*note: each brand of scientific calculator is specific in how you must enter
→ Log
→ antilog

Name:____________________
Blk:___Date:______________

CHEMISTRY 12
ACID BASES UNIT
Lesson #11
pH and pOH

Important Formula's:

pH = $-\log[H_3O^+]$ [H_3O^+] = antilog (−pH)

pOH = $-\log[OH^-]$ [OH^-] = antilog (−pOH)

NOTE: In Chem 12 all log's will be to a base of 10 **!!!!**

Example 1. What is the log of 0.01?

$\log(0.01) = -2.0$

Example 2. What is the log of 10^{-7} ?

$\log(1E(-)7) = -7.0$

Example 3. What is the antilog of 4.0?

antilog (4.0) = 10000 or 1E4

Example 4. If [H_3O^+] = 3.94 x 10^{-4} M , what is the pH?

$pH = -\log[H_3O^+]$

$= -\log[3.94E(-)4]$

$= 3.405$

NOTE: Significant Figures in LOG VALUES * only the decimal places of a log value that are significant

Example 5. If [OH^-] = 9.5 x 10^{-12} M, what is the pOH?

(2)

$pOH = -\log[OH^-]$

$= -\log(9.5 \cdot 10^{-12})$

$= 11.02$

Example 6. If pH = 3.405, what is the [H_3O^+] ?

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

$= \text{antilog}(-3.405) \Rightarrow [H_3O^+] = 3.94 \cdot 10^{-4}$

Example 7. If pOH = 11.68, what is the [OH^-] ?

$[OH^-] = \text{antilog}(-pOH)$

$= \text{antilog}(-11.68) \longrightarrow 2.1 \cdot 10^{-12}$

LOG LAW: Log (A x B) = log (A) + Log (B)

Recall: Kw--> 1.00×10^{-14} = [H_3O^+] [OH^-]

Therefore: $\log(1.00E(-)14) = \log[H_3O^+] + \log[OH^-]$

$-1(-14.000 = \log[H_3O^+] + \log[OH^-])$

pKw $+14.000 = -\log[H_3O^+] + -\log[OH^-]$

$\therefore 14.000 = pH + pOH \longrightarrow pKw = pH + pOH$

Example 8. If pH = 9.355, what is the pOH?

$pKw = pH + pOH$

$14.000 = 9.355 + pOH \rightarrow 14.000 - 9.355 = pOH$

$4.645 = pOH$

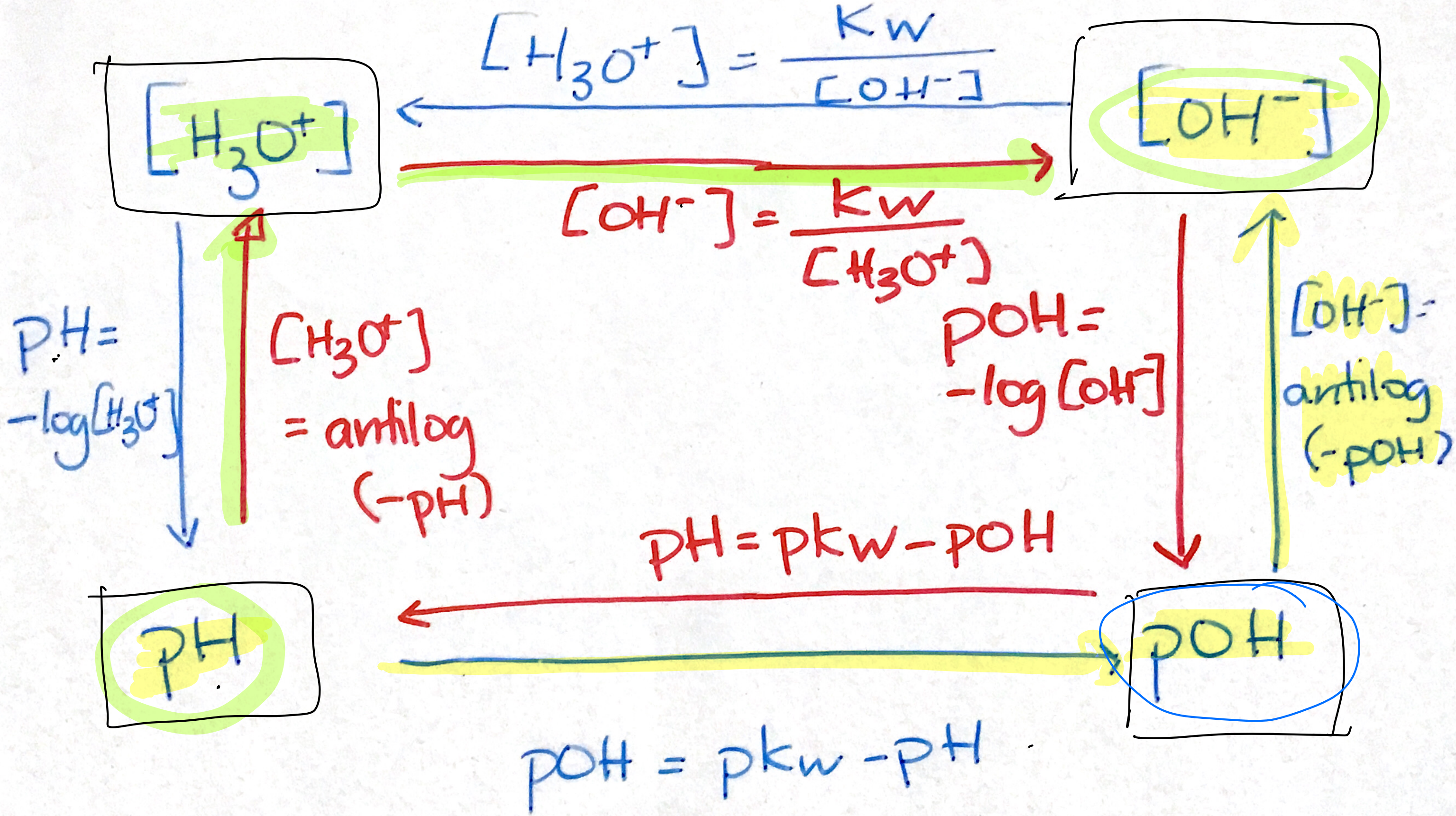
Example 9. If pOH = 2.35, what is the pH?

