ACIDS & BASES

BRONSTED-LOWRY ACIDS & BASES

Bronsted-Lowry acid = **proton donor (H^+ = proton)**

Bronsted-Lowry base = **proton acceptor (H^+ = proton)**

Bronsted-Lowry acid-base reaction = **reaction involving the transfer of a proton**

e.g. $\overset{\text{donates } H^+ \ \therefore \text{ acid}}{HCl} + \overset{OH^- \text{ accepts } H^+ \ \therefore \text{ base}}{NaOH} \rightarrow NaCl + H_2O$ (transfer: H^+)

$$HCl + NaOH \xrightarrow{} NaCl + H_2O$$

HCl: donates H^+ ∴ acid; NaOH: OH^- accepts H^+ ∴ base

e.g.

$$HNO_3 + NH_3 \rightarrow NH_4^+ + NO_3^-$$

HNO_3: donates H^+ ∴ acid; NH_3: NH_3 accepts H^+ ∴ base

TASK 1 – Bronsted-Lowry acids & bases

Identify the Bronsted-Lowry acid and base in each of the following reactions.

	acid	base
i) $H_2O + NH_3 \rightarrow OH^- + NH_4^+$		
ii) $H_2O + HCl \rightarrow H_3O^+ + Cl^-$		
iii) $KOH + HCOOH \rightarrow HCOOK + H_2O$		
iv) $CH_3COOH + HCl \rightarrow CH_3COOH_2^+ + Cl^-$		
v) $NH_3 + HCl \rightarrow NH_4Cl$		
vi) $HCO_3^- + OH^- \rightarrow CO_3^{2-} + H_2O$		
vii) $HCO_3^- + H^+ \rightarrow CO_2 + H_2O$		
viii) $H_2SO_4 + HNO_3 \rightarrow HSO_4^- + H_2NO_3^+$		

pH OF STRONG ACIDS

Monoprotic acid = **acid that releases one H^+ ion per molecule**

e.g. HCl (hydrochloric acid), HNO_3 (nitric acid), CH_3COOH (ethanoic acid)

Diprotic acid = **acid that releases two H^+ ions per molecule**

e.g. H_2SO_4 (sulfuric acid), $H_2C_2O_4$ (ethanedioic acid)

Moles of acid	Moles of H^+
3 moles of HNO_3	
2 moles of HCl	
4 moles of H_2SO_4	
0.3 moles of HNO_3	
0.3 moles of H_2SO_4	

Moles of acid	Moles of H^+
0.1 moles of H_2SO_4	
0.2 moles of HCl	
0.08 moles of HNO_3	
0.08 moles of H_2SO_4	
0.35 moles of HCl	

Definition of pH $\boxed{\mathbf{pH = -\log[H^+]}}$ Useful rearrangement $\boxed{\mathbf{[H^+] = 10^{-pH}}}$

ALWAYS give pH to 2 DECIMAL PLACES

$[H^+]$	0.00100		1.50		2.5×10^{-4}		4.5×10^{-12}		
pH		2.75		3.30		13.70		1.85	-0.70

Calculating the pH of a strong acid

pH of 0.500 mol dm^{-3} HNO_3?

$[H^+]$ = 0.500
pH = -log 0.500
pH = 0.30

pH of 0.200 mol dm^{-3} H_2SO_4?

$[H^+]$ = 2 x 0.200 = 0.400 (diprotic acid!)
pH = -log 0.400
pH = 0.40

[HCl] with pH 1.70?

$[H^+] = 10^{-1.70}$ = 0.0200
[HCl] = 0.0200 mol dm^{-3}

$[H_2SO_4]$ with pH 1.30?

$[H^+] = 10^{-1.30}$ = 0.0501
$[H_2SO_4]$ = 0.0501 / 2 = 0.251 mol dm^{-3}

Dilution of a strong acid

Calculate the pH of the solution formed when 100 cm^3 of water is added to 50 cm^3 of 0.100 mol dm^{-3} HNO_3.

$[H^+]$ in original HNO_3 solution = 0.100

$[H^+]$ in diluted solution = $0.100 \times \frac{50}{150} = 0.0333$

pH = -log 0.0333 = <u>1.48</u>

Calculate the pH of the solution formed when 250 cm^3 of 0.300 mol dm^{-3} H_2SO_4 is made up to 1000 cm^3 solution with water.

$[H^+]$ in original H_2SO_4 solution = 2 x 0.300 = 0.600

$[H^+]$ in diluted solution = $0.600 \times \frac{250}{1000} = 0.150$

pH = -log 0.150 = <u>0.82</u>

TASK 2 – pH of strong acids

1) Calculate the pH of the following solutions.

a) 0.2 mol dm^{-3} HCl

b) 0.05 mol dm^{-3} HNO_3

c) 0.04 mol dm^{-3} H_2SO_4

d) 2.00 mol dm^{-3} HNO_3

2) Calculate the concentration of the following acids.

a) HCl with pH 3.55

b) H_2SO_4 with pH 1.70

c) HNO_3 with pH 1.30

d) H_2SO_4 with pH -0.50

3) Calculate the pH of the solutions formed in the following way.

a) addition of 250 cm^3 of water to 50 cm^3 of 0.200 mol dm^{-3} HNO_3

b) addition of 25 cm^3 of water to 100 cm^3 of 0.100 mol dm^{-3} H_2SO_4

c) adding water to 100 cm^3 of 2.00 mol dm^{-3} H_2SO_4 to make 500 cm^3 of solution

d) adding water to 25 cm^3 of 1.50 mol dm^{-3} HCl to make 250 cm^3 of solution

4) Calculate the pH of the following solutions.

a) 10 g dm^{-3} HCl

b) 20 g dm^{-3} H_2SO_4

c) 50 g dm^{-3} HNO_3

d) 100 g dm^{-3} H_2SO_4

THE IONIC PRODUCT OF WATER, K_W

The ionic product of water (K_w) $H_2O \rightleftharpoons H^+ + OH^-$ ΔH = endothermic

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

$\therefore$ $K_c[H_2O] = [H^+][OH^-]$

As $[H_2O]$ is very much greater than $[H^+]$ and $[OH^-]$, then $[H_2O]$ is effectively a constant number

$\therefore$ $K_c[H_2O]$ = a constant = K_w

$$\boxed{K_w = [H^+][OH^-]}$$

The effect of temperature on the pH of water and the neutrality of water

As the temperature increases, the equilibrium moves right to lower the temperature

$\therefore$ $[H^+]$ and $[OH^-]$ increase

$\therefore$ K_w increases and $\therefore$ pH increases

However, the water is still neutral as $[H^+] = [OH^-]$ (and the definition of neutral is $[H^+] = [OH^-]$)

Calculating the pH of water

In pure water, $[H^+] = [OH^-]$

$\therefore$ $K_w = [H^+]^2$

$\therefore$ $[H^+] = \sqrt{K_w}$

Calculate the pH of water at 40°C when $K_w = 2.09 \times 10^{-14}$ mol^2 dm^{-6}.

