ACIDS & BASES

CHEMSHEETS.co.uk

SECTION 1 – 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**

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

H^+

e.g. $HCl + NaOH \rightarrow NaCl + H_2O$

donates H^+ ∴ acid

OH^- accepts H^+ ∴ base

H^+

e.g. $HNO_3 + NH_3 \rightarrow NH_4^+ + NO_3^-$

donates H^+ ∴ acid

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) $H_2SO_4 + HNO_3 \rightarrow HSO_4^- + H_2NO_3^+$		

SECTION 2 – pH of strong acids

Number of protons released

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

Definition of pH **$pH = -\log [H^+]$**

Useful rearrangement **$[H^+] = 10^{-pH}$**

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

Examples – 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?

$[H^+] = 2 \times 0.200 = 0.400$ (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}

Examples – 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 = $0.100 \times \frac{\text{old volume}}{\text{new volume}} = 0.100 \times \frac{50}{150} = 0.0333$

pH = -log 0.0333 = 1.48

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 = $0.600 \times \frac{\text{old volume}}{\text{new volume}} = 0.600 \times \frac{250}{1000} = 0.150$

pH = -log 0.150 = 0.82

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

SECTION 3 – 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 oppose the increase in temperature

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

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

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, $[H^+] = [OH^-]$

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

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

e.g. 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**

SECTION 4 – The pH of strong bases

Examples – Calculating the pH of a strong base

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

$[OH^-] = 0.200$

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

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

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

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

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

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

[KOH] with pH 12.70?

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

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

[KOH] = **0.05 mol dm^{-3}**

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

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

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

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

Examples – 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 = $0.200 \times \frac{\text{old volume}}{\text{new volume}} = 0.200 \times \frac{100}{150} = 0.1333$

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

pH = $-\log(7.50 \times 10^{-14})$ = 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 = $0.250 \times \frac{\text{old volume}}{\text{new volume}} = 0.250 \times \frac{50}{250} = 0.0500$

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

pH = $-\log(2.00 \times 10^{-13})$ = 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} $Sr(OH)_2$

SECTION 5 – The pH of mixtures of strong acids and strong bases

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

Example – with excess H^+

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^+ = $2 \times {}^{50}/_{1000} \times 0.100 = 0.0100$

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

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

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

pH = -log (0.0833) = $\underline{1.08}$

Example – with excess OH^-

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^+ = $2 \times {}^{25}/_{1000} \times 0.250 = 0.0125$

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

$\therefore$ **XS mol OH^- = $0.0200 - 0.0125 = 0.0075$**

$\therefore$ **XS $[OH^-]$ = $\dfrac{0.0075}{{}^{125}/_{1000}} = 0.0600$**

$\therefore$ **$[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.0600} = 1.67 \times 10^{-13}$**

pH = -log (1.67×10^{-13}) = $\underline{12.78}$

TASK 4 – pH of mixtures of strong acids and strong bases

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 variety of pH calculations so far

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 $K_w = 5.48 \times 10^{-14}$ 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.

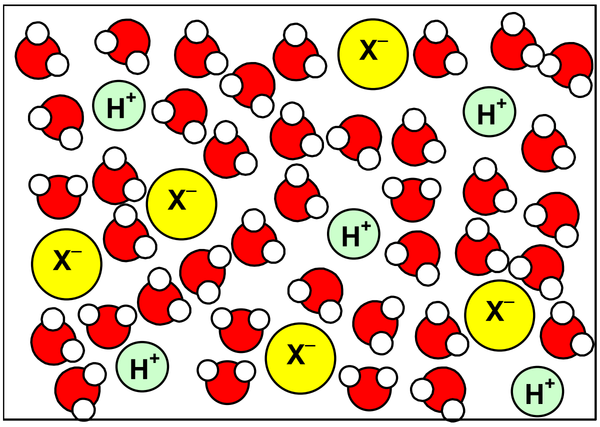
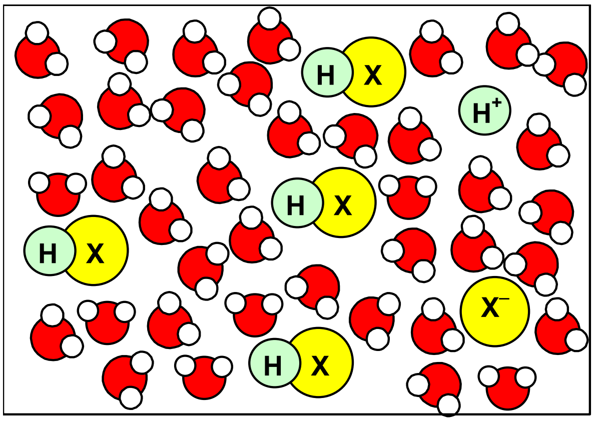
4 a) Calculate the pH of the solution formed when 100 cm^3 of 0.100 mol dm^{-3} H_2SO_4 is added to 50 cm^3 of 0.500 mol dm^{-3} NaOH.

