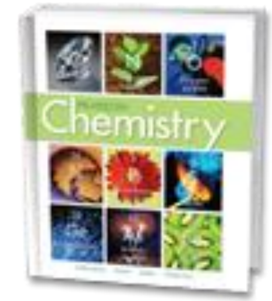
PEARSON Chemistry



Chapter 19
Acids, Bases, and Salts

19.1 Acid-Base Theories

19.2 Hydrogen Ions and Acidity

19.3 Strengths of Acids and Bases
19.4 Neutralization Reactions
19.5 Salts in Solution

PEARSON

What factors do you need to control so a fish has healthy water to live in?

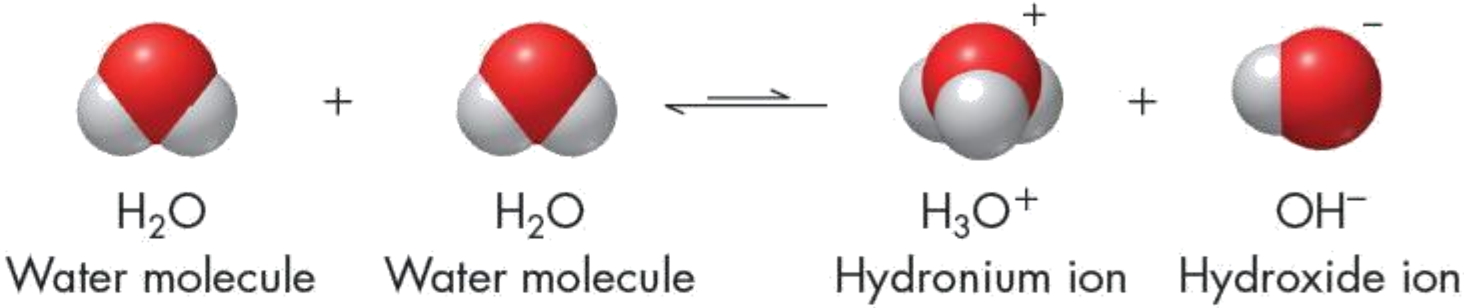
Goldfish can live for 20 years or more in an aquarium if the conditions are right.

Hydrogen Ions from Water

How are $[H^+]$ and $[OH^-]$ related in an aqueous solution?

Water molecules are highly polar and are in constant motion, even at room temperature.

- On occasion, the collisions between water molecules are energetic enough for a reaction to occur.



  - A water molecule that loses a hydrogen ion becomes a hydronium ion (H_3O^+).
  - A water molecule that loses a hydrogen ion becomes a hydroxide ion (OH^-).

Self-Ionization of Water

The reaction in which water molecules produce ions is called the **self-ionization** of water.

$$H_2O(l) \rightleftharpoons \underset{\text{Hydrogen ion}}{H^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

- In an aqueous solution, hydrogen ions are always joined to water molecules as hydronium ions.

Self-Ionization of Water

In pure water at 25° C, the concentration of hydrogen ions is only $1 \times 10^{-7} M$.

- The concentration of OH^- is also $1 \times 10^{-7} M$ because the numbers of H^+ and OH^- ions are equal in pure water.
- Any aqueous solution in which $[H^+]$ and $[OH^-]$ are equal is a **neutral solution**.

Ion-Product Constant for Water

The ionization of water is a reversible reaction, so Le Châtelier’s principle applies.

- Adding either hydrogen ions or hydroxide ions to an aqueous solution is a stress on the system.
- In response, the equilibrium will shift toward the formation of water.
  - The concentration of the other ion will decrease.

Ion-Product Constant for Water

The ionization of water is a reversible reaction, so Le Châtelier's principle applies.

- In any aqueous solution, when $[H^+]$ increases, $[OH^-]$ decreases. Likewise, when $[H^+]$ decreases, $[OH^-]$ increases.

