ACIDS & BASES

- Many common substances in our daily lives are **acids and bases**. **Oranges, lemons and vinegar** are examples of **acids**. In addition, our stomachs contain acids that help digest foods. **Antacid** tablets taken for heartburn and **ammonia** cleaning solutions are examples of **bases**.

- General properties associated with acids include the following:

  - sour taste
  - change color of litmus from blue to red
  - react with metals to produce H_2 gas
  - react with bases to produce salt and water

- General properties associated with bases include the following:

  - bitter taste
  - slippery soapy feeling
  - change color of litmus from red to blue
  - react with acids to produce salt and water

- The most common definition of acids and bases was formulated by the Swedish chemist Svante **Arrhenius** in 1884.

- According to the **Arrhenius** definition,

  - Acids are substances that produce hydronium ion (H_3O^+) in aqueous solution

$$HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)$$

Commonly written as

$$\color{red}{\mathbf{H}}Cl\,(g) \xrightarrow{H_2O} \color{red}{\mathbf{H^+}}\,(aq) + Cl^-\,(aq)$$

Polar Covalent

  - Bases are substances that produce hydroxide ion (OH^-) in aqueous solution

$$NaOH\,(s) \xrightarrow{H_2O} Na^+\,(aq) + OH^-\,(aq)$$

Ionic compound

$$NH_3\,(aq) + H_2O\,(l) \longrightarrow NH_4^+\,(aq) + OH^-\,(aq)$$

BRØNSTED-LOWRY ACIDS & BASES

- The **Arrhenius** definition of acids and bases is **limited to aqueous solutions**.

- A broader definition of acids and bases was developed by **Brønsted and Lowry** in the early 20th century.

- According to **Brønsted-Lowry** definition, an **acid** is a **proton donor**, and a **base is a proton acceptor**.

$HCl\ (g) + H_2O\ (l) \rightarrow H_3O^+\ (aq) + Cl^-\ (aq)$
(acid) **(base)**

$NH_3\ (aq) + H_2O\ (l) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$
(base) **(acid)**

- A substance that can act as a **Brønsted-Lowry acid and base** (such as **water**) is called **amphiprotic**.

- In Brønsted-Lowry definition, any pair of molecules or ions that can be **interconverted by transfer of a proton** is called **conjugate acid-base pair**.

$HCl\ (g) + H_2O\ (l) \rightarrow Cl^-\ (aq) + H_3O^+\ (aq)$
(acid) **(base)** **(conjugate base of HCl)** **(conjugate acid of H_2O)**

$NH_3\ (aq) + H_2O\ (l) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$
(base) **(acid)** **(conjugate acid of NH_3)** **(conjugate base of H_2O)**

BRØNSTED-LOWRY ACIDS & BASES

Examples:

1. Identify the conjugate acid-base pairs for each reaction shown below:

$$H_2O \quad + \quad Cl^- \quad \rightarrow \quad HCl \quad + \quad OH^-$$

$$C_6H_5OH \quad + \quad C_2H_5O^- \quad \rightarrow \quad C_6H_5O^- \quad + \quad C_2H_5OH$$

2. Write the formula for the conjugate acid for each base shown:

HS^-

NH_3

CO_3^{2-}

3. Write the formula for the conjugate base for each acid shown:

HI

CH_3OH

HNO_3

ACID & BASE STRENGTH

- According to the **Arrhenius** definition, the **strength** of acids and bases is based on the amount of their **ionization** in water.

- **Strong** acids and bases are those that **ionize completely** in water.
- Strong acids and bases are **strong electrolytes**.

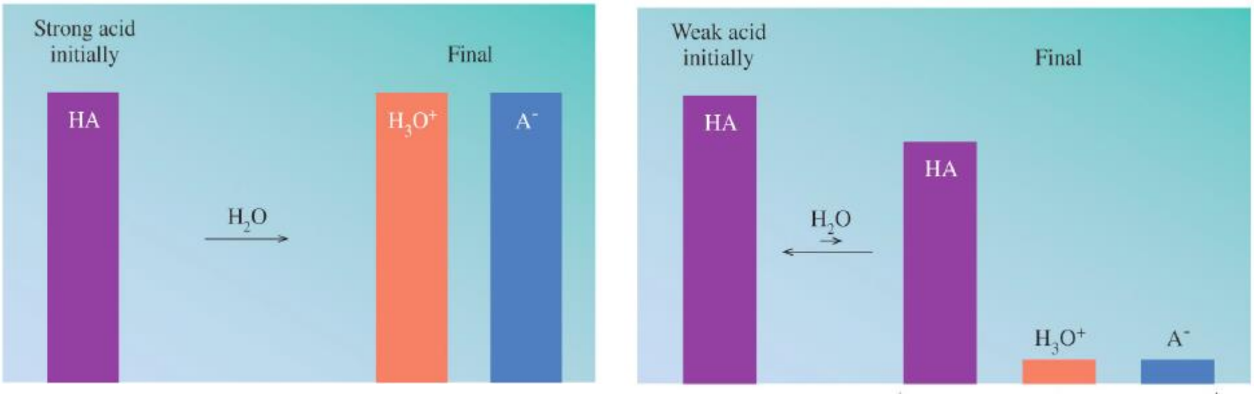
$$HCl\ (aq) \longrightarrow H^+\ (aq) + Cl^-\ (aq)$$