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.

SECTION 6 – Weak acids

What is the difference between strong and 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
	
$HX \rightarrow H^+ + X^-$	$HX \rightleftharpoons H^+ + X^-$

Common acids and bases

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

The acid dissociation constant, K_a

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

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

$$pK_a = -\log K_a$$

$$K_a = 10^{-pK_a}$$

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 smaller the value, the stronger 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)

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

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

Example – finding the pH of a weak acid

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

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

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

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

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

Example – finding the concentration of a weak acid from pH

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 – The 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 (K_a = 1.35 x 10^{-5} mol dm^{3}) or propenoic acid (K_a = 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.

Reactions between weak acids and strong bases

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^-	→	A^-	+	H_2O
before reaction	3		1				
after reaction	2 left				1 made		

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

TASK 7 – Reactions of weak acids

When the following weak acids react with strong bases, calculate
- 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

Finding the pH in reactions between weak acids and strong bases

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	***If XS OH^-***	***If mol HA = OH^-***
4) Calculate moles HA left and A^- formed	4) Calculate $[OH^-]$	4) pH = pK_a of weak acid
5) Calculate [HA] leftover and $[A^-]$ formed	5) Use K_w to find $[H^+]$	
6) Use K_a to find $[H^+]$	6) Find pH	
7) 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!

Example – with excess OH^-

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^- = $0.0100 - 0.00600 = 0.00400$

∴ XS $[OH^-]$ = $\frac{0.00400}{^{130}/_{1000}} = 0.0308$

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

pH = $-\log(3.25 \times 10^{-13}) = \underline{12.49}$

Example – with excess HA

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 = $^{50}/_{1000} \times 0.500 = 0.0250$

mol OH^- = $^{75}/_{1000} \times 0.200 = 0.0150$ ∴ HA is in XS

	HA	+	OH^-	→	A^-	+	H_2O
before reaction	**0.0250**		**0.0150**				
after reaction	**0.0100**		**-**		**0.0150**		

∴ left over [HA] = $\frac{0.0100}{^{125}/_{1000}} = 0.0800$

∴ formed $[A^-]$ = $\frac{0.0150}{^{125}/_{1000}} = 0.120$

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

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

pH = $-\log(1.16 \times 10^{-5}) = \underline{4.94}$

Example – half neutralisation of a weak acid

When half of the HA molecules have reacted with OH^-, $[HA] = [A^-]$. ∴ $K_a = [H^+]$ 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 = $^{100}/_{1000} \times 0.200 = 0.0200$

mol OH^- = $^{40}/_{1000} \times 0.250 = 0.0100$ ∴ HA is in XS

	HA	+	OH^-	→	A^-	+	H_2O
before reaction	**0.0200**		**0.0100**				
after reaction	**0.0100**		**-**		**0.0100**		

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

∴ pH = pK_a = $\underline{4.76}$

TASK 8 – pH of mixtures of weak acids & strong bases

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.

Finding K_a for propanoic acid (Chemsheets A2 1082 or 1083)

TASK 9 – A variety of pH calculations so far

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.

SECTION 7 – Titration calculations

Remember these ionic equations which help a great deal in titration calculations.

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

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

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

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

TASK 10 – 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.

