ChemQuest 72

The Common-Ion Effect

Name: ______________________

Date: ______________

Hour: _____

Information: The Common-Ion Effect

If you have something in common with someone else that means that the two of you have something that is the same about you. For example, if you and your friend both like cookie dough ice cream more than chocolate ice cream then you have something in common.

The common ion effect is similar to having something in common with a friend. If two substances both contain nitrate (NO_3^-) then they have that ion "in common." The common ion effect is based on LeChatelier's Principle. One way it occurs is when there is a weak acid or base in a solution along with a second soluble substance that has the same ion as the weak acid or base. You may not know exactly what the common ion effect is yet, but you will…

Critical Thinking Questions

1. Consider a solution of acetic acid ($HC_2H_3O_2$). Why will sodium acetate ($NaC_2H_3O_2$) cause the "common ion effect," but sodium chloride (NaCl) will not?

   Acetate, $C_2H_3O_2^-$, is present in both acetic acid and sodium acetate.

2. Let's say we have a solution of the weak base ammonia (NH_3).

   a) Complete the equilibrium equation below:

   $$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$

   b) Will NH_4Cl or will KNO_3 cause a "common ion effect?" EXPLAIN.

   NH_4Cl because NH_4^+ is an ion "in common" with the equilibrium

3. Let's consider a solution of the weak acid acetic acid, $HC_2H_3O_2$. Calculate the pH of a 0.075 M solution of acetic acid. The equilibrium can be written like this:

   $$HC_2H_3O_2 + H_2O \leftrightarrow C_2H_3O_2^- + H_3O^+$$

   $$1.7x10^{-5} = \frac{[x][x]}{[0.075 - x]} \rightarrow 1.7x10^{-5} = \frac{x^2}{0.075} \rightarrow x = 0.0011 = [H^+]$$

   pH = -log(0.0011) = **2.95**

4. Questions like the previous question we have done before. We assume that the acetic acid is in pure water. The problem becomes a little different if the water has a common ion dissolved in it. For example… Calculate the pH of a solution that is 0.075 M acetic acid AND 0.033 M sodium acetate. The following table may be helpful:

	$HC_2H_3O_2$ + H_2O ←→ $C_2H_3O_2^-$ + H_3O^+		
Initial:	0.075 M	0.033 M	0 M
Change:	-x	+x	+x
Equilibrium:	0.075 – x	0.033 + x	x

$1.7x10^{-5} = \frac{(0.033+x)(x)}{0.075-x} \rightarrow 1.7x10^{-5} = \frac{(0.033)(x)}{0.075} \rightarrow x = 3.86x10^{-5}M$= [H$^+$]

pH = -log(3.86x10^{-5}) = **4.41**

5. In question 4 the initial concentration of all of the products was not zero. Why not?

   There was some sodium acetate and acetic acid starting out in the solution.

6. Calculate the pH of a solution that is 0.052 M NH_3 and 0.028 M NH_4Cl.

	NH_3 + H_2O ←→ NH_4^+ + OH^-		
Initial:	0.052 M	0.028 M	0 M
Change:	-x	+x	+x
Equilibrium:	0.052 – x	0.028 + x	x

$$1.8x10^{-5} = \frac{(0.028+x)(x)}{0.052-x} \rightarrow 1.8x10^{-5} = \frac{(0.028)(x)}{0.052} \rightarrow x = 3.34x10^{-5}M = [OH]$$

pOH = -log(3.34x10^{-5}) = 4.48 → pH = 14 – 4.48 = **9.52**

7. Calculate the pH of a solution that is 0.15 M acetic acid AND 0.023 M calcium acetate, $Ca(C_2H_3O_2)_2$.

	$HC_2H_3O_2$ + H_2O ←→ $C_2H_3O_2^-$ + H_3O^+		
Initial:	0.015 M	0.023 M	0 M
Change:	-x	+x	+x
Equilibrium:	0.015 – x	0.023 + x	x

$1.7x10^{-5} = \frac{(0.023+x)(x)}{0.015-x} \rightarrow 1.7x10^{-5} = \frac{(0.023)(x)}{0.015} \rightarrow x = 1.11x10^{-5}M$= [H$^+$]

pH = -log(1.11x10^{-5}) = **4.96**

ChemQuest 73

Mixing Acids and Bases

Name: ______________________________

Date: _______________

Hour: _____

Information: Dilutions

When water is added to a solution, the concentration decreases. It is often desirable to be able to calculate the concentration of solutions that have been diluted. For doing this, keep in mind that molarity is equal to the moles of solute divided by the total liters of solution. Therefore, the following equations are valid where M is molarity, $L_{solution}$ is liters of solution and mol_{solute} is the moles of solute:

Equation #1: $$M = \frac{mol_{solute}}{L_{solution}}$$

Equation #2: $$mol_{solute} = (M)(L_{solution})$$

Critical Thinking Questions

For the following questions, assume that liquid volumes are additive.

