

Chemistry 12
Acid - Base #2, Pretest

- Which of the following species is amphiprotic?
 - CN^-
 - HCl
 - HCO_3^-
 - H_2SO_4
- A student tests a 0.10 M solution with different pH indicators and finds that:
 - alizarin yellow remains yellow in the solution. $\text{pH} < 10.1$
 - thymol blue remains blue in the solution. $\text{pH} > 9.6$
 As a result of these observations it is correct to say the solution is a
 - weak base
 - weak acid
 - strong base $\rightarrow \text{pH} = 13$
 - strong acid $\rightarrow \text{pH} = 1$
- What is the pH of a 1.0 M AlCl_3 solution?
 - 11.57
 - 7.00
 - 3.40
 - 2.43
$$K_a = 1.4 \times 10^{-5}$$
- In which of the following equations is water acting as a Brønsted-Lowry base?
 - $\text{NaOH} + \text{H}_2\text{O} \rightleftharpoons \text{Na}^+ + \text{OH}^- + \text{H}_2\text{O}$
 - $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$
 - $\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$
 - $\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{HPO}_4^{2-} + \text{H}_3\text{O}^+$
- The indicator HIn has a yellow acid form and a red base form. The equation for its ionization is:

$$\text{HIn} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{In}^-$$
 When KOH is added to the above system, the equilibrium
 - shifts left and the solution turns red
 - shifts left and the solution turns yellow
 - shifts right and the solution turns red
 - shifts right and the solution turns yellow
- Which of the following substances is a salt that produces an acidic aqueous solution?
 - KNO_3
 - Na_2CO_3
 - NH_4NO_3
 - NaOH
- Consider the following data:

Solution A has a $\text{pH} = 2.0$
Solution B has a $\text{pH} = 5.0$

 By what factor is the $[\text{H}_3\text{O}^+]$ in solution B different from that in solution A?
 - smaller by a factor of 3
 - greater by a factor of 3
 - smaller by a factor of 10^3
 - greater by a factor of 10^3

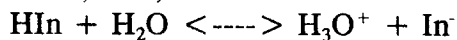
8. Consider the following equilibrium expression:

$$K_{eq} = \frac{[OH^-][H_2PO_4^-]}{[HPO_4^{2-}]}$$

The above K_{eq} expression represents the

- A. K_a expression for the dihydrogen phosphate ion
B. K_b expression for the dihydrogen phosphate ion
C. K_a expression for the monohydrogen phosphate ion
D. K_b expression for the monohydrogen phosphate ion
9. The value of K_b for $Fe(H_2O)_5(OH)^{2+}$ is
A. 1.7×10^{-14}
B. 1.7×10^{-12} $\frac{1 \times 10^{-14}}{6.0 \times 10^{-3}}$
C. 6.0×10^{-3}
D. 1.7×10^2
10. Which of the following is a salt that dissolves in water to form an acidic solution?
A. NH_4CH_3COO
B. $CrCl_3$
C. K_3PO_4
D. $Ca(C_6H_5COO)_2$
11. During a titration, a 25.00 mL sample of 0.40 M NaOH requires 20.00 mL of $H_2C_2O_4$ solution to reach the equivalence (stoichiometric) point. What is the concentration of the acid?
A. 0.80 M
B. 0.50 M
C. 0.40 M
D. 0.25 M
12. Consider the equilibrium: $HX^- + HB^- \rightleftharpoons H_2B + X^{2-}$
If the products are favoured, what is the formula of the strongest base?
A. HB^-
B. HX^-
C. H_2B
D. X^{2-}
13. A vinegar solution is reported to be 6.0% pure. If a 10.0 mL sample of the vinegar solution (density 1.0 g/mL) is titrated against 0.500 M NaOH to the equivalence point, what volume of base will be required?
A. 40 mL
B. 20 mL
C. 12 mL
D. 2.0 mL
14. Which of the following sets of salts is ordered according to increasing pH?
A. $CaCO_3$, CsF, KNO_2 , $NaClO_4$
B. $NaClO_4$, CsF, KNO_2 , $CaCO_3$
C. $CaCO_3$, KNO_2 , CsF, $NaClO_4$
D. $NaClO_4$, KNO_2 , CsF, $CaCO_3$
15. Which equation shows the acid form of an indicator HIn reacting in a basic solution?
A. $In^- + H_3O^+ \rightleftharpoons HIn + H_2O$
B. $In^- + OH^- \rightleftharpoons HIn + O^{2-}$
C. $HIn + H_2O \rightleftharpoons In^- + H_3O^+$
D. $HIn + OH^- \rightleftharpoons In^- + H_2O$

16. An indicator undergoes transition from colourless to blue at pH 10.0. What is the K_a of the indicator?
- A. 10^{-14}
 B. 10^{-10}
 C. 10^{-4}
 D. 10^{10}
17. The following indicators were added to separate samples of a solution, giving the following results:
- | | | |
|------------------|---|------------|
| bromthymol blue | = | green |
| phenolphthalein | = | colourless |
| methyl red | = | yellow |
| chlorophenol red | = | red |
| phenol red | = | yellow |
- The pH of the solution is approximately
- A. 8.2
 B. 6.8
 C. 6.1
 D. 5.4
18. Which of the following indicators is the strongest acid?
- A. orange IV
 B. alizarin yellow
 C. indigo carmine
 D. methyl violet
19. The indicator "macho pink" has a K_a of 6.7×10^{-7} . The indicator has a pink acid form and a blue base form. What is the result of adding macho pink indicator to a solution of pH 9?
- A. the solution is pink
 B. the solution is blue
 C. the [acid form] > [base form]
 D. the transition colour purple form
20. When the salt NH_4HSO_4 is dissolved in water
- A. one ion only hydrolyzes and $\text{pH} < 7$
 B. both ions hydrolyze and $\text{pH} > 7$
 C. both ions hydrolyze and $\text{pH} = 7$
 D. both ions hydrolyze and $\text{pH} < 7$
21. Which of the following substances is a salt that produces a basic solution?
- A. KBr
 B. NH_3
 C. NH_4Cl
 D. Na_3PO_4
22. An indicator, HIn, is found to establish the following equilibrium:



When the indicator was added to solutions of various pH values, the following data were recorded:

pH 6	pH 8	pH 10	pH 12
yellow	yellow	green	blue

At pH 10:

- A. $[\text{H}_3\text{O}^+] = K_a$
 B. $[\text{H}_3\text{O}^+] = \text{pH}$
 C. $[\text{H}_3\text{O}^+] = [\text{In}^-]$
 D. $[\text{H}_3\text{O}^+] = [\text{HIn}]$

23. Which of the following will have the greatest $[H_3O^+]$?
- A. 0.1 M NaF
 B. 0.1 M NaIO₃
 C. 0.1 M Na₂SO₃
 D. 0.1 M NaCH₃COO
24. The approximate K_a value of thymolphthalein indicator is
- A. 1×10^{-10}
 B. 1×10^{-4}
 C. 4
 D. 10
25. The value of the K_b for $HC_6H_5O_7^{2-}$ is
- A. 5.9×10^{-10}
 B. 2.4×10^{-8}
 C. 4.1×10^{-7}
 D. 1.7×10^{-5}

Open Ended. Calculations need to be done to the correct significant figure for full marks.

1. Calculate the pH of

- (a) the strong acid, 0.25 M HBr

$$[H_3O^+] = 0.25 M$$

$$pH = -\log(0.25) = 0.60$$

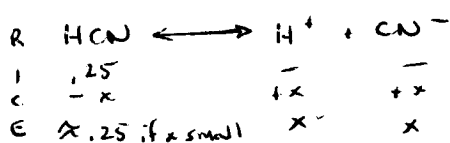
- (b) the strong base, 0.25 M KOH

$$[OH^-] = 0.25 M$$

$$pOH = -\log(0.25) = 0.60$$

$$pH = 14.00 - 0.60 = 13.40$$

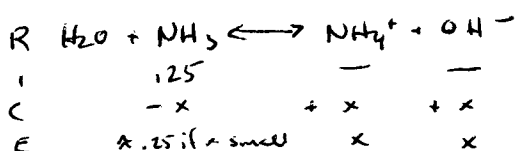
- (c) the weak acid, 0.25 M HCN



$$K_a = \frac{[H^+][CN^-]}{[HCN]} = \frac{x^2}{.25} = 4.9 \times 10^{-10}$$

$$x = [H^+] = 1.11 \times 10^{-5} M \quad pH = 4.96$$

- (d) the weak base, 0.25 M NH₃

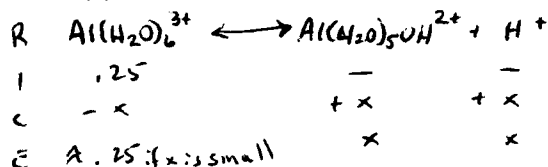


$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.79 \times 10^{-5}$$

$$= \frac{x^2}{.25} \quad x = [OH^-] = 2.11 \times 10^{-3} M$$

$$pOH = 2.675 \quad pH = 11.33$$

- (e) the acidic salt, 0.25 M Al(NO₃)₃

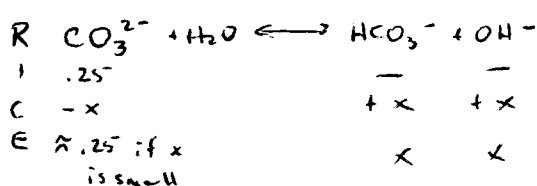


NO_3^- doesn't hydrolyze.

$$K_a = \frac{[Al(H_2O)_5OH^{2+}][H^+]}{[Al(H_2O)_6^{3+}]} = 1.4 \times 10^{-5} = \frac{x^2}{.25}$$

$$x = [H_3O^+] = 1.87 \times 10^{-3} M \quad pH = 2.73$$

- (f) the basic salt, 0.25 M K₂CO₃ K^+ doesn't hydrolyze.



$$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-11}} = 1.79 \times 10^{-4} = \frac{x^2}{.25}$$

$$[OH^-] = 6.68 \times 10^{-3} M \quad pOH = 2.18 \quad pH = 11.83$$

2. Determine by calculation the relative acidity of

(a) NaHC₂O₄ $K_a = 6.4 \times 10^{-5}$
 $K_b = \frac{1.0 \times 10^{-14}}{5.7 \times 10^{-2}} = 1.7 \times 10^{-13}$ $K_a > K_b$ acidic

(b) NH₄H₂PO₄ $K_a = 5.6 \times 10^{-10}$
 $K_a = 6.2 \times 10^{-8}$
 $K_b = \frac{1.0 \times 10^{-14}}{7.5 \times 10^{-3}} = 1.3 \times 10^{-12}$ $K_a > K_b$ acidic

Chemistry 12 Self Test #1

1. Identify the acid-base reactions

- | | | |
|-----|---|--------------------|
| (a) | $\text{H}_2\text{SO}_3 + \text{HCO}_3^- \rightarrow \text{HSO}_3^- + \text{H}_2\text{CO}_3$ | Acid/Base |
| (b) | $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow 2\text{NaCl} + \text{BaSO}_4$ | Double replacement |
| (c) | $2\text{HCl} + \text{Zn} \rightarrow \text{ZnCl}_2 + \text{H}_2$ | Single replacement |
| (d) | $\text{KHSO}_3 + \text{Na}_3\text{PO}_4 \rightarrow \text{K}^+ + 3\text{Na}^+ + \text{SO}_3^{2-} + \text{HPO}_4^{2-}$ | Acid/Base |
| (e) | $\text{Cu} + 2\text{Ag}^+ \rightarrow 2\text{Ag} + \text{Cu}^{2+}$ | Single replacement |
| (f) | $\text{HNO}_2 + \text{NH}_3 \rightarrow \text{NO}_2^- + \text{NH}_4^+$ | Acid/Base |
| (g) | $2\text{KI} + \text{Pb}(\text{NO}_3)_2 \rightarrow 2\text{KNO}_3 + \text{PbI}_2$ | Double replacement |
| (h) | $\text{H}_2\text{CO}_3 + \text{K}_2\text{S} \rightarrow \text{KHCO}_3 + \text{KHS}$ | Acid/Base |
| (i) | $2\text{C} + 3\text{H}_2 \rightarrow \text{C}_2\text{H}_6$ | Synthesis |
| (j) | $\text{H}_2\text{O} + \text{NH}_2^- \rightarrow \text{NH}_3 + \text{OH}^-$ | Acid/Base |
| (k) | $\text{Cl}_2 + 2\text{NaI} \rightarrow 2\text{NaCl} + \text{I}_2$ | Single replacement |
| (l) | $\text{NaClO}_4 + \text{HI} \rightarrow \text{NaI} + \text{HClO}_4$ | Acid/Base |

2. Write the conjugate acid of:

- | | | |
|-----|--------------------------|----------------------------|
| (a) | NO_2^- | HNO_2 |
| (b) | HCO_3^- | H_2CO_3 |
| (c) | HPO_4^{2-} | H_2PO_4^- |
| (d) | CH_3NH_2 | CH_3NH_3^+ |

3. Write the conjugate base of:

- | | | |
|-----|--------------------------------|----------------------------|
| (a) | HF | F^- |
| (b) | HCO_3^- | CO_3^{2-} |
| (c) | NH_3 | NH_2^- |
| (d) | N_2H_5^+ | N_2H_4 |
| (e) | HPO_4^{2-} | PO_4^{3-} |
| (f) | $(\text{CH}_3)_2\text{NH}_2^+$ | $(\text{CH}_3)_2\text{NH}$ |

4. Identify the two acids and the two bases.

- | | | |
|-----|---|--|
| (a) | $2\text{HBr} + \text{S}^{2-} \rightleftharpoons \text{H}_2\text{S} + \text{Br}_2$ | $\text{A} + \text{B} \rightleftharpoons \text{A} + \text{B}$ |
| (b) | $\text{CO}_3^{2-} + \text{HNO}_2 \rightleftharpoons \text{NO}_2^- + \text{HCO}_3^-$ | $\text{B} + \text{A} \rightleftharpoons \text{B} + \text{A}$ |
| (c) | $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ | $\text{B} + \text{A} \rightleftharpoons \text{A} + \text{B}$ |

5. Write the Bronsted-Lowry acid base equilibria which occur when the following pairs of substances are mixed in solution. Identify the conjugate pairs formed.

