Chemistry 12

Acid-Base Equilibrium III

IV 11-12
pg 134-143

Name:
Date:
Block:

1. pH and pOH

pH and pOH

Complete the following table:

re: $K_w = [H_3O^+][OH^-]$

Solution	**$[H_3O^+]$**	**$[OH^-]$**
1.0 M NaOH	$[H_3O^+] = \frac{K_w}{[OH^-]} = 1.0 \times 10^{-14}$ M	1.0 M
1.0 M HCl $[SA] = [H_3O^+]$	1.0 M	$[OH^-] = \frac{K_w}{[H_3O^+]} = 1.0 \times 10^{-14}$ M

- Concentration of acids and bases can range from extremely high to extremely low.
- It is easier to express these concentrations as **logarithms.** ↳ pH

Logarithms

- "Power of 10" way to specify the concentration of hydronium or hydroxide ions in a solution
- The logarithm of a number is the power to which 10 must be raised to obtain that number.

$$10^y = x \Leftrightarrow \log_{10} x = y$$

(concentration)

ex. $10^3 = 1000 \leftrightarrow \log_{10}(1000) = 3$ (↖ pH)

Practice. Take the log of the following numbers.

(concentration)

$\log(1.0 \times 10^{-9})$ (2sf) = -9.00 (2sf)	$\log(1.0 \times 10^{-7})$ = -7.00	$\log(1.0 \times 10^{-5})$ = -5.00	$\log(1.0 \times 10^{-3})$ = -3.00 ← pH
$\log(5.0 \times 10^{-9})$ = -8.30	$\log(2.4 \times 10^{-7})$ = -6.62	$\log(1.6 \times 10^{-5})$ = -4.80	$\log(7.9 \times 10^{-3})$ = -2.10

We want to avoid negative numbers so we **multiply by –1**. This is called taking the **"negative log".**

$-\log(7.9 \times 10^{-3})$ = 2.10

Practice. Take the NEGATIVE log of the following numbers.

$-\log(1.0 \times 10^{-8})$ (2sf) = 8.00 (2sf)	$-\log(1.0 \times 10^{-6})$ = 6.00	$-\log(1.0 \times 10^{-4})$ = 4.00	$-\log(1.0 \times 10^{-2})$ = 2.00
$-\log(5.0 \times 10^{-8})$ = 7.30	$-\log(2.4 \times 10^{-6})$ = 5.62	$-\log(1.6 \times 10^{-4})$ = 3.80	$-\log(7.9 \times 10^{-2})$ = 1.10

- The reverse of "taking the log" is to "take the antilog" → EXPONENTIAL FORM
- It just simply means to write the number as a power of 10.

pH   Concentration

Antilog (2.0) = $10^{2.0}$ = 100   Antilog (-2.0) = $10^{-2.0}$ = 0.01

Practice. Calculate the following.

$10^{-4.23}$ 2 sf = 5.9×10^{-5}	$10^{-0.34}$ 2 sf = 0.46	$10^{-6.89}$ = 1.3×10^{-7}	$10^{-5.790}$ 3 sf = 1.62×10^{-6}
$10^{-2.1}$ 1 sf = 8×10^{-3}	$10^{-6.71}$ = 1.9×10^{-7}	$10^{-5.33}$ 2 sf = 4.7×10^{-6}	$10^{-1.1}$ = 0.08 1 sf

Significant Figures for Logs:

(pH)

- Only the digits after the decimal place of a log value is significant:

  Ex: –log (**5.28** x 10^{-5}) = [–log (**5.28**)] + [–log (10^{-5})]

  = **–0.723** + (5)

  = 4.**277**

  *Molarity: **5.28** x 10^{-5} M (3 SF)*   *pH = 4.**277** (3 SF)*

Practice. Which solutions have the correct number of significant figures? For the incorrect solutions, write the correct answer below.

