pH CALCULATIONS SUMMARY

Key Equations	
Finding pH from $[H^+]$	$pH = -\log[H^+]$
Finding $[H^+]$ from pH	$[H^+] = 10^{-pH}$
Ionic product water to find $[H^+]$ from $[OH^-]$ or vice versa ($K_w = 10^{-14}$ mol^2 dm^{-6} at 298K)	$K_w = [H^+][OH^-]$
Equation for any solution containing a weak acid (HA) (including acidic buffer solutions)	$K_a = \frac{[H^+][A^-]}{[HA]}$
Finding pK_a from K_a	$pK_a = -\log K_a$
Finding K_a from pK_a	$K_a = 10^{-pK_a}$

Always give pH to 2 decimal places

Common acids and bases

	ACIDS		BASES	
	Strong	Weak	Strong	Weak
Monoprotic / basic	HCl hydrochloric acid HNO_3 nitric acid	carboxylic acids (e.g. ethanoic acid)	NaOH sodium hydroxide KOH potassium hydroxide	NH_3 ammonia amines (e.g. methylamine)
Diprotic / basic	H_2SO_4 sulfuric acid	dicarboxylic acids (e.g. ethanedioic acid)	$Ba(OH)_2$ barium hydroxide	diamines (e.g. 1,2-diaminoethane)

Reactions to consider	
Strong acid + strong base	H^+ from the acid reacts with OH^- from the base $H^+ + OH^- \rightarrow H_2O$
Weak acid + strong base	The weak acid HA reacts with OH^- from the base to form A^- $HA + OH^- \rightarrow A^- + H_2O$
Weak acid + salt	These are in equilibrium with each other $HA \rightleftharpoons H^+ + A^-$
H^+ added to mixture of weak acid and its salt (i.e. to an acidic buffer)	H^+ reacts with A^- to form more HA $H^+ + A^- \rightarrow HA$
OH^- added to mixture of weak acid and its salt (i.e. to an acidic buffer)	OH^- reacts with HA to form more A^- $HA + OH^- \rightarrow A^- + H_2O$

Common pH calculations		
Finding the pH of a strong acid from its concentration	• The strong acid is fully dissociated into H^+ ions Use acid concentration to get $[H^+]$ ↓ $pH = -\log[H^+]$	***Find the pH of 0.250 mol dm^{-3} H_2SO_4*** H_2SO_4 is a strong diprotic acid $[H^+] = 2 \times 0.250 = 0.500$ $pH = -\log[H^+] = -\log 0.500 = \underline{\mathbf{0.30}}$
Finding the concentration of a strong acid from its pH	• The strong acid is fully dissociated into H^+ ions $[H^+] = 10^{-pH}$ ↓ Use $[H^+]$ to find acid concentration	***Find the concentration of H_2SO_4 with pH 1.70*** $[H^+] = 10^{-pH} = 10^{-1.70} = 0.0200$ H_2SO_4 is a strong diprotic acid (so for every one H_2SO_4 there are two H^+ ions) $[H_2SO_4] = \frac{0.0200}{2} = \underline{\mathbf{0.0100\ mol\ dm^{-3}}}$
Finding the pH of a strong acid that is diluted	• The strong acid is fully dissociated into H^+ ions Use the dilution factor to get $[H^+]$ ↓ $pH = -\log[H^+]$	***100 cm^3 of water is added to 25 cm^3 of 0.500 H_2SO_4. Find the pH of the diluted acid.*** new $[H_2SO_4]$ after dilution = $old\ conc\ x\ \frac{old\ volume}{new\ volume}$ $= 0.500\ x\ \frac{25}{125} = 0.100$ H_2SO_4 is a strong diprotic acid $[H^+] = 2 \times 0.100 = 0.200$ $pH = -\log[H^+] = -\log 0.200 = \underline{\mathbf{0.70}}$
Finding the pH of a strong alkali from its concentration	• The strong alkali is fully dissociated into OH^- ions • K_w holds for all solutions in water Use alkali conc to get $[OH^-]$ ↓ Use K_w to find $[H^+]$ from $[OH^-]$ ↓ $pH = -\log[H^+]$	***Find the pH of 0.250 mol dm^{-3} NaOH*** NaOH is a strong monobasic alkali $[OH^-] = 0.250$ $K_w = [H^+][OH^-]$ $[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.25} = 4.0 \times 10^{-14}$ $pH = -\log[H^+] = -\log\ 4.0 \times 10^{-14} = \underline{\mathbf{13.40}}$