- (a) HNO_2 and $\text{NH}_3 \rightleftharpoons \text{NO}_2^- + \text{NH}_4^+$ $\text{HNO}_2 / \text{NO}_2^-$ and $\text{NH}_3 / \text{NH}_4^+$
 (b) CO_3^{2-} and $\text{HF} \rightleftharpoons \text{HCO}_3^- + \text{F}^-$ $\text{CO}_3^{2-} / \text{HCO}_3^-$ and HF / F^-
 (c) HCO_3^- and $\text{S}^{2-} \rightleftharpoons \text{CO}_3^{2-} + \text{HS}^-$ $\text{HCO}_3^- / \text{CO}_3^{2-}$ and $\text{S}^{2-} / \text{HS}^-$
 (d) H^+ and $\text{H}_2\text{O} \rightleftharpoons \text{H}_2 + \text{OH}^-$ H^+ / H_2 and $\text{H}_2\text{O} / \text{OH}^-$
 (e) H_2S and $\text{HO}_2^- \rightleftharpoons \text{HS}^- + \text{H}_2\text{O}_2$ $\text{H}_2\text{S} / \text{HS}^-$ and $\text{HO}_2^- / \text{H}_2\text{O}_2$
 (f) O^{2-} and $\text{H}_2\text{O} \rightleftharpoons 2\text{OH}^-$ $\text{O}^{2-} / \text{OH}^-$ and $\text{H}_2\text{O} / \text{OH}^-$
 (g) H_2O and $\text{H}_2\text{SO}_3 \rightleftharpoons \text{H}_3\text{O}^+ + \text{HSO}_3^-$ $\text{H}_2\text{O} / \text{H}_3\text{O}^+$ and $\text{H}_2\text{SO}_3 / \text{HSO}_3^-$

6. Classify each of the following as strong acid, weak acid, strong base, weak base, salt.

- (a) NaCl salt
 (b) KOH strong base
 (c) H_2O weak acid and weak base
 (d) CH_3COOH weak acid
 (e) H_2SO_4 strong acid
 (f) NH_3 weak base
 (g) KI salt
 (h) CaCl_2 salt

7. List the following acids in decreasing order of strength.

CH_3COOH , HNO_2 , HClO_4 and HCO_3^- HClO_4 , HNO_2 , CH_3COOH , HCO_3^-

8. List the conjugate bases of the following acids in decreasing order of strength.

CH_3COOH , HNO_2 , HClO_4 and HCO_3^- CO_3^{2-} , CH_3COO^- , NO_2^- , ClO_4^-

9. Complete the following acid-base equilibria after identifying the stronger acid.

- (a) H_2O_2 and HSO_3^- $\text{H}_2\text{O}_2 + \text{HSO}_3^- \rightleftharpoons \text{H}_3\text{O}_2^+ + \text{SO}_3^{2-}$
 (b) H_2PO_4^- and HCO_3^- $\text{H}_2\text{PO}_4^- + \text{HCO}_3^- \rightleftharpoons \text{HPO}_4^{2-} + \text{H}_2\text{CO}_3$

10. Complete the following acid-base equilibria after identifying the stronger base.

- (a) HCO_3^- and HS^- $\text{HCO}_3^- + \text{HS}^- \rightleftharpoons \text{H}_2\text{S} + \text{CO}_3^{2-}$
 (b) OH^- and NH_3 $\text{OH}^- + \text{NH}_3 \rightleftharpoons \text{H}_2\text{O} + \text{NH}_2^-$
 (c) HPO_4^{2-} and HS^- $\text{HPO}_4^{2-} + \text{HS}^- \rightleftharpoons \text{H}_2\text{PO}_4^- + \text{S}^{2-}$
 (d) HS^- and HSO_3^- $\text{HS}^- + \text{HSO}_3^- \rightleftharpoons \text{H}_2\text{S} + \text{SO}_3^{2-}$

11. Classify the following statements as true or false.

- (a) A strong electrolyte is more completely ionized than a weak electrolyte. TRUE
 (b) In solutions of equal concentration, a weak electrolyte will have a lower electrical conductivity than a strong electrolyte. TRUE
 (c) The stronger an acid, the stronger its conjugate base. FALSE

12. Explain why HF is a stronger acid than H_2O but a weaker acid than HCl.

Ans: The dissociation of HF lies between that of HCl (100%) and that of H_2O . Since strength is related to dissociation, then the strength of HF is between that of HCl (strong) and that of H_2O (weak).

13. For the following, state whether reactants or products are favored.
- | | | |
|-----|---|-----------|
| (a) | $\text{HPO}_4^{2-} + \text{HCO}_3^- \rightleftharpoons \text{H}_2\text{PO}_4^- + \text{CO}_3^{2-}$ | Reactants |
| (b) | $\text{HCl} + \text{CH}_3\text{COO}^- \rightleftharpoons \text{CH}_3\text{COOH} + \text{Cl}^-$ | Products |
| (c) | $\text{NH}_4^+ + \text{F}^- \rightleftharpoons \text{HF} + \text{NH}_3$ | Reactants |
| (d) | $\text{CH}_3\text{COOH} + \text{CO}_3^{2-} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{HCO}_3^-$ | Products |
| (e) | $\text{HCN} + \text{F}^- \rightleftharpoons \text{HF} + \text{CN}^-$ | Reactants |
| (f) | $\text{HCO}_3^- + \text{OH}^- \rightleftharpoons \text{CO}_3^{2-} + \text{H}_2\text{O}$ | Products |
| (g) | $\text{HF} + \text{SO}_4^{2-} \rightleftharpoons \text{F}^- + \text{HSO}_4^-$ | Reactants |
| (h) | $\text{H}_2\text{CO}_3 + \text{F}^- \rightleftharpoons \text{HCO}_3^- + \text{HF}$ | Reactants |
| (i) | $\text{HSO}_4^- + \text{HCO}_3^- \rightleftharpoons \text{SO}_4^{2-} + \text{H}_2\text{CO}_3$ | Products |
14. Write an equation to show the ionization of water. $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$
15. Write the mathematical definition for pH $-\log [\text{H}_3\text{O}^+]$
16. Calculate the $[\text{H}_3\text{O}^+]$ and the $[\text{OH}^-]$ in a solution in which:
- | | | |
|-----------------|---------------------------------|---------------------------------|
| (a) pH = 9.00 | $1.0 \times 10^{-9} \text{ M}$ | $1.0 \times 10^{-5} \text{ M}$ |
| (b) pH = 4.75 | $1.8 \times 10^{-5} \text{ M}$ | $5.6 \times 10^{-10} \text{ M}$ |
| (c) pH = -1.10 | 13M | $7.9 \times 10^{-16} \text{ M}$ |
| (d) pH = 0.00 | 1.0M | $1.0 \times 10^{-14} \text{ M}$ |
| (e) pH = 12.35 | $4.5 \times 10^{-13} \text{ M}$ | $2.2 \times 10^{-2} \text{ M}$ |
| (f) pH = 14.80 | $1.6 \times 10^{-15} \text{ M}$ | 6.3M |
| (g) pOH = 9.00 | $1.0 \times 10^{-5} \text{ M}$ | $1.0 \times 10^{-9} \text{ M}$ |
| (h) pOH = 4.75 | $5.6 \times 10^{-10} \text{ M}$ | $1.8 \times 10^{-5} \text{ M}$ |
| (i) pOH = -1.10 | $7.9 \times 10^{-16} \text{ M}$ | 13M |
| (j) pOH = 0.00 | $1.0 \times 10^{-14} \text{ M}$ | 1.0M |
| (k) pOH = 12.35 | $2.2 \times 10^{-2} \text{ M}$ | $4.5 \times 10^{-13} \text{ M}$ |
| (l) pOH = 14.80 | 6.3M | $1.6 \times 10^{-15} \text{ M}$ |
17. Calculate the $[\text{H}_3\text{O}^+]$, the $[\text{OH}^-]$, the pH and the pOH for
- | | | | | | |
|-----|--|---------------------------------|---------------------------------|-------|-----------------------|
| (a) | 0.0010 M HCl | $1.0 \times 10^{-3} \text{ M}$ | $1.0 \times 10^{-11} \text{ M}$ | 3.00 | 11.00 |
| (b) | 4.0 M NaOH | $2.5 \times 10^{-15} \text{ M}$ | 4.0M | 14.60 | -6.0×10^{-1} |
| (c) | $2.5 \times 10^{-3} \text{ M NaNH}_2$ | $4.0 \times 10^{-12} \text{ M}$ | $2.5 \times 10^{-3} \text{ M}$ | 11.40 | 2.60 |
| (d) | $6.0 \times 10^{-3} \text{ M Ca(OH)}_2$ | $8.3 \times 10^{-13} \text{ M}$ | $1.2 \times 10^{-2} \text{ M}$ | 12.08 | 1.92 |
| (e) | Saturated solution of Mg(OH)_2 ; $K_{sp} = 5.6 \times 10^{-12}$ | $4.5 \times 10^{-11} \text{ M}$ | $2.2 \times 10^{-4} \text{ M}$ | 10.35 | 3.65 |

Key.

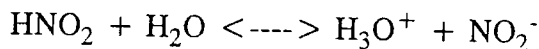
Chemistry 12 Acid-Base #1 Review

- At 30°C , $K_w = 3.0 \times 10^{-14}$. Therefore, a solution at 30°C in which the $[\text{OH}^-] = 1.732 \times 10^{-7} \text{ M}$ is best described as
 - acidic
 - basic
 - neutral
 - amphoteric
 - amphiprotic
- What is the pH of a neutral solution at 10°C when $k_w = 3.0 \times 10^{-15}$?
 - 3.0
 - 7.0
 - 7.3
 - 12
 - 14.5
- The equation: $\text{pH} + \text{pOH} = 14.00$ is true
 - for all solvents
 - at all temperatures
 - at 20°C
 - at 25°C
 - for all solvents and at all temperatures
- A 0.03 M borax solution has a pH of about 9.2. Which statement about this solution is true?
 - $[\text{H}^+] = 6.3 \times 10^{-10} \text{ M}$
 - $[\text{H}^+] = 1.6 \times 10^{-5} \text{ M}$
 - $[\text{H}^+] = 2.0 \times 10^{-2} \text{ M}$
 - $[\text{H}^+] = 9.2 \text{ M}$
 - the solution is basic
- The concentration of hydroxide ion in an aqueous solution is $3.3 \times 10^{-6} \text{ M}$. What is the pH of the solution?
 - 3.03×10^{-9}
 - 3.3
 - 5.48
 - 8.52
 - 5.48
- The pH of a soft drink is 5.67. What is the concentration of the OH^- ?
 - $4.7 \times 10^{-9} \text{ M}$
 - $2.1 \times 10^{-6} \text{ M}$
 - $6.7 \times 10^{-5} \text{ M}$
 - 5.67 M
 - 8.33 M

7. A solution was made by dissolving 0.0788 g $\text{Ca}(\text{OH})_2$, a strong base, in 100 mL water. The pH of this solution was
- A. 1.67
 - B. 1.97
 - C. 9.33
 - D. 12.03
 - E. 12.33

8. The pH of a solution of hydrochloric acid was found to be 2.55. What was the concentration of the acid?
- A. 3.6×10^{-12} M
 - B. 2.8×10^{-3} M
 - C. 5.5×10^{-2} M
 - D. 2.6 M
 - E. 11 M

9. Nitrous acid is a weak acid that ionizes in water according to the following equilibrium:



The expression for the acid ionization constant, K_a , is

- A. $K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{H}_2\text{O}][\text{HNO}_2]}$
 - B. $K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$
 - C. $K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$
 - D. $K_a = \frac{[\text{HNO}_2]}{[\text{H}^+][\text{NO}_2^-]}$
 - E. $K_a = \frac{[\text{H}_2\text{O}][\text{HNO}_2]}{[\text{H}_3\text{O}^+][\text{NO}_2^-]}$
10. The pH of a 0.10 M lactic acid solution at 25°C is 2.43. What is the value of K_a for this acid?
- A. 1.3×10^{-4}
 - B. 1.4×10^{-4}
 - C. 3.7×10^{-3}
 - D. 7.4×10^{-3}
 - E. 0.24

11. The K_a of glycollic acid at 25°C is 1.5×10^{-4} . Calculate the pH of a 0.14 M solution of this acid.
- A. 1.06
 - B. 1.49
 - C. 2.34
 - D. 2.96
 - E. 4.68
12. The K_b of strychnine at 25°C is 1.8×10^{-6} . What is the pH of a 0.15 M solution of this base?
- A. 3.28
 - B. 5.74
 - C. 8.26
 - D. 10.72
 - E. 11.54
13. At 25°C , the pH of a 0.12 M solution of morphine, a base, is 10.64. What is the value of the K_b for morphine?
- A. 4.6×10^{-21}
 - B. 2.3×10^{-8}
 - C. 1.9×10^{-7}
 - D. 1.6×10^{-6}
 - E. 1.5×10^{-4}
14. The cyanide ion, CN^- , is a Bronsted base. In a dilute solution of this ion together with the sodium ion, which of the following relationships is true where K_b refers to the base ionization constant for the cyanide ion?
- A. $[\text{OH}^-] = \frac{[\text{CN}^-]}{[\text{HCN}]} \times K_b$
 - B. $[\text{HCN}] = \frac{[\text{OH}^-]}{[\text{CN}^-]} \times K_b$
 - C. $K_b = \frac{[\text{H}_2\text{O}][\text{CN}^-]}{[\text{HCN}][\text{OH}^-]}$
 - D. $[\text{CN}^-] = \frac{[\text{HCN}]}{[\text{OH}^-]} \times K_b$
 - E. $K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{H}_2\text{O}][\text{CN}^-]}$
15. Calculate the % ionization of the acid in 0.075 M acetic acid solution at 25°C .
- A. 0.018 %
 - B. 0.075 %
 - C. 0.12 %
 - D. 1.6 %
 - E. 100 %

