Outside the pH scale (below 0 and above 14)

What is the pH of 1.00M HCl?

$pH = -\log[H_3O^+]$ ∴ $pH = -\log(1.00M) = 0.000$

1.00×10^{0} ↑ exponent = pH

Therefore, what would the pH be if [HCl] > 1.00M?

The pH would be a negative number.

* Very concentrated acidic solutions (solutions that have $[H_3O^+] >$ 1.0 M) have pH values less than 0. (ie: extremely acidic)
* Very concentrated basic solutions (solutions that have $[OH^-] >$ 1.0 M; $[H_3O^+] <$ 1.0×10^{-14} M) have pH values greater than 14. (extremely basic)

Example: Find the pH and pOH of:

a) 5.0M HNO_3

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

5.0M → 5.0M

$pH = -\log[H_3O^+] = -\log(5.0M) = -0.70$

∴ $pOH = 14.70$

b) 1.5M $Sr(OH)_2$

$Sr(OH)_2 + H_2O \rightarrow Sr^{2+} + 2OH^-$

1.5M → 3.0M (2 × 1.5M)

$pOH = -\log[OH^-] = -\log(3.0M)$

$pOH = -0.48$

∴ $pH = 14.48$

Practice Questions:

1. Find the pH of a 0.0020M solution of HNO_3.

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

0.0020M → 0.0020M (2 s.f.)

$pH = -\log(0.0020M)$

$\boxed{pH = 2.70}$

2. Calculate the pH of a 0.010M NaOH solution.

$NaOH \rightarrow Na^+ + OH^-$

0.010M → 0.010M

$pOH = -\log(0.010M) = 2.00$

$\boxed{\therefore pH = 14 - 2.00 = 12.00}$

3. If the pH is decreased from 5 to 2, what happens to the $[H_3O^+]$ and $[OH^-]$? every change in pH unit by 1 = 10 times Δ in [conc.]

pH Δ 5 → 2 (3 units = 10 × 10 × 10 = 1000)

∴ $[H_3O^+]$ ↑ by 1000 (b/c pH ↓)

∴ $[OH^-]$ ↓ by 1000 (b/c pH ↓)

4. If pH is increased from 7.2 to 8.9, what happens to the $[H_3O^+]$?

pH Δ 7.2 → 8.9 = 1.7 difference.

∴ $10^{1.7} = 50$

∴ because pH is ↑ (more basic) the $[H_3O^+]$ will decrease by 50×

5. Calculate the pH of the final solution if 100.0mL of a strong acid at pH =

4.500 is diluted by adding 50.0mL of water.

Before dilution:
$[H_3O^+] = 10^{-(4.500)}$
$= 3.16 \times 10^{-5} M$

After Dilution.
$M_1V_1 = M_2V_2$
$(3.16 \times 10^{-5})(0.1000) = (0.1500) M_2$
$\therefore M_f = 2.1081 \times 10^{-5} M$
$pH = -\log(2.1081 \times 10^{-5} M) = \boxed{4.676}$

6. By how many pH units does the pH change if 80.0mL of 0.0200M HCl is diluted to a final volume of 160.0mL?

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$
0.0200M 0.0200M

$[H_3O^+] = 0.0200M$
$\therefore pH = -\log(0.0200M)$
$pH_{before} = 1.699$

$M_1V_1 = M_fV_f$
$M_f = \frac{(0.0200)(0.0800)}{(0.1600)}$
$M_f = 0.0100M$

$pH_{after} = -\log(0.0100M) = 2.000$
$pH\ \Delta = 2.000 - 1.699 = \boxed{0.301 \text{ increase}}$
dilution = pH↑ always.

7a) Using your acid/base table for assistance, which has lower pH, a 0.01M solution of HF or a 0.01M solution of CH_3COOH? Why?

HF is a stronger, weak acid ∴ it reacts to a greater extent to produce more $[H_3O^+]$
∴ HF will have a lower pH than CH_3COOH.

b) Which of the solutions above will conduct better? Why? 0.01M HF because it dissociates into ions to a greater extent ∴ ↑[ion] = ↑ conductivity.

Assignment 9:

1. Hebden p. 139 #49deh, 50ef
2. Calculate the pH, pOH and $[OH^-]$ of a 0.00100M solution of HNO_3.
3. Calculate the pOH, pH, and $[H_3O^+]$ of a 2.34×10^{-4}M solution of $Ca(OH)_2$.
4. If the pH is increased from 1 to 6, how do the $[H_3O^+]$ and $[OH^-]$ change?
5. If the pH decreases from 9.3 to 6.5, how does $[H_3O^+]$ change?
6. What is the pH of the final solution if 35.00mL of a strong acid at pH 3.56 is diluted by adding 100.0mL of water?
7. You have 50.00mL of a 0.00345M solution of $HClO_4$. How does the pH change if you dilute the solution to a final volume of 175.0mL?
8. Hebden p.139 #53 & p. 141 #57
9. You dissolve 0.4g of $Ca(OH)_2$ in 500mL of solution. What is the pH?