SECTION 8 – pH curves & indicators

What are indicators and how do they work?

- 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.

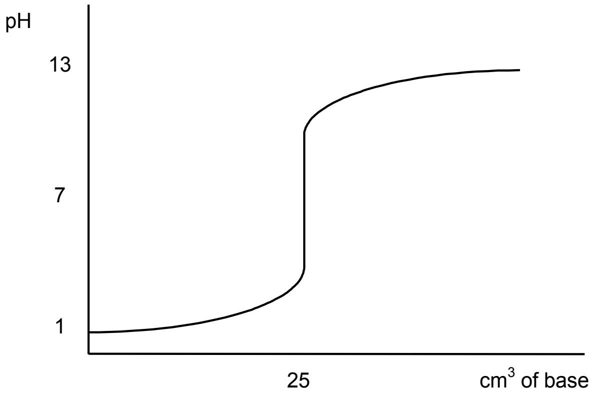
pH curves

- These show how the pH changes as an alkali is added to an acid (or vice versa).
- The equivalence point is when the moles of alkali added equals the moles of acid present – but the pH is not always 7 at the equivalence point.
- In most acid-alkali reactions, the pH curve shows a rapid change in pH around the equivalence point.
- The end point of a titration is when the indicator changes colour, and if a suitable indicator is used then the end point should coincide with the equivalence point.

Plotting pH curves (Chemsheets A2 1086)

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:

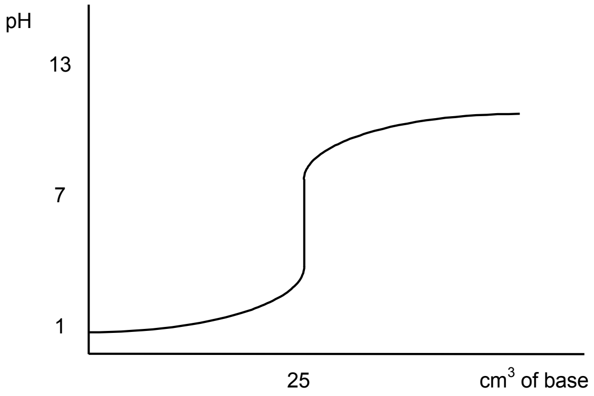
a) strong acid - strong base



pH at equivalence:

suitable indicators:

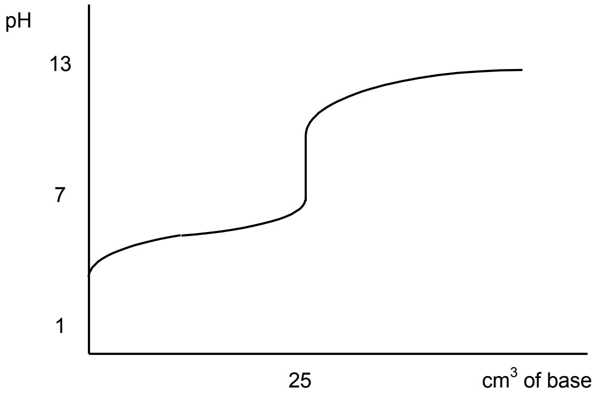
b) strong acid - weak base



pH at equivalence:

suitable indicators:

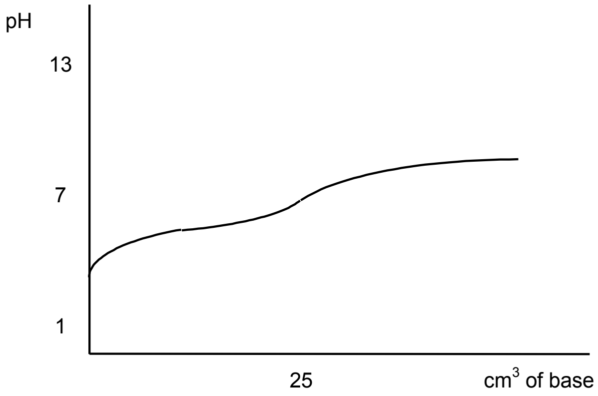
c) weak acid - strong base



pH at equivalence:

suitable indicators:

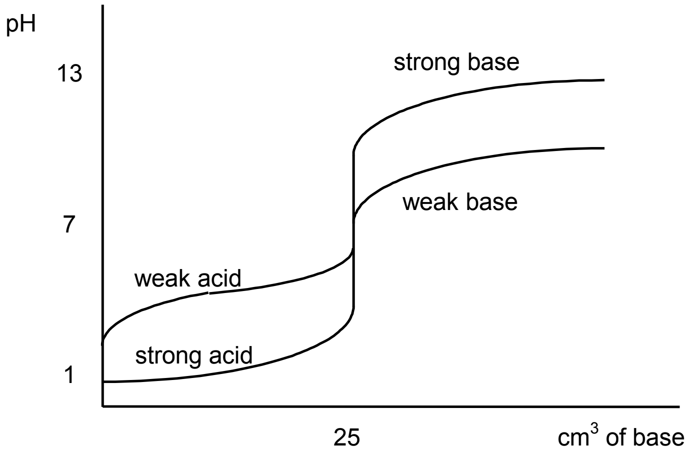
d) weak acid - weak base



pH at equivalence:

suitable indicators:

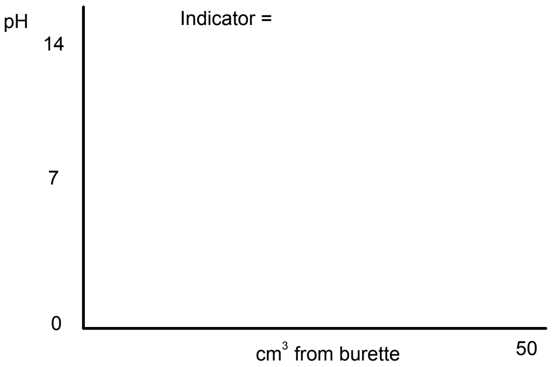
Summary:



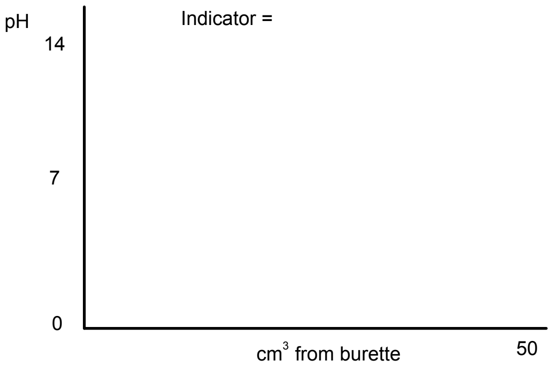
TASK 11 – Sketching pH curves

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

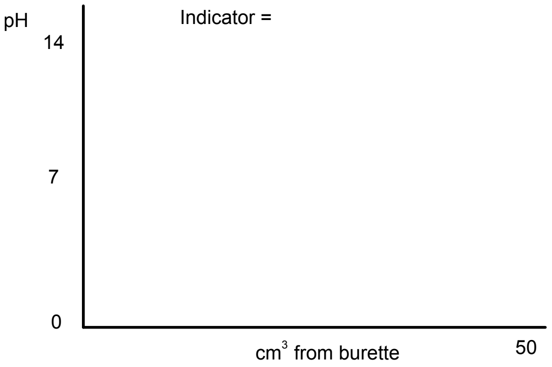
(1) Flask 25 cm^3 0.10 mol dm^{-3} HNO_3
Burette 50 cm^3 0.20 mol dm^{-3} NaOH



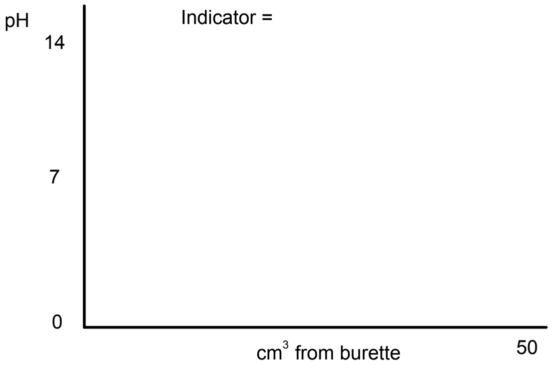
(2) Flask 20 cm^3 0.10 mol dm^{-3} NaOH
Burette 50 cm^3 0.10 mol dm^{-3} HCl