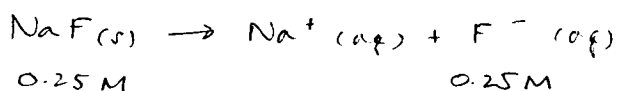
16. An acid that is a weaker acid than acetic acid might have a pK_a of
- A. 1.78×10^{-6}
 - B. 8.5×10^{-4}
 - C. 4.65
 - D. 4.85
 - E. -4.74

17. Considering the data in the following table, which species is the strongest base?

<u>compound</u>	<u>pK_a</u>
acetic acid	4.74
barbituric acid	4.01
tyrosine	8.40
phenol	9.89

- A. phenol
 - B. acetate ion
 - C. anion of barbituric acid
 - D. conjugate base of phenol
 - E. conjugate base of tyrosine
18. An acid with a K_a of 1.5×10^{-9} has a pK_a equal to
- A. 1.59
 - B. 5.18
 - C. 8.82
 - D. 9.15
 - E. 9.18
19. If an acid has a pK_a equal to 6.63, its K_a value is
- A. 4.3×10^{-8}
 - B. 2.3×10^{-7}
 - C. 5.8×10^{-7}
 - D. 6.3×10^{-6}
 - E. 7.37
20. The K_a of a monoprotic acid is 3.6×10^{-6} at 25°C . What is the pK_b of its conjugate base?
- A. 2.79×10^{-9}
 - B. 5.24
 - C. 5.44
 - D. 7.66
 - E. 8.56

Acid / Base Review Sheet (#3) p. 1 of 3



$K_a = 3.5 \times 10^{-4}$
 $K_b = 2.86 \times 10^{-11}$

$K_b = \frac{[HF][OH^-]}{[F^-]}$ Let $x = [OH^-] = [HF]$
 $[F^-] = 0.25 - x \approx 0.25$ if x is small
 $2.86 \times 10^{-11} = \frac{x^2}{0.25}$

$x = 2.67 \times 10^{-6} M = [OH^-]$

$pOH = 5.57$

$pH = 8.43$

2) 5.5% acetic acid 0.100 M NaOH

10.0 mL

$D = 1.05 g/mL$

$5.5\% = \frac{\text{mass of acid}}{\text{mass of sample}} = \frac{x}{10.5g}$

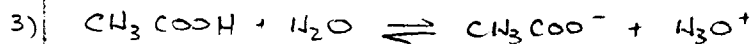
note:

$\text{mass of acid} = 0.5775g / \text{FW}$

$= 0.5775g / 60.0g/mol$

$= 9.625 \times 10^{-3} \text{ mol } CH_3COOH = \text{mol } H^+ = \text{mol } OH^- = \text{mol } NaOH$

$V = \frac{n}{M} = \frac{9.625 \times 10^{-3} \text{ mol } NaOH}{0.100M} = 9.6 \times 10^{-2} L$ or 96 mL



$pH = 4.50$

$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$

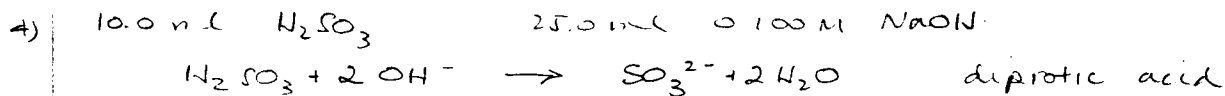
$[H_3O^+] = 3.16 \times 10^{-5} M$

$K_a = 1.8 \times 10^{-5}$

$1.8 \times 10^{-5} = \frac{[CH_3COO^-]}{[CH_3COOH]} \times 3.16 \times 10^{-5}$

$0.5696 = \frac{[CH_3COO^-]}{[CH_3COOH]} \therefore \frac{[CH_3COOH]}{[CH_3COO^-]} = 1.8$

Acid / Base Review Sheet (#3) p. 2 of 3

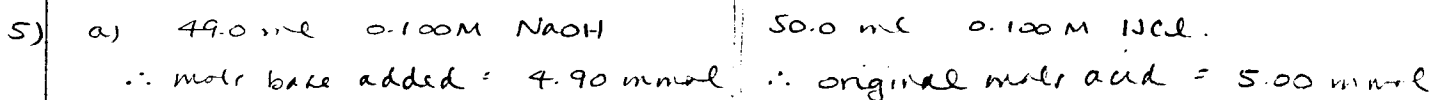


mol NaOH = $0.100 \text{ M} \times 25.0 \text{ mL} = 2.5 \text{ mmol NaOH} = \text{mol OH}^-$

mol $\text{OH}^- = \text{mol H}^+ = 2 (\text{mol acid})$

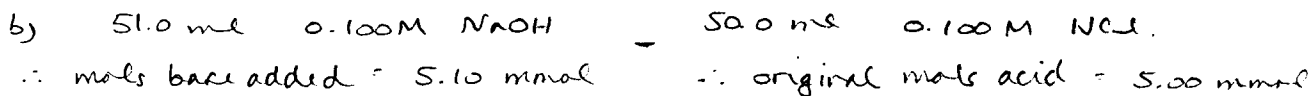
mol acid = $\frac{\text{mol H}^+}{2} = \frac{2.5 \text{ mmol}}{2} = 1.25 \text{ mmol}$

$[\text{H}_2\text{SO}_3] = \frac{1.25 \text{ mmol}}{10.0 \text{ mL}} = 0.125 \text{ M}$



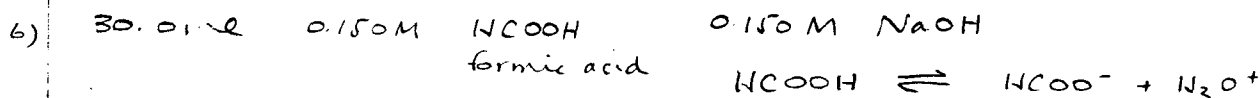
$[\text{H}^+] = \frac{5.00 \text{ mmol acid} - 4.90 \text{ mol base added}}{99.0 \text{ mL}} = 1.01 \times 10^{-3} \text{ M}$

pH = 2.996 (3 sf)



$[\text{OH}^-] = \frac{5.10 \text{ mmol base added} - 5.00 \text{ mmol acid}}{101.0 \text{ mL}} = 9.90 \times 10^{-4} \text{ M}$

pOH = 3.004 pH = 10.996



a) Let $x = [\text{H}_3\text{O}^+] = [\text{HCOO}^-]$ ($K_a = 1.8 \times 10^{-4}$)
 $[\text{HCOOH}] = 0.150 - x \approx 0.150 \text{ M}$ if x is small.

$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$

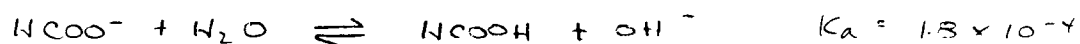
$1.8 \times 10^{-4} = \frac{x^2}{0.150}$ $x = 5.196 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+]$

pH = 2.284

Acid / Base Review sheet (#3) p. 3 of 3

b) continued

b) all HCOOH "used up" $\rightarrow \text{HCOO}^-$ ($\text{HCOOH} + \text{OH}^- \rightarrow \text{HCOO}^- + \text{H}_2\text{O}$)



$$K_b = \frac{[\text{HCOOH}][\text{OH}^-]}{[\text{HCOO}^-]} \quad K_b = 5.56 \times 10^{-11}$$

$$[\text{salt}] = \frac{4.50 \text{ mmol}}{60.0 \text{ mL}} = 0.0750 \text{ M}$$

$$\text{Let } x = [\text{OH}^-] = [\text{HCOOH}]$$

$$[\text{HCOO}^-] = 0.0750 \text{ M} - x \approx 0.0750 \text{ if } x \text{ is small.}$$

$$5.56 \times 10^{-11} = \frac{x^2}{0.0750}$$

$$x = 2.04 \times 10^{-6} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = 5.690$$

$$\text{pH} = 8.310$$

c) original moles acid $30.0 \text{ mL} \times 0.150 \text{ M} = 4.50 \text{ mmol}$
 moles base added $40.0 \text{ mL} \times 0.150 \text{ M} = 6.00 \text{ mmol}$

$$[\text{OH}^-] = \frac{6.00 \text{ mmol} - 4.50 \text{ mmol}}{70.0 \text{ mL}} = 0.0214 \text{ M}$$

$$\text{pOH} = 1.669$$

$$\text{pH} = 12.331$$

Acid Base III Worksheet No. 1

Key.

60.0 mL of 0.10 M HCl are titrated with 0.20 M NaOH solution. Calculate the pH and plot the titration curve for the process when the following volumes of NaOH are added.

Mol H ⁺	vol. NaOH added	Moles OH ⁻	mol excess	[conc]	pH
6.00 mmol	0.0 mL	0.00	6.00 mmol	0.10	1.00
	5.0 mL	1.00 mmol	5.00 mmol	0.077	1.11
	10.0 mL	2.00 mmol	4.00 mmol	0.057	1.24
	15.0 mL	3.00 mmol	3.00 mmol	0.040	1.40
	20.0 mL	4.00 mmol	2.00 mmol	0.025	1.60
	25.0 mL	5.00 mmol	1.00 mmol	0.012	1.93
	28.0 mL	5.60 mmol	0.40 mmol	4.5×10^{-3}	2.34
	29.0 mL	5.80 mmol	0.20 mmol	2.2×10^{-3}	2.65
	29.5 mL	5.90 mmol	0.10 mmol	1.1×10^{-3}	2.95
	30.0 mL	6.00 mmol	0.00 mmol	1.0×10^{-7} M	7.00
	30.5 mL	6.10 mmol	0.10 mmol	1.1×10^{-3}	11.04
	31.0 mL	6.20 mmol	0.20 mmol	2.2×10^{-3}	11.34
	35.0 mL	7.00 mmol	1.00 mmol	0.011	12.02
✓	60.0 mL	12.0 mmol	6.00 mmol	0.050	12.70

What is the pH at the equivalence point?

7.00

What is the volume of NaOH added at the midpoint

15.0 mL

What is the pH at the midpoint?

1.40

What is the K_a of the indicator you should use for this titration?

1.0×10^{-7}

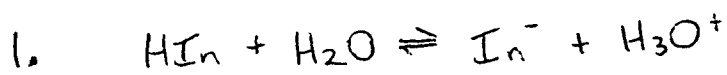
Which indicator should you choose?

bromthymol blue

or phenol red

or neutral red.

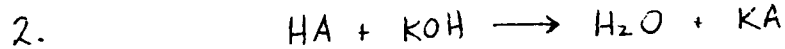
Acid Base III Worksheet 2



(a) yellow in acid solution - eq shifts left - HIn is predominant form

(b) blue in base solution - eq shifts right - In^- is predominant form

(c) in solutions with high $[\text{H}_3\text{O}^+]$ yellow HIn is in greater concentration. As $[\text{H}_3\text{O}^+]$ decreased solution reached a point where $[\text{HIn}] = [\text{In}^-]$ at this point the amount of yellow and blue is equal and green colour is observed. As $[\text{H}_3\text{O}^+]$ decreases further the blue In^- is the predominant form and the solution is blue.



(a) mol KOH = $18.4 \text{ mL} \times 0.100 \text{ M} = 1.84 \text{ mmol}$

ratio is 1:1 so mol HA = 1.84 mmol

$$[\text{HA}] = \frac{1.84 \text{ mmol}}{10.0 \text{ mL}} = 0.184 \text{ M}$$

(b) $0.220 \text{ g} \div 1.84 \text{ mmol} = 120 \text{ g/mole}$

(c) pH at stoichiometric endpoint is 9.600 \therefore acid is weak.

3. mol $\text{Ba}(\text{OH})_2 = 50.0 \text{ mL} \times 0.0200 \text{ M} = 1.00 \text{ mmol Ba}(\text{OH})_2$

mol $\text{OH}^- = 2 \times 1.00 \text{ mmol} = 2.00 \text{ mmol OH}^-$

mol $\text{H}^+ = 30.0 \text{ mL} \times 0.0100 \text{ M} = 0.300 \text{ mmol H}^+$

mol OH^- in excess = $2.00 - 0.300 = 1.70 \text{ mmol}$

$$[\text{OH}^-] = \frac{1.70 \text{ mmol}}{(50.0 + 30.0) \text{ mL}} = 0.0213 \text{ M} \quad \text{pOH} = 1.673 \quad \text{pH} = 12.327$$

4. $[\text{HCl}] = 10^{-2.63} = 2.34 \times 10^{-3} \text{ M}$

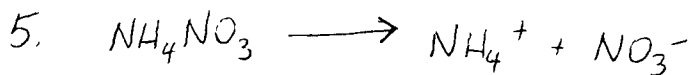
$$[\text{NaOH}] = 10^{-(14 - 10.82)} = 6.61 \times 10^{-4} \text{ M}$$

mol $\text{H}^+ = 2.34 \times 10^{-3} \text{ M} \times 30.0 \text{ mL} = 7.02 \times 10^{-2} \text{ mmol}$

mol $\text{OH}^- = 6.61 \times 10^{-4} \text{ M} \times 70.0 \text{ mL} = 4.63 \times 10^{-2} \text{ mmol}$

mol H^+ in excess = $7.02 \times 10^{-2} - 4.63 \times 10^{-2} = 2.39 \times 10^{-2} \text{ mmol}$

$$[\text{H}^+] = \frac{2.39 \times 10^{-2} \text{ mmol}}{(30.0 + 70.0) \text{ mL}} = 2.39 \times 10^{-4} \text{ M} \quad \text{pH} = 3.62$$



NO_3^- does not hydrolyze.

