ACID-BASE CALCULATIONS

1. In each of the following solutions, determine: i) hydronium ion concentration; ii) hydroxide ion concentration; iii) pH; iv) pOH

a) 0.030 mol/dm^3 HCl $\rightarrow H^+ + Cl^-$ strong acid

$\therefore [H^+] = 0.030 \frac{mol}{L}$

i) hydronium ion concentration = 0.030 M   iii) pH = 1.5

ii) hydroxide ion concentration = 3.3×10^{-13} M   iv) pOH = 12

b) 0.125 mol/dm^3 HBr $\rightarrow H^+ + Br^-$ strong acid

$\therefore [H^+] = 0.125$ M

i) hydronium ion concentration = 0.125 M   iii) pH = 0.903

ii) hydroxide ion concentration = 8.00×10^{-14} M   iv) pOH = 13.1

c) 0.025 mol/dm^3 NaOH $NaOH \rightarrow Na^+ + OH^-$ strong base

$\therefore [OH^-] = 0.025$ M

i) hydronium ion concentration = 4.0×10^{-13} M   iii) pH = 12

ii) hydroxide ion concentration = 0.025 M   iv) pOH = 1.6

d) 0.050 mol/dm^3 $Ca(OH)_2$ $Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$ strong base

$\therefore [OH^-] = 2(0.050) = 0.10$ M

i) hydronium ion concentration = 1.0×10^{-13} M   iii) pH = 13

ii) hydroxide ion concentration = 0.10 M   iv) pOH = 1.0

e) 0.100 mol/dm^3 H_2CO_3 $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ weak acid

	H_2CO_3	H^+	HCO_3^-
init	0.100	0	0
Δ	−x	+x	+x
eq	0.100 − x	x	x

$$K_a = \frac{[H^+][HCO_3^-]}{[H_2CO_3]}$$

$$4.4 \times 10^{-7} = \frac{(x)(x)}{(0.100 - x)}$$

$$4.4 \times 10^{-7} = \frac{x^2}{0.100}$$

$$x^2 = 4.4 \times 10^{-8}$$

$$x = 2.1 \times 10^{-4}$$

i) hydronium ion concentration 2.1×10^{-4} M   iii) pH 3.7

ii) hydroxide ion concentration 4.8×10^{-11} M   iv) pOH 10

f) 0.250 mol/dm³ HCN – weak acid

$$HCN \rightleftharpoons H^+ + CN^-$$

	HCN	H^+	CN^-
init	0.250	0	0
Δ	−x	+x	+x
eq	0.250 − x	x	x

$$K_a = \frac{[H^+][CN^-]}{[HCN]}$$

$$6.2 \times 10^{-10} = \frac{(x)(x)}{(0.250 - x)}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.250}$$

