Acids and Bases

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Properties of Acids

What is an Acid?

Acids release hydrogen ions into solutions

Acids neutralize bases in a neutralization reaction.

Acids corrode active metals.

Acids turn blue litmus to red.

Acids taste sour.

Properties of Bases

Bases release hydroxide ions into a water solution.

Bases neutralize acids in a neutralization reaction.

Bases denature protein.

Bases turn red litmus to blue.

Bases taste bitter.
Arrhenius Acids and Bases

Arrhenius's definition of acids and bases dates back to the 1800's. It is now considered obsolete since it only relates to reactions in water - aqueous solutions.

Arrhenius defined acids and bases this way:

An **acid** is a substance that, when dissolved in water, increases the concentration of hydrogen ions.

\[ 	ext{HCl} + 	ext{H}_2	ext{O} \rightarrow 	ext{H}_3	ext{O}^+ + 	ext{Cl}^- \]

H\(_3\)O\(^+\) is called a hydrated proton or a hydronium ion.

A **base** is a substance that, when dissolved in water, increases the concentration of hydroxide ions.

\[ 	ext{NH}_3 + 	ext{H}_2	ext{O} \rightarrow 	ext{NH}_4^+ + 	ext{OH}^- \]

Brønsted-Lowry Acids and Bases

The Brønsted-Lowry definition dates back to the early 1900's and is considered the modern definition of acids and bases.

This definition is more general and it works for all reactions; not just in those in water.

An **acid** is a proton, H\(^+\), donor.

A **base** is a proton, H\(^+\), acceptor.

\[ \text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4^+ + \text{Cl}^- \]

Lewis Acids

Brønsted-Lowry acids replaced Arrhenius acids because the former were more general: Arrhenius acids could only be defined in aqueous (water) solutions. Brønsted-Lowry acids don't have that limitation.

Arrehenius acids - only substances dissolved in H\(_2\)O

Similarly, Brønsted-Lowry acids are limited to substances that gain or lose hydrogen.

Brønsted-Lowry acids - only substances that gain or lose H\(^+\)

The most general approach is that of Lewis acids; which do not require an aqueous environment or an exchange of hydrogen.

**Lewis acids** are defined as electron-pair acceptors. Atoms with an empty valence orbital can be Lewis acids.

\[ 	ext{H}_3\text{C} + \text{OH}_2 \rightarrow 	ext{H}_3\text{C} - \text{O}^- \]

Lewis Acid
Lewis Bases

Lewis bases are defined as electron-pair donors. Anything that could be a Brønsted-Lowry base is a Lewis base.

Lewis bases can interact with things other than protons, however. Therefore, this definition is the broadest of the three.

C\[\text{CH}_3\]H\[\text{CH}_3\]H\[\text{OH}\]\[\text{CH}_3\text{CH}_3\]H

Lewis Base

1. A Brønsted-Lowry base is defined as a substance that __________.
   - A increases [H+] when placed in H\(_2\)O
   - B decreases [H+] when placed in H\(_2\)O
   - C increases [OH\(-\)] when placed in H\(_2\)O
   - D acts as a proton acceptor
   - E acts as a proton donor

2. A Brønsted-Lowry acid is defined as a substance that __________.
   - A increases K\(_a\) when placed in H\(_2\)O
   - B decreases [H\(^+\)] when placed in H\(_2\)O
   - C increases [OH\(-\)] when placed in H\(_2\)O
   - D acts as a proton acceptor
   - E acts as a proton donor

3. Which of the following compounds could never act as an acid?
   - A SO\(_4^{2-}\)
   - B HSO\(_4^{-}\)
   - C H\(_2\)SO\(_4\)
   - D NH\(_3\)
   - E CH\(_3\)COOH

4. According to the following reaction model, reactant is acting like an acid?
   - H\(_2\)O + H\(_2\)SO\(_4\) \rightarrow H\(_3\)O\(^+\) + HSO\(_4^{-}\)
   - A H\(_2\)SO\(_4\)
   - B H\(_2\)O
   - C H\(_3\)O\(^+\)
   - D HSO\(_4^{-}\)
   - E None of the above

5. According to the following reaction, which reactant is acting like a base?
   - H\(_3\)O\(^+\) + HSO\(_4^{-}\) \rightarrow H\(_2\)O + H\(_2\)SO\(_4\)
   - A H\(_2\)SO\(_4\)
   - B H\(_2\)O
   - C H\(_3\)O\(^+\)
   - D HSO\(_4^{-}\)
   - E None of the above
6 For the following reaction, identify whether the compound in bold is behaving as an acid or a base.

\[ \text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{PO}_4^- + \text{H}_3\text{O}^+ \]

- A Acid
- B Base
- C Neither
- D Both
- E None of the above

7 For the following reaction identify whether the compound in bold is behaving as an acid or a base.

\[ \text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{PO}_4^- + \text{H}_3\text{O}^+ \]

- A Acid
- B Base
- C Both
- D Neither
- E None of the above

8 Which of the following cannot act as a Lewis base?

- A Cl\(^-\)
- B NH\(_3\)
- C BF\(_3\)
- D CN\(^-\)
- E H\(_2\)O

9 In the reaction BF\(_3\) + F\(^-\) → BF\(_4\)^-, BF\(_3\) acts as a/an ____________ acid.

- A Arrhenius
- B Bronsted-Lowry
- C Lewis
- D Arrhenius, Bronsted-Lowry, and Lewis
- E Arrhenius and Bronsted-Lowry

Conjugate Acids and Bases

The term conjugate comes from the Latin word “conjugare,” meaning “to join together.”

Reactions between acids and bases always yield their conjugate bases and acids.
Conjugate Acids and Bases

- After the acid donates a proton, the result is called its conjugate base.
- After the base accepts a proton, the result is called its conjugate acid.

\[
\text{HNO}_2(aq) + \text{H}_2\text{O}(l) \rightarrow \text{NO}_2^-(aq) + \text{H}_3\text{O}^+(aq)
\]

Acids in Water

What Happens When an Acid Dissolves in Water?

- Water acts as a Brønsted-Lowry base and takes a proton (H+) from the acid.
- As a result, the conjugate base of the acid and a hydronium ion are formed.

\[
\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-
\]

Acids in Water

- Which is the acid?
- Which is the base?

10. In liquid ammonia, the reaction represented below occurs. In the reaction \(\text{NH}_4^+\) acts as

\[
2\text{NH}_3 \rightarrow \text{NH}_4^+ + \text{NH}_2
\]

- a catalyst
- both an acid and a base
- the conjugate acid of \(\text{NH}_3\)
- the reducing agent
- the oxidizing agent

Question from the College Board

11. What are the conjugate bases of \(\text{HClO}_4\), \(\text{H}_2\text{S}\), \(\text{PH}_3^+\), \(\text{HCO}_3^-\)?

- \(\text{ClO}_4^-, \text{HS}^-, \text{PH}_3^-, \text{CO}_3^2\)
- \(\text{ClO}_4^-, \text{HS}^-, \text{PH}_3^+, \text{CO}_2^2\)
- \(\text{ClO}_2^-, \text{HS}^2-, \text{PH}_3^+, \text{CO}_2^2\)
- \(\text{ClO}_3^-, \text{H}_2\text{S}_2^-, \text{PH}_3^-, \text{CO}_2^2\)
- \(\text{ClO}_4^-, \text{H}_2\text{S}_2^-, \text{PH}_3^-, \text{CO}_2^2\)

Acid and Base Strength

- Strong acids completely dissociate in water.
- Their conjugate bases are quite weak.