R	$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$
I	1.00 0 0
C	-x +x +x
E	$1.00 - x \approx 1.00$ if x is small

$$[\text{NH}_4^+]_{\text{initial}} = \frac{2.0\text{g} \div 80.0\text{g/mol}}{0.0250\text{L}} = 1.00\text{M}$$

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.6 \times 10^{-10} = \frac{x^2}{1.00}$$

$$x = [\text{H}_3\text{O}^+] = 2.37 \times 10^{-5}\text{M} \quad \text{pH} = -\log(2.37 \times 10^{-5}\text{M}) = 4.62$$

6.

R	$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$
I	2.0M 0 0
C	-x +x +x
E	$2.0\text{M} - x \approx 2.0\text{M}$ if x is small x x

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.79 \times 10^{-5}$$

$$K_b = \frac{x^2}{2.0} = 1.79 \times 10^{-5} \quad [\text{OH}^-] = x = 5.98 \times 10^{-3}\text{M}$$

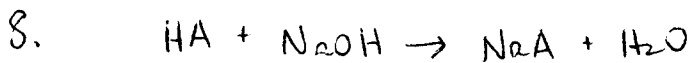
$$\text{pOH} = -\log(5.98 \times 10^{-3}) = 2.224$$

$$\text{pH} = \text{pK}_w - \text{pOH} = 14.00 - 2.224 = 11.78$$

7. Several answers apply:

eg. titration

pH at endpoint of titration



(a) $\text{mol OH}^- = 0.10\text{M} \times 20.0\text{mL} = 2.00\text{mmol}$

$\text{mol HA} = \text{mol OH}^- = 2.00\text{mmol}$

$$[\text{HA}] = \frac{2.00\text{mmol}}{10.0\text{mL}} = 0.20\text{M}$$

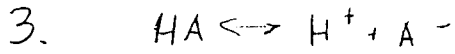
(b) $\text{pH at midpoint} = 5.67 = \text{pK}_a$

$$K_a = 10^{-5.67} = 2.1 \times 10^{-6}$$

Worksheet III

1. $[H_3O^+] = K_a \frac{[acid]}{[salt]}$ adding water decreases both acid and salt concentration
 buffer solution \uparrow weak acid
 so pH does not change appreciably.

2. $[H_3O^+] = 1.8 \times 10^{-5} \frac{[acid]}{[salt]}$ if $[acid] = [salt]$ then $[H_3O^+] = 1.8 \times 10^{-5}$
 pH = 4.74



I 0.100 M

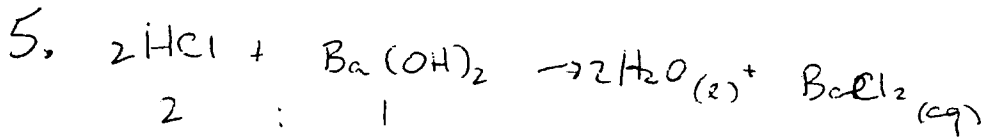
C -20% +20% +20%

E 0.080 M 0.020 0.020

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(0.020)^2}{0.080} = 5.0 \times 10^{-3}$$

4. methyl orange and. thymolphthalein

	3.0	6.0	12.0
	red	yellow	yellow
	colorless	colorless	blue } green.



100.0 mL x 0.20 M HCl \rightarrow 20.0 mmol H^+

300.0 mL x 0.050 M $Ba(OH)_2 \times 2 \rightarrow$ 30.0 mmol OH^-

10.0 mmol OH^- in excess $[OH^-] = \frac{10.0 \text{ mmol}}{100 + 300} = 0.025 \text{ M}$

pOH = 1.60

pH = 12.40

6. "Normal" rain water contains dissolved CO_2 .

$\text{CO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{CO}_3$ (which is acidic) $\leftrightarrow \text{H}^+ + \text{HCO}_3^-$
dissolved CO_2 in the air forms an acidic H_2CO_3 sol'n with rain water.

7. HCl and NH_3 (strong acid / weak base)

8. CH_3COOH and NaOH (weak acid / strong base)

9. NH_3

10. $\text{NaOH} + \text{HCl}$.

$$\text{mol H}^+ = 25.0 \text{ mL} \times 1.0 \text{ M} = 25.0 \text{ mmol} = \text{mol OH}^- = \text{mol NaOH}$$

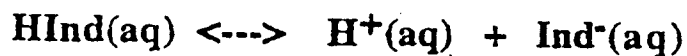
$$\text{grams NaOH} = 2.50 \times 10^{-2} \text{ mol} \times 40 \text{ g/mol} = 1.0 \text{ g}$$

$$\% \text{ completion} = \frac{1.0 \text{ g}}{10.0 \text{ g}} \times 100 = 10.0\%$$

Hydrolysis / Chemical Indicators: Practice Problems.

1. Which of the following substances is a salt that produces a basic aqueous solution?
- a. KBr
 - b. NH₃
 - c. NH₄Cl
 - d. Na₃PO₄

2. An indicator, HInd, is found to establish the following equilibrium:



When the indicator was added to solutions of various pH values, the following data were collected:

pH	2	4	6	8	10	12
colour	yellow	yellow	yellow	yellow	green	blue

At pH 10,

- a. $[\text{H}^+] = K_a$
 - b. $[\text{H}^+] = \text{pH}$
 - c. $[\text{H}^+] = [\text{Ind}^-]$
 - d. $[\text{H}^+] = [\text{HInd}]$
3. Which of the following will have the greatest $[\text{H}_3\text{O}^+]$?
- a. 0.10 M NaF
 - b. 0.10 M NaIO₃
 - c. 0.10 M Na₂SO₃
 - d. 0.10 M NaCH₃COO

4. A solution of unknown pH is tested with various indicators resulting in the following data:

Indicator	Colour
methyl violet	blue
bromcresol green	blue
methyl red	yellow
bromthymol blue	yellow
phenolphthalein	colourless

From the above data, the pH of the solution is approximately

- a. 1.6
b. 5.4
 c. 6.0
d. 8.2
5. Consider the following data:

Indicator	Color of Acid form	Color of Base form	K_a value
'A'	red	yellow	1.0×10^{-3}
'B'	yellow	blue	1.0×10^{-9}

- At a pH of 7.0, the predominant color of
- a. indicator 'A' and indicator 'B' will both be yellow.
b. indicator 'A' will be red and indicator 'B' will be yellow.
c. indicator 'A' will be yellow and indicator 'B' will be blue.
d. indicator 'A' will be orange and indicator 'B' will be green.
6. The approximate K_a value for the indicator thymolphthalein is
- a. 1×10^{-10}
b. 1×10^{-4}
c. 4
d. 10

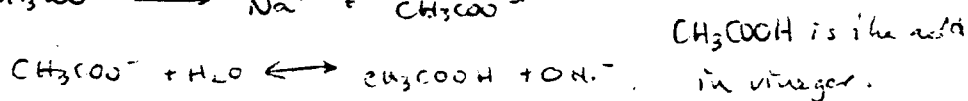
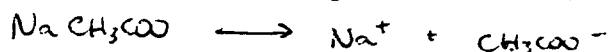
7. The value of K_b for HSO_3^- is ⁻¹³
- (a) ~~6.7×10^{-13}~~ 6.7×10^{-10}
- b. 6.2×10^{-8}
- c. 1.6×10^{-7}
- d. 1.7×10^{-2}

8. Two indicators were added to separate samples of a solution, giving the following results:

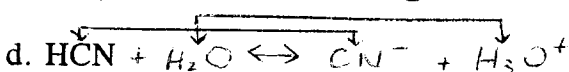
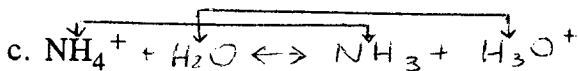
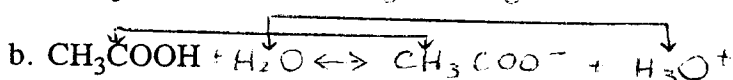
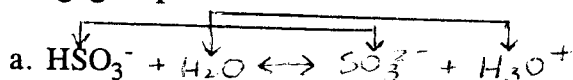
Indicator	Colour
Chlorophenol red	Red
Thymol blue	Yellow

The pH of the solution is approximately

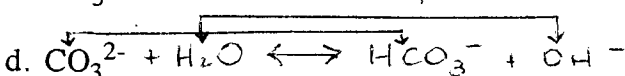
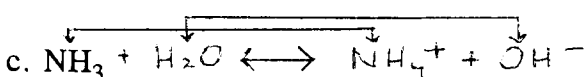
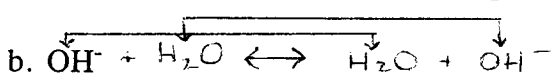
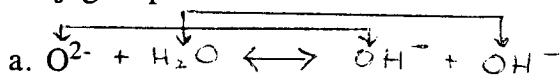
- a. 3
- b. 5
- (c) 7
- d. 10
9. Find the pOH and pH of the following solutions.
- a. 0.50 M KCN $\text{pOH} = 2.50$ $\text{pH} = 11.50$
- b. 1.0 M Na_2CO_3 $\text{pOH} = 1.87$ $\text{pH} = 12.13$
- c. 0.05 M $\text{NaC}_6\text{H}_5\text{COO}$ (sodium benzoate) $\text{pOH} = 5.56$ $\text{pH} = 8.44$
- d. 0.2 M AlCl_3 $\text{pOH} = 11.22$ $\text{pH} = 2.78$
10. When a 0.1 M solution of NH_3 is titrated against a 0.1 M solution of HCl , the endpoint is reached at pH 5.1. Which indicator would be best to use in the titration of NH_3 with HCl ?
methyl red or bromocresol green
11. Calculate the $[\text{OH}^-]$ midway through the colour change (or transition point) for the indicator indigo carmine.
 $[\text{OH}^-] = 0.016 \text{ M} = 0.02 \text{ M}$
12. When the amphiprotic anion, HPO_4^{2-} , is added to water, does it act as an acid or as a base? Support your answer with calculations.
 $K_a = 2.2 \times 10^{-13}$
 $K_b = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7}$ \therefore basic
13. A 0.60 M base solution, $\text{NaX}(\text{aq})$ is found to have a $[\text{OH}^-]$ of 0.12 M. Determine the K_b for the base.
 $K_b = \frac{[\text{HX}][\text{OH}^-]}{[\text{X}^-]} = \frac{(0.12)(0.12)}{(0.60 - 0.12)} = 3.0 \times 10^{-2}$
14. Discuss, in terms of hydrolysis, the use of sodium acetate to produce the vinegar flavour on some potato chips.



1. Write the equations for the hydrolysis of the following acids in water and identify the conjugate pairs.



2. Write the equations for the hydrolysis of the following bases in water and identify the conjugate pairs.



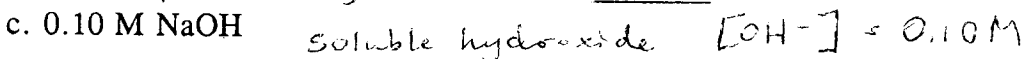
3. Calculate the pH of the following solutions:



$$\text{pH} = -\log(0.0020) = \underline{2.70}$$



$$\text{pH} = -\log(0.10) = \underline{1.00}$$

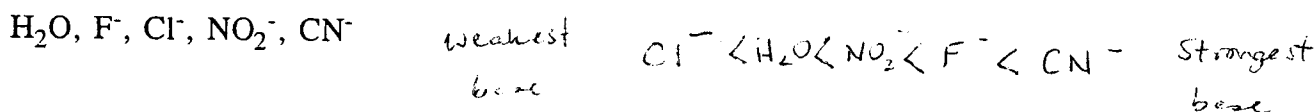


$$[\text{H}_3\text{O}^+] = K_w / [\text{OH}^-] = 1.0 \times 10^{-14} \text{ M} \quad \text{pH} = -\log(1.0 \times 10^{-13}) = \underline{13.00}$$

4. If a solution of acidified water has a pH of 3.00 what is its $[\text{H}_3\text{O}^+]$?

$$[\text{H}_3\text{O}^+] = \text{antilog}(-3.00) = 1.0 \times 10^{-3} \text{ M}$$

5. Arrange the following species according to their strength as bases:

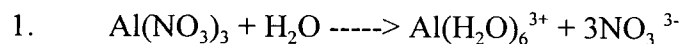


6. At 60°C the value of K_w is 1.0×10^{-13} . Calculate $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ at 60°C.

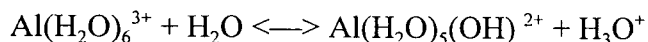
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \quad [\text{H}_3\text{O}^+] = \sqrt{K_w} = 3.2 \times 10^{-7} \text{ M}$$

in water $[\text{H}_3\text{O}^+] = [\text{OH}^-] \quad [\text{OH}^-] = [\text{H}_3\text{O}^+] = 3.2 \times 10^{-7} \text{ M}$

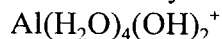
Acid # 10



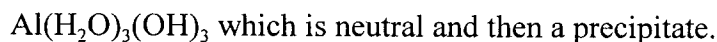
Then:



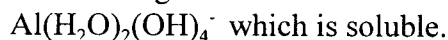
which can hydrolyze further into:



and further into:



Now adding additional OH^- causes further hydrolyzation into:



2. The Colour at the transition point equals that point where K_a of the indicator is the $[\text{H}_3\text{O}^+]$ in the solution.

$$\text{Then: } K_a = [\text{H}_3\text{O}^+] = 5.2 \times 10^{-4}$$

$$\text{The pH} = \text{p}K_a = 3.28$$

3. $K_a = K_{a1} \times K_{a2} \times K_{a3}$

$$(7.5 \times 10^{-3}) (6.2 \times 10^{-8}) (2.2 \times 10^{-13}) = 1.0 \times 10^{-22}$$

