Review: Acid-Base Chemistry

1. Plot a pH curve for the titration of 20.0 mL of 0.400 M nitric acid with 0.200 M potassium hydroxide. Calculate the coordinates of sufficient points on the curve to draw an accurate pH curve (minimum of 10 points).

$HNO_3 + KOH \rightarrow KNO_3 + H_2O$

endpoint $\frac{C_aV_a}{R_a} = \frac{C_bV_b}{R_b} \quad \therefore \quad Vb = \frac{C_aV_aR_b}{C_bR_a}$

$$V_b = \frac{(0.400\ M)(0.0200\ L)(1)}{(0.200\ M)(1)} = 0.0400\ L \text{ or } 40.0\ mL$$

Volume of base	pH
0 mL	0.398
30 mL	1.40
39 mL	2.47
39.5 mL	2.77
39.9 mL	3.48
39.99 mL	4.48
40 mL	7
40.01 mL	9.52
40.1 mL	10.5
40.5 mL	11.2
41 mL	11.5
50 mL	12.5
80 mL	12.9

0 mL of base added

$[H^+] = 0.400\ M \quad \therefore \quad pH = -\log(0.400) = \mathbf{0.398}$

30 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0300 L) = <u>0.00600 mol OH^-</u>
excess = 0.00200 mol H^+

$[H^+] = 0.00200\ mol \div 0.0500\ L = 0.0400\ M$

$\therefore \quad pH = -\log(0.0400) = \mathbf{1.40}$

39 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0390 L) = <u>0.00780 mol OH^-</u>
excess = 0.00020 mol H^+

$[H^+] = 0.00020\ mol \div 0.0590\ L = 0.00339\ M$

$\therefore \quad pH = -\log(0.00339) = \mathbf{2.47}$

39.5 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0395 L) = <u>0.00790 mol OH^-</u>
excess = 0.00010 mol H^+

$[H^+] = 0.00010\ mol \div 0.0595\ L = 0.00168\ M$

$\therefore \quad pH = -\log(0.00168) = \mathbf{2.77}$

39.9 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0399 L) = <u>0.00798 mol OH^-</u>
excess = 0.00002 mol H^+

$[H^+]$ = 0.00002 mol ÷ 0.0599 L = 0.000334 M

∴ pH = -log(0.000334) = **3.48**

39.99 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.03999 L) = <u>0.007998 mol OH^-</u>
excess = 0.000002 mol H^+

$[H^+]$ = 0.000002 mol ÷ 0.05999 L = 0.0000333 M

∴ pH = -log(0.0000333) = **4.48**

40.01 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.04001 L) = <u>0.008002 mol OH^-</u>
excess = 0.000002 mol OH^-

$[OH^-]$ = 0.000002 mol ÷ 0.06001 L = 0.0000333 M

∴ pOH = -log(0.0000333) = 4.48 & pH = 14.00 – 4.48 = **9.52**

40.1 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0401 L) = <u>0.00802 mol OH^-</u>
excess = 0.00002 mol OH^-

$[OH^-]$ = 0.00002 mol ÷ 0.0601 L = 0.000333 M

∴ pOH = -log(0.000333) = 3.48 & pH = 14.00 – 3.48 = **10.5**

40.5 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0405 L) = <u>0.00810 mol OH^-</u>
excess = 0.00010 mol OH^-

$[OH^-]$ = 0.00010 mol ÷ 0.0605 L = 0.00165 M

∴ pOH = -log(0.00165) = 2.78 & pH = 14.00 – 2.78 = **11.2**

41.0 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0410 L) = <u>0.00820 mol OH^-</u>
excess = 0.00020 mol OH^-

$[OH^-]$ = 0.00002 mol ÷ 0.0610 L = 0.00328 M

∴ pOH = -log(0.00328) = 2.48 & pH = 14.00 – 2.48 = **11.5**

50.0 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0500 L) = <u>0.0100 mol OH^-</u>
excess = 0.00200 mol OH^-

