Acids and Bases II

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Info

Study Part

Attention: This study sheet is only a COMPLEMENT to the previous one!!! So: do not wonder if this one's rather small AND do not complain about "missing learning aims"! Thanks ;)
The other one can be found here http://j.mp/limenet-a_b_1

You can explain why pH-indicators change colour, if the pH-value changes in a certain range

Definition: pH-indicators are colored acids (bases) that can react with OH^- (H_30^+), thereby forming the conjugae base, which has a different color that the indicator acid (base).

You know the changes of the colour of cabbage sap in the pH range from 3 to 11

Cabbage sap (see picture above; "A gradient of red cabbage extract pH indicator from acidic solution on the left to basic on the right.") changes as follows:

pH	color
2 | acidic	red
4 | acidic	pink
6 | neutral	blue-violett
8 | basic	blue
10 | basic	blue-green
12 | basic	greenish-yellow

You are able to explain what buffers are and you can tell what the importance of buffers for living beings is

Definition

- A buffer softens the effect of adding acid or alkali – they resist sudden large pH-changes.
- The acid consumes any add base, and the base consumes any added acid.
- *Rule:* If the pH-value of a buffer solution is equal to its pK_A-value, the concentration of the buffer acid is equal to the concentration of its conjugate base. Buffer solution work only for weak acids and bases
- It always consists of a weak acid and its conjugate base *or* a weak base and its conjugate acid

Importance in biology

Certain processes in our body are very sensitive, such as our blood's pH, which has to remain at constant at 7.4 – a change of half a pH is lethal. Therefore strong buffers are in use to prevent this.

You can formulate the equations which are relevant for the function of a buffer solution

Example

$CH_3CO_2H + OH^- \rightarrow CH_3CO_2^- + H_2O$
$CH_3CO_2^- + H_3O^+ \rightarrow CH_3CO_2H + H_2O$

You can explain how we can make any buffer solution

Example: You want a buffer, which keeps the pH close to 7.21: $H_2PO_4^-$ has a pK_A is exactly at 7.21, its conjugate base is HPO_4^{2-}. This results in the following buffer:

$0.1\frac{mol}{l}(H_2PO_4^- + Na^+)\ and\ 0.1\frac{mol}{l}(HPO_4^{2-} + 2Na^+)$

How it buffers:

$against\ a\ base: H_2PO_4^- + OH^- \leftrightarrows HPO_4^{2-} + H_2O$
$against\ an\ acid: HPO_4^{2-} + H_3O^+ \leftrightarrows H_2PO_4^- + H_2O$

If the acid/base is an ion, a resp. other changed ion has to be added.

You can explain whether the pH-value is changed or not, if a salt is added to distilled water

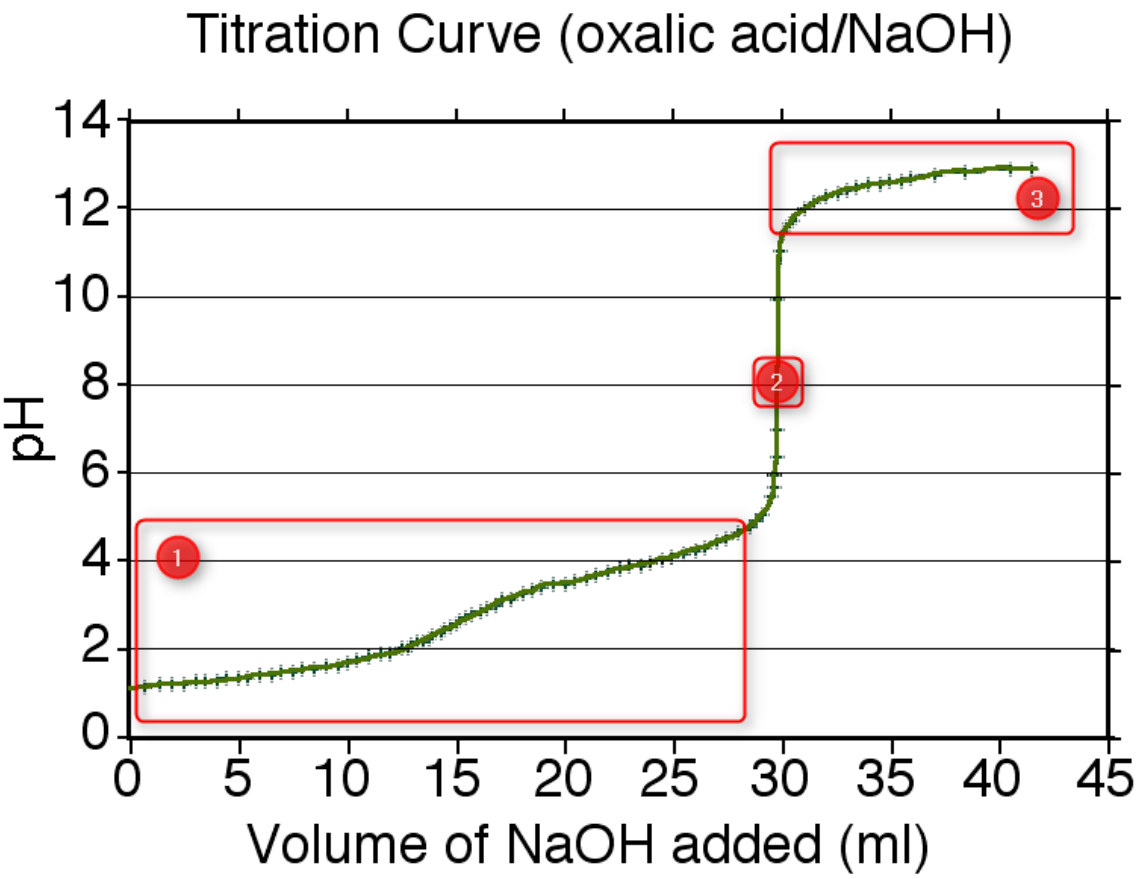
Yes, it is changed, as the ions react to form acids and bases, respectively.