4. $\text{Ca}(\text{OH})_2 + \text{H}_2\text{O} \rightleftharpoons \text{Ca}^{2+} + 2\text{OH}^-$

$$\text{a pH of } 12.32 \Rightarrow [\text{OH}^-] = 2.09 \times 10^{-2}$$

since 2OH^- are produced for each Ca^{2+} , then $[\text{Ca}^{2+}] = 1.05 \times 10^{-2} \text{ M}$

$$\begin{aligned} K_{sp} &= [\text{Ca}^{2+}][\text{OH}^-]^2 \\ &= (1.05 \times 10^{-2})(2.09 \times 10^{-2})^2 \\ &= 4.6 \times 10^{-6} \end{aligned}$$

5. Moles $\text{OH}^- = (42.55\text{mL})(0.50\text{M}) = 21.28 \text{ m mol}$

at the equivalence point, $\text{mol H}^+ = \text{mol OH}^-$

moles $\text{H}^+ = 21.28 \text{ mmol}$

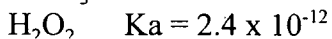
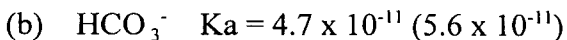
Since the acid is monoprotic, then moles acid = 21.28 mmol

$$\text{molecular mass} = \text{mass/mol} = 2.6\text{g}/21.28 \text{ mmol} = 1.2 \times 10^2 \text{g/mol}$$

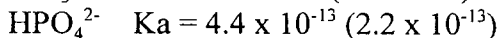
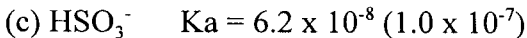
- 6.(a) HSO_3^- $K_a = 6.2 \times 10^{-8}$ (1.0×10^{-7})

$$\text{H}_3\text{PO}_4 \quad K_a = 7.1 \times 10^{-3}$$
 (7.5×10^{-3})

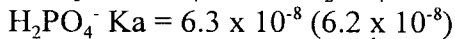
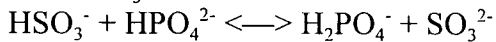
Since 7.1×10^{-3} (7.5×10^{-3}) $>$ 6.2×10^{-8} (1.0×10^{-7}), H_3PO_4 is stronger acid and equilibrium shifts to favor reactants.



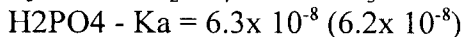
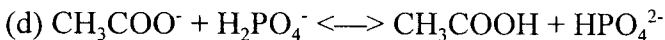
Since $4.7 \times 10^{-11} (5.6 \times 10^{-11}) > 2.4 \times 10^{-12}$, HCO_3^- is the stronger acid and products are favored



Then HSO_3^- is the acid



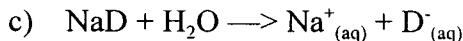
Since $6.3 \times 10^{-8} (6.2 \times 10^{-8})$ is slightly greater than $6.2 \times 10^{-8} (1.0 \times 10^{-7})$ then the products are favored.



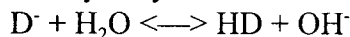
Since $1.8 \times 10^{-5} > 6.3 \times 10^{-8}$, then the reactants are favored.

- 7.a) HC Strongest acid
H₂A
HB
HA
HD weakest acid

b) The conjugate of the weakest acid is the strongest base.
The strongest base is D⁻



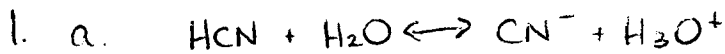
Na⁺ do not hydrolyze.



1. Fill in the missing values:

Solution	$[H_3O^+]$	$[OH^-]$	pH	pOH
Water	1.0×10^{-7}	1.0×10^{-7}	7.00	7.00
0.10 M HCl	1.0×10^{-1}	1.0×10^{-13}	1.00	13.00
0.010 M NaOH	1.0×10^{-12}	1.0×10^{-2}	12.00	2.00
1.0 M HNO_3	1.0	1.0×10^{-14}	0.00	14.00
1.0 M H_2SO_4	2.0	5.0×10^{-15}	-0.30	14.30
blood	3.98×10^{-8}	2.51×10^{-7}	7.400	6.600
16 M HNO_3	16.0	6.25×10^{-16}	-1.204	15.204
6.0 M NaOH	1.67×10^{-15}	5.99	14.777	-0.777
0.10 M Na_2O	5.0×10^{-14}	0.20	13.30	0.70
0.010 M $Ca(OH)_2$	5.0×10^{-13}	0.020	12.30	1.70
2.0 M NaOH	5.0×10^{-15}	2.0	14.30	-0.30

2. Calculate the concentration of OH^- in a water solution in which the H_3O^+ concentration is 1.0×10^{-11} M. Then calculate the pH, and tell whether the solution is acidic, basic, or neutral. $[OH^-] = \frac{k_w}{[H_3O^+]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-11}} = 1.0 \times 10^{-3} M$ $pH = -\log(1.0 \times 10^{-11}) = 11.00$ basic
3. Calculate the concentration of H_3O^+ in a water solution in which the OH^- concentration is 1.0×10^{-8} M. Then calculate the pH, and tell whether the solution is acidic, basic, or neutral. $[H_3O^+] = \frac{k_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-8}} = 1.0 \times 10^{-6} M$ $pH = -\log(1.0 \times 10^{-6}) = 6.00$ acidic
4. Calculate the concentration of H_3O^+ and OH^- in a water solution in which the $pH = 5.00$. Tell whether the solution is acidic, basic, or neutral. $[H_3O^+] = \text{antilog}(-5) = 1.0 \times 10^{-5} M$ $[OH^-] = \frac{k_w}{[H_3O^+]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} M$ acidic
5. Calculate the pH of a solution in which $pOH = 13.00$. Then calculate the concentration of OH^- and of H_3O^+ in the solution. Tell whether the solution is acidic, basic, or neutral. $[OH^-] = \text{antilog}(-13.00) = 1.0 \times 10^{-13} M$ $pH = pK_w - pOH = 14.00 - 13.00 = 1.00$ $[H_3O^+] = \text{antilog}(-1.00) = 1.0 \times 10^{-1} M$ acidic
6. Calculate the concentration of OH^- in a water solution in which the H_3O^+ is 1.0×10^{-12} M. Then calculate the pH and tell whether the solution is acidic, basic, or neutral, and support your answer. $[OH^-] = \frac{k_w}{[H_3O^+]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-12}} = 1.0 \times 10^{-2} M$ $pH = -\log(1.0 \times 10^{-12}) = 12.00$ basic because $[OH^-] > [H_3O^+]$
7. Calculate the concentration of H_3O^+ in a water solution in which the OH^- concentration is 1.0×10^{-6} M. Then calculate the pH, and tell whether the solution is acidic, basic, or neutral and support your answer. $[H_3O^+] = \frac{k_w}{[OH^-]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-6}} = 1.0 \times 10^{-8} M$ basic because $[OH^-] > [H_3O^+]$
 $pH = -\log(1.0 \times 10^{-8}) = 8.00$



c. $K_a = \frac{[\text{CN}^-][\text{H}_3\text{O}^+]}{[\text{HCN}]}$

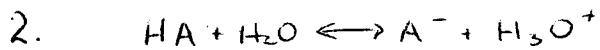
d. let $x = [\text{H}_3\text{O}^+]$, then $[\text{CN}^-] = x$ and $[\text{HCN}] = 0.25\text{M} - x \approx 0.25\text{M}$ if x is small
 $K_a = 4.9 \times 10^{-10} = \frac{x^2}{0.25}$ $x = [\text{H}_3\text{O}^+] = \underline{1.1 \times 10^{-5}\text{M}}$

e. $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.1 \times 10^{-5}\text{M}) = \underline{4.96}$

f. $[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$ $[\text{OH}^-] = \frac{1.00 \times 10^{-14}}{1.1 \times 10^{-5}\text{M}} = \underline{9.0 \times 10^{-10}\text{M}}$

g. $[\text{HCN}]_{\text{eqm}} = [\text{HCN}]_{\text{initial}} - 1.1 \times 10^{-5}\text{M} = \underline{0.25\text{M}}$

h. % dissociation = $\frac{1.1 \times 10^{-5}\text{M}}{0.25\text{M}} \times 100 = \underline{4.4 \times 10^{-3}\%}$



$K_a = \frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]}$

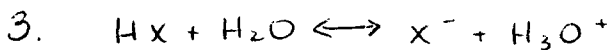
$[\text{H}_3\text{O}^+] = \text{antilog}(-3.45) = 3.55 \times 10^{-4}\text{M}$

$[\text{A}^-] = [\text{H}_3\text{O}^+] = 3.55 \times 10^{-4}\text{M}$

$[\text{HA}] = 0.20\text{M} - 3.55 \times 10^{-4}\text{M} = 0.20\text{M}$

$K_a = \frac{(3.55 \times 10^{-4})^2}{0.20} = \underline{6.3 \times 10^{-7}}$

% dissociation = $\frac{3.55 \times 10^{-4}\text{M}}{0.20\text{M}} \times 100 = \underline{1.8 \times 10^{-1}\%}$



$K_a = \frac{[\text{X}^-][\text{H}_3\text{O}^+]}{[\text{HX}]} = 1.0 \times 10^{-8}$

let $x = [\text{H}_3\text{O}^+]$, then $[\text{X}^-] = x$

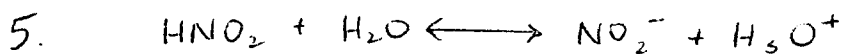
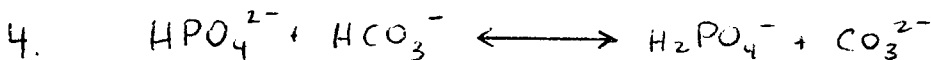
and $[\text{HX}] = 1.0\text{M} - x \approx 1.0\text{M}$ if x is small

$1.0 \times 10^{-8} = \frac{x^2}{1.0}$

$x = [\text{H}_3\text{O}^+] = 1.0 \times 10^{-4}\text{M}$

$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.0 \times 10^{-4})$

$\text{pH} = 4.00$



$K_a = \frac{[\text{NO}_2^-][\text{H}_3\text{O}^+]}{[\text{HNO}_2]} = 4.6 \times 10^{-4}$

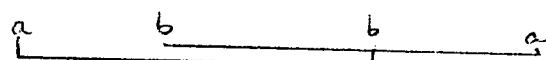
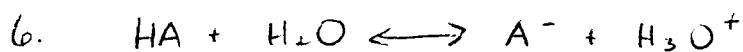
let $x = [\text{H}_3\text{O}^+]$, then $[\text{NO}_2^-] = x$

and $[\text{HNO}_2] = 0.40\text{M} - x$

$\approx 0.40\text{M}$ if x is small

$4.6 \times 10^{-4} = \frac{x^2}{0.40}$

$x = [\text{H}_3\text{O}^+] = \underline{1.4 \times 10^{-2}\text{M}}$



$[H_3O^+] = 4.5 \times 10^{-4} M$

$[A^-] = [H_3O^+] = 4.5 \times 10^{-4} M$

$[HA] = 0.35 - 4.5 \times 10^{-4} M = 0.35 M$

$K_a = \frac{[A^-][H_3O^+]}{[HA]}$

$K_a = \frac{(4.5 \times 10^{-4} M)^2}{0.35} = 5.8 \times 10^{-7}$

this acid is between H_2CO_3 and $Al(H_2O)_6^{3+}$ on the Strengths of Acids table.



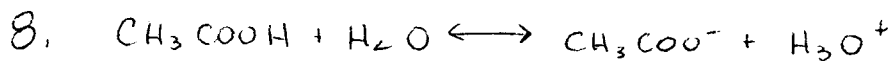
let $x = [H_3O^+]$, then $[CN^-] = x$

and $[HCN] = \frac{1.6 \text{ mols}}{2.0 L} - x \approx 0.80 M$ if x is small

$K_a = \frac{[CN^-][H_3O^+]}{[HCN]} = 4.9 \times 10^{-10}$

$4.9 \times 10^{-10} = \frac{x^2}{0.80}$

$x = [CN^-] = \underline{2.0 \times 10^{-5} M}$



$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = 1.8 \times 10^{-5}$

$[H_3O^+] = \text{antilog}(-2.456) = 3.499 \times 10^{-3} M$

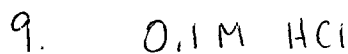
$[CH_3COO^-] = [H_3O^+] = 3.499 \times 10^{-3} M$

$1.8 \times 10^{-5} = \frac{(3.499 \times 10^{-3})^2}{[CH_3COOH]_{\text{initial}} - 3.499 \times 10^{-3} M}$

$[CH_3COOH] = [CH_3COOH]_{\text{initial}} - 3.499 \times 10^{-3} M$

$[CH_3COOH]_{\text{initial}} = 6.8 \times 10^{-1} M$ mols $CH_3COOH = 0.68 M \times 2.00 L = 1.4 \text{ mol}$

mass $CH_3COOH = 1.4 \text{ mol} \times 60.0 g/mol = \underline{82 g.}$



$\frac{1.0 \text{ mL} \times 0.1 M}{100. \text{ mL}} = 1.0 \times 10^{-3} M \text{ HCl}$

pH = 1.0

pH = 3.0

changes by 2.0 pH units.



