

MC Practice Test G.1

Acid/Base

Name _____ Per _____

- *This is practice – Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.*
- *DO NOT USE A CALCULATOR. You may use ONLY the blue periodic table.*
- *Try to work at a pace of 1.3 min per question. Time yourself. It is important that you practice working for speed.*

1. A 0.500 mole sample of which compound below, when placed in water gives the lowest concentration of anions.
 - a. NaNO_2
 - b. HNO_2
 - c. $\text{Pb}(\text{NO}_3)_2$
 - d. $\text{Al}(\text{NO}_3)_3$
 - e. a and b would produce the same amount
2. The conjugate base of HSO_4^- is
 - a. OH^-
 - b. H_2SO_4
 - c. SO_4^{2-}
 - d. HSO_4^-
 - e. H_3SO_4^+
3. What is the pH of an aqueous solution at 25.0 °C in which $[\text{H}^+]$ is 0.00250 M?
 - a. 2.60
 - b. -2.60
 - c. 3.40
 - d. -3.40
 - e. 11.4
4. What is the $[\text{OH}^-]$ in a solution that has a pH of 5
 - a. $1 \times 10^{-14} \text{ M}$
 - b. $1 \times 10^{-5} \text{ M}$
 - c. $1 \times 10^{-9} \text{ M}$
 - d. $1 \times 10^{+5} \text{ M}$
 - e. $1 \times 10^{+9} \text{ M}$
5. What is the concentration of hydronium ions, H_3O^+ , in a solution at 25.0 °C with $\text{pOH} = 4.282$?
 - a. 4.28 M
 - b. 9.72 M
 - c. $1.92 \times 10^{-10} \text{ M}$
 - d. $5.22 \times 10^{-5} \text{ M}$
 - e. $1.66 \times 10^4 \text{ M}$
6. What is the pH of a 0.00005 M solution of barium hydroxide?
 - a. 4.0
 - b. 4.3
 - c. 5.0
 - d. 9.7
 - e. 10.0
7. HZ is a weak acid. An aqueous solution of HZ is prepared by dissolving 0.020 mol of HZ in sufficient water to yield 1.0 L of solution. The pH of the solution was 5.00 at 25.0 °C. Calculate the K_a of HZ.
 - a. 2.0×10^{-2}
 - b. 1.0×10^{-5}
 - c. 1.0×10^{-8}
 - d. 5.0×10^{-9}
 - e. 1.0×10^{-12}
8. What is the concentration of the hydroxide ion in pure water at 25°C?
 - a. 0 M
 - b. $1 \times 10^{-7} \text{ M}$
 - c. $1 \times 10^{-14} \text{ M}$
 - d. 14 M
 - e. 7.00 M
9. In a basic solution, _____.
 - a. $[\text{H}_3\text{O}^+] = [\text{OH}^-]$
 - b. $[\text{H}_3\text{O}^+] > [\text{OH}^-]$
 - c. $[\text{H}_3\text{O}^+] < [\text{OH}^-]$
 - d. $[\text{H}_3\text{O}^+] = 0 \text{ M}$
 - e. $[\text{OH}^-] > 7.00$
10. Which one of the following statements regarding K_w is false?
 - a. $\text{p}K_w$ is 14.00 at 25°C
 - b. The value of K_w is always 1.0×10^{-14}
 - c. K_w changes with temperature.
 - d. The value of K_w shows that water is a weak acid.
 - e. K_w is known as the ion product of water.
11. Of the following substances, which of the following will form basic aqueous solutions?
 NH_4Cl $\text{Cu}(\text{NO}_3)_2$ K_2CO_3 NaF
 - a. NH_4Cl , $\text{Cu}(\text{NO}_3)_2$,
 - b. NH_4Cl , K_2CO_3
 - c. NaF only
 - d. K_2CO_3 , NaF
 - e. K_2CO_3 , NaF , $\text{Cu}(\text{NO}_3)_2$
 - f. NH_4Cl only
 - g. None of the above will be basic, they are all just salt water with a pH of 7

12. Which of the following salts will have a pH of 7.0 for an aqueous 0.10 M solution at 25.0°C.

NaOCl KNO₂ NH₄Cl Ca(OCl)₂

- NaOCl
- KNO₂
- NH₄Cl
- Ca(OCl)₂
- KCl and NH₄Cl
- None of them would have a pH of 7, they all change the pH

13. What is the pOH of a 0.002 M solution of NaNO₂?

HNO₂, $K_a = 5.0 \times 10^{-4}$

- 3.15
 - 3.60
 - 6.70
 - 10.85
 - 13.40
14. A 0.1-molar solution of acetic acid (CH₃COOH) has a pH of about
- 1
 - 3
 - 7
 - 10
 - 14
15. K_a , the acid dissociation constant, for an acid is 9×10^{-4} at room temperature. At this temperature, what is the approximate percent dissociation of the acid in a 1.0 M solution?
- 0.03%
 - 0.09%
 - 3%
 - 5%
 - 9%
16. What is the ionization constant, K_a for a weak monoprotic acid if a 0.30-molar solution has a pH of 4.0?
- 9.7×10^{-10}
 - 4.7×10^{-2}
 - 1.7×10^{-6}
 - 3.0×10^{-4}
 - 3.3×10^{-8}
17. Phenol, C₆H₅OH, has a $K_a = 1.0 \times 10^{-10}$ What is the pH of a 0.010 M solution of phenol?
- 2
 - 5
 - 6
 - 10
 - 12

18. Determine the OH⁻ concentration in 1.0 M aniline (C₆H₅NH₂) solution. (K_b for aniline is 4.0×10^{-10} .)

- 2.0×10^{-5} M
- 4.0×10^{-10} M
- 3.0×10^{-6} M
- 5.0×10^{-7} M
- 1.0×10^0 M

19. $\text{IO}_3^- + \text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{HIO}_3 + \text{C}_2\text{H}_3\text{O}_2^-$

The above equation has an equilibrium constant that is less than 1. What are the relative strengths of the acids and bases?