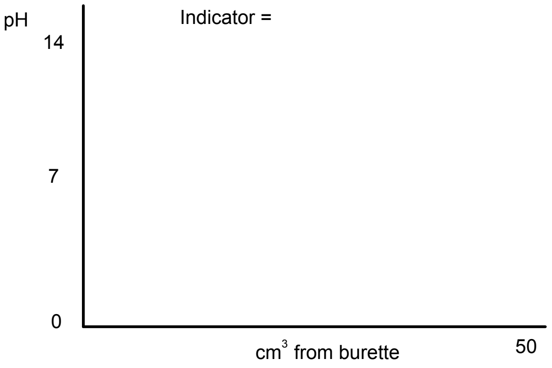
(3) Flask 10 cm^3 0.20 mol dm^{-3} HNO_3
Burette 50 cm^3 0.05 mol dm^{-3} NaOH



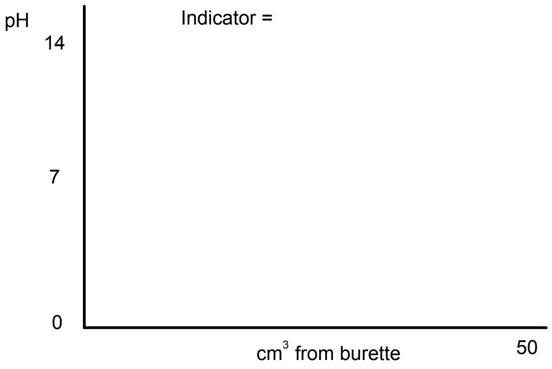
(4) Flask 30 cm^3 1.00 mol dm^{-3} NH_3
Burette 50 cm^3 1.00 mol dm^{-3} HCl



(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



SECTION 9 – 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 buffers

- 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^-.
- The key in an acidic buffer solution is that the [acid] and [salt] are much higher than $[H^+]$.

Basic buffers

- 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
- The key in a basic buffer solution is that the [base] and [salt] are much higher than $[OH^-]$.

Type of buffer	**Acidic buffer**		**Basic buffer**	
Components	Weak acid + one of its salts [acid] & [salt] >> $[H^+]$		Weak base + one of its salts [base] & [salt] >> $[OH^-]$	
Route 1	Mixture of weak acid and one of its salts	*e.g. ethanoic acid + sodium ethanoate*	Mixture of weak base and one of its salts	*e.g. ammonia + ammonium chloride*
Route 2	Mixture of an excess of weak acid and a strong base	*e.g.* ***excess*** *ethanoic acid + sodium hydroxide*	Mixture of an excess of weak base and a strong acid	*e.g.* ***excess*** *ammonia + hydrochloric acid*

How do buffer solutions work?

- The pH of an acidic buffer solution is found using the K_a expression:

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

- Therefore the pH of an acidic buffer depends on the ratio of [HA] to $[A^-]$ (i.e. the ratio of [acid] to [salt]).
- In a similar way, the pH of a basic buffer depends on the ratio of [base] to [salt]
- When small amounts of acid or alkali are added, the ratio remains roughly constant and so the pH hardly changes. If large amounts of acid or alkali are added, the ratio would change significantly and so the pH would change significantly.