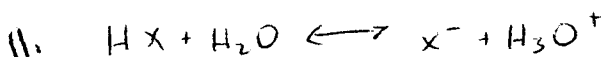
$K_a = \frac{[X^-][H_3O^+]}{[HX]}$

$[H_3O^+] = \text{antilog}(-1.40) = 3.98 \times 10^{-2} M$

$[X^-] = [H_3O^+] = 3.98 \times 10^{-2} M$

$[HX] = 0.18 M - 3.98 \times 10^{-2} M = 0.140 M$

$K_a = \frac{(3.98 \times 10^{-2} M)^2}{0.140 M} = \underline{1.1 \times 10^{-2}}$



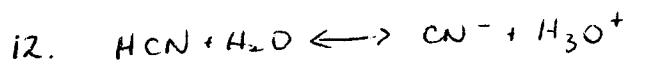
$K_a = \frac{[X^-][H_3O^+]}{[HX]}$

$[H_3O^+] = \text{antilog}(-4.26) = 5.50 \times 10^{-5} M$

$[X^-] = [H_3O^+] = 5.50 \times 10^{-5} M$

$[HX] = 1.0 M - 5.50 \times 10^{-5} M = 1.0 M$

$K_a = \frac{(5.50 \times 10^{-5})^2}{1.0} = \underline{3.0 \times 10^{-9}}$



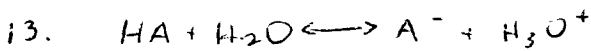
$$K_a = \frac{[\text{CN}^-][\text{H}_3\text{O}^+]}{[\text{HCN}]} = 4.9 \times 10^{-10}$$

let $[\text{H}_3\text{O}^+] = x$, then $[\text{CN}^-] = x$
and $[\text{HCN}] = 0.20\text{M} - x$
 $\approx 0.20\text{M}$ if x is small

$$4.9 \times 10^{-10} = \frac{x^2}{0.20}$$

$$[\text{H}_3\text{O}^+] = x = 9.90 \times 10^{-6}\text{M}$$

$$\% \text{ dissociation} = \frac{9.90 \times 10^{-6}\text{M}}{0.20\text{M}} \times 100 = \underline{4.9 \times 10^{-3} \%}$$



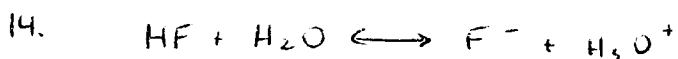
$$K_a = \frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]}$$

$$K_a = \frac{(2.5 \times 10^{-5})^2}{0.10} = \underline{6.3 \times 10^{-9}}$$

$$[\text{H}_3\text{O}^+] = 0.025\% \times 0.10\text{M} = 2.5 \times 10^{-5}\text{M}$$

$$[\text{A}^-] = [\text{H}_3\text{O}^+] = 2.5 \times 10^{-5}\text{M}$$

$$[\text{HA}] = 0.10\text{M} - 2.5 \times 10^{-5}\text{M} \approx 0.10\text{M}$$



$$K_a = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]} = 3.5 \times 10^{-4}$$

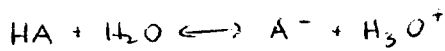
$$3.5 \times 10^{-4} = \frac{x^2}{0.50}$$

$$x = [\text{H}_3\text{O}^+] = 1.32 \times 10^{-2}\text{M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log (1.32 \times 10^{-2}\text{M}) = \underline{1.88}$$

15.



$$K_a = \frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]}$$

$$K_a = \frac{(2.51 \times 10^{-2})^2}{7.49 \times 10^{-2}} = \underline{8.4 \times 10^{-3}}$$

$$[\text{H}_3\text{O}^+] = \text{antilog}(-1.60) = 2.51 \times 10^{-2}\text{M}$$

$$[\text{A}^-] = [\text{H}_3\text{O}^+] = 2.51 \times 10^{-2}\text{M}$$

$$[\text{HA}] = 0.10\text{M} - 2.51 \times 10^{-2}\text{M} = 7.49 \times 10^{-2}\text{M}$$

16. $\text{p}K_a = -\log K_a = -\log (1.5 \times 10^{-9}) = \underline{8.82}$

17. a) HIn_1 yellow

In_1^- pink

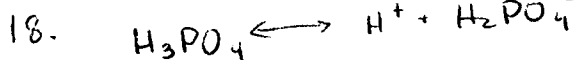
HIn_2 blue

In_2^- yellow

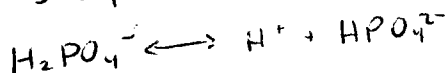
HIn_3 pink

In_3^- colorless

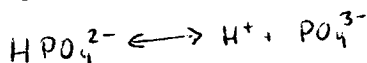
b) HA_1
 HA_2
 HIn_2
 HA_2



$$K_{a1} = \frac{[\text{H}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} = 7.5 \times 10^{-3}$$

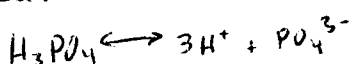


$$K_{a2} = \frac{[\text{H}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} = 6.2 \times 10^{-8}$$



$$K_{a3} = \frac{[\text{H}^+][\text{PO}_4^{3-}]}{[\text{HPO}_4^{2-}]} = 2.2 \times 10^{-13}$$

overall:



$$K_T = \frac{[\text{H}^+]^3 [\text{PO}_4^{3-}]}{[\text{H}_3\text{PO}_4]} = K_1 \times K_2 \times K_3 = \underline{1.0 \times 10^{-22}}$$

$$19. [H^+] = \text{antilog}(-7.43) = 3.7 \times 10^{-8} \text{ M}$$

$$[OH^-] = \frac{k_w}{[H^+]} = \frac{1.00 \times 10^{-14}}{3.7 \times 10^{-8}} = \underline{2.7 \times 10^{-7} \text{ M}}$$

20. HCl is a strong acid so $[H_3O^+] = [HCl] = 2.0 \text{ M}$

$$\text{pH} = -\log [H_3O^+] = -\log(2.0) = \underline{-0.30}$$

21. NaOH is highly soluble, so $[OH^-] = 0.15 \text{ M}$

$$\text{pOH} = -\log [OH^-] = -\log(0.15 \text{ M}) = 0.824$$

$$\text{pH} + \text{pOH} = \text{p}K_w \quad \text{p}K_w = 14.000 \text{ at } 25^\circ \text{C}$$

$$\text{pH} = 14.000 - 0.824 = \underline{13.18}$$

22. $\text{Ba(OH)}_2 \rightarrow 2\text{OH}^-$ $[OH^-] = 2 \times 0.025 \text{ M} = 0.050 \text{ M}$

$$\text{pOH} = -\log [OH^-] = 1.301$$

$$\text{pH} = 14.000 - 1.301 = \underline{12.70}$$

Key.

Acid #4

1. What is the concentration of a NaOH solution when 30 mL of 0.50 M HCl are needed to neutralize 50 mL of the base?
2. What is the concentration of acetic acid in vinegar when 32.5 mL of 0.56 M NaOH are required to neutralize 15 mL of vinegar?
3. What is the concentration of NH_3 in household ammonia when 48.25 mL of 0.525 M HCl are needed to neutralize 22.0 mL of the ammonia solution?
4. What is the concentration of an H_2SO_4 solution when 23 mL of 0.15 M KOH are needed to neutralize 15 mL of the acid?
5. A 5.0 g tablet of $\text{Mg}(\text{OH})_2$ neutralizes 450 mL of HCl acid. What is the molarity of the HCl acid?
6. What mass of $\text{Ca}(\text{OH})_2$ can be neutralized by 23 mL of 0.25 M HNO_3 ?

1. $\text{mol H}^+ = 30 \text{ mL} \times 0.50 \text{ M} = 15 \text{ mmol H}^+$

$\text{mol OH}^- = \text{mol H}^+ = 15 \text{ mmol OH}^-$

$[\text{NaOH}] = 15 \text{ mmol} / 50 \text{ mL} = 0.30 \text{ M}$

2. $\text{mol OH}^- = 32.5 \text{ mL} \times 0.56 \text{ M} = 18.2 \text{ mmol OH}^-$

$\text{mol H}^+ = \text{mol OH}^- = 18.2 \text{ mmol H}^+$

$[\text{CH}_3\text{COOH}] = 18.2 \text{ mmol} / 15 \text{ mL} = 1.2 \text{ M}$

3. $\text{mol H}^+ = 48.25 \text{ mL} \times 0.525 \text{ M} = 25.33 \text{ mmol H}^+$

$\text{mol OH}^- = \text{mol H}^+ = 25.33 \text{ mmol OH}^-$

$[\text{NH}_3] = 25.33 \text{ mmol} / 22.0 \text{ mL} = 1.15 \text{ M}$

4. $\text{mol OH}^- = 23 \text{ mL} \times 0.15 \text{ M} = 3.45 \text{ mmol OH}^-$

$\text{mol H}^+ = \text{mol OH}^- = 3.45 \text{ mmol H}^+$

$\text{mol H}_2\text{SO}_4 = 3.45 \text{ mmol H}^+ \left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol H}^+} \right) = 1.73 \text{ mmol}$

$[\text{H}_2\text{SO}_4] = \frac{1.73 \text{ mmol}}{15 \text{ mL}} = 0.12 \text{ M}$

5. $\text{mol OH}^- = 5.0 \text{ g} \div 58.3 \text{ g/mol} \times 2 = 0.172 \text{ mol}$

$\text{mol H}^+ = \text{mol OH}^- = 0.172 \text{ mol}$

$[\text{HCl}] = 0.172 \text{ mol} / 450 \text{ L} = 0.38 \text{ M}$

6. $\text{mol H}^+ = 23 \text{ mL} \times 0.25 \text{ M} = 5.75 \times 10^{-3} \text{ mol H}^+$

$\text{mol OH}^- = \text{mol H}^+ = 5.75 \times 10^{-3} \text{ mol OH}^-$

$\text{mol Ca}(\text{OH})_2 = 5.75 \times 10^{-3} \text{ mol OH}^- \left(\frac{1 \text{ mol Ca}(\text{OH})_2}{2 \text{ mol OH}^-} \right) = 2.88 \times 10^{-3} \text{ mol}$

$\text{mass Ca}(\text{OH})_2 = 2.88 \times 10^{-3} \text{ mol} \times 74.1 \text{ g/mol} = 0.21 \text{ g}$

Acid Base Worksheet No. 4

1a) $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{K_w}{K_a(\text{NH}_4^+)} = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.79 \times 10^{-5}$$

let $[\text{OH}^-] = x$, then $[\text{NH}_4^+] = x$
 and $[\text{NH}_3] = 0.30\text{M} - x \approx 0.30\text{M}$ if x is small

$$1.79 \times 10^{-5} = \frac{x^2}{0.30} \quad x = [\text{OH}^-] = \underline{2.3 \times 10^{-3}\text{M}}$$

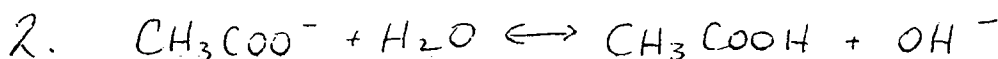
b) $\text{pOH} = -\log(2.3 \times 10^{-3}\text{M}) = 2.636$

$\text{pH} = \text{pK}_w - \text{pOH} = 14.000 - 2.636 = \underline{11.36}$

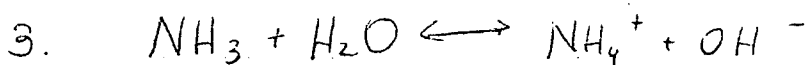
c) $[\text{H}_3\text{O}^+] = \text{antilog}(-\text{pH}) = \text{antilog}(-11.36) = \underline{4.3 \times 10^{-12}\text{M}}$

d) $[\text{NH}_3] = 0.30\text{M} - 2.3 \times 10^{-3}\text{M} = \underline{0.30\text{M}}$

e) % dissociation = $\frac{2.3 \times 10^{-3}\text{M}}{0.30\text{M}} \times 100 = \underline{0.77\%}$



$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} = \frac{K_w}{K_a(\text{CH}_3\text{COOH})} = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-5}} = \underline{5.6 \times 10^{-10}}$$



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{K_w}{K_a(\text{NH}_4^+)} = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.79 \times 10^{-5}$$

let $[\text{OH}^-] = x$, then $[\text{NH}_4^+] = x$ and $[\text{NH}_3] = 1.0\text{M} - x$
 $[\text{NH}_3] \approx 1.0\text{M}$ if x is small

$$1.79 \times 10^{-5} = \frac{x^2}{1.0}$$

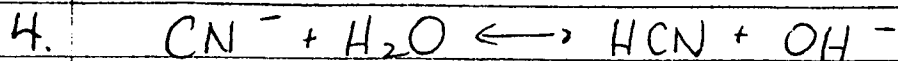
$$x = [\text{OH}^-] = 4.23 \times 10^{-3}\text{M}$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.00 \times 10^{-14}}{4.23 \times 10^{-3}} = 2.36 \times 10^{-12}\text{M}$$

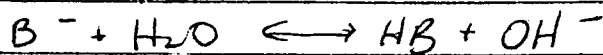
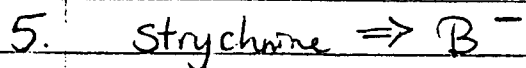
$$\text{pH} = -\log(2.36 \times 10^{-12})$$

$$\underline{\underline{\text{pH} = 11.63}}$$

Acid Base Worksheet No. 4



$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} = \frac{K_w}{K_a(\text{HCN})} = \frac{1.00 \times 10^{-14}}{4.9 \times 10^{-10}} = \underline{2.0 \times 10^{-5}}$$



$$K_b = \frac{[\text{HB}][\text{OH}^-]}{[\text{B}^-]} = 1.8 \times 10^{-6}$$

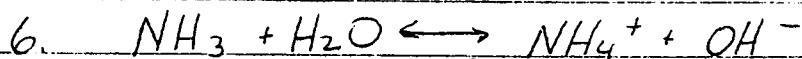
let $x = [\text{OH}^-]$, then $[\text{HB}] = x$ and $[\text{B}^-] = 0.15\text{M} - x$