$K_w = [H^+]^2$

$\therefore$ $[H^+] = \sqrt{K_w} = \sqrt{(2.09 \times 10^{-14})} = 1.45 \times 10^{-7}$

$\therefore$ pH = $-\log(1.45 \times 10^{-7})$ = **6.84**

Calculate the pH of water at 100°C when $K_w = 51.3 \times 10^{-14}$ mol^2 dm^{-6}.

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THE pH OF STRONG BASES

Monobasic base =	base that releases one OH^- ion	e.g.	NaOH, KOH, NH_3
Dibasic base =	base that releases two OH^- ions	e.g.	$Ba(OH)_2$, $Ca(OH)_2$

Calculating the pH of a strong base

pH of 0.200 mol dm^{-3} NaOH?

$[OH^-] = 0.200$

$[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.200} = 5 \times 10^{-14}$

$pH = -\log[H^+] = -\log(5 \times 10^{-14}) = \underline{\mathbf{13.30}}$

pH of 0.0500 mol dm^{-3} $Ba(OH)_2$?

$[OH^-] = 2 \times 0.0500 = 0.100$

$[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.100} = 1 \times 10^{-13}$

$pH = -\log[H^+] = -\log(1 \times 10^{-13}) = \underline{\mathbf{13.00}}$

[KOH] with pH 12.70?

$[H^+] = 10^{-pH} = 10^{-12.70} = 2.00 \times 10^{-13}$

$[OH^-] = \dfrac{K_w}{[H^+]} = \dfrac{10^{-14}}{2.00 \times 10^{-13}} = 0.05$

$[KOH] = \underline{\mathbf{0.05\ mol\ dm^{-3}}}$

$[Ba(OH)_2]$ with pH 13.30?

$[H^+] = 10^{-pH} = 10^{-13.30} = 5.01 \times 10^{-14}$

$[OH^-] = \dfrac{K_w}{[H^+]} = \dfrac{10^{-14}}{5.01 \times 10^{-14}} = 0.200$

$[Ba(OH)_2] = \underline{\mathbf{0.100\ mol\ dm^{-3}}}$

Dilution of a strong base

Calculate the pH of the solution formed when 50 cm^3 of water is added to 100 cm^3 of 0.200 mol dm^{-3} NaOH.

$[OH^-]$ in original NaOH solution = 0.200

$[OH^-]$ in diluted solution = $0.200 \times \dfrac{100}{150} = 0.1333$

$[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.1333} = 7.50 \times 10^{-14}$

$pH = -\log(7.50 \times 10^{-14}) = \underline{13.12}$

Calculate the pH of the solution formed when 50 cm^3 of 0.250 mol dm^{-3} KOH is made up to 250 cm^3 solution with water.

$[OH^-]$ in original KOH solution = 0.250

$[OH^-]$ in diluted solution = $0.250 \times \dfrac{50}{250} = 0.0500$

$[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.0500} = 2.00 \times 10^{-13}$

$pH = -\log(2.00 \times 10^{-13}) = \underline{12.70}$

TASK 3 – pH of strong bases

1) Calculate the pH of the following solutions.

   a) 0.15 mol dm^{-3} KOH
   b) 0.05 mol dm^{-3} NaOH
   c) 0.20 mol dm^{-3} $Ba(OH)_2$

2) Calculate the concentration of the following acids.

   a) NaOH with pH 14.30
   b) $Ba(OH)_2$ with pH 12.50
   c) KOH with pH 13.70

3) Calculate the pH of the solutions formed in the following way.

   a) addition of 100 cm^3 of water to 25 cm^3 of 0.100 mol dm^{-3} NaOH
   b) addition of 25 cm^3 of water to 100 cm^3 of 0.100 mol dm^{-3} $Ba(OH)_2$
   c) adding water to 100 cm^3 of 1.00 mol dm^{-3} KOH to make 1 dm^3 of solution

4) Calculate the pH of the following solutions.

   a) 20 g dm^{-3} NaOH
   b) 100 g dm^{-3} KOH
   c) 1 g dm^{-3} $Sr(OH)_2$

Reaction between a strong acid and a strong base

1) Calculate moles H^+
2) Calculate moles OH^-
3) Calculate moles XS H^+ or OH^-
4) Calculate XS $[H^+]$ or XS $[OH^-]$
5) Calculate pH

Calculate the pH of the solution formed when 50 cm^3 of 0.100 mol dm^{-3} H_2SO_4 is added to 25 cm^3 of 0.150 mol dm^{-3} NaOH.

mol H^+ = 2 x $^{50}/_{1000}$ x 0.100 = 0.0100

mol OH^- = $^{25}/_{1000}$ x 0.150 = 0.00375

$\therefore$ **XS mol H^+ = 0.0100 – 0.00375 = 0.00625**

$\therefore$ **XS $[H^+]$ = $\dfrac{0.00625}{^{75}/_{1000}}$ = 0.0833**

pH = -log (0.0833) = <u>1.08</u>

Calculate the pH of the solution formed when 25 cm^3 of 0.250 mol dm^{-3} H_2SO_4 is added to 100 cm^3 of 0.200 mol dm^{-3} NaOH.

mol H^+ = 2 x $^{25}/_{1000}$ x 0.250 = 0.0125

mol OH^- = $^{100}/_{1000}$ x 0.200 = 0.0200

∴ XS mol OH^- = 0.0200 – 0.0125 = 0.0075

∴ XS [OH^-] = $\frac{0.0075}{^{125}/_{1000}}$ = 0.0600

∴ XS [H^+] = $\frac{K_w}{[OH^-]}$ = $\frac{10^{-14}}{0.0600}$ = 1.67 x 10^{-13}

pH = -log (1.67 x 10^{-13}) = 12.78

TASK 4 – strong acid + strong base

1) Calculate the pH of the solution formed when 20 cm^3 of 0.100 mol dm^{-3} HNO_3 is added to 30 cm^3 of 0.050 mol dm^{-3} KOH.