$[OH^-]$ = 0.00200 mol ÷ 0.0700 L = 0.0286 M

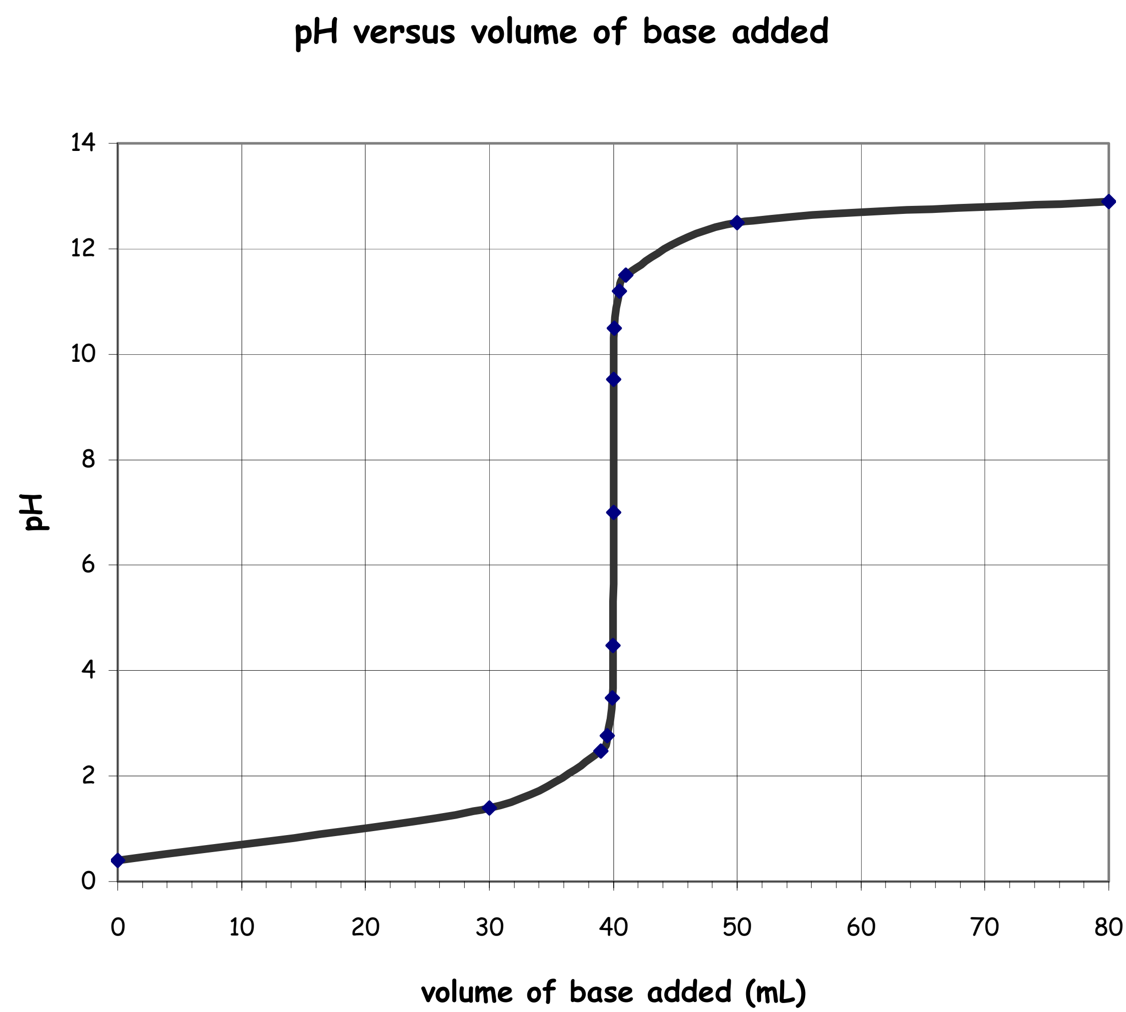
$\therefore$ pOH = -log(0.0286) = 1.54 & pH = 14.00 – 1.54 = **12.5**