- Weak acids only partially dissociate in water.
- Their conjugate bases are weak bases.
In this example, H_2O is a much stronger base than Cl\(^-\) so the proton moves from HCl to H_2O.

Conversely, HCl is a much stronger acid than H_2O. Their conjugate bases are very strong.

K >> 1

Acid and Base Strength

Substances with negligible acidity do not dissociate in water. They will not readily give up protons.

Their conjugate bases are very strong.

In any acid-base reaction, equilibrium will favor the reaction in which the proton moves toward the stronger base.

In other words, a stronger base will "hold onto" its proton whereas a strong acid easily releases its proton(s).

HCl (aq) + H_2O (l) \rightarrow H_3O^+ (aq) + Cl^- (aq)

Acid base conj. acid conj. base

An alternative way to consider equilibrium is that it will favor the reaction AWAY from the stronger acid.

Acid and Base Strength

Consider this equilibrium between acetic acid and acetate ion:

CH_3COOH (aq) + H_2O (l) \rightarrow H_3O^+ (aq) + CH_3COO^- (aq)

Does equilibrium lie to the left (K<1) or to the right (K>1)?

If you look for the stronger acid: Equilib lies away from the stronger acid.

If you look for the stronger base: Equilib favors this base accepting a proton.

Acid and Base Strength

Acetic acid is a weak acid. This means that only a small percent of the acid will dissociate.

The double headed arrow is used only in weak acid or weak base dissociation equations.

CH_3COOH (aq) + H_2O (l) \rightarrow H_3O^+ (aq) + CH_3COO^- (aq)

A single arrow is used for strong acid or strong bases which dissociate completely.

NaOH \rightarrow Na^+ (aq) + OH^- (aq)
12 Strong acids have ___________ conjugate bases.

- A strong
- B weak
- C neutral
- D negative

13 HBr, hydrobromic acid is a strong acid. This means that it ___________.

- A aqueous solutions of HBr contain equal concentrations of H⁺ and OH⁻
- B does not dissociate at all when it is dissolved in water
- C cannot be neutralized by a base
- D dissociates completely to H⁺ and Br⁻ when it dissolves in water

14 For the following reaction, determine which side of the equilibrium is favored.

\[ \text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{PO}_4^- + \text{H}_3\text{O}^+ \]

- A the right side
- B the left side
- C Neither side is favored

15 For the following reaction, determine which side of the equilibrium is favored.

\[ \text{H}_3\text{O}^+ + \text{HSO}_4^- \rightarrow \text{H}_2\text{O} + \text{H}_2\text{SO}_4 \]

- A the right side
- B the left side
- C Neither side is favored

16 For the following reaction, determine which side of the equilibrium is favored.

\[ \text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^- \]

- A the right side
- B the left side
- C Neither side is favored
If a substance can act both as an acid and base, it is known as amphoteric. For example, water can act as a base or acid depending on the situation.

\[ \text{HCl} + \text{H}_2\text{O} \rightarrow \text{Cl}^- + \text{H}_3\text{O}^+ \]

Above, water accepts a proton, thus acting as a base.

\[ \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^- \]

Above, water donates a proton, thus acting as an acid.

Another term for amphoteric is amphiprotic.

For each of the following substances, write two equations, one showing it as a Bronsted-Lowry acid and another showing it as a Bronsted-Lowry base.

17. A substance that is capable of acting as both an acid and as a base is _________.

- A. autosomal
- B. conjugated
- C. amphoteric
- D. saturated
- E. miscible

18. Write the equations and equilibrium expressions for HS\(^-\) when it is acting like a Brønsted-Lowry acid and when it is acting like a Brønsted-Lowry base.

Strong Acids and Bases

Recall, strong acids completely ionize in solution.

The seven strong acids are:

- HCl  hydrochloric acid
- HBr  hydrobromic acid
- HI   hydroiodic acid
- HNO\(_3\)  nitric acid
- H\(_2\)SO\(_4\)  sulfuric acid
- HClO\(_3\)  chloric acid
- HClO\(_4\)  perchloric acid

Memorize this list.
**Strong Acids**

The seven strong acids are strong electrolytes because they are 100% ionized. In other words, these compounds exist totally as ions in aqueous solution.

For the monoprotic strong acids (acids that donate only one proton per molecule of the acid), the hydronium ion concentration equals the acid concentration.

\[ [\text{Acid}] = [\text{H}_3\text{O}^+] \]

So, if you have a solution of 0.5 M HCl, then \([\text{H}_3\text{O}^+] = 0.5 \text{ M}\)

**Strong Bases**

All strong bases are group of compounds called "metal hydroxides."

- All alkali metals in Group I form hydroxides that are strong bases: LiOH, NaOH, KOH, etc.
- Only the heavier alkaline earth metals in Group II form strong bases: Ca(OH)₂, Sr(OH)₂, and Ba(OH)₂.

Again, these substances dissociate completely in aqueous solution. In other words, NaOH exists entirely as Na⁺ ions and OH⁻ ions in water.

---

19. What is the hydroxide ion concentration of a 0.22 M calcium hydroxide solution?

- A 0.11
- B 0.22
- C 0.44
- D 0.88
- E Not enough information.

20. What is the concentration of H⁺ in a 25ml solution of 0.05M HCl when diluted to final volume of 100ml?

21. What is the [H⁺] ion concentration of a 50 ml solution of 0.025M H₂SO₄, when diluted with 150 ml of water?

22. A solution of 25 ml of 0.1M HCl and 50 ml of 0.5M HNO₃ are mixed together. What is the [H⁺] ion concentration of the resulting solution?
Auto-ionization of Water

As we have seen, water is amphoteric, meaning that it can act as either an acid or a base.

In pure water, a few molecules act as bases and a few act as acids, in a process referred to as autoionization.

\[ \text{H}_2\text{O} (\text{l}) + \text{H}_2\text{O} (\text{l}) \rightarrow \text{H}_3\text{O}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \]

The double headed arrow indicates that both the forward and reverse reactions occur simultaneously.

When there is an equilibrium state, the ratio of products to reactants yields a constant. This value is known as the equilibrium constant, \( K \) and will be discussed in more depth later in this unit.

All concentrations are in M, molarity, as designated by brackets, [ ]:

\[ K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2} \]

This special equilibrium constant, \( K_w \) is referred to as the ion-product constant for water.

At 25°C, \( K_w = 1.0 \times 10^{-14} \). Since this is such a small number, we conclude that pure water contains relatively very few ions.

