



## New Jersey Center for Teaching and Learning Progressive Mathematics Initiative®

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## Acids and Bases

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## Properties of 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**.



$\text{H}_3\text{O}^+$  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**.



## 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,  $\text{H}^+$ , donor.

A **base** is a proton,  $\text{H}^+$ , acceptor.



## Brønsted-Lowry Acids and Bases

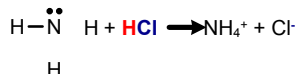
A Brønsted-Lowry acid:

- must have a removable (acidic) proton
- or
- must transfer a proton to another substance



A Brønsted-Lowry base:

- must have a pair of nonbonding electrons
- or
- must accept a proton



## Brønsted-Lowry Acids and Bases



HCl donates the proton and acts as a Brønsted-Lowry acid.

$\text{H}_2\text{O}$  accepts the proton and acts as a Brønsted-Lowry base.

## 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.

### Arrhenius acids - only substances dissolved in $\text{H}_2\text{O}$

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

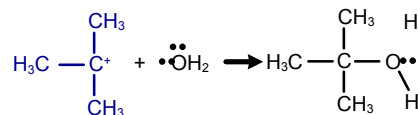
### Brønsted-Lowry acids - only substances that gain or lose $\text{H}^+$

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

## Lewis Acids

Lewis acids are defined as electron-pair acceptors.

Atoms with an empty valence orbital can be Lewis acids.



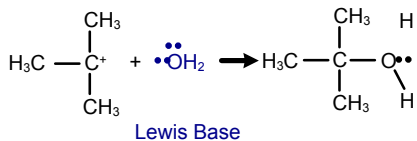
Lewis Acid

## 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.



1 A Brønsted-Lowry base is defined as a substance that \_\_\_\_\_.

- A increases  $[H^+]$  when placed in  $H_2O$
- B decreases  $[H^+]$  when placed in  $H_2O$
- C increases  $[OH^-]$  when placed in  $H_2O$
- D acts as a proton acceptor
- E acts as a proton donor

1 A Brønsted-Lowry base is defined as a substance that \_\_\_\_\_.

- A increases  $[H^+]$  when
- B decreases  $[H^+]$  when
- C increases  $[OH^-]$  when
- D acts as a proton acceptor
- E acts as a proton donor

D  
The Brønsted-Lowry definition  
of a base is a substance that  
is a proton acceptor.

2 A Brønsted-Lowry acid is defined as a substance that \_\_\_\_\_.

- A increases  $K_a$  when placed in  $H_2O$
- B decreases  $[H^+]$  when placed in  $H_2O$
- C increases  $[OH^-]$  when placed in  $H_2O$
- 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
- B decreases  $[H^+]$  when
- C increases  $[OH^-]$  when
- D acts as a proton acceptor
- E acts as a proton donor

E  
The Brønsted-Lowry definition  
of an acid is a substance that  
is a proton donor.

3 Which of the following compounds could never act as an acid?

- A  $SO_4^{2-}$
- B  $HSO_4^-$
- C  $H_2SO_4$
- D  $NH_3$
- E  $CH_3COOH$

- 3 Which of the following compounds could never act as an acid?

- A  $\text{SO}_4^{2-}$   
 B  $\text{HSO}_4^-$   
 C  $\text{H}_2\text{SO}_4$   
 D  $\text{NH}_3$   
 E  $\text{CH}_3\text{COOH}$

Answer

A  
In order to be an acid you must be either be a substance that increases the  $\text{H}^+$  concentration or a substance that donates a proton or an electron pair acceptor.  $\text{SO}_4^{2-}$  fits none of these definitions.

- 4 According to the following reaction model, reactant is acting like an acid?



- A  $\text{H}_2\text{SO}_4$   
 B  $\text{H}_2\text{O}$   
 C  $\text{H}_3\text{O}^+$   
 D  $\text{HSO}_4^-$   
 E None of the above

- 4 According to the following reaction model, reactant is acting like an acid?

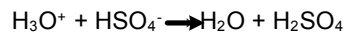


- A  $\text{H}_2\text{SO}_4$   
 B  $\text{H}_2\text{O}$   
 C  $\text{H}_3\text{O}^+$   
 D  $\text{HSO}_4^-$   
 E None of the above

Answer

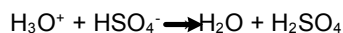
A  
 $\text{H}_2\text{SO}_4$  is donating a proton so it is the reactant that is acting like an acid.

- 5 According to the following reaction, which reactant is acting like a base?



- A  $\text{H}_2\text{SO}_4$   
 B  $\text{H}_2\text{O}$   
 C  $\text{H}_3\text{O}^+$   
 D  $\text{HSO}_4^-$   
 E None of the above

- 5 According to the following reaction, which reactant is acting like a base?

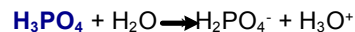


- A  $\text{H}_2\text{SO}_4$   
 B  $\text{H}_2\text{O}$   
 C  $\text{H}_3\text{O}^+$   
 D  $\text{HSO}_4^-$   
 E None of the above

Answer

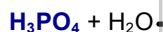
D  
 $\text{HSO}_4^-$  is accepting a proton so it is the reactant that is acting like a base.

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



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

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

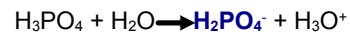


- A Acid  
 B Base  
 C Neither  
 D Both  
 E None of the above

Answer

A  
 $H_3PO_4$  is donating a proton so it is acting like an acid.

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



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

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



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

Answer

B  
 $H_2PO_4^-$  is accepting a proton so it is acting like a base.

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

- A  $Cl^-$   
 B  $NH_3$   
 C  $BF_3$   
 D  $CN^-$   
 E  $H_2O$

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

- A  $Cl^-$   
 B  $NH_3$   
 C  $BF_3$   
 D  $CN^-$   
 E  $H_2O$

Answer

C  
 A Lewis Base is an electron pair donor and any substance that is a Bronsted Lowry base would be a Lewis base.  $BF_3$  is not a Bronsted Lowry base and can not act as an electron pair donor.

9 In the reaction  $BF_3 + F^- \rightleftharpoons BF_4^-$

$BF_3$  acts as a/an \_\_\_\_\_ acid.

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

9 In the reaction  $\text{BF}_3 + \text{F}^- \rightarrow \text{BF}_4^-$

$\text{BF}_3$  acts as a/an \_\_\_\_\_

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

Answer

C  
 $\text{BF}_3$  is accepting an electron pair from  $\text{F}^-$  and is acting as a Lewis Acid.

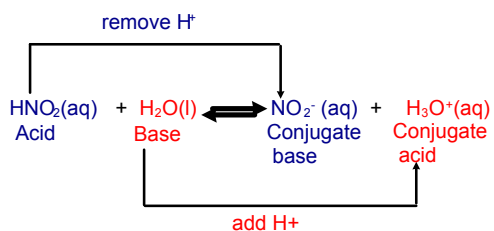
## Conjugate Acid Base Pairs

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## 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.



## Acids in Water

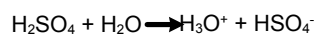
What Happens When an Acid Dissolves in Water?

Which is the acid?

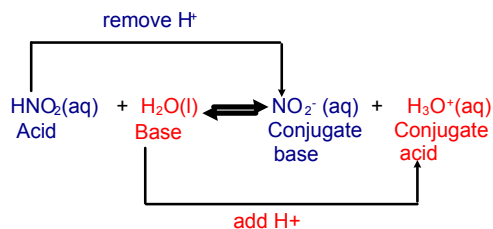
Which is the base?

Water acts as a Brønsted-Lowry base and takes a proton ( $\text{H}^+$ ) from the acid.

As a result, the conjugate base of the acid and a hydronium ion are formed.



## 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.

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



- A a catalyst  
 B both an acid and a base  
 C the conjugate acid of  $\text{NH}_3$   
 D the reducing agent  
 E the oxidizing agent

Question from the College Board

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



- A a catalyst
- B both an acid and
- C the conjugate acid
- D the reducing agent
- E the oxidizing agent

**Answer**

C

NH<sub>4</sub><sup>+</sup> acts as the conjugate acid of NH<sub>3</sub>.

Question from the College Board

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

- A  $\text{ClO}_4^+$ ,  $\text{HS}^-$ ,  $\text{PH}_3^-$ ,  $\text{CO}_3^-$
- B  $\text{ClO}_4^-$ ,  $\text{HS}^-$ ,  $\text{PH}_3$ ,  $\text{CO}_3^{2-}$
- C  $\text{ClO}_4^{2-}$ ,  $\text{HS}^{2-}$ ,  $\text{PH}_3^{3-}$ ,  $\text{CO}_3^{2-}$
- D  $\text{ClO}_4^-$ ,  $\text{H}_2\text{S}^-$ ,  $\text{PH}_3$ ,  $\text{CO}_3^{2-}$

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

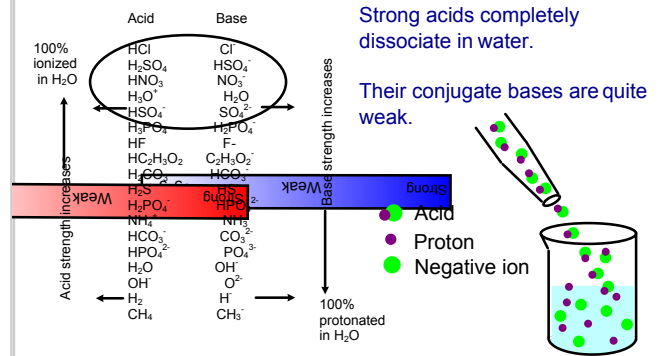
- A  $\text{ClO}_4^+$ ,  $\text{HS}^-$ ,  $\text{PH}_3^-$
- B  $\text{ClO}_4^-$ ,  $\text{HS}^-$ ,  $\text{PH}_3$
- C  $\text{ClO}_4^{2-}$ ,  $\text{HS}^{2-}$ ,
- D  $\text{ClO}_4^-$ ,  $\text{H}_2\text{S}^-$ ,  $\text{PH}_3$ ,

**Answer**

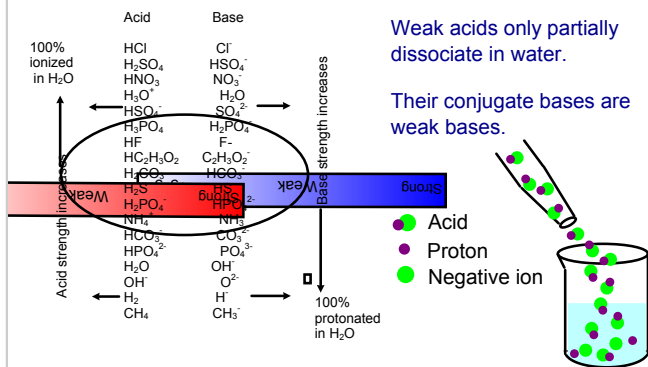
B

To create a conjugate base for an acid you must remove an H and decrease the charge by 1.  
 $\text{HClO}_4$ 's conjugate base is  $\text{ClO}_4^-$ ,  
 $\text{H}_2\text{S}$ 's is  $\text{HS}^-$ , and  $\text{PH}_4^+$  is  $\text{PH}_3$  and  $\text{HCO}_3^-$  is  $\text{CO}_3^{2-}$ .