80.0 mL of base added

mol H^+ = (0.400 M)(0.0200 L) = 0.00800 mol H^+
mol OH^- = (0.200 M)(0.0800 L) = <u>0.0160 mol OH^-</u>
excess = 0.00800 mol OH^-

$[OH^-]$ = 0.00800 mol ÷ 0.100 L = 0.0800 M

$\therefore$ pOH = -log(0.0800) = 1.10 & pH = 14.00 – 1.10 = **12.9**



2. Identify the acid, the base, the conjugate acid, and the conjugate base in each of the following reactions.
   a) $HC_2H_3O_2^- + F^- \leftrightarrows C_2H_3O_2^- + HF$

      Acid base conjugate Base conjugate acid

   b) $HCO_3^- + NH_3 \leftrightarrows CO_3^{2-} + NH_4^+$

      Acid base conjugate Base conjugate acid

   c) $HSO_4^- + H_2O \leftrightarrows H_3O^+ + SO_4^{2-}$

      Acid base conjugate acid conjugate base

3. Aqueous solutions of nitric acid and nitrous acid of the same concentration are prepared.
   a) How do their pH values compare?

      The pH of nitric acid is lower

   b) Explain your answer using equations.

      $HNO_3 + H_2O \rightarrow H_3O^+ \; NO_3^-$ (complete ionization)
      $HNO_2 + H_2O \leftrightarrows H_3O^+ + NO_2^-$ (partial ionization)

      Because the HNO_2 is only partially ionized, the concentration of H^+ is less in HNO_2 than in HNO_3.

4. State two examples of conjugate acid-base pairs, each involving the hydrogen sulfite ion.

   HSO_3^-/SO_3^{2-}, and H_2SO_3/HSO_3^-

5. If the pH of a solution is 6.8, what is the colour of each of the following indicators in this solution?
   a) methyl red ➔ yellow
   b) chlorophenol red ➔ red
   c) bromothymol blue ➔ green
   d) phenolphthalein ➔ colourless
   e) methyl orange ➔ yellow
   f) alizarin yellow ➔ yellow

6. PH curves provide information about acid-base reaction systems.
   a) What is a buffering action?

      A relatively constant pH when small amount of acid or base are added

   b) Where does buffering action appear on a pH curve?

      Buffering action is most noticeable when the volume of the titrant is half of what is required for the equivalence point, or halfway between successive equivalence points

   c) How are quantitative reactions represented on a pH curve?

      At the midpoint of the vertical section

   d) Define pH endpoint and equivalence point.

      pH endpoint ➔ midpoint of the vertical section
      Equivalence point ➔ quantity of titrant at the endpoint

   e) How is a suitable indicator chosen for a titration?

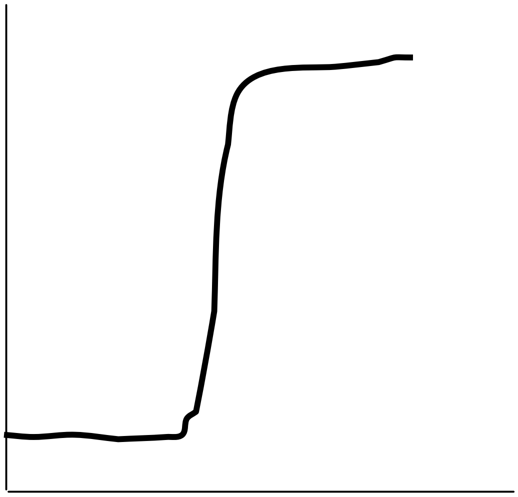
      The midpoint of the indicator change should equal the pH of the endpoint, and the colour change of the indicator should be complete within the pH range of the vertical section of the pH curve.

   f) Do non-quantitative reactions have an endpoint? Explain your answer briefly.