The magnitude of \( K_w \) indicates that ________ .

- A water ionizes to a very small extent
- B the autoionization of water is exothermic
- C water ionizes very quickly
- D water ionizes very slowly
24. The ion-product constant for water, $K_w$, is represented by

- **A** $[H_2O]^2$
- **B** $[H_3O^+] x [OH^-]$
- **C** $[H_3O^+] + [OH^-]$
- **D** $[H_3O^+] - [OH^-]$

25. What is the $[H^+]$ ion concentration of a solution with an $[OH^-]$ ion concentration of $1 \times 10^{-9}$?

*Try to solve without a calculator!*

- **A** $1 \times 10^{-14}$
- **B** $1 \times 10^{5}$
- **C** $1 \times 10^{-5}$
- **D** $1.23 \times 10^{-5}$
- **E** $1 \times 10^{-7}$

26. What is the $[OH^-]$ ion concentration of a solution with an $[H^+]$ ion concentration of $1 \times 10^{-4}$?

*Try to solve without a calculator!*

- **A** $1 \times 10^{-14}$
- **B** $1 \times 10^{-10}$
- **C** $1 \times 10^{-5}$
- **D** $1.23 \times 10^{-5}$
- **E** $1 \times 10^{-10}$

27. What is the $[OH^-]$ ion concentration of a solution with an $[H^+]$ ion concentration of $1.23 \times 10^{-9}$?

- **A** $1 \times 10^{-14}$
- **B** $4.34 \times 10^{-6}$
- **C** $1.23 \times 10^{-5}$
- **D** $1.23 \times 10^{-5}$
- **E** $8.13 \times 10^{-6}$

**pH**

It is a measure of hydrogen ion concentration, $[H^+]$ in a solution, where the concentration is measured in moles $H^+$ per liter, or molarity.

The pH scale ranges from 0-14.

$pH$ is defined as the negative base-10 logarithm of the concentration of hydronium ion.

$$pH = -\log [H_3O^+]$$
**pH**

pH is defined as the negative base-10 logarithm of the concentration of hydronium ion.

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

<table>
<thead>
<tr>
<th>Hydrogen ion concentration, [H+] in moles/Liter</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0 x 10^{-1}</td>
<td>1</td>
</tr>
<tr>
<td>1.0 x 10^{-2}</td>
<td>2</td>
</tr>
<tr>
<td>1.0 x 10^{-10}</td>
<td>10</td>
</tr>
</tbody>
</table>

Is the relationship between [H+] and pH a direct or an inverse one?

**pH**

Because of the base-10 logarithm, each 1.0-point value on the pH scale differs by a value of ten.

A solution with pH = 9 has a hydrogen ion concentration, [H+], that is ten times more than a pH = 10 solution.

A solution with pH = 8 has a hydrogen ion concentration, [H+], that is 10^2 or 100 times more than a pH = 10 solution.

A solution with pH = 7 has a hydrogen ion concentration, [H+], that is 10^3 or 1000 times more than a pH = 10 solution.

---

28 The molar concentration of hydronium ion, [H_3O^+], in pure water at 25 °C is ________.

- **A** 0
- **B** 1
- **C** 7
- **D** 10^{-7}
- **E** 10^{-14}

---

29 A solution with pH = 3 has a hydrogen ion concentration that is ________ than a solution with pH = 5.

- **A** 2x more
- **B** 2x less
- **C** 100x more
- **D** 100x less

---

30 A solution with pH = 14 has a hydrogen ion concentration that is ________ than a solution with pH = 11.

- **A** 3x more
- **B** 3x less
- **C** 1000x more
- **D** 1000x less
Therefore, in pure water

\[
pH = -\log [H_3O^+] = 7.00
\]

\[
pH = -\log (1.0 \times 10^{-7}) = 7.00
\]

An acid has a higher $[H_3O^+]$ than pure water, so its pH is $<7$.

A base has a lower $[H_3O^+]$ than pure water, so its pH is $>7$.

<table>
<thead>
<tr>
<th>Solution type</th>
<th>$[H^+]$ (M)</th>
<th>$[OH^-]$ (M)</th>
<th>pH value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acidic</td>
<td>$&gt;1.0 \times 10^{-7}$</td>
<td>$&lt;1.0 \times 10^{-7}$</td>
<td>$&lt;7.00$</td>
</tr>
<tr>
<td>Neutral</td>
<td>$=1.0 \times 10^{-7}$</td>
<td>$=1.0 \times 10^{-7}$</td>
<td>$=7.00$</td>
</tr>
<tr>
<td>Basic</td>
<td>$&lt;1.0 \times 10^{-7}$</td>
<td>$&gt;1.0 \times 10^{-7}$</td>
<td>$&gt;7.00$</td>
</tr>
</tbody>
</table>

For a basic solution, the hydrogen ion concentration is __________ than the hydroxide ion concentration.

- A greater than
- B less than
- C equal to
- D Not enough information.

For an acidic solution, the hydroxide ion concentration is __________ than the hydrogen ion concentration.

- A greater than
- B less than
- C equal to
- D Not enough information.

Which solution below has the highest concentration of hydroxide ions?

- A pH = 3.21
- B pH = 7.00
- C pH = 8.93
- D pH = 12.6

These are the pH values for several common substances.
34 Which solution below has the lowest concentration of hydrogen ions?

- A  pH = 1.98
- B  pH = 8.53
- C  pH = 5.91
- D  pH = 11.4

35 For a 1.0-M solution of a weak base, a reasonable pH would be_____.

- A  2
- B  6
- C  7
- D  9
- E  13

36 For a 1.0-M solution of a strong acid, a reasonable pH would be_____.

- A  2
- B  6
- C  7
- D  9
- E  13

37 The pH of a solution with a concentration of 0.01M hydrochloric acid is

- A  $10^{-2}$
- B  12
- C  2
- D  $10^{-12}$

38 A solution with the pH of 5.0

- A  is basic
- B  has a hydrogen ion concentration of 5.0M
- C  is neutral
- D  has a hydroxide-ion concentration of $1 \times 10^{-6}$

How Do We Measure pH?

For less accurate measurements, one can use Litmus paper:
- "Red" litmus paper turns blue above ~pH = 8
- "Blue" litmus paper turns red below ~pH = 5

Or an indicator (usually an organic dye) such as one of the following:

- Methyl violet
- Thymol blue
- Methyl orange
- Methyl red
- Bromothymol blue
- Phenolphthalein
- Alizarin yellow R

pH range for color change
How Do We Measure pH?

For more accurate measurements, one uses a pH meter, which measures the voltage in the solution.

How Do We Calculate pH?

Recall that pH is defined as the negative base-10 logarithm of the concentration of hydronium ion (or hydrogen ion).

\[ \text{pH} = -\log [\text{H}_3\text{O}^+] \quad \text{or} \quad \text{pH} = -\log [\text{H}^+] \]

How Do We Calculate pH?

