Name: Beth "Key" Period: ____________

Grade 12 pH Calculations Practice

1. Calculate pH and pOH for the following solutions :

	pH	pOH		pH	pOH
a) $[H^+] = 1.0 \times 10^{-5}$ M	5.00	9.00	b) $[OH^-] = 3.0 \times 10^{-8}$ M	6.48	7.52
c) $[H^+] = 2.5 \times 10^{-2}$	1.60	12.40	d) $[OH^-] = 7.5 \times 10^{-3}$ M	11.88	2.12
e) $[H^+] = 1.2 \times 10^{-14}$ M	13.92	0.08	f) $[H^+] = 6.0$ M	−0.78	14.78

2. Calculate $[H^+]$ and $[OH^-]$ for the following :

a) pH = 3.0
$[H^+] = 1 \times 10^{-3}M$
$[OH^-] = 1 \times 10^{-11}M$

b) pOH = 2.60
$[H^+] = 4.0 \times 10^{-12}M$
$[OH^-] = 2.5 \times 10^{-3}M$

c) pOH = 5.63
$[H^+] = 4.3 \times 10^{-9}M$
$[OH^-] = 2.3 \times 10^{-6}M$

d) pH = 7.51
$[H^+] = 3.1 \times 10^{-8}M$
$[OH^-] = 3.2 \times 10^{-7}M$

e) pOH = -1.13
$[H^+] = 7.4 \times 10^{-16}M$
$[OH^-] = 13M$

f) pH = 0.03
$[H^+] = 0.93M$
$[OH^-] = 1.1 \times 10^{-14}M$

3. Calculate the pH and the pOH of the following acids :

a) 0.50M Perchloric acid, $HClO_4$

pH = 0.30 $\quad 14 - 0.30 = $ pOH = 13.70

b) 1.3M Hydrochloric acid, HCl

pH = − 0.11 $\quad 14 - (-0.11) = $ pOH = 14.11

c) 0.257M Nitric acid, HNO_3

pH = 0.590 $\quad 14 - 0.590 = $ pOH = 13.410

d) 0.750M Sulfuric Acid, H_2SO_4

$0.750M\ H_2SO_4 = 1.50M\ H^+$

pH = −0.176 $\quad 14 - (-0.176) = $ pOH = 14.176

5. Calculate the pH of the solution formed when 35.0 mL of 1.00 M HCl is mixed with 175.0 mL of 0.25M NaOH.

$HCl + NaOH \rightarrow NaCl + H_2O$

$M = \frac{mol}{L}$ 0.0350mol 0.04375mol

− 0.0350 mol used

0.0088 mol remain

M·L = mol

HCl) 1.00M · 0.0350L = 0.0350mol

NaOH) 0.25M · 0.1750L = 0.04375mol

175.0 mL + 35.0 mL = 210mL = 0.210L

$M_{NaOH} = \frac{0.0088\ mol}{0.210L} = 0.042M$

pOH = −log(0.042M) = 1.38

14.00 − 1.38 = pH = 12.62

Calculations Involving pH

pH is a logarithmic scale so a pH of 9 is 100 times greater than a pH of 7.

On the pH scale: 7 = neutral, <7= acid, >7= base

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

[ ] = indicate concentration in molarity (M)

$$M = \frac{\text{moles of solute}}{\text{L of solution}}$$

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

Water undergoes autoionization which means a water molecule can react with a second water molecule.

$$H_2O(l) + H_2O(l) \leftrightarrow H_3O^+(aq) + OH^-(aq)$$

The equilibrium constant for this reaction is K_w .

$$\mathbf{K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}}$$

H_2O is not included in this expression because it is a pure liquid.

The very tiny value of K_w indicates that very little ionization actually occurs and very few ions will form.

At equilibrium in <u>pure water</u> $[H_3O^+]=[OH^-] = X$ so:

$$1 \times 10^{-14} = X^2$$

$$X = 1 \times 10^{-7}M = [H_3O^+]=[OH^-]$$

$$pH = pOH = 7$$

For all aqueous solutions containing acid or base:

$$\mathbf{[H_3O^+][OH^-] = 1 \times 10^{-14} \quad \text{and} \quad pH + pOH = 14}$$