$$1.55 \times 10^{-10} = x^2$$

$$x = 1.2 \times 10^{-5}$$

i) hydronium ion concentration 1.2×10^{-5} M iii) pH 4.9

ii) hydroxide ion concentration 8.0×10^{-10} M iv) pOH 9.1

g) 0.125 mol/dm³ HNO_2 – weak acid

$$HNO_2 \rightleftharpoons H^+ + NO_2^-$$

HNO_2	H^+	NO_2^-
0.125	0	0
−x	+x	+x
0.125 − x	x	x

$$K_a = \frac{[H^+][NO_2^-]}{[HNO_2]}$$

$$7.2 \times 10^{-4} = \frac{(x)(x)}{(0.125 - x)}$$

$$7.2 \times 10^{-4} = \frac{x^2}{0.125}$$

$$9.0 \times 10^{-5} = x^2$$

$$x = 9.5 \times 10^{-3}$$

i) hydronium ion concentration 9.5×10^{-3} M iii) pH 2.0

ii) hydroxide ion concentration 1.1×10^{-12} M iv) pOH 12

h) 0.150 mol/dm³ HOOCCOOH – weak acid

$$H_2C_2O_4 \rightleftharpoons H^+ + HC_2O_4^-$$

$H_2C_2O_4$	H^+	$HC_2O_4^-$
0.150	0	0
−x	+x	+x
0.150 − x	x	x

$$K_a = \frac{[H^+][HC_2O_4^-]}{[H_2C_2O_4]}$$

$$5.4 \times 10^{-2} = \frac{(x)(x)}{(0.150 - x)}$$

$$(5.4 \times 10^{-2})(0.150 - x) = x^2$$

$$0.0081 - 0.054x = x^2$$

$$0 = x^2 + 0.054x - 0.0081$$

$$x = \frac{-0.054 \pm \sqrt{(0.054)^2 - (4)(1)(-0.0081)}}{2(1)}$$

$$x = \frac{-0.054 \pm 0.188}{2}$$

$$x = -0.12 \text{ or } \underline{0.067}$$

i) hydronium ion concentration 0.067 M iii) pH 1.2

ii) hydroxide ion concentration 1.5×10^{-13} M iv) pOH 13

i) 0.175 mol/dm^3 H_2SO_3 – weak acid

$$H_2SO_3 \rightleftharpoons H^+ + HSO_3^-$$

	H_2SO_3	H^+	HSO_3^-
	0.175	0	0
	$-x$	$+x$	$+x$
	$0.175-x$	x	x

$$K_a = \frac{[H^+][HSO_3^-]}{[H_2SO_3]}$$

$$1.3\times10^{-2} = \frac{(x)(x)}{(0.175-x)}$$

$$(0.013)(0.0175-x) = x^2$$

$$0.002275 - 0.013x = x^2$$

$$0 = x^2 + 0.013x - 0.002275$$

$$x = \frac{-0.013 \pm \sqrt{(0.013)^2 - (4)(1)(-0.002275)}}{2(1)}$$

$$x = \frac{-0.013 \pm \sqrt{0.009269}}{2}$$

$$x = -0.055 \text{ or } 0.042$$

i) hydronium ion concentration 0.042 M iii) pH 1.4

ii) hydroxide ion concentration 2.4×10^{-3} M iv) pOH 13

2. Calculate $[H_3O^+]$, $[ClO_4^-]$ and $[OH^-]$ in an aqueous solution that is 0.150 M in $HClO_4(aq)$. Is the solution acidic or basic?

$$HClO_4 \rightarrow H^+ + ClO_4^-$$

	$HClO_4$	H^+	ClO_4^-
	0.150	0	0
	0	0.150	0.150

$[H_3O^+] = 0.150$ M

$[ClO_4^-] = 0.150$ M

$[OH^-] = 6.67\times10^{-14}$ M

ACIDIC

3. Calculate , $[OH^-]$, $[K^+]$ and $[H_3O^+]$ in an aqueous solution that is 0.250 M in KOH(aq). Is the solution acidic or basic?

$$KOH \rightarrow K^+ + OH^-$$

	KOH	K^+	OH^-
	0.250	0	0
	0	0.250	0.250

$[OH^-] = 0.250$ M

$[K^+] = 0.250$ M

$[H^+] = 4.00\times10^{-14}$ M

BASIC

4. Compute $[Ca^{2+}]$, $[OH^-]$ and $[H_3O^+]$ for a solution that is prepared by dissolving 0.600 g of $Ca(OH)_2$ in enough water to make 0.500 dm^3 of solution.

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

	$Ca(OH)_2$	Ca^{2+}	$2OH^-$
	0.0162	0	0
	0	0.0162	2(0.0162)

$[OH^-] = 0.0324$ M

$[Ca^{2+}] = 0.0162$ M

$[H_3O^+] = 3.09\times10^{-13}$ M

$$mol = \frac{0.600\,g}{74.09268\,g/mol} = 0.00810 \text{ mol}$$

$$C = \frac{n}{V} = \frac{0.00810 \text{ mol}}{0.500 \text{ L}} = 0.0162 \text{ M}$$