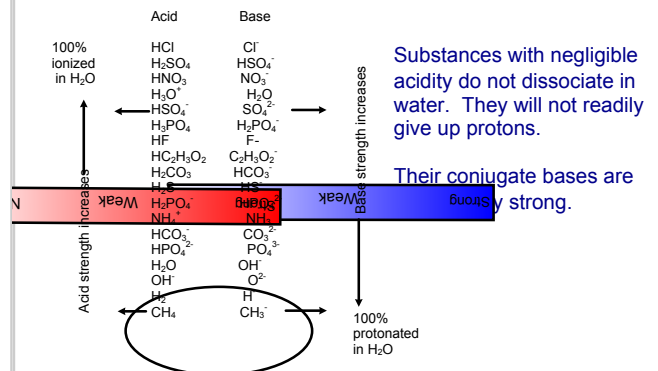
### Acid and Base Strength



### Acid and Base Strength



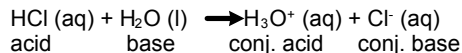
### Acid and Base Strength



### Acid and Base Strength

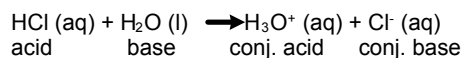
In any acid-base reaction, equilibrium will favor the reaction in which the proton moves toward the stronger base.

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



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

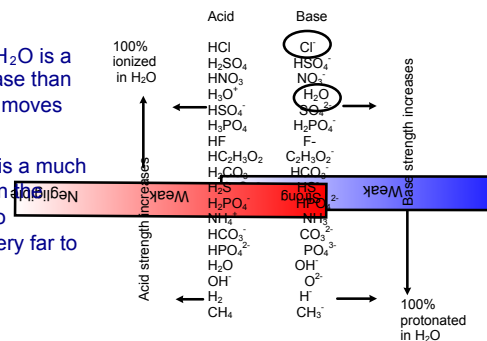
### Acid and Base Strength



In this example, H<sub>2</sub>O is a much stronger base than Cl<sup>-</sup> so the proton moves from HCl to H<sub>2</sub>O

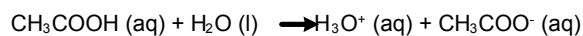
Conversely, HCl is a much stronger acid than the hydronium ion, so equilibrium lies very far to the right

$$K \gg 1$$



### Acid and Base Strength

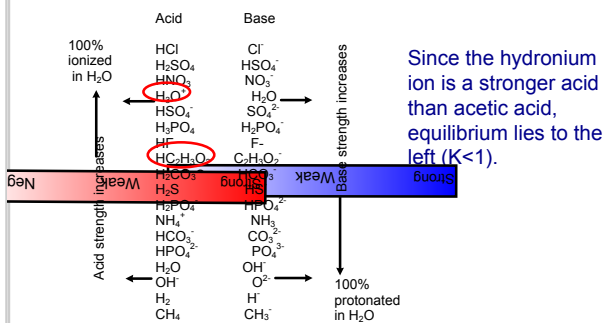
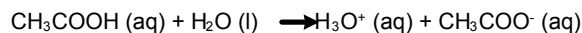
Consider this equilibrium between acetic acid and acetate ion:



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

If you look for the stronger acid:	If you look for the stronger base:
Equilib lies away from the stronger acid.	Equilib favors this base accepting a proton.

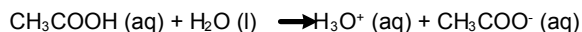
### Acid and Base Strength



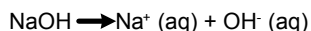
### Acid and Base Strength

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.



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



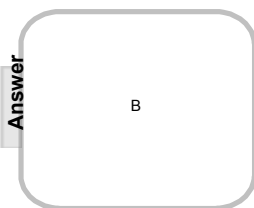
12 Strong acids have \_\_\_\_\_ conjugate bases.

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



12 Strong acids have \_\_\_\_\_ conjugate bases.

- A strong
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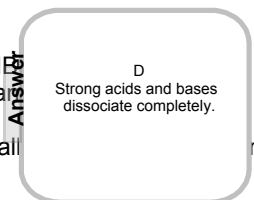


13 HBr, hydrobromic acid is a strong acid. This means that it \_\_\_\_\_.

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

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

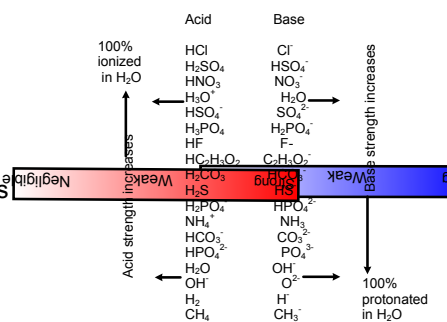
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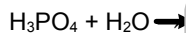
14 For the following reaction, determine which side of the equilibrium is favored.



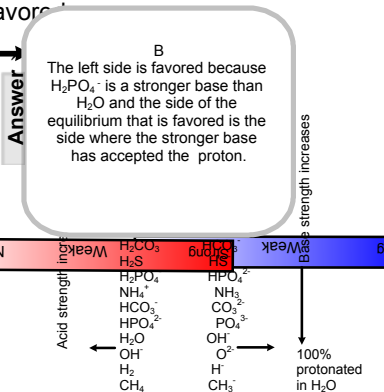
- A the right side
- B the left side
- C Neither side is favored



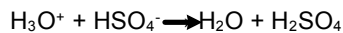
14 For the following reaction, determine which side of the equilibrium is favored.



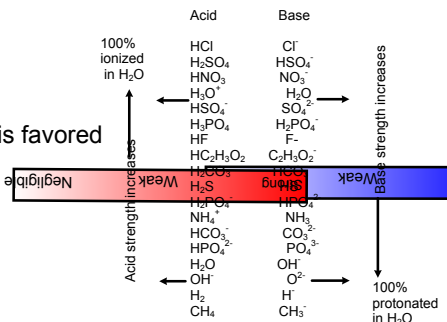
- A the right side
- B the left side
- C Neither side is favored



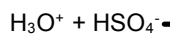
15 For the following reaction, determine which side of the equilibrium is favored.



- 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.

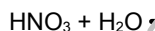


- A the right side
- B the left side
- C Neither side is favored

**Answer**

**A**  
The right side is favored because H<sub>2</sub>O is a stronger base than HSO<sub>4</sub><sup>-</sup> and the side of the equilibrium that is favored is the side where the stronger base has accepted the proton.

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



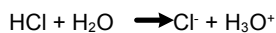
- A the right side
- B the left side
- C Neither side is favored

**Answer**

**A**  
The right side is favored because H<sub>2</sub>O is a stronger base than NO<sub>3</sub><sup>-</sup> and the side of the equilibrium that is favored is the side where the stronger base has accepted the proton.

### Amphoteric Substances

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.

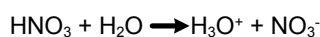


Above, water accepts a proton, thus acting as a base.



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

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



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

### Amphoteric Substances

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### Amphoteric Substances

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

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

Answer

C

18 Write the equations and equilibrium expressions for HS<sup>-</sup> when it is acting like a Brønsted-Lowry acid and when it is acting like a Brønsted-Lowry base.

## Strong Acids and Bases

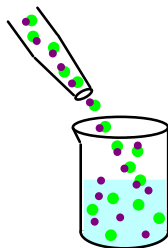
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### Strong Acids

Recall, strong acids completely ionize in solution.

The seven strong acids are:

- HCl hydrochloric acid
- HBr hydrobromic acid
- HI hydroiodic acid
- HNO<sub>3</sub> nitric acid
- H<sub>2</sub>SO<sub>4</sub> sulfuric acid
- HClO<sub>3</sub> chloric acid
- HClO<sub>4</sub> 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)<sub>2</sub>, Sr(OH)<sub>2</sub>, and Ba(OH)<sub>2</sub>.

Again, these substances dissociate completely in aqueous solution. In other words, NaOH exists entirely as Na<sup>+</sup> ions and OH<sup>-</sup> 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.

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

Answer

C  
Calcium hydroxide's formula is Ca(OH)<sub>2</sub>. For every 1 mole of Ca(OH)<sub>2</sub> there are 2 moles of OH<sup>-</sup>. If you have a 0.22M solution of Ca(OH)<sub>2</sub> then the concentration of OH<sup>-</sup> is 2x as much or 0.44M.

