

**Textbook:** pg. 673 #98, 100, 104, 106\*, 107\*, 108, 109, 112, 117, 120\*, 123\*, 125, 129, 133, 138, 139.

*Hints for questions with \*:*

- **#106:** ammonium acetate consists of ammonium (weak acid,  $\text{NH}_4^+$ ) and acetate (a weak base,  $\text{CH}_3\text{COO}^-$ ). Look up the  $K_a$  of ammonium and the  $K_b$  of acetate.
- **#107:** The question is asking for the ratio of the Novacaine (a base) compared to its conjugate acid. Thus you are looking for  $[\text{Nov}] / [\text{HNov}^+]$ . Write out the reaction of Nov and  $\text{H}_2\text{O}$ . Write out the  $K_b$  expression. Then, plug in the  $K_b$  and the  $[\text{OH}^-]$  and solve for the  $[\text{Nov}] / [\text{HNov}^+]$  ratio. (*No ICE table is needed. You have all equilibrium concentrations.*)
- **#120:**  $\text{Na}_2\text{CO}_3$  consists of  $\text{Na}^+$  and  $\text{CO}_3^{2-}$  ions.  $\text{CO}_3^{2-}$  is a base. It reacts with water in 2 stages. Write the reactions for each stage. (Stage 1: add on one  $\text{H}^+$ ; Stage 2: add on another  $\text{H}^+$ ) **Do ICE tables** and calculations for each stage. Look up any needed  $K_b$  values. You will need to solve for the concentrations of all species in the final solution.
- **#123:**  $\text{NaCN}$  consists of  $\text{Na}^+$  and  $\text{CN}^-$ .  $\text{CN}^-$  is a base. Write the reaction of  $\text{CN}^-$  in water and look up its  $K_b$ . You can figure out the  $[\text{OH}^-]$  from pH. **Use an ICE table** and fill in what you know and solve for the initial  $[\text{CN}^-]$ . Use that concentration to find grams of  $\text{NaCN}$ .

### **Acid-Base Reactions WKS**

**Optional Study Guide:** pp. 317-319 #1, 12, 14, 15, 18, 19, 26, 29, 31, 32, 34, 36;  
pp. 332-333 #8, 9, 12, 13, 14, 15, 27

### **Topics:**

- 1) **Bronsted Acids and Bases:** proton donor/proton acceptor, write reactions, label conjugate acid/base pairs
- 2) **Acids and Bases in water:**
  - Water is amphoteric, so it can act as an acid or a base depending on what it's with.
  - Use  $K_w$  expression to find  $[\text{H}_3\text{O}^+]$  or  $[\text{OH}^-]$ .  $K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$ .
  - Solve for pH or pOH.  $\text{pH} = -\log [\text{H}_3\text{O}^+]$   $\text{pOH} = -\log [\text{OH}^-]$ .  $\text{pH} + \text{pOH} = 14$
- 3) **Strengths of Acids and Bases:**
  - Know the strong acids (the other acids are weak); Hydroxides are strong bases.  $\text{NH}_3$  is a weak base.
  - Predict whether reactants or products are favored in an acid/base reaction by knowing strengths. (Stronger species always react to form weaker species.)
- 4) **Calculations involving weak acids or weak bases in water.** (Use  $K_a$  and  $K_b$ )
  - larger  $K_a$ : stronger acid (larger  $K_b$  = stronger base)
  - Given initial concentration of a weak acid, calculate equilibrium concentrations or  $K_a$  (Use ICE).
  - Find percent ionization of the acid.
  - Find pH from equilibrium  $[\text{H}^+]$  or  $[\text{OH}^-]$ .  
(Remember that  $[\text{H}^+] = K_a$  only in the special case that we did in the lab when one is at the "half-neutralization point" where  $[\text{HA}] = [\text{A}^-]$ .)
- 5) **Connection between the  $K_a$  of an acid and the  $K_b$  of its conjugate base.** ( $K_a)(K_b) = K_w$ 
  - The larger the  $K_a$  of the acid, the smaller the  $K_b$  of its conjugate base
- 6) **Calculations involving diprotic or triprotic acids.** (Do 2 or 3 ICE calculations to find equilibrium conc.)
- 7) **Molecular structure and strengths of acids:**
  - a) **Binary acids:** (H-X)
    1. **Going down a column:** Size of X gets larger (larger bond distance), so weaker H-X bond. Thus, more likely to lose  $\text{H}^+$ , so acids get stronger as go down column  
Ex:  $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$   
*smallest X-----> largest X*  
*Weakest acid -----> strongest acid*

2. **Going across a row:** Electronegativity of X increases across row, so bond polarity increases. This weakens the bond, so acid is more likely to lose  $H^+$ . Thus, acids get stronger.