1. A certain solution is prepared by dissolving 4.0 moles of salt (NaCl) in enough water to make 400 mL of solution. Later, the solution was diluted with enough water so that the volume of the solution was 650 mL. Calculate the molarity of the solution before and after dilution.

   M_{before} = 4.0mol ÷ 0.400L = 10.0 M

   M_{after} = 4.0mol ÷ 0.650L = 6.15 M

2. A 6.0 M solution of salt has a volume of 500 mL. Later, 275 mL of water is added. Confirm that the molarity of the resulting is approximately 3.87 M. (Hint: first find the moles of salt present before the additional 275 mL of water was added by using equation #2 and then find the new molarity using equation #1.)

   moles of salt = 6.0M x 0.5L = 3.0mol

   M = 3.0mol ÷ (0.500 + 0.275)L = 3.87 M

3. Calculate the molarity of the solution formed by taking 350 mL of 2.25 M HCl and adding 420 mL of water.

   mol solute = 2.25M x 0.350L = 0.788 mol

   M = 0.788mol ÷ (0.350 + 0.420) = 1.02 M

4. Imagine that you have 300 mL of a stock solution of 2.8 M HCl solution. Describe how I could prepare 50 mL of 1.2 M HCl solution by using some of stock solution and diluting it with water. Be specific.

   mol HCl needed = 1.2M x 0.050L = 0.060 mol

   (2.8M)(*x*L) = 0.060 → *x* = 0.0214 L = 21.4 mL

   Take 21.4 mL of the stock solution and dilute it to 50 mL.

Information: Mixing Strong Acids and Strong Bases

Usually when we speak of "salt" we mean table salt, which is sodium chloride (NaCl). A salt is a general term for an ionic compound formed when an acid and a base mix. Whenever an acid and a base react, water and a salt are formed. For example consider the following reactions in which nitric acid (HNO_3) and hydrochloric acid (HCl) react with the base sodium hydroxide (NaOH):

$$HCl + NaOH \rightarrow NaCl + H_2O$$
$$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$$

Notice that in each reaction water and a salt (sodium chloride one reaction and sodium nitrate in another) were formed.

If equal moles of strong acid and strong base react, then they neutralize each other and form a solution of salt water. If there are more moles of acid than base then the resulting solution will be acidic. If there are more moles of base than acid, then the resulting solution will be basic.

Critical Thinking Questions

5. Consider the reaction of 2.5 moles of hydrochloric acid with 1.9 moles of sodium hydroxide.
   a) If this reaction took place in 2.0 L of solution, what is the concentration of leftover hydrochloric acid after the reaction?

      mol leftover = 2.5 – 1.9 = 0.6 mol ÷ 2.0L = 0.3 M

   b) From your answer to part a, verify that the pH of the solution after the reaction is approximately 0.52.

      pH = -log(0.3) = 0.52

6. Question 5 could be rewritten like this: *Consider the reaction of 1.0 L of _2.5_ M hydrochloric acid with 1.0 L of _1.9_ M sodium hydroxide.* Fill in the blanks with the appropriate numbers indicating the molarity.

7. 320 mL of 3.1 M HCl is mixed with 240 mL of 4.1 M NaOH. Use the following steps to find the pH of the resulting solution.
   a) Calculate the moles of HCl and the moles of NaOH that are reacting using Equation #2.
      mol HCl = (3.1M)(0.320L) = 0.992 mol
      mol NaOH = (4.1M)(0.240L) = 0.984 mol

   b) Find out which substance is left over and find out how many moles of this substance is left over.
      mol HCl leftover = 0.992 – 0.984 = 0.008 mol

   c) Divide the moles left over by the total volume in liters to get the concentration of the left over substance.
      0.008 mol ÷ (0.320+0.240) = 0.014 M

   d) Your answer to part c is also the concentration of H^+ (if the acid is left over) or the concentration of OH^- (if the base is left over). From this information calculate the pH of the solution. You should get approximately 1.85 for your answer.
      $[H^+]$ = [HCl] = 0.014M
      pH = -log(0.014) = 1.85