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

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

Answer

To determine the new concentration you can use the formula  $M_1V_1=M_2V_2$   
 $M_1 = .05M$ ,  $V_1 = 25 \text{ ml}$  and  $V_2 = 100 \text{ ml}$   
 $M_2 = M_1V_1/V_2$   
 $M_2 = .05M \times 25 \text{ ml} / 100 \text{ ml}$   
 $M_2 = 0.0125M \text{ HCl}$   
 The concentration of H<sup>+</sup> equals the concentration of the HCl solution and therefore equals 0.0125M.

21 What is the [H<sup>+</sup>] ion concentration of a 50 ml solution of 0.025M H<sub>2</sub>SO<sub>4</sub>, when diluted with 150 ml of water?

21 What is the  $[H^+]$  ion concentration of a 50 ml solution of 0.025M  $H_2SO_4$  of water?

Answer

Using  $M_1V_1 = M_2V_2$ . You are given:  
 $M_1 = 0.025M$ ,  $V_1 = 50$  ml and  $V_2 = 100$  ml.  
 $M_2 = M_1V_1/V_2$   
 $M_2 = 0.025M \times 50 \text{ ml} / 100 \text{ ml}$   
 $M_2 = 0.0125M$   
 For every 1 mole of  $H_2SO_4$  there are 2 moles of  $H^+$ .  
 If you have a 0.0125M solution of  $H_2SO_4$ , then the concentration of  $H^+$  is 2x as much or 0.0250M

## Auto-ionization of Water

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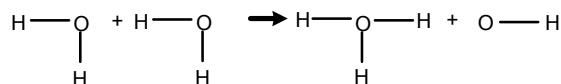
22 A solution of 25 ml of 0.1M HCl and 50 ml of 0.5M  $HNO_3$  are mixed together. What is the  $[H^+]$  ion concentration of the resulting solution?

Answer

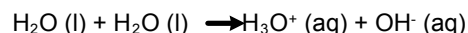
## Auto-ionization of Water

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

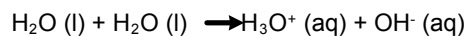
In pure water, a few molecules act as bases and a few act as acids, in a process referred to as autoionization.



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



## Autoionization of Water



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,  $[ ]$ .



$$K = \frac{[H_3O^+] \times [OH^-]}{[H_2O] \times [H_2O]}$$

## Ion-Product Constant

In most dilute acid and base solutions, the concentration of undissociated water, remains more or less a constant. We can thus disregard the denominator in the equilibrium expression.

$$K = \frac{[H_3O^+] \times [OH^-]}{[H_2O] \times [H_2O]} \text{ becomes } K_w = [H_3O^+] \times [OH^-]$$

**Ion-Product Constant**

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-]$$

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.

23 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

23 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

Answer

A

$K_w$  is a very small number that means the amount of reactants is much greater than the amount of products. Water ionizes to very small extent.

24 The ion-product constant for water,  $K_w$  is represented by

- A  $[\text{H}_2\text{O}]^2$
- B  $[\text{H}_3\text{O}^+] \times [\text{OH}^-]$
- C  $[\text{H}_3\text{O}^+] + [\text{OH}^-]$
- D  $[\text{H}_3\text{O}^+] - [\text{OH}^-]$

24 The ion-product constant for water,  $K_w$  is represented by

- A  $[\text{H}_2\text{O}]^2$
- B  $[\text{H}_3\text{O}^+] \times [\text{OH}^-]$
- C  $[\text{H}_3\text{O}^+] + [\text{OH}^-]$
- D  $[\text{H}_3\text{O}^+] - [\text{OH}^-]$

Answer

B

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-]$$

25 What is the  $[\text{H}^+]$  ion concentration of a solution with an  $[\text{OH}^-]$  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  $1 \times 10^{-7}$

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  $1 \times 10^{-7}$

Answer

C

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

$$[H^+] = K_w / [OH^-]$$

$$[H^+] = 1 \times 10^{-14} / 1 \times 10^{-9}$$

$$[H^+] = 1 \times 10^{-5}$$

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  $1.23 \times 10^{-9}$
- E  $1 \times 10^{-10}$

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  $1.23 \times 10^{-9}$
- E  $1 \times 10^{-10}$

Answer

E

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

$$[H^+] = K_w / [OH^-]$$

$$[H^+] = 1 \times 10^{-14} / 1 \times 10^{-4}$$

$$[H^+] = 1 \times 10^{-10}$$

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  $8.13 \times 10^{-6}$

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  $8.13 \times 10^{-6}$

Answer

E

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

$$[H^+] = K_w / [OH^-]$$

$$[H^+] = 1 \times 10^{-14} / 1.23 \times 10^{-9}$$

$$[H^+] = 8.13 \times 10^{-6}$$

pH

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**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.

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

Hydrogen ion concentration, $[H^+]$ in moles/Liter	pH
$1.0 \times 10^{-1}$	1
$1.0 \times 10^{-2}$	2
$1.0 \times 10^{-10}$	10

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  $10^{-14}$

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  $10^{-14}$

Answer

D  
If the solution is neutral the concentration of  $[H_3O^+] = [OH^-]$  and each has a concentration of  $1 \times 10^{-7}$ .



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

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

Answer

C  
 The solution with a pH of 3 has a hydrogen ion concentration is 100 x more than a solution with a pH of 5, because pH is a log function and the difference is  $1 \times 10^3$  vs  $1 \times 10^5$  which is 100 x more.

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

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

Answer

D  
 The solution with a pH of 14 has a hydrogen ion concentration is 1000 x less because pH is a log function and the difference is  $1 \times 10^{14}$  vs  $1 \times 10^{11}$  which is 1000 x less.

## pH

Therefore, in pure water

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = 7.00$$

$$\text{pH} = -\log (1.0 \times 10^{-7}) = 7.00$$

An acid has a higher  $[\text{H}_3\text{O}^+]$  than pure water, so its pH is <7.

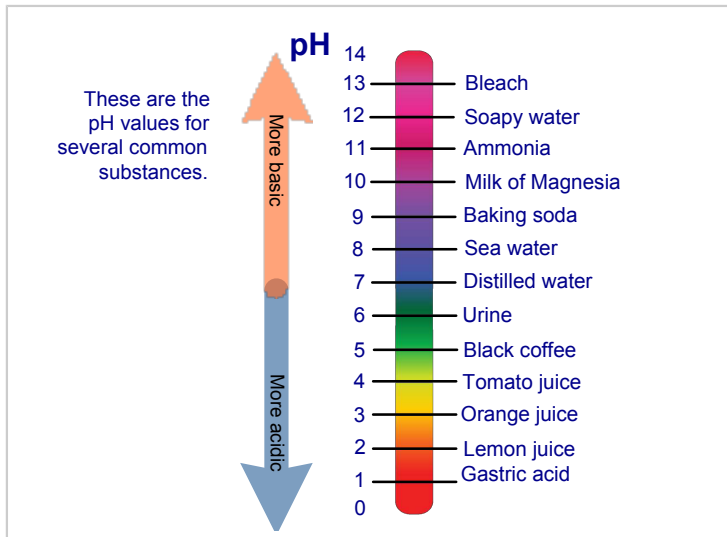
A base has a lower  $[\text{H}_3\text{O}^+]$  than pure water, so its pH is >7.

Solution type	$[\text{H}^+](\text{M})$	$[\text{OH}^-](\text{M})$	pH value
Acidic	$>1.0 \times 10^{-7}$	$<1.0 \times 10^{-7}$	<7.00
Neutral	$=1.0 \times 10^{-7}$	$=1.0 \times 10^{-7}$	=7.00
Basic	$<1.0 \times 10^{-7}$	$>1.0 \times 10^{-7}$	>7.00

## pH

ACID	BASE
$[\text{H}^+] > [\text{OH}^-]$ There are excess hydrogen ions in solution.	$[\text{H}^+] < [\text{OH}^-]$ There are excess hydroxide ions in solution.

Solution type	$[\text{H}^+](\text{M})$	$[\text{OH}^-](\text{M})$	pH value
Acidic	$>1.0 \times 10^{-7}$	$<1.0 \times 10^{-7}$	<7.00
Neutral	$=1.0 \times 10^{-7}$	$=1.0 \times 10^{-7}$	=7.00
Basic	$<1.0 \times 10^{-7}$	$>1.0 \times 10^{-7}$	>7.00



31 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.

31 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.

Answer

B  
In a basic solution the concentration of hydrogen ion is less than the concentration of hydroxide ion.

32 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.

32 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.

Answer

B  
In an acidic solution, the concentration of hydroxide ion is less than the concentration of hydrogen ion.

33 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

33 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

Answer

D

The smaller the pH the more acidic the solution. The larger the pH the more basic the solution. The solution with the highest concentration of hydroxide ion, the most basic solution, has the largest pH in this case 12.6.

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

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

Answer

A

The smaller the pH the more acidic the solution. The larger the pH the more basic the solution. The solution with the highest concentration of hydrogen ion, the most acidic solution, has the smallest pH in this case 1.98.

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

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

Answer

D

A weak base would have a relatively high pH but not a very high pH. A strong base would have a much higher pH

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

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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

Answer

A  
A reasonable pH for a strong acid is 2.

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

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

Slide 82 (Answer) / 208

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

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

Answer

C  
 $\text{pH} = -\log [\text{H}^+]$   
 $\text{pH} = -\log (0.010)$   
 $\text{pH} = 2$

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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  $1 \times 10^{-9}$

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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  $1 \times 10^{-9}$

Answer

D  
 The hydrogen ion concentration of a solution with a pH of 5.0 is  $1 \times 10^{-5}$ .  
 $[\text{OH}^-] = K_w / [\text{H}^+]$   
 $[\text{OH}^-] = 1 \times 10^{-14} / 1 \times 10^{-5} = 1 \times 10^{-9}$ .

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### How Do We Measure pH?