Finding the concentration of a strong alkali from its pH	• The strong alkali is fully dissociated into OH^- ions • K_w holds for all solutions in water $[H^+] = 10^{-pH}$ Use K_w to find $[OH^-]$ from $[H^+]$ Use $[OH^-]$ to find alkali conc	***Find the concentration of $Ba(OH)_2$ with pH 13.30*** $[H^+] = 10^{-pH} = 10^{-13.30} = 5.01 \times 10^{-14}$ $K_w = [H^+][OH^-]$ $[OH^-] = \frac{K_w}{[H^+]} = \frac{10^{-14}}{5.01 \times 10^{-14}} = 0.200$ $Ba(OH)_2$ is a strong dibasic alkali (so there are two OH^- ions for each $Ba(OH)_2$) $[Ba(OH)_2] = \frac{0.200}{2}$ = **0.100 mol dm^{-3}**
Finding the pH of a strong alkali that is diluted	• The strong alkali is fully dissociated into OH^- ions • K_w holds for all solutions in water Use the dilution factor to get $[OH^-]$ Use K_w to find $[H^+]$ from $[OH^-]$ $pH = -\log[H^+]$	***250 cm^3 of water is added to 100 cm^3 of 0.600 NaOH. Find the pH of the diluted alkali.*** [NaOH] after dilution = $old\ conc\ x\ \frac{old\ volume}{new\ volume}$ $= 0.600\ x\ \frac{100}{350}$ = 0.171 NaOH is a strong monobasic alkali $[OH^-]$ = 0.171 $K_w = [H^+][OH^-]$ $[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.171} = 5.83 \times 10^{-14}$ $pH = -\log[H^+] = -\log 5.83 \times 10^{-14}$ = **13.23**
Finding the pH of water	• In pure water, $[H^+] = [OH^-]$ • Therefore $K_w = [H^+]^2$ for pure water • The pH of water changes with temperature, but it stays neutral as $[H^+] = [OH^-]$ Use $K_w = [H^+]^2$ to find $[H^+]$ $pH = -\log[H^+]$	***Find the pH of water at 50℃ given that K_w = 6.63 x 10^{-14} mol^2 dm^{-6}*** As this is pure water, $[H^+] = [OH^-]$ $K_w = [H^+]^2$ $[H^+] = \sqrt{K_w} = \sqrt{6.63 \times 10^{-14}} = 2.57 \times 10^{-7}$ $pH = -\log[H^+] = -\log 2.57 \times 10^{-7}$ = **6.59**

Finding the pH of a weak acid	• In a weak acid dissolved in water, if nothing else is added then $[H^+] = [A^-]$ • As it is a weak acid, we assume that the equilibrium concentration of HA is the same as the concentration if none of it had dissociated at all when added to water Use $K_a = \frac{[H^+]^2}{[HA]}$ to find $[H^+]$ $pH = -\log[H^+]$	***Find the pH of 0.200 mol dm^{-3} ethanoic acid given that pKa = 4.76*** $K_a = \frac{[H^+][A^-]}{[HA]}$ As this is a weak acid in water with nothing else added, $[H^+] = [A^-]$ $K_a = \frac{[H^+]^2}{[HA]}$ $[H^+]^2 = K_a\,[HA]$ $[H^+] = \sqrt{K_a\,[HA]} = \sqrt{10^{-4.76} \times 0.200} = 1.86 \times 10^{-3}$ $pH = -\log[H^+] = -\log 1.86 \times 10^{-3} =$ **<u>2.73</u>**
Finding the concentration of a weak acid from its pH	• In a weak acid dissolved in water, if nothing else is added then $[H^+] = [A^-]$ • As it is a weak acid, we assume that the equilibrium concentration of HA is the same as the concentration if none of it had dissociated at all when added to water $[H^+] = 10^{-pH}$ Use $K_a = \frac{[H^+]^2}{[HA]}$ to find [HA]	***Find the concentration of ethanoic acid with pH 3.28 given that pKa = 4.76*** $[H^+] = 10^{-pH} = 10^{-3.28} = 5.25 \times 10^{-4}$ $K_a = \frac{[H^+][A^-]}{[HA]}$ As this is a weak acid in water with nothing else added, $[H^+] = [A^-]$ $K_a = \frac{[H^+]^2}{[HA]}$ $[HA] = \frac{[H^+]^2}{K_a} = \frac{[5.25 \times 10^{-4}]^2}{10^{-4.76}}$ [HA] = **<u>0.0158 mol dm^{-3}</u>**

