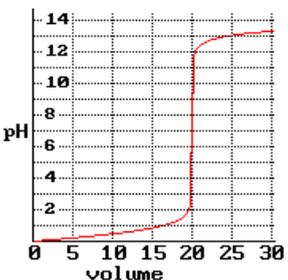
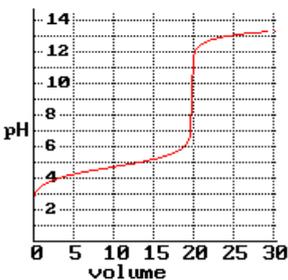
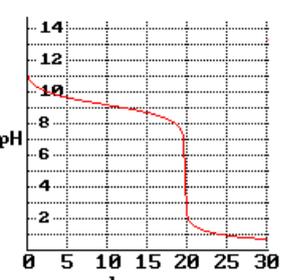
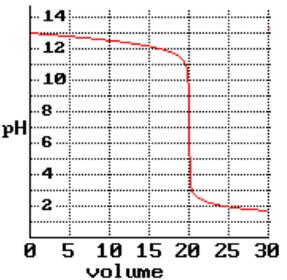


$\text{pH} = -\log [\text{H}^+]$	$[\text{H}^+] = \text{inv log} (-\text{pH})$	1.45×10^{-3} is entered into your calculator as 1.45 EXP - 3	$K_w = 1 \times 10^{-14}$ $K_w = [\text{H}^+] \times [\text{OH}^-]$
calculate pH from $[\text{H}^+]$	calculate $[\text{H}^+]$ from pH	entering values into your calculator	ionic product for water
$\text{pOH} = -\log [\text{OH}^-]$	$[\text{OH}^-] = \text{inv log} (-\text{pOH})$	$\text{pH} + \text{pOH} = 14$	$[\text{H}^+] = K_w / [\text{OH}^-]$ and $[\text{OH}^-] = K_w / [\text{H}^+]$
calculate pOH from $[\text{OH}^-]$	calculate $[\text{OH}^-]$ from pOH	relationship between pH & pOH	using K_w to find $[\text{H}^+]$ or $[\text{OH}^-]$
For a weak acid $K_a = \frac{[\text{H}^+]^2}{[\text{acid}]}$	For a weak base $K_b = \frac{[\text{OH}^-]^2}{[\text{base}]}$	$K_w = K_a \times K_b$ $\text{p}K_w = \text{p}K_a + \text{p}K_b$	$[\text{H}^+] = \text{inv log} (-\text{pH})$ If pH is 3.45 enter shift log - 3.34 to find $[\text{H}^+]$
calculate the pH of a weak acid	calculate the pH of a weak base	relationship between K_a, K_b and K_w and $\text{p}K_a, \text{p}K_b$ and $\text{p}K_w$	entering values into your calculator
$\text{p}K_a = -\log K_a$ $\text{p}K_b = -\log K_b$ $\text{p}K_w = -\log K_w$	$K_a = \text{inv log} (-\text{p}K_a)$ $K_b = \text{inv log} (-\text{p}K_b)$ $K_w = \text{inv log} (-\text{p}K_w)$	$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$ Or, more simply, $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$	$\text{p}K_w = 14$ $\text{p}K_a + \text{p}K_b = 14$
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	equation for the self ionisation of water	relationship between $\text{p}K_a, \text{p}K_b$ and $\text{p}K_w$

Ks or Ksp	Sparingly soluble AgCl, CaCO ₃ , and BaSO ₄ are examples of ___ type compounds	Sparingly soluble Mg(OH) ₂ , PbCl ₂ and Ag ₂ S are examples of ___ type compounds	$K_s = [Ag^+][Cl^-]$
solubility product	AB	AB ₂ or A ₂ B	solubility product expression for AgCl
$K_s = s^2$	$K_s = 4s^3$	$K_s = [Ag^+]^2[S^{2-}]$	$K_s = [Pb^{2+}][Cl^-]^2$
calculating Ks from s (AB type)	calculating Ks from s (AB ₂ or A ₂ B) type	solubility product expression for Ag ₂ S	solubility product expression for PbCl ₂
$s = \sqrt{K_s}$	$s = \sqrt[3]{K_s/4}$	s measured in mol L ⁻¹	Multiply by M, the mass of one mole of the substance
calculating s from Ks (AB type)	Calculating s from Ks (AB ₂ or A ₂ B) type	solubility	to convert mol L ⁻¹ into g L ⁻¹
If I.P. < Ks precipitation will	If I.P. > Ks precipitation will.....	$n = m/M$ n = amount, in moles m = mass, in g M = molar mass, in g mol ⁻¹	Divide by M, the mass of one mole of the substance
not occur	occur	formula relating n, m and M	to convert g L ⁻¹ into mol L ⁻¹

multiply solubility, s , given in mol L^{-1} $\times M$ to find...	divide solubility, given in g L^{-1} by M to find...	If you're not sure what will form, read the question as the known K_s of the compound must be given. That is what might or might not form depending on the IONIC PRODUCT	What you get when you substitute the ACTUAL VALUES of the concentrations of the ions that you have
solubility in g L^{-1}	solubility in mol L^{-1}	knowing <u>what</u> might (or not) precipitate	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	same form as the K_s expression but uses the concentrations of ions at the start of the reaction eg on mixing 2 solutions	The smaller the value of K_s , the lower the.....
common ion effect	K_s (also known as K_{sp})	ionic product	solubility
precipitation will continue until the ion concentrations _____ to the equilibrium levels	formation of a solid from solution	compound with a small K_{sp} value is only _____ in water	a solution that contains as much dissolved solute as it can under a given set of conditions
decrease	precipitation	slightly soluble	saturated solution

