Buffer Solutions

1. What is a buffer solution?

   A buffer solution is one that is able to resist (or minimize) changes in pH. In order to identify if you have a buffer solution you are looking for a solution that contains a weak acid/base conjugate pair in solution.

   For example:

   If a solution contained both HOCl and NaOCl – it would be a buffer as it meets both criteria – a weak acid (HOCl) and it's conjugate base (^{-}OCl). Keep in mind that the base will be added in the form of a conjugate salt, in this cae NaOCl. The Na^{+} is just a spectator ion and its presence does not make or break the formation of the buffer.

2. When solving problems for a buffer solution you can bypass the ICE chart and use the Henderson-Hasselbalch equation instead. What is the Henderson-Hasselbalch equation?

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

   Remember that "p" means "- log of …"

3. Identify the buffer?

   When doing these problems you are looking to see

   1. Are all substances in solution weak?
   2. Is there a conjugate pair?

a. HCl and $NaClO_4$

Answer the questions

1. No. HCl is a strong acid.
This automatically disqualifies this solution as being a buffer.

b. HClO and $LiClO_2$

1. Yes. Both HClO and ClO_2^- are weak.
2. No. The conjugate base of HClO is ClO^- not ClO_2^-.
This disqualifies this solution as being a buffer.

c. HF and KF

1. Yes. Both HF and F^- are weak.
2. Yes. The conjugate base of HF is F^-

Because both criteria are met, this solution is a buffer.

d. CH_3NH_2 and CH_3NH_3Cl

1. Yes. Both CH_3NH_2 and $CH_3NH_3^+$ are weak.
2 Yes. The conjugate acid of CH_3NH_2 is $CH_3NH_3^+$

Because both criteria are met, this solution is a buffer.

4. Determine the pH of the following solutions

a. 1.00 L solution of 1.00M HNO_2 and 1.50M $NaNO_2$.
($K_a = 4 \times 10^{-4}$)

This is a buffer solution. HNO_2 is a weak acid and the salt contains its conjugate base NO_2^-. This means that we can is the Henderson-Hasselbalch equation to solve for the pH.

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

$$pH = -\log(4 \times 10^{-4}) + \log \frac{(1.50M)}{(1.00M)}$$

$$\boxed{pH = 3.57}$$

b. 25.0g of NH_3 and 40.0g NH_4NO_3 in 501.0 mL of water.
($K_a = 5.6 \times 10^{-10}$)

This is a buffer solution. NH_3 is a weak base and the salt contains its conjugate acid NH_4^+. This means that we can use the Henderson-Hasselbalch equation to solve for the pH. We will just need to convert to the correct units first.

$$25.0\ g\ NH_3 \quad \frac{1\ mol}{17.04g} = 1.467\ mol\ NH_3$$

$$40.0\ g\ NH_4NO_3 \quad \frac{1\ mol}{80.06\ g} = 0.50\ mol\ NH_4NO_3$$

Because the NO_3^- is just a spectator ion, we are only really interested in the amount of conjugate base this salt provides, we convert from moles of NH_4NO_3 to moles of NH_4^+.

$$pH = -\log(5.6 \times 10^{-10}) + \log \frac{\left(\frac{1.47\ mol}{\cancel{0.501\ L}}\right)}{\left(\frac{0.50\ mol}{\cancel{0.501\ L}}\right)} = -\log(5.6 \times 10^{-10}) + \log \frac{(1.47\ mol)}{(0.50\ mol)}$$

$$pH = \boxed{9.72}$$

As you can see the volumes cancel out, since the volumes will ultimately cancel out, it is valid to simply plug in moles rather than molarity when using this formula.

5. What is buffer capacity?

The amount of acid or base a buffer can absorb without having a significant change in pH.

6. Consider:

0.1 M NaF and 0.1M HF
1.0M NaF and 1.0M HF
0.01M NaF and 0.01M HF

a. Which has the highest pH?

They all have the same pH because they all have the same K_a value and ratio of [base]: [acid] . So in a buffer the pH doesn't depend on how concentrated the acid and base are; it depends on their ratios to one another.

b. Which has the highest buffer capacity?

1.0 M NaF and 1.0 M HF has the highest buffer capacity as it has the greatest concentration of acid/base.

7. Consider a solution that contains both NH_3 and NH_4NO_3. If a solution has a pH = 8.5, calculate the ratio of

$$\frac{[NH_3]}{[NH_4^+]}$$

$K_a = 5.6 \times 10^{-10}$

Looking at the solution you will notice that it contains a weak base and the salt contains its conjugate acid. That means that this solution is a buffer and we can use the Henderson - Hasselbalch equation.

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

All we have to do is plug the information provided into the equation and solve for the $\frac{[base]}{[acid]}$.

$$8.5 = -\log(5.6 \times 10^{-10}) + \log \frac{[base]}{[acid]}$$

$$\frac{[base]}{[acid]} = \boxed{0.177}$$

8. What volumes of 0.22M CH_3COOH and 0.46M $NaCH_3COO$ must be mixed to prepare a 1.00L solution buffered at pH = 5.00? (K_a = 1.76 x 10^{-5})

Once again we start by determining whether or not this solution is a buffer. As it contains CH_3COOH (a weak acid) and a salt that contains its conjugate base (CH_3COO^-), this solution would be a buffer. This means that we can use the Henderson-Hasselbalch equation – plugging in all the relevant data.

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

$$5.00 = -\log(1.76 \times 10^{-5}) + \log \frac{[base]}{[acid]}$$

$$\frac{[base]}{[acid]} = 1.76$$

because they are in the same volume of solution

$$\frac{\text{moles base}}{\text{moles acid}} = 1.76$$

Remember:

$$Molarity \times Volume = moles$$

Since the volume of the solution is equal to 1.00L we know that the sum of the volume of base added plus the volume of acid added is equal to 1.00 L.

$$\text{Let } x = \text{volume of base added}$$
$$\text{Let } (1.00 - x) = \text{volume of acid added.}$$

Thus we can set up the following equation

$$\frac{(x)(0.46\ M)}{(1.00 - x)(0.22)} = 1.76$$

$$0.46x = 0.387 - 0.387x$$

$$0.847x = 0.387$$

$$x = \boxed{0.457L \rightarrow \text{volume of base added}}$$

$$1.00 - 0.457 = \boxed{0.543L \rightarrow \text{volume of acid added}}$$