8. Calculate the pH of a solution formed by mixing 450 mL of 0.79 M HCl with 430 mL of 1.2 M NaOH. Hint: this is very similar to question 7.
   13.26
   mol HCl = (0.79M)(0.450L) = 0.356mol; mol NaOH = (1.2M)(0.430L) = 0.516mol
   mol NaOH leftover = 0.516 – 0.356 = 0.16 mol ÷ (0.45+0.43) = 0.182 M = $[OH^-]$
   pOH = -log(0.182) = 0.74 → pH = 14 – 0.74 = 13.26

9. Calculate the pH of a solution formed by mixing 820 mL of 1.2 M HNO_3 with 700 mL of 0.9 M NaOH.
   0.633
   mol HNO_3 = (1.2)(0.820) = 0.984 mol; mol NaOH = (0.9)(0.700) = 0.630 mol
   mol HNO_3 leftover = 0.984 – 0.630 = 0.354 mol ÷ (0.700+0.820) = 0.233 M = $[H^+]$
   pH = -log(.233) = 0.633

10. Consider 400 mL of a 2.5 M HCl solution. How many milliliters of 1.25 M NaOH will be needed to neutralize the HCl?
    800 mL
    mol NaOH must equal mol HCl = (2.5M)(0.400L) = 1.00 mol
    L = mol ÷ M = 1.00 ÷ 1.25 = 0.800 L → 800 mL

ChemQuest 74

Acidic Properties of SALTS

Name: ______________________________
Date: ________________
Hour: _____

Information: Review Definitions of Acids and Bases

Recall the definitions of acids and bases according to Bronsted-Lowry:

1) acid: substance that donates a proton, H^+, in a reaction
2) base: substance that accepts a proton, H^+, in a reaction

Recall also, that we have seen that it is possible for ions to act as acids or bases. We will now examine the acidity or basicity of ions.

Critical Thinking Questions

1. If the chloride ion (Cl^-) or the cyanide ion (CN^-) were dissolved in water, would you expect the ions act as acids or bases? Ask yourself: could it accept a H^+? Does it have a H^+ to give away? Explain.

   Since they don't have a H^+ to give away, we can guess that they would act like bases and receive an H^+.

2. If the ammonium ion (NH_4^+) were dissolved in water, would you expect it to act as an acid or base? Explain.

   We can expect NH_4^+ to act like an acid since it appears to have at least one H^+ to give away.

3. Although not all ions act as acids or bases, we can generalize from the above two questions that **IF** an ion acts as an acid or a base the following rules will be observed:

   a) When in water, positive ions may act like ____acids_______.

   b) When in water, negative ions may act like ____bases_______.

Information: Determining Whether an Ion Affects the pH

It was mentioned above that not all ions will act as an acid or a base. Some, in fact, will not affect the pH of a solution at all. We are now going to look at what determines whether an ion affects the pH or not.

Consider the chloride ion. If it affects the pH, it will react with water in the following way:

Equation 1: $Cl^- + H_2O \leftrightarrow HCl + OH^-$

Now consider the cyanide ion. If it affects the pH, it will react with water in the following way:

Equation 2: $CN^- + H_2O \leftrightarrow HCN + OH^-$

Hopefully, these equations confirm what you wrote for the questions above: negative ions, if they affect the pH, will act like bases because they produce OH^- ions. In fact, the cyanide ion (CN^-) is the conjugate base of the weak acid HCN. The base ionization constants (K_b) for equations 1 and 2 are shown below:

$$K_{b1} = \frac{[HCl][OH^-]}{[Cl^-]} \qquad K_{b2} = \frac{[HCN][OH^-]}{[CN^-]}$$

Critical Thinking Questions

4. The expression for K_{b1} assumes that HCl exists in the solution without breaking into ions completely. Similarly, the expression for K_{b2} assumes that HCN exists in solution without breaking into ions completely. However, one of these—HCl or HCN—completely breaks up into ions when dissolved. Which substance breaks up completely into ions? Explain.

   HCl completely breaks up into ions and exists as H^+ and Cl^- in water. We know this because HCl is a strong acid.

5. One of the ions—Cl^- or CN^-—does not affect the pH because it will not react as depicted in equation 1 or 2 above. Given your answer to question 4, which ion do you think will not affect the pH of a solution?

   Cl^- will not affect the pH because it will not accept a H^+ to form HCl.