$pKw = pH + pOH$

$14.000 = pH + 2.35 \rightarrow 14.000 - 2.35 = pH$

$11.65 = pH$

Use the following diagram to work back and forth between [H_3O^+], [OH^-], pH and pOH:


[H3O+]
[OH-]
pH
pOH
[H3O+] = Kw / [OH-]
[OH-] = Kw / [H3O+]
pH = -log[H3O+]
[H3O+] = antilog (-pH)
pOH = -log[OH-]
[OH-] = antilog (-pOH)
pH = pKw - pOH
pOH = pKw - pH

Example 10. If pH = 6.330, what is the $[OH^-]$?

Route #1: pH → pOH → $[OH^-]$

$pOH = 14.000 - 6.330 = 7.670$

$[OH^-] = \text{antilog}(-7.670) = 2.14 \cdot 10^{-8}$ M OH^-

Route #2: pH → $[H_3O^+]$ → $[OH^-]$

$[H_3O^+] = \text{antilog}(-6.330) = 4.68 \cdot 10^{-7}$ M H_3O^+

$$[OH^-] = \frac{1.00 \cdot 10^{-14}}{4.68 \cdot 10^{-7}} = 2.14 \cdot 10^{-8} \text{ M } OH^-$$

The pH SCALE :

The relationships between $[H_3O^+]$, $[OH^-]$, pH and pOH are shown on the following diagram.

pH :	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
$[H_3O^+]$:	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
$[OH^-]$:	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1
pOH :	14	13	12	11	10	9	8	7	6	5	4	3	2	1	0

You should note the following about the diagram.

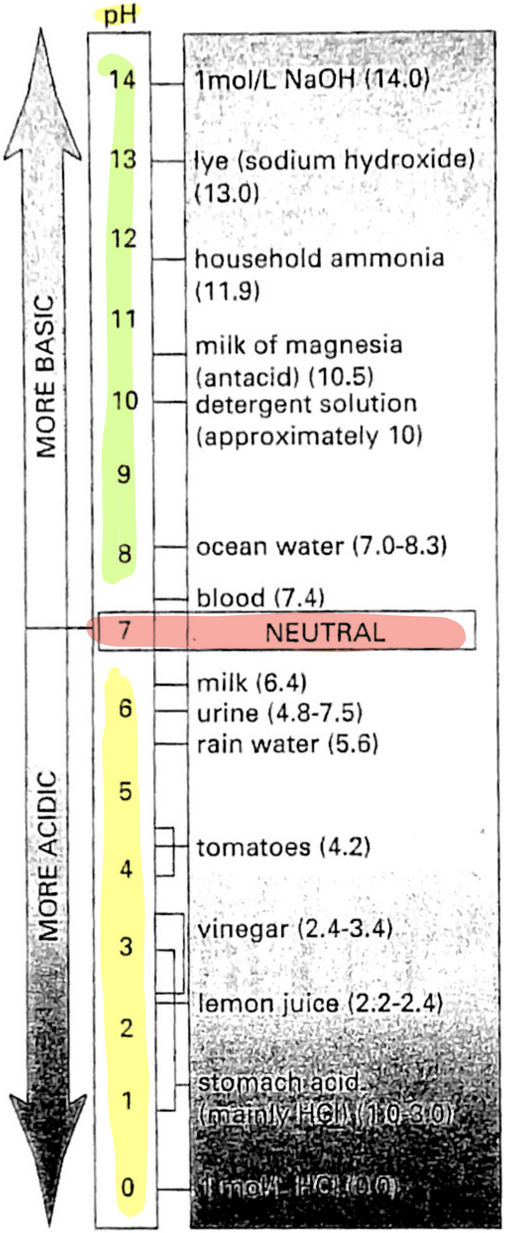
a) The pH scale INCREASES as the pOH scale DECREASES.

b) A solution is ACIDIC when its pH is LESS THAN 7; a solution is BASIC when its pOH is LESS THAN 7. Conversely, a solution is BASIC when its pH is GREATER THAN 7; a solution is ACIDIC when its pOH is GREATER THAN 7. A NEUTRAL solution has pH = pOH = 7.

c) At any point along the horizontal scale it is found, as expected, that $pH + pOH = 14$
and $[H_3O^+][OH^-] = 10^{-14}$.

d) While it is possible for pH to have a value of −1 or 15, say, the pH scale is meant for use in the range 0 to 14. A pH of −1.00 is better handled in terms of its molar concentration: $[H_3O^+] = 10$ M.

The pH Scale with Common Household Substances



The pH values of many common solutions fall within a range from 0 to 14, as shown on this pH scale. The table above the pH scale relates the positive pH values to their hydronium ion concentrations and their logarithms.

Seatwork/Homework: Exercises 47-57 in Hebden
PLO's: L8-L12