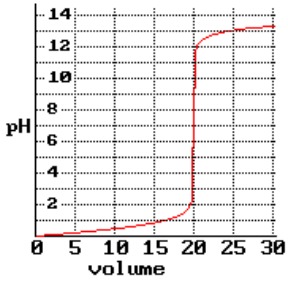
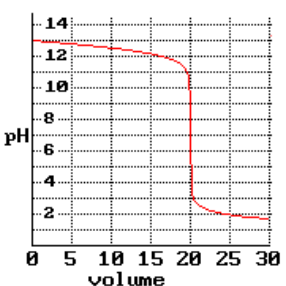
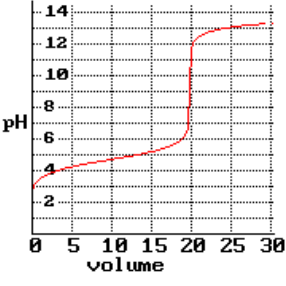
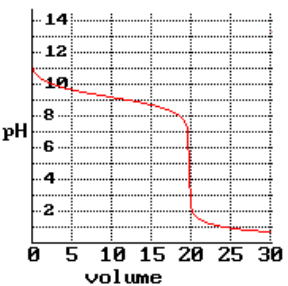
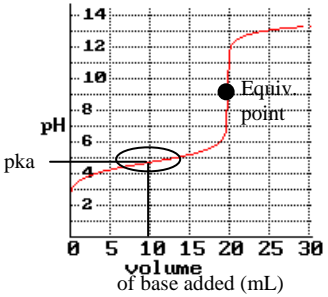


$\text{pH} = -\log [\text{H}^+]$ <p>and</p> $[\text{H}^+] = \text{inv log} (-\text{pH})$	$K_w = 1 \times 10^{-14}$ $K_w = [\text{H}^+] \times [\text{OH}^-]$	$[\text{H}^+] = K_w / [\text{OH}^-]$ <p>and</p> $[\text{OH}^-] = K_w / [\text{H}^+]$	$\text{pH} + \text{pOH} = 14$
pH formulae	ionic product for water	using K_w to find $[\text{H}^+]$ or $[\text{OH}^-]$	relationship between pH & pOH
$\text{pOH} = -\log [\text{OH}^-]$ <p>and</p> $[\text{OH}^-] = \text{inv log}(-\text{pOH})$	$K_w = K_a \times K_b$ $\text{p}K_w = \text{p}K_a + \text{p}K_b$ $\text{p}K_w = 14$	$\text{p}K_a = -\log K_a$ $\text{p}K_b = -\log K_b$ $\text{p}K_w = -\log K_w$	$K_a = \text{inv log} (-\text{p}K_a)$ $K_b = \text{inv log} (-\text{p}K_b)$ $K_w = \text{inv log} (-\text{p}K_w)$
pOH formulae	relationship between K_a, K_b and K_w and $\text{p}K_a, \text{p}K_b$ and $\text{p}K_w$	turning K_a, K_b, K_w into $\text{p}K_a, \text{p}K_b, \text{p}K_w$	turning $\text{p}K_a, \text{p}K_b, \text{p}K_w$ into K_a, K_b, K_w
<p>For a weak acid</p> $K_a = \frac{[\text{H}^+]^2}{[\text{acid}]}$	<p>For a weak base</p> $K_b = \frac{[\text{OH}^-]^2}{[\text{base}]}$	$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$ $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$	K_s or K_{sp}
calculate the pH of a weak acid	calculate the pH of a weak base	equation for the self ionisation of water	solubility product
$K_s = [\text{Ag}^+][\text{Cl}^-]$	$K_s = [\text{Ag}^+]^2[\text{S}^{2-}]$	$K_s = [\text{Pb}^{2+}][\text{Cl}^-]^2$	$K_s = s^2$ $s = \sqrt{K_s}$
solubility product expression for AgCl (AB type)	solubility product expression for Ag ₂ S (A ₂ B type)	solubility product expression for PbCl ₂ (AB ₂ type)	AB type calculations

$K_s = 4s^3$ $s = \sqrt[3]{K_s/4}$	Solubility can be measured in mol L ⁻¹ or g L ⁻¹ BUT s is mol L ⁻¹	mol L ⁻¹ into g L ⁻¹ , × M g L ⁻¹ into mol L ⁻¹ , ÷ M	$n = m/M$ n = amount, in moles m = mass, in g M = molar mass, in g mol ⁻¹
AB ₂ or A ₂ B type calculations	solubility	conversions	formula relating n, m and M
If I.P. < K _s precipitation will	If I.P. > K _s precipitation will.....	precipitation will continue until the ion concentrations _____ to the equilibrium levels	What you get when you substitute the ACTUAL VALUES of the concentrations of the ions that you have after mixing
not occur	occur	decrease	ionic product
solubility of one salt is reduced by the presence of another having a common ion	Equilibrium constant describing the solubility of compounds that are slightly (sparingly) soluble in water	a solution that contains as much dissolved solute as it can under a given set of conditions	$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$
common ion effect	K _s (also known as K _{sp})	saturated solution	Buffer pH equation
solution whose pH resists change upon addition of small amounts of H ⁺ and OH ⁻	Examples of buffers CH ₃ COOH / CH ₃ COO ⁻ and NH ₃ /NH ₄ ⁺	buffer solutions are most effective when concentrations of acid & base are _____	the pH when equivalent amounts of acid and alkali have been mixed
buffer solution	acidic buffer basic buffer	Equal / equimolar	Equivalence point