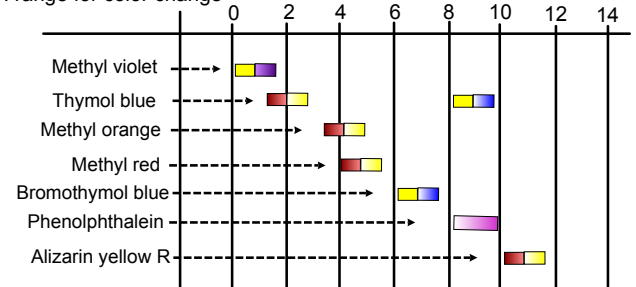
For less accurate measurements, one can use Litmus paper

"Red" litmus paper turns blue above  $\sim \text{pH} = 8$

"Blue" litmus paper turns red below  $\sim \text{pH} = 5$

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

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).

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

## 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: 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*

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

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

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

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

C  
 $\text{pH} = -\log \text{H}^+$   
 $\text{pH} = -\log (1 \times 10^{-5})$   
 $\text{pH} = 5$

40 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

40 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

C

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

$$[H^+] = 1 \times 10^{-14} / [OH^-]$$

$$[H^+] = 1 \times 10^{-14} / 1 \times 10^{-12}$$

$$[H^+] = 1 \times 10^{-2}$$

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

$$pH = -\log (1 \times 10^{-2})$$

$$pH = 2$$

41 What is the pH of an aqueous solution at  $25.0^\circ\text{C}$  in which  $[H^+]$  is  $0.0025$  M?

- A 3.4  
 B 2.6  
 C -2.6  
 D -3.4  
 E 2.25

41 What is the pH of an aqueous solution at  $25.0^\circ\text{C}$  in which  $[H^+]$  is  $0.0025$  M?

- A 3.4  
 B 2.6  
 C -2.6  
 D -3.4  
 E 2.25

B

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

$$pH = -\log (0.0025)$$

$$pH = 2.60$$

### 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  $[H^+] = 10^{-pH}$

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  $7.14 \times 10^{-3}$  M?

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

Answer

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

$$pH = -\log (7.14 \times 10^{-3} \text{M})$$

$$pH = 2.15$$

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

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

Answer

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ \text{pH} &= -\log (1.92 \times 10^{-9} \text{M}) \\ \text{pH} &= 8.72 \end{aligned}$$

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

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

Answer

$$\begin{aligned} [\text{H}^+] &= 10^{-\text{pH}} \\ [\text{H}^+] &= 10^{-4.29} \\ [\text{H}^+] &= 5.13 \times 10^{-5} \text{ M} \end{aligned}$$

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

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

Answer

$$\begin{aligned} [\text{H}^+] &= 10^{-\text{pH}} \\ [\text{H}^+] &= 10^{-4.29} \\ [\text{H}^+] &= 5.13 \times 10^{-5} \text{ M} \\ K_w &= [\text{H}^+][\text{OH}^-] \\ [\text{OH}^-] &= K_w / [\text{H}^+] \\ [\text{OH}^-] &= 1 \times 10^{-14} / 5.13 \times 10^{-5} \text{ M} \\ [\text{OH}^-] &= 1.95 \times 10^{-10} \end{aligned}$$

## 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

$$\text{pOH} = -\log [\text{OH}^-]$$

$$\text{pK}_w = -\log K_w$$

$$\text{pK}_a = -\log K_a$$

$$\text{pK}_b = -\log K_b$$

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

- 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

## Relationship between pH and pOH

Because

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1.0 \times 10^{-14},$$

we know that

$$-\log [\text{H}_3\text{O}^+] + -\log [\text{OH}^-] = -\log K_w = 14.00$$

or, in other words,

$$\text{pH} + \text{pOH} = \text{pK}_w = 14.00$$

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

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

Answer

$$\begin{aligned} & \text{C} \\ \text{pH} + \text{pOH} &= 14 \\ \text{pH} &= 14 - \text{pOH} \\ \text{pH} &= 14 - 6.12 \\ \text{pH} &= 7.88 \end{aligned}$$

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

Answer

$$\begin{aligned} & \text{A} \\ \text{pH} &= -\log[\text{H}^+] \\ \text{pH} &= -\log(2.5 \times 10^{-4} \text{ M}) \\ \text{pH} &= 3.60 \\ \text{pH} + \text{pOH} &= 14 \\ \text{pOH} &= 14 - 3.60 \\ \text{pOH} &= 10.4 \end{aligned}$$



48 An aqueous solution of an acid has pH of 4.11 what is the [OH<sup>-</sup>] 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  $4.03 \times 10^7$

48 An aqueous solution of an acid has pH of 4.11 what is the [OH<sup>-</sup>] 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  $4.03 \times 10^7$

Answer

D

$$\begin{aligned} \text{pH} + \text{pOH} &= 14 \\ \text{pOH} &= 14 - \text{pH} \\ \text{pOH} &= 14 - 4.11 \\ \text{pOH} &= 9.89 \\ [\text{OH}^-] &= 10^{-\text{pOH}} \\ [\text{OH}^-] &= 10^{-9.89} \\ [\text{OH}^-] &= 1.29 \times 10^{-10} \end{aligned}$$

49 An aqueous solution of an base has pOH of 3.33 what is the [H<sup>+</sup>] 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  $4.88 \times 10^7$

49 An aqueous solution of an base has pOH of 3.33 what is the [H<sup>+</sup>] 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  $4.88 \times 10^7$

Answer

B

$$\begin{aligned} \text{pH} + \text{pOH} &= 14 \\ \text{pH} &= 14 - \text{pOH} \\ \text{pH} &= 14 - 3.33 \\ \text{pH} &= 10.67 \\ [\text{H}^+] &= 10^{-\text{pH}} \\ [\text{H}^+] &= 10^{-10.67} \\ [\text{H}^+] &= 2.14 \times 10^{-11} \end{aligned}$$

## Weak Acids and Bases

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### Acid Dissociation Constants, $K_a$



For a generalized acid dissociation, the equilibrium expression is

$$K_c, \text{ AKA } K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

**Acid Dissociation Constants,  $K_a$** 

The greater the value of  $K_a$ , the stronger is the acid.

Acid	Proton Transfer	$K_a$ Value
Hydrochloric acid	$\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$	Large
Sulfuric acid	$\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$	Large
Nitric acid	$\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$	About 20
Hydrofluoric acid	$\text{HF} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{F}^-$	$1.2 \times 10^{-4}$
Carbonic acid	$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HCO}_3^-$	$4.3 \times 10^{-7}$
Hydrogen cyanide	$\text{HCN} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{CN}^-$	$4.9 \times 10^{-16}$

50 The acid dissociation constant ( $K_a$ ) of HF is  $6.7 \times 10^{-4}$ . Which of the following is true of a 0.1M solution of HF?

- A [HF] is greater than  $[\text{H}^+][\text{F}^-]$
- B [HF] is less than  $[\text{H}^+][\text{F}^-]$
- C [HF] is equal to  $[\text{H}^+][\text{F}^-]$
- D [HF] is equal to  $[\text{H}^+][\text{F}^-]$

50 The acid dissociation constant ( $K_a$ ) of HF is  $6.7 \times 10^{-4}$ . Which of the following is true of a 0.1M solution of HF?

- A [HF] is greater than  $[\text{H}^+][\text{F}^-]$
- B [HF] is less than  $[\text{H}^+][\text{F}^-]$
- C [HF] is equal to  $[\text{H}^+][\text{F}^-]$
- D [HF] is equal to  $[\text{H}^+][\text{F}^-]$

Answer

A  
Because the  $K_a$  value is so small it means that the amount of reactants is much greater than the amount of products therefore [HF] is greater than  $[\text{H}^+][\text{F}^-]$

**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



or, without water

**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.



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

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  $\text{H}_3\text{O}^+$ ?

**Calculating  $K_a$  from the pH**

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

$$2.38 = -\log [\text{H}_3\text{O}^+]$$

$$-2.38 = \log [\text{H}_3\text{O}^+]$$

$$10^{-2.38} = 10^{\log [\text{H}_3\text{O}^+]} = [\text{H}_3\text{O}^+]$$

$$4.2 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+] = [\text{HCOO}^-]$$

Note that this is a monoprotic acid, so  
[acid] = [conjugate base]

**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.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

	[HCOOH], M	[HCOO <sup>-</sup> ], M	[H <sub>3</sub> O <sup>+</sup> ], M
Initially	0.10	0	0
Change	-4.2 x 10 <sup>-3</sup>	+4.2 x 10 <sup>-3</sup>	+4.2 x 10 <sup>-3</sup>
At Equilibrium	0.10 - 4.2 x 10 <sup>-3</sup> ≈ 0.10	4.2 x 10 <sup>-3</sup>	4.2 x 10 <sup>-3</sup>

**Calculating  $K_a$  from the pH**

Now, we substitute values into the  $K_a$  expression and solve:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$K_a = \frac{[4.2 \times 10^{-3}][4.2 \times 10^{-3}]}{[0.10]}$$

$$K_a = 1.8 \times 10^{-4}$$

**Calculating Percent Ionization**

$$\text{Percent Ionization} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HA}]_{\text{initial}}} \times 100\%$$

In this example

$$\begin{aligned} [\text{H}_3\text{O}^+]_{\text{eq}} &= 4.2 \times 10^{-3} \text{ M} \\ [\text{HCOOH}]_{\text{initial}} &= 0.10 \text{ M} \end{aligned}$$

$$\begin{aligned} \text{Percent Ionization} &= \frac{4.2 \times 10^{-3}}{0.10} \times 100\% \\ &= 4.2\% \end{aligned}$$

**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.

*Important note:*

In the case of weak acids and bases, we assume little ionization and therefore we ignore the amount of ionization and write 0.10 and not 0.10 - x.

If you find that the % ionization is greater than 5% you can't ignore the x term and you must use the quadratic equation to solve.