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

Ion-Product Constant for Water

For aqueous solutions, the product of the hydrogen-ion concentration and the hydroxide-ion concentration equals 1.0×10^{-14}.

$$[H^+] + [OH^-] = 1.0 \times 10^{-14}$$

Ion-Product Constant for Water

The product of the concentrations of the hydrogen ions and the hydroxide ions in water is called the **ion-product constant for water (K_w)**.

$$K_w = [H^+] \times [OH^-] = 1.0 \times 10^{-14}$$

Ion-Product Constant for Water

Acidic Solutions

- When some substances dissolve in water, they release hydrogen ions.

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

- The hydrogen-ion concentration is greater than the hydroxide-ion concentration.
- A solution in which $[H^+]$ is greater than $[OH^-]$ is an **acidic solution**.
  - The $[H^+]$ is greater than $1 \times 10^{-7}M$.

Ion-Product Constant for Water

Basic Solutions

- When sodium hydroxide dissolves in water, it forms hydroxide ions in solution.

$$NaOH(aq) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

- The hydrogen-ion concentration is less than the hydroxide-ion concentration.
- A **basic solution** is one in which [H+] is less than [OH−].
  - Basic solutions are also known as alkaline solutions.

Using the Ion-Product Constant for Water

If the $[H^+]$ in a solution is $1.0 \times 10^{-5} M$, is the solution acidic, basic, or neutral? What is the $[OH^-]$ of this solution?

1 Analyze List the knowns and the unknowns.

Use the expression for the ion-product constant for water and the known concentration of hydrogen ions to find the concentration of hydroxide ions.

KNOWNS	UNKNOWNS
$[H^+] = 1.0 \times 10^{-5} M$	Is the solution acidic, basic, or neutral?
$K_w = 1.0 \times 10^{-14}$	$[OH^-] = ?M$

2 Calculate **Solve for the unknowns.**

Use $[H^+]$ to determine whether the solution is acidic, basic, or neutral.

- $[H^+]$ is $1.0 \times 10^{-5}M$, which is greater than $1.0 \times 10^{-7}M$.
- Thus, the solution is acidic.

2 Calculate **Solve for the unknowns.**

Rearrange the expression for the ion-product constant to solve for $[OH^-]$.

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

$$[OH^-] = \frac{K_w}{[H^+]}$$

2 Calculate Solve for the unknowns.

Substitute the known values of $[H^+]$ and K_w. Then solve for $[OH^-]$.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}}$$

$$= 1.0 \times 10^{-9} M$$

When you divide numbers written in scientific notation, subtract the exponent in the denominator from the exponent in the numerator.

3 Evaluate Does the result make sense?

- If $[H^+]$ is greater than $1.0 \times 10^{-7}M$, then $[OH^-]$ must be less than $1.0 \times 10^{-7}M$.
- $1.0 \times 10^{-9}M$ is less than $1.0 \times 10^{-7}M$.
- To check your calculation, multiply the values for $[H^+]$ and $[OH^-]$ to make sure that the result equals 1.0×10^{-14}.

Checkpoint

A solution does not have an equal number of H^+ and OH^- ions. What do you know about this solution?

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A solution does not have an equal number of H^+ and OH^- ions. What do you know about this solution?

You know that the solution is not neutral. Without knowing more information, you cannot say if it is acidic or basic (alkaline).

The pH Concept

How is pH used to classify a solution as neutral, acidic, or basic?

Expressing hydrogen-ion concentration in molarity is not practical. A more widely used system for expressing $[H^+]$ is the pH scale.

- The pH scale ranges from 0 to 14.

Hydrogen Ions and pH

The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration.

$$pH = -\log[H^+]$$

Hydrogen Ions and pH

In pure water or a neutral solution, the $[H^+] = 1 \times 10^{-7}M$, and the pH is 7.

$$\begin{aligned} pH &= -\log(1 \times 10^{-7}) \\ &= -(\log 1 + \log 10^{-7}) \\ &= -(0.0 + (-7.0)) = 7.0 \end{aligned}$$

- If the $[H^+]$ of a solution is greater than $1 \times 10^{-7}M$, the pH is less than 7.0.
- If the $[H^+]$ is less than $1 \times 10^{-7}M$, the pH is greater than 7.0.