–log (5.61x10^{-8}) 3 sf = 7.25 2 sf 7.251 ✓	–log (8.9x10^{-5}) 2 sf = 4.051 3 sf 4.05 ✓	–log (3.0912x10^{-2}) 5 sf = 1.509872895 9 sf 1.50987 ✓	–log (1.0x10^{-10}) 2 sf = 10.00 2 sf ✓
$10^{-4.52}$ 2 sf = 3.02 x 10^{-5} 3 sf 3.0×10^{-5} ✓	$10^{-3.1}$ 1 sf = 8 x 10^{-4} 1 sf ✓	$10^{-1.11}$ 2 sf = 0.078 2 sf ✓	$10^{-0.96}$ 2 sf = 0.1096 4 sf 0.11 ✓

$K_w = [H_3O^+][OH^-]$

Take the negative log…

$-\log K_w = -\log [H_3O^+] + -\log [OH^-]$

Results in…

$pK_w = pH + pOH$

-log = p

Which means:

$pH = -\log [H_3O^+]$ *and* $pOH = -\log [OH^-]$

and

$[H_3O^+] = 10^{-pH}$ and $[OH^-] = 10^{-pOH}$

Practice:

1. What is the pH of a 0.010 M nitric acid solution?

re: [SA] = [H3O+]

$pH = -\log [H_3O^+] = -\log(0.010) = 2.00$ 2 sf *no unit

2. What is the pH of a solution with $[H_3O^+] = 3.2 \times 10^{-4}$ M?

$pH = -\log(3.2 \times 10^{-4}) = 3.49$ 2 sf

3. What is $[H_3O^+]$ of a solution with pH = 2.31?

$[H_3O^+] = 10^{-pH} = 10^{-2.31} = 4.9 \times 10^{-3}$ M 2 sf

4. What is the pOH of a 0.05M NaOH solution?

$pOH = -\log [OH^-] = -\log(0.05) = 1.3$ 1 sf

5. What is the pOH of a solution with $[OH^-] = 2.08 \times 10^{-12}$?

$pOH = -\log(2.08 \times 10^{-12}) = 11.682$ 3 sf

pH + pOH = 14

6. What is the pH of a solution with a pOH = 11.022?

$[OH^-] = 10^{-11.022} = 9.51 \times 10^{-12}$ M

$[H_3O^+] = \frac{K_w}{[OH^-]} = 1.1 \times 10^{-3}$ M

$pH = -\log(1.1 \times 10^{-3}) = 2.958$

14 - 11.022 = 2.978

$$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$$

Take the negative log…

$$-\log K_w = = -\log [H_3O^+] + -\log [OH^-] = -\log (1.00) + -\log (10^{-14})$$

Results in…

$$pK_w = pH + pOH = 0 + 14$$

-log = p

Which means:

$$pH + pOH = 14$$

Practice:

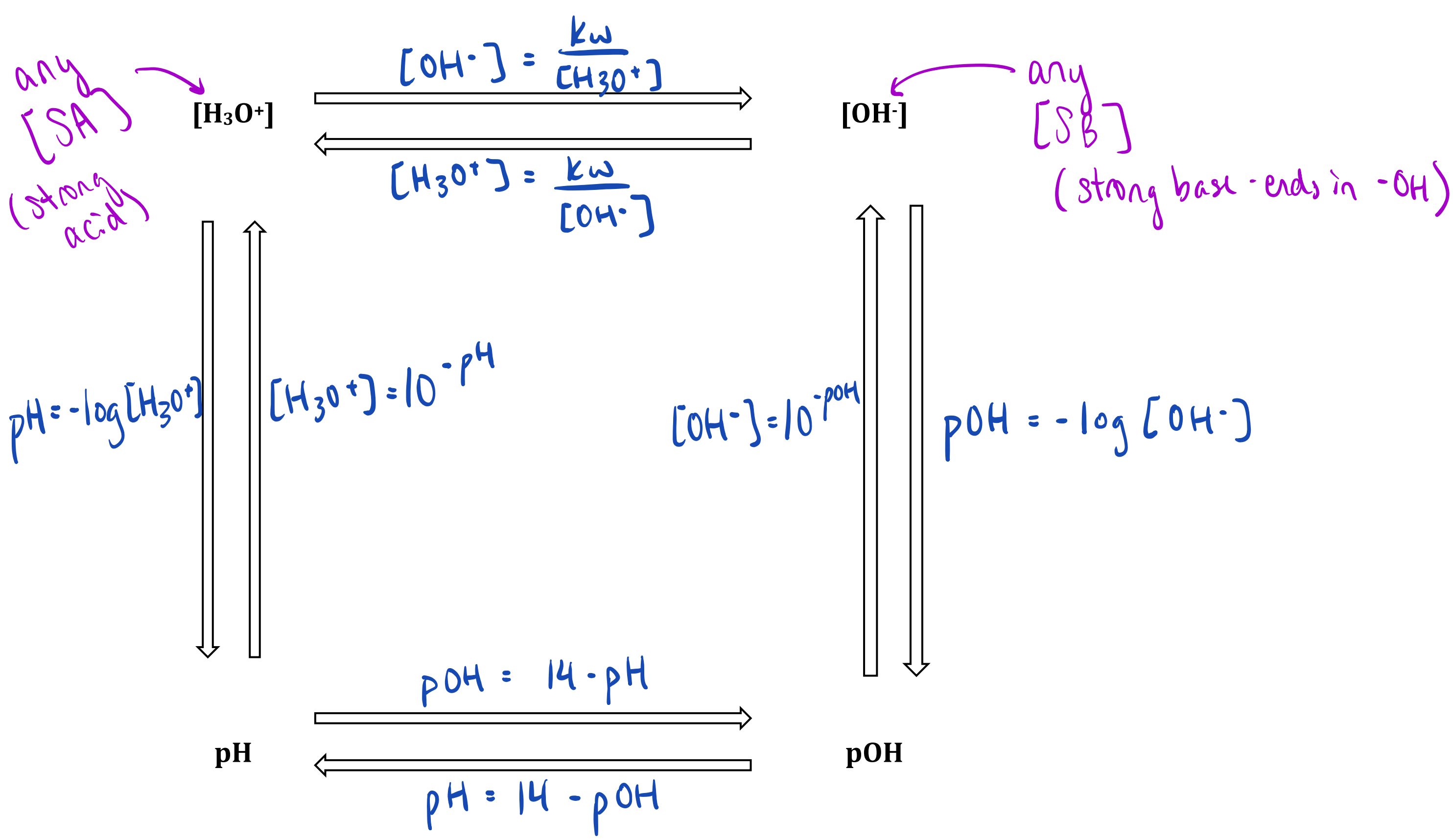
1. If pH = 0.355, what is pOH?