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<td>Finding the pH of a weak acid with a strong alkali added (excess HA)</td>
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- Excess HA: the HA reacts with the OH^- to form A^-

$$HA + OH^- \rightarrow A^- + H_2O$$

- Note that an acidic buffer has been formed as there is a mixture of HA and A^-

Find moles HA and OH^- to see which is in excess

↓

Excess HA: find moles HA and A^-

↓

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

↓

$pH = -\log[H^+]$

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Find the pH of a mixture of 25 cm³ of 0.200 mol dm⁻³ ethanoic acid (pKa = 4.76) with 10 cm³ of 0.100 mol dm⁻³ NaOH

mol HA = $0.200\ x\ \frac{25}{1000}$ = 0.00500

mol OH^- = $0.100\ x\ \frac{10}{1000}$ = 0.00100

HA is in excess

	HA	+	OH^-	→	A^-	+	H_2O
before reaction	0.005		0.001		0		lots
after reaction	0.004		0		0.001		lots

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

$$[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{10^{-4.76}\ x\ \frac{0.004}{\frac{35}{1000}}}{\frac{0.001}{\frac{35}{1000}}} = 6.95 \times 10^{-5}$$

(note that the $\frac{35}{1000}$, which is the new total volume, cancel out)

$pH = -\log[H^+] = -\log 6.95 \times 10^{-5}$ = **<u>4.16</u>**

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<td>Finding the pH of a weak acid with a strong alkali added (with half the acid neutralised)</td>
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- Excess HA where half of the acid is neutralise: as [HA] = [A^-], then pH = pK_a

Find moles HA and OH^- to see which is in excess

↓

Half of HA neutralised

↓

$pH = pK_a$

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Find the pH of a mixture of 20 cm³ of 0.200 mol dm⁻³ ethanoic acid (pKa = 4.76) with 20 cm³ of 0.100 mol dm⁻³ NaOH

mol HA = $0.200\ x\ \frac{20}{1000}$ = 0.00400

mol OH^- = $0.100\ x\ \frac{20}{1000}$ = 0.00200

HA is in excess with exactly half of the HA neutralised

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

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<td>Finding the pH of a weak acid with a strong alkali added (excess OH^-)</td>
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- Excess OH^-: all the HA has gone and so we are dealing with a diluted strong alkali

Find moles HA and OH^- to see which is in excess

Excess OH^-: find mol excess OH^-

Find $[OH^-]$ of the excess OH^-

Use K_w to find $[H^+]$ from $[OH^-]$

$\mathrm{pH} = -\log[\mathrm{H}^+]$
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Find the pH of a mixture of 25 cm³ of 0.200 mol dm⁻³ ethanoic acid (pKa = 4.76) with 50 cm³ of 0.400 mol dm⁻³ NaOH

mol HA = $0.200\ x\ \frac{25}{1000}$ = 0.00500

mol OH^- = $0.400\ x\ \frac{50}{1000}$ = 0.0200

OH^- is in excess

mol excess OH^- = 0.0200 – 0.00500 = 0.0150

$[OH^-] = \frac{0.0150}{\frac{75}{1000}} = 0.200$

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

$[\mathrm{H}^+] = \frac{\mathrm{K_w}}{[\mathrm{OH}^-]} = \frac{10^{-14}}{0.200} = 5.0 \times 10^{-14}$

$\mathrm{pH} = -\log[\mathrm{H}^+] = -\log 5.0 \times 10^{-14}$ = **<u>13.30</u>**
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<td>Finding the pH of a mixture of a strong acid and one of its salts</td>
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- The HA and A^- are in equilibrium with each other in a buffer solution

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

Find moles HA and A^-

Use $\mathrm{K_a} = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{HA}]}$ to find $[H^+]$