Hydrogen Ions and pH

- A solution with a pH less than 7.0 is acidic.
- A solution with a pH of 7.0 is neutral.
- A solution with a pH greater than 7.0 is basic.

Hydrogen Ions and pH

Relationships Among $[H^+]$, $[OH^-]$, and pH

	$[H^+]$ (mol/L)	$[OH^-]$ (mol/L)	pH	
Increasing acidity	1×10^{0}	1×10^{-14}	0.0	1*M* HCl
	1×10^{-1}	1×10^{-13}	1.0	0.1*M* HCl
	1×10^{-2}	1×10^{-12}	2.0	Gastric juice; Lemon juice
	1×10^{-3}	1×10^{-11}	3.0	
	1×10^{-4}	1×10^{-10}	4.0	Tomato juice; Black coffee
	1×10^{-5}	1×10^{-9}	5.0	
	1×10^{-6}	1×10^{-8}	6.0	Milk
Neutral	1×10^{-7}	1×10^{-7}	7.0	Pure water; Blood
Increasing basicity	1×10^{-8}	1×10^{-6}	8.0	Seawater
	1×10^{-9}	1×10^{-5}	9.0	
	1×10^{-10}	1×10^{-4}	10.0	Milk of magnesia
	1×10^{-11}	1×10^{-3}	11.0	Household ammonia
	1×10^{-12}	1×10^{-2}	12.0	
	1×10^{-13}	1×10^{-1}	13.0	0.1*M* NaOH
	1×10^{-14}	1×10^{0}	14.0	1*M* NaOH

When $[H^+]$ is given in the format 1×10^{-n}, it's easy to find the pH. It's just the absolute value of the exponent n. Also, note that $[H^+] \times [OH^-]$ always equals 1×10^{-14}.

CHEMISTRY & YOU

In an aquarium, the pH of water is another factor that affects the ability of fish to survive. Most freshwater fish need a slightly acidic or neutral pH. For a saltwater tank, the ideal pH is slightly basic. What might explain this difference in the ideal pH range?

CHEMISTRY & YOU

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Natural freshwater is slightly acidic, while natural saltwater typically contains compounds that make it slightly basic. Fish have adapted to these conditions in the wild and need them replicated in their tanks.

Calculating pH from $[H^+]$

Expressing $[H^+]$ in scientific notation can make it easier to calculate pH.

- You would rewrite $0.0010M$ as $1.0 \times 10^{-3}M$.
  - The coefficient 1.0 has two significant figures.
- The pH for this solution is 3.00.
  - The two numbers to the right of the decimal point represent the two significant figures in the concentration.

Calculating pH from $[H^+]$

It is easy to find the pH for solutions when the coefficient is 1.0.

- The pH of the solution equals the exponent, with the sign changed from minus to plus.
  - A solution with $[H^+] = 1 \times 10^{-2} M$ has a pH of 2.0.
- When the coefficient is a number other than 1, you will need to use a calculator with a log function key to calculate pH.

Calculating pH from [H^+]

What is the pH of a solution with a hydrogen-ion concentration of $4.2 \times 10^{-10} M$?

1 Analyze **List the known and the unknown.**

To find the pH from the hydrogen-ion concentration, you use the equation $pH = -\log[H^+]$.

KNOWN	UNKNOWN
$[H^+] = 4.2 \times 10^{-10} M$	pH = ?

2 Calculate Solve for the unknown.

Start with the equation for finding pH from $[H^+]$.

$$pH = -\log[H^+]$$

2 Calculate **Solve for the unknown.**

Substitute the known $[H^+]$ and use the log function on your calculator to calculate the pH.

$$pH = -\log(4.2 \times 10^{-10})$$

$$= -(-9.37675)$$

$$= 9.37675$$

$$= 9.38$$

Round the pH to two decimal places because the hydrogen-ion concentration has two significant figures.