$$NaOH\ (s) \xrightarrow{H_2O} Na^+\ (aq) + OH^-\ (aq)$$

- **Weak** acids and bases are those that **ionize partially** in water, and are therefore written as ***reversible*** reactions (indicated by $\rightleftarrows$).
- Weak acids and bases are **weak electrolytes**.

$$HC_2H_3O_2 \overset{H_2O}{\rightleftarrows} H^+\ (aq) + C_2H_3O_2^-\ (aq)$$

$$NH_3\ (aq) + H_2O\ (l) \rightleftarrows NH_4^+\ (aq) + OH^-\ (aq)$$



Ionization of Strong vs. Weak acids

ACID & BASE STRENGTH

- Listed below are the formulas and names of common acids and bases, and comparison of their characteristics.

COMMON STRONG ACIDS & BASES			
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO_3	Nitric acid	$Ca(OH)_2$	Calcium hydroxide
H_2SO_4	Sulfuric acid	$Ba(OH)_2$	Barium hydroxide

COMMON WEAK ACIDS & BASES			
$HC_2H_3O_2$	Acetic acid	NH_3	Ammonia
H_2CO_3	Carbonic acid	$CO(NH_2)_2$	Urea
H_3PO_4	Phosphoric acid		
HF	Hydrofluoric acid		
H_2S	Hydrosulfuric acid		

Characteristic	Acids	Bases
Reaction: Arrhenius	Produce H^+	Produce OH^-
Reaction: Brønsted–Lowry	Donate H^+	Accept H^+
Electrolytes	Yes	Yes
Taste	Sour	Bitter, chalky
Feel	May sting	Slippery
Litmus	Red	Blue
Phenolphthalein	Colorless	Pink
Neutralization	Neutralize bases	Neutralize acids

IONIZATION OF WATER

- As noted previously, **water** can act both as an **acid and a base**.
- In pure water, one **water** molecule **donates a proton** to another water molecule to produce **ions**.

$$\underset{\substack{\text{Base} \\ \left[\begin{array}{c}\text{Proton}\\ \text{acceptor}\end{array}\right]}}{H_2O} + \underset{\substack{\text{Acid} \\ \left[\begin{array}{c}\text{Proton}\\ \text{donor}\end{array}\right]}}{H_2O} \rightleftharpoons \underset{\substack{\text{Acid} \\ \left[\begin{array}{c}\text{Proton}\\ \text{donor}\end{array}\right]}}{H_3O^+} + \underset{\substack{\text{Base} \\ \left[\begin{array}{c}\text{Proton}\\ \text{acceptor}\end{array}\right]}}{OH^-}$$

- In pure water, the **transfer of protons** between water molecules produces **equal numbers of H_3O^+ and OH^- ions**. However, the number of ions produced in pure water is very small, as indicated below:

$$\text{Pure water} \Rightarrow [H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ M$$

- When the **concentrations of H_3O^+ and OH^- are multiplied** together, the **ion-product constant (K_w)** is formed.

$$K_w = [H_3O^+] \times [OH^-]$$

$$= (1.0 \times 10^{-7}\ M) \times (1.0 \times 10^{-7}\ M) = 1.0 \times 10^{-14}$$

- All aqueous solutions have **H_3O^+ and OH^- ions**. An **increase in the concentration** of one of the ions will cause an equilibrium shift that causes a **decrease in the other one**.

ACIDIC & BASIC SOLUTIONS

- When **[H_3O^+]** and **[OH^-] are equal** in a solution, it is **neutral**.
- When **[H_3O^+]** is **greater** than **[OH^-]** in a solution, it is **acidic**.