	ACIDS	BASES
A.	$\text{HIO}_3 < \text{HC}_2\text{H}_3\text{O}_2$	$\text{IO}_3^- < \text{C}_2\text{H}_3\text{O}_2^-$
B.	$\text{HIO}_3 < \text{HC}_2\text{H}_3\text{O}_2$	$\text{IO}_3^- > \text{C}_2\text{H}_3\text{O}_2^-$
C.	$\text{HIO}_3 > \text{HC}_2\text{H}_3\text{O}_2$	$\text{IO}_3^- > \text{C}_2\text{H}_3\text{O}_2^-$
D.	$\text{HIO}_3 > \text{HC}_2\text{H}_3\text{O}_2$	$\text{IO}_3^- < \text{C}_2\text{H}_3\text{O}_2^-$
E.	$\text{HIO}_3 = \text{HC}_2\text{H}_3\text{O}_2$	$\text{IO}_3^- = \text{C}_2\text{H}_3\text{O}_2^-$

20. $\text{H}_2\text{C}_3\text{H}_2\text{O}_4 + 2\text{H}_2\text{O} \rightleftharpoons 2\text{H}_3\text{O}^+ + \text{C}_3\text{H}_2\text{O}_4^{2-}$

As shown above, malonic acid is a diprotic acid. The successive equilibrium constants are 1.5×10^{-3} (K_{a1}) and 2.0×10^{-6} (K_{a2}). What is the equilibrium constant for the above reaction?

- 1.0×10^{-14}
- 2.0×10^{-6}
- 4.0×10^{-12}
- 3.0×10^{-9}
- 1.5×10^{-3}

21. $\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HPO}_4^{2-}$

Which species, in the above equilibrium, behave as bases?

- HPO₄²⁻
- H₂PO₄⁻
- H₂O

- I only
- I and II
- II and III
- I and III
- III only

22. $\text{HC}_3\text{H}_5\text{O}_2 + \text{HCOO}^- \rightleftharpoons \text{HCOOH} + \text{C}_3\text{H}_5\text{O}_2^-$

The equilibrium constant, K , for the above equilibrium is 7.2×10^{-2} . This value implies which of the following?

- A solution with equimolar amounts of HC₃H₅O₂ and HCOO⁻ is neutral
- C₃H₅O₂⁻ is a stronger base than HCOO⁻
- HC₃H₅O₂ is a stronger acid than HCOOH
- HCOO⁻ is a stronger base than C₃H₅O₂⁻
- The value of the equilibrium does not depend on the temperature

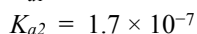
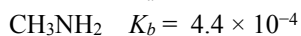
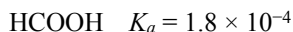
23. The calculation of $[H^+]$ concentration and pH for weak acids is more complex than for strong acids due to
- the incomplete ionization of weak acids.
 - the low K_a value for strong acids.
 - the more complex atomic structures of strong acids.
 - the low percent ionization of strong acids.
 - the inconsistent K_b value for strong acids.
24. The general reaction of an acid dissolving in water may be shown as
- $$HA + HOH \rightleftharpoons H_3O^+ + A^-$$
- One of the two conjugate acid base pair for this reaction is
- HA and HOH.
 - HA and A^-
 - HOH and A^-
 - H_3O^+ and A^-
 - HA and H_3O^+
25. Strong acids are those which
- have an equilibrium lying far to the left.
 - yield a really weak ("pathetic" in fact) conjugate base when reacting with water.
 - have a conjugate base which is a stronger base than water.
 - readily remove the H^+ ions from water.
 - are only slightly dissociated (ionized) at equilibrium.
26. When calculating the pH of a hydrofluoric acid solution ($K_a = 7.2 \times 10^{-4}$) from its concentration, the contribution of water ionizing ($K_w = 1.0 \times 10^{-14}$) is usually ignored because
- hydrofluoric acid is such a weak acid.
 - hydrofluoric acid can dissolve glass.
 - the ionization of water provides relatively few H^+ ions
 - the $[OH^-]$ for pure water is unknown.
 - the conjugate base of HF is such a strong base.
27. The percent dissociation (percent ionization) for weak acids
- is always the same for a given acid, no matter what the concentration.
 - usually increases as the acid becomes more concentrated.
 - compares the amount of acid that has dissociated at equilibrium with the initial concentration of the acid.
 - may only be used to express the dissociation of weak acids.
 - has no meaning for polyprotic acids.
28. The $[OH^-]$ of a certain aqueous solution is $1.0 \times 10^{-5} M$. The pH of this same solution must be
- 1.0×10^{-14}
 - 5.00
 - 7.00
 - 9.00
 - 12.00
29. In many calculations for the pH of a weak acid from the concentration of the acid, an assumption is made that often takes the form $[HA]_0 - x = [HA]_0$. This
- is valid because x is very small compared to the initial concentration of the weak acid.
 - is valid because the concentration of the acid changes by such large amounts.
 - is valid because the actual value of x cannot be known.
 - is valid because pH is not dependent upon the concentration of the weak acid.
 - approximation is always shown to be valid and so need not be checked.
30. HA is a weak acid which is 4.0% dissociated at 0.100M. Determine the K_a for this acid.
- 0.0040
 - 0.00016
 - 0.040
 - 1.6
 - 16.5
31. What is the pH of a 0.01-molar solution of NaOH?
- 1
 - 2
 - 8
 - 10
 - 12
32. What is the volume of 0.05-molar HCl that is required to neutralize 50 ml of a 0.10-molar $Mg(OH)_2$ solution?
- 100 ml
 - 200 ml
 - 300 ml
 - 400 ml
 - 500 ml
33. Which of the following best describes the pH of a 0.01-molar solution of $HBrO_2$ ($K_a = 2 \times 10^{-9}$)
- Less than or equal to 2
 - Between 2 and 7
 - 7
 - Between 7 and 11
 - Greater than or equal to 11