6. Complete the following sentence: The negative ions formed from a _____strong_________ (strong/weak) acid will not affect the pH, but the negative ion formed from a _____weak________ (strong/weak) acid will raise the pH and act like a(n) _____base_________ (acid/base).

7. Similar reasoning applies to the positive ions formed from strong and weak bases. Using question six as a pattern, complete the following sentence:

   The positive ions formed from a _______strong_______ (strong/weak) base will not affect the pH, but the positive ion formed from a _____weak_______ (strong/weak) base will lower the pH and act like a(n) ______acid_______ (acid/base).

8. Determine whether each of the following will act like an acid (A) or a base (B). If the ion will not affect the pH, place an X in the blank.

__X__ a) Na^+ (from the strong base NaOH) __X__ b) NO_3^- (from the strong acid HNO_3)

__A__ c) NH_4^+ (from the weak base NH_3) __X__ d) K^+ (from the strong base KOH)

__B__ e) NO_2^- (from the weak acid HNO_2) __B_ f) $C_2H_3O_2^-$ (from the weak acid $HC_2H_3O_2$)

9. Write the chemical equation for acetic acid dissociating.
$HC_2H_3O_2 \leftrightarrow H^+ + C_2H_3O_2^-$ **or** $HC_2H_3O_2 + H_2O \leftrightarrow H_3O^+ + C_2H_3O_2^-$

10. From the K_a of acetic acid ($1.7x10^{-5}$), calculate the K_b for the acetate ion.
$K_b = 1x10^{-14} \div 1.7x10^{-5} = 5.88x10^{-10}$

11. When ammonia (NH_3) reacts with water, the ammonium ion is formed (NH_4^+). Write the chemical equation for this process.
$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

12. Given the K_b for ammonia ($1.8x10^{-5}$), calculate the K_a for the ammonium ion.
$K_a = 1x10^{-14} \div 1.8x10^{-5} = 5.56x10^{-10}$

Information: Salts

Salts are ionic compounds. They are similar to strong acids and bases because they dissociate completely in water. However, not all salts affect the pH. Consider the salt sodium acetate ($NaC_2H_3O_2$). If you place 2.0 moles of sodium acetate in 1.0 L of water, it will dissociate completely as follows:

$$NaC_2H_3O_2 \rightarrow Na^+ + C_2H_3O_2^-$$

Each mole of $NaC_2H_3O_2$ breaks up into a mole of Na^+ and a mole of $C_2H_3O_2^-$. Therefore, because you started out with 2.0 moles of $NaC_2H_3O_2$ in 1.0 L of water, the concentration of Na^+ is 2.0 M and the concentration of $C_2H_3O_2^-$ is 2.0 M. We have seen already that Na^+ will not affect the pH since it is derived from the strong base NaOH. $C_2H_3O_2^-$ (derived from the weak acid $HC_2H_3O_2$) will affect the pH by producing OH^- ions: $C_2H_3O_2^- + H_2O \leftrightarrow HC_2H_3O_2 + OH^-$.

Critical Thinking Questions

13. Determine whether each of the following salts will act like an acid (A) or a base (B). If the salt will not affect the pH, place an X in the blank.

__X__ a) NaCl __A__ b) NH_4Cl __B__ c) NaCN __X__ d) KNO_3

14. What is the pH of a 0.45 M solution of $NaC_2H_3O_2$? (Follow the following steps.)

a) Will Na^+ or will $C_2H_3O_2^-$ affect the pH?

$C_2H_3O_2^-$

b) Hopefully your answer to part a is $C_2H_3O_2^-$. Find the concentration of $C_2H_3O_2^-$ in the solution. (Note: remember that salts dissociate completely.)

$[C_2H_3O_2^-] = [NaC_2H_3O_2] = 0.45$ M

c) Now that you know $[C_2H_3O_2^-]_{initial}$ complete the necessary equilibrium calculations using the K_b for $C_2H_3O_2^-$ (see question 10) to calculate $[OH^-]$.

$$K_b = 5.88\text{x}10^{-10} = \frac{[HC_2H_3O_2][OH^-]}{C_2H_3O_2^-} = \frac{(x)(x)}{(0.45 - x)} = \frac{x^2}{0.45} \rightarrow x = 1.63\text{x}10^{-5}\text{ M}$$

d) Calculate the pOH of the solution.

pOH = $-\log(1.63\text{x}10^{-5}) = 4.79$

e) From the pOH, verify that the pH is 9.2.

