

1. Calculate $[H^+]$ in a 2.00 L solution of hydrogen chloride in which 3.65 g of HCl is dissolved. K_a for HCl is very large.

We will need to know the concentration of the acid, HCl.

Concentration is measured as molarity which has the units $\frac{mol}{L}$

Since the question tells us the *mass* of HCl in grams, we will need to convert grams to moles. To do so we need to know the molar mass of HCl.

element	Molar Mass (g/mol)	×	No. Atoms	=		
H	1.0	×	1	=	1.0	
Cl	35.5	×	1	=	35.5	
Molar Mass compound (g/mol)					=	36.5 $\frac{g}{mol}$

Unit analysis may be used to calculate the concentration of the acid solution:

$$[HCl] = 3.65g \times \frac{mol}{36.5g} \times \frac{1}{2.00L} = 0.05M$$

Next write a balanced equation: $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

Since HCl is a strong acid, it breaks down completely into ions. BECAUSE OF THIS, since there is a 1:1 relationship between HCl and H^+ , the $[HCl]$ will equal the $[H^+]$:

$$[H^+] = [HCl] = 0.05 M$$

Answer

2. Calculate $[H^+]$ in a solution containing 3.20 g of HNO_3 in 250 mL of solution. Nitric acid is a very strong acid.

This is a similar question – again we need to find $[H^+]$ for a strong acid. Begin by calculating the concentration of the acid solution.

This will involve converting grams HNO_3 into moles HNO_3 :

element	Molar Mass (g/mol)	×	No. Atoms	=		
H	1.0	×	1	=	1.0	
N	14.0	×	1	=	14.0	
O	16.0	×	3	=	48.0	
Molar Mass compound (g/mol)					=	63.0 $\frac{g}{mol}$

Unit analysis may be used to calculate the concentration of the acid solution:

$$[HNO_3] = 3.20g \times \frac{mol}{63.0g} \times \frac{1}{0.250L} = 0.203M$$

Next write a balanced equation: $\text{HNO}_{3(\text{aq})} \rightarrow \text{H}^+_{(\text{aq})} + \text{NO}_3^-_{(\text{aq})}$

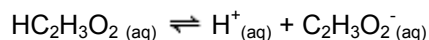
Since HNO_3 is a strong acid, it breaks down completely into ions. BECAUSE OF THIS, since there is a 1:1 relationship between HNO_3 and H^+ , the concentration of HNO_3 will equal the concentration of H^+ :

$$[\text{H}^+] = [\text{HNO}_3] = 0.203 \text{ M}$$

Answer

3. An acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) solution is 0.25 M. Given that K_a for acetic acid is 1.8×10^{-5} , find $[\text{H}^+]$.

The question gives us the concentration of the acid (0.25 M), so we may begin by writing a balanced equation. We need this in order to set up the K_a expression.



Because this is a **weak acid**, $[\text{H}^+]$ **WILL NOT EQUAL** $[\text{HC}_2\text{H}_3\text{O}_2]$.

We will need to use the K_a expression to find $[\text{H}^+]$:

Set up the equation based on the balanced equation:

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

Substitute in known values, and let the unknown values = x:

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.25]}$$

Rearrange the equation

$$x^2 = (1.8 \times 10^{-5})(0.25)$$

$$x^2 = 4.5 \times 10^{-6}$$

Take the square root to find x:

$$\sqrt{x^2} = \sqrt{4.5 \times 10^{-6}}$$

$$x = [\text{H}^+] = 2.1 \times 10^{-3} \quad \text{Answer}$$

$$\text{also } [\text{C}_2\text{H}_3\text{O}_2] = 2.1 \times 10^{-3}$$

4. A solution of acetic acid contains 12.0 g of $\text{HC}_2\text{H}_3\text{O}_2$ in 500 mL of solution. Calculate $[\text{H}^+]$.