   pOH = 14 − 0.355 = **13.645**

2. If pH = 6.330, what is $[OH^-]$?

   pOH = 14 − 6.330 = 7.670

   $[OH^-] = 10^{-7.670} =$ **2.14×10^{-8} M**



Hebden Workbook: Pg. 139 #49-53, Pg. 141 #55 - 56:

Every other part (a, c, e …)

Determine the pH of the solution that results when 50.0 mL of 0.200 M H_2SO_4 is mixed with 100.0 mL of 0.400 M NaOH.

① Dilution

$[H_2SO_4]$

$C_1V_1 = C_2V_2$

$(0.200M)(50.0mL) = C_2(150.0mL)$

$C_2 = 0.0667M = [H_3O^+]_i$

$[NaOH]$

$(0.400M)(100.0mL) = C_2(150.0mL)$

$C_2 = 0.267M = [OH^-]_i$

bigger! basic solution

② Neutralization

$[OH^-]_f = [OH^-]_i - [H_3O^+]_i$

$= 0.267M - 0.0667M = 0.200M = [OH^-]_f$

$pOH = -\log(0.200) = 0.699$

$pH = 14 - 0.699$

$= 13.301$

2sf

pOH = 2.00

A student adds 35.0 mL of an HCl solution with a pH of 2.00 to 15.0 mL of NaOH solution with a pH of 12.00. Calculate the pH of the final solution.

$[HCl]$

$10^{-2.00} = 0.010M = [H_3O^+]$ ← C_1

$[NaOH]$

$10^{-2.00} = 0.010M = [OH^-]$ ← C_1

① Dilution

$C_2 = \frac{(0.010)(35.0)}{(50.0)} = 0.0070M = [H_3O^+]_i$ ← acidic solution!

$C_2 = \frac{(0.010)(15.0)}{(50.0)} = 0.0030M = [OH^-]_i$

$[H_3O^+]_f = 0.0070M - 0.0030M = 0.0040M$

$pH = -\log(0.0040) = 2.40$

2sf

What mass of NaOH must be added to 500.0 mL of a solution of 0.020 M HI to obtain a solution with a pH of 2.50?

↳ $[H_3O^+]_f$ ← we know there will be an excess of $[H_3O^+]$ b/c pH = 2.50 → acidic

↳ $[H_3O^+]_i$

$[H_3O^+]_f = 10^{-2.50} = 0.0032M$

$[H_3O^+]_f = [H_3O^+]_i - [OH^-]_i$

$[OH^-]_i = 0.020M - 0.0032M$

$= 0.017M$

$0.5000L \times \frac{0.017\,mol}{1L} \times \frac{40.0g}{1\,mol} = 0.34g$

Hebden Workbook Pg. 143 #58 - 68 ☆