$Example\ K_2S(aq) = 2K^+ + S^{2-} + H_2O$
$S^{2-} + H_2O \rightarrow HS^- + OH^-$
$HS^-is\ a\ medium\ base\ (equilibrium\ to\ the\ left)$

You can estimate the concentration of any particle in a solution of a certain amount of salt in water

$Example\ K_2S(aq) = 2K^+ + S^{2-} + H_2O$
$1.: H_2O, 2.: 0.2\frac{mol}{l}K^+, 3.: S^{2-}, 4.: OH^- = HS^-$

You can interpret the shape of the pH-curve (pH against amount of OH^-) and you can explain the expressions equivalence point and buffer region



1. **Buffer region** in this region, the mixture is buffered, thus not very seneitive to the addition of the $NaOH$
2. **Equivalence point** the buffer was "killed" by the added base and the pH drastically jumps
3. **Buffer region** the mixture is buffered again and approaches slowly its maximum pH.

You know what can be found out by an acid-alkali titration experiment

You can determine the concentration of an acid of a solution. *Example:* "How many vinegar molecules are in normal vinegar?"

$$concentration\ of\ an\ acid = [HA]_0\ (before\ it\ was\ added\ to\ water)$$

You can explain how an acid-alkali titration experiment is done and how we can calculate the amount of acid in a solution out of the result

Explanation

A base drops from a burette with a scale slowly into a constantly stirred solution, containing the acid and an indicator. Upon color change, the addition of the base is stopped and volume of added base is read off.

Calculation

1. Concentration of the acetic acid

$$[HA] = [OH^-] * \frac{V_{OH^-}}{V_{HA}} = 0.1\frac{mol}{l} * \frac{17ml}{\frac{1}{5} * 10ml = 2ml} = 0.85\frac{mol}{l}$$

2. Concentration of the pure vinegar in g/l
   $mass\ of\ CH_3CO_2H: 2 * C + 4 * H + 2 * 0 = 2 * 12 + 4 * 1 + 2 * 16 = 60u$
   $1mol = 60g$
   $0.85mol = x \rightarrow x = 51\frac{g}{l}$

You can explain how limestone is formed and how it is dissolved in water

It's an equilibrium which tries to form limestone, but is dissolved constantly again.

$CaCO_3 + H_2O + CO_2 \leftrightarrows Ca^{2+} + 2HCO_3^-$

It can be influenced…

- On the bottom of a pan, limestone is formed: reaction from right to left, because due to heat, as H_2O evaporates (and therefore "makes place" for more products) *or* the CO_2 comes out of the solution

Lab: detoxification of hydrogen-peroxide

- $2H_2O_2 \rightarrow 2H_2O + O_2(g)$
- The reaction is done with a catalyst to increase the surface, using "Katalase" or Fe^{3+}
- Higher concentrations lead to strongly exothermic (explosive) reactions
- The color change of the iron is because Fe^{4+} is formed as an intermediate (changed central ion in the complex)
- Exponential increase of reaction speed of the sodium-thiosulphate (using $\frac{1}{t}$) – linear increase of the reaction speed with respect to the concentration of the educts

Lab: Neutralization

Example $CuO = Cu^{2+} + O^{2-}$
$O^{2-} + HCl = O^{2-} + 2H_3O^+Cl^- \rightarrow H_2O + 2H_2O + Cl^-$