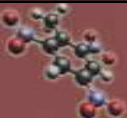


# Acids and Bases

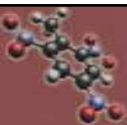
Keywords  
 $K_a$  and  $pK_a$   
pH and pOH  
 $K_a$ ,  $K_b$  and  $K_w$   
Buffers

Author: J R Reid



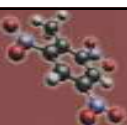
## Keywords

- Acid and Base: A proton ( $H^+$ ) donator is an acid, a proton acceptor is a base.
- Weak and strong acids: An acid that dissociates it's hydrogen readily is strong. Weak ones don't dissociate well.
- Concentrated and dilute acids: High concentrations = concentrated, low concentration = dilute.
- Acid Dissociation Constant ( $K_a$ ): The equilibrium constant based on the acid dissociation.
- Conjugate base: The what the weak acid turns into once it has lost it's hydrogen ion e.g.  $NH_4^+$  becomes  $NH_3$
- Buffer: A mixture of acid and base that absorbs acids without large changes in pH



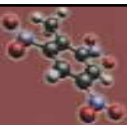
## K<sub>a</sub> and pK<sub>a</sub>

- As we know the K<sub>a</sub> is the acid dissociation constant (see Level 2 *Equilibrium - Acids and Bases*)
- We have also seen that pH is a way of looking at the hydrogen/hydronium ion concentrations in a solution
 
$$\text{pH} = -\log[\text{H}_3\text{O}^+_{(\text{aq})}]$$
- K<sub>a</sub> can also be converted to a log version called pK<sub>a</sub>:
 
$$\text{pK}_a = -\log K_a$$
- For example a K<sub>a</sub> of 1x10<sup>-11</sup> becomes a pK<sub>a</sub> of 11



## pH and pOH

- In Level 2 Chemistry we also saw that K<sub>w</sub> (the dissociation constant for water) could be used to calculate the concentrations of hydronium and hydroxide in a solution:
 
$$\begin{aligned} K_w &= [\text{H}_3\text{O}^+][\text{OH}^-] \\ &= 1 \times 10^{-14} \end{aligned}$$
- And the shortcut for calculating the concentrations is to say:
 
$$14 = \text{pH} + \text{pOH}$$
 Then we convert the pH or pOH into concentrations e.g.
  - pH 4 means a pOH of 10
  - pH 4 means [H<sub>3</sub>O<sup>+</sup>] = 1x10<sup>-4</sup> molL<sup>-1</sup>
  - pOH 10 means [OH<sup>-</sup>] = 1x10<sup>-10</sup> molL<sup>-1</sup>



## K<sub>a</sub>, K<sub>b</sub> and K<sub>w</sub>

- K<sub>b</sub> (the dissociation of a base constant) is a new term for this year. It is derived from the following expression:

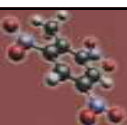
$$K_b = \frac{[HA_{(aq)}][OH_{(aq)}^-]}{[A_{(aq)}^-]}$$

- This is where A<sup>-</sup> is the reactant (a base) and it is producing OH<sup>-</sup> in water.
- There is another way of calculating K<sub>b</sub>, by using the following formulae:

$$K_w = K_a \times K_b \quad \text{rearranged to form...}$$

$$K_b = K_w / K_a$$

Note: They don't give you formulae in the Chemistry Exams

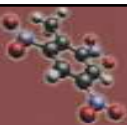


## Buffers

- Buffers are solutions that 'resist' the addition of acids and bases. Even though we are adding the acid/base the pH doesn't alter to the degree that it should.
- This is because buffers are made from a mixture of:
  - A weak acid and its conjugate base (in the form of a salt) or...
  - A weak base and its conjugate acid
- What we end up getting is a balanced equilibrium between the acid and base pair and according to Le Chatelier's principle, and change will be resisted e.g.

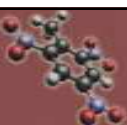
$$CH_3COOH + A^- \rightleftharpoons CH_3COO^- + HA$$

Adding base (A<sup>-</sup>)
Adding acid (HA)



