

Therefore the pH of water at 298K is 7, however this isn't the case when you change the temperature though.

The value of K_w will change as temperature changes. It is unlikely you will be required to calculate pH of a solution as temperature changes however you need to be aware that the pH of pure water changes though. This doesn't mean that the water becomes more acidic/alkaline though. Remember what I said, a neutral solution is a solution whereby $[H_3O^+] = [OH^-]$, regardless

We can use Le Chatelier's principle to predict this change. The dissociation of water is endothermic so increasing the temperature would cause the equilibrium to favour the product sides (the right-hand side) and therefore $[H_3O^+]$ will increase so the pH would drop. Decreasing the temperature would cause the equilibrium to favour the reactant's side (the left-hand side) and therefore $[H_3O^+]$ decreases, resultantly pH increases.

So water has higher pH at lower temperatures and lower pH at higher temperature but is STILL neutral!

Calculating pH of a strong acid:

I mentioned earlier that a strong acid is an acid which completely dissociates to its ions. For example, hydrochloric acid is regarded as a strong acid.

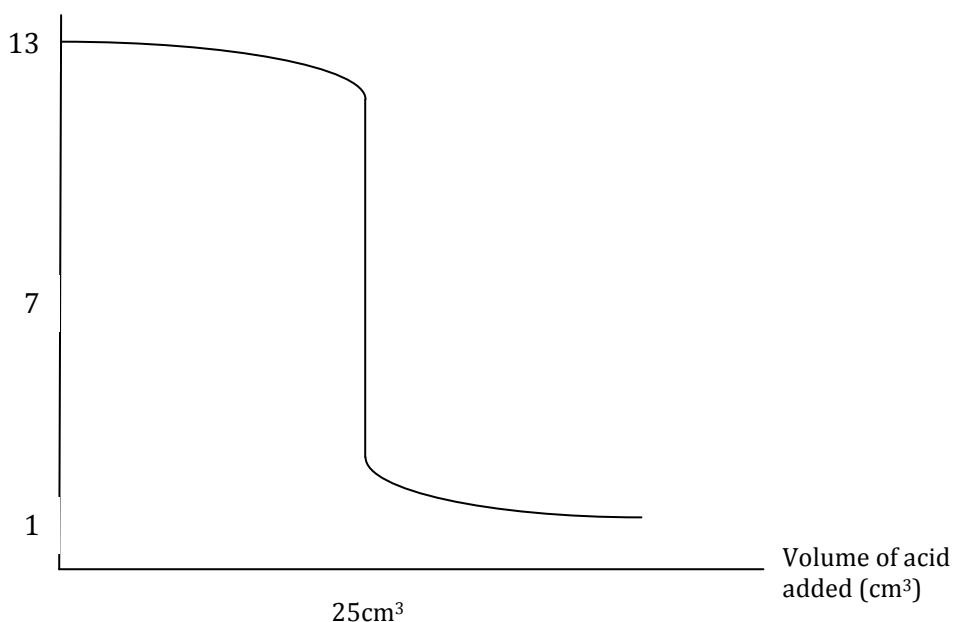
To calculate the pH of a strong acid, we make an important assumption. The concentration of hydrogen ions (or hydronium) in a strong monoprotic acid will be the same as the concentration of the acid itself.

Therefore to calculate the pH of $0.1 \text{ mol dm}^{-3} \text{ HCl}$, we can use the following logarithm:

$$\log_{10}(0.1) = 1.00$$

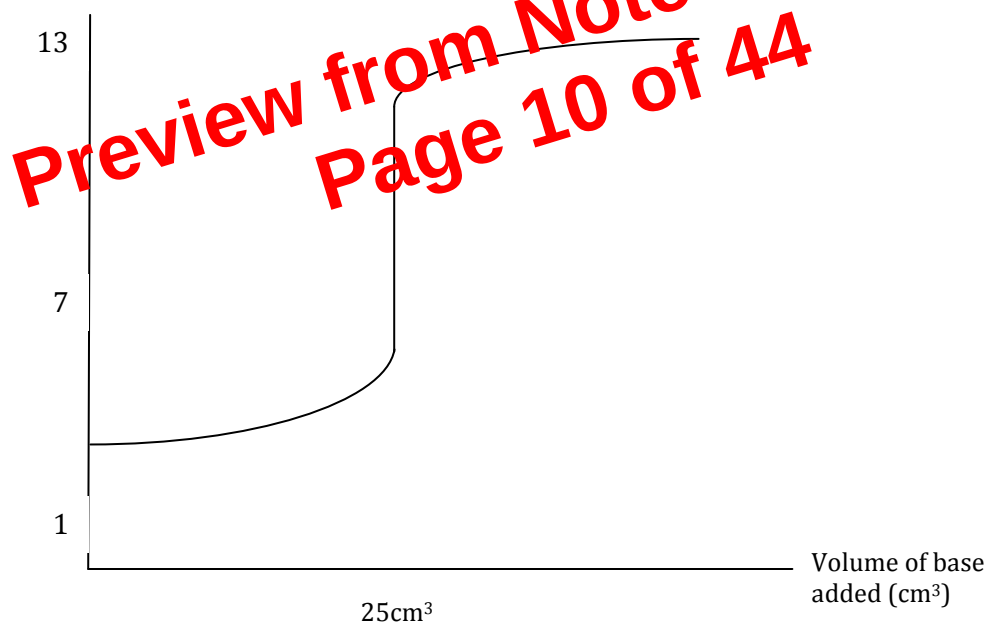
As mentioned before, always give pH values to 2 decimal places unless otherwise stated.

Now, for polyprotic acid, you can't exactly make the same assumption like sulphuric acid because they donate two moles of protons per mole of acid so therefore the concentration of hydrogen ions will not equal the concentration of the acid.



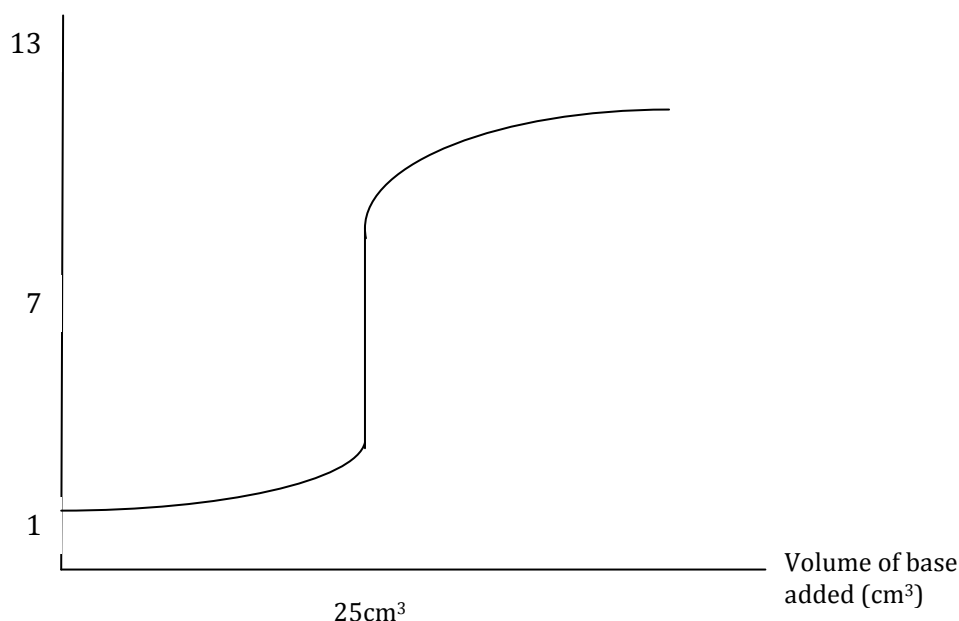
This particular titration curve is slightly different; here the acid is run through the burette rather than the alkali. However it essentially follows the same pattern, the equivalence point can be calculated from reading the midpoint of the vertical portion of the graph. This would be pH 7 again.

