**How Buffers Work**

**http://www.sparknotes.com/chemistry/acidsbases/buffers/section1.html**

Only a small amount of a strong acid is necessary to drastically alter the pH of a solution or water.

For certain experiments, however, it is desirable **to keep a fairly constant pH while acids or bases are added to the solution** either by the reaction or by the experimenter.

Buffers are designed to fill that role.

Chemists use buffers routinely to moderate the pH of a reaction.

Biology finds manifold uses for buffers which range from **controlling blood pH** to ensuring that **urine does not reach painfully acidic levels**.

**Buffer:**  a mixture of a \_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_ and its conjugate base

 or

 a mixture of a weak base and its \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Buffers work by reacting with any added acid or base to control the pH.

e.g. weak base ammonia NH3 + \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**NH3(aq) + H2O(l) ↔ NH4+(aq) + OH-(aq)**

How do we make the buffer?

**Add some HCl:**

When HCl is added to that buffer, the NH3 "soaks up" the acid's proton to become NH4+. Because that proton is locked up in the ammonium ion, its proton does not serve to significantly increase the pH of the solution.

**Add some NaOH:**

When NaOH is added to the same buffer, the ammonium ion donates a proton to the base to become ammonia and water. Here the buffer also serves to neutralize the base.

As the above example shows, a buffer works by replacing a strong acid or base with a weak one.

The strong acid's proton is replaced by ammonium ion, a weak acid.

The strong base OH- was replaced by the weak base ammonia.

These replacements of strong acids and bases for weaker ones give buffers their ability to moderate pH.

**Calculating the pH of Buffered Solutions**

Buffers must be chosen for the appropriate pH range that they are called on to control. The pH range of a buffered solution is given by the Henderson- Hasselbalch equation. For the purpose of the derivation, we will imagine a buffer composed of an acid, HA, and its conjugate base, A-. We know that the acid ionization constant Ka of the acid is given by this expression:



The equation can be rearranged as follows:



A buffer problem can be fairly simple to solve, provided you don't get confused by all the other chemistry you know.

For example, let's calculate the pH of a solution that is 0.5 M acetic acid and 0.5 sodium acetate.

Calculate the pH of the buffer.

**Another example:**

A buffer solution containing 0.10 mol dm-3 of ethanoic acid and 0.20 mol dm-3 of sodium ethanoate. How do you calculate its pH?

Sketch of buffer solution:

|  |
| --- |
|  |

In any solution containing a weak acid, there is an equilibrium between the un-ionised acid and its ions. So for ethanoic acid, you have the equilibrium:



The presence of the ethanoate ions from the sodium ethanoate will have moved the equilibrium to the left, but the equilibrium still exists.

That means that you can write the equilibrium constant, Ka, for it:



*Where you have done calculations using this equation previously with a weak acid, you will have assumed that the concentrations of the hydrogen ions and ethanoate ions were the same. Every molecule of ethanoic acid that splits up gives one of each sort of ion.*

**That's no longer true for a buffer solution:**



If the equilibrium has been pushed even further to the left, the number of ethanoate ions coming from the ethanoic acid will be completely negligible compared to those from the sodium ethanoate.

We therefore assume that the ethanoate ion concentration is the same as the concentration of the sodium ethanoate - in this case, 0.20 mol dm-3.

In a weak acid calculation, we normally assume that so little of the acid has ionised that the concentration of the acid at equilibrium is the same as the concentration of the acid we used. That is even more true now that the equilibrium has been moved even further to the left.

So the assumptions we make for a buffer solution are:



Now, if we know the value for Ka, we can calculate the hydrogen ion concentration and therefore the pH.

Ka for ethanoic acid is 1.74 x 10-5 mol dm-3.

Remember that we want to calculate the pH of a buffer solution containing 0.10 mol dm-3 of ethanoic acid and 0.20 mol dm-3 of sodium ethanoate.



Then all you have to do is to find the pH using the expression:

**pH = -log10 [H+]**