**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<sup>+</sup>] ions that are produced, compared to the original acid concentration.

$$\text{Percent Ionization} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HA}]_{\text{initial}}} \times 100\%$$

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

2.42 x 10<sup>-4</sup>

2.65 x 10<sup>-9</sup>

3.71 x 10<sup>-10</sup>

2.13 x 10<sup>-5</sup>

6.78 x 10<sup>-12</sup>

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

- $2.42 \times 10^{-4}$   
  $2.65 \times 10^{-9}$   
  $3.71 \times 10^{-10}$   
  $2.13 \times 10^{-5}$   
  $6.78 \times 10^{-12}$

Answer

$$\begin{aligned} & \text{B} \\ & K_a = x^2/0.125 \\ & x = [\text{H}^+] = 10^{-4.74} \\ & x = 1.82 \times 10^{-5} \\ & K_a = (1.82 \times 10^{-5})^2 / 0.125 \\ & K_a = 2.65 \times 10^{-9} \end{aligned}$$

52 What is % ionization of a 0.125M solution of hypobromous acid (HBrO) that has a pH of 4.74.

- 1.45%  
 68.89%  
 .0145%  
 .00145%  
 .6889%

52 What is % ionization of a 0.125M solution of hypobromous acid (HBrO) that has a pH of 4.74.

- 1.45%  
 68.89%  
 .0145%  
 .00145%  
 .6889%

Answer

$$\begin{aligned} & \text{C} \\ & [\text{H}^+] = 10^{-4.74} \\ & = 1.82 \times 10^{-5} \\ & \% \text{ ionization} = [\text{H}^+]/[\text{HBrO}] \times 100 \% \\ & = 1.82 \times 10^{-5} / 0.125 \times 100 \% \\ & = .0145\% \end{aligned}$$

53 What is the  $K_a$  of a 0.20M solution of nitrous acid ( $\text{HNO}_2$ ) that has a pH of 2.02?

- $1.3 \times 10^{-5}$   
  $9.9 \times 10^{-2}$   
  $1.2 \times 10^{-8}$   
  $4.2 \times 10^{-7}$   
  $4.6 \times 10^{-4}$

53 What is the  $K_a$  of a 0.20M solution of nitrous acid ( $\text{HNO}_2$ ) that has a pH of 2.02?

- $1.3 \times 10^{-5}$   
  $9.9 \times 10^{-2}$   
  $1.2 \times 10^{-8}$   
  $4.2 \times 10^{-7}$   
  $4.6 \times 10^{-4}$

Answer

$$\begin{aligned} & \text{E} \\ & K_a = x^2/0.20 \\ & x = [\text{H}^+] = 10^{-2.02} \\ & x = 9.55 \times 10^{-3} \\ & K_a = (9.55 \times 10^{-3})^2 / 0.20 \\ & K_a = 4.6 \times 10^{-4} \end{aligned}$$

54 What is % ionization of a 0.20M solution of nitrous acid ( $\text{HNO}_2$ ) that has a pH of 2.02?

- .0477  
 4.77%  
 5.99%  
 .0599%  
 .6889%

54 What is % ionization of a 0.20M solution of nitrous acid ( $\text{HNO}_2$ ) that has a pH of 2.02?

- .0477  
 4.77%  
 5.99%  
 .0599%  
 .6889%

Answer

$$\begin{aligned} \text{B} \\ [\text{H}^+] &= 10^{-2.02} \\ &= 9.55 \times 10^{-3} \\ \% \text{ ionization} &= \frac{[\text{H}^+]}{[\text{HNO}_2]} \times 100 \% \\ &= \frac{9.55 \times 10^{-3}}{0.20} \times 100 \% \\ &= 4.77\% \end{aligned}$$

55 What is the  $K_a$  of a 0.115M solution of a weak acid that has a 2.55% ionization?

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

55 What is the  $K_a$  of a 0.115M solution of a weak acid that has a 2.55% ionization?

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

Answer

$$\begin{aligned} \text{C} \\ \% \text{ ionization} &= \frac{[\text{H}^+]}{[\text{HA}]} \times 100 \% \\ [\text{H}^+] &= \% \text{ ionization} / 100 \times [\text{HA}] \\ [\text{H}^+] &= 2.55 / 100 \times 0.115 \\ [\text{H}^+] &= 2.93 \times 10^{-3} \\ K_a &= [\text{H}^+]^2 / [\text{HA}] \\ K_a &= (2.93 \times 10^{-3})^2 / 0.115 \\ K_a &= 7.48 \times 10^{-5} \end{aligned}$$

56 If the acid dissociation constant,  $K_a$  for an acid HA is  $8 \times 10^{-4}$  at  $25^\circ\text{C}$ , 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

56 If the acid dissociation constant,  $K_a$  for an acid HA is  $8 \times 10^{-4}$  at  $25^\circ\text{C}$ , 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%

Answer

$$\begin{aligned} \text{E} \\ K_a &= \frac{[\text{H}^+]^2}{[\text{HA}]} \\ [\text{H}^+] &= \sqrt{K_a \times [\text{HA}]} \\ [\text{H}^+] &= \sqrt{8 \times 10^{-4} \times 0.50} \\ [\text{H}^+] &= 0.02 \\ \% \text{ ionization} &= \frac{[\text{H}^+]}{[\text{HA}]} \times 100\% \\ \% \text{ ionization} &= 0.02 / 0.50 \times 100\% \\ \% \text{ ionization} &= 4\% \end{aligned}$$

Question from the College Board

### 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  $25^\circ\text{C}$ .  $K_a$  for acetic acid at  $25^\circ\text{C}$  is  $1.8 \times 10^{-5}$ .

First, we write the dissociation equation for acetic acid



**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 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  $1.8 \times 10^{-5}$ .

We next set up an ICE chart...

	$[\text{HC}_2\text{H}_3\text{O}_2], \text{ M}$	$[\text{H}_3\text{O}^+], \text{ M}$	$[\text{C}_2\text{H}_3\text{O}_2^-], \text{ M}$
Initial	0.30 M	0	0
Change	- x	+ x	+ x
Equilibrium	about 0.30 M	x	x

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

**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  $1.8 \times 10^{-5}$ .

Now, substituting values from the ICE chart into the  $K_a$  expression yields

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

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

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

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

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

**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  $1.8 \times 10^{-5}$ .

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ \text{pH} &= -\log (2.3 \times 10^{-3}) \\ \text{pH} &= 2.64 \end{aligned}$$

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.

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  $1 \times 10^{-3}$

Answer

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

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

- A 2.00  
 B 2.20  
 C 2.50  
 D 2.17

Answer

$$\begin{aligned} & \text{B} \\ K_a &= [\text{H}^+]^2 / [\text{HA}] \\ 3.2 \times 10^{-4} &= x^2 / 0.122 \\ x^2 &= 3.2 \times 10^{-4} \times 0.122 \\ x^2 &= 3.904 \times 10^{-5} \\ x &= 6.25 \times 10^{-3} \\ \text{pH} &= -\log(6.25 \times 10^{-3}) \\ \text{pH} &= 2.20 \end{aligned}$$

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  $1 \times 10^{-1}$

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  $1 \times 10^{-1}$

Answer

$$\begin{aligned} & \text{D} \\ K_a &= [\text{H}^+]^2 / [\text{HA}] \\ [\text{HA}] &= [\text{H}^+]^2 / K_a \\ [\text{H}^+] &= 10^{-\text{pH}} \\ [\text{H}^+] &= 10^{-4} \\ [\text{HA}] &= (1 \times 10^{-4})^2 / 1 \times 10^{-7} \\ [\text{HA}] &= 1 \times 10^{-1} \end{aligned}$$

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

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%?

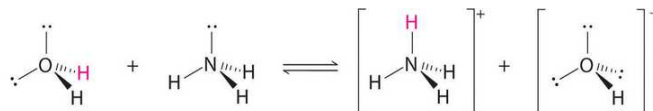
- 1.15  
 4.22  
 3.21  
 2.13  
 6.78

Answer

$$\begin{aligned} & \text{C} \\ \% \text{ ionization} &= [\text{H}^+] / [\text{HA}] \times 100\% \\ [\text{H}^+] &= \% \text{ ionization} / 100 \times [\text{HA}] \\ [\text{H}^+] &= 1.22 / 100 \times 0.05\text{M} \\ [\text{H}^+] &= 6.1 \times 10^{-4} \\ \text{pH} &= -\log [\text{H}^+] \\ \text{pH} &= 3.21 \end{aligned}$$

## 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.

**Weak Bases**

The equilibrium constant expression for this reaction is

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

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  $[OH^-]$  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

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

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

**D**  
 Ammonia is the only weak base so its  $K_b$  is very small.

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.

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

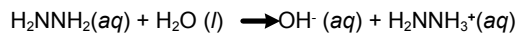
- A This is a concentrated base.  
 B This base ionizes slightly in aqueous solution.  
 C This is a strong base.  
 D An aqueous solution of this would be acidic.

**B**  
 This base ionizes slightly in aqueous solution. Its  $K_b$  value is very small which means the amount of reactants is greater than the amount of the products.

**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



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

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



### Calculating $K_b$ from the pH

$$\begin{aligned} \text{pOH} &= 14 - \text{pH} \\ \text{pOH} &= 14 - 10.9 \\ \text{pOH} &= 3.1 \end{aligned}$$

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

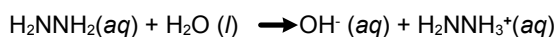
$$-3.1 = \log [\text{OH}^-]$$

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

$$7.94 \times 10^{-4} = [\text{OH}^-] = [\text{H}_2\text{NNH}_3^+]$$

### Calculating $K_b$ from the pH

What is the  $K_b$  of a 0.20 M solution of hydrazine  $\text{H}_2\text{NNH}_2$  at 25°C that has a pH of 10.9?