2) Calculate the pH of the solution formed when 25 cm^3 of 0.150 mol dm^{-3} H_2SO_4 is added to 50 cm^3 of 0.100 mol dm^{-3} NaOH.

3) Calculate the pH of the solution formed when 100 cm^3 of 0.050 mol dm^{-3} HCl is added to 50 cm^3 of 0.500 mol dm^{-3} KOH.

4) Calculate the pH of the solution formed when 10 cm^3 of 1.00 mol dm^{-3} H_2SO_4 is added to 25 cm^3 of 1.00 mol dm^{-3} NaOH.

5) Calculate the pH of the solution formed when 50 cm^3 of 0.250 mol dm^{-3} HNO_3 is added to 50 cm^3 of 0.100 mol dm^{-3} $Ba(OH)_2$.

6) Calculate the pH **change** to 100 cm^3 of 0.200 mol dm^{-3} HCl solution in a flask if 50 cm^3 of 0.100 mol dm^{-3} NaOH is added.

7) Calculate the pH **change** to 50 cm^3 of 0.150 mol dm^{-3} KOH solution in a flask if 50 cm^3 of 0.100 mol dm^{-3} H_2SO_4 is added.

TASK 5 – a mixture!

1) Calculate the pH of the following solutions:
   a) 0.150 mol dm^{-3} $Ba(OH)_2$
   b) 0.200 mol dm^{-3} HNO_3
   c) 1.500 mol dm^{-3} H_2SO_4
   d) 0.0500 mol dm^{-3} NaOH

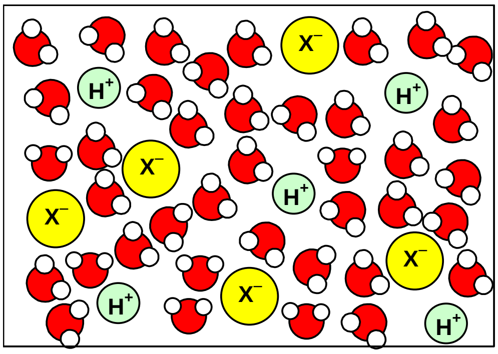
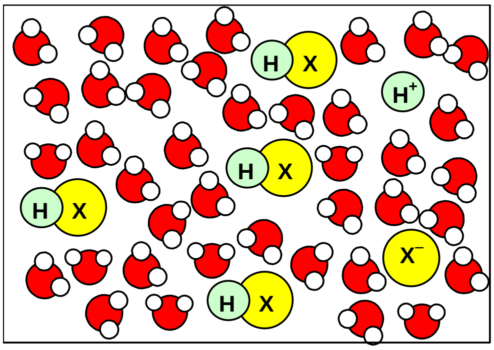
2) a) Calculate the pH of water at 50°C when K_w = 5.48 x 10^{-14} mol^2 dm^{-6}.
   b) is the water neutral? Explain your answer.

3) a) 20 cm^3 of 1.0 M H_2SO_4 with water added to make the volume up to 100 cm^3.
   b) 50 cm^3 of 0.05 M KOH with 200 cm^3 of water added.

4) a) Calculate the pH of the solution formed when 100 cm^3 of 0.100 mol dm^{-3} H_2SO_4 is added to 50 cm^3 of 0.500 mol dm^{-3} NaOH.
   b) Calculate the pH of the solution formed when 25 cm^3 of 0.250 mol dm^{-3} HCl is added to 15 cm^3 of 0.100 mol dm^{-3} KOH.

5) Calculate the pH of the solution formed when 3.5 g of impure sodium hydroxide (98.7 % purity) is dissolved in water and made up to 100 cm^3, and then 25 cm^3 of 0.35 mol dm^3 diprotic acid is added.

WEAK ACIDS

Strong acid	**Weak acid**
all the molecules break apart to form ions	only a small fraction of the molecules break apart to form ions
	
$HX \rightarrow H^+ + X^-$	$HX \rightleftharpoons H^+ + X^-$

Some common acids and bases

	Strong acids	Weak acids	Strong bases	Weak bases
Monoprotic / basic	HCl hydrochloric acid HNO_3 nitric acid	carboxylic acids (e.g. ethanoic acid)	NaOH sodium hydroxide KOH potassium hydroxide	NH_3 ammonia
Diprotic / basic	H_2SO_4 sulphuric acid		$Ba(OH)_2$ barium hydroxide	

The acid dissociation constant (K_a)

$HA \rightleftharpoons H^+ + A^-$

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \& \quad pK_a = -\log K_a \quad \& \quad K_a = 10^{-pKa}$$

These expressions hold for weak acids at all times

Note

- K_a – has units mol dm^{-3}
- K_a – the bigger the value, the stronger the acid
- pK_a – the bigger the value, the weaker the acid

In a solution of a weak acid in water, with nothing else added:

a) $[H^+] = [A^-]$

b) $[HA] \approx [HA]_{initial}$ (i.e. the concentration of HA at equilibrium is virtually the same as it was before any of it dissociated as so little dissociates, e.g. in a 0.100 mol dm^{-3} solution of HA, there is virtually 0.100 mol dm^{-3} of HA)

$$K_a = \frac{[H^+]^2}{[HA]}$$

This expression ONLY holds for weak acids in aqueous solution with nothing else added

Calculating the pH of a weak acid

Calculate the pH 0.100 mol dm^{-3} propanoic acid (pK_a = 4.87).

$$K_a = \frac{[H^+]^2}{[HA]}$$

$$[H^+]^2 = K_a\,[HA]$$

$$[H^+] = \sqrt{(K_a\,[HA])} = \sqrt{(K_a\,[HA])} = \sqrt{(10^{-4.87} \times 0.100])} = 1.16 \times 10^{-3}$$

$$pH = -\log(1.16 \times 10^{-3}) = \underline{2.94}$$

Calculate the concentration of a solution of methanoic acid with pH 4.02 (K_a = 1.35 x 10^{-5} mol dm^{-3}).

$$[H^+] = 10^{-4.02} = 9.55 \times 10^{-5}$$

$$[HA] = \frac{[H^+]^2}{K_a} = \frac{(9.55 \times 10^{-5})^2}{1.35 \times 10^{-5}} = 6.76 \times 10^{-4} \text{ mol dm}^{-3}$$

TASK 6 – pH of weak acids

1) Calculate the pH of the following weak acids:
   a) 0.150 mol dm^{-3} benenecarboxylic acid (pK_a = 4.20)
   b) 0.200 mol dm^{-3} butanoic acid (K_a = 1.51 x 10^{-5} mol dm^{-3})
   c) 1.00 mol dm^{-3} methanoic acid (K_a = 1.78 x 10^{-4} mol dm^{-3})

2) Calculate the concentration of the following weak acids.
   a) ethanoic acid with pH 4.53 (K_a = 1.74 x 10^{-5} mol dm^{-3})
   b) pentanoic acid with pH 3.56 (pK_a = 4.86)

3) a) Which is the stronger acid, ethanoic acid (pK_a = 4.76) or propanoic acid (pK_a = 4.87)?
   b) Which is the stronger acid, propanoic acid (1.35 x 10^{-5} mol dm^{3}) or propenoic acid (5.50 x 10^{-5} mol dm^{3})?

