Buffers

If the $[H^+]$ (or pH) of a solution is not appreciably affected by the addition of small amounts of acids and bases, the solution is said to be buffered. A solution will have this property if it contains relatively large amounts of both a weak acid and a weak base.

If a small amount of a strong acid is added to this solution, most of the added H+ will combine with an equivalent amount of the weak base of the buffer to form the conjugate acid of that weak base; thus the $[H^+]$ of the solution remains almost constant. If a small amount of a strong base is added to the buffer solution, most of the OH^- will combine with an equivalent amount of the weak acid of the buffer to form the conjugate base of that weak acid. In this way, the $[H^+]$ of the buffer solution is not appreciably affected by the addition of small amounts of acid or base.

The buffer capacity is the amount of acid or base that may be added to a buffer solution before a significant change in pH occurs. When too much acid or base is added, the buffer capacity of a solution is exceeded and a significant change in pH is observed.

Any pair of weak acid, weak base can be used to form a buffer solution, as long as each can form its conjugate base or conjugate acid in water solution.

All pH curves have at least one region where buffering action occurs. The curves in these relatively constant pH regions are most nearly horizontal at a volume of titrant which is one-half the first equivalence point, or halfway between successive equivalence point for polyprotic substances.

Write Brønsted-Lowry equations to show what happens when acid or base is added to the following buffers.

1. ethanoic acid/ethanoate buffer (acetic acid/acetate)

2. ammonium ion/ammonia buffer

3. phosphoric acid/dihydrogen phosphate ion buffer

4. methanoic acid/methanoate buffer (formic acid/formate)