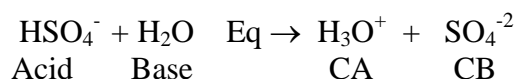
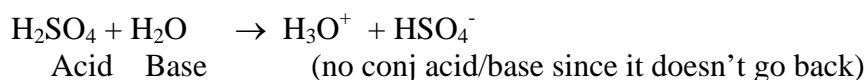
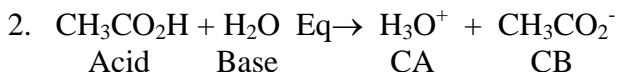


Answers*

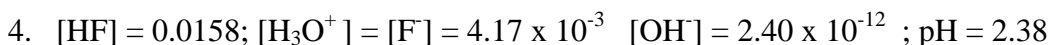
*I will use "Eq→" to represent an equilibrium arrow.

1. "Increasing acidity" means to start with the species that has the lowest dissociation of H_3O^+ and continue in order up to the strongest acid. Check the equilibrium constants if you are not sure how much each acid actually dissociates.
 $\text{HCO}_3^- < \text{H}_2\text{O} < \text{CH}_3\text{CO}_2\text{H} < \text{HCl}$



3. Compare equilibrium constants of the two acids and the two bases in each reaction. The reaction will only go from stronger to weaker. So if the stronger acid is on the reactant side, the reaction will go forward. If the stronger acid/base is on the product side, the reactant side will be favored.

First reaction favors products, second reaction favors reactants.



5. a) Carbon in center; 3 oxygens around carbon. Each H attached to a different O. both Hs are acidic (would be boxed).

b) N in center, 2 Os attached to N, one H attached to one O. This is an acidic H.

c) First 6 carbon atoms connected in a ring (hexagon) with alternating double bonds. One H attached to each C atom except for the 6th Carbon atom which attaches to the last C. That last C is attached by a single bond to one O and attached by a double bond to the second O. The last H is attached to the single-bonded O. This last H is the only acidic H.

d) Attach each C in a row, and attach the O at the end of the Carbon row and the H should be attached to the O at the end. The other Hs go around each C to give each a full octet. Only the H attached to the O atom is acidic.

6. 1) yes, this could be a buffer: a weak acid and its conjugate base.
 2) yes, this could be a buffer: a weak acid and its conjugate base.
 3) no, strong acid cannot be a buffer
 4) no, strong acid cannot be a buffer. Also, note that Na_2SO_4 wouldn't be the CB of H_2SO_4 if the strong acid could have a CB.

7. a) $\text{pH} = 2.05$ b) Na^+ , H_2SO_3 , HSO_3^- , SO_3^{2-} , H_3O^+ , H_2O , OH^-
 c) H_3O^+ (from HCl) + $\text{HSO}_3^- \rightarrow \text{H}_2\text{SO}_3 + \text{H}_2\text{O}$
 d) $\text{pH} = 1.71$ (I did this problem using moles first, and then converted all numbers to Molarity using .105 L as the final volume.)

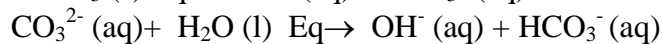
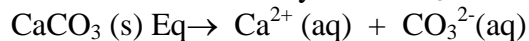
Important NOTE: Regarding answer to the Titration problem on the Quiz #4. and problem 8 below. The Titration problem on the quiz should have been done the way the problem #8 is explained; however, the problem was done because I was thinking of it more as pouring 2 beakers together (as in the worksheets given in class) rather than a true titration problem. This fact does not affect the grading in any way. In both cases, the first thing that needs to be considered is the neutralization of the acid or base.

The difference between a true titration and the pouring of two solutions together into one beaker is that the titration doesn't end (equivalence point) until all of the acid or base poured into the flask has reacted (gone to completion) whereas if you have two solutions in beakers, and if you don't have enough of one solution you will not attain the equivalence point. On the exam, I will give the amounts of both solutions so as to avoid this confusion since I wasn't able to clarify it in person during class.

8. At the equivalence point, all of the moles of whatever is in the flask are used up. There should only be conjugate acid or conjugate base left in the flask, and then if a hydrolysis reaction occurs, the pH is based on the dissociation of the conjugate acid or conjugate base. Again, I used moles in this problem: in each case, 0.1 moles of species in the flask will be neutralized by 0.1 moles of species in the buret, figure out the volume required of solution in the buret, and then convert all species to molarity before using the Equilibrium expression.
- 1) $\text{pH} = 7.00$ Strong acid, strong base gives a neutral salt and water.
 - 2) $\text{pH} = 4.71$
 - 3) $\text{pH} = 10.64$
9. The midpoint is half way to the equivalence point: half of the moles in the flask have reacted with the solution in the buret and there is no more solution from the buret because all of it has reacted; however, some new conjugate species has been introduced because of the titration reaction. Remember to consider the added volume.
- 1) $\text{pH} = 0.30$
 - 2) $\text{pH} = 9.25$
 - 3) $\text{pH} = 7.46$
10. The titration curve starts slow because there is either a dominant acid or base in the flask, and as titrant (solution in the buret) is added, it reacts and is used up. Once all of the original solution in the flask is completely used up (limiting reactant) then the pH rapidly changes to neutral and then the opposite solution (acid or base) dominates in the flask so the curve changes rapidly in the opposite direction.

11. 0.62 g/L BaF₂

12. 5.83×10^{-5} M solubility of CaCO₃. When dissolves:



In an acidic solution, the free hydronium ions would react with the hydroxide ions and this would cause some of the product of the carbonate ion hydrolysis reaction to be used up, and by Le Chatelier's principle, more product would be formed. So an acidic solution would make the calcium carbonate more soluble than it normally would be in pure water.

13. K for reaction = 3.75×10^{-6} . Rounded to one s.f. K = 4×10^{-6}

14.

- Sodium chloride would have no effect
- Magnesium chloride would decrease the solubility by introducing another source of Mg ion.
- NaOH would decrease the solubility by introducing another source of hydroxide ion.
- HCl would increase the solubility because it would react with the hydroxide ion produced by the dissociation equation, using up a product and causing more products to be produced.
- Water would have no effect at all on the solubility.