$\mathrm{pH} = -\log[\mathrm{H}^+]$
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Find the pH of a mixture of 50 cm³ of 0.100 mol dm⁻³ ethanoic acid (pKa = 4.76) with 100 cm³ of 0.250 mol dm⁻³ sodium ethanoate

mol HA = $0.100\ x\ \frac{50}{1000}$ = 0.0050

mol A^- = $0.250\ x\ \frac{100}{1000}$ = 0.0250

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

$[\mathrm{H}^+] = \frac{\mathrm{K_a}\,[\mathrm{HA}]}{[\mathrm{A}^-]} = \frac{10^{-4.76}\ x\ \frac{0.0050}{\frac{150}{1000}}}{\frac{0.0250}{\frac{150}{1000}}} = 3.48 \times 10^{-6}$

(note that the $\frac{150}{1000}$, which is the new total volume, cancel out)

$\mathrm{pH} = -\log[\mathrm{H}^+] = -\log 3.48 \times 10^{-6}$ = **<u>5.46</u>**
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Adding H^+ to an acidic buffer solution	• The H^+ reacts with the A^- to form more HA $H^+ + A^- \rightarrow HA$ Find moles HA and A^- before adding H^+ ↓ Find moles H^+ added ↓ Find moles HA and A^- after reaction of H^+ with A^- ↓ Use $K_a = \frac{[H^+][A^-]}{[HA]}$ to find [H^+] ↓ $pH = -\log[H^+]$	***A buffer solution is made by mixing 50 cm³ of 0.100 mol dm⁻³ ethanoic acid (pKa = 4.76) with 100 cm³ of 0.250 mol dm⁻³ sodium ethanoate.*** ***Calculate the pH after 1.0 cm³ of 0.50 mol dm⁻³ hydrochloric acid is added to this buffer.*** mol HA before H^+ added = $0.100 \ x \ \frac{50}{1000}$ = 0.0050 mol A^- before H^+ added = $0.250 \ x \ \frac{100}{1000}$ = 0.0250 mol H^+ added = $0.50 \ x \ \frac{1}{1000}$ = 0.0005 $H^+ \ + \ A^- \ \rightarrow \ HA$ before reaction: 0.0005, 0.0250, 0.0050 after reaction: used up, 0.0245, 0.0055 $K_a = \frac{[H^+][A^-]}{[HA]}$ $[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{10^{-4.76} \ x \ \frac{0.0055}{\frac{151}{1000}}}{\frac{0.0245}{\frac{151}{1000}}}$ = 3.90 x 10⁻⁶ (note that the $\frac{151}{1000}$, which is the new total volume, cancel out) $pH = -\log[H^+] = -\log 3.90 \ x \ 10^{-6}$ = <u>**5.41**</u>
Adding OH^- to an acidic buffer solution	• The OH^- reacts with the HA to form more A^- $HA + OH^- \rightarrow A^- + H_2O$ Find moles HA and A^- before adding OH^- ↓ Find moles OH^- added ↓ Find moles HA and A^- after reaction of OH^- with HA ↓ Use $K_a = \frac{[H^+][A^-]}{[HA]}$ to find [H^+] ↓ $pH = -\log[H^+]$	***A buffer solution is made by mixing 50 cm³ of 0.100 mol dm⁻³ ethanoic acid (pKa = 4.76) with 100 cm³ of 0.250 mol dm⁻³ sodium ethanoate.*** ***Calculate the pH after 1.0 cm³ of 0.20 mol dm⁻³ sodium hydroxide is added to this buffer.*** mol HA before OH^- added = $0.100 \ x \ \frac{50}{1000}$ = 0.0050 mol A^- before OH^- added = $0.250 \ x \ \frac{100}{1000}$ = 0.0250 mol OH^- added = $0.20 \ x \ \frac{1}{1000}$ = 0.0002 $HA \ + \ OH^- \ \rightarrow \ A^- \ + \ H_2O$ before reaction: 0.0050, 0.0002, 0.0250, lots after reaction: 0.0048, used up, 0.0252, lots $K_a = \frac{[H^+][A^-]}{[HA]}$ $[H^+] = \frac{K_a\,[HA]}{[A^-]} = \frac{10^{-4.76} \ x \ \frac{0.0048}{\frac{151}{1000}}}{\frac{0.0252}{\frac{151}{1000}}}$ = 3.31 x 10⁻⁶ (note that the $\frac{151}{1000}$, which is the new total volume, cancel out) $pH = -\log[H^+] = -\log 3.31 \ x \ 10^{-6}$ = <u>**5.48**</u>