working range of indicator is approximately equal to pK_a of indicator ± 1 pH unit	pH change is so gradual at the end point, that colour changes are difficult to observe for a titration between ___ and ___		
Choosing an indicator	WA/WB (or vice versa)	SB added to SA Eg NaOH to HCl	SA added to a SB Eg HCl to NaOH
		Long vertical region Equivalence point at $pH = 7$	Short vertical region. Equivalence point at $pH > 7$
SB added to WA Eg NaOH to CH_3COONa	SA added to WB Eg HCl to NH_3	SA-SB titration curve	WA - SB titration curve
Short vertical region. Equivalence point at $pH < 7$		Region where acid/base conjugate pair are in approx equal concentrations	When acid and conjugate base (or base and conjugate acid) are equal, $pH =$
SA - WB titration curve	Buffer region and pK_a (about 4.8) and $EP > 7$	Buffer region	$pH = pK_a + \log 1/1$ so $pH = pK_a$
If a buffer is diluted its pH is changed but its ability to act as a buffer.....	Curve begins at high pH typical of strong base and ends at low pH typical of strong acid. There is a large rapid change in pH near the equivalence point which is at pH 7	Curve begins at low pH typical of strong acid, and ends at high pH typical of strong base. There is a large rapid change in pH near the equivalence point which is at pH 7	Curve begins at a higher acidic pH and ends at high basic pH. The pH at the equivalence point is > 7 .
Decreases (as lower conc. of acid and base)	HCl (SA) added to NaOH (SB)	NaOH (SB) added to HCl (SA)	NaOH (SB) added to ethanoic acid CH_3COOH (WA)

Curve begins at a high basic pH and ends a higher acidic pH. The pH at the equivalence point is > 7 .	Curve begins at low pH and ends at a less high basic pH. The pH at the equivalence point is < 7	Curve begins a less high basic pH and ends at a low pH. The pH at the equivalence point is < 7	Curve begins at higher acidic pH and ends at low basic pH. There is not a great pH change at the equivalence point (pH ~ 7) making this a very difficult titration to perform.
CH_3COOH (WA) added to NaOH (SB)	Ammonia NH_3 (WB) added to HCl (SA)	HCl (SA) added to Ammonia NH_3 (WB)	Ammonia NH_3 (WB) added to ethanoic acid CH_3COOH (WA)
<p><u>Most</u> H_2O $\text{H}^+ = \text{Cl}^-$ OH^- <u>Least</u> pH (1), high electrical conductivity</p>	<p><u>Most</u> H_2O $\text{Na}^+ = \text{OH}^-$ H^+ <u>Least</u> pH (13), high electrical conductivity</p>	<p><u>Most</u> H_2O $\text{Na}^+ = \text{Cl}^-$ $\text{H}^+ = \text{OH}^-$ <u>Least</u> Neutral pH, high electrical conductivity</p>	<p><u>Most</u> H_2O Cl^- Ca^{2+} $\text{H}^+ = \text{OH}^-$ <u>Least</u> Neutral pH, high electrical conductivity</p>
0.1 mol L^{-1} HCl species in solution	0.1 mol L^{-1} NaOH species in solution	0.1 mol L^{-1} NaCl species in solution	0.1 mol L^{-1} CaCl_2 species in solution
<p><u>Most</u> H_2O $\text{C}_2\text{H}_5\text{OH}$ $\text{H}^+ = \text{OH}^-$ <u>Least</u> Neutral pH, low electrical conductivity</p>	<p><u>Most</u> H_2O CH_3COOH $\text{H}^+ = \text{CH}_3\text{COO}^-$ OH^- <u>Least</u> pH (3), low electrical conductivity</p>	<p><u>Most</u> H_2O NH_3 $\text{NH}_4^+ = \text{OH}^-$ H^+ <u>Least</u> pH (11), low electrical conductivity</p>	<p><u>Most</u> H_2O $\text{NH}_4^+ = \text{Cl}^-$ $\text{NH}_3 = \text{H}^+$ OH^- <u>Least</u> pH (4), high electrical conductivity</p>
0.1 mol L^{-1} $\text{C}_2\text{H}_5\text{OH}$ species in solution	0.1 mol L^{-1} CH_3COOH species in solution	0.1 mol L^{-1} NH_3 species in solution	0.1 mol L^{-1} NH_4Cl species in solution
<p><u>Most</u> H_2O $\text{CH}_3\text{COO}^- = \text{Na}^+$ $\text{CH}_3\text{COOH} = \text{OH}^-$ H^+ <u>Least</u> pH (9), high electrical conductivity</p>	<p>SALT HYDROLYSIS $\text{NH}_4\text{Cl} + \text{aq} \rightarrow \text{NH}_4^+ + \text{Cl}^-$ $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$ pH of solution will be</p>	<p>SALT HYDROLYSIS $\text{CH}_3\text{COONa} + \text{aq} \rightarrow \text{CH}_3\text{COO}^- + \text{Na}^+$ $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$ pH of solution will be</p>	<p>when salts are dissolved in water the solution is not always neutral due to reversible reactions between the salt ions and the water – this is called...</p>
0.1 mol L^{-1} CH_3COONa species in solution	pH < 7 / acidic	pH > 7 / alkaline	salt hydrolysis