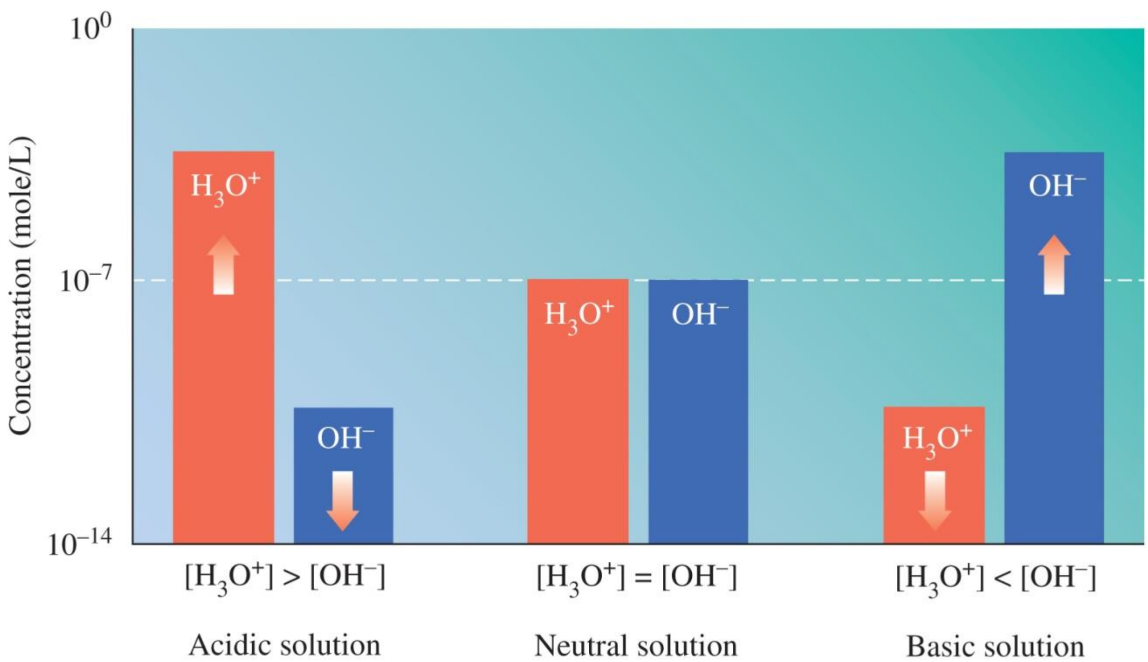
For example, if [H_3O^+] is 1.0 x 10^{-4} M, then [OH^-] would be 1.0 x 10^{-10} M.

$$\left[OH^-\right] = \frac{K_w}{\left[H_3O^+\right]} = \frac{1.0 \text{ x } 10^{-14}}{1.0 \text{ x } 10^{-4}} = 1.0 \text{ x } 10^{-10} \text{ M}$$

- When **[OH^-]** is **greater** than **[H_3O^+]** in a solution, it is **basic**.

For example, if [OH^-] is 1.0 x 10^{-6} M, then [H_3O^+] would be 1.0 x 10^{-8} M.

$$\left[H_3O^+\right] = \frac{K_w}{\left[OH^-\right]} = \frac{1.0 \text{ x } 10^{-14}}{1.0 \text{ x } 10^{-6}} = 1.0 \text{ x } 10^{-8} \text{ M}$$



ACIDIC & BASIC SOLUTIONS

Examples:

1. Calculate the $[OH^-]$ in a solution with $[H_3O^+] = 2.3 \times 10^{-4}$ M. Classify the solution as acid or basic.

$$\left[\mathbf{OH^-}\right] = \frac{\mathbf{K_w}}{\left[\mathbf{H_3O^+}\right]} =$$

2. Calculate the $[H_3O^+]$ in a solution with $[OH^-] = 3.8 \times 10^{-6}$ M. Classify the solution as acid or basic.

$$\left[\mathbf{H_3O^+}\right] = \frac{\mathbf{K_w}}{\left[\mathbf{OH^-}\right]} =$$

3. Calculate the $[OH^-]$ in a solution with $[H_3O^+] = 5.8 \times 10^{-8}$ M. Classify the solution as acid or basic.

4. Calculate $[H_3O^+]$ in solution prepared by dissolving 2.8 g KOH to make 45 mL of solution.

THE pH SCALE

- The **acidity** of a solution is commonly measured on a **pH scale**.

$$\mathbf{pH = -\log [H_3O^+]}$$

- The **pH scale** ranges from **0-14**, where **acidic** solutions are **less than 7** and **basic** solutions are **greater than 7**.