We will need to know the concentration of the acid, $\text{HC}_2\text{H}_3\text{O}_2$

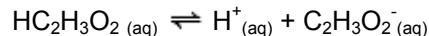
Since the question tells us the *mass* of $\text{HC}_2\text{H}_3\text{O}_2$ in grams, we will need to convert grams to moles. To do so we need to know the molar mass of $\text{HC}_2\text{H}_3\text{O}_2$.

element	Molar Mass (g/mol)	×	No. Atoms	=	
H	1.0	×	4	=	4.0
C	12.0	×	2	=	24.0
O	16.0	×	2	=	32.0
Molar Mass compound (g/mol)					= 60.0 $\frac{\text{g}}{\text{mol}}$

Unit analysis may be used to calculate the concentration of the acid solution:

$$[HC_2H_3O_2] = 12.0g \times \frac{mol}{60.0g} \times \frac{1}{0.500L} = 0.400M$$

Next we need a balanced equation which will be needed in order to set up the K_a expression.



Because this is a **weak acid**, $[H^+]$ **WILL NOT EQUAL** $[HC_2H_3O_2]$.

We will need to use the K_a expression to find $[H^+]$. K_a for this acid was given in the previous question or can be found in the Table of Acid Strengths:

Set up the equation based on the balanced equation:

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

Substitute in known values, and let the unknown values = x:

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.400]}$$

Rearrange the equation

$$x^2 = (1.8 \times 10^{-5})(0.400)$$

$$x^2 = 7.2 \times 10^{-6}$$

Take the square root to find x:

$$\sqrt{x^2} = \sqrt{7.2 \times 10^{-6}}$$

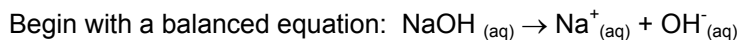
$$x = [H^+] = 2.7 \times 10^{-3}$$

Answer

also $[C_2H_3O_2^-] = 2.7 \times 10^{-3}$

5. Calculate $[H^+]$ and $[OH^-]$ at 25° C in:

a. a 5.0 M NaOH solution. NaOH is a strong base.



Since this is a strong base and there is a 1:1 ratio between NaOH and OH^- ,

$[OH^-] = [NaOH] = 5.0 M$ first part of the answer

Find $[H^+]$ using K_w :

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

Rearrange this equation to find the unknown; in this case $[H^+]$:

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

Substitute in values and solve for $[H^+]$:

$$[H^+] = \frac{1.0 \times 10^{-14}}{[5.0]}$$

K_w will always equal 1.0×10^{-14}

$$[H^+] = 2.0 \times 10^{-15} M$$

answer

- b. a 0.025 M $\text{Ca}(\text{OH})_2$ solution. $\text{Ca}(\text{OH})_2$ is a strong base.

Begin with a balanced equation: $\text{Ca}(\text{OH})_2 (\text{aq}) \rightarrow \text{Ca}^{2+} (\text{aq}) + 2 \text{OH}^- (\text{aq})$

Since this is a strong base and there is a 1:2 ratio between $\text{Ca}(\text{OH})_2$ and OH^- ,

$$[\text{OH}^-] = 2 \times [\text{Ca}(\text{OH})_2] = 2 \times 0.025 = 0.050 \text{ M} \quad \text{first part of the answer}$$

Find $[\text{H}^+]$ using K_w :

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

Rearrange this equation to find the unknown; in this case $[\text{H}^+]$:

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

Substitute in values and solve for $[\text{H}^+]$:

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[0.050]}$$

K_w will always
equal
 1.0×10^{-14}

$$[\text{H}^+] = 2.0 \times 10^{-13} \text{ M}$$

answer

- c. a 0.10 M HCl solution. HCl is a strong acid

Begin with a balanced equation: $\text{HCl} (\text{aq}) \rightarrow \text{H}^+ (\text{aq}) + \text{Cl}^- (\text{aq})$

Since this is a strong acid and there is a 1:1 ratio between HCl and H^+ ,

$$[\text{H}^+] = [\text{HCl}] = 0.10 \text{ M} \quad \text{first part of the answer}$$

Find $[\text{OH}^-]$ using K_w :

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

Rearrange this equation to find the unknown; in this case $[\text{OH}^-]$:

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

Substitute in values and solve for $[\text{OH}^-]$:

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[0.10]}$$

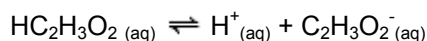
K_w will always
equal
 1.0×10^{-14}

$$[\text{OH}^-] = 1.0 \times 10^{-13} \text{ M}$$

answer

- d. a 0.01 M $\text{HC}_2\text{H}_3\text{O}_2$ solution. $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid with $K_a = 1.8 \times 10^{-5}$.