3 Evaluate Does the result make sense?

- The value of the hydrogen-ion concentration is between $1 \times 10^{-9}M$ and $1 \times 10^{-10}M$.
- So, the calculated pH should be between 9 and 10, which it is.

Calculating $[H^+]$ from pH

You can calculate the hydrogen-ion concentration of a solution if you know the pH.

- If the pH is an integer, it is easy to find $[H^+]$.
  - For a pH of 9.0, $[H^+] = 1 \times 10^{-9} M$.
- Most pH values are not whole numbers.
  - When the pH value is not a whole number, you will need a calculator with an antilog (10^x) function to get an accurate value for the hydrogen-ion concentration.

Calculating [H^+] from pH

The pH of an unknown solution is 6.35. What is the hydrogen-ion concentration of the solution?

1 Analyze List the known and the unknown.

You will use the antilog function of your calculator to find the concentration.

KNOWN	UNKNOWN
pH = 6.35	$[H^+] = ?M$

2 Calculate Solve for the unknown.

First, simply swap the sides of the equation for finding pH and substitute the known value.

$$pH = -\log[H^+]$$

$$-\log[H^+] = pH$$

$$-\log[H^+] = 6.35$$

2 Calculate **Solve for the unknown.**

Change the signs on both sides of the equation and then solve for the unknown.

$$\log[H^+] = -6.35$$

$$[H^+] = \text{antilog}(-6.35)$$

2 Calculate **Solve for the unknown.**

Use the antilog (10^x) function on your calculator to find $[H^+]$. Report the answer in scientific notation.

$$[H^+] = 4.5 \times 10^{-7}$$

On most calculators, use the 2nd or INV key followed by log to get the antilog.

3 Evaluate Does the result make sense?

- The pH is between 6 and 7.
- So, the hydrogen-ion concentration must be between $1 \times 10^{-6}M$ and $1 \times 10^{-7}M$.
- The answer is rounded to two significant figures because the pH was measured to two decimal places.

Calculating pH from $[OH^-]$

If you know the $[OH^-]$ of a solution, you can find its pH.

- You can use the ion-product constant to determine $[H^+]$ for a known $[OH^-]$.
- Then you use $[H^+]$ to calculate the pH.

Calculating pH from $[OH^-]$

What is the pH of a solution if $[OH^-] = 4.0 \times 10^{-11} M$?

1 Analyze List the knowns and the unknown.

- To find $[H^+]$, divide K_w by the known $[OH^-]$.
- Then calculate pH as you did in Sample Problem 19.3.

KNOWNS	UNKNOWN
$[OH^-] = 4.0 \times 10^{-11} M$	pH = ?
$K_w = 1.0 \times 10^{-14}$	

2 Calculate Solve for the unknown.

- Start with the ion-product constant to find $[H^+]$.
- Rearrange the equation to solve for $[H^+]$.

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

$$[H^+] = \frac{K_w}{[OH^-]}$$

2 Calculate **Solve for the unknown.**

Substitute the values for K_w and $[OH^-]$ to find $[H^+]$.

$$[H^+] = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-11}} = 0.25 \times 10^{-3}M$$

$$= 2.5 \times 10^{-4}M$$

2 Calculate **Solve for the unknown.**

- Next use the equation for finding pH.
- Substitute the value for $[H^+]$.

$$pH = -\log[H^+]$$
$$= -\log(2.5 \times 10^{-4})$$

- Use a calculator to find the log.

$$= -(-3.60205)$$
$$= 3.60$$

Round the pH to two decimal places because the $[OH^-]$ has two significant figures.

3 Evaluate Does the result make sense?

- A solution in which $[OH^-]$ is less than $1 \times 10^{-7}M$ is acidic because $[H^+]$ is greater than $1 \times 10^{-7}M$.
- The hydrogen-ion concentration is between $1 \times 10^{-3}M$ and $1 \times 10^{-4}M$.
- Thus, the pH should be between 3 and 4.

