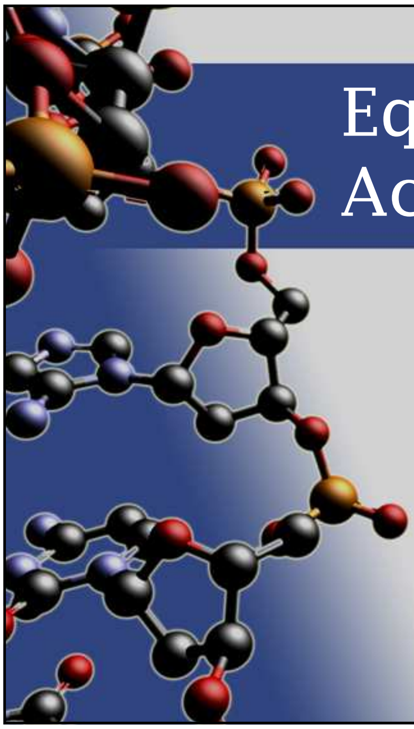

Equilibrium – Acids and Bases

Definition – Acids and Bases
The Dissociation of Acids (K_a)
The Dissociation of Water (K_w)
Conjugate Acids and Bases
Using K_w
Calculating pH

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Acids and Bases

- We have used acids and bases for a number of years now. We know that they:
  - Turn litmus and universal indicator red (acids) and blue (bases)
  - They produce hydrogen gas with reactive metals
  - They produce carbon dioxide with carbonates
  - They have an H (hydrogen) in there chemical formula
- Now we have a more complex set of definitions:
  - Acids are proton (H^+) donators (they give them away)
  - Bases are proton (H^+) acceptors (they take them)

Note: When H^+ comes in contact with water it forms an ion called hydronium (H_3O^+). So technically $H^+_{(aq)}$ doesn't exist.

Dissociation of Acids

- As mentioned earlier acids donate protons (H^+). This means that there must be a hydrogen available somewhere in the compound.
- When a compound dissociates, breaks apart. When an acid dissociates it breaks off an H^+. This happens when an acidic compound enters water and it ionises.
- There are different levels of dissociation:
  - Some compounds of hydrogen dissociate very easily in water, these are strong acids
  - Some do not dissociate readily, these are weak acids
  - Some do not dissociate, they are not acids

Acid Dissociation Constant (K_a)

- With weak acids the dissociation is an equilibrium situation:

$$HA \leftrightharpoons H^+ + A^-$$

- This means that while some of the HA is breaking up into H^+ and A^-, some of the reverse reaction is occurring (HA is being remade)
- If we always use a $1 molL^{-1}$ solution of acid then we can change or original formula for working out the equilibrium constant to this:

$$K_a = \frac{[H^+]\,[A^-]}{[HA]}$$

- This equilibrium constant is called an acid dissociation constant (or K_a for short)

Strong Acids – Weak Acids

- A strong acid as a high K_a value, a weak acid has a small K_a value.

Examples:

- HCl, H_2SO_4, HNO_3 – these are strong acids because they dissociate fully (there is none of the original compound intact)
- NH_4^+ - a weak acid because most of the NH_4^+ remains intact

- Note: 'Concentrated' and 'dilute' refer to the concentration of the acid. For example you can have a *dilute* strong acid or a *concentrated* weak acid.

The Dissociation of Water (K_w)

- Water also dissociates like and acid. This is what happens:

$$2H_2O_{(l)} \leftrightharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

- You may have noticed that it broke up into and acid (H^+) and a base (OH^-). These two components reacts again to make water.
- Because this equilibrium situation is used a lot it is given a special name and symbol:

$$K_w = [H_3O^+][OH^-]$$

- K_w is the Dissociation constant for water. It has a value of:

$$K_w = 1x10^{-14}$$

Conjugate Acids and Bases

- When a weak acid dissociates in water the piece that the H^+ breaks off still has the ability to bond again with the H^+. This make the fragment a proton acceptor, and therefore a base.
- This base is called a conjugate base. All weak acids can form conjugate bases e.g:

$$CH_3COOH \leftrightharpoons CH_3COO^- + H^+$$

Weak Acid — Conjugate Base

- Also, any weak base can have a conjugate acid:

$$NH_3 \leftrightharpoons NH_4^+$$

Weak Base — Conjugate Acid

What is pH?

- pH is a measure of the levels H_3O^+ in the solution. It is a log scale – like the Richter scale – every whole number is actually an change by a factor of **ten.**
- It is calculated using the following formula:

$$pH = -\log[H_3O^+_{(aq)}]$$

- In other words we:
  1. Find the H_3O^+ concentration
  2. Log it (using a calculator)
  3. Then change the + or – sign at the front of the number

Example I: Calculating pH

- Joan has a dilute acid solution. It has a Hydronium concentration of $1x10^{-11}molL^{-1}$. What is the pH?
  1. On my calculator I press the **log** button
  2. Now I type in $1x10^{-11}$ (**1** then **exp** then **11** then +/-)
  3. Press = then the +/- button again
- My answer is: pH = 11
- Notice how the answer is the exponent with the – sign changed? This shows us that for every increase by a pH of 1, we actually have a **decrease in acid** (by a factor of 10)

Example II: Calculating $[H_3O^+]$

- Joan now has an acid solution. It has a pH of 3.0. What is the $[H_3O^+]$?
  1. First I press the **2nd Fnct / inv / shft** button (one of those three will probably be there) then the **log** button
  2. I now type in **3.0** then the +/- button
  3. Press =
- My answer is: $[H_3O^+] = 1x10^{-3}molL^{-1}$
- Again, please note that the pH and the $[H_3O^+]$ are very closely related

Using Kw

- The water dissociation constant can be used to calculate the OH^- or H_3O^+ of a solution because now matter how much acid or base we add the following formula always works:

  $$1x10^{-14} = [H_3O^+][OH^-]$$

- It looks just like the K_w equation but when we change the concentration of acid or base in a solution it can no longer be called an equilibrium constant calculation (the conditions are no longer *standard*). This is now an 'Ionic Product' calculation.
- Using some maths we can do some substitution if we know one of the new acid/base concentrations.

Example: Using K_w

- Jeff has a solution with a hydronium concentration of $1.2x10^{-3}$ $molL^{-1}$. What is the hydroxide concentration of this solution?
  - Write down this: $1x10^{-14} = [H_3O^+][OH^-]$
  - Substitute in what we know: $[H_3O^+] = 1.2x10^{-3}$

    $$1x10^{-14} = 1.2x10^{-3} \text{ x } [OH^-]$$

  - Now rearrange the numbers to make $[OH^-]$ the subject of the equation. Remember the rule, change the side – change the sign

    $$[OH^-] = 1x10^{-14} / 1.2x10^{-3}$$

    $$[OH^-] = 8.3x10^{-12} \text{ molL}^{-1}$$

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