Because this is a weak acid, we need to do some calculations in order to determine $[\text{H}^+]$. But we still begin with a balanced equation:



We need to set up the K_a equation in order to find $[\text{H}^+]$ for this weak acid:

Set up the K_a equation based on the balanced equation:

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

Substitute in known values, and let the unknown values = x :

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.010]}$$

Rearrange the equation

$$x^2 = (1.8 \times 10^{-5})(0.010)$$

$$x^2 = 1.8 \times 10^{-7}$$

$$x = [H^+] = 4.2 \times 10^{-4}$$

first part of the answer

Find $[OH^-]$ using K_w :

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

Rearrange this equation to find the unknown; in this case $[OH^-]$:

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

Substitute in values and solve for $[OH^-]$:

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[4.20 \times 10^{-4}]}$$

K_w will always equal 1.0×10^{-14}

$$[OH^-] = 2.4 \times 10^{-11} \text{ M}$$

answer

6. A mass of 1.4 g of KOH is dissolved in water to form 500 mL of solution. What is the concentration of H^+ ions in this solution if the temperature of the solution is 25°C ? KOH is a strong base.

We realize that if we have a base we will first determine $[OH^-]$. Then we will use K_w to find $[H^+]$.

To calculate the concentration of KOH, we need to convert grams to moles. The molar mass of KOH is 56.1 g/mol

$$[KOH] = 1.4 \text{ g} \times \frac{\text{mol}}{56.1 \text{ g}} \times \frac{1}{0.500 \text{ L}} = 0.05 \text{ M}$$

Write the balanced equation: $KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

Since KOH is a strong base and there is a 1:1 ratio between OH^- and KOH: $[OH^-] = [KOH] = 0.05 \text{ M}$

Find $[H^+]$ using K_w :

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

Rearrange this equation to find the unknown; in this case $[H^+]$:

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

Substitute in values and solve for $[H^+]$:

$$[H^+] = \frac{1.0 \times 10^{-14}}{[0.050]}$$

K_w will always equal 1.0×10^{-14}

$$[H^+] = 2.0 \times 10^{-13} \text{ M}$$

answer

7. A mass of 4.0 g of NaOH is dissolved in water to form 500 mL of solution with a temperature of 25° C. Calculate the hydronium ion concentration in this solution.

The concentration of the hydronium ion, H_3O^+ , is found by finding $[\text{H}^+]$.

Since we are given a base, we will first find $[\text{OH}^-]$. Then we use K_w to find $[\text{H}^+]$, which will also equal $[\text{H}_3\text{O}^+]$.

The previous questions show the steps that need to be done. Key steps and answers:

Find $[\text{NaOH}]$.	$[\text{NaOH}] = 0.20 \text{ M}$	
Find $[\text{OH}^-]$	$[\text{OH}^-] = 0.20 \text{ M}$	
Find $[\text{H}^+] = [\text{H}_3\text{O}^+]$	$[\text{H}_3\text{O}^+] = 5.0 \times 10^{-14} \text{ M}$	final answer

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8. Calculate the pH of a solution of nitric acid that is:

- a. $1.0 \times 10^{-4} \text{ M}$

To find pH we must first find $[\text{H}^+]$

The question gives us the concentration of nitric acid, HNO_3 $[\text{HNO}_3] = 1.0 \times 10^{-4} \text{ M}$

The balanced equation: $\text{HNO}_{3(\text{aq})} \rightarrow \text{H}^+_{(\text{aq})} + \text{NO}_3^-_{(\text{aq})}$

Since nitric acid is a strong acid,
and there is a 1:1 ratio between HNO_3 & H^+ : $[\text{H}^+] = 1.0 \times 10^{-4} \text{ M}$

$\text{pH} = -\log[\text{H}^+] = -\log(1.0 \times 10^{-4}) = 4.0$ answer

-
- b. consists of 6.3 g of solute dissolved in 1.00 L of solution?