Checkpoint

Why do we use the pH scale to express hydrogen-ion concentration?

Checkpoint

Why do we use the pH scale to express hydrogen-ion concentration?

It is more convenient and practical to use the pH scale. Expressing hydrogen-ion concentration in molarity takes up a lot of space and is not as easy to work with.

Measuring pH

What are two methods that are used to measure pH?

Either acid-base indicators or pH meters can be used to measure pH.

Acid-Base Indicators

An indicator (HIn) is an acid or a base that dissociates in a known pH range.

- Indicators work because their acid form and base form have different colors in solution.

$$\underset{\text{Acid form}}{HIn(aq)} \rightleftharpoons H^+(aq) + \underset{\text{Base form}}{In^-(aq)}$$

- The acid form of the indicator (HIn) is dominant at low pH and high $[H^+]$.
- The base form (In^-) is dominant at high pH and high $[OH^-]$.

Acid-Base Indicators

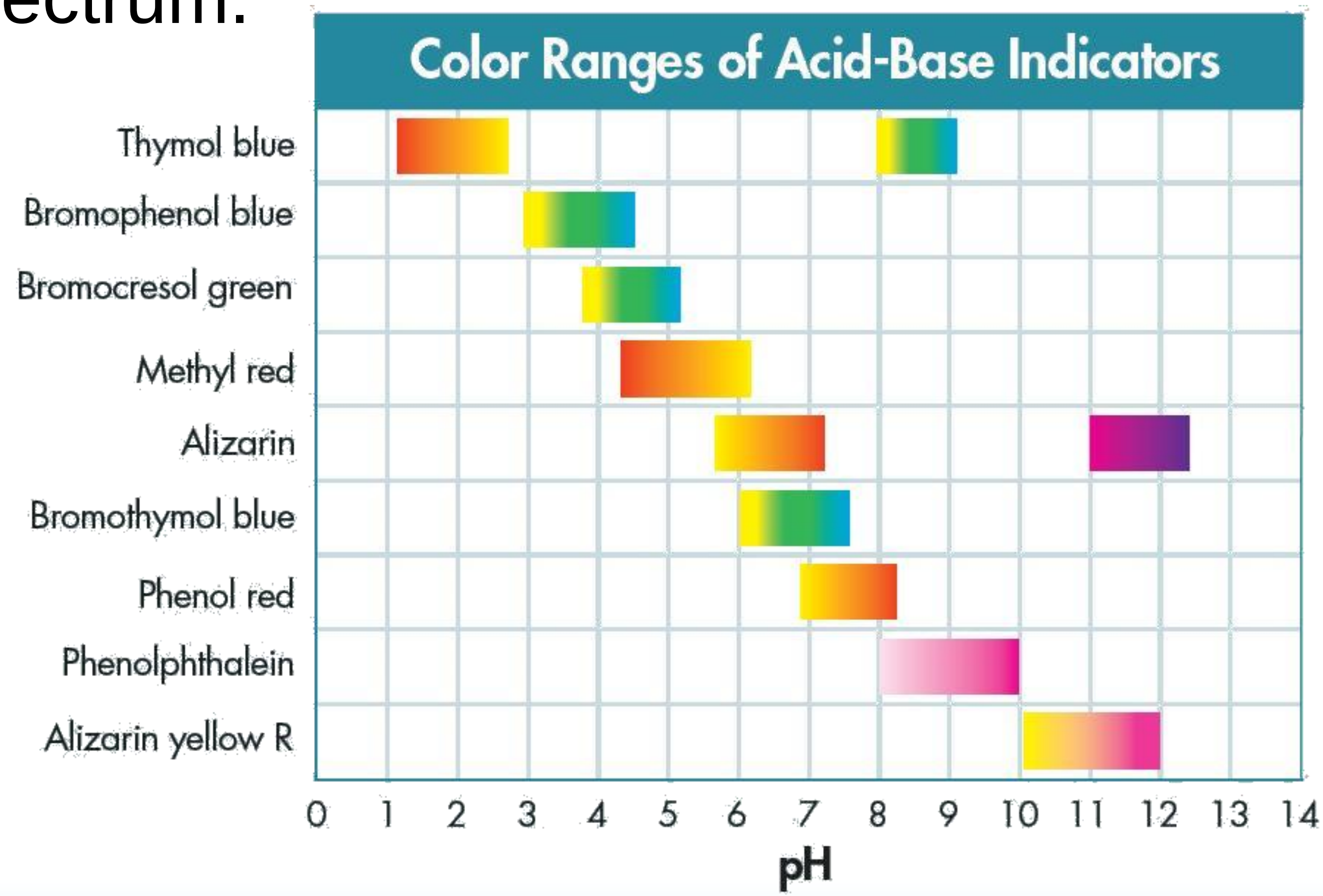
The change from dominating acid form to dominating base form occurs within a narrow range of about two pH units.

