Buffer and Media Preparation – What Is a Buffer?

<u>Your Objectives:</u>

At the end of the lesson you should be will able to explain how to make a buffer solution.

A **buffer solution** (more precisely, pH buffer or hydrogen ion buffer) is an aqueous solution consisting of a mixture of a weak acid and its conjugate base, or vice versa). Its pH changes insignificantly when a small amount of strong acid or base is added to it. Buffer solutions are used as a means of keeping pH at a nearly constant value in a wide variety of chemical applications. In nature, there are many systems that use buffering for pH regulation. For example, the bicarbonate buffering system is used to regulate the pH of blood.

pH

In chemistry, pH (denoting 'potential of hydrogen' or 'power of hydrogen') is a scale used to specify the acidity or basicity of an aqueous solution. Lower pH values correspond to solutions which are more acidic in nature, while higher values correspond to solutions which are more basic or alkaline. At room temperature (25 C or 77 F), pure water is neutral (neither acidic nor basic) and has a pH of 7.

Substance	pH range	Type
Battery acid	< 1	
Gastric acid	1.0 – 1.5	
Vinegar	2.5	Acid
Orange juice	3.3 – 4.2	
Black coffee	5 – 5.03	
Milk	6.5 – 6.8	
Pure water	7	Neutral
Sea water	7.5 – 8.4	
Ammonia	11.0 – 11.5	Base
Bleach	12.5	
Lye	13.0 – 13.6	

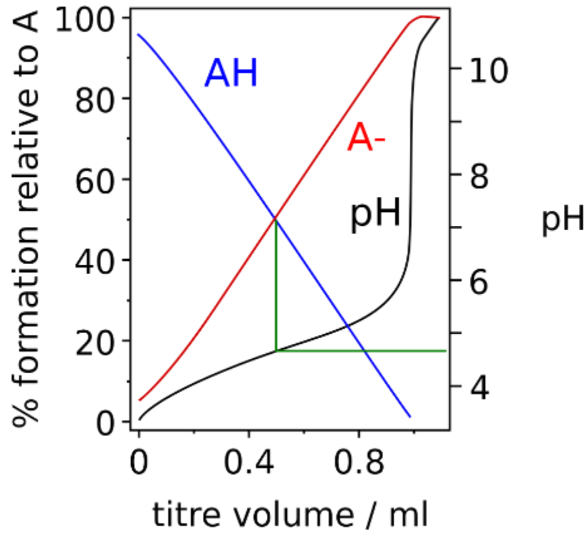
The pH scale is logarithmic and inversely indicates the concentration of hydrogen ions in the solution (a lower pH indicates a higher concentration of hydrogen ions). More precisely, pH is the negative of the base-10 logarithm of the activity of the hydrogen ion.

At 25°C, solutions with a pH less than 7 are acidic, and solutions with a pH greater than 7 are basic. The neutral value of the pH depends on the temperature being lower than 7 if the temperature increases. The pH value can be less than 0 for very strong acids, or greater than 14 for very strong bases. The pH scale is traceable to a set of standard solutions whose pH is established by international agreement.

Buffer solutions achieve their resistance to pH change because of the presence of an equilibrium between the weak acid HA and its conjugate base A^-:

$$HA \rightleftharpoons H^+ + A^-$$

When some strong acid is added to an equilibrium mixture of the weak acid and its conjugate base, hydrogen ions (H+) are added, and the equilibrium is shifted to the left, in accordance with Le Châtelier's principle. Because of this, the hydrogen ion concentration increases by less than the amount expected for the quantity of strong acid added. Similarly, if strong alkali is added to the mixture, the hydrogen ion concentration decreases by less than the amount expected for the quantity of alkali added. The effect is illustrated by the simulated titration of a weak acid with pKa = 4.7. The relative concentration of undissociated acid is shown in blue, and of its conjugate base in red.



The pH changes relatively slowly in the buffer region, pH = pKa ± 1, centered at pH = 4.7, where [HA] = [A−]. The hydrogen ion concentration decreases by less than the amount expected because most of the added hydroxide ion is consumed in the reaction

$OH^- + HA \rightarrow H_2O + A^-$

and only a little is consumed in the neutralization reaction (which is the reaction that results in an increase in pH)

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

Once the acid is more than 95% deprotonated, the pH rises rapidly since most of the added alkali is consumed in the neutralization reaction.

Helpful link: https://www.khanacademy.org/science/ap-chemistry/buffers-titrations-solubility-equilibria-ap/buffer-solutions-tutorial-ap/v/buffer-system

The pH of a solution containing a buffering agent can only vary within a narrow range, regardless of what else may be present in the solution. In biological systems this is an essential condition for enzymes to function correctly. In human blood, for instance, we find a mixture of carbonic acid (H_2CO_3) and bicarbonate (HCO_3^-) present in the plasma fraction, which constitutes the major mechanism for maintaining the pH level of blood at between 7.35 and 7.45. Outside this narrow range (7.40 ± 0.05 pH unit), acidosis and alkalosis metabolic conditions rapidly develop, ultimately leading to death if the correct buffering capacity is not rapidly restored.

If the pH value of a solution rises or falls too much, the effectiveness of an enzyme decreases in a process, known as denaturation, which is usually irreversible. The majority of biological samples that are used in research are kept in a buffer solution, often phosphate buffered saline (PBS) at pH 7.4.

In industry, buffering agents are used in fermentation processes as well as in setting the correct conditions for dyes used in colouring fabrics. They are also used in chemical analysis and calibration of pH meters.

Simple buffering agents

Buffering agent	Useful pH range
Citric acid	2.1–7.4
Acetic acid	3.8–5.8
KH_2PO_4	6.2–8.2
CHES	8.3–10.3
Borate	8.25–10.25

For buffers in acidic regions, the pH may be adjusted to a desired value by adding a strong acid, such as hydrochloric acid, to the particular buffering agent. For alkaline buffers, a strong base, such as sodium hydroxide, may be added. Alternatively, a buffer mixture can be made from a mixture of an acid and its conjugate base. For example, an acetate buffer can be made from a mixture of acetic acid and sodium acetate. Similarly, an alkaline buffer can be made from a mixture of the base and its conjugate acid.