Weak acid-strong base titration:



The above titration curve is for a weak acid-strong base titration for instance between ethanoic acid and potassium hydroxide. 25cm^3 of ethanoic acid is added to a conical flask and potassium hydroxide is run through a burette. As the base is added, the pH of the acid gradually increases from pH 3 to pH 7 until 25cm^3 . When 25cm^3 of base is added to the solution, the pH drastically increases from pH 7 to pH 11 and resultantly

Strong acid-weak base titration:



The above titration curve is for a strong acid-weak base titration for instance between hydrochloric acid and ammonia. 25cm³ of hydrochloric acid is added to a conical flask and ammonia is run through a burette. As the base is added, the pH of the acid gradually increases from pH 1 to pH 3 until 25cm³. When 25cm³ of base is added to the solution, the pH drastically increases from pH 3 to pH 7 and the solution becomes more alkaline than acidic. After 25cm³, the pH gradually increases until it reaches pH 11.

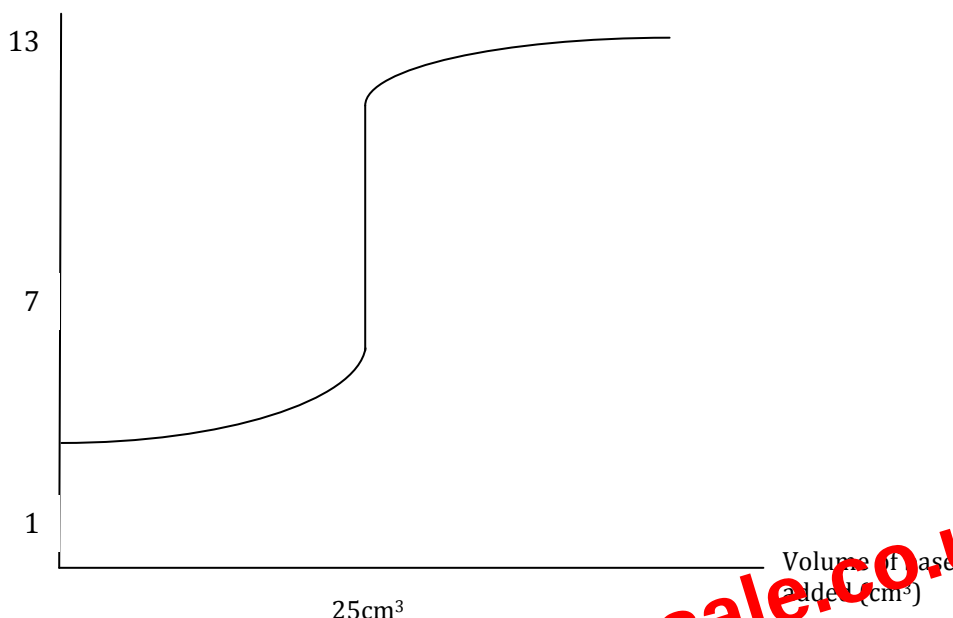
Key points to make out:

1. The pH of the acid starts at pH 1 and increases until pH 3 upon the addition of alkali until about when the volume of the acid almost equals the volume of the base
2. The pH drastically increases from 3 to 7 when the volumes of both acid and base equal each other.
3. The pH continues to increase as the base is added in excess. The pH will go from 9 to 11.
4. The vertical portion of the graph is from pH 3 to 7 and therefore the equivalence point is at pH 7.

Consider the following equation:



This reaction is a weak acid-strong base reaction between propanoic acid and sodium hydroxide. The titration curve will look a bit like this:



This will be expressed in the K_a expression like such:

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{CH}_2\text{COO}^-]}{[\text{CH}_3\text{CH}_2\text{COOH}]}$$

As you increase the alkali, you are producing more and more of the carboxylate anion. Eventually at the exact middle (so 12.5cm^3 of the base is added in the titration earlier), the amount of the weak acid and its conjugate base will equal each other.

