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 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
$[H^+] = 7.4 \times 10^{-16} M$
$[OH^-] = 13 M$ | f) pH = 0.03
$[H^+] = 0.93 M$
$[OH^-] = 1.1 \times 10^{-14} M$ |

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

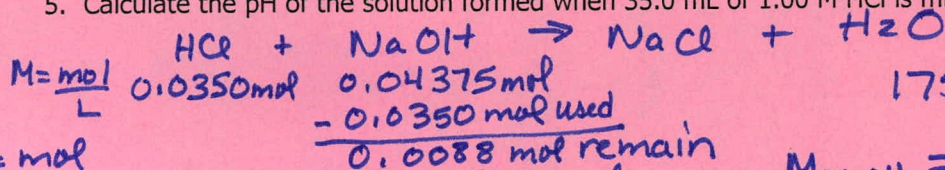
$$pH = 0.590 \quad 14 - 0.590 = pOH = 13.410$$

d) 0.750M Sulfuric Acid, H_2SO_4

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

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

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.



$$M = \frac{mol}{L} \quad \begin{array}{r} 0.0350 mol \\ - 0.0350 mol \text{ used} \\ \hline 0.0088 mol \text{ remain} \end{array}$$

$$175.0 mL + 35.0 mL = 210 mL = 0.210 L$$

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

$$pOH = -\log(0.042 M) = 1.38$$

$$pH = 14.00 - 1.38 = 12.62$$

$$M \cdot L = mol$$

$$HCl) 1.00 M \cdot 0.0350 L = 0.0350 mol$$

$$NaOH) 0.25 M \cdot 0.1750 L = 0.04375 mol$$

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

$$\text{pH} = -\log[\text{H}^+] \quad [] = \text{indicate concentration in molarity (M)}$$

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

$$\text{pOH} = -\log[\text{OH}^-]$$

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



The equilibrium constant for this reaction is K_w .

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

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 pure water $[\text{H}_3\text{O}^+] = [\text{OH}^-] = X$ so:

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

$$X = 1 \times 10^{-7} \text{M} = [\text{H}_3\text{O}^+] = [\text{OH}^-]$$

$$\text{pH} = \text{pOH} = 7$$

For all aqueous solutions containing acid or base:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14} \quad \text{and} \quad \text{pH} + \text{pOH} = 14$$