Chem 1 Dr. Casagrande / Mrs. Rinaldi

Chapter 19 Review Sheet: Acids & Bases

1. Identify the physical and chemical properties of acids and bases.
   Acids: corrosive to metal and tissue, electrolytes, sour taste, turn blue litmus red, react with metal to form H_2, react with CO_3^{2-} to form CO_2; Bases: corrosive to tissue, slippery, electrolytes, bitter taste, turn red litmus blue
2. Classify solutions as acidic, basic, or neutral based on the relative levels of $[H^+]$ and $[OH^-]$.
   Acidic: $[H^+] > [OH^-]$; Basic: $[OH^-] > [H^+]$; Neutral: $[H^+] = [OH^-]$
3. Describe what an acid-base neutralization reaction is.
   A reaction of an acid and a base to form a salt (any ionic compound not containing H^+ or OH^-) and water.
4. Write a neutralization reaction for a given acid-base system.
   $HX\ (aq) + YOH\ (aq) \rightarrow H_2O\ (l) + YX\ (aq)$
5. Explain what titration is and how neutralization reactions are used in acid-base titrations (Lab!).
   A solution of acid or base of known concentration is used to neutralize a solution of base or acid of unknown concentration, and knowing the volume of each solution and the concentration of the known, the concentration of the unknown is calculated (stoichiometry!)
6. Explain why, at the equivalence point, mol H^+ = mol OH^-
   In a neutralization reaction, 1 mol H^+ reacts with 1 mol OH^- to for 1 mol H_2O.
7. Describe what an indicator is and explain how it determines the endpoint of a titration.
   It is a dye that changes color at a particular pH; it is used to change color at or near the equivalence point to indicate the end point of the titration.
8. Determine the concentration an unknown solution being titrated given its volume and the volume and concentration of the known solution.
   Use Chart G: vol Known × M known → mol Known → mol Unknown ÷ vol Unknown = M unknown.
9. Describe the Arrhenius and Brønsted models for acids and bases.
   Arrhenius acids dissociate to give H^+ in aqueous solution, Arrhenius bases dissociate to give OH^- in aqueous solution; Brønsted acids donate H^+ (proton), Brønsted bases accept H^+.
10. Identify the acid, base, conjugate acid, and conjugate base in a reaction.
    Acid gives H^+ to base (on reactants side). Conjugate acid is what is formed when the base has gained an H^+, conjugate base is what is formed when the acid has lost H^+. (See, e.g. #50 below.)
11. ~~Determine whether an acid is mono-, di-, or triprotic and write the ionization reactions for a polyprotic acid.~~
    ~~HX: monoprotic, H_2X diprotic, H_3X: triprotic. Each proton is removed stepwise to form H_3O^+ and the conjugate base with one fewer H^+.~~
12. Relate the strength of an acid or base (strong or weak) to its degree of ionization (fully or partially ionized).
    Strong acids and bases ionize (dissociate) fully (100%). Weak acids and bases only partially ionize.
13. ~~Write an acid ionization (dissociation) reaction equation and write the acid ionization constant expression (K_a) for it.~~
    ~~$HX\ (aq) \rightarrow H^+\ (aq) + X^-(aq)$; $K_a = \frac{[H^+][X^-]}{[HX]}$, where X is any monatomic or polyatomic ion.~~
14. Compare the strengths of weak acids or bases from the values of their acid or base ionization constants (K_a or K_b).
    The higher the K_a or K_b, the stronger the acid or base.
15. Relate the strength of an acid or base to its strength as an electrolyte.
    Stronger acids or bases are stronger electrolytes because there are more ions in solution.

16. Given the $[H^+]$ or $[OH^-]$, calculate the other from $K_w = [H^+][OH^-] = 1.0\times10^{-14}$.
   Use the equations $[H^+] = \dfrac{1.0\times10^{-14}}{[OH^-]}$ or $[OH^-] = \dfrac{1.0\times10^{-14}}{[H^+]}$
17. Given the $[H^+]$ calculate pH or given the $[OH^-]$ calculate pOH.
   Use the equations $pH = -\log[H^+]$ or $pOH = -\log[OH^-]$.
18. Classify solutions as acidic, neutral, or basic based on their pH.
   Acid: pH < 7; Base: pH > 7; Neutral: pH = 7
19. Given pH or pOH, determine the other from pH + pOH = 14.
   Use the equations pH = 14 – pOH or pOH = 14 – pH
20. ~~Describe the dangers of hydrofluoric acid (*An Invisible Fire* article) and the general treatment for accidental exposure.~~
   ~~See answer key for Intro to Acids & Bases Lab~~
21. Describe the causes and effects of acid rain and explain what has been done to reduce it.
   See answer key for Titration of Vinegar Lab

Chapter 19 Chapter Assessment pp. 630-632 #42, ~~47,~~ 50, 52, 53, 58, 62, 63, 64, 83, ~~85,~~ 87, 88, 89, 90, 95, 96.