	Acidic buffer **(e.g. $CH_3COOH + CH_3COO^-$)**	**Basic buffer** **(e.g. NH_3, NH_4^+)**
Add a little H^+	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ The added H^+ is removed by reaction with CH_3COO^- to form CH_3COOH. The $[CH_3COO^-]$ falls slightly and the $[CH_3COOH]$ rises slightly, but as $[CH_3COOH]$ & $[CH_3COO^-] >> [H^+]$, the ratio of $[CH_3COOH]/[CH_3COO^-]$ remains roughly constant.	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ The added H^+ is removed by reaction with OH^-, so some NH_3 reacts to replace the OH^-. The $[NH_3]$ falls slightly and the $[NH_4^+]$ rises slightly, but as $[NH_3]$ & $[NH_4^+] >> [OH^-]$, the ratio of $[NH_3]/[NH_4^+]$ remains roughly constant.
Add a little OH^-	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ The added OH^- reacts with H^+, and so some CH_3COOH breaks down to replace that H^+. The $[CH_3COO^-]$ rises slightly and the $[CH_3COOH]$ falls slightly, but as $[CH_3COOH]$ & $[CH_3COO^-] >> [H^+]$, the ratio of $[CH_3COOH]/[CH_3COO^-]$ remains roughly constant.	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ The added OH^- is removed by reaction with NH_4^+ to form NH_3. The $[NH_3]$ rises slightly and the $[NH_4^+]$ falls slightly, but as $[NH_3]$ & $[NH_4^+] >> [OH^-]$, the ratio of $[NH_3]/[NH_4^+]$ remains roughly constant.
Add water	The ratio of $[CH_3COOH]$ to $[CH_3COO^-]$ remains constant and so the pH remains constant.	The ratio of $[NH_3]$ to $[NH_4^+]$ remains constant and so the pH remains constant.

Making a buffer solution (Chemsheets A2 1089)

Calculating the pH of acidic buffers

1) 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 = 82.0

mol CH_3COONa = 2.05 / 82.0 = 0.0250

mol CH_3COO^- = 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}) = 5.46

2) a) A buffer solution was made by mixing 25.0 cm^3 of 1.00 mol dm^{-3} ethanoic acid with 25 cm^3 of 0.400 mol dm^{-3} sodium hydroxide. Find the pH of this buffer. (K_a for ethanoic acid = 1.74 x 10^{-5} mol dm^{-3}).

mol HA = $^{25}/_{1000}$ x 1.00 = 0.0250

mol OH^- = $^{25}/_{1000}$ x 0.400 = 0.0100 ∴ **HA is in XS**

	HA	+	OH^-	→	A^-	+	H_2O
before reaction	0.0250		0.0100				
after reaction	0.0150		-		0.0100		

∴ **left over [HA]** $= \frac{0.0150}{^{50}/_{1000}}$

∴ **formed [A$^-$]** $= \frac{0.0100}{^{50}/_{1000}}$

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

$$[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{1.74 \times 10^{-5} \times \frac{0.0150}{^{50}/_{1000}}}{\frac{0.0100}{^{50}/_{1000}}} = 2.61 \times 10^{-5}$$

pH = -log (2.61 x 10^{-5}) = 4.58

b) Calculate the new pH of the buffer if 0.2 cm^3 of 0.50 mol dm^{-3} sulfuric acid is added to the sample from part (a).

mol H^+ added = $2 \times {}^{0.2}/_{1000} \times 0.50$ = 0.0002

	A^- +	H^+ →	HA
before reaction	**0.0100**	**0.0002**	**0.0150**
after reaction	**0.0098**		**0.0152**

∴ **[HA] = $\dfrac{0.0152}{{}^{50.2}/_{1000}}$**

∴ **[A$^-$] = $\dfrac{0.0098}{{}^{50.2}/_{1000}}$**

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

$$[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{1.74 \times 10^{-5} \times \dfrac{0.0152}{{}^{50.2}/_{1000}}}{\dfrac{0.0098}{{}^{50.2}/_{1000}}} = 2.70 \times 10^{-5}$$

pH = -log (2.70×10^{-6}) = <u>4.57</u>

c) Calculate the new pH of the buffer if 1.0 cm^3 of 0.100 mol dm^{-3} sodium hydroxide is added to the sample from part (a).

mol OH^- added = ${}^{1.0}/_{1000} \times 0.100$ = 0.0001

	HA +	OH^- →	A^- +	H_2O
before reaction	**0.0150**	**0.0001**	**0.0100**	
after reaction	**0.0149**	**-**	**0.0101**	

