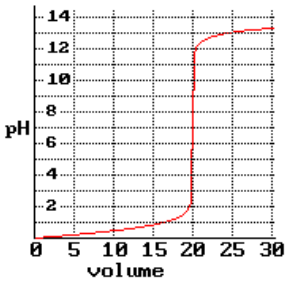
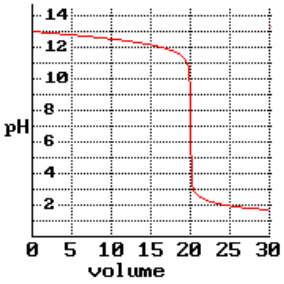
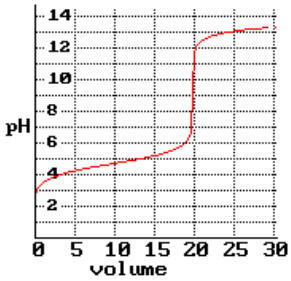
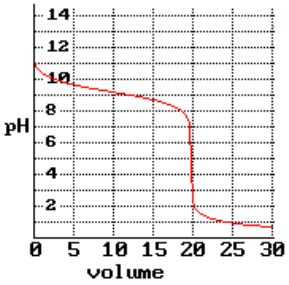
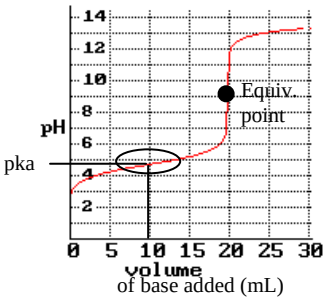
$pH = - \log [H^+]$ and $[H^+] = \text{inv log } (- pH)$	$K_w = 1 \times 10^{-14}$ and $K_w = [H^+] \times [OH^-]$	$[H^+] = K_w / [OH^-]$ and $[OH^-] = K_w / [H^+]$	$pH + pOH = 14$
pH formulae	ionic product for water	using Kw to find $[H^+]$ or $[OH^-]$	relationship between pH & pOH
$pOH = - \log [OH^-]$ and $[OH^-] = \text{inv log}(- pOH)$	$K_w = K_a \times K_b$ and $pK_w = pK_a + pK_b$ $pK_w = 14$	$pK_a = - \log K_a$ $pK_b = - \log K_b$ $pK_w = - \log K_w$	$K_a = \text{inv log } (- pK_a)$ $K_b = \text{inv log } (- pK_b)$ $K_w = \text{inv log } (- pK_w)$
pOH formulae	relationship between Ka,Kb and Kw and pKa, pKb and pKw	turning Ka, Kb, Kw into pKa, pKb, pKw	turning pKa, pKb, pKw into Ka, Kb, Kw
For a weak acid $K_a = \frac{[H^+]^2}{[\text{acid}]}$	For a weak base $K_b = \frac{[OH^-]^2}{[\text{base}]}$	$2H_2O \rightleftharpoons H_3O^+ + OH^-$ $H_2O \rightleftharpoons H^+ + OH^-$	Ks or Ksp
calculate the pH of a weak acid	calculate the pH of a weak base	equation for the self ionisation of water	solubility product
$K_s = [Ag^+][Cl^-]$	$K_s = [Ag^+]^2[S^{2-}]$	$K_s = [Pb^{2+}][Cl^-]^2$	$K_s = s^2$ $S = \sqrt{K_s}$
solubility product expression for AgCl (AB type)	solubility product expression for Ag_2S (A_2B type)	solubility product expression for $PbCl_2$ (AB_2 type)	AB type calculations

$Ks = 4s^3$ $s = \sqrt[3]{K_s/4}$	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}, $\times$ M g L^{-1} into mol L^{-1}, $\div$ M	$n = m/M$ n = amount, in moles m = mass, in g M = molar mass, in g mol^{-1}
AB_2 or A_2B type calculations	solubility	conversions	formula relating n, m and M
If I.P. < Ks precipitation will ……	If I.P. > Ks 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 = pKa + log [base]/[acid]
common ion effect	Ks (also known as Ksp)	saturated solution	Buffer pH equation
solution whose pH resists change upon addition of small amounts of H^+ and OH^-	Examples of buffers CH_3COOH / CH_3COO^- 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 pKa 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 pKa (about 4.8) and EP > 7	Buffer region	pH =pKa + log 1/1 so pH = pKa
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)
<u>Most</u> H_2O $H^+ = Cl^-$ OH^- <u>Least</u> pH (1), high electrical conductivity	<u>Most</u> H_2O $Na^+ = OH^-$ H^+ <u>Least</u> pH (13), high electrical conductivity	<u>Most</u> H_2O $Na^+ = Cl^-$ $H^+ = OH^-$ <u>Least</u> Neutral pH, high electrical conductivity	<u>Most</u> H_2O Cl^- Ca^{2+} $H^+ = OH^-$ <u>Least</u> Neutral pH, high electrical conductivity
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
<u>Most</u> H_2O C_2H_5OH $H^+ = OH^-$ <u>Least</u> Neutral pH, low electrical conductivity	<u>Most</u> H_2O CH_3COOH $H^+ = CH_3COO^-$ OH^- <u>Least</u> pH (3), low electrical conductivity	<u>Most</u> H_2O NH_3 $NH_4^+ = OH^-$ H^+ <u>Least</u> pH (11), low electrical conductivity	<u>Most</u> H_2O $NH_4^+ = Cl^-$ $NH_3 = H^+$ OH^- <u>Least</u> pH (4), high electrical conductivity
0.1 mol L^{-1} C_2H_5OH 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
<u>Most</u> H_2O $CH_3COO^- = Na^+$ $CH_3COOH = OH^-$ H^+ <u>Least</u> pH (9), high electrical conductivity	SALT HYDROLYSIS $NH_4Cl + aq \rightarrow NH_4^+ + Cl^-$ $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ pH of solution will be	SALT HYDROLYSIS $CH_3COONa + aq \rightarrow CH_3COO^- + Na^+$ $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ pH of solution will be	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…
0.1 mol L^{-1} CH_3COONa species in solution	pH <7 / acidic	pH >7 / alkaline	salt hydrolysis