	$[\text{H}_2\text{NNH}_2], M$	$[\text{OH}^-], M$	$[\text{H}_2\text{NNH}_3^+], M$
Initially	0.20	0	0
Change	$-7.94 \times 10^{-4}$	$+7.94 \times 10^{-4}$	$+7.94 \times 10^{-4}$
At Equilibrium	$0.20 - 7.94 \times 10^{-4} \approx 0.20$	$7.94 \times 10^{-4}$	$7.94 \times 10^{-4}$

### Calculating $K_b$ from the pH

What is the  $K_b$  of a 0.20 M solution of hydrazine  $\text{H}_2\text{NNH}_2$  at 25°C that has a pH of 10.9?

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

$$K_b = \frac{[\text{OH}^-][\text{H}_2\text{NNH}_3^+]}{[\text{H}_2\text{NNH}_2]}$$

$$K_b = \frac{[7.94 \times 10^{-4}][7.94 \times 10^{-4}]}{[0.20]}$$

$$K_b = 3.15 \times 10^{-6}$$

### Calculating Percent Ionization

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

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

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  $2.44 \times 10^{-10}$

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- 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  $2.44 \times 10^{-10}$

Answer

$$\begin{aligned} E \\ K_b &= \frac{[\text{OH}^-]^2}{[\text{B}]} \\ [\text{OH}^-] &= 10^{-\text{pOH}} \\ [\text{OH}^-] &= 10^{-4.98} \\ [\text{OH}^-] &= 1.05 \times 10^{-5} \\ K_b &= (1.05 \times 10^{-5})^2 / 0.450 \\ K_b &= 2.44 \times 10^{-10} \end{aligned}$$

64 Calculate the  $K_b$  of a 0.724M solution of hypobromite ion ( $\text{BrO}^-$ ) that has a pH of 11.23.

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

64 Calculate the  $K_b$  of a 0.724M solution of hypobromite ion ( $\text{BrO}^-$ ) that has a pH of 11.23.

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

Answer

$$\begin{aligned}
 & \text{B} \\
 [\text{H}^+] &= 10^{-\text{pH}} \\
 [\text{H}^+] &= 10^{-11.23} \\
 [\text{H}^+] &= 5.89 \times 10^{-12} \\
 [\text{OH}^-] &= 1 \times 10^{-14} / [\text{H}^+] \\
 [\text{OH}^-] &= 1 \times 10^{-14} / 5.89 \times 10^{-12} \\
 [\text{OH}^-] &= 1.70 \times 10^{-3} \\
 K_b &= [\text{OH}^-]^2 / [\text{BrO}^-] \\
 K_b &= (1.70 \times 10^{-3})^2 \times 0.724 \\
 K_b &= 4.00 \times 10^{-6}
 \end{aligned}$$

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

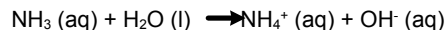
Students type their answers here

Answer

### Calculating pH from the $K_b$

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

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

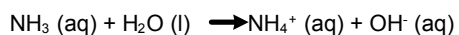


First write the equilibrium expression for the dissociation equation. Obtain the  $K_b$  value from an earlier page.

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

### Calculating pH from the $K_b$

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



	$[\text{NH}_3], M$	$[\text{NH}_4^+], M$	$[\text{OH}^-], M$
Initially	0.15	0	0
Change	-x	+x	+x
At Equilibrium	$0.15 - x \approx 0.15$	x	x

### Calculating pH from the $K_b$

$$\begin{aligned}
 1.8 \times 10^{-5} &= \frac{(x)^2}{(0.15)} \\
 (1.8 \times 10^{-5})(0.15) &= x^2 \\
 2.7 \times 10^{-6} &= x^2 \\
 1.6 \times 10^{-3} &= x
 \end{aligned}$$

Again, remember what your "x" is!

**Calculating pH from the  $K_b$** 

In this case, "x" is  $[\text{OH}^-]$ .

Therefore,

$$\begin{aligned} [\text{OH}^-] &= 1.6 \times 10^{-3} \text{ M} \\ \text{pOH} &= -\log(1.6 \times 10^{-3}) \\ \text{pOH} &= 2.80 \end{aligned}$$

So, now solving for pH:

$$\begin{aligned} \text{pH} &= 14.00 - 2.80 \\ \text{pH} &= 11.20 \end{aligned}$$

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

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

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

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

Answer

$$\begin{aligned} K_b &= \frac{[\text{OH}^-]^2}{[\text{B}]} \\ [\text{OH}^-] &= \sqrt{K_b \times [\text{B}]} \\ [\text{OH}^-] &= \sqrt{4.3 \times 10^{-10} \times 0.550} \\ [\text{OH}^-] &= 1.54 \times 10^{-5} \end{aligned}$$

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

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

Answer

$$\begin{aligned} K_b &= \frac{[\text{OH}^-]^2}{[\text{B}]} \\ [\text{OH}^-] &= \sqrt{K_b \times [\text{B}]} \\ &= \sqrt{4.3 \times 10^{-10} \times 0.550} \\ [\text{OH}^-] &= 1.54 \times 10^{-5} \\ \text{pOH} &= -\log[\text{OH}^-] \\ \text{pOH} &= 4.81 \\ \text{pH} &= 14 - \text{pOH} \\ \text{pH} &= 9.19 \end{aligned}$$

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

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

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

- A 8.94  
 B 11.47  
 C 2.53  
 D 5.06  
 E 3.45

Answer

$$\begin{aligned} & \text{B} \\ K_b &= [\text{OH}^-]^2/[\text{B}] \\ [\text{OH}^-] &= \sqrt{K_b \times [\text{B}]} \\ &= \sqrt{6.4 \times 10^{-5} \times 0.135} \\ &= \sqrt{8.64 \times 10^{-6}} \\ &= 2.94 \times 10^{-3} \\ \text{pOH} &= 2.53 \\ \text{pH} &= 14 - 2.53 \\ \text{pH} &= 11.47 \end{aligned}$$

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

Students type their answers here

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

Students type their answers here

Answer

$$\begin{aligned} & \text{D} \\ K_b &= [\text{OH}^-]^2/[\text{B}] \\ [\text{OH}^-] &= \sqrt{K_b \times [\text{B}]} \\ &= \sqrt{6.4 \times 10^{-4} \times 0.112} \\ &= \sqrt{7.168 \times 10^{-5}} \\ &= 8.47 \times 10^{-3} \\ \text{pOH} &= 1.67 \\ \text{pH} &= 14 - 1.67 \\ \text{pH} &= 12.33 \end{aligned}$$

## Polyprotic Acids

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### Polyprotic Acids

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

Here are some examples:

Sulfuric acid	$\text{H}_2\text{SO}_4$
Phosphoric acid	$\text{H}_3\text{PO}_4$
Carbonic acid	$\text{H}_2\text{CO}_3$
Oxalic acid	$\text{H}_2\text{C}_2\text{O}_4$

### 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

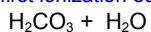


Write the  $K_a$  expression for this equation. This is referred to as  $K_{a1}$ .

## Polyprotic Acids

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The first ionization equation is



Write the  $K_a$  expression for this equation.

Answer

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

## Polyprotic Acids

Now, we examine both dissociation constants.

The second ionization equation for carbonic acid is

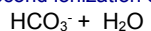


Write the  $K_a$  expression for this equation. This is referred to as  $K_{a2}$ .

## Polyprotic Acids

Now, we examine both dissociation constants.

The second ionization equation for carbonic acid is



Write the  $K_a$  expression for this equation.

Answer

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

## Polyprotic Acids

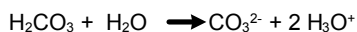
Notice that the bicarbonate ion,  $\text{HCO}_3^-$ , appears in each expression; in the numerator for  $K_{a1}$  and in the denominator for  $K_{a2}$ .

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

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

## Polyprotic Acids

The equation for the complete ionization of carbonic acid is



and the  $K_a$  expression for this reaction is

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

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

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

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

## Polyprotic Acids

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

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- A good hydrogen-ion acceptor
- B good hydrogen-ion donor
- C poor hydrogen-ion acceptor
- D poor hydrogen-ion donor

D  
 Carbonic acid is a weak acid so it poorly donates the hydrogen ion

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  $6.0 \times 10^{24}$

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

B  
 $K = 2.0 \times 10^{-4} \times 3.0 \times 10^{-6}$   
 $K = 6.0 \times 10^{-10}$

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

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

A  
 $K_a = \frac{[\text{H}^+]^2}{[\text{HA}]}$   
 $[\text{H}^+] = \sqrt{K_a \times [\text{HA}]}$   
 $[\text{H}^+] = \sqrt{6.0 \times 10^{-10} \times 0.10}$   
 $= 7.74 \times 10^{-6}$   
 $\text{pH} = -\log(7.74 \times 10^{-6})$   
 5.11

73 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

73 Absorbic acid,  $\text{H}_2\text{C}_2\text{H}_6\text{O}_6$  has the following dissociation constants:

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$$K_{a2} = 1.6 \times 10^{-12}$$

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

A  
 Because the difference between  $K_{a1}$  and  $K_{a2}$  is more than  $10^3$ , only  $K_{a1}$  is relevant.  
 $K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$   
 $[\text{H}^+] = \sqrt{K_a \times [\text{HA}]}$   
 $= \sqrt{8.0 \times 10^{-5} \times .20}$   
 $= 4 \times 10^{-3}$   
 $\text{pH} = -\log(4 \times 10^{-3})$   
 $= 2.40$

## 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

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

### $K_a$ and $K_b$



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 " $K_b$ "?

### $K_a$ and $K_b$



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

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

What do these expressions have in common?

**$K_a$  and  $K_b$** 

$$K_b \times K_a = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

$$K_b \times K_a = \frac{[\cancel{\text{NH}_4^+}][\text{OH}^-]}{[\text{NH}_3]} \frac{[\text{NH}_3][\cancel{\text{H}_3\text{O}^+}]}{[\cancel{\text{NH}_4^+}]}$$

$$K_b \times K_a = [\text{OH}^-][\text{H}_3\text{O}^+] = K_w$$

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

 **$K_a$  and  $K_b$** 

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.

