

Definitions and Concepts for Edexcel Chemistry A-level

Topic 12: Acid-Base Equilibria

Bronsted-Lowry Acid: A proton donor, e.g. $\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-$.

Bronsted-Lowry Base: A proton acceptor, e.g. $\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^-$.

Conjugate acid/base pair: Two species that differ by H^+ . We can have an acid and its conjugate base (e.g. H_3O^+ and H_2O), or a base and its conjugate acid (e.g. NH_3 , NH_4^+).

pH: A figure expressing the acidity or alkalinity of a solution on a logarithmic scale on which 7 is neutral, lower values are more acidic and higher values more alkaline.

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

K_a: The acidic dissociation constant, $\text{K}_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$. Taking the negative logarithm and rearranging for pH gives: $\text{pH} = \text{pK}_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$

pK_a = $-\log_{10}[\text{K}_a]$, the lower the value, the stronger the acid (as equilibrium is shifted more towards ionised products)

K_w: The ionic product of water, $\text{K}_w = [\text{H}^+][\text{OH}^-]$, at 298K, $\text{K}_w = 1.0 \times 10^{-14} \text{ mol}^2\text{dm}^{-6}$

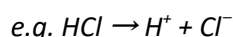
pK_w = $-\log_{10}[\text{K}_w]$, and so $\text{pH} + \text{pOH} = 14$ at 298 K

Lewis Acid: Electron pair acceptor.

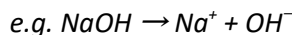
Lewis Base: Electron pair donor.

e.g. NH_3 (Lewis base, donates a free electron pair) + BH_3 (Lewis acid, has no electron pairs, accepts an electron pair from ammonia) $\rightarrow \text{H}_3\text{N} \cdots \text{BH}_3$

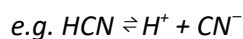
Strong acid: Acid which completely dissociates in water.



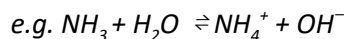
Strong base: Base which completely dissociates in water.



Weak acid: Acid which dissociates only slightly in water (reversible reaction).



Weak base: Base which is only slightly protonated in water (reversible reaction).



pH of a weak acid: Can be calculated from the K_a . We define K_a as:

$$\text{K}_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Then, if the H^+ produced by dissociation of water ($\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$) are negligible, we can assume:

$$[\text{H}^+] = [\text{A}^-]$$



In addition, we can postulate that the acid is weak, so the concentration of undissociated acid molecules, $[HA]$, will be approximately the same as the initial acid concentration, c . These two assumptions lead to:

$$K_a = [H^+]^2/c$$

$$\text{from this, } [H^+] = (K_a \times c)^{0.5}$$

$$\text{And } pH = -0.5 \log(K_a \times c).$$

Amphoteric: Refers to a substance that can act as an acid or a base, e.g. HCO_3^- ; it can accept a proton and form H_2O and CO_2 , or donate a proton and form CO_3^{2-} .

Monoprotic acid: Can release only one H^+ upon dissociation, e.g. HCl .

Polyprotic acid: Can release more than one H^+ upon dissociation, e.g. H_2SO_4 .

Equivalence point: The point when full neutralisation occurs, e.g. when titrating an acid with a base, it is the point when all acid has been neutralised, and so $[H^+] = [OH^-]$.

End point: The point during the titration when the indicator changes the colour. A suitable indicator should change colour near the equivalence point (so should have a pH range within the pH change during the equivalence point).

Buffer: A solution which resists change in pH when small amounts of strong acid/base are added.

Acidic buffer: A buffer containing a weak acid and its conjugate base, e.g. a solution of acetic acid and sodium acetate.

Alkaline buffer: A buffer containing a weak base and its conjugate acid, e.g. a solution of ammonia and ammonium chloride.

Preparation of buffers: E.g. when preparing an acetate buffer (CH_3COOH/CH_3COONa), one can either add a suitable amount of salt to the suitable solution of the acid, or react an acid with a strong base as a limiting reagent.