34. Which of the following species is amphoteric?
- H^+
 - CO_3^{2-}
 - HCO_3^-
 - H_2CO_3
 - H_2
35. How many liters of distilled water must be added to 1 liter of an aqueous solution of HCl with a pH of 1 to create a solution with a pH of 2?
- 0.1 L
 - 0.9 L
 - 2 L
 - 9 L
 - 10 L
36. A 1-molar solution of a very weak monoprotic acid has a pH of 5. What is the value of K_a for the acid?
- $K_a = 1 \times 10^{-10}$
 - $K_a = 1 \times 10^{-7}$
 - $K_a = 1 \times 10^{-5}$
 - $K_a = 1 \times 10^{-2}$
 - $K_a = 1 \times 10^{-1}$
37. The value of K_a for HSO_4^- is 1×10^{-2} . What is the value of K_b for SO_4^{2-} ?
- $K_b = 1 \times 10^{-12}$
 - $K_b = 1 \times 10^{-8}$
 - $K_b = 1 \times 10^{-2}$
 - $K_b = 1 \times 10^2$
 - $K_b = 1 \times 10^5$
38. How much 0.1-molar NaOH solution must be added to 100 milliliters of a 0.2-molar H_2SO_3 solution to neutralize all of the hydrogen ions in H_2SO_3 ?
- 100 ml
 - 200 ml
 - 300 ml
 - 400 ml
 - 500 ml
39. After adding the required amount of 0.01-molar NaOH to 100 milliliters of a 0.02-molar H_2SO_3 solution to neutralize, what would be the pH of the resulting solution?
- 2 or below
 - between 3 and 6
 - 7
 - between 8 and 10
 - above 10
40. The concentrations of which of the following species will be increased when HCl is added to a solution of $\text{HC}_2\text{H}_3\text{O}_2$ in water?
- H^+
 - $\text{C}_2\text{H}_3\text{O}_2^-$
 - $\text{HC}_2\text{H}_3\text{O}_2$
- I only
 - I and II only
 - I and III only
 - II and III only
 - I, II, and III
41. Which of the following species is amphoteric?
- HNO_3
 - $\text{HC}_2\text{H}_3\text{O}_2$
 - HSO_4^-
 - H_3PO_4
 - ClO_4^-
42. If 0.630 grams of HNO_3 (molecular weight 63.0) are placed in 1 liter of distilled water at 25°C , what will be the pH of the solution? (Assume that the volume of the solution is unchanged by the addition of the HNO_3 .)
- 0.01
 - 0.1
 - 1
 - 2
 - 3
43. Which of the following is the strongest acid?
- H_2SO_4
 - HSO_4^-
 - H_2SO_3
 - HSO_3^-
 - H_2S
44. The first acid dissociation constant for tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$, is 1.0×10^{-3} . What is the base dissociation constant, K_b , for $\text{HC}_4\text{H}_4\text{O}_6^-$?
- 1.0×10^{-13}
 - 1.0×10^{-11}
 - 1.0×10^{-7}
 - 1.0×10^{-4}
 - 1.0×10^{-1}
45. Which of the following expressions is approximately equal to the hydrogen ion concentration of a 1-molar solution of a very weak monoprotic acid, HA with an ionization constant K_a ?
- K_a
 - K_a^2
 - $2K_a$
 - $2K_a^2$
 - $(K_a)^{1/2}$

The next 8 questions refer to the following descriptions of chemical solutions.

- (A) a solution with $\text{pH} = 7$
- (B) a solution with a $\text{pH} < 7$ which is not a buffer
- (C) a solution with a $\text{pH} < 7$ which is a buffer
- (D) a solution with a $\text{pH} > 7$ which is not a buffer
- (E) a solution with a $\text{pH} > 7$ which is a buffer

Ionization constants



(Contrary to the appearance of its formula, phosphorous acid is diprotic, not triprotic.)

- 46. A solution with an initial KCOOH concentration of 1 M and an initial K_2HPO_3 concentration of 1 M. Equal volumes of both solutions mixed together.
- 47. A solution with an initial H_3PO_3 concentration of 1 M and an initial KH_2PO_3 concentration of 1 M. Equal volumes of both solutions mixed together.
- 48. A solution with an initial CH_3NH_2 concentration of 1 M and an initial $\text{CH}_3\text{NH}_3\text{Cl}$ concentration of 1 M. Equal volumes of both solutions mixed together.
- 49. A solution made of equal volumes of 0.5 M CH_3NH_2 and 0.25 M HCl
- 50. A solution made of 10 ml of 0.1 M H_3PO_3 and 20 ml of 0.1 M NaOH .
- 51. A solution made of 20 ml of 0.1 M HCl and 10 ml of 0.1 M NaOH
- 52. A solution made with 0.10 M NaCl
- 53. A solution made with equal volumes of 0.5 M KOH and HNO_3
- 54. A solution made with equal volumes of 1 M HCOOH and 1 M Na_3PO_3

The next 4 questions refer to the following descriptions of equal quantities of each aqueous solution mixed together in which each of the two species listed have concentrations of 1 M.

- (A) $\text{H}_2\text{C}_2\text{O}_4$, oxalic acid and KHC_2O_4 , potassium hydrogen oxalate
- (B) KNO_3 , potassium nitrate and HNO_3 , nitric acid
- (C) NH_3 , ammonia and NH_4NO_3 ammonium nitrate
- (D) $\text{C}_2\text{H}_5\text{NH}_2$, ethylamine and KOH , potassium hydroxide
- (E) HCl , hydrochloric acid and KOH , potassium hydroxide

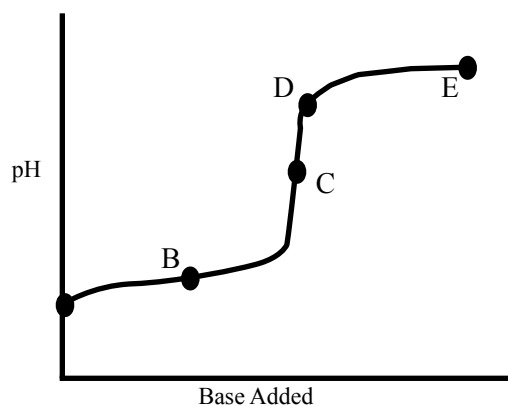
55. The most acidic solution.