Ex: *less acidic*  $CH_4 < NH_3 < H_2O < HF$  *more acidic*

- b) **Oxoacids and Carboxylic Acids:** Stronger the bond polarity, the stronger the acid. Ex:  $Z-O-H$
- When Z has a higher electronegativity, the Z pulls electrons more strongly from O. This makes the O-H bond more polar, so it is weakened. This makes the acid stronger.

- When there are more oxygens on Z (thus, making the oxidation number on Z higher), the stronger the electrons are pulled from Z. The O-H bond weakens and the acid is stronger.

8) **Salts and the hydrolysis of salts**

- salts are ionic compounds formed from reactions between an acid and a base.
- If either the cation or the anion of the salt reacts with water, the salt is said to undergo hydrolysis  
Ex: acidic cations:  $NH_4^+$  and some small highly charged metallic cations (but not alkali metal cations or  $Mg^{2+}$ ,  $Ca^{2+}$ ,  $Sr^{2+}$ ,  $Ba^{2+}$  or  $Ra^{2+}$ )  
Ex: basic anions: any conjugate bases of weak acids

9) **Lewis Acids and Bases**

- Lewis Acid: accepts pair of electrons. Examples include: any Bronsted acid; if has incomplete octet; if has cation that can form a complex ion; a non-metallic oxide.
- Lewis Base: donates a pair of electrons (Has a lone pair of electrons. Ex:  $NH_3$ ,  $H_2O$ )

10) **Acid Base Properties of oxides and hydroxides**

- Most metallic oxides are basic (All alkali and alkaline earth metal oxides are basic except for  $BeO$ ).
- Non-metallic oxides are acidic ( $CO_2$ ,  $SO_3$ ,  $N_2O_5$ ,  $P_4O_{10}$ ,  $Cl_2O_7$ )
- There are some amphoteric oxides. Ex:  $Al_2O_3$ ,  $BeO$
- Alkali and alkaline earth metal hydroxides are basic, but if it can form a complex, it is amphoteric.  
Examples of amphoteric hydroxides:  $Cr(OH)_3$ ,  $Zn(OH)_2$ ,  $Al(OH)_3$ ,  $Sn(OH)_2$ ,  $Be(OH)_2$
- Oxides act as anhydrous acids or bases and can react to form salts:
  - $2 NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l)$
  - $CaO(s) + 2 HBr(aq) \rightarrow CaBr_2(aq) + H_2O(l)$
  - $K_2O(s) + N_2O_5(g) \rightarrow 2 KNO_3(s)$

11) **Predicting products of reactions involving complex ions.**

- a) Reactions between a metal cation (Lewis acid) and a coordinating ligand (Lewis base), a molecule or ion that surrounds the metal ion
- *The number of ligands is usually  $2 \times$  the charge on the ion.*
  - *Common complex ion metals (Lewis acids):*  $Fe, Co, Ni, Cr, Cu, Zn, Ag, Al$
  - *Common ligands (Lewis bases):*  $NH_3, CN^-, OH^-, SCN^-, Cl^-$
- b) Complex formation can enable insoluble compounds [e.g.  $AgCl$ ,  $Cu(OH)_2$ ] to dissolve.
- c) Usually complex ions form if excess or concentrated amounts of a ligand are reacted.
- d) Hydration: when dissolved in water, all metals are surrounded by 6  $H_2O$  molecules
- e) Ammonia:
- Dilute  $NH_3$  will cause metal hydroxide to precipitate (A/B of  $NH_3$ ):  
 $Fe^{3+}(aq) + 3 NH_3(aq) + 3 H_2O \rightarrow Fe(OH)_3(s) + 3 NH_4^+$
  - Concentrated  $NH_3$  will cause ammine complex to form:  
 $Fe^{3+}(aq) + 6 NH_3(aq) \rightarrow [Fe(NH_3)_6]^{3+}(aq)$   
*All metals except  $Al^{3+}$  will complex with ammonia*