pH = $14 - 4.79 = 9.21 \sim 9.2$

15. Find the pH of a 0.32 M solution of NH_4NO_3.

4.88

NH_4^+ affects pH: $NH_4^+ \leftrightarrow H^+ + NH_3$; $K_a = 5.56\text{x}10^{-10}$ (see question 12)

$$K_a = 5.56\text{x}10^{-10} = \frac{[H^+][NH_3]}{[NH_4^+]} = \frac{(x)(x)}{0.32-x} = \frac{x^2}{0.32} \rightarrow x = 1.33\text{x}10^{-5} = [H^+]$$

pH = $-\log(1.33\text{x}10^{-5}) = 4.88$

ChemQuest 75

Titration

Name: ______________________________
Date: ______________
Hour: _____

Information: Titrating a Strong Acid and a Strong Base

A titration is an experimental means for determining the concentration of an unknown acid or a base. If you do not know the concentration of a certain acid you can determine it experimentally using a titration. A titration depends on an indicator that will change color when the equivalence point is reached. (The equivalence point is when the moles of acid equal the moles of base in a solution.) The following questions will walk you through a hypothetical titration experiment…

Critical Thinking Questions

1. Students attempted to determine the concentration of an unlabeled bottle of hydrochloric acid (HCl) by titration. First, they poured 40 mL of the acid solution into a flask. Then they added a few drops of phenolphthalein to serve as an indicator. The students found that it took 28 mL of 0.075 M NaOH to make the phenolphthalein change color.
   a) How many moles of NaOH were added to the HCl by the students?

      (0.075 M)(0.028 L) = 0.0021 mol

   b) Remember that at the equivalence point, the moles of base equal the moles of acid. You now know the moles of acid that were in the original 40 mL. Calculate the concentration of the acid.

      0.0021 mol ÷ 0.040 L = 0.0525 M

2. In a titration experiment it was determined that it took 42 mL of 0.125 M HCl to neutralize 120 mL of KOH. Calculate the molarity of the KOH.

   (0.125 M)(0.042 L) = 0.00525 mol

   0.00525 mol ÷ 0.120 L = 0.044 M

3. 50 mL of 0.25 M NaOH was titrated with 0.40 M HNO_3. How many mL of the HNO_3 were needed before reaching the equivalence point?

   (0.25 M)(0.050 L) = 0.0125 mol
   0.0125 mol ÷ 0.40 M = 0.03125 L → 31.25 mL

Information: Weak Acids or Bases

So far, the titration calculations you have done all involved strong acids and bases. What if one of the substances is weak? The following example will help you to learn about such titrations.

Question: **What is the pH at the equivalence point when 45 mL of 0.095 M ethylamine is titrated with 0.12 M HCl?**

Step 1: Find out what substances are strong and weak.

- If you have a strong acid and a strong base, the pH is 7 at the equivalence point.
- If you have a strong acid and a weak base, the pH will be acidic at the equivalence point
- If you have a weak acid and a strong base, the pH will be basic at the equivalence point.

Look at your K_b and K_a charts. HCl does not appear on any of the charts and therefore it is strong. Ethylamine appears on the K_b chart so it is weak. Its K_b is $4.7x10^{-4}$.

Step 2: If one of the substances in the question has a K_b then you need to find the K_a. If it has a K_a then you need to find the K_b. In this example, we need to find the K_a that corresponds to the K_b of $4.7x10^{-4}$.

$$K_a = \frac{1x10^{-14}}{4.7x10^{-4}} = 2.13x10^{-11} = \frac{x^2}{y-x}$$

x is the [H^+] since we are using K_a. y is the concentration of the conjugate acid. We can typically ignore the "–x" in the denominator since the K_a or K_b is sufficiently small.

Step 3: Find the moles.

Moles = (M)(L) = (0.095 M)(0.045 L) = 0.004275 mol

Step 4: Find Liters of the missing substance. In this problem we don't have the liters of HCl

$$M = \frac{mol}{L} \quad \rightarrow \quad 0.12M = \frac{0.004275mol}{L} \quad \rightarrow \quad L = \frac{0.004275mol}{0.12M} = 0.0356L$$

Step 5: Find y

$$y = \frac{mol}{total\,liters} = \frac{0.004275mol}{(0.045+0.0356)} = 0.0532M$$

Step 6: Find x from Step 2.

$$2.13x10^{-11} = \frac{x^2}{y-x} \quad \rightarrow \quad 2.13x10^{-11} = \frac{x^2}{0.0532} \quad \rightarrow \quad x = 1.06x10^{-6}$$

x is the [H^+] since we are using K_a. Therefore pH = -log($1.06x10^{-6}$) = ***5.99***

Critical Thinking Questions

4. Consider titrating 0.35 L of 0.20 M $HC_2H_3O_2$ with 0.10 M NaOH.
   a) How many L of NaOH will be required? (Hint: see step 4 above.)