56. The solution with pH closest to 7

57. A buffer with $\text{pH} > 7$

58. A buffer with a $\text{pH} < 7$

The diagram below shows the titration of a weak monoprotic acid with a strong base.



- 59. At this point in the titration the pH of the solution is equal to the pK_a of the acid.
- 60. This is the equivalence point of the titration.
- 61. Of the points shown on the graph, this is the point at which the most excess base has been added
- 62. At this point the solution is a balanced buffer.

63. The K_a of hypochlorous acid (HClO) is 3.0×10^{-8} at 25.0°C . Calculate the pH of a 0.030 M hypochlorous acid solution.
- 8.30
 - 7.54
 - 5.30
 - 4.52
 - 3.00
64. The K_b of hydroxylamine, HONH_2 is 1.0×10^{-8} at 25.0°C . What is the pH of 100 ml of 0.050 M aqueous solution of hydroxylamine, to which 0.35 g of hydroxylamide chloride, HONH_3Cl has been added? Assume no volume change to the solution.
- 4
 - 6
 - 7
 - 8
 - 10
65. A buffer that has ten times as many moles of lactic acid as moles of sodium lactate has a pH of 5.0, what is the K_a for lactic acid?
- 1×10^{-4}
 - 5×10^{-4}
 - 1×10^{-5}
 - 2×10^{-5}
 - 1×10^{-6}
66. Calculate the pH of solution made by combining 100.0 ml of $0.28\text{ M HC}_2\text{H}_3\text{O}_2$ and 50.0 ml of 0.36 M NaOH
- K_a for $\text{HC}_2\text{H}_3\text{O}_2$ 1.8×10^{-5}
- 3.00
 - 3.68
 - 4.74
 - 5.00
 - 6.00
67. **Acid** **K_a**
- | | |
|---------------------------|-----------------------|
| H_3PO_4 | 7.2×10^{-3} |
| H_2PO_4^- | 6.3×10^{-8} |
| HPO_4^{2-} | 4.2×10^{-13} |
- Using the information above, choose the best answer for preparing a buffer with $\text{pH} = 7$
- $\text{K}_2\text{HPO}_4 + \text{KH}_2\text{PO}_4$
 - H_3PO_4
 - $\text{K}_2\text{HPO}_4 + \text{K}_3\text{PO}_4$
 - K_3PO_4
 - $\text{K}_2\text{HPO}_4 + \text{H}_3\text{PO}_4$
68. A solution of a weak base is titrated with a solution of a standard strong acid. The progress of the titration is followed with a pH meter. Which of the following observations would occur?
- The pH of the solution gradually decreases throughout the experiment.
 - Initially the pH drops slowly, and then it drops much more rapidly.
 - At the equivalence point the pH is 7
 - After the equivalence point, the pH becomes constant because this is the buffer region.
 - The pOH at the equivalence point equals the $\text{p}K_a$ of the base.
69. You are given equimolar solutions of each of the following. Which has the lowest pH?
- NH_4Cl
 - NaCl
 - K_3PO_4
 - Na_2CO_3
 - KNO_3
70. When sodium nitrite is dissolved in water,
- The solution is acidic because of hydrolysis of the sodium ion.
 - The solution is neutral
 - The solution is basic because of hydrolysis of the sodium ion.
 - The solution is acidic because of hydrolysis of the NO_2^- ion.
 - The solution is basic because of hydrolysis of the NO_2^- ion.
71. Which of the solutions below would have a pH above 7
- NH_4NO_3
 - AlCl_3
 - KClO_4
 - K_2SO_3
 - HCl
72. The addition of nitric acid would increase the solubility of which of the following solid compounds?
- KCl
 - $\text{Pb}(\text{CN})_2$
 - $\text{Cu}(\text{NO}_3)_2$
 - NH_4NO_3
 - FeSO_4

73. Which statement below is true about the soluble salts listed below?
- KNO_3 forms a basic solution
 - NaCl forms an acidic solution
 - KClO forms a neutral solution
 - NH_4NO_3 forms a basic solution
 - Na_2CO_3 forms a basic solution
74. The equilibrium expression for the hydrolysis $\text{Na}_2\text{C}_2\text{O}_4$ of is best represented by which of the following?
- $$K = \frac{[\text{OH}^-][\text{C}_2\text{O}_4^{2-}]}{[\text{HC}_2\text{O}_4^-]}$$
 - $$K = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{O}_4^{2-}]}{[\text{HC}_2\text{O}_4^-]}$$
 - $$K = \frac{[\text{OH}^-][\text{HC}_2\text{O}_4^-]}{[\text{C}_2\text{O}_4^{2-}]}$$
 - $$K = \frac{[\text{C}_2\text{O}_4^{2-}]}{[\text{HC}_2\text{O}_4^-][\text{OH}^-]}$$
 - $$K = \frac{[\text{C}_2\text{O}_4^{2-}]}{[\text{HC}_2\text{O}_4^-][\text{H}_3\text{O}^+]}$$
75. Which of the following procedures will produce a buffered solution?
- Equal volume of 0.5 M NaOH and 1 M HCl solutions are mixed.
 - Equal volumes of 0.5 M NaOH and 1 M $\text{HC}_2\text{H}_3\text{O}_2$ solutions are mixed.
 - Equal volumes of 1 M $\text{NaC}_2\text{H}_3\text{O}_2$ and 1 M $\text{HC}_2\text{H}_3\text{O}_2$ solutions are mixed.
- I only
 - III only
 - I and II only
 - II and III only
 - I, II, and III
76. Which of the solutions below would have a pH = 7
- NH_4Cl
 - AlBr_3
 - KClO_4
 - K_2SO_3
 - HCl
77. Which of substances can act as a Lewis acid?
- Na^+
 - Fe^{3+}
 - Cl^-
- I only
 - II only
 - I and II only
 - II and III only
 - I, II, and III
78. Which statement is true for the reaction between boron trifluoride and ammonia?
- BF_3 is a Lewis acid.
 - BF_3 is an Arrhenius base
 - NH_3 is an Arrhenius base.
 - NH_3 is a Bronsted-Lowry base
 - NH_3 accepts a pair of electrons from BF_3
79. Which of the following acids can be oxidized to form a stronger acid?
- $\text{H}_2\text{C}_2\text{O}_4$
 - HNO_2
 - H_2SO_4
 - H_3PO_4
 - $\text{H}_2\text{C}_2\text{H}_3\text{O}_2$
80. During the preparation of a solution of nitric acid, a student accidentally spills about few milliliters of 12 M HNO_3 on the bench top. The student finds three bottles containing liquids sitting near the spill: a bottle of distilled water, a bottle of 5 percent $\text{NaHCO}_3(\text{aq})$, and a bottle of saturated $\text{NaCl}(\text{aq})$. Which of the liquids is best to use in cleaning up the spill?
- 16 M HNO_3
 - 5 % NaHCO_3 solution
 - saturated solution of NaCl
 - 5 M NaOH
 - distilled water
81. A 0.5-molar solution of which of the following salts will have the lowest pH?
- KCl
 - $\text{Cu}(\text{NO}_3)_2$
 - NaI
 - KNO_3
 - $\text{NaC}_2\text{H}_3\text{O}_2$

