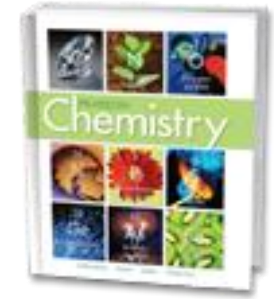
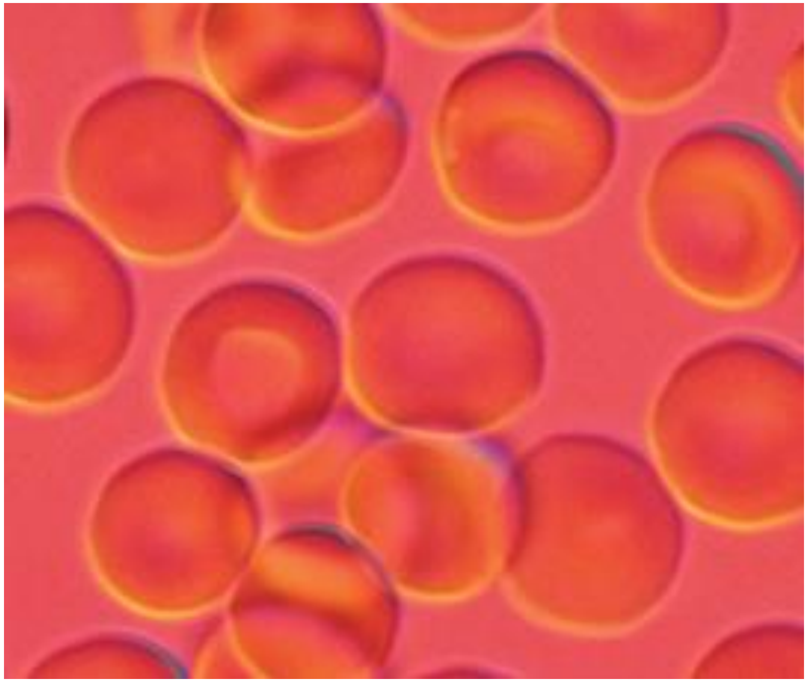


Capítulo 19

Acidos, Bases, y Sales

¿Cómo es controlado el pH de la sangre?

El pH de la sangre humana necesita permanecer cercano a 7.4. Una persona no puede sobrevivir más que unos minutos si el pH de la sangre cae por debajo de 6.8 o aumenta por encima de 7.8.



Hidrólisis de sales

¿ Cuándo una solución de una sal es ácida o básica?

Una sal es uno de los productos de una reacción de neutralización.

- Una sal consiste en un anión procedente de un ácido y un catión procedente de una base.
- Las soluciones acuosas de muchas sales son neutras.

Algunas sales dan soluciones ácidas o básicas.

- Se adiciona un indicador universal a soluciones acuosas de 3 sales 0.10*M*. Clasifiquelas:

Ammonium chloride, NH_4Cl(*aq*), is acidic (pH of about 5.3).

Sodium chloride, NaCl(*aq*), is neutral (pH of 7).

Sodium ethanoate, CH_3COONa(*aq*), is basic (pH of about 8.7).

Las sales que producen soluciones ácidas poseen iones positivos que liberan protones al medio.

Las sales que poseen iones negativos que capturan iones hidrógeno del agua.

Sodium ethanoate (CH_3COONa) is the salt of a weak acid and a strong base.

- In solution, the salt is completely ionized.

$$\underset{\text{Sodium ethanoate}}{CH_3COONa(aq)} \rightarrow \underset{\text{Ethanoate ion}}{CH_3COO^-(aq)} + \underset{\text{Sodium ion}}{Na^+(aq)}$$

El anion acetato es una base de Brønsted-Lowry, lo que significa que es un acceptor de H+

- Reacciona con el agua liberando oxidrilos
- En el equilibrio los reacctivos estan favorecidos.

$$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

CH_3COO^-: H^+ acceptor, Brønsted-Lowry base

H_2O: H^+ donor, Brønsted-Lowry acid

OH^-: (makes the solution basic)

$$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

H^+ aceptor
Brønsted-Lowry base

H^+ donor
Brønsted-Lowry
ácido

(hace la
solución
básica)

Este proceso es llamado hidrólisis debido a la ruptura de la molécula de agua en sus iones.