4) Calculate the K_a value for phenylethanoic acid given that a 0.100 mol dm^{-3} solution has a pH of 2.66.

Reaction between a weak acid and a strong base

When a weak acid reacts with a strong base, for every mole of OH^- added, one mole of HA is used up and one mole of A^- is formed.

e.g.	HA	+	OH^-	$\rightarrow$	A^-	+	H_2O
before reaction	3		1				
after reaction	2 left				2 made		

e.g.	HA	+	OH^-	$\rightarrow$	A^-	+	H_2O
before reaction	3		10				
after reaction			7 left		3 made		

TASK 7 – Reaction of weak acids

When the following weak acids react with strong bases
- the moles of HA left after reaction
- the moles of OH^- left after reaction
- the moles of A^- formed in the reaction

1) 4 moles of HA with 2.5 moles of NaOH
2) 6 moles of HA with 1.3 moles of $Ba(OH)_2$
3) 0.15 moles of HA with 0.25 moles of KOH
4) 0.30 moles of HA with 0.15 moles of NaOH
5) 100 cm^3 of 0.100 mol dm^{-3} HA with 50 cm^3 0.050 mol dm^{-3} NaOH
6) 25 cm^3 of 0.500 mol dm^{-3} HA with 40 cm^3 of 1.0 mol dm^{-3} KOH
7) 10 cm^3 of 0.100 mol dm^{-3} HA with 10 cm^3 of 0.080 mol dm^{-3} NaOH

Calculating the pH for the solution formed from reaction between a weak acid and a strong base

When a weak acid reacts with a strong base, for every mole of OH^- added, one mole of HA is used up and one mole of A^- is formed.

1) Calculate moles HA (it is still HA and not H^+ as it is a weak acid)
2) Calculate moles OH^-
3) Calculate moles XS HA or OH^-

If XS HA

4) Calculate moles HA left and A^- formed
5) Calculate [HA] leftover and $[A^-]$ formed
6) Use K_a to find $[H^+]$
7) Find pH

If XS OH⁻

4) Calculate $[OH^-]$
5) Use K_w to find $[H^+]$
6) Find pH

Note – if there is XS base, then in terms of working out the pH it is irrelevant whether it was a strong or weak acid as it has all reacted!

Calculate the pH of the solution formed when 30 cm^3 of 0.200 mol dm^{-3} ethanoic acid (pK_a = 4.76) is added to 100 cm^3 of 0.100 mol dm^{-3} NaOH.

mol HA = $^{30}/_{1000} \times 0.200 = 0.00600$

mol OH^- = $^{100}/_{1000} \times 0.100 = 0.0100$

OH^- is in XS

XS mol OH^- = 0.0100 – 0.00600 = 0.00400

$\therefore$ **XS $[OH^-]$ = $\dfrac{0.00400}{^{130}/_{1000}} = 0.0308$**

$\therefore$ **XS $[H^+]$ = $\dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.0308} = 3.25 \times 10^{-13}$**

pH = -log (3.25×10^{-13}) = <u>12.49</u>

Calculate the pH of the solution formed when 50 cm^3 of 0.500 mol dm^{-3} ethanoic acid (pK_a = 4.76) is added to 75 cm^3 of 0.200 mol dm^{-3} NaOH.

mol HA = $^{50}/_{1000}$ x 0.500 = 0.0250
mol OH^- = $^{75}/_{1000}$ x 0.200 = 0.0150
HA is in XS

	HA	+	OH^-	$\rightarrow$	A^-	+	H_2O
before reaction	**0.0250**		**0.0150**				
after reaction	**0.0100**		**-**		**0.0150**		

$\therefore$ **left over [HA]** $= \dfrac{0.0100}{^{125}/_{1000}} = 0.0800$

$\therefore$ **formed [A^-]** $= \dfrac{0.0150}{^{125}/_{1000}} = 0.120$

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

$$[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{10^{-4.76} \times 0.0800}{0.120} = 1.16 \times 10^{-5}$$

pH = -log (1.16 x 10^{-5}) = <u>4.94</u>

Half neutralisation of a weak acid

When half of the HA molecules have reacted with OH^-, [HA] = [A^-]. $\therefore$ K_a = [H^+] or pK_a = pH

Calculate the pH of the solution formed when 100 cm^3 of 0.200 mol dm^{-3} ethanoic acid (pK_a = 4.76) is added to 40 cm^3 of 0.250 mol dm^{-3} KOH.

mol HA = $^{100}/_{1000}$ x 0.200 = 0.0200
mol OH^- = $^{40}/_{1000}$ x 0.250 = 0.0100
HA is in XS

	HA	+	OH^-	$\rightarrow$	A^-	+	H_2O
before reaction	**0.0200**		**0.0100**				
after reaction	**0.0100**		**-**		**0.0100**		

half neutralisation and so [HA] = [A^-]

$\therefore$ **pH = pK_a = <u>4.76</u>**

TASK 8 – weak acid + strong base

1) Calculate the pH of the solution formed when 20 cm^3 of 0.100 mol dm^{-3} methanoic acid (K_a = 1.7 x 10^{-4} mol dm^{-3}) is added to 40 cm^3 of 0.080 mol dm^{-3} KOH.

2) Calculate the pH of the solution formed when 50 cm^3 of 0.500 mol dm^{-3} propanoic acid (pK_a = 4.87) is added to 100 cm^3 of 0.080 mol dm^{-3} KOH.

3) Calculate the pH of the solution formed when 50 cm^3 of 0.500 mol dm^{-3} ethanoic acid (pK_a = 4.76) is added to 50 cm^3 of 0.250 mol dm^{-3} KOH.