$\approx 0.15\text{M}$ if x is small

$$1.8 \times 10^{-6} = \frac{x^2}{0.15}$$

$$x = [\text{OH}^-] = 5.20 \times 10^{-4}\text{M} \quad \text{pOH} = -\log 5.20 \times 10^{-4} = 3.284$$

$$\text{pH} = \text{p}K_w - \text{pOH} = 14.000 - 3.284 = \underline{10.72}$$



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.79 \times 10^{-5}$$

let $x = [\text{OH}^-]$, then $[\text{NH}_4^+] = x$ and $[\text{NH}_3] = 0.10\text{M} - x \approx 0.10$

if x is small

$$1.79 \times 10^{-5} = \frac{x^2}{0.10\text{M}} \quad x = [\text{OH}^-] = 1.34 \times 10^{-3}\text{M}$$

$$\% \text{ dissociation} = \frac{1.34 \times 10^{-3}}{0.10\text{M}} \times 100 = \underline{1.3\% \text{ dissociation}}$$

7. $\text{p}K_b = -\log K_b = -\log (2.0 \times 10^{-5}\text{M}) = \underline{6.70}$



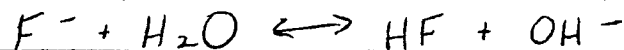
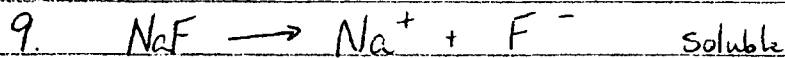
$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-5}} = \underline{5.56 \times 10^{-10}}$$

let $x = [\text{OH}^-]$, then $[\text{CH}_3\text{COOH}] = x$ and $[\text{CH}_3\text{COO}^-] = 0.25 - x \approx 0.25\text{M}$ if x is small

$$5.56 \times 10^{-10} = \frac{x^2}{0.25} \quad x = [\text{OH}^-] = 1.18 \times 10^{-5} \quad \text{pOH} = 4.929$$

$$\text{pH} = 14.000 - 4.929 = 9.07$$

Acid Base Worksheet No. 4

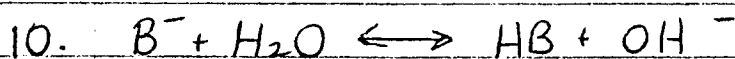


$$K_b = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]} = \frac{K_w}{K_{a(\text{HF})}} = \frac{1.00 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.86 \times 10^{-11}$$

Let $[\text{OH}^-] = x$, then $[\text{HF}] = x$ and $[\text{F}^-] = 0.50\text{M} - x \approx 0.50\text{M}$ if x is small

$$2.86 \times 10^{-11} = \frac{x^2}{0.50} \quad [\text{OH}^-] = x = 3.78 \times 10^{-6}\text{M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (3.78 \times 10^{-6}) = \underline{\underline{5.42}}$$



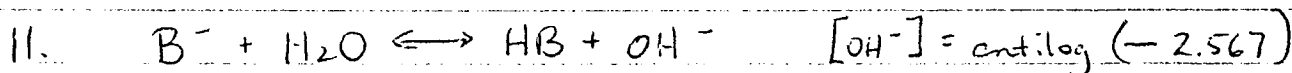
$$K_b = \frac{[\text{HB}][\text{OH}^-]}{[\text{B}^-]} \quad [\text{OH}^-] = \text{antilog}(-(\text{pOH} - \text{pK}_b))$$

$$= 4.17 \times 10^{-3}\text{M}$$

$$[\text{HB}] = [\text{OH}^-] = 4.17 \times 10^{-3}\text{M}$$

$$K_b = \frac{(4.17 \times 10^{-3})^2}{9.58 \times 10^{-2}} = 1.81 \times 10^{-4} \quad [\text{B}^-] = 0.10\text{M} - 4.17 \times 10^{-3}\text{M} = 9.58 \times 10^{-2}$$

$$K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{1.81 \times 10^{-4}} = \underline{\underline{5.5 \times 10^{-11}}}$$



$$[\text{OH}^-] = \text{antilog}(-2.567)$$

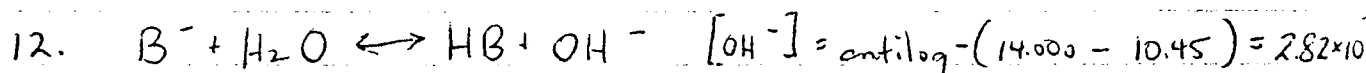
$$= 2.710 \times 10^{-3}\text{M}$$

$$K_b = \frac{[\text{HB}][\text{OH}^-]}{[\text{B}^-]}$$

$$[\text{HB}] = [\text{OH}^-] = 2.710 \times 10^{-3}\text{M}$$

$$[\text{B}^-] = 0.250 - 2.710 \times 10^{-3}\text{M} = 0.2473$$

$$K_b = \frac{(2.710 \times 10^{-3})^2}{0.2473} = 2.97 \times 10^{-5}$$



$$[\text{OH}^-] = \text{antilog}(-(\text{pOH} - \text{pK}_b)) = 2.82 \times 10^{-4}$$

$$[\text{HB}] = [\text{OH}^-] = 2.82 \times 10^{-4}\text{M}$$

$$K_b = \frac{[\text{HB}][\text{OH}^-]}{[\text{B}^-]}$$

$$[\text{B}^-] = 0.20\text{M} - 2.82 \times 10^{-4}\text{M} = 0.200\text{M}$$

$$K_b = \frac{(2.82 \times 10^{-4})^2}{0.200} = 4.0 \times 10^{-7}$$

Acid #5

Key.

1. Complete the table by filling in the blank spaces.

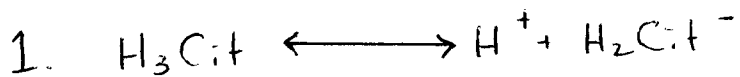
Acid	Base	Salt 0.10 M	Ion hydrolyzing	Relative acidity	pH
HCl	NaOH	NaCl	neither	neutral	7.000
HBr	KOH	KBr	neither	neutral	7.000
HNO ₃	Ca(OH) ₂	Ca(NO ₃) ₂	neither	neutral	7.000
HCl	NH ₃	NH ₄ Cl	NH ₄ ⁺	acidic	5.13
H ₂ SO ₄	KOH	KHSO ₄	HSO ₄ ⁻	acidic	omit
H ₂ CO ₃	NaOH	Na ₂ CO ₃	CO ₃ ²⁻	basic	11.63
H ₂ CO ₃	NaOH	NaHCO ₃	HCO ₃ ⁻	basic	omit
HF	KOH	KF	F ⁻	basic	8.23
HCl	Al(OH) ₃	AlCl ₃	Al(H ₂ O) ₆ ³⁺	acidic	omit
HCl	Fe(OH) ₃	FeCl ₃	Fe(H ₂ O) ₆ ³⁺	acidic	omit
HCOOH	KOH	HCOOK	HCOO ⁻	basic	8.37
HCN	Ca(OH) ₂	Ca(CN) ₂	CN ⁻	basic	11.31

2. Predict the relative acidity of each salt

AlCl₃ acidicNH₄Br acidicFe(NO₃)₃ acidicCa(NO₃)₂ neutralNa₃PO₄ basicNa₂CO₃ basicNH₄CH₃COO neutralK₂SO₄ basicK₂HPO₄ basic(NH₄)₂C₂O₄ acidicNaHCO₃ basic(NH₄)₂CO₃ basicNaHSO₃ acidicKH₂PO₄ acidicAlI₃ acidicKHSO₄ acidic

Acid #6

Answer key.



$$K_a = \frac{[\text{H}^+][\text{H}_2\text{Cit}^-]}{[\text{H}_3\text{Cit}]} = 8.4 \times 10^{-4}$$

$$[\text{H}^+] = \text{antilog}(-2.10) = 7.94 \times 10^{-3} \text{ M}$$

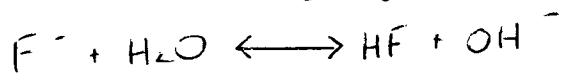
$$[\text{H}_2\text{Cit}^-] = [\text{H}^+] = 7.94 \times 10^{-3} \text{ M}$$

$$[\text{H}_3\text{Cit}] = [\text{H}_3\text{Cit}]_{\text{initial}} - [\text{H}^+]$$

$$K_a = \frac{(7.94 \times 10^{-3})^2}{[\text{H}_3\text{Cit}]_{\text{initial}} - 7.94 \times 10^{-3}} = 8.4 \times 10^{-4}$$

$$[\text{H}_3\text{Cit}]_{\text{initial}} = \underline{\underline{8.3 \times 10^{-2} \text{ M}}}$$

2. NaF: Na^+ does not hydrolyze



$$K_b = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]} = \frac{K_w}{K_a(\text{HF})} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.86 \times 10^{-11}$$

let $x = [\text{OH}^-]$

then $[\text{HF}] = x$

and $[\text{F}^-] = 1.0 \text{ M} - x$

$\approx 1.0 \text{ M}$ if x is small

$$2.86 \times 10^{-11} = \frac{x^2}{1.0}$$

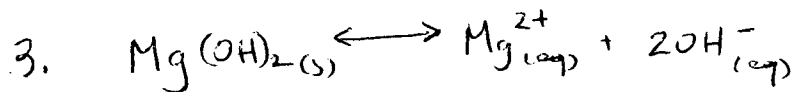
$$[\text{OH}^-] = x = 5.35 \times 10^{-6} \text{ M}$$

$$\text{pOH} = 5.272$$

$$\underline{\underline{\text{pH} = 8.73}}$$

Acid 6 cont'd.

Answer key:



$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^{-}]^2 = 1.8 \times 10^{-11}$$

$$\text{let } [\text{Mg}^{2+}] = x$$

$$\text{then } [\text{OH}^{-}] = 2x$$

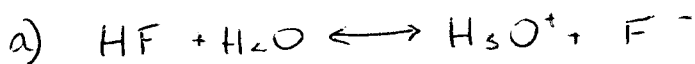
$$K_{sp} = 4x^3 = 1.8 \times 10^{-11}$$

$$x = 1.65 \times 10^{-4} \text{ M}$$

$$[\text{OH}^{-}] = 2x = 3.30 \times 10^{-4} \text{ M}$$

$$\text{pOH} = 3.481$$

$$\underline{\underline{\text{pH} = 10.52}}$$



$$K_a = \frac{[\text{H}_3\text{O}^{+}][\text{F}^{-}]}{[\text{HF}]} = 3.5 \times 10^{-4}$$

$$[\text{H}_3\text{O}^{+}] = \text{antilog}(-2.21) = 6.17 \times 10^{-3} \text{ M}$$

$$[\text{F}^{-}] = [\text{H}_3\text{O}^{+}] = 6.17 \times 10^{-3} \text{ M}$$

$$[\text{HF}] = (2.09/\text{L} \div 20 \text{ g/mol}) = 6.17 \times 10^{-3} \text{ M} = 9.38 \times 10^{-2} \text{ M}$$

$$K_a = \frac{(6.17 \times 10^{-3})^2}{9.38 \times 10^{-2}} = \underline{\underline{4.1 \times 10^{-4}}}$$

$$b) \quad \% \text{ dissociation} = \frac{6.17 \times 10^{-3}}{(2.09/\text{L} \div 20 \text{ g/mol})} \times 100 = \underline{\underline{6.2\%}}$$

$$5. \quad \text{pH range } 8.2 - 10.0$$

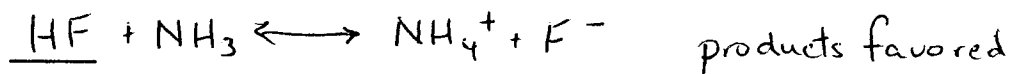
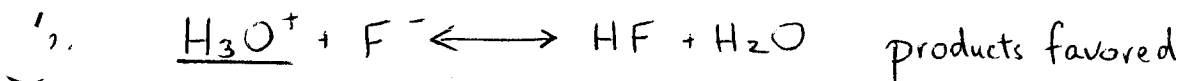
$$\text{pOH range } 5.8 - 4.0$$

$$[\text{OH}^{-}] \text{ range } 1.5 \times 10^{-6} \text{ M to } 1.0 \times 10^{-4} \text{ M a difference of } 9.35 \times 10^{-5} \text{ M}$$

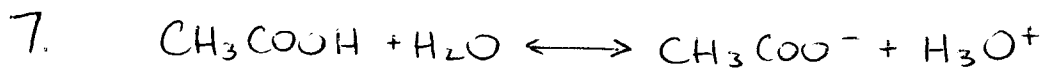
$$\underline{\underline{\text{or } 1 \times 10^{-4} \text{ M}}}$$

Acid # 6 cont'd

Answer key



Stronger acid underlined



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5}$$

$$[\text{H}_3\text{O}^+] = \text{anti. log}(-3.2) = 6.3 \times 10^{-4} \text{ M}$$

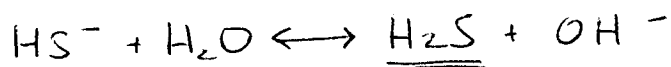
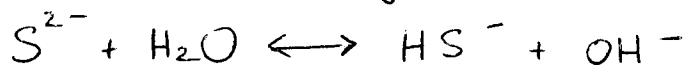
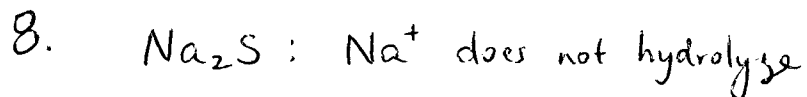
$$[\text{CH}_3\text{COO}^-] = [\text{H}_3\text{O}^+] = 6.3 \times 10^{-4} \text{ M}$$

$$[\text{CH}_3\text{COOH}] = [\text{CH}_3\text{COOH}]_{\text{initial}} - 6.3 \times 10^{-4} \text{ M}$$

$$1.8 \times 10^{-5} = \frac{(6.3 \times 10^{-4})^2}{[\text{CH}_3\text{COOH}]_{\text{ini}} - 6.3 \times 10^{-4}}$$

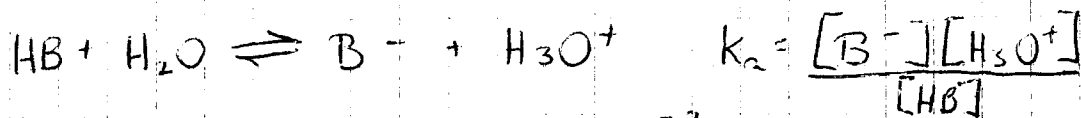
$$[\text{CH}_3\text{COOH}]_{\text{initial}} = 2.3 \times 10^{-2} \text{ M}$$

$$\text{mass} = 2.3 \times 10^{-2} \text{ mol/L} \times 60.0 \text{ g/mol} = \underline{\underline{1.4 \text{ g per litre}}} = \underline{\underline{1 \text{ g/L}}}$$



as H_2S gas escapes the reactions proceed in the forward direction to produce more OH^- (strongly basic) and H_2S gas (odour.)