74 For the acid HCN, what is the equation on which the  $K_a$  expression is based?

- A  $\text{HCN} \leftrightarrow \text{OH}^- + \text{HCN}$
- B  $\text{HCN} \leftrightarrow \text{OH}^- + \text{CN}^-$
- C  $\text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{CN}^- + \text{H}_3\text{O}^+$
- D  $\text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{CN}^-$

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- C  $\text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{CN}^- + \text{H}_3\text{O}^+$
- D  $\text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{CN}^-$

Answer

C  
For the acid dissociation constant  $K_a$  for HCN, HCN must be a proton donor. Only C shows HCN acting as a proton donor.

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.

- A  $5.0 \times 10^{-11}$
- B  $2.0 \times 10^{-10}$
- C  $5.0 \times 10^{-10}$
- D  $8.0 \times 10^{-10}$
- E  $2.0 \times 10^{-4}$

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No calculator.

- A  $5.0 \times 10^{-11}$
- B  $2.0 \times 10^{-10}$
- C  $5.0 \times 10^{-10}$
- D  $8.0 \times 10^{-10}$
- E  $2.0 \times 10^{-4}$

Answer

A  
 $K_a \times K_b = K_w$   
 $K_b = K_w / K_a$   
 $K_b = 1 \times 10^{-14} / 2.0 \times 10^{-4}$   
 $.50 \times 10^{-10}$  or  $5.0 \times 10^{-11}$



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

- A  $2.5 \times 10^{-3}$   
 B  $2.5 \times 10^{-4}$   
 C  $2.5 \times 10^{-5}$   
 D  $4.3 \times 10^{-4}$   
 E  $5.6 \times 10^{-11}$

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

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 B  $2.5 \times 10^{-4}$   
 C  $2.5 \times 10^{-5}$   
 D  $4.3 \times 10^{-4}$   
 E  $5.6 \times 10^{-11}$

Answer

$$\begin{aligned}
 &E \\
 &K_a \times K_b = K_w \\
 &K_a = K_w / K_b \\
 &K_a = 1 \times 10^{-14} / 1.8 \times 10^{-5} \\
 &K_a = 5.6 \times 10^{-10}
 \end{aligned}$$

77 Which of the substances below is the strongest acid ?

- A  $\text{HClO}$   
 B  $\text{HCO}_3^-$   
 C  $\text{H}_2\text{S}$   
 D  $\text{NH}_2\text{CH}_3^+$   
 E  $\text{H}_2\text{S}$  and  $\text{HClO}$

77 Which of the substances below is the strongest acid ?

- A  $\text{HClO}$   
 B  $\text{HCO}_3^-$   
 C  $\text{H}_2\text{S}$   
 D  $\text{NH}_2\text{CH}_3^+$   
 E  $\text{H}_2\text{S}$  and  $\text{HClO}$

Answer

C  
The strongest acid would have the weakest conjugate base.  $\text{HS}^-$  is the weakest base. It has the smallest  $K_b$  value. Which makes  $\text{H}_2\text{S}$  the strongest acid. Another explanation is that the acid with the highest  $K_a$  would be the strongest acid.  $\text{H}_2\text{S}$  has the highest  $K_a$ .

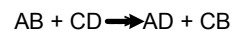
## Acid-Base Properties of Salt Solutions

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## Neutralization Reactions

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:



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

## Neutralization Reactions

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



*Note that the term "salt" refers to any ionic compound that does not include H<sup>+</sup> or OH<sup>-</sup>.*

## 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.

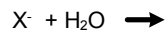
Hydrolyze means to interact with water molecule and dissociate.

## 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<sup>-</sup>. X<sup>-</sup> is the conjugate base of the acid in the neutralization reaction. This anion may undergo hydrolysis as shown below:



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<sup>-</sup>) 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.*

## Anion Effect on pH

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

The conjugate base (NO<sub>2</sub><sup>-</sup>) of a weak acid (HNO<sub>2</sub>) will hydrolyze, it will react with water because HNO<sub>2</sub> does not dissociate completely and a dynamic equilibrium exists.

Write the equilibrium expression for the reaction of NO<sub>2</sub><sup>-</sup> and water.

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

## Anion Effect on pH

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

The conjugate base (HSO<sub>3</sub><sup>-</sup>) of a weak acid (H<sub>2</sub>SO<sub>3</sub>) is a special case because HSO<sub>3</sub><sup>-</sup> is amphoteric. It can behave as either an acid or as a base.

How it behaves is determined by the relative magnitude of K<sub>a</sub> for its acid and K<sub>b</sub> for its base. If K<sub>a</sub> > K<sub>b</sub> it will behave as an acid if K<sub>b</sub> > K<sub>a</sub> then it will behave as a base.

*When amphoteric anions hydrolyze they will make acidic solutions if their K<sub>a</sub> > K<sub>b</sub> or make basic solutions if their K<sub>b</sub> > K<sub>a</sub>.*

**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.*

**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.*

**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  $Fe^{3+}$  and  $Al^{3+}$ .

The effect of these cations increases as ionic charge increases.

The  $K_a$  for  $Fe^{2+}$  is  $3.2 \times 10^{-10}$  and for  $Fe^{3+}$  is  $6.3 \times 10^{-3}$ . These values are in the range for some weak acids we are familiar with like acetic acid  $K_a = 1.8 \times 10^{-5}$ .

**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.

**Acid-Base Properties of Salt Solutions**

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

Compare the following:

attraction between a  $1^+$  cation and  $H_2O$  molecule.

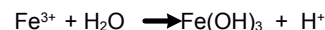
attraction between a  $3^+$  cation and  $H_2O$  molecule.

The cation with the greater charge will make the solution more acidic.

**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:



The increased  $H^+$  concentration reduces the pH.

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

**Combined Effect of Cation and Anion**

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

Why do  $\text{Cl}^-$  and  $\text{Na}^+$  not react with water?

**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?

**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  $\text{Fe}(\text{ClO})_3$ ,  $K_a$  for  $\text{Fe}^{3+}$   $6.3 \times 10^{-3}$  is the  $K_b$  for  $\text{ClO}^-$  is  $3.33 \times 10^{-7}$ . This solution would be acidic because  $K_a > K_b$ .

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

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

Answer

NaCl contains a cation from a strong base,  $\text{Na}^+$  and the conjugate base of a strong acid,  $\text{Cl}^-$ . Neither will hydrolyze and the resulting solution will be neutral with a pH of 7.

79 The pH of a solution of 0.1 M solution of  $K_3PO_4$  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  $K_3PO_4$  will be approximately \_\_\_\_\_.

- A 2  
 B 5  
 C 7  
 D 11  
 E 13

Answer

E  
 $K_3PO_4$  contains the cation of a strong base,  $K^+$ , that will not react with water and the conjugate base,  $PO_4^{3-}$ , of a weak acid,  $HPO_4^{2-}$ . The  $PO_4^{3-}$  will hydrolyze increasing the  $OH^-$  concentration, making a basic solution of pH approximately 13.

80 The pH of a solution of 0.1M solution of  $LiHCO_3$  at 25° C 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° C will be approximately \_\_\_\_\_.

- A 2  
 B 5  
 C 7  
 D 11  
 E 13

Answer

D  
 $LiHCO_3$  contains the cation of a strong base,  $Li^+$ , that will not react with water and the amphoteric, conjugate base,  $HCO_3^-$ , of a weak acid,  $H_2CO_3$ . The  $K_a$  for  $HCO_3^-$  is  $5.6 \times 10^{-11}$  and the  $K_b = 1 \times 10^{-14} / K_a$ .  $K_b = 1.79 \times 10^{-4}$ . The  $K_b$  is greater than  $K_a$  therefore in water  $HCO_3^-$  will act like a base with a pH approximately 11.

81 A 0.1M aqueous solution of \_\_\_\_\_ will have a pH of 7.0 at 25°C.

- A NaOCl  
 B BaBr<sub>2</sub>  
 C NH<sub>4</sub>Cl  
 D Ca(C<sub>2</sub>H<sub>3</sub>O<sub>2</sub>)<sub>2</sub>  
 E none of these

81 A 0.1M aqueous solution of \_\_\_\_\_ will have a pH of 7.0 at 25°C.

- A NaOCl  
 B BaBr<sub>2</sub>  
 C NH<sub>4</sub>Cl  
 D Ca(C<sub>2</sub>H<sub>3</sub>O<sub>2</sub>)<sub>2</sub>  
 E none of these

Answer

B  
BaBr<sub>2</sub> contains the cation of a strong base, Ba<sup>2+</sup> and the conjugate base of a strong acid, Br<sup>-</sup>. Neither will hydrolyze and the resulting solution will be neutral.

82 A 0.1M aqueous solution of \_\_\_\_\_ has a pH of less than 7.0 at 25°C.

- A  $\text{Na}_2\text{S}$   
 B  $\text{KF}$   
 C  $\text{NaNO}_3$   
 D  $\text{NH}_4\text{Cl}$   
 E  $\text{NaF}$

82 A 0.1M aqueous solution of \_\_\_\_\_ has a pH of less than 7.0 at 25°C.

- A  $\text{Na}_2\text{S}$   
 B  $\text{KF}$   
 C  $\text{NaNO}_3$   
 D  $\text{NH}_4\text{Cl}$   
 E  $\text{NaF}$

Answer

D  
 $\text{NH}_4\text{Cl}$  contains the conjugate acid of a weak base,  $\text{NH}_4^+$  and the conjugate base of a strong acid,  $\text{Cl}^-$ . Only the  $\text{NH}_4^+$  will hydrolyze and the resulting solution will be acidic.