If you are still struggling for ways to make the “easy” math easy, please come in and we can discuss strategies that will work for you.

1. b The lead and aluminum nitrate salts produce LOTS of nitrate anions in solution. The substance that dissolves to give the least number of nitrites in solution would at first appear to be both a, and b since they have only one nitrate compared to c and d, however, all of a ionizes, but only some of b ionizes because it is a WA.
2. c To get a conjugate base, you must react the HSO_4^- as an acid, and the remaining ion will be its conjugate base.
 $\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$ Thus the SO_4^{2-} is the conjugate base.
3. a If pH 2 is $-\log(1 \times 10^{-2})$ and pH 3 is $-\log(1 \times 10^{-3})$ then it would stand to reason that the pH of some concentration between 1×10^{-2} and 1×10^{-3} would give a pH in between 2 and 3. Since $[\text{H}^+] = 0.0025$ is between 1×10^{-2} and 1×10^{-3} , the only appropriate pH option would be choice a
4. c First calc the pOH ($14 - 5 = 9$) Then “undo” the pOH of 9 to get c. (solve: $10^{-9} = [\text{OH}^-]$)
5. c First calc the pH ($14 - 4.282 = 9.718$) Since we know that pH = 9 when $[\text{H}^+] = 1 \times 10^{-9}$ and we know that pH = 10 when $[\text{H}^+] = 1 \times 10^{-10}$ so we know that the actual $[\text{H}^+]$ must be less than 1×10^{-9} and greater than 1×10^{-10} thus the only option even close is c (If you had a calculator, which you don’t, it would be an easy solve: $10^{-9.718} = [\text{H}^+]$)
6. e Since $\text{Ba}(\text{OH})_2$ dissolves to produce 2 OH^- ions in solution, $[\text{OH}^-]$ is twice the molarity given, 0.00010 which gives a pOH of 4, thus a pH of 10.
7. d “Undo” the pH ($10^{-5.00}$) to get the $[\text{H}^+] = 1.0 \times 10^{-5}$. The $[\text{H}^+] = [\text{Z}^-]$ because they are related to each other in 1:1 ratio. The equilibrium $[\text{HZ}]$ is $0.020 - x$ (which is too small to worry about from a sig fig point of view). Then solve the K_a expression. $K_a = \frac{[\text{H}^+][\text{Z}^-]}{[\text{HZ}]}$ so $K_a = \frac{[1.0 \times 10^{-5}][1.0 \times 10^{-5}]}{[0.020]}$
8. b Since $[\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$ in room temperature pure water, we know that $[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7}$
9. c In a basic solution, $[\text{OH}^-]$ is greater than in water, while $[\text{H}^+]$ is less than in water.
10. b K_w like all other K_{eq} ’s are temperature dependent and thus K_w is not always 1.0×10^{-14}
11. d Salts are basic (pH above 7) when the negative anion is the conjugate base of a weak acid. Thus in K_2CO_3 , the CO_3^{2-} is the weak conjugate base of H_2CO_3 and in NaF , the F^- is the weak conjugate base of HF . Further, NH_4Cl contains a conjugate weak acid, NH_4^+ and $\text{Cu}(\text{NO}_3)_2$ contains the (Lewis) acid forming Cu^{2+} ion which hydrolyses one of the waters that solvates around it when in solution.
12. f Salts are neutral only if they contain the “pathetic” ions, i.e. anions from strong acids (Cl^- , Br^- , I^- , NO_3^- , SO_4^{2-} , ClO_4^- , ClO_3^-) and cations from strong bases (group I and II metal ions), thus none of the salts presented meet those criteria.
13. c It is important to notice that this is a salt that is pH and thus pOH changing. The anion, NO_2^- is the conjugate weak base from the WA HNO_2 . Thus this is an “x²” problem of a WB. First you must determine the K_b . Since $K_a \times K_b = K_w$,
 $K_b = \frac{[1.0 \times 10^{-14}]}{[5.0 \times 10^{-4}]}$ then $K_b = 2 \times 10^{-11}$ then solve $2.0 \times 10^{-11} = \frac{[x^2]}{[0.002]}$ to get 2.0×10^{-7} will give a pOH in the 6-ish range,
 option c
14. b Even without being given a K_a , you would know that it must be acidic, thus pH = 1 or pH = 3 are the only reasonable options. But a 0.1 M SA would be $[\text{H}^+] = 0.1$ for a pH of 1, and a WA must be less than this, thus b is the only reasonable option.
15. c This is an “x²” problem for a WA, to calculate $[\text{H}^+]$ and then convert to percent dissociation which is
 $\%dissociation = \frac{[\text{H}^+]}{[\text{HA}]}$ So solve $9 \times 10^{-4} = \frac{[x^2]}{[1]}$ to get $[\text{H}^+] = 0.03$ M which would be $\%dissoc = \frac{[0.03]}{[1]}$ equal to 3 %
16. e First you must “unlog” to solve for $[\text{H}^+] = 1 \times 10^{-4}$, then plug into the K_a expression and solve $K_a = \frac{[1.0 \times 10^{-4}]^2}{[0.30]}$
17. c this is an “x²” problem, when solving for $x = [\text{H}^+]$ from which you can determine the pH. $1.0 \times 10^{-10} = \frac{[x^2]}{[0.01]}$ $x = 1 \times 10^{-6}$
 which solves for a pH of 6.00
18. a this also is an “x²” problem, when solving for $x = [\text{OH}^-]$. $4.0 \times 10^{-10} = \frac{[x^2]}{[1.0]}$ $x = 2 \times 10^{-5}$ M
19. d Since the K for the reaction is less than 1, we know that the equilibrium is “more to the left.” This tells us that the acid and base on the left are more likely to be weaker than the acid and base on the right. You can tell which are acids and bases by remembering that an acid donates H^+ and bases accept H^+ .
20. d Since this is a diprotic acid, when the complete ionization occurs, the dissociation constant for the total ionization would be the product of each of the two dissociations. For $\text{H}_2\text{A} \rightleftharpoons \text{H}^+ + \text{HA}^-$ K_{a1} and for $\text{HA}^- \rightleftharpoons \text{H}^+ + \text{A}^{2-}$ K_{a2} , thus the K for the overall equation would be K_a
21. d Remember, the bases are the proton acceptors.