So you can now cancel out the acid and the conjugate base in the K_a expression:

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{CH}_2\text{COO}^-]}{[\text{CH}_3\text{CH}_2\text{COOH}]}$$

$$K_a = [\text{H}^+]$$

Therefore the value of K_a equals the concentration of hydrogen in the solution.

$$-\log_{10}[\text{H}^+] = \text{pH}$$

$$-\log_{10}(K_a) = \text{p}K_a$$

As pH is the negative logarithm of the concentration of hydrogen and pK_a shares the same relationship with the value of K_a , therefore it can be said that pK_a equals pH:

$$pK_a = pH$$

Generally, you can find the pH of the acidic solution by reading the pH at the half-equivalence point and from there; you can calculate the value of K_a or $[H_3O^+]$

In the exam, you may receive a question giving you a titration curve and expecting you to calculate the K_a for a weak acid. This doesn't exactly mean that you have to use half-equivalence, rather if they refer to the acid alone (without suggesting any neutralisation occurs or half-equivalence) without providing you the pH to calculate K_a , you need to determine the pH by reading pH at the very beginning of the titration curve (before any alkali is added) and from there, calculate K_a by using the "regular" method.

Titration and indicators:

For different acid-base titrations, you will have different indicators which can be used. You need to be aware of the different indicators you can use and why you use them.

In any strong acid-strong base titration, you can use either methyl orange or phenolphthalein because the colour change of both occur during the vertical portion of the graph (from 3 to 11)

In any weak acid-strong base titration, you use phenolphthalein because the colour change for phenolphthalein occurs during the vertical portion of the graph (from about pH 5 to 9)

In any strong acid-weak base titration, you use methyl orange because the colour change for methyl orange occurs during the vertical portion of the graph (from about pH 3 to 7)

In any weak acid-weak base titration, you do not use any indicator because there is no visible vertical portion of the graph.

Sometimes in the exam, they will give a selection of unknown indicators and will expect you to indicate what is the most appropriate indicator based on the type of titration.

Again, look at the pH range for the indicator; if the pH range is within the vertical portion of the selected titration, then the indicator will be suitable as the colour change will occur during the vertical portion of the titration. If however the pH range is not within the vertical portion of the selected titration, then the indicator will not be suitable. Never ever mention universal indicator as a suitable indicator!

Adding a salt to a weak acid to form a buffer solution

Probably the hardest of the three but so long as you can recall molar calculations, you should be fine

Nitrous acid, HNO_2 , is a weak acid with an acid dissociation constant of 4.7×10^{-4} at 4°C .

Calculate the pH of the buffer solution made by adding 1.38g of sodium nitrite (NaNO_2) to 100cm^3 of the 0.12 mol dm^{-3} solution of nitrous acid at this temperature.

Okay, you have a salt which you're adding to a solution of a weak acid. The first you do is calculate moles of this salt using mass/ M_r

Using a periodic table, you'll find out that the molar mass of sodium nitrite is 69.

$$\frac{1.38}{6.9} = 0.02 \text{ moles}$$

Now, you have to calculate the concentration of this salt in the solution. Here you make the assumption that the volume of the solution does not change. Before you calculate the concentration, you need to convert the volume into the appropriate units.

$$100/1000 = 0.1 \text{ dm}^3$$

Calculate the concentration of the salt.