What is the pH of the solution with hydrogen ion concentration of \(5.67 \times 10^{-8}\) M (molar)?

\[ \text{pH} = -\log [\text{H}^+] \]

First, take the log of \(5.67 \times 10^{-8}\) = -7.246

Now, change the sign from - to +

Answer: \(\text{pH} = 7.246\)

If you take the log of \(-5.67 \times 10^{-8}\), you will end up with an incorrect answer.

The order of operations: 1. Take the log
2. Switch the sign

What is the pH of the solution with a hydrogen ion concentration of \(1 \times 10^{-5}\) M?

- A \(1 \times 10^{-5}\)
- B -5
- C 5
- D 9

What is the pH of a solution with a hydroxide ion concentration of \(1 \times 10^{-12}\) M?

- A \(1 \times 10^2\)
- B 12
- C 2
- D -12

What is the pH of an aqueous solution at 25.0 °C in which [H⁺] is 0.0025 M?

- A 3.4
- B 2.6
- C -2.6
- D -3.4
- E 2.25
**Additional pH Calculations**

If you are given the pH and asked to find the \([H^+]\) (or \([H_3O^+]\)) in a solution, use the inverse log.

Since \(pH = -\log [H^+]\), then \([H^+] = 10^{-\text{pH}}\)

What is the hydrogen ion concentration (M) in a solution of Milk of Magnesia whose \(pH = 9.8\)?

\[ [H^+] = 10^{-9.8} \]
\[ [H^+] = 1.58 \times 10^{-10} \text{ M or mol/Liter} \]

42 What is the pH of a solution whose hydronium ion concentration is \(7.14 \times 10^{-3} \text{ M}\)?

43 What is the pH of a solution whose hydronium ion concentration is \(1.92 \times 10^{-9} \text{ M}\)?

44 What is the hydronium ion concentration in a solution whose \(pH = 4.29\)?

45 What is the hydroxide ion concentration in a solution whose \(pH = 4.29\)?

**Other “p” Scales**

The "p" in pH tells us to take the negative base-10 logarithm of the quantity (in this case, hydronium ions).

Some similar examples are

\[ pOH = -\log [OH^-] \]
\[ pK_w = -\log K_w \]
\[ pK_a = -\log K_a \]
\[ pK_b = -\log K_b \]
Relationship between pH and pOH

Because 
\[ [\text{H}_3\text{O}^+] [\text{OH}^-] = \text{K}_w = 1.0 \times 10^{-14}, \]
we know that

\[ -\log [\text{H}_3\text{O}^+] + -\log [\text{OH}^-] = -\log \text{K}_w = 14.00 \]
or, in other words,

\[ \text{pH} + \text{pOH} = \text{pK}_w = 14.00 \]

46 An aqueous solution of a base has a pH 6.12 what is its pOH?

- A 4.22
- B 8.88
- C 7.88
- D 2.11
- E 7.59 \times 10^{-7}

47 An aqueous solution of an acid has a hydrogen ion concentration of 2.5 \times 10^{-4}. What is the pOH of this solution?

- A 10.4
- B 6.13
- C 8.4
- D 7.5
- E 9.4

48 An aqueous solution of an acid has pH of 4.11 what is the \([\text{OH}^-]\) concentration of this solution?

- A 4.55 \times 10^{11}
- B 5.78 \times 10^{11}
- C 4.25 \times 10^{8}
- D 1.29 \times 10^{-10}
- E 4.03 \times 10^7

49 An aqueous solution of an base has pOH of 3.33 what is the \([\text{H}^+]\) concentration of this solution?

- A 4.68 \times 10^4
- B 2.14 \times 10^{11}
- C 5.67 \times 10^8
- D 9.07 \times 10^9
- E 4.88 \times 10^7

Weak Acids and Bases

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Acid Dissociation Constants, $K_a$

For a generalized acid dissociation, the equilibrium expression is

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

This equilibrium constant is called the acid-dissociation constant, $K_a$.

**Calculating $K_a$ from the pH**

The pH of a 0.10 M solution of formic acid, HCOOH, at 25°C is 2.38. Calculate $K_a$ for formic acid at this temperature.

The dissociation equation for formic acid may be written as a reaction with water

$$\text{HCOOH} + \text{H}_2\text{O} \rightleftharpoons \text{HCOO}^- + \text{H}_3\text{O}^+$$

or, without water

$$\text{HCOOH} \rightleftharpoons \text{HCOO}^- + \text{H}^+$$

From this dissociation equation, write the $K_a$ expression:

$$K_a = \frac{[H_3O^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

To calculate $K_a$, we need the equilibrium concentrations of all three species.

We know the concentration of HCOOH, how do we determine the concentration of $H_3O^+$?

**Calculating $K_a$ from the pH**

The pH of a 0.10 M solution of formic acid, HCOOH, at 25°C is 2.38. Calculate $K_a$ for formic acid at this temperature.

The dissociation equation for formic acid may be written as a reaction with water

$$\text{HCOOH} + \text{H}_2\text{O} \rightleftharpoons \text{HCOO}^- + \text{H}_3\text{O}^+$$

or, without water

$$\text{HCOOH} \rightleftharpoons \text{HCOO}^- + \text{H}^+$$

From this dissociation equation, write the $K_a$ expression:

$$K_a = \frac{[H_3O^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

To calculate $K_a$, we need the equilibrium concentrations of all three species.

We know the concentration of HCOOH, how do we determine the concentration of $H_3O^+$?
Calculating $K_a$ from the pH

The pH of a 0.10 M solution of formic acid, HCOOH, at 25°C is 2.38. Calculate $K_a$ for formic acid at this temperature.

$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$

<table>
<thead>
<tr>
<th>Initially</th>
<th>[HCOOH], M</th>
<th>[HCOO^-], M</th>
<th>[H_3O^+], M</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.10</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-4.2 x 10^{-3}</td>
<td>+4.2 x 10^{-3}</td>
<td>+4.2 x 10^{-3}</td>
</tr>
<tr>
<td>At Equilibrium</td>
<td>0.10 - 4.2 x 10^{-3} = 0.10</td>
<td>4.2 x 10^{-3}</td>
<td>4.2 x 10^{-3}</td>
</tr>
</tbody>
</table>

Calculating Percent Ionization

One way to compare the strength of two acids is by the extent to which each one ionizes. This is done by calculating percent ionization, or the ratio of $[H^+]$ ions that are produced, compared to the original acid concentration.

Percent Ionization = \frac{[H_3O^+]_{eq}}{[HA]_{initial}} \times 100\%

Calculating Percent Ionization

In this example

$[H_3O^+]_{eq} = 4.2 \times 10^{-3}$ M

$[HCOOH]_{initial} = 0.10$ M

Percent Ionization = \frac{4.2 \times 10^{-3}}{0.10} \times 100\% = 4.2\%

51. What is the $K_a$ of a 0.125M solution of hypobromous acid (HBrO) that has a pH of 4.74.

- $2.42 \times 10^{-4}$
- $2.65 \times 10^{-9}$
- $3.71 \times 10^{-10}$
- $2.13 \times 10^{-5}$
- $6.78 \times 10^{-12}$
### Slide 115 / 208
52. What is % ionization of a 0.125M solution of hypobromous acid (HBrO) that has a pH of 4.74.