22. b Since the K given is smaller than 1, the equilibrium position lies more on the reactant than product side, thus the product base is stronger, and the reactant base is weaker. In addition, you can also conclude that the product acid is stronger, and the reactant acid is weaker, however, that was not an option.
23. a $[H^+]$ concentration of SA is simple because the complete ionization means the molarity of the acid will equal the molarity of the $[H^+]$ but the incomplete ionization of acids means that these calculations require the use of the K_a and the equilibrium expression.
24. b The other conjugate acid base pair would be HOH and H_3O^+
25. b The conjugate of a SA is "pathetic" because it has is not able to hydrolyze with water and change pH in any way.
26. c Compared to the quantity of H^+ ions contributed by HF, although small itself, is far more than the amount of H^+ ions contributed by water.
27. c Percent is always part out of total, $\%dissociation = \frac{[H^+]}{[HA]}$
28. d As you know if $[OH^-] = 1 \times 10^{-5}$, then $pOH = 5$ and $pH = 9$ since $pH + pOH = 14$
29. a In most cases, and certainly any that AP will present, the $-x$ is always small enough to be considered insignificant compared to the original molarity.
30. b Since $\%dissociation = \frac{[H^+]}{[HA]}$, 4% or 0.04 of 0.1 M will be $[H^+] = 0.004$. Use this in the equilibrium expression

$$K_a = \frac{[H^+][A^-]}{[HA]} \text{ and } K_a = \frac{[0.004]^2}{[0.10]} \text{ then solve.}$$
31. e For SB, the $[OH^-]$ is the same as the molarity of the base since it completely dissociates, thus $-\log [OH^-]$, pOH would be 2 and $pH = 12$
32. b You can use the dilution equation for this. $M_a V_a = M_b V_b$ and do not forget that the base produces twice as much $[OH^-]$, thus $(0.05 M)(V_a) = 2(0.1 M)(50 \text{ ml})$ and solve.
33. b This is essentially the same question as #14. A 0.01 M SA would yield a pH of 2, since hypobromous acid is weak, it will have a less acidic pH, but still lower than 7.
34. c Amphoteric means the ability to react both as an acid or a base. Any diprotic weak acid that has lost one of its protons has the ability to lose a second, and act as an acid, or accept the first proton back, and act as a base.
35. d pH of 1 means $[H^+] = 0.1$ and a pH of 2 means $[H^+] = 0.01$, thus we can work the dilution equation $M_1 V_1 = M_2 V_2$ so, $(0.1 M)(1 \text{ L}) = (0.01 M)(V)$ solve for V to get 10 L. Since there is already 1 L, you only need to add 9 L of water to bring the total volume to 10 L
36. a pH of 5 means $[H^+] = 1 \times 10^{-5}$, and since $HA \rightleftharpoons H^+ + A^-$, the $[H^+] = [A^-]$ and since $K_a = \frac{[H^+][A^-]}{[HA]}$ substitute

$$K_a = \frac{[1 \times 10^{-5}][1 \times 10^{-5}]}{[1]} \text{ and solve for } K_a = 1 \times 10^{-10}$$
37. a Remember that SO_4^{2-} is the conjugate base of the not so weak acid HSO_4^- (we usually consider this to be a strong acid....)
 and since $K_a \times K_b = K_w$, $K_b = \frac{[1.0 \times 10^{-14}]}{[1 \times 10^{-2}]}$ then solve.
38. d It is important to remember that sulfurous acid is a diprotic acid. Using $M_a V_a = M_b V_b$ substitute $2(0.2 M)(100 \text{ ml}) = (0.1 M)V_b$ and solve for 400 ml.
39. d At the equivalence point, there is water and salt in the solution, however the soluble salt, Na_2SO_3 is a weak base. Specifically the sulfite ion is the conjugate weak base of the weak sulfurous acid.
40. c Consider the equilibrium reaction $HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$ and use LeChatelier's principle to realize that adding HCl, a source of H^+ will cause the equilibrium position to shift to the left, but it will never be able to shift far enough to reduce the incoming $[H^+]$ to a value below what it started at.
41. c This question is a repeat of #34 but with a different diprotic acid.
42. d This is a strong acid. Thus 0.63 g/63 g/mol yields 0.01 mole/L of ions, this gives a pH of 2
43. a c, d, and e are weak acids, and HSO_4^- is the second dissociation of H_2SO_4 , thus HSO_4^- is a weaker acid
44. b Since $HC_4H_4O_6^-$ is the conjugate base of $H_2C_4H_4O_6$ and $K_a \times K_b = K_w$, $K_b = \frac{[1.0 \times 10^{-14}]}{[1 \times 10^{-3}]}$ then solve
45. e $HA \rightleftharpoons H^+ + A^-$, the $[H^+] = [A^-]$ and since $K_a = \frac{[H^+][A^-]}{[HA]}$ solving for $[H^+]$ when $[HA] = 1$, is simply a square root.