4) Calculate the pH of the solution formed when 50 cm^3 of 0.500 mol dm^{-3} chloroethanoic acid (pK_a = 2.86) is added to 25 cm^3 of 0.100 mol dm^{-3} $Ba(OH)_2$.

5) Calculate the pH of the solution formed when 50 cm^3 of 1.50 mol dm^{-3} dichloroethanoic acid (K_a = 0.0513 mol dm^{-3}) is added to 100 cm^3 of 2.00 mol dm^{-3} KOH.

6) Calculate the pH of the solution formed when 25 cm^3 of 1.00 mol dm^{-3} benzenecarboxylic acid (pK_a = 4.20) is added to 50 cm^3 of 0.0400 mol dm^{-3} NaOH.

TASK 9 – Another mixture

1) Calculate the pH of 0.100 mol dm^{-3} H_2SO_4.
2) Calculate the pH of 0.250 mol dm^{-3} methanoic acid (K_a = 1.70 x 10^{-4} mol dm^{-3})
3) Calculate the pH of 0.20 mol dm^{-3} $Sr(OH)_2$.
4) Calculate the pH of a mixture of 20 cm^3 of 0.500 mol dm^{-3} NaOH and 80 cm^3 of 0.200 mol dm^{-3} HNO_3.
5) Calculate the pH of the solution formed when 100 cm^3 of water is added to 25 cm^3 of 0.100 mol dm^{-3} NaOH.
6) Calculate the pH of a mixture of 25 cm^3 0.200 mol dm^{-3} ethanoic acid (pK_a = 4.76) and 25 cm^3 0.100 mol dm^{-3} NaOH.
7) Calculate the pH of a mixture of 100 cm^3 0.100 mol dm^{-3} ethanoic acid (pK_a = 4.76) and 50 cm^3 0.150 mol dm^{-3} NaOH.
8) Calculate the pH of a mixture of 50 cm^3 0.200 mol dm^{-3} propanoic acid (pK_a = 4.87) and 25 cm^3 1.00 mol dm^{-3} KOH.

pH CURVES & INDICATORS

- Indicators are weak acids where HA and A^- are different colours. $HA \rightleftharpoons H^+ + A^-$
- At low pH, HA is the main species present. At high pH, A^- is the main species present.
- The pH at which the colour changes varies from one indicator to another.
- Note that universal indicator is a mixture of indicators and so shows many colours at different pHs.

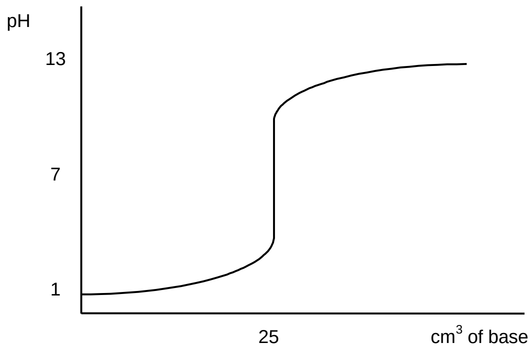
indicator	colour of HA	pH range of colour change	colour of A^-
methyl orange	red	3.2 - 4.4	yellow
phenolphthalein	colourless	8.2 - 10.0	pink

- In a titration, the pH changes rapidly at the end point as the last drop of acid/alkali is added. For an indicator to change colour at this moment where the moles of acid = moles of base, the indicator must change colour within the range of the rapid change in pH at the end point.

pH curves for monoprotic acids (e.g. HCl, $HNO_3$with monoprotic bases

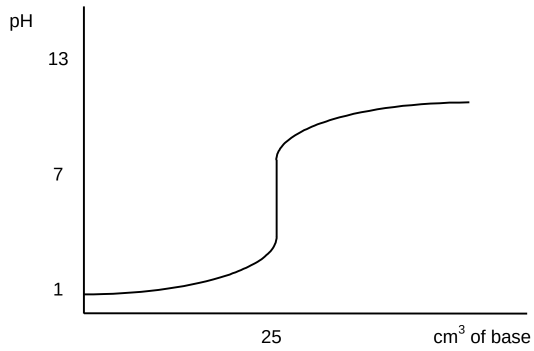
The curves below show the pH as 0.100 mol dm^{-3} base is added to 25.0 cm^3 of 0.100 mol dm^{-3} acid:

a) strong acid - strong base



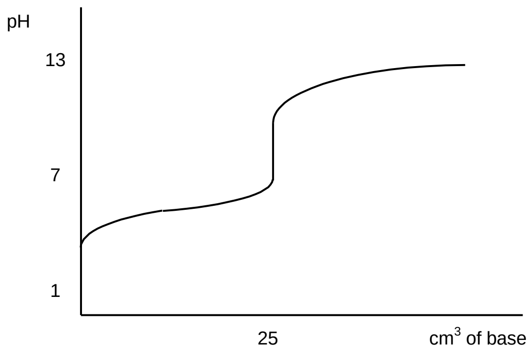
suitable indicators:

b) strong acid - weak base



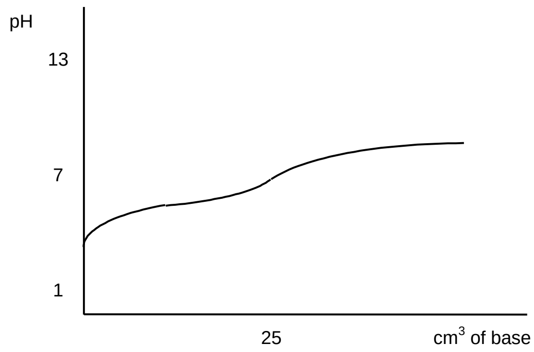
suitable indicators:

c) weak acid - strong base



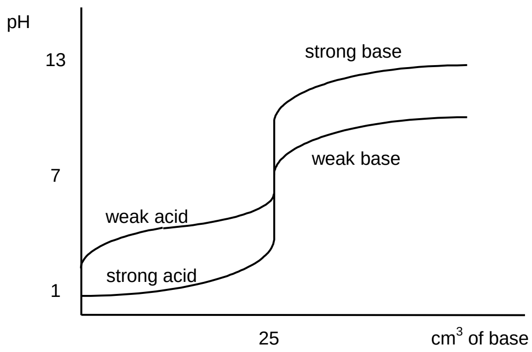
suitable indicators:

d) weak acid - weak base



suitable indicators:

