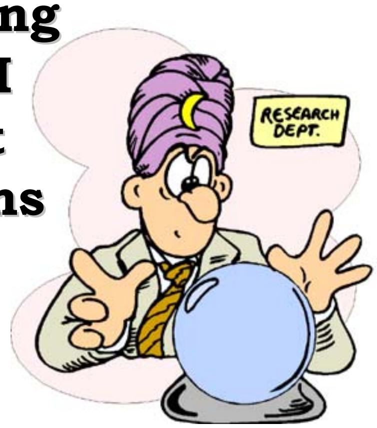
Predicting the pH of salt solutions



Hydrolysis of ions

Hydrolysis refers to a reaction with water (e.g. splitting water into H^+ and $OH^–$)

When salts are added to water, pH can change

E.g. when Na_3PO_4 is added to water, ions form

$Na_3PO_4(aq) \rightarrow 3Na^+(aq) + PO_4^{3–}(aq)$

These ions may react with H_2O, affecting the pH

$PO_4^{3–}(aq) + H^+(aq) \leftrightarrow HPO_4^{2–}(aq)$

$Na^+(aq) + OH^–(aq) \leftrightarrow NaOH\ (aq)$

If the anion (-ve) reacts to remove lots of H^+ but the cation (+ve) removes very little $OH^–$, then H^+ will decrease and the solution will be basic.

The degree of hydrolysis

$PO_4^{3–}(aq) + H^+(aq) \leftrightarrow HPO_4^{2–}(aq)$

$Na^+ + OH^– \leftrightarrow NaOH$

- The problem with writing equilibria this way is we do not know the strength of the reactions
- However, if we reverse the reaction we can look up Ka and Kb values (pg. 608, 615):

  $HPO_4^{2–} \leftrightarrow PO_4^{3–} + H^+$ Ka= 4.5 x $10^{–13}$

  $NaOH \leftrightarrow Na^+ + OH^–$ Kb= 55
- Small Ka: few products; adding $PO_4^{3–}$ = shift left
- Large Kb: mostly products; Na^+ has little affect
- Thus, adding Na_3PO_4 will cause more H^+ to be removed, resulting in a basic solution

Accuracy of predictions

Theoretically, using Ka and Kb values you could predict the exact pH resulting from a certain salt being added to distilled water.

However, you only need to be able to predict if a solution will be acidic, basic, or neutral.

Note: you can't judge the pH change solely on the difference between Ka and Kb. Other factors are involved (e.g. the formula of the compound and its molar mass both affect [ ])

Note: hydrolysis refers to reactions with water. Several variations for writing equilibriums exist. However, focusing on how the $H^+/OH^–$ balance of water is affected is easiest.

Steps in determining pH

1. Write the ions that form: e.g. NH_4CN
2. Determine the reaction ions have with water:
3. Look up the Ka of the conjugate acid and the Kb of the conjugate base:
4. Determine if more H^+ or $OH^–$ is removed:

Buffers - lab

Read 15.6 (621-623) up to and including special topic 15.2 (carbonate buffer)

Calibrate pH meter, get a plastic bottle with distilled H_2O to rinse your pH meter btw tests

You will use 4 solutions (≈20 mL of each):

distilled water, water + $NaC_2H_3O_2$ (5 scoops), 0.2 M $HC_2H_3O_2$, 0.2 M $HC_2H_3O_2$ + $NaC_2H_3O_2$

For each, record the initial pH and the pH upon addition of 5, 10, and 15 drops of 1 M HCl

Remake the 4 solutions

For each, record the initial pH and the pH upon addition of 5, 10, and 15 drops of 1 M NaOH

HCl	H_2O	$NaC_2H_3O_2$	$HC_2H_3O_2$	$NaC_2H_3O_2$ + $HC_2H_3O_2$
0				
5				
10				
15				
NaOH	H_2O	$NaC_2H_3O_2$	$HC_2H_3O_2$	$NaC_2H_3O_2$ + $HC_2H_3O_2$
0				
5				
10				
15				

Buffers - summary

Solutions with buffers resist changes in pH, when small amounts of acid or base are added

Buffers are important in blood, cells, resisting the effects of acid rain on lake ecosystems.

A buffer is created when a weak acid is mixed with a salt that contains the identical ion.

Two equilibria contribute to the consistent $[H^+]$

$$HA \leftrightarrow H^+ + A^–$$

$$Na^+ + A^– \leftrightarrow NaA$$