- El sufijo *-lysis* proviene del Griego y significa "separar" o "perder."
- En la solución, la concentración de iones hidroxilo es mayor que la de protones.
- Así, la solución es básica.

El Cloruro de amonio (NH_4Cl) es la sal formada entre el HCl (ácido fuerte) y la base débil amoníaco (NH_3).

- Se ioniza completamente en solución.

$$NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

El ion amonio es lo suficientemente ácido como para donar iones H+ a la molécula de agua.

- Los productos son amoníaco e iones hidronio.
- Los reactivos están favorecidos en el proceso, esto se representa con el tamño relatico de las flechas.

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

H^+ donor
Brønsted-Lowry acid

H^+ acceptor
Brønsted-Lowry base

(makes the solution acidic)

This process is another example of hydrolysis.

- At equilibrium the $[H_3O^+]$ is greater than the $[OH^-]$.
- Thus, a solution of ammonium chloride is acidic.

$$\underset{\substack{\text{H}^+\text{ donor} \\ \text{Brønsted-Lowry acid}}}{NH_4^+(aq)} + \underset{\substack{\text{H}^+\text{ acceptor} \\ \text{Brønsted-Lowry base}}}{H_2O(l)} \rightleftharpoons NH_3(aq) + \underset{\substack{\text{(makes the} \\ \text{solution acidic)}}}{H_3O^+(aq)}$$

Para determiner si una sal producirá medio ácido o alcalino en agua, pueden seguirse las siguientes reglas:

Anión	Catión		Medio
Ácido fuerte +	**Base fuerte**	→	Neutro
Ácido fuerte +	Base débil	→	**Ácido**
Ácido débil +	**Base fuerte**	→	**Básico**

Checkpoint

Does the fact that a weak acid-strong base titration is basic mean that there is some base “left over” at the equivalence point?

Checkpoint

Does the fact that a weak acid-strong base titration is basic mean that there is some base “left over” at the equivalence point?

No. All of the acid and base has been converted to a salt solution at the equivalence point. The solution is basic because the salt hydrolyzes. The salt has negative ions that attract hydrogen ions from water.

Buffers o Soluciones amortiguadoras

¿Cuáles son los componentes de un buffer?

Suppose you add 10 mL of 0.10*M* sodium hydroxide to 1 L of pure water.

- The pH will increase about 4 pH units—from 7.0 to about 11.0.
- This change is a relatively large increase in pH.

Now consider a solution containing 0.20*M* each of ethanoic acid and sodium ethanoate. This solution has a pH of 4.76.

- If you add 10 mL of 0.10*M* sodium hydroxide to 1 L of this solution, the pH increases 0.01 pH units—from 4.76 to 4.77.
  - This is a relatively small change in pH.
- If 10 mL of acid had been added instead of the base, the amount of change in pH would also have been small.

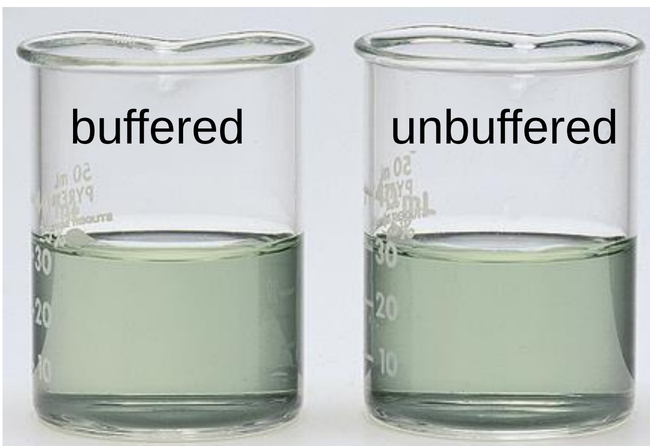
The solution of ethanoic acid and sodium ethanoate is an example of a buffer.

- A **buffer** is a solution in which the pH remains fairly constant when small amounts of acid or base are added.