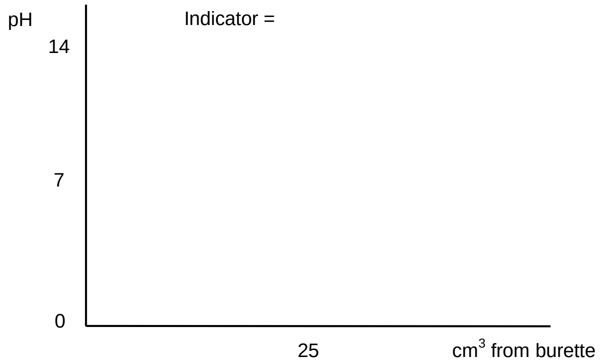
Summary:



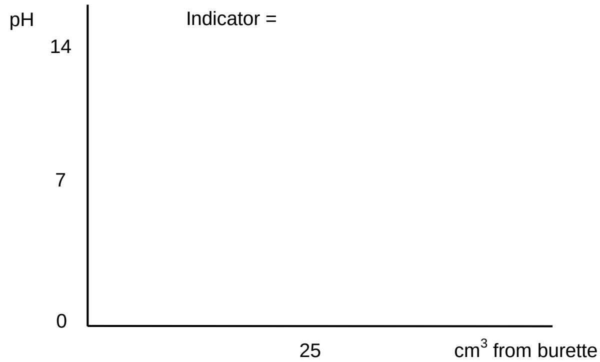
TASK 10 – sketching pH curves

Sketch each of the following pH curves on the grids shown, and name a suitable indicator.

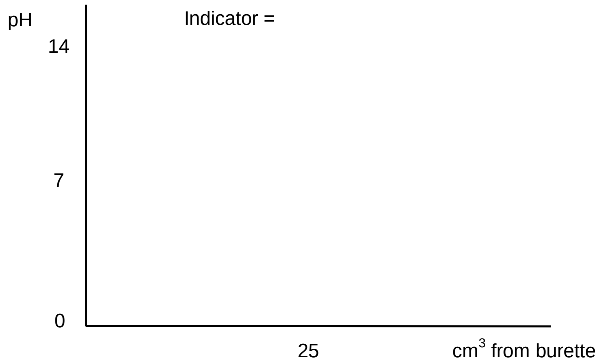
(1) Flask 25 cm^3 0.10 mol dm^{-3} HNO_3
Burette 50 cm^3 0.20 mol dm^{-3} NaOH



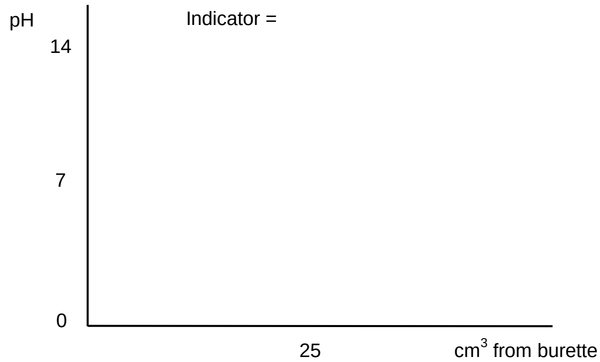
(2) Flask 20 cm^3 0.10 mol dm^{-3} NaOH
Burette 50 cm^3 0.10 mol dm^{-3} HCl



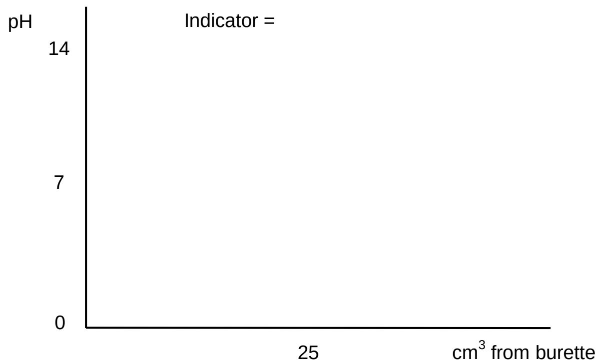
(3) Flask 10 cm^3 0.20 mol dm^{-3} HNO_3
Burette 50 cm^3 0.05 mol dm^{-3} NaOH



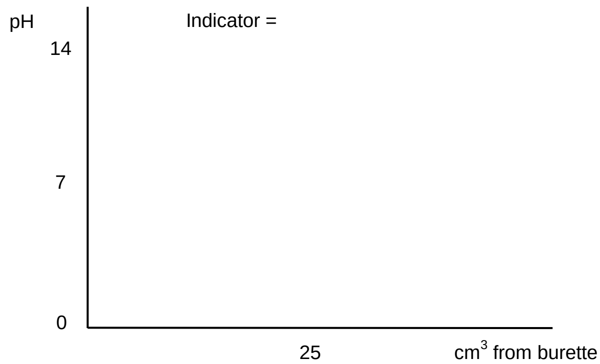
(4) Flask 30 cm^3 1.00 mol dm^{-3} NH_3
Burette 50 cm^3 1.00 mol dm^{-3} HCl



(5) Flask 20 cm^3 0.20 mol dm^{-3} CH_3COOH
Burette 50 cm^3 0.05 mol dm^{-3} NaOH



(6) Flask 50 cm^3 0.500 mol dm^{-3} NH_3
Burette 50 cm^3 1.00 mol dm^{-3} methanoic acid



Titration calculations

Remember these ionic equations

$H^+ + OH^- \rightarrow H_2O$

$2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$

$H^+ + HCO_3^- \rightarrow H_2O + CO_2$

$H^+ + NH_3 \rightarrow NH_4^+$

TASK 11 – Titration calculations

1) 25.0 cm^3 of a solution of sodium hydroxide required 18.8 cm^3 of 0.0500 mol dm^{-3} H_2SO_4.

$$H_2SO_4 + 2\,NaOH \rightarrow Na_2SO_4 + 2\,H_2O$$

a) Find the concentration of the sodium hydroxide solution in mol dm^{-3}.

b) Find the concentration of the sodium hydroxide solution in g dm^{-3}.

2) 25.0 cm^3 of arsenic acid, H_3AsO_4, required 37.5 cm^3 of 0.100 mol dm^{-3} sodium hydroxide for neutralisation.

$$3\,NaOH(aq) + H_3AsO_4(aq) \rightarrow Na_3AsO_4(aq) + 3\,H_2O(l)$$

a) Find the concentration of the acid in mol dm^{-3}.

b) Find the concentration of the acid in g dm^{-3}.

