# Acid/Base Equilbrium

**Arrhenius acid and Arrhenius base**

**An acid is a substance which, when dissolved in water, increases the concentration of hydrogen ions, H+(aq).  For example:**

**HCl (aq) 🡪 H+(aq) + Cl-(aq)**

**A base is a substance which, when added to water, increases the concentration of hydroxide ion, OH-(aq). For example:**

**NaOH (aq)🡪 Na+(aa) + OH-(aq)**

**Water is the basis of acid-base chemistry and undergoes auto ionization to a small extent**

**H20(l) ⬄H+(aq) + OH-(aq)**

**Kw=[H+][OH-]**

**Or another way: H20(l) + H20(l) ⬄H3O+(aq) + OH-(aq)**

**Kw=[H3O+][OH-]**

**Kw = 1.0 x 10-14**

**Because the concentrations above are so small, we use a scale based on logs.**

**pH = -log[H+] pOH = -log[OH-] [OH-][H+] = 1 x 10-14**

**For neutral aqueous solutions, the pH =7.**

**pHRange**

**0-6.99 acidic**

**7 neutral**

**7.01-14 basic**

**1. For aqueous solutions of the following substances, write the dissociation reaction and indicate whether the substance behaves as an Arrhenius acid or base.**

**a) HF(aq)**

**b) HC6H5O(aq)**

**c) Ba(OH)2(aq)**

**d) LiOH (aq)**

**e) H2O (aq)**

**f) H2CO3**

**2. Calculate the pH and pOH in each of the following aqueous solutions. Indicate whether the solution is acidic or basic.**

**a) [H+]=1.4 x 1 0-3M d) [H+]=7.11 x 10-11M**

**b) [OH-]= 2.08 X 10-7M e) [H+]=4.0 M**

**c) [OH-] = 6.44 X 10-14M f) [OH-]=10.1M**

**3. Calculate the [H+] and [OH-] in each of the following aqueous solutions.**

**a) pH=7.41**

**b)pH=11.0**

**c) pOH=0.230**

**d) pOH = 7.00**

**e) pH=14.9**

**f) pH=-0.543**

***Bronsted-Lowry acid and Bronsted –Lowry base***

**An acid is defined as a substance that donates a proton and a base as a substance that accepts a proton.**

**In the equation below, HCL is a Bronsted acid because it donates a proton (H+) to H2O. H2O is a Bronsted base because it accepts the proton (H+) from HCl.**

**On the right-hand side of the reaction , Cl- is a Bronsted base because it accepts the proton (H+) from H3O+. H3O+  is a Bronsted base because it donates a proton (H+) to Cl-.**

**HCl(aq) + H2O(l) <-> Cl-(aq) + H3O+(aq)**

**Acid base conj base conj acid**

**A strong acid completely dissociates in water producing a high concentration of ions. Strong**

**acids are electrolytes. For strong acids , the [H+] is equal to the concentration of the acid. A**

**weak acid does not completely dissociate in water. Weak acids are weak electrolytes. The [H+]**

**is less than concentration of the acid. Use the rice table.**

**A strong base completely dissociates in water producing a high concentration of ions. Strong bases are strong electrolytes . for strong bases the [OH-] is equal to the concentration of the base. A weak base does not completely dissociate in water. Weak bases are weak electrolytes and the [OH-] is less than the concentration of the base.**

**4. For the following compounds, write the reaction with water and indicate the Bronsted acids, bases, the conjugate acid, and conjugate bases.**

**a. HCl(g)**

**b. NH3**

**c. HCN**

**d. HC6H5O**

**e. CH3CH2NH2**

**5. For each of the following acids, write the formula for the conjugate base.**

**A. H2PO-**

**B. HClO2**

**C. H2O**

**D. CH3NH3-**

**E. Oh-**

**F. NH4-**

**6. For each of the following bases, write the formula for the conjugate acid.**

**A. OH-**

**B. Cl-**

**C. CO3-2**

**D.H2O**

**E.CH3CH2NH2**

**F. (CH3)3N**

**7. Calculate the pH of a 0.450 M HCL solution.**

**8. Calculate the pH of a 0.710 m KOH solution.**

**9. Calculate the magnitude of the equilibrium constant for benzoic acid, HC7H5O2, if a 0.100 m solution has a pH=2.59.**

**10. Calculate the magnitude of the equilibrium constant for an aqueous solution of ammonia if 0.100 M solution has a pH=11.13.**

**11. Calculate the pH of a solution, which is 0.712M CH3NH2 (methylamine) Kb=4.4x10-4**

**Hydrolysis of Salts**

**Salts are ionic compounds which are formed in a neutralization reaction between an acid and a base. Salts can be characterized from the type of acid and base which combine in the neutralization reaction. Salts can be formed from the reaction of a strong acid and a strong base, a strong acid and a weak base, a weak acid and a strong base, or a weak acid and a weak base.**

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| **NaCl is a neutral salt. In general, salts are formed from the reaction between a strong acid and a strong base are neutral.**  **To explain this look at the reaction each ion will undergo in water and decide if it is physically possible. The two ions produced in water Na+ and Cl-.**  **Na+ + H2O ⬄ NaOH + H+**  **Cl - +H2O ⬄HCl + OH-**  **Notice that both of these ions undergo hydrolysis (reaction with water) to form a strong base and a strong acid. Since these species completely break up in solution these reactions will not happen, and the solution will be neutral.** |

