Chemistry 30

Unit 5: Acids & Bases — Answers

Practice Set 2: 2-1 to 2-4 K_a, K_b, K_w and pH

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)	$\times$	No. Atoms		
H	1.0	$\times$	1	=	1.0
Cl	35.5	$\times$	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**

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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)	$\times$	No. Atoms		
H	1.0	$\times$	1	=	1.0
N	14.0	$\times$	1	=	14.0
O	16.0	$\times$	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: $HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$

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^+:

$[H^+] = [HNO_3] = 0.203$ M **Answer**

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3. An acetic acid ($HC_2H_3O_2$) solution is 0.25 M. Given that K_a for acetic acid is 1.8×10^{-5}, find $[H^+]$.

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.

$HC_2H_3O_2{}_{(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$

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 $[H^+]$:

Set up the equation based on the balanced equation: $K_a = \dfrac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$

Substitute in known values, and let the unknown values = *x*: $1.8 \times 10^{-5} = \dfrac{[x][x]}{[0.25]}$

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

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

also $[C_2H_3O_2^-] = 2.1 \times 10^{-3}$

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4. A solution of acetic acid contains 12.0 g of $HC_2H_3O_2$ in 500 mL of solution. Calculate $[H^+]$.

We will need to know the concentration of the acid, $HC_2H_3O_2$

Since the question tells us the *mass* of $HC_2H_3O_2$ in grams, we will need to convert grams to moles. To do so we need to know the molar mass of $HC_2H_3O_2$.

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 $\frac{g}{mol}$

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.

$HC_2H_3O_{2\,(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$

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

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

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

   Begin with a balanced equation: $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

   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 $Ca(OH)_2$ solution. $Ca(OH)_2$ is a strong base.

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

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

$[OH^-] = 2 \times [Ca(OH)_2] = 2 \times 0.025 = 0.050$ 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 $[H^+]$: $$[H^+] = \frac{1.0 \times 10^{-14}}{[0.050]}$$

K_w will always equal 1.0×10^{-14}

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

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

Begin with a balanced equation: $HCl_{\,(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

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

$[H^+] = [HCl] = 0.10$ M **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 $[OH^-]$: $$[OH^-] = \frac{1.0 \times 10^{-14}}{[0.10]}$$

K_w will always equal 1.0×10^{-14}

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

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d. a 0.01 M $HC_2H_3O_2$ solution. $HC_2H_3O_2$ 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 $[H^+]$. But we still begin with a balanced equation:

$$HC_2H_3O_{2\,(aq)} \rightleftharpoons H^{+}_{(aq)} + C_2H_3O_2^{-}{}_{(aq)}$$

We need to set up the K_a equation in order to find $[H^+]$ 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}

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

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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

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

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: $[OH^-] = [KOH] = 0.05$ 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}

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

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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 [H^+].

Since we are given a base, we will first find [OH^-]. Then we use K_w to find [H^+], which will also equal [H_3O^+].

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

Find [NaOH].	[NaOH] = 0.20 M	
Find [OH^-]	[OH^-] = 0.20 M	
Find [H^+] = [H_3O^+]	**[H_3O^+] = 5.0 × 10^{-14} M**	**final answer**

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

a. 1.0 × 10^{-4} M

To find pH we must first find [H^+]

The question gives us the concentration of nitric acid, HNO_3 [HNO_3] = 1.0 × 10^{-4} M

The balanced equation: $HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$

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

pH = – log[H^+] = – log(1.0 × 10^{-4}) = 4.0 answer

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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)

$$[HNO_3] = 6.3g \times \frac{mol}{63.0g} \times \frac{1}{1.0L} = 0.10M$$

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

pH = – log[H^+] = – log(0.1) = 1.0 answer

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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^+: $[H^+] = 0.55$ M

pH = – log[H^+] = – log(0.55) = 0.26 answer

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10. What is the [H^+] of a solution with a pH of 10.00 at 25°C? ***1.00×10^{-10}***

[H^+] = antilog(-pH) = antilog(-10.00) = **1.00×10^{-10} M answer**

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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 × [$Ba(OH)_2$] = 2 × 0.0020 = 0.0040 M or **4.00×10^{-3} 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(2.5×10^{-12}) = 11.6 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 (4.3×10^{-3}) = 2.4

Then: pH = 14 – pOH

pH = 14 – 2.4

pH = 11.6 **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: $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

$[OH^-] = [NaOH] = 0.15\text{ M}$

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

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b) a blood sample with a pH of 7.40

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

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