3) A 250 cm^3 solution of NaOH was prepared. 25.0 cm^3 of this solution required 28.2 cm^3 of 0.100 mol dm^{-3} HCl for neutralisation. Calculate what mass of NaOH was dissolved to make up the original 250 cm^3 solution.

4) 3.88 g of a monoprotic acid was dissolved in water and the solution made up to 250 cm^3. 25.0 cm^3 of this solution was titrated with 0.095 mol dm^{-3} NaOH solution, requiring 46.5 cm^3. Calculate the relative molecular mass of the acid.

5) A 1.575 g sample of ethanedioic acid crystals, $H_2C_2O_4.nH_2O$, was dissolved in water and made up to 250 cm^3. One mole of the acid reacts with two moles of NaOH. In a titration, 25.0 cm^3 of this solution of acid reacted with exactly 15.6 cm^3 of 0.160 mol dm^{-3} NaOH. Calculate the value of n.

6) A solution of a metal carbonate, M_2CO_3, was prepared by dissolving 7.46 g of the anhydrous solid in water to give 1000 cm^3 of solution. 25.0 cm^3 of this solution reacted with 27.0 cm^3 of 0.100 mol dm^{-3} hydrochloric acid. Calculate the relative formula mass of M_2CO_3 and hence the relative atomic mass of the metal M.

7) A 1.00 g sample of limestone is allowed to react with 100 cm^3 of 0.200 mol dm^{-3} HCl. The excess acid required 24.8 cm^3 of 0.100 mol dm^{-3} NaOH solution. Calculate the percentage of calcium carbonate in the limestone.

8) An impure sample of barium hydroxide of mass 1.6524 g was allowed to react with 100 cm^3 of 0.200 mol dm^{-3} hydrochloric acid. When the excess acid was titrated against sodium hydroxide, 10.9 cm^3 of sodium hydroxide solution was required. 25.0 cm^3 of the sodium hydroxide required 28.5 cm^3 of the hydrochloric acid in a separate titration. Calculate the percentage purity of the sample of barium hydroxide.

BUFFER SOLUTIONS

What is a buffer solution?

- Buffer solution = solution that resists changes in pH when small amounts of acid or alkali are added.
- Note – the pH does change, just not by much!)
- Acidic buffer solutions have a pH less than 7.
- Basic buffer solutions have a pH less than 7.

Examples of buffer solutions

- Acidic buffer solutions are made from a mixture of a weak acid and one of its salts (i.e. HA and A^-) (e.g. ethanoic acid & sodium ethanoate).
- An acidic buffer solution can also be made by mixing an excess of a weak acid with a strong alkali as it results in a mixture of HA and A^-.
- Basic buffer solutions are made from a mixture of a weak alkali and one of its salts (e.g. ammonia & ammonium chloride).
- A basic buffer solution can also be made by mixing an excess of a weak alkali with a strong acid

How acidic buffer solutions work

- This equilibrium exists in an acidic buffer solution: $HA \rightleftharpoons H^+ + A^-$
- There is a lot of HA and a lot of A^-
- If a small amount of H^+ is added: the equilibrium moves left to remove the H^+ added; as the added H^+ is removed, the pH remains roughly constant.
- If a small amount of OH^- is added: the OH^- reacts with and removes some H^+; the equilibrium moves right to replace the OH^- removed; as the removed H^+ is replaced, the pH remains roughly constant.

How basic buffer solutions work

- This equilibrium exists in a basic buffer solution: e.g. $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
- There is a lot of NH_3 and a lot of NH_4^+
- If a small amount of OH^- is added: the equilibrium moves left to remove the OH^- added; as the added OH^- is removed, the pH remains roughly constant.
- If a small amount of H^+ is added: the H^+ reacts with and removes some OH^-; the equilibrium moves right to replace the OH^- removed; as the removed OH^- is replaced, the pH remains roughly constant.

Calculating the pH of acidic buffer solutions

A buffer solution was made by adding 2.05 g of sodium ethanoate to 0.500 dm^3 of 0.01 mol dm^{-3} ethanoic acid. Calculate the pH of this solution (K_a for ethanoic acid = 1.74 x 10^{-5} mol dm^{-3}).

M_r CH_3COONa = 82.1

mol CH_3COONa = 2.05 / 82.1 = 0.0250

mol CH_3COO^- = 2.05 / 82.1 = 0.0250

$$[A^-] = \frac{0.0250}{0.500} = 0.0500$$

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

$$[H^+] = \frac{K_a[HA]}{[A^-]} = \frac{1.74 \times 10^{-5} \times 0.01}{0.050} = 3.48 \times 10^{-6}$$

pH = -log (3.48 x 10^{-6}) = <u>5.46</u>

TASK 12 – Buffer solution calculations

1) Calculate the pH of the following buffer solutions made by mixing weak acids with their salts.

a) 50 cm^3 of 1.0 mol dm^{-3} methanoic acid (K_a = 1.78 x 10^{-4} mol dm^{-3}) mixed with 20 cm^3 of 1.0 mol dm^{-3} sodium methanoate.

b) 25 cm^3 of 0.100 mol dm^{-3} butanoic acid (pK_a = 4.82) mixed with 20 cm^3 of 0.100 mol dm^{-3} sodium butanoate.

c) 1.00 g of potassium ethanoate is dissolved in 50 cm^3 of 0.200 mol dm^{-3} ethanoic acid (K_a = 1.74 x 10^{-5} mol dm^{-3}).

2) Calculate the pH of the following buffer solutions made by mixing an excess of weak acids with strong bases.

a) 25 cm^3 of 0.5 mol dm^{-3} methanoic acid (K_a = 1.78 x 10^{-4} mol dm^{-3}) is mixed with 10 cm^3 of 1.0 mol dm^{-3} sodium hydroxide.

b) 100 cm^3 of 1.0 mol dm^{-3} ethanoic acid (K_a = 1.78 x 10^{-4} mol dm^{-3}) is mixed with 50 cm^3 of 0.8 mol dm^{-3} sodium hydroxide.

3) a) What mass of sodium methanoate should be dissolved in 250 cm^3 of 0.100 mol dm^{-3} methanoic acid to form a buffer solution with a pH of 5.20 (K_a for methanoic acid = 1.78 x 10^{-4} mol dm^{-3}).

b) What mass of sodium ethanoate should be dissolved in 25 cm^3 of 0.100 mol dm^{-3} ethanoic acid to form a buffer solution with a pH of 3.50 (K_a for ethanoic acid = 1.74 x 10^{-5} mol dm^{-3}).