      (0.35 L)(0.20 M) = 0.07 mol
      0.07 mol ÷ 0.10 M = **0.70 L**

   b) Calculate the pH at the equivalence point.

$$K_b = \frac{1x10^{-14}}{1.7x10^{-5}} = 5.88x10^{-10} = \frac{x^2}{0.0667} \rightarrow x = 6.26x10^{-6}M = [OH]$$

$$\frac{(0.35)(0.20)}{(0.7 + 0.35)} = 0.0667$$

      pOH = -log($6.26x10^{-6}$) = 5.20 → pH = 14 – 5.20 = **8.80**

5. Consider 0.50 L of 0.25 M NH_3 that is titrated with 0.20 M HCl. Find the pH at the equivalence point.

   L_{HCl} = (0.50L)(0.25M) ÷0.20 M = 0.625 L

$$K_a = \frac{1x10^{-14}}{1.8x10^{-5}} = 5.56x10^{-10} = \frac{x^2}{0.111} \rightarrow x = 7.856x10^{-6}M = [H]$$

$$\frac{(0.5)(0.25)}{(0.5 + 0.625)} = 0.111$$

   pH = -log($7.856x10^{-6}$) = **5.10**

6. A 1.49 g sample of propionic acid ($HC_3H_5O_2$) was dissolved in water to give 45.0 mL of solution. The solution was then titrated with 0.210 M NaOH. What was the pH of the solution at the equivalence point? The K_a of propionic acid is $1.34x10^{-5}$.

1.49g ÷ 74 g/mol = 0.020 mol
0.020 mol ÷ 0.210 M = 0.095 L

$$K_b = \frac{1x10^{-14}}{1.34x10^{-5}} = 7.46x10^{-10} = \frac{x^2}{0.0367} \rightarrow x = 5.23x10^{-6}M = [OH]$$

$$\frac{(0.020)}{(0.095 + 0.45)} = 0.0367$$

pOH = -log($5.23x10^{-6}$) = 5.28 → pH = 14 – 5.20 = **8.72**

7. Find the pH when 48 mL of 0.092 M ethylamine is titrated to the equivalence point with 0.15 M HCl. K_b for ethylamine is $5.6x10^{-4}$.

(0.48 L)(0.092 M) = 0.0442 mol
0.0442 mol ÷ 0.15 M = 0.294 L

$$K_a = \frac{1x10^{-14}}{5.6x10^{-4}} = 1.79x10^{-11} = \frac{x^2}{0.0571} \rightarrow x = 1.011x10^{-6}M = [H]$$

$$\frac{(0.48)(0.092)}{(0.48 + 0.294)} = 0.0571$$

pH = -log($1.011x10^{-6}$) = **6.00**

ChemQuest 76

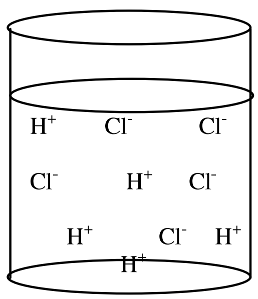
Buffered Solutions

Name: ____________________________
Date: ________________
Hour: _____

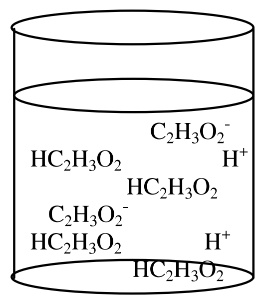
Information: What is a Buffer?

At times it is very helpful to have a solution that will resist changes in pH. For example, our bodies need to maintain a certain pH for our survival. Various organs and body processes do not work if the pH changes. Thus, our blood is a special solution that maintains a pH of around 7.4 even if small amounts of acid or base are added to it. A solution that resists changes in pH is called a "buffered solution."

Beaker A: HCl (aq)



Beaker B: $HC_2H_3O_2$ (aq)



Critical Thinking Questions

1. Which beaker--A or B--contains a strong acid? If you had not memorized which acids are strong, what is it about the contents of the beaker, as shown in the diagram, that tells you it is strong?

   A, because it exists in solution completely as ions.