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| **NH4Cl is an acidic salt. In general, salts formed from the reaction between a strong acid and a weak base are acidic.**  **The two ions produced in solution from Nh4Cl are NH4+ and Cl-. We already know that Cl- will not undergo hydrolysis so let’s look at NH4+.**  **NH4++H20 ⬄ NH3+H3O+**  **This forms a weak base, which is physically possible. Remember that since a weak base barely breaks up, this means that it will generally form. Therefore this produces an acidic solution.** |

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| **NaC2H3O2 is a basic salt. In general, salts formed from the reaction between a weak acid and a strong base are basic.**  **The two ions produced in an aqueous solution are Na+ and C2H3O2-. We know that the OH-will not undergo hydrolysis, so we should look at the acetate anion.**  **C2H3O2-(aq)+H2O(aq) ⬄ HC2H3O2-(aq)+OH-(aq)**  **Since this forms a weak acid, this is physically possible. These salts will form a basic solution.** |

**Doing hydrolysis calculations**

**A+ + H2O ⬄ HA + OH-**

**Kb =**

**However, if you go to look u a Kb for this in your book, like in the tables in your index, you won’t find it. You can, however, for a Ka for HA (the conjugate acid).**

**HA + H2O ⬄ A- + H3O+**

**Ka =**

**Ka\*Kb =\*= [H3O+] [OH-] = Kw**

**This is an extremely important relationship in hydrolysis problems:**

**Ka\*Kb= Kw**

**Polyprotic acids are compounds containing more than one ionizable proton. A polyprotic acid with the general formula H2A has two ionizable protons. The loss of the protons occur stepwise. The two equations which describe the dissociation of the protons in H2A are:**

**H2A(aq) ⬄ H+(aq) + HA-(aq)**

**HA- ⬄ H+(aq) + A-2(aq)**

**One important point about polyprotic acids is that after a proton is lost, the next proton is lost much less easily. This means that the acid becomes weaker and weaker.**

**12. Predict the products when KCN(s) is added to water. Will the pH of the solution formed when the salt is added to water be greater or less than 7?**

**13. Write the dissociation equation that describes what happens when CH3NH3Cl(s) is added to water.**

**14. Write the equation which describes the acidic character of CH3NH3+.**

**Write the equilibrium expression and calculate Ka for CH3NH3+.**

**15. Calculate the pH of a 0.700 M NaC2H3O2 solution. Ka for HC2H3O2 = 1.8 x 10-5**

**16. Calculate the pH of the following salt solutions.**

**a. 0.355 M KClO**

**b. 0.777 M NH4Cl**

**c. 0.0345 M KCl**

**d. 0.411 M KC2H3O2**

**Hydrolysis of salts of polyprotic acids**

**Consider the salt NaH2PO4, which is a salt of the polyprotic acid H3PO4. Na+ will not undergo hydrolysis, so we should only deal with H2PO4-2. However, in this case, two possible hydrolysis reactions could happen.**

**H2PO4-1 + H2O ⬄H3O+ + HPO4-2  Ka2= 6.2 x 10-8**

**H2PO4-1 + H2O ⬄ OH- + H3PO4Kb = Kw/Ka1 = 1.3 x 10-12**

**Notice this is an example of amphoteric behavior. Which one will occur? Both will in reality, but the one with the larger magnitude of K is the only reaction we have to deal with. In this case, the first reaction occurs to a greater extent and the only reaction we need to deal with (Ka is larger than Kb).**

**One problem students run across is which Ka for the polyprotic acid to use when calculating a Kb value. Just look at the species (compounds) in the hydrolysis reaction and make sure they all appear in that Ka expression.**

**What makes an acid an acid and a base a base?**

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**The greater the difference in electronegativity between two elements, the weaker the bond will be and the greater the chance that the bond will break.**

**Which bond is more likely to break is sulfuric acid, the S-O bond or the O-H bond?**

**Which bond is more likely to break in sodium hydroxide, the Na-O bond or the O-H bond?**

**Strength of acids**

The strength of a binary (two element) acids are determined by the size of the anion. The larger the anion is, the more spread out the bond is, and therefore the more easily the acid ionizes.

Summary: Larger ions make stronger acids because their bond is more easily broken.

Weaker HF < HCl < HBr < HI Stronger

Oxyacids (acids containing oxygen) don’t have the hydrogen directly attached to the central atom, so a different affect is at work there. There can be two reasons for different acid strength.

I. A different number of oxygens on an acid having the same central atom.

The greater the number of oxygens, the stronger the acid.

Weaker HClO < HClO2 < HClO3 < HClO4 stronger

II. A different central atom on acids having the same number of oxygens.

The more electronegative the central atom is, the stronger the acid will be (electronegativity weakens the bond)

Weaker HIO3 < HBrO3 < HClO3 Stronger

**17. Carbonic acid H2CO3 is a diprotic acid.**

**i) Write the two dissociation reactions, demonstrating the diprotic behavior.**

**ii) If the initial concentration of H2CO3 is 0.100 M, calculate [H+]. In your calculation, assume only the first dissociation occurs. (Note: the equilibrium constant for the first dissociation, Ka1 is 4.3 x 10-7).**

**iii) Now consider the second dissociation equation for which Ka2 = 5.6 x 10-11What is the initial concentration of [HCO3-]? What is the initial concentration of [H+]? Calculate the final [H+] assuming the second dissociation occurs.**

**18. Determine the pH of a 0.100 M NaH2PO4 solution.**

**19. Determine the pH of a 0.100 M Na2HPO4 solution .**