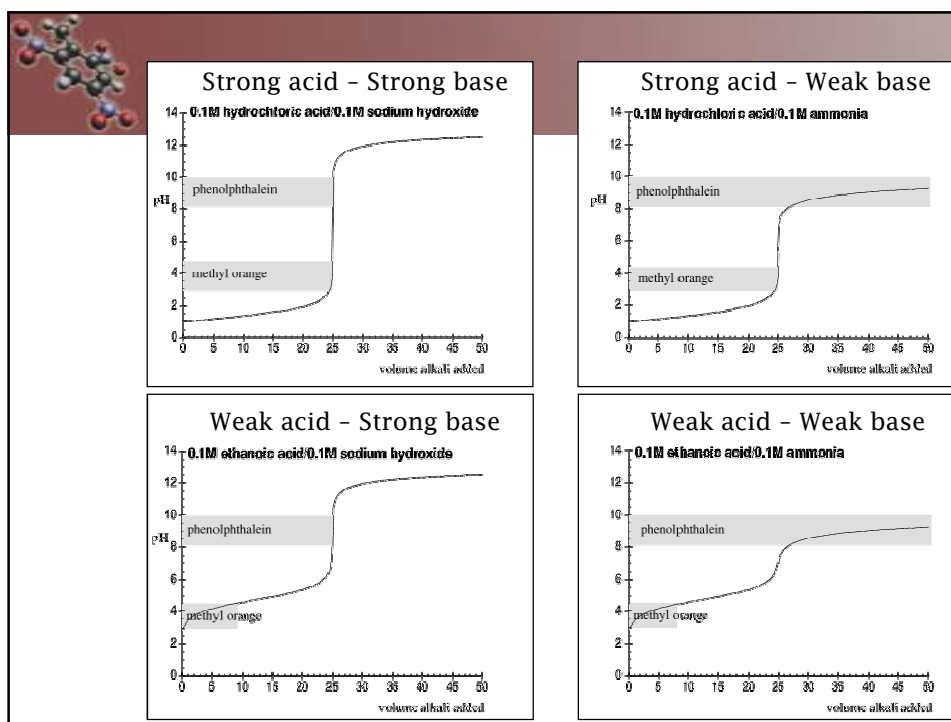
## Calculating Buffer pH

- The normal expression for the  $K_a$  of an acid is:
 
$$K_a = \frac{[H_3O_{(aq)}^+][A_{(aq)}^-]}{[HA_{(aq)}]}$$
- When we rearrange this equation we get this:
 
$$[H_3O_{(aq)}^+] = \frac{K_a [HA_{(aq)}]}{[A_{(aq)}^-]}$$
- As we can see this 'new' equation has the formula built in for the calculation of  $[H_3O^+]$  and therefore pH.
- By the way, the  $[HA]$  is the acid concentration, the  $A^-$  is the base concentration and the  $K_a$  is the constant for the type of acid you used.



## Titration Curves

- During an acid - base titration we are adding an acid to a base (strangely enough). As we do this the base neutralises the acid, and the pH changes.
- If we use a weak acid, we get increasing amounts of its conjugate acid being produced and we create our own buffer, where we add more acid but the pH doesn't change a lot.
- If we plot these pH changes we have a titration curve. Titration curves look very different depending on the types of chemicals we are adding together e.g.
  - Strong acid to Strong base
  - Strong acid to weak base
  - Weak acid to strong base
  - Weak acid to weak base
- Here are examples of the curves:



## Titration Curves – Characteristics and Keywords

- Buffer Zones – where acid/base is being added but there is no major shift in pH – these occur half way to the equivalence points on the weak acid/strong base or strong acid/weak base curves – they also occur on the weak side (i.e. the weak **base** side or the weak **acid** side)
- Endpoint – Where the acid/base has been neutralised
- Equivalence point – Where balanced amounts (moles) of the acid and base have been added together

## Choosing an Indicator

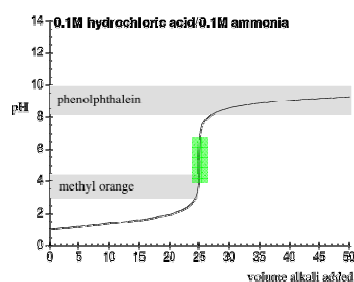
- Indicators are chemicals that change colour in acidic and alkaline conditions. They are also weak acids or bases and have their own  $K_a$  values.
- Each of the examples of titration curves have different pH range for an equivalence point. Therefore we need different indicators that match that pH range.
- To choose the right indicator we need to make sure that the  $pK_a$  of the indicator matches the pH range of the titration's equivalence points (see Chart on next page)

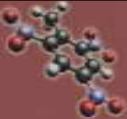
## Indicator Chart

Equivalence Point	Indicator	Indicators $pK_a$
pH 3-5	Methyl Orange	4
pH 4-6	Methyl Red	5
pH 6-8	Bromothymol Blue	7
pH 8-10	Phenolphthalein	9
pH 10-12	Cresyl Blue	11

### Example - How to use the chart:

- Find the pH range of the titration's equivalence zone (see the green box)
- Find the indicator whose  $pK_a$  falls within this range





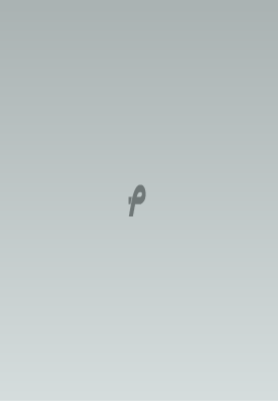
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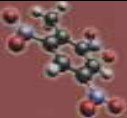
- One
- Three
- Four

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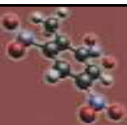
- One
- Three

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- One
- Two a)
- Three and Four

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