$$\frac{1.38}{6.9} = 0.2 \text{ M}$$

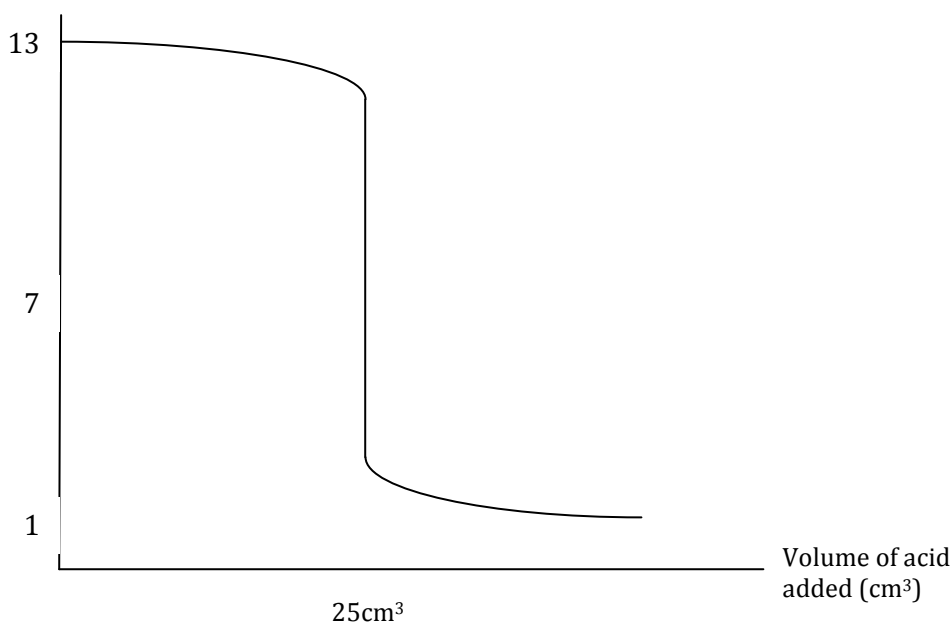
Once you calculate the concentration of the salt, you can now insert everything into the K_a expression:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

$$4.7 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][0.2]}{[0.12]}$$

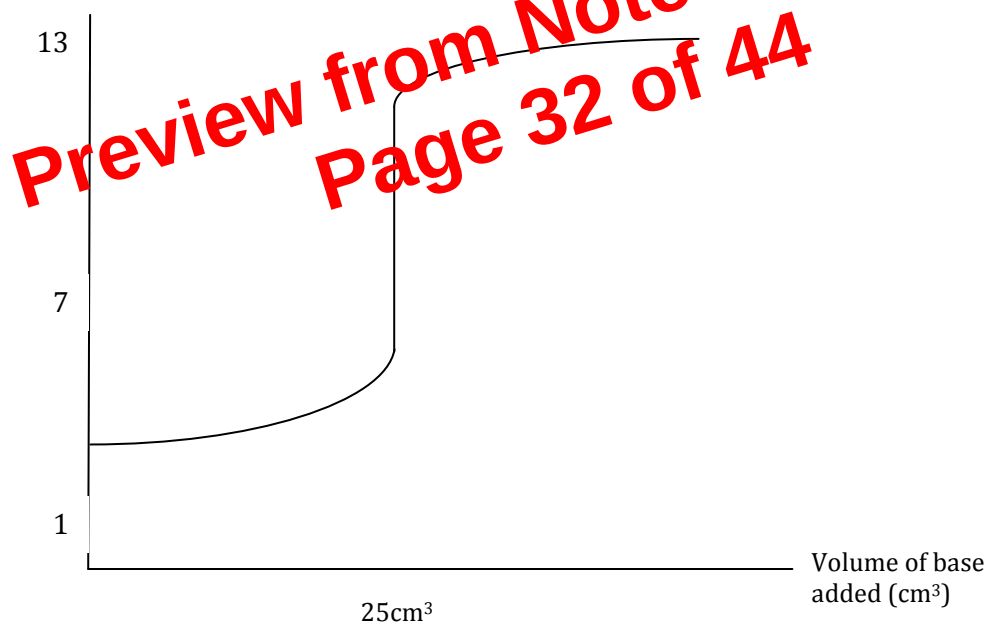
Rearrange the expression to give:

$$\frac{K_A \times [\text{HCOOH}]}{[\text{HCOO}^-]} = [\text{H}_3\text{O}^+]$$



This particular titration curve is slightly different; here the acid is run through the burette rather than the alkali. However it essentially follows the same pattern, the equivalence point can be calculated from reading the midpoint of the vertical portion of the graph. This would be pH 7 again.

Weak acid-strong base titration:



The above titration curve is for a weak acid-strong base titration for instance between ethanoic acid and potassium hydroxide. 25cm^3 of ethanoic acid is added to a conical flask and potassium hydroxide is run through a burette. As the base is added, the pH of the acid gradually increases from pH 3 to pH 7 until 25cm^3 . When 25cm^3 of base is added to the solution, the pH drastically increases from pH 7 to pH 11 and resultantly