Acidic solutions	**pH < 7**	$[H_3O^+] > 1.0 \times 10^{-7}$
Neutral solutions	**pH = 7**	$[H_3O^+] = 1.0 \times 10^{-7}$
Basic solutions	**pH > 7**	$[H_3O^+] < 1.0 \times 10^{-7}$

Examples:

1. The $[H_3O^+]$ of a liquid detergent is 1.4×10^{-9} M. Calculate its pH.

   $$pH = -\log [H_3O^+] = -\log [1.4 \times 10^{-9}] = -(-8.85) = 8.85$$

2. The pH of black coffee is 5.3. Calculate its $[H_3O^+]$.

   $$[H_3O^+] = \text{antilog}(-pH) = 10^{-pH} = 10^{-5.3} = 5 \times 10^{-6}$$

3. The $[H_3O^+]$ of a solution is 3.5×10^{-3} M. Calculate its pH.

4. The pH of tomato juice is 4.1. Calculate its $[H_3O^+]$.

5. The $[OH^-]$ of a cleaning solution is 1.0×10^{-5} M. What is the pH of this solution?

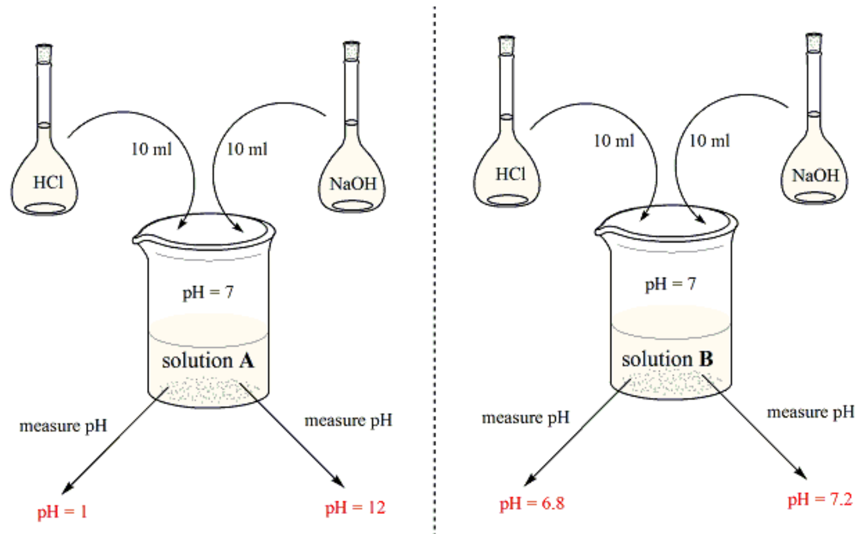
   $[H_3O^+] =$ pH=

6. Calculate the pH of a 1.2×10^{-3} M KOH solution.

BUFFERS

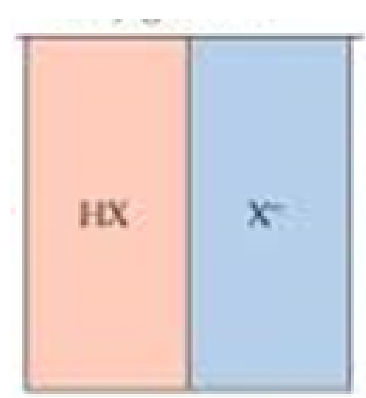
- The pH of water and most solutions change dramatically when a small amount of acid or base are added to it. However, when a small amount of acid or base is added to a buffer, the pH does not change very much.

- A **buffer** solution **maintains the pH** of a solution by neutralizing the added acid or base. Proper physiological functioning in the human body requires a very tight balance between the concentrations of acids and bases in the blood. A variety of buffering systems permits blood and other bodily fluids to maintain a narrow pH range.



addition of acid/base to water addition of acid/base to buffer

- A buffer solution consists of a weak acid and its conjugate base or a weak base and its conjugate acid. Both buffer systems work similarly, but we will concentrate on the former type to understand how a buffer works.



- An example of such buffer is a mixture of acetic acid (weak acid) and sodium acetate (conjugate base). In this buffer, acetic acid is represented as HX and the acetate ion is represented as X^-. The sodium ion is the cation associated with the acetate ion and is not part of the buffer system.

- In an ideal buffer, the mixture of weak acid and conjugate base are equal in amounts, but could be slightly different from one another.