- 1.45%
- 68.89%
- 0.145%
- 0.0145%
- 6.889%

### Slide 116 / 208
53. What is the $K_a$ of a 0.20M solution of nitrous acid (HNO$_2$) that has a pH of 2.02?

- $1.3 \times 10^{-5}$
- $9.9 \times 10^{-2}$
- $1.2 \times 10^{-4}$
- $4.2 \times 10^{-7}$
- $4.6 \times 10^{-4}$

### Slide 117 / 208
54. What is % ionization of a 0.20M solution of nitrous acid (HNO$_2$) that has a pH of 2.02?

- 0.0477
- 4.77%
- 5.99%
- 0.599%
- 6.889%

### Slide 118 / 208
55. What is the $K_a$ of a 0.115M solution of a weak acid that has a 2.55% ionization?

- $4.6 \times 10^{-3}$
- $7.45 \times 10^{-1}$
- $7.48 \times 10^{-5}$
- $9.0 \times 10^{-4}$
- $1.2 \times 10^{-6}$

### Slide 119 / 208
56. If the acid dissociation constant, $K_a$ for an acid HA is $8 \times 10^{-4}$ at 25°C, what percent of the acid is dissociated in a 0.50-molar solution of HA?

- A 0.08%
- B 0.2%
- C 1%
- D 2%
- E 4%

*Question from the College Board*

### Slide 120 / 208
**Calculating pH from $K_a$**

Calculate the pH of a 0.30 M solution of acetic acid, HCO$_2$H$_2$O$_2$ at 25°C. $K_a$ for acetic acid at 25°C is $1.8 \times 10^{-5}$.

First, we write the dissociation equation for acetic acid

HCO$_2$H$_2$O$_2$ (aq) + H$_2$O (l) $\rightarrow$ H$_3$O$^+$ (aq) + C$_2$H$_3$O$_2^-$ (aq)
### Calculating pH from $K_a$

Calculate the pH of a 0.30 M solution of acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, at 25°C. $K_a$ for acetic acid at 25°C is $1.8 \times 10^{-5}$.

From the dissociation equation, we obtain the equilibrium constant expression:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

#### Calculating pH

We next set up an ICE chart...

<table>
<thead>
<tr>
<th>Initial</th>
<th>$\text{HC}_2\text{H}_3\text{O}_2$, M</th>
<th>$[\text{H}_3\text{O}^+]$, M</th>
<th>$[\text{C}_2\text{H}_3\text{O}_2^-]$, M</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>0.30 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>- $x$</td>
<td>+ $x$</td>
<td>+ $x$</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>about 0.30 M</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

We are assuming that $x$ will be very small compared to 0.30 and can, therefore, be ignored.

#### Calculating pH expression

$$1.8 \times 10^{-5} = \frac{(x)^2}{(0.30)}$$

$$(1.8 \times 10^{-5}) (0.30) = x^2$$

$5.4 \times 10^{-6} = x^2$  

$2.3 \times 10^{-3} = x$  

Remember what your "$x$" is! In this case, it’s $[\text{H}_3\text{O}^+]$, but other times it’s $[\text{OH}]$.

### Significant figure rules for pH on the AP exam:

- The calculated pH value should have as many DECIMAL places as the $[\text{H}^+]$ has sig figs.
- So if the $[\text{H}^+]$ has 2 sig figs, report the pH to the 0.01 place value.
- If the $[\text{H}^+]$ has 3 sig figs, report the pH to the 0.001 place value.

### Question 57

What is the hydrogen ion concentration of a weak acid that has a dissociation constant is $1 \times 10^{-6}$ and a concentration of 0.01M?

- **A**  $1 \times 10^{-6}$
- **B**  $1 \times 10^{-5}$
- **C**  $1 \times 10^{-4}$
- **D**  $1 \times 10^{-3}$

### Question 58

What is the pH of an acid that has an acid dissociation constant $K_a$ of $3.2 \times 10^4$ and the acid concentration is 0.122M?

- **A**  2.00
- **B**  2.20
- **C**  2.50
- **D**  2.17
59. What is the concentration of the acid if the pH is 4 and the $K_a$ is $1 \times 10^{-7}$?

- A. $1 \times 10^{-7}$
- B. $1 \times 10^{-5}$
- C. $1 \times 10^{-3}$
- D. $1 \times 10^{-1}$

60. What is the pH of a 0.05M solution of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) that has a % ionization of 1.22%?

- 1.15
- 4.22
- 3.21
- 2.13
- 6.78

---

**Weak Bases**

Bases react with water to produce a hydroxide ion.

Even though $\text{NH}_3$ does not have the hydroxide ion, $\text{OH}^-$, in its formula, it is a base according to both the Arrhenius and Bronsted-Lowry definitions.

$\text{B}^- + \text{H}_2\text{O} \rightleftharpoons \text{HB} + \text{OH}^-$

The equilibrium constant expression for this reaction is

$$K_b = \frac{[\text{HB}][\text{OH}^-]}{[\text{B}^-]}$$

where $K_b$ is the base-dissociation constant.

Just as for $K_a$, the stronger a base is, it will have a higher $K_b$ value.

In fact, since the strong bases dissociate 100%, their $K_b$ values are referred to as "very large".

$K_b$ can be used to find $[\text{OH}^-]$ and, ultimately, pH.

---

61. Which base has the smallest base dissociation constant, $K_b$?

- A. potassium hydroxide
- B. sodium hydroxide
- C. calcium hydroxide
- D. ammonia

62. A base has a dissociation constant, $K_b = 2.5 \times 10^{-11}$. Which of the following statements is true?

- A. This is a concentrated base.
- B. This base ionizes slightly in aqueous solution.
- C. This is a strong base.
- D. An aqueous solution of this would be acidic.
Calculating $K_b$ from pH

What is the $K_b$ of a 0.20 M solution of hydrazine $H_2NNH_2$ at 25°C that has a pH of 10.9?

First, we write the dissociation equation for hydrazine

$$H_2NNH_2(aq) + H_2O (l) \rightarrow OH^- (aq) + H_2NNH_3^+(aq)$$

From the dissociation equation, we obtain the equilibrium constant expression:

$$K_b = \frac{[OH^-][H_2NNH_3^+]}{[H_2NNH_2]}$$

Next, we substitute values into the $K_b$ expression and solve:

$$\text{pOH} = 14 - \text{pH}$$

$$\text{pOH} = 14 - 10.9$$

$$\text{pOH} = 3.1$$

$$3.1 = -\log [OH^-]$$

$$-3.1 = \log [OH^-]$$

$$10^{-3.1} = 10^{\log[OH^-]} = [OH^-]$$

$$7.94 \times 10^{-4} = [OH^-] = [H_2NNH_3^+]$$

Calculating Percent Ionization

What would be the analogous formula to calculate percent ionization for a base?

$$\text{Percent Ionization} = \frac{[OH^-]_{eq}}{[\text{Base}]_{initial}} \times 100\%$$

63 Calculate the $K_b$ of a 0.450M solution of weak base solution with a pOH of 4.98.

- A. $3.22 \times 10^{-7}$
- B. $2.11 \times 10^{-5}$
- C. $2.03 \times 10^{-18}$
- D. $2.33 \times 10^{-5}$
- E. $2.44 \times 10^{-10}$
64. Calculate the $K_b$ of a 0.724M solution of hypobromite ion ($\text{BrO}^-$) that has a pH of 11.23.