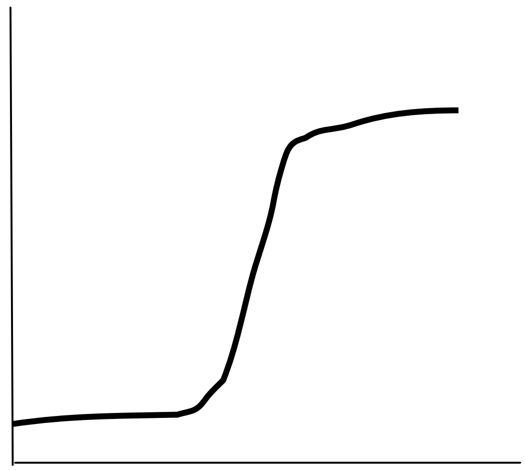
      Nonquantitative reactions do not have distinct endpoints because pH changes gradually in the region where the equivalence point is reached.

7. Sketch a pH curve for the following:

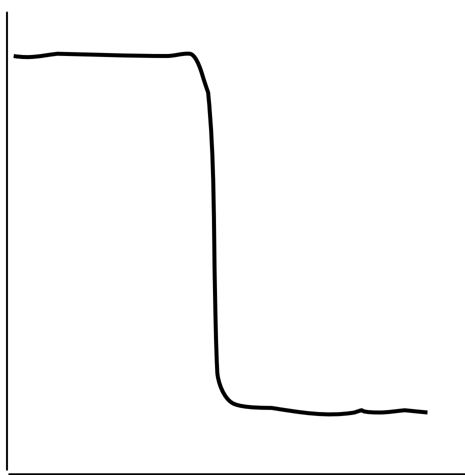
a) A strong acid titrated with a strong base.



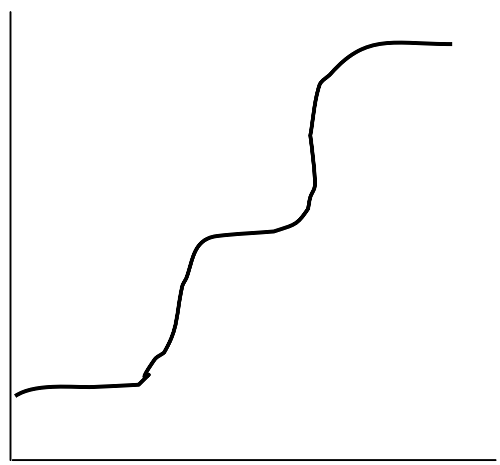
d) A strong acid titrated with a weak base.



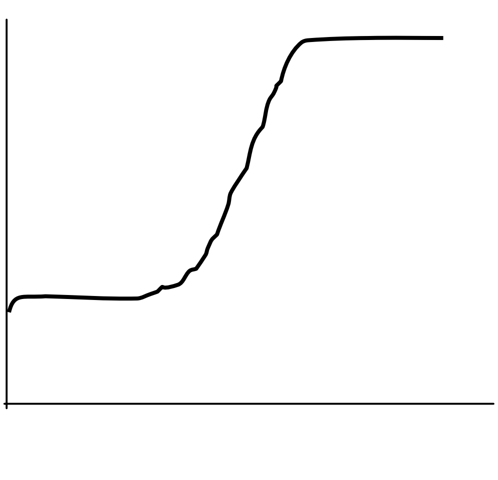
b) A strong base titrated with a strong acid.



e) A diprotic acid titrated with a strong base.



c) A weak acid titrated with a strong base.



8. Write two experimental designs to rank a group of bases in order of strength.

Solutions of equal concentration of several bases are prepared and the pH is measured (using pH paper or a pH meter) for each solution. The higher the pH of the solution, the stronger the base. The manipulated variable is the base, and the responding variable is the pH. The controlled variables are the temperature and the concentration.

Solutions of equal concentration of several bases are prepared and the electrical conductivity of the aqueous solutions are measured. The higher the conductivity, the stronger the base. The manipulated variable is the base, and the responding variable is the electrical conductivity. The controlled variables are the temperature and the concentration.

9. Separate samples of an unknown solution were tested with indicators. Thymol blue was yellow and bromothymol blue was blue. Estimate the approximate pH and hydronium ion concentration of the solution.