4) a) 2 cm^3 of 0.10 mol dm^{-3} NaOH is added to 100 cm^3 of water. Calculate the **change** in pH of the water.

b) 2 cm^3 of 0.10 mol dm^{-3} NaOH is added to 100 cm^3 of a buffer solution containing 0.15 mol dm^{-3} ethanoic acid and 0.10 mol dm^{-3} sodium ethanoate (K_a ethanoic acid = 1.74 x 10^{-5} mol dm^{-3}). Calculate the **change** in pH of the buffer solution.

c) Explain why the pH of the buffer solution only changes slightly compared to water.

GENERAL WORK

TASK 13 – A mixture of pH calculations

1) Calculate the pH of 0.100 mol dm^{-3} H_2SO_4.

2) Calculate the pH of the solution formed when 200 cm^3 of water are added to 50 cm^3 of 0.500 mol dm^{-3} HCl.

3) Calculate the pH of 0.500 mol dm^{-3} NaOH (K_w = 1.00 x 10^{-14} mol^2 dm^{-6}).

4) Calculate the pH change when water is added to 25 cm^3 of 0.250 mol dm^{-3} NaOH to prepare 1.00 dm^3 of solution (K_w = 1.00 x 10^{-14} mol^2 dm^{-6}).

5) Calculate the pH of 0.100 mol dm^{-3} chloroethanoic acid given that K_a = 1.38 x 10^{-3} mol dm^{-3}.

6) Find the pH of 0.100 mol dm^{-3} benzenecarboxylic acid (K_a = 6.31 x 10^{-5} mol dm^{-3}) when it has been half neutralised by NaOH.

7) Calculate the pH of water at 50ºC given that K_w = 5.476 x 10^{-14} mol^2 dm^{-6} at 50ºC and state and explain whether the water is still neutral.

8) Find the pH of the buffer solution prepared by adding 1.00 g of sodium ethanoate to 250 cm^3 of 0.100 mol dm^{-3} ethanoic acid (K_a = 1.74 x 10^{-5} mol dm^{-3}).

9) Calculate the pH of the solution formed when 25 cm^3 of 0.100 mol dm^{-3} NaOH is added to 50 cm^3 of 0.250 mol dm^{-3} HNO_3.

10) Calculate the pH of the solution formed when 100 cm^3 of 0.100 mol dm^{-3} NaOH is added to 20 cm^3 of 0.150 mol dm^{-3} H_2SO_4.

11) Calculate the pH of the solution formed when 50 cm^3 of 0.100 mol dm^{-3} NaOH is added to 100 cm^3 of 0.300 mol dm^{-3} CH_3COOH (K_a = 1.74 x 10^{-5} mol dm^{-3}).

12) Calculate the pH of the solution formed when 50 cm^3 of 0.0500 mol dm^{-3} $Ba(OH)_2$ is added to 20 cm^3 of 0.100 mol dm^{-3} HCOOH (K_a = 1.78 x 10^{-4} mol dm^{-3}).

FULL WORKED SOLUTIONS are available to subscribers of www.chemsheets.co.uk.

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TASK 1 – Bronsted-Lowry acids & bases

1	Acid = H_2O, base = NH_3	2	Acid = HCl, base = H_2O	3	Acid = HCOOH, base = KOH
4	Acid = HCl, base = CH_3COOH	5	Acid = HCl, base = NH_3	6	Acid = HCO_3^-, base = OH^-
7	Acid = H^+, base = HCO_3^-	8	Acid = H_2SO_4, base = HNO_3		

TASK 2 – pH of strong acids

1	a	0.70	b	1.30	c	1.10	d	-0.30
2	a	2.82×10^{-4}	b	0.0100	c	0.0501	d	1.58 mol dm^{-3}
3	a	1.48	b	0.80	c	0.10	d	0.82
4	a	0.56	b	0.39	c	0.10	d	-0.31

TASK 3 – pH of strong bases

1	a	13.18	b	12.70	c	13.60
2	a	2.00	b	0.0158	c	0.501 mol dm^{-3}
3	a	12.30	b	13.20	c	13.00
4	a	13.70	b	14.25	c	12.22

TASK 4 – Strong acid + strong base

1	2.00	2	1.48	3	13.12	4	13.15	5	1.60

6 new pH = 1.00, increase by 0.30

7 new pH = 1.60, decrease by 11.58

TASK 5 – a mixture

1 a 13.48 b 0.70 c -0.48 d 12.70

2 a 6.63 b still neutral as $[H^+] = [OH^-]$

3 a 0.40 b 12.00

4 a 12.52 b 0.93

5 13.74

TASK 6 – pH of weak acids

1 a 2.51 b 2.76 c 1.87

2 a 5.00×10^{-5} b 5.50×10^{-3} mol dm^{-3}

3 a ethanoic acid b propanoic acid

4 4.79×10^{-5} mol dm^{-3}

TASK 7 – Reactions of weak acids

1) HA = 1.5, OH^- = 0, A^- = 2.5
2) HA = 3.4, OH^- = 0, A^- = 2.6
3) HA = 0, OH^- = 0.10, A^- = 0.15
4) HA = 0.15, OH^- = 0, A^- = 0.15
5) HA = 0.0075, OH^- = 0, A^- = 0.0025
6) HA = 0, OH^- = 0.0275, A^- = 0.0125
7) HA = 0.0002, OH^- = 0, A^- = 0.0008

TASK 8 – Weak acid + strong base

1	12.30	2	4.54	3	4.76	4	2.26	5	13.92	6	3.14

TASK 9 – Another mixture

1	0.70	2	2.19	3	13.60	4	1.22	5	12.30
6	4.76	7	5.24	8	13.30				

TASK 11 – Titration calculations

1	a 0.0752 mol dm^{-3}			b	3.01 g dm^{-3}						
2	a 0.050 mol dm^{-3}			b	7.10 g dm^{-3}						
3	0.113 g	4	87.8	5	2	6	K	7	87.7%	8	90.8%

TASK 12 – Buffer solution calculations

1	a	3.35	b	4.72	c	4.77
2	a	4.35	b	3.57		
3	a	47.9 g	b	0.0113 g		
4	a	pH = 11.29, change = 4.29			b	pH = 4.60, change = 0.02

TASK 13 – A mixture of pH calculations

1	0.70	2	1.00	3	13.70	4	1.60	5	1.93	6	4.20
7	6.63, neutral as $[H^+] = [OH^-]$			8	4.45	9	0.88	10	12.52	11	4.06
12	12.63										