∴ **[HA] = $\dfrac{0.0149}{{}^{51.0}/_{1000}}$**

∴ **[A$^-$] = $\dfrac{0.0101}{{}^{51.0}/_{1000}}$**

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

$$[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{1.74 \times 10^{-5} \times \dfrac{0.0149}{{}^{51.0}/_{1000}}}{\dfrac{0.0101}{{}^{51.0}/_{1000}}} = 2.57 \times 10^{-5}$$

pH = -log (2.57×10^{-5}) = <u>4.59</u>

TASK 12 – Buffer solution calculations

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

a) 50.0 cm^3 of 1.00 mol dm^{-3} methanoic acid ($K_a = 1.78 \times 10^{-4}$ mol dm^{-3}) mixed with 20.0 cm^3 of 1.00 mol dm^{-3} sodium methanoate.

b) 25.0 cm^3 of 0.100 mol dm^{-3} butanoic acid ($pK_a = 4.82$) mixed with 20.0 cm^3 of 0.100 mol dm^{-3} sodium butanoate.

c) 1.00 g of potassium ethanoate is dissolved in 50.0 cm^3 of 0.200 mol dm^{-3} ethanoic acid ($K_a = 1.74 \times 10^{-5}$ mol dm^{-3}).

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

a) 25.0 cm^3 of 0.500 mol dm^{-3} methanoic acid ($K_a = 1.78 \times 10^{-4}$ mol dm^{-3}) is mixed with 10.0 cm^3 of 1.00 mol dm^{-3} sodium hydroxide.

b) 100 cm^3 of 1.00 mol dm^{-3} ethanoic acid ($K_a = 1.78 \times 10^{-4}$ mol dm^{-3}) is mixed with 50.0 cm^3 of 0.800 mol dm^{-3} sodium hydroxide.

3 a) Calculate the pH of a buffer solution formed by mixing 20.0 cm^3 of 1.20 mol dm^{-3} methanoic acid ($K_a = 1.78 \times 10^{-4}$ mol dm^{-3}) with 20.0 cm^3 of 0.500 mol dm^{-3} sodium methanoate.

b) Calculate the pH of this buffer solution if 1.2 cm^3 of 0.40 mol dm^{-3} sodium hydroxide is added.

4 a) Calculate the pH of a buffer solution formed by mixing 50.0 cm^3 of 0.500 mol dm^{-3} ethanoic acid ($K_a = 1.74 \times 10^{-5}$ mol dm^{-3}) with 10.0 cm^3 of 0.800 mol dm^{-3} sodium hydroxide.

b) Calculate the pH of this buffer solution if 2.0 cm^3 of 0.20 mol dm^{-3} hydrochloric acid is added.

5 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 = 1.78×10^{-4} mol dm^{-3}).

b) What mass of sodium ethanoate should be dissolved in 25.0 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 = 1.74×10^{-5} mol dm^{-3}).

6 a) 2.00 cm^3 of 0.100 mol dm^{-3} NaOH is added to 100.0 cm^3 of water. Calculate the **change** in pH of the water.

b) 2.00 cm^3 of 0.100 mol dm^{-3} NaOH is added to 100 cm^3 of a buffer solution containing 0.150 mol dm^{-3} ethanoic acid and 0.100 mol dm^{-3} sodium ethanoate (K_a ethanoic acid = 1.74×10^{-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.

TASK 13 – One final lovely mixture of calculations just for fun

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.

Subscribe for many more exercises with answers.

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 – pH of mixtures of strong acids and strong bases

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 variety of pH calculations so far

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 – The pH of weak acids

1 a 2.51 b 2.76 c 1.87
2 a 5.01×10^{-5} b 5.50×10^{-3} mol dm^{-3}
3 a ethanoic acid b propenoic 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 – pH of mixtures of weak acids & strong bases

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

TASK 9 – A variety of pH calculations so far

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 10 – 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 1.13 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 3.37 b 3.40
4 a 4.43 b 4.40
5 a 47.9 g b 0.0113 g
6 a pH = 11.29, change = 4.29 b pH = 4.60, change = 0.02

TASK 13 – One final lovely mixture of calculations just for fun

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