- $4.80 \times 10^{-23}$
- $4.00 \times 10^{-6}$
- $2.35 \times 10^{-3}$
- $1.22 \times 10^{-4}$
- $6.10 \times 10^{-5}$

65. What is the $K_b$ of a 0.125M solution of a weak base that is 1.25% ionized?

**Calculating pH from the $K_b$**

The same process we followed to determine $K_a$ from the pH we can use to determine $K_b$ from the pH.

What is the pH of a 0.15 M solution of $\text{NH}_3$?

$\text{NH}_3 \ (\text{aq}) + \text{H}_2\text{O} \ (\text{l}) \ \rightarrow \ \text{NH}_4^+ \ (\text{aq}) + \text{OH}^- \ (\text{aq})$

First write the equilibrium expression for the dissociation equation. Obtain the $K_b$ value from an earlier page.

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$

$$[\text{OH}^-] = 1.6 \times 10^{-3} \text{ M}$$

$$pOH = -\log (1.6 \times 10^{-3})$$

$$pOH = 2.80$$

So, now solving for pH:

$$pH = 14.00 - 2.80$$

$$pH = 11.20$$
66 What is the OH\(^-\) concentration of a 0.550 M solution of the weak base aniline, C\(_6\)H\(_5\)NH\(_2\)? (K\(_b\) = 4.3 \times 10^{-10})

- 2.37 \times 10^{-11}
- 1.54 \times 10^{-5}
- 7.69 \times 10^{-10}
- 1.30 \times 10^{-11}

67 What is the pH of a 0.550 M solution of the weak base aniline, C\(_6\)H\(_5\)NH\(_2\)? (K\(_b\) = 4.3 \times 10^{-10})

- 2.37 \times 10^{-11}
- 1.54 \times 10^{-5}
- 7.69 \times 10^{-10}
- 1.30 \times 10^{-11}

68 What is the pH of a 0.135 M solution of trimethylamine, (CH\(_3\))\(_3\)N? (K\(_b\) = 6.4 \times 10^{-5})

- A 8.94
- B 11.47
- C 2.53
- D 5.06
- E 3.45

69 What is the pH of a 0.112 M solution of ethylamine (C\(_5\)H\(_5\)NH\(_2\))? (K\(_b\) = 6.4 \times 10^{-4})

- Students type their answers here

---

**Polyprotic Acids**

Polyprotic acids are characterized by having more than one acidic proton.

Here are some examples:

<table>
<thead>
<tr>
<th>Acid</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sulfuric acid</td>
<td>H(_2)SO(_4)</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H(_3)PO(_4)</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>H(_2)CO(_3)</td>
</tr>
<tr>
<td>Oxalic acid</td>
<td>H(_2)C(_2)O(_4)</td>
</tr>
</tbody>
</table>
**Polyprotic Acids**

Polyprotic acids have a $K_a$ value for each proton that can be removed. For example, consider carbonic acid, $\text{H}_2\text{CO}_3$.

The first ionization equation for carbonic acid is:

$$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightarrow \text{HCO}_3^- + \text{H}_3\text{O}^+$$

Write the $K_a$ expression for this equation. This is referred to as $K_{a1}$.

Now, we examine both dissociation constants.

The second ionization equation for carbonic acid is:

$$\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{CO}_3^{2-} + \text{H}_3\text{O}^+$$

Write the $K_a$ expression for this equation. This is referred to as $K_{a2}$.

**Polyprotic Acids**

Notice that the bicarbonate ion, $\text{HCO}_3^-$, appears in each expression; in the numerator for $K_{a1}$ and in the denominator for $K_{a2}$.

$$K_{a1} \times K_{a2} = \frac{[\text{HCO}_3^-][\text{H}_3\text{O}^+][\text{CO}_3^{2-}][\text{H}_2\text{O}]}{[\text{H}_2\text{CO}_3][\text{HCO}_3^-]}$$

$$K_{a1} \times K_{a2} = \frac{[\text{H}_2\text{O}][\text{CO}_3^{2-}][\text{H}_3\text{O}^+][\text{H}_2\text{CO}_3]}{[\text{HCO}_3^-][\text{RSQ}_2]}$$

$K_{a1} \times K_{a2} = \text{overall K}$

So, the product of $K_{a1} \times K_{a2}$ for a diprotic acid yields the overall $K$ for complete dissociation.

Generally, the $pH$ of polyprotic acids depends only on the removal of the first proton.

This holds true when the difference between the $K_{a1}$ and $K_{a2}$ values is at least $10^3$.

70 The $K_a$ of carbonic acid is $4.3 \times 10^{-7}$

$$\text{H}_2\text{CO}_3 \rightarrow \text{H}^+ + \text{HCO}_3^-$$

This means that $\text{H}_2\text{CO}_3$ is a ______.

- A good hydrogen-ion acceptor
- B good hydrogen-ion donor
- C poor hydrogen-ion acceptor
- D poor hydrogen-ion donor
71 A diprotic acid, $\text{H}_2\text{X}$, has the following dissociation constants:
\[ K_{a1} = 2.0 \times 10^{-4} \]
\[ K_{a2} = 3.0 \times 10^{-6} \]

What is the overall $K$ value for this acid?

- A $5.0 \times 10^{-10}$
- B $6.0 \times 10^{-10}$
- C $5.0 \times 10^{-24}$
- D $6.0 \times 10^{-24}$
- E $6.0 \times 10^{-24}$

72 A diprotic acid, $\text{H}_2\text{X}$, has the following dissociation constants:
\[ K_{a1} = 2.0 \times 10^{-4} \]
\[ K_{a2} = 3.0 \times 10^{-6} \]

What is the pH of a 0.10 M solution of $\text{H}_2\text{X}$?

- A 5.11
- B 6.11
- C 10.22
- D 2.11
- E 5.21

Absorbic acid, $\text{H}_2\text{C}_2\text{H}_6\text{O}_6$ has the following dissociation constants:
\[ K_{a1} = 8.0 \times 10^{-5} \]
\[ K_{a2} = 1.6 \times 10^{-12} \]

What is the pH of a 0.20 M solution of Absorbic acid?

- A 2.40
- B 8.30
- C 3.00
- D 4.20
- E 11.2

The Relationship Between $K_a$ and $K_b$

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$K_a$ and $K_b$

For a conjugate acid-base pair, $K_a$ and $K_b$ are related in a special way.