42. In terms of ion concentrations, distinguish between acidic, neutral, and basic solutions.
   Acidic: $[H^+]>[OH^-]$; neutral: $[H^+] = [OH^-]$; basic $[H^+]<[OH^-]$

47. ~~Explain the difference between a monoprotic acid, a diprotic acid, and a triprotic acid. Give an example of each.~~
   ~~A monoprotic acid: can donate one H^+ (e.g. HCl); diprotic acid: can donate 2 H^+ (e.g. H_2SO_4); triprotic acid: can donate 3 H^+ (e.g. H_3PO_4)~~

50. Identify the conjugate acid-base pairs in the equilibrium equation.
   $$\underset{A}{HC_2H_3O_2} + \underset{B}{H_2O} \rightleftharpoons \underset{CA}{H_3O^+} + \underset{CB}{C_2H_3O_2^-}$$

52. Explain the difference between a strong acid and a weak acid.
   In dilute aqueous solution, a strong acid ionizes completely while a weak acid only slightly ionizes.

53. Why are strong acids and bases also strong electrolytes?
   Because strong acids and bases ionize completely, the high concentration of ions produced causes the solution to have high electrical conductivity.

58. Explain why the base ionization constant (K_b) is a measure of the strength of a base.
   The greater the value of K_b, the greater the $[OH^-]$ and stronger the base.

62. What is the relationship between the pOH and the hydroxide-ion concentration of a solution?
   $pOH = -\log[OH^-]$

63. Solution A has a pH of 2.0. Solution B has a pH of 5.0. Which solution is more acidic? Based on the hydrogen-ion concentrations in the two solutions, how many times more acidic?
   Solution A is more acidic. It is 10^3 or 1000 times more acidic than solution B.

64. If the concentration of hydrogen ions in an aqueous solution decreases, what must happen to the concentration of hydroxide ions? Why?

[OH^-] must increase because [H^+][OH^-] = $1.0\times10^{-14} = K_w$.

83. Write the balanced chemical equation for the ionization of perchloric acid ($HClO_4$) in water.

$HClO_4\,(aq) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + ClO_4^-\,(aq)$

A B CA CB

~~85. Write the equation for the ionization reaction and the acid ionization constant expression for the HS^- ion in water.~~

$\cancel{HS^-\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + S^{2-}\,(aq);\ K_a = \frac{[H_3O^+][S^{2-}]}{[HS^-]}}$

87. Given the concentration of either hydrogen ion or hydroxide ion, use the ion product constant of water to calculate the other ion at 298 K.

a. [H^+] = 1.0×10^{-4} M $\quad [OH^-] = \dfrac{1.0\times10^{-14}}{1.0\times10^{-4}} = 1.0\times10^{-10}\ \text{M}$

b. [OH^-] = 1.3×10^{-2} M $\quad [H^+] = \dfrac{1.0\times10^{-14}}{1.3\times10^{-2}} = 7.7\times10^{-13}\ \text{M}$

H^+ — OH^-

pH — pOH

88. Calculate the pH at 298 K of solutions having the following ion concentrations.

a. [H^+] = 1.0×10^{-4} M $\quad$ pH = –log(1.0×10^{-4}) = 4.00

b. [H^+] = 5.8×10^{-11} M $\quad$ pH = –log(5.8×10^{-11}) = 10.24

89. Calculate the pOH and pH at 298 K of solutions having the following ion concentrations.

a. [OH^-] = 1.0×10^{-12} M $\quad$ pOH = –log(1.0×10^{-12}) = 12.00; pH = 14 – 12.00 = 2.00

b. [OH^-] = 1.3×10^{-2} M $\quad$ pOH = –log(1.3×10^{-2}) = 1.89; pH = 14 – 1.89 = 12.11

90. Calculate the pH of each of the following strong acid or strong base solutions at 298 K.

Since these are strong acids & bases, [H^+] = concentration of acid or [OH^-] = concentration of base

a. 2.6×10^{-2} M HCl $\quad$ pH = –log(2.6×10^{-2}) = 1.59

b. 0.28 M HNO_3 $\quad$ pH = –log(0.28) = 0.55

c. 7.5×10^{-3} M NaOH $\quad$ pOH = –log(7.5×10^{-3}) = 2.13; pH = 14 – 2.13 = 11.87

d. 0.44 M KOH $\quad$ pOH = –log(0.44) = 0.36; pH = 14 – 0.36 = 13.64

$Sr(OH)_2$ $\quad [OH^-] = 2[Sr(OH)_2]$

95. In a titration, 33.21 ml 0.3020 M rubidium hydroxide (RbOH) solution is required to exactly neutralize 20.00 mL hydrofluoric acid (HF) solution. What is the molarity of the hydrofluoric acid solution?

$$\text{mol HF} = 0.03321\ \text{L RbOH} \times \frac{0.3020\ \text{mol}}{1\ \text{L}} \times \frac{1\ \text{mol HF}}{1\ \text{mol RbOH}} = 0.01003\ \text{mol HF}$$

* Stoich

$$M_{HF} = \frac{0.01003\ \text{mol HF}}{0.02000\ \text{L}} = 0.5015\ \text{M}$$

$M = \dfrac{mol}{L}$

96. A 35.00 mL-sample of NaOH solution is titrated to an endpoint by 14.76 mL 0.4122 M HBr solution. What is the molarity of the NaOH solution?

$$\text{mol NaOH} = 0.01476\ \text{L HBr} \times \frac{0.4122\ \text{mol}}{1\ \text{L}} \times \frac{1\ \text{mol NaOH}}{1\ \text{mol HBr}} = 0.006084\ \text{mol NaOH}$$

$$M_{NaOH} = \frac{0.006084\ \text{mol NaOH}}{0.03500\ \text{L}} = 0.1738\ \text{M}$$