

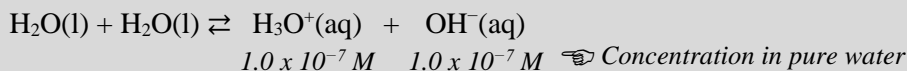
# Exercise 15.1a

## pH & pOH - Answers

Name: \_\_\_\_\_

Date: \_\_\_\_\_ Per: \_\_\_\_\_

In any aqueous solution the water present will self-ionize to form hydronium ( $\text{H}_3\text{O}^+$ ) and hydroxide ( $\text{OH}^-$ ) ions according to the equation below. In pure water these concentrations must be equal due to the mole ratio in which they form during ionization. The amounts produced of each ion in pure water at  $25^\circ\text{C}$  are listed below the ionization equation. As the concentration of either ion changes, the concentration of the other changes to maintain equilibrium. pH and pOH are based on these concentrations. As pH increases, pOH decreases and vice versa.



In this chapter and those that follow, the symbols [ ] placed around a formula or number mean "the molar concentration of".

**DIRECTIONS: Complete the following in the spaces provided:**

1.  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = \boxed{1.0 \times 10^{-7}} \boxed{1.0 \times 10^{-7}} = \boxed{1.0 \times 10^{-14}}$

2. Complete the formulas for calculating the following from  $[\text{OH}^-]$  &  $[\text{H}_3\text{O}^+]$  concentrations:

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

3. What type of relationship exists between the  $[\text{H}_3\text{O}^+]$  and  $[\text{OH}^-]$  in an aqueous solution? inversely proportional.

4. If the concentration of  $[\text{H}_3\text{O}^+]$  in an aqueous solution doubles the  $[\text{OH}^-]$  is halved.

5. Complete the formulas for calculating the following from  $[\text{H}_3\text{O}^+]$  &  $[\text{OH}^-]$  concentrations:

$\text{pH} = \boxed{-\log [\text{H}^+]}$        $\text{pOH} = \boxed{-\log [\text{OH}^-]}$

6. Without using a calculator, find the:

	$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-5} \text{ M}$	$[\text{OH}^-] = 1.0 \times 10^{-3} \text{ M}$	$[\text{OH}^-] = 1.0 \times 10^{-11} \text{ M}$	$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-8} \text{ M}$
pH of a solution with:	5	11	3	8
pOH of a solution with:	9	3	11	6
This solution is:	acid    base	acid    base	acid    base	acid    base

7. Without using a calculator, find the:

	$\text{pH} = 4$	$\text{pOH} = 12$	$\text{pH} = 1$	$\text{pOH} = -1$
$[\text{H}_3\text{O}^+]$ in a solution with:	$1.0 \times 10^{-4}$	$1.0 \times 10^{-2}$	$1.0 \times 10^{-1}$	$1.0 \times 10^{-15}$
$[\text{OH}^-]$ in a solution with:	$1.0 \times 10^{-10}$	$1.0 \times 10^{-12}$	$1.0 \times 10^{-13}$	$1.0 \times 10^1$
This solution is:	acid    base	acid    base	acid    base	acid    base

8. Complete the formulas for calculating the following from pH & pOH:

$[\text{H}_3\text{O}^+] = \boxed{\text{antilog}(-\text{pH})}$        $[\text{OH}^-] = \boxed{\text{antilog}(-\text{pOH})}$

9. Complete the formulas for calculating the following from pOH & pH:

$\text{pH} = \boxed{14.0 - \text{pOH}}$        $\text{pOH} = \boxed{14.0 - \text{pH}}$

(OVER)

# Exercise 15.1a

## pH & pOH - Answers

Name: \_\_\_\_\_

Date: \_\_\_\_\_ Per: \_\_\_\_\_

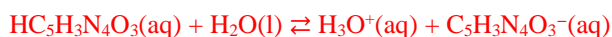
**DIRECTIONS:** Using the logarithm/antilog functions of a calculator, complete the following:

	[H <sup>+</sup> ]	pH	[OH <sup>-</sup> ]	pOH	Acid/Base
10.	$3.50 \times 10^{-3} M$	$-\log(3.50 \times 10^{-3}) = 2.45\bar{6}$	$1.0 \times 10^{-14} / 3.50 \times 10^{-3} = 2.8\bar{57} \times 10^{-12}$	$-\log(2.8\bar{57} \times 10^{-12}) = 11.54\bar{4}$	Acid
11.	$\text{antilog}(-5.30) = 5.0\bar{1} \times 10^{-6}$	5.30	$1.0 \times 10^{-14} / 5.0\bar{1} \times 10^{-6} = 2.0 \times 10^{-9}$	$14.00 - 5.30 = 8.70$	Acid
12.	$1.0 \times 10^{-14} / 6.70 \times 10^{-3} = 1.49\bar{2} \times 10^{-12}$	$-\log(1.49\bar{2} \times 10^{-12}) = 11.82\bar{6}$	$6.70 \times 10^{-3} M$	$-\log(6.70 \times 10^{-3}) = 2.17\bar{4}$	Base
13.	$1.0 \times 10^{-14} / 2.0 \times 10^{-10} = 5.0 \times 10^{-5}$	$14.00 - 9.70 = 4.30$	$\text{antilog}(-9.70) = 2.0 \times 10^{-10}$	9.70	Acid
14.	$4.50 \times 10^{-9} M$	$-\log(4.50 \times 10^{-9}) = 8.34\bar{6}$	$1.0 \times 10^{-14} / 4.50 \times 10^{-9} = 2.22\bar{2} \times 10^{-6}$	$-\log(2.22\bar{2} \times 10^{-6}) = 5.65\bar{3}$	Base
15.	$\text{antilog}(-11.2) = 6.3 \times 10^{-12} *$	11.2	$1.0 \times 10^{-14} / 6.31 \times 10^{-12} = 2.0 \times 10^{-3} *$	$14.0 - 11.2 = 2.8$	Base
16.	$1.0 \times 10^{-14} / 9.10 \times 10^{-11} = 1.09\bar{8} \times 10^{-4}$	$-\log(1.09\bar{8} \times 10^{-4}) = 3.95\bar{9}$	$9.10 \times 10^{-11} M$	$-\log(9.10 \times 10^{-11}) = 10.04\bar{0}$	Acid
17.	$\text{antilog}(-12.60) = 2.5 \times 10^{-13}$	$14.00 - 1.40 = 12.60$	$\text{antilog}(-1.40) = 0.039\bar{8}$	1.40	Base

