Chemistry 12

Acid-Base Equilibrium VI

Name:
Date:
Block:

1. Buffers

Buffers

An acid-base <u>buffer</u> is a solution that resists changes in pH following the addition of relatively small amounts of a strong acid or base.

Example: Consider a solution of 1.0 M acetic acid.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

- Acetic acid is a weak acid – only a small percent of the weak acid is ionized

	$CH_3COOH_{(aq)}$	+	$H_2O_{(l)}$	$\leftrightharpoons$	$CH_3COO^-_{(aq)}$	+	$H_3O^+_{(aq)}$
I	1.0M		––		0 M		0 M
C	– x		––		+ x		+ x
E	1.0M – x		––		x		x

$K_a =$

$[CH_3COOH]$ $[CH_3COO^-]$

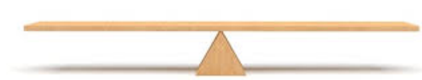
If a **strong base** was added to a solution, acetic acid will be there to neutralize the base.

If a **strong acid** was added, there would be no species to neutralize it.

In order for a buffer solution to be effective, EQUIVALENT CONCENTRATIONS of a weak acid and a conjugate base must be in solution.

A **regular** acetic acid solution…

An acetic acid **buffer** solution…



Regular CH_3COOH solution…	Buffer CH_3COOH solution…
$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$	$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

Consider the following pairs of solutions:

a) Circle the pairs of chemical species below that could be used to prepare a buffer solution.

b) For the pairs that you circled, write the buffer equation.

HNO_3 and $NaNO_3$	KF and HF	HNO_2 and HNO_3	HCOOH and LiHCOO
$NaHSO_4$ and Na_2SO_4	K_2CO_3 and $K_2C_2O_4$	HCl and NaCl	KH_2PO_4 and K_2HPO_4

(Acid) Buffer Equation:

$$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

Because …

Then …

Or in general…

The hydronium ion concentration (and therefore the pH) of a buffer solution depends on:

1. the K_a value.
2. the **ratio** of the concentration of the weak acid to its conjugate base.

Using our acetic acid example…

$$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

$[H_3O^+]$ =

pH =

Now, let's shift our system!

We now **add 0.10 mol HCl** to 1.0 M buffer solution with no volume change.

- Would we expect the pH to decrease or increase?
- Would the system shift left or right?

$$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

1.0 M 1.0 M

$[H_3O^+]$ =

pH =

We now **add 0.10 mol NaOH** to 1.0 M buffer solution with no volume change.

- Would we expect the pH to decrease or increase?
- Would the system shift left or right?

$$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

1.0 M 1.0 M

$[H_3O^+]$ =

pH =

Add 0.10M HCl Add 0.10M NaOH

<u>Unbuffered System</u>

pH =

Add 0.10M HCl Add 0.10M NaOH

<u>Buffer System</u>

pH =

Practice.

Consider a 1.0 M hydrofluoric acid buffer system.

a. What is the equation?

b. Calculate the pH of an undisturbed system.

c. Calculate the pH when 0.050M hydrobromic acid is added to the original buffer system.

d. Calculate the pH when 0.050M $Mg(OH)_2$ is added to the original buffer system.

Hebden Workbook Pg. 181 # 131-133, 136, 137