5. The pH of human muscle fluid is 6.8. Compute the value of $[H_3O^+]$ in muscle fluid at 25°C.

$$[H_3O^+] = 10^{-6.8} = 1.6 \times 10^{-7}\ M$$

6. The pH of household ammonia is about 12. Calculate the value of $[OH^-]$ for the ammonia solution.

$$[OH^-] = 10^{-2} = 1 \times 10^{-2}\ M \text{ or } 0.01\ M$$

7. The pH of human blood is fairly constant at 7.4. Compute the hydronium ion and the hydroxide ion concentrations in human blood.

$$[H_3O^+] = 10^{-7.4} = 4.0 \times 10^{-8}\ M$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-8}} = 2.5 \times 10^{-7}\ M$$

8. The pH of the world's oceans is remarkably constant at 8.15. Compute the hydronium ion and hydroxide ion concentrations in the oceans.

$$[H_3O^+] = 10^{-8.15} = 7.08 \times 10^{-9}\ M$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.00 \times 10^{-14}}{7.08 \times 10^{-9}} = 1.41 \times 10^{-6}\ M$$

9. Normal rainwater has a pH of about 5.6, whereas what is called acid rain has been observed to have pH values as low as 3.0. Compute the ratio of $[H_3O^+]$ in acid rain to that in normal rain.

$$\frac{\text{acid}}{\text{normal}} = \frac{10^{-3.0}}{10^{-5.6}} = \frac{1.0 \times 10^{-3}}{2.5 \times 10^{-6}} = 398$$

Acid rain is approximately 400 times more acidic than normal rain

10. The value of K_a in water at 25°C for benzoic acid, $HC_7H_5O_2$ is 6.46×10^{-5}. Calculate the pH of an aqueous solution with a total concentration of $HC_7H_5O_2$ of 0.0200 M.

$HC_7H_5O_2$	$+ H_2O \rightleftharpoons$	H_3O^+	$+ C_7H_5O_2^-$
0.0200		0	0
$-x$		$+x$	$+x$
$0.0200 - x$		x	x

$$K_a = \frac{[H_3O^+][C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

$$6.46 \times 10^{-5} = \frac{(x)(x)}{(0.0200 - x)}$$

$$6.46 \times 10^{-5} = \frac{x^2}{0.0200}$$

$$1.292 \times 10^{-6} = x^2$$

$$x = 1.14 \times 10^{-3}$$

$$pH = -\log(1.14 \times 10^{-3}) = 2.94$$

11. Sulfamic acid, HO_3SNH_2, is used as a stabilizer for chlorine in swimming pools. Calculate the pH of a 0.0400 M sulfamic acid solution given that $K_a = 0.100$.