\* Nasty rounding issues due to sig figs.

**DIRECTIONS:** Using the logarithm/antilog functions of a calculator, complete the following:

18. Uric acid is a chemical created when the body breaks down substances called purines. A 0.128 M solution of uric acid (HC<sub>5</sub>H<sub>3</sub>N<sub>4</sub>O<sub>3</sub>) has a pH of 2.39. Calculate the K<sub>a</sub> of uric acid.



$$[\text{H}_3\text{O}^+]_{\text{equil}} = \text{antilog}(-2.39) = 4.0\bar{7} \times 10^{-3} M$$

Reaction:	HC <sub>5</sub> H <sub>3</sub> N <sub>4</sub> O <sub>3</sub> (aq)	+	H <sub>2</sub> O(l)	⇌	H <sub>3</sub> O <sup>+</sup> (aq)	+	C <sub>5</sub> H <sub>3</sub> N <sub>4</sub> O <sub>3</sub> <sup>-</sup> (aq)	
Initial:	[0.128]		na		0		0	Picture the reaction completely shifted to the left - there would only be HC <sub>5</sub> H <sub>3</sub> N <sub>4</sub> O <sub>3</sub> (aq) and no products.
Change:	- x		na		+ x		+ x	As the reaction proceeds HC <sub>5</sub> H <sub>3</sub> N <sub>4</sub> O <sub>3</sub> (aq) is consumed and produces equal amounts of each product. (1:1 mole ratio)
Equilibrium:	[0.128 - 4.07 × 10 <sup>-3</sup> ]		na		[4.07 × 10 <sup>-3</sup> ]		[4.07 × 10 <sup>-3</sup> ]	The amount of H <sub>3</sub> O <sup>+</sup> (x) can be found based on the pH and substituted for the changes that occurred.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_5\text{H}_3\text{N}_4\text{O}_3^-]}{[\text{HC}_5\text{H}_3\text{N}_4\text{O}_3]} = \frac{[4.0\bar{7} \times 10^{-3}][4.0\bar{7} \times 10^{-3}]}{[0.012\bar{39}]} = 1.33\bar{6} \times 10^{-3}$$

$1.34 \times 10^{-3}$

19. Codeine (C<sub>18</sub>H<sub>21</sub>NO<sub>3</sub>) is a weak organic base. A 5.0 × 10<sup>-3</sup> M solution of codeine has a pH of 9.95. Calculate the value of K<sub>b</sub> for this substance.



If pH = 9.95, then (14.00 - 9.95 = 4.05) pOH = 4.05

$$[\text{OH}^-]_{\text{equil}} = \text{antilog}(-4.05) = 8.9\bar{1} \times 10^{-5} M$$

Reaction:	C <sub>18</sub> H <sub>21</sub> NO <sub>3</sub>	+	H <sub>2</sub> O(l)	⇌	OH <sup>-</sup> (aq)	+	HC <sub>18</sub> H <sub>21</sub> NO <sub>3</sub> <sup>+</sup> (aq)	
Initial:	[5.0 × 10 <sup>-3</sup> ]		na		0		0	Picture the reaction completely shifted to the left - there would only be C <sub>18</sub> H <sub>21</sub> NO <sub>3</sub> (aq) and no products.
Change:	- x		na		+ x		+ x	As the reaction proceeds C <sub>18</sub> H <sub>21</sub> NO <sub>3</sub> (aq) is consumed and produces equal amounts of each product. (1:1 mole ratio)
Equilibrium:	[5.0 × 10 <sup>-3</sup> - 8.91 × 10 <sup>-5</sup> ]		na		[8.91 × 10 <sup>-5</sup> ]		[8.91 × 10 <sup>-5</sup> ]	The amount of OH <sup>-</sup> (x) can be found based on the pH (really pOH) and substituted for the changes that occurred.

$$K_b = \frac{[\text{OH}^-][\text{HC}_{18}\text{H}_{21}\text{NO}_3^+]}{[\text{C}_{18}\text{H}_{21}\text{NO}_3]} = \frac{[8.9\bar{1} \times 10^{-5}][8.9\bar{1} \times 10^{-5}]}{[4.9\bar{1} \times 10^{-3}]} = 1.6\bar{1} \times 10^{-6}$$

$1.6 \times 10^{-6}$