83 An aqueous solution of \_\_\_\_\_ will produce a basic solution at 25°C.

- A  $\text{NH}_4\text{ClO}_4$   
 B  $\text{KBr}$   
 C  $\text{NaCl}$   
 D  $\text{CaCO}_3$   
 E  $\text{Fe}(\text{ClO}_3)_2$

83 An aqueous solution of \_\_\_\_\_ will produce a basic solution at 25°C.

- A  $\text{NH}_4\text{ClO}_4$   
 B  $\text{KBr}$   
 C  $\text{NaCl}$   
 D  $\text{CaCO}_3$   
 E  $\text{Fe}(\text{ClO}_3)_2$

Answer

D  
 $\text{CaCO}_3$  contains the cation of a strong base,  $\text{Ca}^{2+}$  and the conjugate base of a weak acid,  $\text{CO}_3^{2-}$ . The  $\text{CO}_3^{2-}$  will hydrolyze with water and produce a basic solution.  $\text{NH}_4\text{ClO}_4$  and  $\text{Fe}(\text{ClO}_3)_2$  will make acidic solutions.  $\text{KBr}$  and  $\text{NaCl}$  will make neutral solutions.

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

- A  $\text{NaNO}_3$   
 B  $\text{Na}_2\text{CO}_3$   
 C  $\text{NH}_4\text{Cl}$   
 D  $\text{NaHSO}_4$   
 E  $\text{Na}_2\text{SO}_4$

Question from the College Board

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

- A  $\text{NaNO}_3$   
 B  $\text{Na}_2\text{CO}_3$   
 C  $\text{NH}_4\text{Cl}$   
 D  $\text{NaHSO}_4$   
 E  $\text{Na}_2\text{SO}_4$

Answer

E  
 The only two salts that will result in a basic solution are  $\text{Na}_2\text{CO}_3$  and  $\text{Na}_2\text{SO}_4$ . The  $K_b$  for  $\text{CO}_3^{2-}$  is  $1.79 \times 10^{-4}$ . The  $K_b$  for  $\text{SO}_4^{2-}$  is  $8.33 \times 10^{-13}$ . A 1-molar solution of  $\text{CO}_3^{2-}$  would be the most basic and would have the highest pH.

Question from the College Board

85 An aqueous solution of \_\_\_\_\_ will produce a neutral solution.

- A  $\text{NH}_4\text{ClO}_4$
- B  $\text{AlBr}_3$
- C  $\text{FeCl}_3$
- D  $\text{SrCO}_3$
- E  $\text{BaCl}_2$

85 An aqueous solution of \_\_\_\_\_ will produce a neutral solution.

- A  $\text{NH}_4\text{ClO}_4$
- B  $\text{AlBr}_3$
- C  $\text{FeCl}_3$
- D  $\text{SrCO}_3$
- E  $\text{BaCl}_2$

Answer

E  
 $\text{BaCl}_2$  contains the cation of a strong base and the conjugate base of a strong acid. Neither the cation or anion will hydrolyze and the resulting solution will be neutral.

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

$\text{NH}_4\text{Cl}$ ,  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{K}_2\text{CO}_3$ ,  $\text{NaF}$

- A  $\text{NH}_4\text{Cl}$  and  $\text{Cu}(\text{NO}_3)_2$
- B  $\text{NH}_4\text{Cl}$  and  $\text{K}_2\text{CO}_3$
- C  $\text{K}_2\text{CO}_3$  and  $\text{NaF}$
- D  $\text{NaF}$  only
- E  $\text{NH}_4\text{Cl}$  only

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

$\text{NH}_4\text{Cl}$ ,  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{K}_2\text{CO}_3$ ,  $\text{NaF}$

- A  $\text{NH}_4\text{Cl}$  and  $\text{Cu}(\text{NO}_3)_2$
- B  $\text{NH}_4\text{Cl}$  and  $\text{K}_2\text{CO}_3$
- C  $\text{K}_2\text{CO}_3$  and  $\text{NaF}$
- D  $\text{NaF}$  only
- E  $\text{NH}_4\text{Cl}$  only

Answer

C  
 Only  $\text{K}_2\text{CO}_3$  and  $\text{NaF}$  contain anions that will hydrolyze and produce basic solutions. The other two salts will hydrolyze and form acidic solutions.

## Acid-Base Properties of Salt Solutions

Summary: Anions

The conjugate bases of strong acids will not undergo hydrolysis and not affect pH.  
 (Examples:  $\text{Cl}^-$ ,  $\text{I}^-$ ,  $\text{NO}_3^-$ )

The conjugate bases of weak acids will hydrolyze and will raise pH by creating excess  $\text{OH}^-$  ions.  
 (Examples:  $\text{NO}_2^-$ ,  $\text{ClO}^-$ ,  $\text{CO}_3^{2-}$ )

Amphoteric anions will either act acidic or basic based on the  $K_a$  and  $K_b$  values of the acid and base.  
 (Examples:  $\text{HCO}_3^-$ ,  $\text{H}_2\text{PO}_4^-$ )

## Acid-Base Properties of Salt Solutions

Summary: Cations

Cations from strong bases will not undergo hydrolysis and will not affect pH.  
 (Examples,  $\text{Na}^+$ ,  $\text{Li}^+$ ,  $\text{Ca}^{2+}$ )

The conjugate acids of weak bases will hydrolyze and will lower pH by creating excess  $\text{H}^+$  ions.  
 (Examples:  $\text{NH}_4^+$ ,  $\text{C}_2\text{H}_5\text{NH}_3^+$ )

Small metal cations from weak bases will hydrolyze and lower pH by creating excess  $\text{H}^+$  ions. The higher the charge on the cation the greater the affect on pH.  
 (Examples  $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$ ,  $\text{Al}^{3+}$ )

## Factors Affecting Acid Strength

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## 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



## Factors Affecting Acid Strength

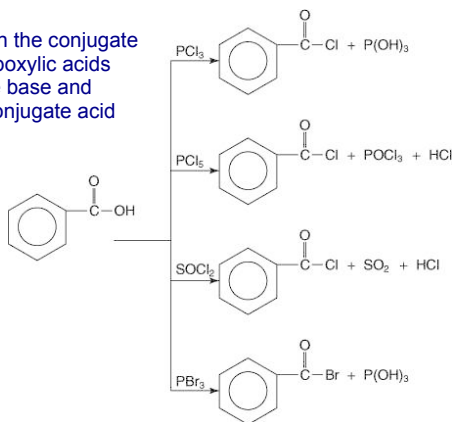
Among oxyacids, the more oxygen atoms in the molecule, the stronger the acid would be.

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



## Slide 204 (Answer) / 208

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

- A HCl
- B HBr
- C HI
- D they would all

Answer

C

HI would be the strongest acid because I is the largest atom. Bond strength decreases as the size of the atom bonded to H increases. Less bond strength makes the substance more readily lose the H<sup>+</sup>.

## Slide 205 / 208

88 H<sub>2</sub>S is a stronger acid than H<sub>2</sub>O because \_\_\_\_\_.

- A The strength of the H-X bond is greater in H<sub>2</sub>O
- B The strength of the H-X bond is less than in H<sub>2</sub>O
- C The H-X bond is more polar in H<sub>2</sub>O than in H<sub>2</sub>S
- D The H-X bond is more polar in H<sub>2</sub>S than in H<sub>2</sub>O

## Slide 205 (Answer) / 208

88 H<sub>2</sub>S is a stronger acid than H<sub>2</sub>O because \_\_\_\_\_.

- A The strength of the H-X bond is greater in H<sub>2</sub>O
- B The strength of the H-X bond is less than in H<sub>2</sub>O
- C The H-X bond is more polar in H<sub>2</sub>O than in H<sub>2</sub>S
- D The H-X bond is more polar in H<sub>2</sub>S than in H<sub>2</sub>O

Answer

A

As you go down a group the size of the atoms increases, as the size increases the bond strength decreases. S is a bigger atom than O and the bond strength is less in H<sub>2</sub>S and in H<sub>2</sub>S would be the stronger acid.

## Slide 206 / 208

89 Order the following acids from weakest to strongest:  
H<sub>2</sub>O, HF, NH<sub>3</sub>, and CH<sub>4</sub>.

- A HF<H<sub>2</sub>O<NH<sub>3</sub><CH<sub>4</sub>
- B HF<NH<sub>3</sub><H<sub>2</sub>O<CH<sub>4</sub>
- C CH<sub>4</sub><HF<NH<sub>3</sub><H<sub>2</sub>O
- D NH<sub>3</sub><CH<sub>4</sub><H<sub>2</sub>O<HF
- E CH<sub>4</sub><NH<sub>3</sub><H<sub>2</sub>O<HF

Answer

## Slide 207 / 208

90 Arrange the following compounds in order of decreasing acid strength: H<sub>2</sub>SO<sub>4</sub>, H<sub>2</sub>SeO<sub>3</sub>, H<sub>2</sub>SeO<sub>4</sub>. Justify your response.

Students type their answers here

## Slide 207 (Answer) / 208

90 Arrange the following compounds in order of decreasing acid strength: H<sub>2</sub>SO<sub>4</sub>, H<sub>2</sub>SeO<sub>3</sub>, H<sub>2</sub>SeO<sub>4</sub>. Justify your response.

Students type their answers here

Answer

The correct ordering is H<sub>2</sub>SO<sub>4</sub>>H<sub>2</sub>SeO<sub>4</sub>>H<sub>2</sub>SeO<sub>3</sub>. For oxyacids like H<sub>2</sub>SO<sub>4</sub> and H<sub>2</sub>SeO<sub>4</sub> which have the same number of O atoms and the same number of OH groups, the electronegativity of the central atom becomes important. Because S is more electronegative than Se, H<sub>2</sub>SO<sub>4</sub> is the stronger acid. For acids with the same central atom such as H<sub>2</sub>SeO<sub>4</sub> and H<sub>2</sub>SeO<sub>3</sub> the acidity increases with the number of oxygen atoms bonded to the central atom, therefore H<sub>2</sub>SeO<sub>4</sub> is a stronger acid than H<sub>2</sub>SeO<sub>3</sub>.