Thymol Blue	➔ pH range = 1.2 – 2.8, red - yellow	➔ pH > 2.8
	➔ pH range = 8.0 – 9.6, yellow - blue	➔ pH < 8.0
Bromothymol Blue	➔ pH range = 6.0 – 7.6, yellow - blue	➔ pH > 7.6

The pH of the solution must be greater than 7.6 and less than 8.0, therefore an approximate pH would be 7.8, which corresponds to a hydronium ion concentration of 1.6×10^{-8} mol/L.

10. In an experimental investigation of amphiprotic substances, samples of baking soda were added to a solution of sodium hydroxide and to a solution of hydrochloric acid. The pH of the sodium hydroxide changed from 13.0 to 9.5 after the addition of the baking soda. The pH of the hydrochloric acid changed form 1.0 to 4.5 after the addition of baking soda. Provide a theoretical explanation of these results by writing chemical equations to describe the reactions.

The bicarbonate ion has the ability to donate a proton and can act as an acid:

$NaOH \rightarrow Na^{+} + OH^{-}$

Add baking soda ($NaHCO_3 \rightarrow Na^{+} + HCO_3^{-}$)

The bicarbonate ion neutralizes some of the base, therefore the pH decreases

$OH- + HCO_3^{-} \leftrightarrows H_2O + CO_3^{2-}$

The bicarbonate ion has the ability to accept a proton and can act as a base:

$HCl + H_2O \rightarrow H_3O^{+} + Cl^{-}$

Add baking soda ($NaHCO_3 \rightarrow Na^{+} + HCO_3^{-}$)

The bicarbonate ion neutralizes some of the acid, therefore the pH increases.

$H_3O^{+} + HCO_3^{-} \leftrightarrows H_2O + H_2CO_3$

11. Each of seven unlabelled beakers was known to contain one of the following 0.10 mol/L solutions: $HCl(aq)$, $CH_3COOH(aq)$, $Ba(OH)_2(aq)$, $NH_3(aq)$, $C_2H_4(OH)_2(aq)$, $H_2SO_4(aq)$, and $NaOH(aq)$. Describe diagnostic test(s) required to distinguish the solutions and label the beakers. Use the "*If* ____, *and* ____, *then* ____" format, a flow chart, or a table to communicate your answer.

If the solutions are tested with a pH meter, ***and*** the pH values are ordered from smallest to largest, ***then*** the solutions are sulfuric acid, hydrochloric acid, acetic acid, ethanediol, ammonia, sodium hydroxide, and barium hydroxide, respectively.

OR

Test with litmus paper

blue litmus turns red = acidic → test with pH paper
- lowest pH → H_2SO_4
- middle pH → HCl
- higher pH → CH_3COOH

no change with blue or red = neutral → $= C_2H_4(OH)_2$ alcohol

red litmus turns blue = basic → test with pH paper
- lowest pH → NH_3
- middle pH → $NaOH$
- highest pH → $Ba(OH)_2$

12. A 25.0 mL sample of diluted rust-removing solution containing phosphoric acid was titrated to the second endpoint using 1.50 mol/L sodium hydroxide. The average equivalence point of the sodium hydroxide solution was 17.9 mL. What is the concentration of phosphoric acid in the rust-removing solution?

$$H_3PO_4 + 2\,NaOH \rightarrow Na_2HPO_4 + 2\,H_2O$$

$$C_a = \frac{C_b V_b R_a}{V_a R_b} \qquad C_a = \frac{(1.50\text{ M})(17.9\text{ mL})(1)}{(25.0\text{ mL})(2)} = 0.537\text{ M } H_3PO_4$$

13. A 50.0 mL volume of 0.560 mol/L hydrochloric acid was spilled on a counter. A student quickly decided to sprinkle calcium hydroxide onto the spill. If 1.00 g of solid calcium hydroxide was used, would it completely neutralize the acid? Justify your answer with calculations.

$$2\ HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2\ H_2O$$

$$(0.0500\ L)(0.560\ mol/L) = 0.0280\ mol\ HCl$$

$$(0.0280\ mol\ HCl)\left[\frac{1\ mol\ Ca(OH)_2}{2\ mol\ HCl}\right](74.09268\ g/mol) = 1.04\ g\ Ca(OH)_2$$

