WKS
K_w, pH and pOH

Name Answer Key
Period ____ Date ______________

Use the Acids and Bases equations from Chart B of your reference packet to answer the following questions and calculations.

1) How does the water dissociation constant for water relate to the concentrations of H^+ and OH^- in aqueous solutions?

The ion product, 1.00×10^{-14}, equals the product of $[H^+]$ and $[OH^-]$: $K_w = [H^+][OH^-] = 1.0\times10^{-14}$

2) pH is a **[linear / logarithmic]** function which means that for every change in pH of 1 the concentration of $[H^+]$ changes by **[one / a factor of 10]**.

3) If the pH of one solution is 5 and another has a pH of 2, which solution has a higher $[H^+]$? By what *factor* is this solution stronger than the other solution? Explain or use a calculation to show why.

The solution with pH = 2 has a higher $[H^+]$, $[H^+] = 10^{-pH} = 1.0\times10^{-2}$ M, while the first solution has $[H^+] = 10^{-pH} = 1.0\times10^{-5}$ M. The second solution is 1000 times stronger (more concentrated) because each pH unit is a factor of 10 so a 3 pH unit *decrease* indicates a 10^3 *increase* in $[H^+]$:

$$\frac{[H^+]_2}{[H^+]_1} = \frac{1.0\times10^{-2}}{1.0\times10^{-5}} = 10^3 = 1000$$

4) If you know the $[OH^-]$ of a solution, how would you determine the pH?

First determine $[H^+] = 1.0\times10^{-14}/[OH^-]$, then pH = $-\log[H^+]$ –OR–
Determine pOH = $-\log[OH^-]$, then pH = 14 – pOH

5) Why can one assume that the hydrogen ion concentration in an aqueous solution of a strong monoprotic acid equals the molarity of the acid or that the concentration of hydroxide ion for a strong base equals the molarity of the base? [Hint: what is true for a strong acid or base?]

Strong acids and bases dissociate 100%, so $[H^+]$ = [acid] or $[OH^-]$ = [base]

6) Calculate the pH and pOH of the following solutions of strong acids and bases at 25°C (298 K). Remember question 4 when answering (c) and (d). Be careful with letter (d)—the subscript is important.

a) 1.0 M HI

pH = $-\log(1.0) = 0.00$

pOH = 14.00 – pH = 14.00 – 0.00 = 14.00

b) 0.050 M HNO_3

pH = $-\log(0.050 \text{ M}) = 1.30$

pOH = 14.00 – 1.30 = 12.70

c) 1.0 M KOH

pOH = $-\log(1.0) = 0.00$

pH = 14.00 – 0.00 = 14.00

d) 2.4×10^{-5} M $Mg(OH)_2$

$[OH^-] = 2[Mg(OH)_2] = 4.8\times10^{-5}$
pOH = $-\log(4.8\times10^{-5}) = 4.32$

pH = 14.00 – 4.32 = 9.68

All of the following solutions are at a temperature of 25°C or 298 K:

7) What is the pH of a solution with $[H^+] = 1.00\times10^{-13}$ M? What is the pOH? What is $[OH^-]$? Is it acidic, basic or neutral?

$$pH = -\log(1.00\times10^{-13}) = 13.00$$

$$pOH = 14.00 - 13.00 = 1.00$$
$$[OH^-] = 10^{-1.00} = 0.10\text{ M}$$

–OR–

$$[OH^-] = \frac{1.0\times10^{-14}}{1.00\times10^{-13}} = 0.10\text{ M}$$
$$pOH = -\log(0.10) = 1.00$$

Solution is ***BASIC***

8) What is the pOH of a solution with $[OH^-] = 3.50\times10^{-2}$ M? What is the pH? What is $[H^+]$? Is it acidic, basic or neutral?

$$pOH = -\log(3.50\times10^{-2}) = 1.46$$

$$pH = 14.00 - 1.46 = 12.54$$
$$[H^+] = 10^{-12.54} = 2.88\times10^{-13}\text{ M}$$

–OR–

$$[H^+] = \frac{1.0\times10^{-14}}{3.50\times10^{-2}} = 2.86\times10^{-13}\text{ M}$$
$$pH = -\log(2.86\times10^{-13}) = 12.54$$

Solution is ***BASIC***

9) What are the $[H^+]$, $[OH^-]$, and pOH of a solution whose pH = 2.38? Is it acidic, basic, or neutral?

$$[H^+] = 10^{-2.38} = 4.17\times10^{-3}\text{ M}$$

$$pOH = 14.00 - 2.38 = 11.62$$
$$[OH^-] = 10^{-11.62} = 2.40\times10^{-12}\text{ M}$$

–OR–

$$[OH^-] = \frac{1.0\times10^{-14}}{4.17\times10^{-3}} = 2.40\times10^{-12}$$
$$pOH = -\log(2.40\times10^{-12}) = 11.62$$

Solution is ***ACIDIC***

10) What are the $[H^+]$, $[OH^-]$, and pH of a solution whose pOH = 9.00? Is it acidic, basic, or neutral?

$$[OH^-] = 10^{-9.00} = 1.00\times10^{-9}\text{ M}$$

$$pH = 14.00 - 9.00 = 5.00$$
$$[H^+] = 10^{-5.00} = 1.00\times10^{-5}\text{ M}$$

–OR–

$$[H^+] = \frac{1.0\times10^{-14}}{1.00\times10^{-9}} = 1.00\times10^{-5}$$
$$pH = -\log(1.00\times10^{-5}) = 5.00$$

Solution is ***ACIDIC***