$$HO_3SNH_2 + H_2O \rightleftharpoons O_3SNH_2^- + H_3O^+$$

	HO_3SNH_2	$O_3SNH_2^-$	H_3O^+
	0.0400	0	0
	$-x$	$+x$	$+x$
	$0.0400-x$	x	x

$$K_a = \frac{[O_3SNH_2^-][H_3O^+]}{[HO_3SNH_2]}$$

$$0.100 = \frac{(x)(x)}{(0.0400-x)}$$

$$(0.100)(0.0400-x) = x^2$$

$$0.00400 - 0.100x = x^2$$

$$0 = x^2 + 0.100x - 0.00400$$

$$x = \frac{-0.1 \pm \sqrt{(0.1)^2 - 4(1)(-0.004)}}{2(1)}$$

$$x = \frac{-0.100 \pm \sqrt{0.0260}}{2}$$

$$x = -0.131 \text{ or } 0.0306$$

$$pH = -\log(0.0306) = 1.51$$

12. Calculate the pH of an aqueous solution of 0.150 M HClO.

$$HClO + H_2O \rightleftharpoons ClO^- + H_3O^+$$

	HClO	ClO^-	H_3O^+
	0.150	0	0
	$-x$	$+x$	$+x$
	$0.150-x$	x	x

$$K_a = \frac{[H_3O^+][ClO^-]}{[HClO]}$$

$$2.9 \times 10^{-8} = \frac{(x)(x)}{0.150-x}$$

$$2.9 \times 10^{-8} = \frac{x^2}{0.150}$$

$$4.35 \times 10^{-9} = x^2$$

$$x = 6.6 \times 10^{-5}$$

$$pH = -\log(6.6 \times 10^{-5}) = 4.2$$

13. Suppose two aspirin tablets (1 tablet = 324 mg) are dissolved in enough water to make 500.0 cm^3 of solution. Compute the pH of the resulting solution if $K_a = 2.75 \times 10^{-5}$ and the molar mass for aspirin (acetylsalicylic acid) is 180.15 g/mol.

2 tablets = 2(324) = 648 mg = 0.648 g

$$mol = \frac{0.648\ g}{180.15\ g/mol} = 0.00360 \text{ mol aspirin}$$

$$C = \frac{0.00360\ mol}{0.5000\ dm^3} = 0.00719\ M$$

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

	HA	H_3O^+	A^-
	0.00719	0	0
	$-x$	$+x$	$+x$
	$0.00719-x$	x	x

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$2.75 \times 10^{-5} = \frac{(x)(x)}{0.00719-x}$$

$$(2.75 \times 10^{-5})(0.00719-x) = x^2$$

$$1.98 \times 10^{-7} - 2.75 \times 10^{-5}x = x^2$$

$$0 = x^2 + 2.75 \times 10^{-5}x - 1.98 \times 10^{-7}$$

$$x = \frac{-2.75 \times 10^{-5} \pm \sqrt{(2.75 \times 10^{-5})^2 - (4)(1)(-1.98 \times 10^{-7})}}{2(1)}$$