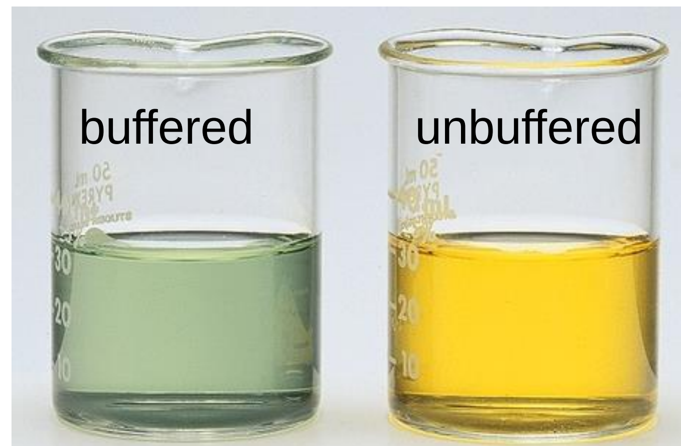
A buffer is a solution of a weak acid and one of its salts or a solution of a weak base and one of its salts.

The figure below compares what happens when 1.0 mL of 0.10*M* HCl solution is added to an unbuffered solution and to a solution with a buffer.

HCl is added to each solution.



The indicator shows that both solutions are basic (pH of about 8).



The indicator shows no visible pH change in the buffered solution.

The color change in the unbuffered solution indicates a change in pH from 8 to about 3.

How Buffers Work

A buffer solution is better able to resist drastic changes in pH than is pure water.

- A buffer solution contains one component that can react with hydrogen ions (hydrogen -ion acceptor) and one that can react with hydroxide ions (hydrogen-ion donor).
  - These components act as reservoirs of neutralizing power that can be tapped when either hydrogen ions or hydroxide ions are added to the solution.

How Buffers Work

The ethanoic acid-ethanoate ion buffer can be used to show how a buffer works.

- When an acid is added to the buffer, the ethanoate ions (CH_3COO^-) act as a hydrogen-ion "sponge."
- As the ethanoate ions react with the hydrogen ions, they form ethanoic acid.

How Buffers Work

Ethanoic acid is a weak acid and does not ionize extensively in water, so the change in pH is very slight.

$$\underset{\text{Ethanoate ion}}{CH_3COO^-(aq)} + \underset{\text{Hydrogen ion}}{H^+(aq)} \rightleftharpoons \underset{\text{Ethanoic acid}}{CH_3COOH(aq)}$$

How Buffers Work

When hydroxide ions are added to the buffer, the ethanoic acid and the hydroxide ions react to produce water and the ethanoate ion.

$$\underset{\text{Ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)} \rightleftharpoons \underset{\text{Ethanoate ion}}{CH_3COO^-(aq)} + \underset{\text{Water}}{H_2O(l)}$$

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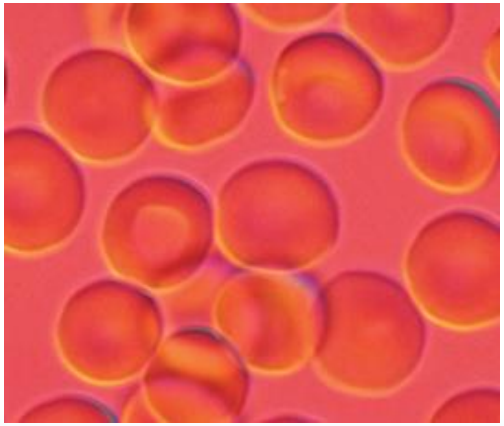
- The ethanoate ion is not a strong enough base to accept hydrogen ions from water to a great extent.
  - Therefore, the reverse reaction is minimal and the change in pH is very slight.

CHEMISTRY & YOU

The equilibrium between carbonic acid (H_2CO_3) and hydrogen carbonate ions (HCO_3^-) helps keep the pH of blood within a narrow range (7.35–7.45). If the pH rises, molecules of carbonic acid donate hydrogen ions. What can happen if the pH drops, that is, if the $[H^+]$ increases?