In the next 8 questions it is important to identify the two substances being combined. You need to think about whether they are SA, SB, WA, WB. Remember to be on alert for salts that contain an ion that is a conjugate WB or conjugate WA. If you are combining a strong with a weak, be on alert for where you are in that combination: equivalence point, beyond equivalence point, or part way in the titration which is to say buffer.

46. d KCOOH is a salt in which COOH^- is a conjugate weak base (the K^+ is a non-pH changing “pathetic”), the other particle, KH_2PO_3 is a salt that contains H_2PO_3^- , the conjugate weak base of the H_3PO_3 acid (the K^+ is a non-pH changing “pathetic”). So, two weak bases do not make a buffer and the pH would certainly be above 7.
47. c H_3PO_3 is a weak acid, and as stated in the previous problem, KH_2PO_3 contains H_2PO_3^- , the conjugate weak base of the H_3PO_3 acid. Thus with an acid and its conjugate base in solution, you have a buffer with an acidic pH (You know that it is acidic because the K_a , 3×10^{-2} of the acid is larger than the K_b , 3.3×10^{-13} (since $K_a \times K_b = K_w$) of the conjugate weak base.)
48. e CH_3NH_2 is a weak base and $\text{CH}_3\text{NH}_3\text{Cl}$ is a salt which contains the weak base's conjugate weak acid CH_3NH_3^+ (the Cl^- is a non-pH changing “pathetic”). So the weak base and its conjugate make a buffer with a pH above 7. (You know that it is basic because the K_b , 4.4×10^{-4} of the weak base is larger than the K_a , 2.3×10^{-11} (since $K_a \times K_b = K_w$) of the conjugate weak acid.)
49. e This solution has a weak base, CH_3NH_2 and a strong acid, HCl , thus you need to be on alert for the possible buffer formation. We can determine this by looking at the quantities of each substance. Equal volumes mean that the concentration is determining the quantities, and you can see that we have twice as much WB as the incoming SA. This means that you are at the halfway point of a titration, which means buffer. The pH will be above 7 because once again you know that it is basic because the K_b , 4.4×10^{-4} of the weak base is larger than the K_a , 2.3×10^{-11} (since $K_a \times K_b = K_w$) of the conjugate weak acid.
50. d This solution is a di-protic weak acid, H_3PO_3 and a strong base, NaOH , thus you need to be on alert for the possible buffer formation. We can determine this by looking at the quantities of each substance. Since we have a concentration that is twice the acid, both of the H^+ 's get ripped off and we are at the equivalence point which is certainly not a buffer and the pH would be above 7 because of the conjugate base that is in the water along with the pathetic Na^+ ions from the SB.
51. b This solution contains a strong acid, HCl , and a strong base NaOH which can never make a buffer. Thus you simply need to consider which substance is in excess. The concentrations are the same, but the HCl is present in double the volume. This means that you still have half of the strong acid remaining in solution causing a pH below 7.
52. a NaCl is a salt made of two “pathetic” ions, thus it is not a pH changing solution, and it is certainly not a buffer.
53. a This solution is made of a strong acid and a strong base. Equal volumes and equal concentrations mean that we have added equal quantities, thus we are at the equivalence point, which will not be a buffer and will have a pH = 7.
54. b This solution may appear to be a buffer because it contains a weak acid, HCOOH and a weak base, Na_2HPO_3 however, for a buffer to be a buffer, the WA and WB (HPO_3^{2-}) must be conjugates of each other, which these two substances are not.
55. b Solution b contains a strong acid, HNO_3 , and a non-pH changing salt, KNO_3 . This can not be a buffer because there is no weak acid or weak base. The KNO_3 has no effect on the strong acid, this means the pH will be very acidic.
56. e When equal quantities of strong acid and strong base are combined, the pH will be 7 since you would be at the equivalence point.
57. c Remember that to have a buffer with a pH above 7, you need a weak base and its conjugate weak acid. NH_3 is a weak base and NH_4NO_3 is a salt which contains the conjugate weak acid, NH_4^+ . This is a buffer with a pH above 7. (You can assume this because you know that NH_3 is a weak base, and if you looked up its K_b it would be greater than the K_a of the conjugate weak acid.)
58. a For a buffer below 7 you need a weak acid and its conjugate weak base. Oxalic acid is the weak acid, and the K is a salt which contains its conjugate weak base, HC_2O_4^- .
59. b Halfway to the equivalence point, in a titration, the WA = WB and thus the pH will = the $[\text{H}^+]$.
60. c The equivalence point on a graph is the middle of the steep vertical rise.
61. e As more base is added beyond the equivalence point, the excess base will prevail.
62. b the term “balanced” buffer implies in the halfway-ish area of the titration.
63. d This is an “x²” problem. Thus $3.0 \times 10^{-8} = \frac{x^2}{[0.03\text{M}]}$ Look for the easy math. $x = [\text{H}^+] = 3 \times 10^{-5}$,
This indicates a pH of 4-ish
64. b This is a buffer made of a weak base and its conjugate weak acid. We can use the acid buffer formula $[\text{H}^+] = K_a \frac{[\text{WA}]}{[\text{WB}]}$, but we will need the K_a of the conjugate weak acid. We can calculate that since $K_a \times K_b = K_w$, $K_a = 1 \times 10^{-6}$. Next we need to calculate the quantities of both the WB and conjugate WA. $(0.05\text{ M})(0.10\text{ L}) = 0.005\text{ moles WB}$ and $35\text{ g} \times \frac{1\text{ mol}}{69\text{ g}} = 0.005\text{ mol}$ of the chloride salt which contains the conjugate weak acid. Then insert into the buffer equation formula and of course

halfway is very easy math. $[H^+] = 1 \times 10^{-6} \frac{[0.005]}{[0.005]}$ $[H^+] = 1 \times 10^{-6}$, thus pH = 6