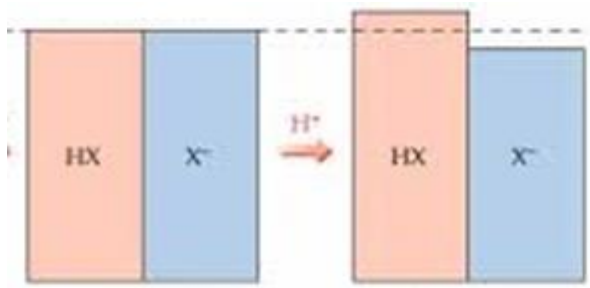
HOW BUFFERS WORK

- The acetic acid dissociates in water and produces a small amount of H_3O^+ and acetate ion, as shown below. The added sodium acetate increases the concentration of the acetate ion necessary for proper buffering capability.

$$HC_2H_3O_2\ (aq) + H_2O\ (l) \rightleftarrows H_3O^+\ (aq) + C_2H_3O_2^-\ (aq)$$

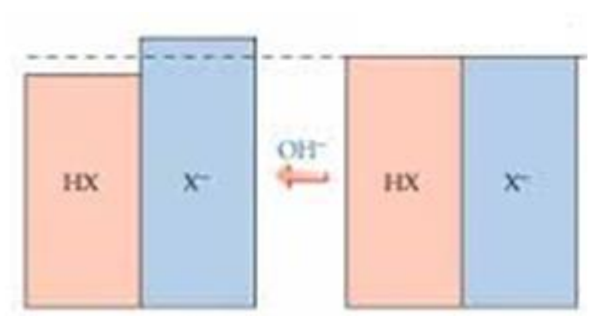
- When a small amount of acid is added to this buffer, the acetate ion neutralizes it to produce the weak acid and water, as shown below. As a result, the concentration of the weak acid (HX) increases slightly while the concentration of the conjugate base (X^-) decreases slightly. However, the pH of the solution changes slightly.

$$H_3O^+\ (aq) + C_2H_3O_2^-\ (aq) \longrightarrow HC_2H_3O_2\ (aq) + H_2O\ (l)$$



- When a small amount of base is added to this buffer, the acetic acid neutralizes it to produce the acetate ion and water, as shown below. As a result, the concentration of the weak acid (HX) decreases slightly while the concentration of the conjugate base (X^-) increases slightly. However, the pH of the solution changes slightly.

$$HC_2H_3O_2\ (aq) + OH^-\ (l) \rightleftarrows H_2O\ (aq) + C_2H_3O_2^-\ (aq)$$



CALCULATING pH OF BUFFERS

- Earlier we discussed that weak acids dissociate partially in solution as described by the equation below:

$$HC_2H_3O_2\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + C_2H_3O_2^-\ (aq)$$

- The degree of dissociation of a weak acid in solution can be quantified by the **acid dissociation constant (K_a)** which can be written as:

$$K_a = \frac{[H_3O^+]\,[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

- The greater the dissociation constant, the more the acid ionizes and the greater the $[H_3O^+]$ it produces in solution. Listed below are acid dissociation constants for some common acids.

Name	Formula	K_a
acetic acid	CH_3COOH	1.8×10^{-5}
benzoic acid	C_6H_5COOH	6.4×10^{-5}
chlorous acid	$HClO_2$	1.2×10^{-2}
formic acid	$HCOOH$	1.8×10^{-4}
hydrocyanic acid	HCN	6.2×10^{-10}
hydrofluoric acid	HF	6.3×10^{-4}

- Rearranging the acid dissociation constant above, we can calculate the $[H_3O^+]$ as:

$$[H_3O^+] = K_a \text{ x } \frac{[HC_2H_3O_2]}{[C_2H_3O_2^-]}$$

- The pH of the buffer solution can then be calculated from the $[H_3O^+]$.

CALCULATING pH OF BUFFERS

Examples:

1. The K_a for acetic acid is 1.8×10^{-5}. Calculate the pH of a buffer that is 1.0 M $HC_2H_3O_2$ and 1.0 M $NaC_2H_3O_2$.

2. One of the buffer systems in blood is $H_2PO_4^-/HPO_4^{2-}$. The K_a for $H_2PO_4^-$ is 6.2×10^{-8}. What is the pH of a buffer that is prepared from 0.10 M $H_2PO_4^-$ and 0.50 M HPO_4^{2-}.

3. One of the buffer systems used in blood is H_2CO_3/HCO_3^-.
   a) Write an equation that shows how this buffer neutralizes added acid.

   b) Write an equation that shows how this buffer neutralizes added base.