First calculate the concentration of HNO_3 (whose molar mass is 63.0 g/mol)

$$[\text{HNO}_3] = 6.3\text{g} \times \frac{\text{mol}}{63.0\text{g}} \times \frac{1}{1.0\text{L}} = 0.10\text{M}$$

Since nitric acid is a strong acid,
and there is a 1:1 ratio between HNO_3 & H^+ : $[\text{H}^+] = 0.10 \text{ M}$

$\text{pH} = -\log[\text{H}^+] = -\log(0.1) = 1.0$ answer

9. Calculate the pH of a solution that consists of 5.0 g of HCl in 250 mL of solution?

First calculate the concentration HCl (whose molar mass is 36.5 g/mol)

$$[HCl] = 5.00g \times \frac{mol}{36.5g} \times \frac{1}{0.250L} = 0.55M$$

The balanced equation: $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

Since hydrochloric acid is a strong acid,

and there is a 1:1 ratio between HCl & H^+ :

$$[H^+] = 0.55 \text{ M}$$

$$pH = -\log[H^+] = -\log(0.55) = 0.26 \text{ answer}$$

10. What is the $[H^+]$ of a solution with a pH of 10.00 at 25°C?

$$1.00 \times 10^{-10}$$

$$[H^+] = \text{antilog}(-pH) = \text{antilog}(-10.00) = 1.00 \times 10^{-10} \text{ M answer}$$

11. What is the pH of an aqueous solution containing 0.0020 M barium hydroxide, $Ba(OH)_2$?

To find pH we must first find $[H^+]$. However, barium hydroxide is a base, so we first find $[OH^-]$ and then use K_w to find $[H^+]$

Balanced equation: $Ba(OH)_{2(aq)} \rightarrow Ba^{2+}_{(aq)} + 2 OH^-_{(aq)}$

From the balanced equation we find that the concentration of OH^- will be twice that of $Ba(OH)_2$:

$$[OH^-] = 2 \times [Ba(OH)_2] = 2 \times 0.0020 = 0.0040 \text{ M or } 4.00 \times 10^{-3} \text{ M}$$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-3}} = 2.5 \times 10^{-12}$$

$$pH = -\log[H^+] = -\log(2.5 \times 10^{-12}) = 11.6 \text{ answer}$$

ALTERNATE METHOD:

When you have a base, there is another way to determine pH other than by first calculating $[H^+]$.

Use the relationship: $pH + pOH = 14$ which rearranges to: $pH = 14 - pOH$

$$pOH = -\log [OH^-] = -\log (4.3 \times 10^{-3}) = 2.4$$

Then: $pH = 14 - pOH$

$$pH = 14 - 2.4$$

$$pH = 11.6 \text{ answer}$$

12. Calculate the hydronium ion concentration of:

a) 100.0 mL of an aqueous solution containing 0.60 g of sodium hydroxide, NaOH.

Finding hydronium ion concentration, $[H_3O^+]$ means finding $[H^+]$.

Again, we are given a base so we will first find $[OH^-]$.

$$[NaOH] = 0.60g \times \frac{mol}{40.0g} \times \frac{1}{0.100L} = 0.15M$$

Balanced equation: $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

$$[OH^-] = [NaOH] = 0.15 M$$

Use K_w to find $[H^+]$:
$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.15} = 6.7 \times 10^{-14} M = [H_3O^+] \text{ answer}$$

b) a blood sample with a pH of 7.40

$$[H_3O^+] = [H^+] = \text{antilog}(-\text{pH})$$

$$[H_3O^+] = \text{antilog}(-7.40) = 4.0 \times 10^{-8} M \text{ answer}$$