**Buffers**

If 1 mL of 0.10 M hydrochloric [acid](https://www.ck12.org/c/physical-science/acid?referrer=crossref) is added to 1 L of pure [water](https://www.ck12.org/c/biology/water?referrer=crossref) the [pH](https://www.ck12.org/c/physical-science/ph?referrer=crossref) drops drastically, from 7 to 4. This is a 1000-fold increase in the acidity of the [solution](https://www.ck12.org/c/physical-science/solution?referrer=crossref)!

For many purposes (like your blood), it is desirable to have a solution which is capable of resisting such large changes in pH when relatively small amounts of acid or [base](https://www.ck12.org/c/physical-science/base?referrer=crossref) are added to them. Such a solution is called a buffer. A **buffer** is a relatively concentrated [solution](https://www.ck12.org/c/physical-science/solution?referrer=crossref) of a weak [acid](https://www.ck12.org/c/physical-science/acid?referrer=crossref) or [base](https://www.ck12.org/c/physical-science/base?referrer=crossref) and its salt. Both components must be present for the system to act as a buffer to resist changes in pH. Commercial buffer [solutions](https://www.ck12.org/c/chemistry/solutions?referrer=crossref), which have a wide variety of pH values, can be obtained. Some common buffer systems are listed [below](https://www.ck12.org/book/CK-12-Chemistry-Concepts-Intermediate/section/21.23/#x-ck12-MTM2NzA0MTA1Njk4NQ..):

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| --- |
| **Some Common Buffers** |
| **Buffer system** | **Buffer components** | **pH of buffer (equal molarities of both components)** |
| Acetic acid / acetate [ion](https://www.ck12.org/c/physical-science/ion?referrer=crossref) | CH3COOH / CH3COO− | 4.74 |
| Carbonic acid / hydrogen carbonate [ion](https://www.ck12.org/c/physical-science/ion?referrer=crossref) | H2CO3 / HCO3− | 6.38 |
| Dihydrogen phosphate [ion](https://www.ck12.org/c/physical-science/ion?referrer=crossref" \o "ion) / hydrogen phosphate ion | H2PO4− / HPO42− | 7.21 |
| Ammonia / ammonium ion | NH3 / NH4+ | 9.25 |

One example of a buffer is a [solution](https://www.ck12.org/c/physical-science/solution?referrer=crossref) made of acetic acid (the weak acid) and sodium acetate (the salt). The pH of a buffer consisting of 0.50 M CH3COOH and 0.50 M CH3COONa is 4.74. If 10 mL of 1 M HCl is added to 1 L of the buffer, the pH only decreases to 4.73. This ability to “soak up” the additional hydrogen ions from the HCl that was added is due to the equilibrium which was already established in the solution.

CH3COOH(aq)  + H2O(l) ⇄ CH3COO−(aq) + H3O+(aq)

Remembering Le Chatelier’s principle – the increase in H+ from the extra acid will cause the equilibrium shown above to shift to the right. However since the amount of extra H+ was small compared to the existing concentrations of acetic acid and acetate [ion](https://www.ck12.org/c/physical-science/ion?referrer=crossref), the concentration change of either of them is minor. Small concentration changes to the equilibrium system for a buffer have only very minor effects on the pH.

If 10 mL of 1 M NaOH were added to another 1 L of the same buffer, the pH would only increase to 4.76. In this case the buffer equilibrium shifts to the right slightly to neutralize the additional hydroxide ions.

OH−(aq) + H3O+(aq) ⇄ 2 H2O(l)

CH3COOH(aq)  + H2O(l) ⇄ CH3COO−(aq) + H3O+(aq)

Again, the ratio of acetate ion to acetic acid changes slightly, again causing a very small change in the pH.

However it is possible to add so much acid or [base](https://www.ck12.org/c/physical-science/base?referrer=crossref) to a buffer that its ability to resist a significant change in pH is overwhelmed. The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs. The buffer capacity is exceeded when the number of moles of H+ or OH− that are added to the buffer exceeds the number of moles of the buffer components.