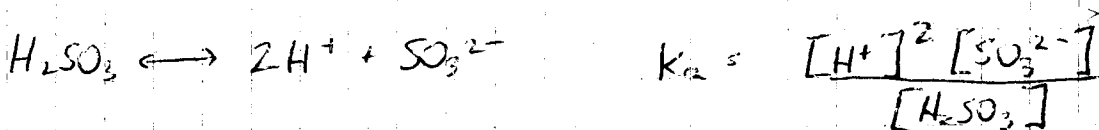
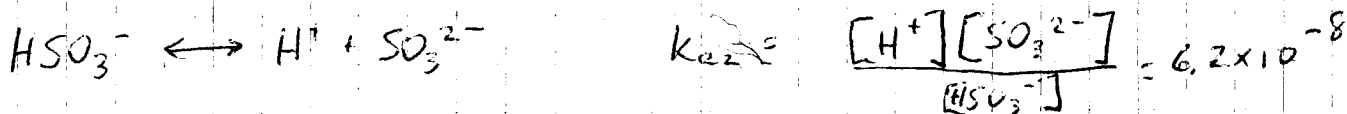
Acid # 1
Acid HB



$$[\text{H}_3\text{O}^+] = [\text{B}^-] \quad K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HB}]}$$

$$K_a [\text{HB}] = [\text{H}_3\text{O}^+]^2$$

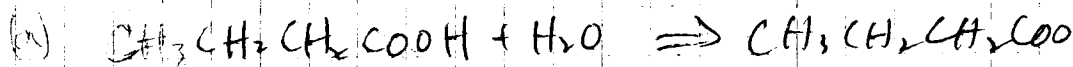
$$[\text{H}_3\text{O}^+] = \sqrt{K_a [\text{HB}]}$$



$$K_a = K_{a1} \times K_{a2}$$

$$\frac{[\text{H}^+][\text{HSO}_3^-]}{[\text{H}_2\text{SO}_3]} \times \frac{[\text{H}^+][\text{SO}_3^{2-}]}{[\text{HSO}_3^-]} = \frac{[\text{H}^+]^2 [\text{SO}_3^{2-}]}{[\text{H}_2\text{SO}_3]} = K_a$$

$$K_a = 1.7 \times 10^{-2} \times 6.2 \times 10^{-8} = 1.1 \times 10^{-9}$$



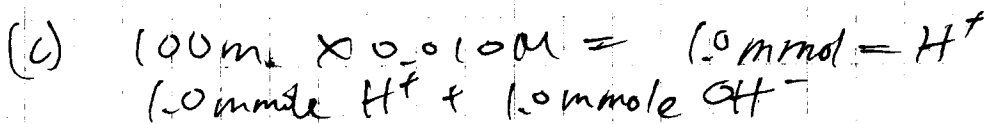
(b) $K_a = 1.54 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-]}{[\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}]}$

Let $x = [\text{H}_3\text{O}^+] = [\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-]$
then $[\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}] = 0.010 \text{ M} - x$

$$1.54 \times 10^{-5} = \frac{x^2}{0.010}$$

$$\therefore x = [\text{H}_3\text{O}^+] = 3.92 \times 10^{-4}$$

pH = 3.41



$$\text{Vol NaOH} = \frac{1.0 \text{ mmol}}{0.35 \text{ M}} = 2.86 \text{ mL} = 2.9 \text{ mL}$$

(d) H_2O water

$NaCH_3CH_2CH_2COO$ sodium butanoate

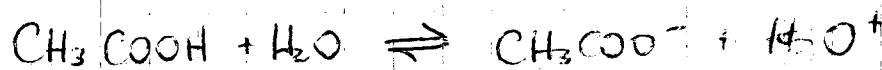
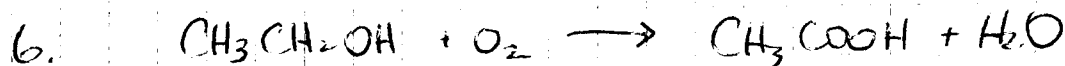
4. # of moles $(50\text{ mL} \times 0.50\text{ M}) + (100\text{ mL} \times 0.25\text{ M}) = 50$ mmoles

$$[H_2SO_4] = \frac{50 \text{ mmoles}}{50\text{ mL} + 100\text{ mL}} = 0.33\text{ M}$$

5. Fumaric Acid \rightarrow HFum.



as sodium hydroxide reacts with H^+ to produce water $[H^+]$ decreases and equilibrium shifts to products causing more fumaric acid to dissolve as the titration proceeds.



$$K_a = 1.8 \times 10^{-5} = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$[H_3O^+] = [CH_3COO^-] = \text{antilog } -2.80 = 1.58 \times 10^{-3}\text{ M}$$

$$[CH_3COOH] = X - 1.58 \times 10^{-3}\text{ M}$$

$$1.8 \times 10^{-5} = \frac{(1.58 \times 10^{-3})^2}{X - 1.58 \times 10^{-3}}$$

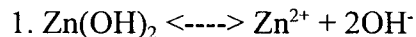
$$X = 1.40 \times 10^{-1}\text{ M}$$

amount of ethanoic acid in 1 litre = 0.140 moles

0.140 moles of ethanoic

0.140 moles \times 46 g/mols = 6.4 grams

Acid #8

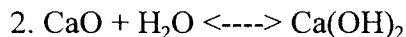


$$K_{sp} = [\text{Zn}^{2+}][\text{OH}^-]^2$$

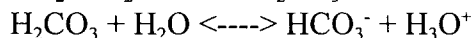
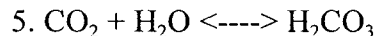
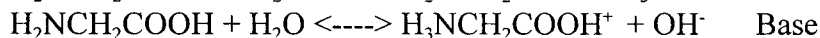
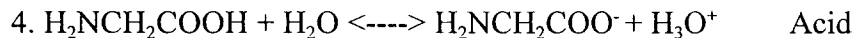
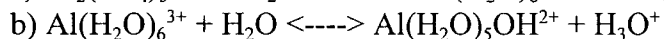
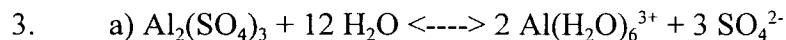
$$[\text{OH}^-] = \text{Antilog}^{-1}(-14.000 - 9.627) = 4.236 \times 10^{-5} \text{ M}$$

$$[\text{Zn}^{2+}] = 1/2 (4.236 \times 10^{-5}) = 2.118 \times 10^{-5} \text{ M}$$

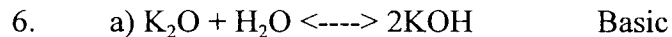
$$K_{sp} = (2.118 \times 10^{-5})(4.236 \times 10^{-5})^2 \\ = 3.80 \times 10^{-14}$$



Calcium hydroxide dissolves to form OH^- ions in solution



In daytime CO_2 levels are low and solution has a high pH at night when CO_2 levels rise the pH drops.



Acid #9 **Answer Key**

- $$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$$

Since the reaction is endothermic, then, lowering the temperature from 25°C to 5.0°C shifts the equilibrium to favour reactants. This shift lowers the $[\text{H}_3\text{O}^+]$ subsequently raising the pH.
- Colour changes occur when $[\text{H}^+]$ in the solution equals the K_a of the indicator. Since different indicators have different K_a values, then they must change colour at different pH.
- Moles diprotic acid = $2.68\text{g} / 126\text{g mol}^{-1} = 0.02127\text{ mol}$
 $[\text{acid}] = \text{mol} / \text{vol} = 0.02127 / 0.250 = 0.08508\text{ M}$
 Moles acid used = $(0.08508\text{M})(25.0\text{ml}) = 2.127\text{mmol}$
 Moles H^+ used = $(2.127\text{mmol}) \times 2 = 4.254\text{ mmol}$
 At the stoichiometric point, $\text{mol H}^+ = \text{mol OH}^-$, then: moles $\text{OH}^- = 4.254\text{ mmol}$
 $[\text{NaOH}] = 4.254\text{mmol} / 19.32\text{ml} = 0.220\text{M}$
- The K_b need to be compared:

NO_2^-	$K_b = 2.0 \times 10^{-11}$ (2.17×10^{-11})
$\text{C}_6\text{H}_5\text{COO}^-$	$K_b = 1.5 \times 10^{-10}$ (1.54×10^{-10})

Since K_b for $\text{C}_6\text{H}_5\text{COO}^-$ is larger, then $\text{C}_6\text{H}_5\text{COO}^-$ is the stronger base.
- $\text{Al}(\text{NO}_3)_3 + \text{H}_2\text{O} \rightarrow \text{Al}(\text{H}_2\text{O})_6^{3+} + \text{NO}_3^-$
 The NO_3^- do not hydrolyze
 The hydrated aluminum ions do hydrolyze and produce H_3O^+ which give the $\text{Al}(\text{NO}_3)_3$ solution its acid property.

$$\text{Al}(\text{H}_2\text{O})_6^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+} + \text{H}_3\text{O}^+$$
- Let the acid be HA.

$$\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$$

Then: $[\text{H}_3\text{O}^+] = [\text{A}^-] = (20\%)(0.100) = 0.0200\text{ M}$
 $[\text{HA}] = 0.100 - 0.0200 = 0.0800\text{M}$
 $K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$
 $K_a = (0.0200)^2 / 0.0800$
 $K_a = 5.00 \times 10^{-3}$
- $\text{H}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HS}^-$
 Let $X = [\text{H}_3\text{O}^+] = [\text{HS}^-]$
 Then $[\text{H}_2\text{S}] = 0.0500 - X \sim 0.0500$ (if X is small)
 $K_a = [\text{H}_3\text{O}^+][\text{HS}^-] / [\text{H}_2\text{S}]$
 $(9.1 \times 10^{-8})(1.0 \times 10^{-7}) = X^2 / 0.0500$
 $[\text{H}_3\text{O}^+] = X = 7.07 \times 10^{-5}$ (6.75×10^{-5})
 $\% \text{dissociation} = [\text{H}_3\text{O}^+] / \text{initial}[\text{H}_2\text{S}] \times 100$
 $\% = 7.07 \times 10^{-5} / 0.0500 \times 100 = 0.14\% \quad 0.13\%$

8. Compare the K_a with the K_b for the ion:

$$K_a \text{ for } \text{HPO}_4^{2-} = 4.4 \times 10^{-13} \quad (2.2 \times 10^{-13})$$

$$K_b \text{ for } \text{HPO}_4^{2-} = 1.6 \times 10^{-7}$$

Since the K_b is far greater than the K_a , the ion and its solution is basic.

9. $[\text{Acid}] = \text{moles} / \text{vol} = ((100.0\text{g}) / 120\text{gmol}^{-1}) / 0.600 = 1.389 \text{ M}$

$$[\text{H}^+] = 2 \times [\text{Acid}] = 2 \times 1.389 = 2.778 \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = -4.437 = -0.444$$

10. (a) In acid solution, the acid form of an indicator predominates; H₂Sg is present in larger concentration.

(b) Because the acid form is green, the indicator would colour the acid solution green.

11. Moles $\text{OH}^- = (50.0)(0.20) = 10.0\text{mmol}$

at the stoichiometric point, $\text{mol H}^+ = \text{mol OH}^-$, then moles $\text{H}^+ = 10.0 \text{ mmol}$

moles oxalic acid = $10.0 \text{ mmol} \times 1/2 = 5.00 \text{ mmol acid}$

mass of acid = $(5.00 \text{ mmol})(126\text{g/mol}) = 0.63 \text{ g}$

12. $K_b = K_w / K_a = 1.00 \times 10^{-14} / 6.6 \times 10^{-5} = 1.5 \times 10^{-10}$

13. $\text{C}_6\text{H}_5\text{COONa} + \text{H}_2\text{O} \rightleftharpoons \text{Na}^+ + \text{C}_6\text{H}_5\text{COO}^-$

Na^+ do not hydrolyze

$\text{C}_6\text{H}_5\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{COOH} + \text{OH}^-$

Let $x = [\text{OH}^-] = [\text{C}_6\text{H}_5\text{COOH}]$, then $[\text{C}_6\text{H}_5\text{COO}^-] = 0.10 - x \sim 0.10\text{M}$

$$K_b = [\text{C}_6\text{H}_5\text{COOH}][\text{OH}^-] / [\text{C}_6\text{H}_5\text{COO}^-]$$

$$1.5 \times 10^{-10} = X^2 / 0.10 \quad X = [\text{OH}^-] = 3.87 \times 10^{-6} \text{ M}$$

$$\text{pH} = 8.59$$

14. $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$

$$K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$$

$$[\text{H}_3\text{O}^+] = [\text{A}^-] = 3.0 \times 10^{-3} \text{ M}$$

$$[\text{HA}] = 0.0100 - 3.0 \times 10^{-3} = 7.00 \times 10^{-3} \text{ M}$$

$$K_a = (3.0 \times 10^{-3})^2 / 7.00 \times 10^{-3}$$

$$K_a = 1.3 \times 10^{-3}$$