$$x = \frac{-2.75 \times 10^{-5} \pm 8.90 \times 10^{-4}}{2}$$

$$x = -4.59 \times 10^{-4} \text{ or } 4.31 \times 10^{-4}$$

$$pH = -\log(4.31 \times 10^{-4}) = 3.37$$

14. What is the hydronium ion concentration in a solution of 0.100 M CH_3COOH and 2.00 M $NaCH_3COO$?

$NaCH_3COO \rightarrow Na^+ + CH_3COO^-$

$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$

CH_3COOH	H^+	CH_3COO^-
0.100	0	2.00
−x	+x	+x
0.100 − x	x	2.00 + x

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{(x)(2.00+x)}{(0.100-x)}$$

$$(0.100-x)(1.8\times10^{-5}) = (x)(2.00+x)$$

$$1.8\times10^{-6} - 1.8\times10^{-5}x = 2.00x + x^2$$

$$0 = x^2 + 2.000018x - 1.8\times10^{-6}$$

$$x = \frac{-2.000018 \pm \sqrt{(2.000018)^2 - 4(1)(-1.8\times10^{-6})}}{2(1)}$$

$$x = \frac{-2.000018 \pm \sqrt{4.0000792}}{2}$$

$$x = -2.00 \text{ or } 9.00\times10^{-7}$$

$$[H_3O^+] = 9.0\times10^{-7}\ M$$

15. What is the hydronium ion concentration in a solution of 0.0875 M HClO and 0.0550 M NaClO?

$NaClO \rightarrow Na^+ + ClO^-$

$HClO \rightleftharpoons H^+ + ClO^-$

$HClO$	H^+	ClO^-
0.0875	0	0.0550
−x	+x	+x
0.0875 − x	x	0.0550 + x

$$K_a = \frac{[H^+][ClO^-]}{[HClO]}$$

$$2.9\times10^{-8} = \frac{(x)(}{(0.0875-x)}$$

$$(2.9\times10^{-8})(0.0875-x) = (x)(0.0550+x)$$

$$2.5375\times10^{-9} - 2.9\times10^{-8}x = 0.0550x + x^2$$

$$0 = x^2 + 0.0550x - 2.5375\times10^{-9}$$

$$x = \frac{-0.0550 \pm \sqrt{(0.0550) - 4(1)(-2.5375\times10^{-9})}}{2(1)}$$

$$x = \frac{-0.0550 \pm \sqrt{0.003025}}{2}$$

$$x = \frac{-0.0550 \pm 0.0550}{2}$$

$$x = -0.055 \text{ or } 4.6\times10^{-8}$$

$$[H_3O^+] = 4.6\times10^{-8}\ \frac{mol}{L}$$

(15) ~~22~~. What is the hydronium ion concentration in a solution of 0.0875 M HClO and 0.0550 M NaClO?

$HClO$	+ H_2O ⇌	H_3O^+	+ ClO^-
0.0875		0	0.0550
$-x$		$+x$	$+x$
$0.0875 - x$		x	$0.0550 + x$

$$2.9 \times 10^{-8} = \frac{(x)(0.0550 - x)}{(0.0875 - x)}$$

$[H_3O^+] = 4.6 \times 10^{-8}$ mol/L

$$(2.9 \cdot 10^{-8})(0.0875 - x) = (x)(0.0550 + x)$$

$$2.5375 \cdot 10^{-9} - 2.9 \cdot 10^{-8} x = 0.0550x + x^2$$

$$0 = x^2 + 0.0550x - 2.5375 \cdot 10^{-9}$$

$$x = \frac{-0.0550 \pm \sqrt{(0.0550)^2 - 4(1)(-2.5375 \cdot 10^{-9})}}{2(1)}$$

$$x = \frac{-0.0550 \pm 0.0550}{2} \qquad x = -0.055 \text{ or } 4.6 \times 10^{-8}$$

(16) ~~23~~. Calculate the $[H_3O^+]$ in a solution that is 1.125 M CH_3COOH if enough $NaCH_3COO$ is added to make the solution 0.5005 M with respect to the CH_3COO^-?

CH_3COOH	+ H_2O ⇌	CH_3COO^-	+ H_3O^+
1.125		y	0
$-x$		$+x$	$+x$
$1.125 - x$		0.5005	x

$$1.8 \times 10^{-5} = \frac{(0.5005)(x)}{1.125 - x}$$

$$(1.8 \cdot 10^{-5})(1.125 - x) = 0.5005x$$

$[H_3O^+] = 4.0 \times 10^{-5}$ M

$$2.025 \cdot 10^{-5} - 1.8 \cdot 10^{-5} x = 0.5005x$$

$$2.025 \cdot 10^{-5} = 0.500482x$$

$$x = 4.046 \times 10^{-5}$$

(17) ~~30~~. A solution was made up to be 0.100 M in chloroacetic acid ($HC_2H_2O_2Cl$) and also 0.00200 M in sodium chloroacetate ($NaC_2H_2O_2Cl$). The K_a for chloroacetic acid is 1.36×10^{-3}. Determine the pH of the solution.

$HC_2H_2O_2Cl$	+ H_2O ⇌	H_3O^+	+ $C_2H_2O_2Cl^-$
0.100		0	0.00200
$-x$		$+x$	$+x$
$0.100 - x$		x	$0.00200 + x$

$$1.36 \cdot 10^{-3} = \frac{(x)(0.00200 + x)}{(0.100 - x)}$$

$$(0.00136)(0.100 - x) = 0.00200x + x^2$$

$$0.000136 - 0.00136x = 0.00200x + x^2$$

$$0 = x^2 + 0.00336x - 0.000136$$

$$x = \frac{-0.00336 \pm \sqrt{(0.00336)^2 - (4)(1)(-0.000136)}}{(2)(1)}$$

$$x = \frac{-0.00336 \pm 0.02356}{2} \qquad x = -0.0135 \text{ or } 0.0101$$

$$pH = -\log(0.0101) = \boxed{2.00}$$

(18) ~~31~~. Find the pH of a liter of solution, in which is dissolved 0.0800 mol of $HC_2H_3O_2$ and 0.100 mol of $NaC_2H_3O_2$.

