral: 7

12. What is the difference between weak and strong acids/bases?

weak acids and bases do not dissociate completely in solution, strong acids and bases do.

13. Calculate the percent ionization in a 1.0M solution of nitrous acid. (Ka = 4.5 x 10^{-4})

R	$HNO_2 \rightleftharpoons$	H^+ +	NO_2^-
I	1.0M	0	0
C	-x	+x	+x
E	1.0-x	x	x

$4.5 \times 10^{-4} = \frac{x^2}{1.0}$

$x = \sqrt{(4.5 \times 10^{-4})(1)}$

$= 2.1 \times 10^{-2} M$

$\% I = \frac{[H^+]}{[1.0]} = 2.1\%$

14. Vinegar is a dilute solution of acetic acid (HC2H3O2). If the concentration of HC2H3O2 in a vinegar solution is 0.240M, calculate the percent ionization of acetic acid. (Ka = 1.75 x 10^{-5})

R	HA	$+ H_2O$	$\rightleftharpoons H^+$	$+ A^-$
I	0.240	—	0	0
C	$-x$	—	$+x$	$+x$
E	0.240	—	x	x

$$1.75 \times 10^{-5} = \frac{x^2}{0.240}$$

$$x = 0.002049$$

$$\%I = \frac{[H^+]}{[HA]} \times 100$$

$$\%I = 0.854\%$$