Quiz 17.3 – Weak Acid/Base Titrations

Name: Key

These questions concern titrating a solution of HNO_2 with NaOH. 25.00 ml of the HNO_2 solution with unknown concentration are placed in an Erlenmeyer flask, and a burette is filled with a 0.575 M solution of NaOH

Question 1 HNO_2 $K_A = 5.6\cdot10^{-4}$ $pK_A = 3.25$ $K_B = 1.786\cdot10^{-11}$

NaOH is added slowly while the pH is monitored. How will you know when the equivalence point of the titration has been reached?

point of steepest pH change

Question 2

The equivalence point is reached after 23.42 ml of the base have been added.
What was the initial acid concentration?

$C_A V_A = C_B V_B$ $C_A \cdot 25.00\,ml = 0.575\,M \cdot 23.42\,ml$

$C_A = 0.539\,M$

Question 3

Now that you know the initial concentration, calculate what the pH should have been before any base was added

$K_A = \frac{x^2}{[HA]_0 - x}$ $5.6\cdot10^{-4} = \frac{x^2}{0.539 - x}$ ↑ assume x is small

$x = \sqrt{5.6\cdot10^{-4}\cdot 0.539}$ $x = 0.0174$ (3.2%) $pH = -\log x = 1.76$

Question 4

What is the pH at the equivalence point, and which acid/base indicator would be appropriate for identifying the end-point in a titrations without a pH meter?

$0.575\,M \cdot 23.42\,ml = 13.47$ mmoles of NO_2^- @ equivalence

$\frac{13.47\,mmoles}{48.42\,ml} = 0.278\,M$

$1.786\cdot10^{-11} = \frac{x^2}{0.278 - x}$ ← definitely small

$x = 2.23\cdot10^{-6}$

$pH = 8.35$ Phenolphtholein

Question 5

What will the pH be after 11.71 ml of NaOH have been added?

↳ 6.735 mmoles

	HNO_2	$+OH^-$	→ NO_2^-
B	13.47	6.735	∅
C	-6.735	-6.735	+6.735
A	6.735	∅	6.735

$pH = pK_A + \log\frac{B}{A} = 3.25 + \log\frac{6.735}{6.735} = 3.25$

Question 6

What will the pH be after you have added 23.67 ml and 23.17 ml of the NaOH solution

	HNO_2	$+OH^-$	→ NO_2^-
B	13.47	13.61	∅
C	-13.47	-13.47	+13.47
A	∅	0.14	13.47

↳ 13.61 mmol

$[OH^-] = \frac{0.14\,mmol}{48.67\,ml}$ $[OH^-] = 2.88\cdot10^{-3}\,M$ → $pH = 11.46$

↳ 13.32 mmol

	HNO_2	$+OH^-$	→ NO_2^-
B	13.47	13.32	∅
C	-13.32	-13.32	+13.32
	0.15	∅	13.32
	48.17 ml → 0.00311 M		48.17 ml → 0.277 M

$K_A = \frac{x([A^-]_0 + x)}{[HA]_0 - x}$ $5.6\cdot10^{-4} = \frac{x(0.277 + x)}{0.00311 - x}$

$x = 6.29\cdot10^{-6}$ → $pH = 5.20$

Question 7

On the back of this page, sketch the titration curve, noting the pH at the most important points