Write the ionization equations for the following:
1) reaction of ammonia ($\text{NH}_3$) and water

\[
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^- 
\]

2) reaction of ammonium ion ($\text{NH}_4^+$) and water

\[
\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+
\]

Write the corresponding equilibrium constant expression for each of these equations.

Which of these expressions is referred to as "$K_a$"?
Which is referred to as "$K_b$"?
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**K_a and K_b**

\[
\begin{align*}
K_a & = \frac{[NH_4^+][OH^-]}{[NH_3]} \\
K_b & = \frac{[NH_4^+][OH^-]}{[NH_3]} \\
K_a \times K_b & = [NH_4^+][OH^-][NH_3][H_2O^+] \\
\end{align*}
\]

What do these expressions have in common?

So, the product of \( K_a \times K_b \) for any conjugate acid-base pair yields the ion-product constant, \( K_w \).

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**K_a and K_b**

\[
\begin{align*}
K_a & = \frac{[NH_4^+][OH^-]}{[NH_3]} \\
K_b & = \frac{[NH_4^+][OH^-]}{[NH_3]} \\
K_a \times K_b & = [NH_4^+][OH^-][NH_3][H_2O^+] = K_w \\
\end{align*}
\]

For a specific conjugate acid-base pair, \( K_a \) and \( K_b \) are related in this way:

\[
K_a \times K_b = K_w
\]

Therefore, if you know the value of one of them, you can calculate the other.

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**K_a and K_b**

74 For the acid HCN, what is the equation on which the \( K_a \) expression is based?

\[ \begin{align*}
& A \quad \text{HCN} \leftrightarrow \text{OH}^- + \text{HCN} \\
& B \quad \text{HCN} \leftrightarrow \text{OH}^- + \text{CN}^- \\
& C \quad \text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{CN}^- + \text{H}_3\text{O}^+ \\
& D \quad \text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{CN}^- \\
\end{align*} \]

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75 The \( K_a \) for an acid, HX, is \( 2.0 \times 10^{-4} \). What is the \( K_b \) for its conjugate base, X^-?

No calculator.

\[ \begin{align*}
& A \quad 5.0 \times 10^{-11} \\
& B \quad 2.0 \times 10^{-10} \\
& C \quad 5.0 \times 10^{-10} \\
& D \quad 8.0 \times 10^{-10} \\
& E \quad 2.0 \times 10^{-4} \\
\end{align*} \]

76 The \( K_b \) for \( \text{NH}_3 \) is \( 1.8 \times 10^{-5} \). What is the \( K_a \) for the ammonium ion, \( \text{NH}_4^+ \)?

\[ \begin{align*}
& A \quad 2.5 \times 10^3 \\
& B \quad 2.5 \times 10^4 \\
& C \quad 2.5 \times 10^5 \\
& D \quad 4.3 \times 10^4 \\
& E \quad 5.6 \times 10^{-11} \\
\end{align*} \]
Which of the substances below is the strongest acid?

- A HClO
- B HCO₃⁻
- C H₂S
- D NH₃CH₃⁺
- E H₂S and HClO

Neutralization reactions are a special class of double-replacement reactions that occur between an acid and a base. Recall from last year that the general formula for a double-replacement reaction is:

\[
AB + CD \rightarrow AD + CB
\]

Double-replacement reactions are also known as ion-exchange or precipitation reactions.

The general formula for any acid-base neutralization reaction is:

\[
\text{acid} + \text{base} \rightarrow \text{salt} + \text{water}
\]

Note that the term "salt" refers to any ionic compound that does not include H⁺ or OH⁻.

Acid-base properties of salt solutions

When an acid and base react together, the resulting solution is not always neutral. The pH of the resulting mixture depends on the relative strengths of the acid and of the base.

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

Hydrolyze means to interact with water molecule and dissociate.

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

Anion effect on pH

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

Each salt contains a positive ion, a cation, and a negative ion, the anion.

Let's consider the anion first which can be denoted, X⁻. X⁻ is the conjugate base of the acid in the neutralization reaction. This anion may undergo hydrolysis as shown below:

\[
X^- + H_2O \rightarrow
\]

Predict the products.

Do you think the resulting solution will be acidic, basic or neutral? Justify your response.
**Anion Effect on pH**

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

The conjugate base (Cl\(^-\)) of a strong acid (HCl) will not hydrolyze, it will not react with water because HCl dissociates completely. In fact, none of the anions from the strong seven acids will hydrolyze. In other words, these anions will not affect pH.

List here the conjugate bases of the strong seven acids.

The conjugate bases of strong acids will not undergo hydrolysis and will not affect pH.

**Cation Effect on pH**

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

The conjugate acid (NH\(_4^+\)) of a weak base (NH\(_3\)) will hydrolyze and the resulting solution will be acidic.

Write the equilibrium expression for the reaction of NH\(_4^+\) and water.

All cations that are conjugate acids of weak bases will undergo hydrolysis and will affect pH.

**Anion Effect on pH**

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

The conjugate base (NO\(_2^-\)) of a weak acid (HNO\(_2\)) will hydrolyze, it will react with water because HNO\(_2\) does not dissociate completely and a dynamic equilibrium exists.

Write the equilibrium expression for the reaction of NO\(_2^-\) and water.

All conjugate bases of weak acids will hydrolyze and the resulting solutions will be basic.

**Cation Effect on pH**

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

The metal cation (K\(^+\)) of a strong base (KOH) will not hydrolyze. In fact, none of the cations from the strong bases will hydrolyze. In other words, these metal cations in solution will not affect pH.

List here the metal cations from Groups 1 and 2 that form strong bases.

None of the cations of strong bases will undergo hydrolysis and affect pH.

**Anion Effect on pH**

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

The conjugate acid (HSO\(_3^-\)) of a weak acid (H\(_2\)SO\(_3\)) is a special case because HSO\(_3^-\) is amphoteric. It can behave as either an acid or as a base.

How it behaves is determined by the relative magnitude of K\(_a\) for its acid and K\(_b\) for its base. If K\(_a\)> K\(_b\), it will behave as an acid if K\(_b\)> K\(_a\), then it will behave as a base.

When amphoteric anions hydrolyze they will make acidic solutions if their K\(_a\)>K\(_b\) or make basic solutions if their K\(_b\)>K\(_a\).

**Cation Effect on pH**

The pH of a salt solution depends on the ability of the salt ions to hydrolyze.

Metal cations other than Group 1 and heavy Group 2 cations can hydrolyze to affect pH. This effect is greatest for small, highly charged cations like Fe\(^{2+}\) and Al\(^{3+}\). The effect of these cations increases as ionic charge increases.

The K\(_a\) for Fe\(^{2+}\) is 3.2 x 10\(^{-19}\) and for Fe\(^{3+}\) is 6.3 x 10\(^{-3}\). These values are in the range for some weak acids we are familiar with like acetic acid K\(_a\) = 1.8 x 10\(^{-5}\).
**Acid-Base Properties of Salt Solutions**

We have observed that metal cations (hydrated in solution) lower the pH of a solution.