$\text{pH} = \text{pKa} + \log \frac{[\text{base}]}{[\text{acid}]}$	solution whose pH changes very little on addition of small amounts of H^+ and OH^-	CH_3COOH mixed with CH_3COO^- (from CH_3COONa) is an _____	NH_3 mixed with NH_4^+ (from NH_4Cl) is a _____
Buffer pH equation	buffer solution	acidic buffer	basic buffer
buffer solutions are most effective when concentrations of acid & base are _____	the pH when equivalent amounts of acid and alkali have been mixed	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 _____
Equal / equimolar	Equivalence point	Choosing an indicator	WA/WB (or vice versa)
substance which changes colour according to the pH. - most are weak acids			
indicator	SB added to SA Eg NaOH to HCl	SB added to WA Eg NaOH to CH_3COONa	SA added to WB Eg HCl to NH_3
	Long vertical region Equivalence point at pH = 7	Short vertical region. Equivalence point at pH > 7	Short vertical region. Equivalence point at pH < 7
SA added to a SB Eg HCl to NaOH	SA-SB titration curve	WA - SB titration curve	SA - WB titration curve

Region either side of point where CH_3COOH and CH_3COO^- are in approx equal concentrations	Region either side of point where NH_3 and NH_4^+ are in approx equal concentrations	When acid and conjugate base (or base and conjugate acid) are equal, $\text{pH} =$	
Buffer region	Buffer region	$\text{pH} = \text{pKa} + \log 1/1$ so $\text{pH} = \text{pKa}$	Buffer region and pKa (about 4.8) and $\text{EP} > 7$
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 $\text{pH} 7$	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 $\text{pH} 7$	Curve begins at a higher acidic pH and ends at high basic pH. The pH at the equivalence point is > 7 .	Curve begins at a high basic pH and ends a higher acidic pH. The pH at the equivalence point is > 7 .
HCl (SA) added to NaOH (SB)	NaOH (SB) added to HCl (SA)	NaOH (SB) added to ethanoic acid CH_3COOH (WA)	CH_3COOH (WA) added to NaOH (SB)
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 ($\text{pH} \sim 7$) making this a very difficult titration to perform.	when salts are dissolved in water the solution is not always neutral due to reversible reactions between the salt ions and the water.
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)	salt hydrolysis
If a buffer is diluted its pH is changed but its ability to act as a buffer.....			
Decreases (as lower conc. of acid and base)			

<p><u>Most</u> H₂O H⁺ = Cl⁻ OH⁻ <u>Least</u> pH (1), high electrical conductivity</p>	<p><u>Most</u> H₂O Na⁺ = OH⁻ H⁺ <u>Least</u> pH (13), high electrical conductivity</p>	<p><u>Most</u> H₂O Na⁺ = Cl⁻ H⁺ = OH⁻ <u>Least</u> Neutral pH, high electrical conductivity</p>	<p><u>Most</u> H₂O Cl⁻ Ca²⁺ H⁺ = 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 CaCl₂ species in solution
<p><u>Most</u> H₂O C₂H₅OH H⁺ = OH⁻ <u>Least</u> Neutral pH, low electrical conductivity</p>	<p><u>Most</u> H₂O CH₃COOH H⁺ = CH₃COO⁻ OH⁻ <u>Least</u> pH (3), low electrical conductivity</p>	<p><u>Most</u> H₂O NH₃ NH₄⁺ = OH⁻ H⁺ <u>Least</u> pH (11), low electrical conductivity</p>	<p><u>Most</u> H₂O NH₄⁺ = Cl⁻ NH₃ = H⁺ OH⁻ <u>Least</u> pH (4), high electrical conductivity</p>
0.1 mol L⁻¹ C₂H₅OH species in solution	0.1 mol L⁻¹ CH₃COOH species in solution	0.1 mol L⁻¹ NH₃ species in solution	0.1 mol L⁻¹ NH₄Cl species in solution
<p><u>Most</u> H₂O CH₃COO⁻ = Na⁺ CH₃COOH = OH⁻ H⁺ <u>Least</u> pH (9), high electrical conductivity</p>	<p>SALT HYDROLYSIS NH₄Cl + aq → NH₄⁺ + Cl⁻ NH₄⁺ + H₂O ⇌ NH₃ + H₃O⁺ pH of solution will be</p>	<p>SALT HYDROLYSIS CH₃COONa + aq → CH₃COO⁻ + Na⁺ CH₃COO⁻ + H₂O ⇌ CH₃COOH + OH⁻ pH of solution will be</p>	
0.1 mol L⁻¹ CH₃COONa species in solution	pH <7 / acidic	pH >7 / alkaline	