$$\underset{\text{Acid form}}{HIn(aq)} \rightleftharpoons H^+(aq) + \underset{\text{Base form}}{In^-(aq)}$$

- At all pH values below the range, you would see only the color of the acid form.
- At all pH values above this range, you would see only the color of the base form.

Acid-Base Indicators

Many indicators are needed to span the entire pH spectrum.



Acid-Base Indicators

An indicator strip is a piece of paper or plastic that has been soaked in an indicator and then dried.

- The paper is dipped into an unknown solution.
- The color that results is compared with a color chart to measure the pH.
- Some indicator paper has absorbed multiple indicators.
  - The colors that result will cover a wide range of pH values.

Acid-Base Indicators

Soil pH can affect how plants develop.

In acidic soils, hydrangeas produce blue flowers.

In basic soils, hydrangeas produce pink flowers.

pH Meters

A pH meter is used to make rapid, continuous measurements of pH.

- The measurements of pH obtained with a pH meter are typically accurate to within 0.01 pH unit of the true pH.
- If the pH meter is connected to a computer or chart recorder, the user will have a record of the pH changes.

pH Meters

A pH meter can be easier to use than liquid indicators or indicator strips.

- The pH reading is visible in a display window on the meter.
- The color and cloudiness of the solution do not affect the accuracy of the pH value obtained.

Checkpoint

You're planting a garden and want to know the approximate pH of your soil. What method should you use?

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You're planting a garden and want to know the approximate pH of your soil. What method should you use?

While you could use a pH meter and have a very accurate reading of your soil's pH, it is also OK to approximate the pH using a pH indicator strip.

- For aqueous solutions, the product of the hydrogen-ion concentration and the hydroxide-ion concentration equals 1×10^{-14}.

- A solution with a pH less than 7.0 is acidic. A solution with a pH of 7 is neutral. A solution with a pH greater than 7.0 is basic.

- Either acid-base indicators or pH meters can be used to measure pH.

- **self-ionization:** a term describing the reaction in which two water molecules react to produce ions
- **neutral solution:** an aqueous solution in which the concentrations of hydrogen and hydroxide ions are equal; it has a pH of 7.0
- **ion-product constant for water (K_w):** the product of the concentrations of hydrogen ions and hydroxide ions in water; it is 1×10^{-14} at 25°C

- **<u>acidic solution</u>:** any solution in which the hydrogen-ion concentration is greater than the hydroxide-ion concentration
- **<u>basic solution</u>:** any solution in which the hydroxide-ion concentration is greater than the hydrogen-ion concentration
- **<u>pH</u>:** a number used to denote the hydrogen-ion concentration, or acidity, of a solution; it is the negative logarithm of the hydrogen-ion concentration of a solution

Reactions

The pH of a solution reflects the hydrogen-ion concentration.

END OF 19.2