# Dilution and pH calculations

$$pH = -\log_{10}[H_3O^+]$$

$$10^{-14} = [H_3O^+][OH^-]$$

$$C_1V_1 = C_2V_2$$

Using the three formulae shown on the right answer the following questions.

1) Calculate the pH of a solution that has an $[H_3O^+]$ :

   i. $10^{-4}$ M

   *$pH = -\log_{10}[10^{-4}] = 4$*

   ii. 0.35 M *$= 10^{-0.456} \Rightarrow pH = -\log_{10}[10^{-0.456}] = 0.456$*

   iii. $4.52 \times 10^{-4}$ M *$= 10^{0.655} \times 10^{-4} = 10^{-3.345} \Rightarrow pH = -\log_{10}[10^{-3.345}] = 3.345$*

2) Calculate the pH of a solution that has an $[OH^-]$ :

   i. $10^{-6}$ M

   *$[H_3O^+] = 10^{-14}/[OH^-] = 10^{-14}/10^{-6} = 10^{-8} \Rightarrow pH = -\log_{10}[10^{-8}] = 8$*

   ii. 0.78 M

   *$[H_3O^+] = 10^{-14}/[OH^-] = 10^{-14}/10^{-0.108} = 10^{-13.89} \Rightarrow pH = -\log_{10}[10^{-13.89}] = 13.9$*

   iii. $3.6 \times 10^{-10}$ M *$= 10^{0.556} \times 10^{-10} = 10^{-9.444}$*

   *$[H_3O^+] = 10^{-14}/[OH^-] = 10^{-14}/10^{-9.444} = 10^{-4.556} \Rightarrow pH = -\log_{10}[10^{-4.556}] = 4.56$*

3) Consider the table below. It represents changes made to an original solution. All solutions are at 25°C. Complete the table.

| Volume of original solution (mL) | pH | Volume of water added (mL) | New pH |
|---|---|---|---|
| 300 | 0.55 | 200 | $C_1V_1 = C_2V_2$<br>$C_2 = (0.300L \times 10^{-055}) / 0.500L$<br>$= 0.300 \times 0.282 / 0.500$<br>$[H_3O^+] = 0.169M$<br>$pH = -\log_{10}[0.169] = 0.772$ |
| 150 | $C_1V_1 = C_2V_2$<br>$C_1 = (0.500L \times 10^{-1.20}) / 0.150L$<br>$= 0.500 \times 0.0631 / 0.150$<br>$[H_3O^+] = 0.210M$<br>$pH = -\log_{10}[0.210] = 0.678$ | 350 | 1.20 |
| 200 | 4.52 | $C_1V_1 = C_2V_2$<br>$V_2 = 10^{-4.52} \times 0.200 / 10^{-6.33}$<br>$V_2 = 3.02 \times 10^{-5} \times 0.200/4.68 \times 10^{-7}$<br>*Final volume = 12.91 L*<br>*Volume added to 0.200L is 12.71L* | 6.33 |
| $V_1 = C_2V_2 / C_1$<br><br>$V_2 = V_1 + 0.100L$<br>$V_1 = 10^{-3.53}(x + 0.1) / 10^{-2.34}$<br>$V_1 = 0.0644V_1 + 0.00644$ | 2.34 | 100 | 3.53 |

| | | | |
|---|---|---|---|
| => $0.936V_1 = 0.00644$<br><br>$V_1 = 6.88$ mL | | | |

4) A 350mL sample of an acid solution has 4.52 grams of HCl dissolved in it.

i. Knowing that HCl is a strong acid what can be assumed about the ionisation of HCl in water?

*It is complete so for every mol of HCl that dissolves one mol of $H_3O^+$ will be produced.*

ii. Calculate the $[OH^-]$ of the resulting solution.

*Step 1 calculate the mol of HCl*
*=> 4.52/36.5 = 0.124 mol*
*Step 2 Find the mol of $H_3O^+$*
*=> 0.124 mol*
*Step 3 find $[H_3O^=]$*
*=> 0.124/ 0.350 = 0.354M*
*Step 4 find $[OH^-]$*
*=> $10^{-14}/0.354 = [OH^-] = 10^{-14}/10^{-0.451} = 10^{-13.55}$*

iii. What is the pH of the solution that results?

*$pH = -\log_{10}[0.354] = 10^{-0.451}$*
*= 0.451*

iv. 150 mL of distilled water is added to the 350 mL acid solution. Calculate the pH of the resulting solution.

*$C_1V_1 = C_2V_2$*
*$10^{-0.451}$ X 0.350 / 0.500 = $[H_3O^+] = 0.248M$*
*$pH = -\log_{10}[0.248] = 0.606$*

5) Consider a 400 mL solution with a $[H_3O^+]$ of $10^{-3.524}$M.

i. Calculate the $[OH^-]$

*$[OH^-] = 10^{-14} / [H_3O^+] = 10^{-10.476}$*

ii. Calculate the pH of the solution.

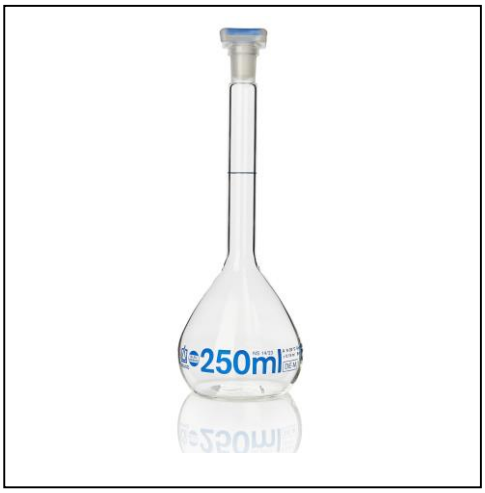
*3.524*

iii. Calculate the pH of the resulting solution when 200 mL of distilled water is added to the 400 mL solution.

*Step 1 find the $[H_3O^+]$ of the final solution*
*=> $C_1V_1 = C_2V_2$*
*=> $10^{-3.524}$ X 0.400 = $C_2$ X 0.600*
*=> $C_2 = 10^{-3.524}$ X 0.400/ 0.600 = 1.995 X $10^{-4}$ = $10^{0.300}$ X $10^{-4}$ = $10^{-3.700}$*

*Step 2 find pH*
*-$\log_{10}[10^{-3.700}]$ = 3.700*

6) Consider a solution that is made up by placing 0.512 g of calcium hydroxide ($Ca(OH)_2$) in a 250 mL volumetric flask and made to the mark with distilled water.




i. Calculate the molarity of the $Ca(OH)_2$ solution.
*Mol of $Ca(OH)_2$*
*=> 0.512 / 74.1 = 0.00691*

ii. Calculate the $[OH^-]$
*Step 1 the dissociation of $Ca(OH)_2$*
*$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$*
*Step 2 mol of $OH^-$ = 2 X 0.00691 = 0.0138 mol*
*Step 3 find $[OH^-]$*
*=> 0.0138 / 0.250 = 0.0553 M*

iii. Calculate the $[H_3O^+]$
$[H_3O^+]$ = $10^{-14}$ / 0.0553 = $10^{-14}$ /$10^{-1.26}$ = $10^{-12.74}$

iv. Calculate the pH of the solution.

*pH = -$\log_{10}$ [$10^{-12.74}$] = 12.74*