$HC_2H_3O_2$	+ H_2O ⇌	H_3O^+	+ $C_2H_3O_2^-$
0.0800		0	0.100
$-x$		$+x$	$+x$
$0.0800 - x$		x	$0.100 + x$

$$1.8 \cdot 10^{-5} = \frac{(x)(0.100 + x)}{(0.0800 - x)}$$

$$(0.0800 - x)(1.8 \cdot 10^{-5}) = 0.100x + x^2$$

$$1.44 \cdot 10^{-6} - 1.8 \cdot 10^{-5}x = 0.100x + x^2$$

$$0 = x^2 + 0.100018x - 1.44 \cdot 10^{-6}$$

$$x = \frac{-.100018 \pm \sqrt{(0.100018)^2 - 4(1)(1.44 \cdot 10^{-6})}}{2(1)}$$

$$x = \frac{-0.100018 \pm 0.100047}{2}$$

$$x = -0.10 \text{ or } 1.4 \cdot 10^{-5}$$

$$pH = -\log(1.4 \cdot 10^{-5}) = \boxed{4.8}$$

(19)

~~22~~. 25.0 mL of a 0.100 M HCOOH is added to 35.0 mL of 0.0500 M HCOONa. Determine the pH.

$HCOOH + H_2O \rightleftharpoons$	H_3O^+	+	$HCOO^-$
0.0417	0		0.0292
$-x$	$+x$		$+x$
$0.0417-x$	x		$0.0292+x$

$C_2 = \frac{C_1V_1}{V_2}$ $[HCOOH] = \frac{(0.100)(25.0)}{60.0} = 0.0417\ M$

$[HCOONa] = \frac{(0.0500)(35.0)}{60.0} = 0.0292\ M$

$$1.8\cdot10^{-4} = \frac{(x)(0.0292+x)}{(0.0417-x)}$$

$$(0.0417-x)(1.8\cdot10^{-4}) = (x)(0.0292+x)$$

$$7.5\times10^{-6} - 1.8\times10^{-4}x = 0.0292x + x^2$$

$$0 = x^2 + 0.0293x - 7.5\cdot10^{-6}$$

$$x = \frac{-0.0293 \pm \sqrt{(0.0293)^2-(4)(1)(-7.5\cdot10^{-6})}}{(2)(1)}$$

$$x = \frac{-0.0293 \pm 0.0299}{2}$$

$x = -0.0296$ or 2.53×10^{-4}

$pH = -\log(2.5\times10^{-4}) = \boxed{3.6}$

(20)

~~23~~. 100.0 mL of 0.300 M CH_3COOH is mixed with 100.0 mL of 0.400 M CH_3COONa. Determine the pH.

$CH_3COOH + H_2O \rightleftharpoons$	CH_3COO^-	+	H_3O^+
0.150	0.200		0
$-x$	$+x$		$+x$
$0.150-x$	$0.200+x$		x

$[CH_3COOH] = \frac{(0.300)(100.0)}{(200.0)} = 0.150\ M$

$[CH_3COONa] = \frac{(0.400)(100.0)}{(200.0)} = 0.200\ M$

$$1.8\cdot10^{-5} = \frac{(0.200+x)(x)}{(0.150-x)}$$

$$(0.150-x)(1.8\cdot10^{-5}) = (0.200x + x^2)$$

$$2.7\times10^{-6} - 1.8\cdot10^{-5}x = 0.200x + x^2$$

$$0 = x^2 + 0.200\,018 - 2.7\times10^{-6}$$

$$x = \frac{-0.200\,018 \pm \sqrt{(0.200\,018)^2 - 4(1)(2.7\cdot10^{-6})}}{2(1)}$$

$$x = \frac{-0.200\,018 \pm 0.200\,045}{2}$$

$x = -0.20$ or 1.3×10^{-5}

$pH = -\log(1.3\times10^{-5}) = \boxed{4.9}$