65. e This obviously is a buffer problem, thus you can use the buffer formula. $[H^+] = K_a \frac{[WA]}{[WB]}$. Since we know the pH = 5, we know that $[H^+] = 1 \times 10^{-5}$, and we can use the 10:1 ratio of WA to conjugate WB. $[1 \times 10^{-5}] = K_a \frac{[10]}{[1]}$ then *look for the easy math*
66. d This problem is a weak acid that is being titrated with strong base. It may be a buffer problem. You will know as soon as you calculate the quantities of each. WA (0.028 M)(100 ml) = 28 mmoles, the calculate incoming invader (0.36 M)(50 ml) = 18 mmoles, thus we have not made it to the equivalence point. $[H^+] = K_a \frac{[WA]}{[WB]}$ The amount of incoming invader will shift the equilibrium: $[H^+] = 1.8 \times 10^{-5} \frac{[28 - 18]}{[-0 + 18]}$ solve $[H^+] = 1.8 \times 10^{-5} \frac{[10]}{[18]}$ and *look for the easy math*.
67. a When buliding a buffer, the most balanced, or ideal buffers have approximately equal quantities of WA and WB. When this is true, $[H^+] = K_a$, and pH of the buffer = pK_a. Thus look for a weak acid with a pK_a near the desired pH. Thus the second K_a of this acid, H₂PO₄⁻ has a pK_a in the 7-ish range. So we need a salt of H₂PO₄⁻ wich we can find as , and we need the conjugate WB of this acid.
68. b Review your titration cuves for this answer.
69. a For this problem you need to look for a molecular acid, or look for a salt that contains a conjugate weak acid ion. NH₄Cl will satisfy that as NH₄⁺ is a conjugate weak acid (of NH₃). Cl⁻ is of course “pathetic”. “b and e” are neutral, and “c and d” contain conjugate weak base ions.
70. e Nitrite is a conjugate weak base of (nitrous acid), thus in water this reaction occurs: $NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$
71. d To get a pH over 7, you need a conjugate weak base, thus you need to look for a salt which contains an anion is *not* an anion of a strong acid. Thus any anion that is *not* Cl⁻, Br⁻, I⁻, NO₃⁻, SO₄²⁻, ClO₃⁻, and ClO₄⁻, will be an anion that is a weak base. Thus SO₃²⁻ is the conjugate weak base of sulfurous acid, H₂SO₃, which is a weak acid. Note that NH₄⁺ is a conjugate weak acid of NH₃, thus “a” would be acidic. “b” would be acidic because metal ions in solution (not groups 1 or 2) hydrolyze with the water (as a Lewis acid) to generate H⁺ ions, “c” has pH = 7, because K⁺ and are “pathetic” and “e” is of course simply acidic.
72. b In this problem nitrate ion will have no effect on any of the salts except Pb(CN)₂. Remember this is a K_{sp} situation. Consider this equilibrium: $Pb(CN)_2 \rightleftharpoons Pb^{2+} + CN^-$ and when nitric acid is added, the H⁺ will react with the CN⁻ base and cause the Pb(CN)₂ to dissolve more and shift right to generate more CN⁻ to make up for what is being removed as HCN
73. e The fastest way to solve this problem is to cross out all the “pathetics” and look for what ions are left and decide if the are weak acid ions or weak base ions. KNO₃ and NaCl would be neutral, KClO would be basic because of the ClO⁻ ion, and NH₄NO₃ would be acidic because of the NH₄⁺ ion. Na₂CO₃ is basic because of the CO₃²⁻ ion is a conjugate weak base of carbonic acid.
74. c Consider the hydrolysis of C₂O₄²⁻ ion: $C_2O_4^{2-} + H_2O \rightleftharpoons HC_2O_4^- + OH^-$ then write the equilibrium expression.
75. d Remember that there are two ways to produce a buffer. By putting both a WA and a salt that contains its conjugate WB in solution (conversely, WB and a salt that contains its conjugate WA) OR by combining a WA and half the moles of SB which would produce a solution that has been “pushed” to a “halfway” titration (conversely, combining a WB and half the moles of SA which would produce a solution that has been “pushed” to a “halfway” titration). “I” is not a buffer because it is a SA SB, and that combo never produces a buffer. “II” is a buffer because of the second reason stated above, and “III” is a buffer because of the first reason stated above.
76. c Only c contains all “pathetics”. Items a, b, e would be acidic and item d would be basic.
77. b negative ions are never Lewis acids, and group I and II cations do not hydrolyze with water and thus are not Lewis acids.
78. a BF₃ is a molecule with a incomplete octet around the B atom, thus when it reacts with NH₃ a synthesis reaction occurs that has been defined as a Lewis acid base reaction because the BF₃ is accepting the N’s unshared pair of electrons. The reaction is: $BF_3 + NH_3 \rightarrow BF_3NH_3$
79. b When HNO₂ is converted to HNO₃, a strong acid – the oxidation number of N changes from 3+ to 5+, oxidation
80. b NaHCO₃, baking soda, in the presence of acid is a weak base. will be dissolved into Na⁺ and HCO₃⁻. The bicarbonate ion is a weak base and will
81. b Like problem #27, the fastest way to solve this problem is to cross out all the “pathetics” and look for what ions are left and decide if the are weak acid ions or weak base ions. KCl, NaI, and KNO₃ contain only “pathetics”. NaC₂H₃O₂ is basic because it contains one pathetic ion, Na⁺ and C₂H₃O₂⁻ which is the conjugate WB of acetic acid. Thus the only acidic forming salt is Cu(NO₃)₂ which contains the pathetic nitrate ion but also the Cu²⁺ which hydrolyzes with the water to generate H⁺ ions. This Cu²⁺ ion is a Lewis acid.