Lone pairs on oxygen are attracted to the cation.

The shift of electron density in water makes the O-H bond more polar.

Hydrogen atoms become more acidic, and thus more likely to become H\(^+\) ions in solution.

Shift in electron density **away from** hydrogen atoms.

---

**Cation Effect on pH**

*The pH of a salt solution depends on the ability of the salt ions to hydrolyze.*

Consider the ability of the aluminum Fe\(^{3+}\) to hydrolyze:

\[
\text{Fe}^{3+} + \text{H}_2\text{O} \rightarrow \text{Fe(OH)}_3 + \text{H}^+ 
\]

The increased H\(^+\) concentration reduces the pH.

*Small positively charged cations will hydrolyze and form acidic solution.*

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**Combined Effect of Cation and Anion**

If a salt contains an anion that does not react with water like Cl\(^-\) and a cation that does not react with water like Na\(^+\) the solution will be neutral and have a pH of 7.

Why do Cl\(^-\) and Na\(^+\) not react with water?

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**Combined Effect of Cation and Anion**

If a salt contains an anion that reacts with water to produce hydroxide ions and a cation that does not react with water then the resulting solution will be basic.

Can you give two examples of this type of salt?

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**Combined Effect of Cation and Anion**

If the salt contains an cation that reacts with water and an anion that does not, the resulting solution will acidic. This results when the cation is the conjugate acid of a weak base or a small metal ion with a 2+ or greater charge.

Can you give two examples of this type of salt?
**Combined Effect of Cation and Anion**

If a salt contains a cation that reacts with water and an anion that reacts with water then the solution's pH will be determined by comparing the cation's $K_a$ to the anion's $K_b$.

If $K_a > K_b$ then the solution will be acidic if $K_b < K_a$ then the solution will be basic.

For example, let's consider a salt solution of Fe(ClO)$_3$.

$K_a$ for Fe$^{3+}$ $6.3 \times 10^{-3}$ is the $K_b$ for ClO$^{-}$ is $3.33 \times 10^{-7}$.

This solution would be acidic because $K_a > K_b$.

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78 The pH of a solution of 0.1M solution of NaCl will be approximately ________.

- A 2
- B 5
- C 7
- D 11
- E 13

---

79 The pH of a solution of 0.1 M solution of K$_3$PO$_4$ will be approximately ________.

- A 2
- B 5
- C 7
- D 11
- E 13

---

80 The pH of a solution of 0.1M solution of LiHCO$_3$ at 25$^\circ$ C will be approximately ________.

- A 2
- B 5
- C 7
- D 11
- E 13

---

81 A 0.1M aqueous solution of ______ will have a pH of 7.0 at 25$^\circ$C.

- A NaOCl
- B BaBr$_2$
- C NH$_4$Cl
- D Ca(C$_2$H$_3$O$_2$)$_2$
- E none of these

---

82 A 0.1M aqueous solution of ______ has a pH of less than 7.0 at 25$^\circ$C.

- A Na$_2$S
- B KF
- C NaNO$_3$
- D NH$_4$Cl
- E NaF
83 An aqueous solution of ____ will produce a basic solution at 25°C.

- A NH₄ClO₄
- B KBr
- C NaCl
- D CaCO₃
- E Fe(ClO₃)₂

84 A 1-molar solution of which of the following salts has the highest pH?

- A NaNO₃
- B Na₂CO₃
- C NH₄Cl
- D NaHSO₄
- E Na₂SO₄

85 An aqueous solution of ____ will produce a neutral solution.

- A NH₄ClO₄
- B AlBr₃
- C FeCl₃
- D SrCO₃
- E BaCl₂

86 Of the following substances, an aqueous solution of ______ will form basic solutions.

- NH₄Cl, Cu(NO₃)₂, K₂CO₃, NaF

- A NH₄Cl and Cu(NO₃)₂
- B NH₄Cl and K₂CO₃
- C K₂CO₃ and NaF
- D NaF only
- E NH₄Cl only

Acid-Base Properties of Salt Solutions

Summary: Anions

The conjugate bases of strong acids will not undergo hydrolysis and not affect pH.
(Examples: Cl⁻, I⁻, NO₃⁻)

The conjugate bases of weak acids will hydrolyze and will raise pH by creating excess OH⁻ ions.
(Examples: NO₂⁻, ClO⁻, CO₃²⁻)

Amphoteric anions will either act acidic or basic based on the Kₐ and Kₐ values of the acid and base.
(Examples: HCO₃⁻, H₂PO₄⁻)

Acid-Base Properties of Salt Solutions

Summary: Cations

Cations from strong bases will not undergo hydrolysis and will not affect pH.
(Examples, Na⁺, Li⁺, Ca²⁺)

The conjugate acids of weak bases will hydrolyze and will lower pH by creating excess H⁺ ions.
(Examples: NH₄⁺, C₂H₅NH₃⁺)

Small metal cations from weak bases will hydrolyze and lower pH by creating excess H⁺ ions. The higher the charge on the cation the greater the affect on pH.
(Examples Fe²⁺, Fe³⁺, Al³⁺)
Factors Affecting Acid Strength

The more polar the H-X bond and/or the weaker the H-X bond, the more acidic the compound.

So acidity increases from left to right across a row and from top to bottom down a group.

Factors Affecting Acid Strength

In oxyacids, in which an -OH is bonded to another atom, Y, the more electronegative Y is, the more acidic the acid.

\[ K_a = 3.0 \times 10^{-8} \]
\[ K_a = 2.3 \times 10^{-11} \]

In HOCl, the electron density will be shifted to the more electronegative Cl atom weakening the O-H bond.

The strength of the acid will be in the order

\[ \text{HOCl} > \text{HOBr} > \text{HOI} \]

Factors Affecting Acid Strength

Among oxyacids, the more oxygen atoms in the molecule, the stronger the acid would be.

The electron density will be more towards the oxygen atoms and the oxidation number increases, the strength of the acid increases.

For a series of oxyacids, acidity increases with the number of oxygen atoms.

Factors Affecting Acid Strength

Resonance in the conjugate bases of carboxylic acids stabilizes the base and makes the conjugate acid more acidic.

87 Which of the following would be the strongest acid: HBr, HCl or HI?

- A HCl
- B HBr
- C HI
- D they would all be equally strong acids
88 H₂S is a stronger acid than H₂O because_____________.

- A The strength of the H-X bond is greater in H₂O
- B The strength of the H-X bond is less than in H₂O
- C The H-X bond is more polar in H₂O than in H₂S
- D The H-X bond is more polar in H₂S than in H₂O

89 Order the following acids from weakest to strongest: H₂O, HF, NH₃, and CH₄.

- A HF<H₂O<NH₃<CH₄
- B HF<NH₃<H₂O<CH₄
- C CH₄<HF<NH₃<H₂O
- D NH₃<CH₄<H₂O<HF
- E CH₄<NH₃<H₂O<HF

90 Arrange the following compounds in order of decreasing acid strength: H₂SO₄, H₂SeO₃, H₂SeO₄. Justify your response.