2. Which of the following ions are capable of acting like a base?

   A) NO_3^- B) F^- C) CN^- D) Cl^-

3. Complete the following chemical equation of an ion acting like a base.

   $ClO^- + H^+ \rightarrow HClO$

4. If a base such as NaOH were added to Beaker A or to Beaker B, the H^+ ions present would react with the OH^- ions to neutralize them. Complete the equation for this process.

   $H^+ + OH^- \rightarrow H_2O$

5. As we see from question 4, both beakers--A and B--have ions that can neutralize a base. Only one of the beakers, however, has negative ions that are capable of neutralizing an acid.

   a) Which ion from the above beaker is capable of neutralizing an acid?

   $C_2H_3O_2^-$

   b) Complete the equation below showing the neutralization.

   $H^+ + \underline{C_2H_3O_2^-} \rightarrow HC_2H_3O_2$

   enter the ion in the blank

6. A buffer needs to be able to neutralize BOTH an acid and a base. Which substances do you think would be better to use in creating buffered solution--strong acids and bases OR weak acids and bases? Why?

   Weak acids and bases because they can neutralize both acids and bases as we found in question 5.

Information: Creating a Buffer

The most effective buffers are composed of a weak acid or base along with its conjugate base or acid so that there will be plenty of ions to neutralize either acids or bases. For example, a solution of the weak acid acetic acid ($HC_2H_3O_2$) and sodium acetate ($NaC_2H_3O_2$) would make an effective buffer. The acetate ion is the conjugate base of the weak acid, acetic acid. (Recall that the sodium ion, of course, does not affect the pH.)

Also, a solution of the weak base ammonia (NH_3) and ammonium chloride (NH_4Cl) would make a good buffer. The ammonium ion (NH_4^+) is the conjugate acid of the weak base, ammonia. (Recall that chloride ions will not affect the pH.)

Critical Thinking Questions

7. As just mentioned in the information section, neither Na^+ ions nor Cl^- ions affect the pH of a solution. Why is this?

   Neither can bond with OH^- or H^+.

8. Consider a generic weak acid equilibrium: $HA \leftrightarrow H^+ + A^-$

   a) Write the K_a expression.

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

   b) Rearrange the expression from part (a) to solve for $[H^+]$. Hint: K_a will appear on the right-hand side in the following expression:

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

   c) If you take the -log of both sides you will have an equation for the pH. Write that equation:

   $$pH = -\log[H^+] = -log\frac{K_a[HA]}{[A^-]}$$

9. pH stands for -log[H^+]. pOH stands for the -log of [OH^-]. What do you think pK_a stands for?

   -logK_a

Information: Henderson-Hasselbalch Equation

At the end of question 8c you should have had an equation similar to the following:

$$\text{pH} = -\log\left(K_a \frac{[HX]}{[X^-]}\right)$$

We can simplify the equation further:

To get this subtraction sign to become an addition sign we can use the properties of logarithms and take the reciprocal of the fraction here

$$\text{pH} = -\log K_a - \log\left(\frac{[HX]}{[X^-]}\right)$$

The final version of the equation is called the Henderson-Hasselbalch Equation and looks like this:

$$\text{pH} = \text{p}K_a + \log\left(\frac{[X^-]}{[HX]}\right) \qquad \text{OR} \qquad \text{pH} = \text{p}K_a + \log\left(\frac{[base]}{[acid]}\right)$$

Critical Thinking Questions

10. Why do you think [X^-] can be replaced with the term [base]?
    The X^- ion is a basic ion.

11. Use the Henderson-Hasselbalch Equation to calculate the pH of a buffer composed of 0.24 M lactic acid ($HC_3H_5O_3$) and 0.22 M potassium lactate ($KC_3H_5O_3$). The K_a of lactic acid is $1.4x10^{-4}$.

$$pH = -\log(1.4x10^{-4}) + \ log\frac{0.22}{0.24} = 3.82$$

12. A solution was made by dissolving 83.5 g of lactic acid ($HC_3H_5O_3$) in 620 mL of water. Another solution was made by dissolving 75 g of sodium lactate ($NaC_3H_5O_3$) in 550 mL of water. The two solutions were then mixed together to form a buffer. What is the pH of the buffer?
    [acid]=(83.5g÷90g/mol)÷0.620L = 1.496 M
    [base]=(75g÷112g/mol)÷0.550L = 1.218 M

$$pH = -\log(1.4x10^{-4}) + \ log\frac{1.218}{1.496} = 3.76$$

13. A buffer is most effective at a pH range near its pK_a. Use the Henderson-Hasselbalch Equation to offer an explanation why.