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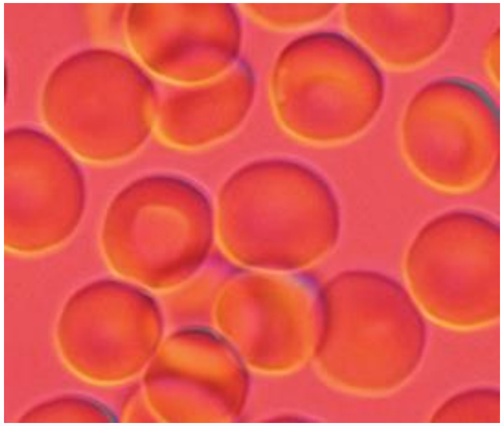
Hydrogen carbonate ions can accept hydrogen ions when the pH drops.



The Capacity of a Buffer

Buffer solutions have their limits.

- As acid is added to an ethanoate buffer, eventually no more ethanoate ions will be present to accept the hydrogen ions.
  - At that point, the buffer can no longer control the pH.

The Capacity of a Buffer

Buffer solutions have their limits.

- The ethanoate buffer also becomes ineffective when too much base is added.
  - No more ethanoic acid molecules are present to donate hydrogen ions.

The Capacity of a Buffer

Adding too much acid or base will exceed the buffer capacity of a solution.

- The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.

The Capacity of a Buffer

This table lists some common buffer systems.

Important Buffer Systems		
Buffer name	**Formulas**	**Buffer pH***
Ethanoic acid-ethanoate ion	CH_3COOH/CH_3COO^-	4.76
Dihydrogen phosphate ion-hydrogen phosphate ion	$H_2PO_4^-/HPO_4^{2-}$	7.20
Carbonic acid-hydrogen carbonate ion (solution saturated with CO_2)	H_2CO_3/HCO_3^-	6.46
Ammonium ion-ammonia	NH_4^+/NH_3	9.25

* Components have concentrations of 0.1*M*.

The Capacity of a Buffer

Two buffer systems help maintain optimal human blood pH.

- One is the carbonic acid-hydrogen carbonate buffer system.
- The other is the dihydrogen phosphate-hydrogen phosphate buffer system.

Describing Buffer Systems

Write balanced chemical equations to show how the carbonic acid-hydrogen carbonate buffer can "mop up" added hydroxide ions and hydrogen ions.

1 Analyze **Identify the relevant concepts.**

A buffer contains two components:

- a hydrogen-ion acceptor (which can react with H^+)
- a hydrogen-ion donor (which can react with OH^-)

2 Calculate Apply the concepts to this problem.

Identify the hydrogen-ion acceptor and the hydrogen-ion donor.

- H_2CO_3, a weak acid, can release hydrogen ions.
- HCO_3^- is the conjugate base, which can accept hydrogen ions.

2 Calculate Apply the concepts to this problem.

Write the equation for the reaction that occurs when a base is added to the buffer.

When a base is added, the hydroxide ions react with H_2CO_3.

$$H_2CO_3(aq) + (OH^-)(aq) \rightleftharpoons HCO_3^-(aq) + H_2O(l)$$

2 Calculate **Apply the concepts to this problem.**

Write the equation for the reaction that occurs when an acid is added to the buffer.

When an acid is added, the hydrogen ions react with HCO_3^-.

$$HCO_3^-(aq) + (H^+)(aq) \rightleftharpoons H_2CO_3(aq)$$

Checkpoint

How is the work of a buffer solution similar to a neutralization reaction?

Checkpoint

How is the work of a buffer solution similar to a neutralization reaction?

A buffer solution contains compounds that are able to neutralize both acids and bases. It performs acid-base neutralization reactions without significant change in pH.

- **Salts that produce acidic solutions have positive ions that release hydrogen ions to water. Salts that produce basic solutions have negative ions that attract hydrogen ions from water.**

- **A buffer is a solution of a weak acid and one of its salts or a weak base and one of its salts.**

- **salt hydrolysis**: a process in which the cations or anions of a dissociated salt accept hydrogen ions from water or donate hydrogen ions to water
- **buffer**: a solution in which the pH remains relatively constant when small amounts of acid or base are added; a buffer can be either a solution of a weak acid and the salt of a weak acid or a solution of a weak base with the salt of a weak base
- **buffer capacity**: a measure of the amount of acid or base that may be added to a buffer solution before a significant change in pH occurs

END OF 19.5