    The pH is calculated based on the pK_a value.

Information: Adding a Strong Acid or Base to a Buffer

A buffer is most effective at a pH near its pK_a. Furthermore, a buffer is most effective when there are nearly equal concentrations of conjugate acid-base pairs. That is, when [base]/[acid] is near a value of one the buffer will work the best.

Adding an acid or base to a buffer solution will change the pH minimally as long as the pH is near the pKa of the solution.

Critical Thinking Questions

14. A certain buffer solution had a concentration of NH_3 of 0.35 M and a concentration of NH_4Cl of 0.41 M. Show the calculations that prove that the pH of the buffer is 9.19.

$$K_a = 1x10^{-14} \div 1.8x10^{-5} = 5.56x10^{-10}$$

$$pH = -\log(5.56x10^{-10}) + \log\frac{0.35}{0.41} = 9.19$$

15. 420 mL of the buffer from the previous question was placed in a beaker. Calculate the pH after 45 mL of 0.25 M NaOH was added to the solution. Follow these steps…
    a) Calculate the initial *moles* of base and of acid in the buffer:

    moles base = mol_{NH3}= (0.420L)(0.35M) = ___0.147___ mol

    moles acid = mol_{NH4+}= (0.420L)(0.41M) = ___0.172___ mol

    b) Calculate the moles of NaOH added using the volume of NaOH and the molarity of NaOH.
    (0.045L)(0.25M) = 0.0113 mol

    c) The substance (NaOH) added to the buffer is a strong ___base___.
    acid OR base

    d) Since NaOH is a base you add the moles of NaOH from part (b) to the moles of base calculated in part (a). Then take the moles of acid from part (a) and *subtract* the moles of NaOH found in part (b).

    new moles base = 0.147 + 0.0113 = 0.158 mol

    new moles acid = 0.172 – 0.0113 = 0.161 mol

    e) Find the new concentrations of base and acid.

    $$[base] = \frac{new\ moles\ from\ part\ (d)}{total\ liters} = \frac{0.158}{(0.420+0.045)} = 0.340\ M$$

    [acid] = 0.161 ÷ (0.420 + 0.045) = 0.346 M

    f) Use the values from part (e) in the Henderson-Hasselbalch Equation to calculate the pH of the buffer after the NaOH has been added.

    $$pH = -\log(5.56x10^{-10}) + \log\frac{0.340}{0.346} = 9.25$$

16. A buffer contains 0.32 mol of acetic acid and 0.35 mol of sodium acetate in 1.25 L of solution. What is the pH of the buffer before and after the addition of 125 mL of 0.75 M HCl?

BEFORE: [base] = 0.35mol ÷ 1.25L = 0.28 M [acid] = 0.32mol ÷ 1.25L = 0.256 M

$$pH = -\log(1.7x10^{-5}) + \ log\frac{0.280}{0.256} = 4.808$$

AFTER: mol acid added: (0.125L)(0.75M) = 0.09375 mol
[base] = (0.35 – 0.09375mol) ÷ (1.25+0.125L) = 0.186 M
[acid] = (0.32 + 0.09375mol) ÷ (1.25+0.125L) = 0.301 M

$$pH = -\log(1.7x10^{-5}) + \ log\frac{0.186}{0.301} = 4.56$$

17. A buffer is made by dissolving 8.2 g of propionic acid ($HC_3H_5O_2$) and and 10.0 g of sodium propionate ($Na\ C_3H_5O_2$) in 975 mL of water. What is the pH before and after 0.9 g of NaOH is added to the solution? K_a for propionic acid is $1.34x10^{-5}$.

BEFORE: mol acid: 8.2g÷74g/mol = 0.1108 mol mol base: 10.0g÷96g/mol = 0.1042 mol
[acid] = 0.1108mol ÷ 0.975L = 0.1136 M [base] = 0.1042mol÷0.975L = 0.1069 M

$$pH = -\log(1.34x10^{-5}) + \ log\frac{0.1069}{0.1136} = 5.81$$

AFTER: mol base added: 0.9g ÷ 40g/mol = 0.0225 mol
mol acid: 0.1108 – 0.0225 = 0.0883 mol mol base: 0.1042 + 0.0225 = 0.1267 mol
[acid] = 0.0883mol ÷ 0.975L = 0.0906 M [base] = 0.1267mol ÷ 0.975L = 0.1299 M

$$pH = -\log(1.34x10^{-5}) + \ log\frac{0.1299}{0.0906} = 5.03$$