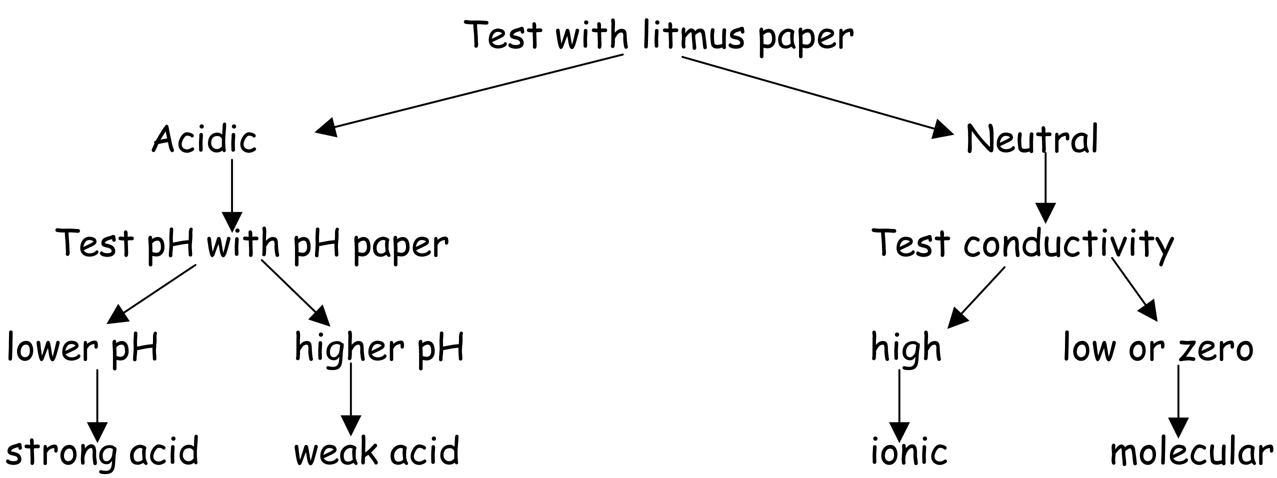
14. Acid rain has a pH less than that of normal rain. The presence of dissolved carbon dioxide, which forms carbonic acid, gives normal rain a pH of 5.6. What is the hydrogen ion concentration in normal rain?

$$pH = 5.6 \quad \therefore \quad [H^+] = antilog(-5.6) = 2.5 \times 10^{-6}\ mol/L$$

15. If the pH of a solution changes by 3 pH units as a result of adding a weak acid, by how much does the hydrogen ion concentration change?

pH changes by 3, therefore $[H^+]$ changes by 10^3 or 10^{-3}

16. Write an experimental design for the identification of four colourless solutions: a strong acid solution, a weak acid solution, a neutral molecular solution, and a neutral ionic solution. Write sentences, create a flow chart, or design a table to describe the required diagnostic tests.



17. Salicylic acid, $C_6H_4OHCOOH$, is an active ingredient of solutions, such as Clearasil®, that are used to treat acne. Since the K_a for this acid was not listed in any convenient references, a student tried to determine the value experimentally. If the pH of a saturated (1.00 g/460.0 mL) solution of salicylic acid was found to be 2.4 at 25°C, calculate the ionization constant for this acid.

1.00 g ÷ 138.122 84 g/mol = 0.007 24 mol $C_6H_4OHCOOH$

0.007 24 mol ÷ 0.4600 L = 0.0157 mol/L

pH = 2.4 ∴ $[H^+]$ = antilog(-2.4) = 0.003 98 mol/L

	$C_6H_4OHCOOH$ + H_2O	⇆	$C_6H_4OHCOO^-$	+ H_3O^+
Initial	0.0157		0	0
Δ	- x		+ x	+ x
equilibrium	0.0157 – 0.003 98		0.003 98	0.003 98

$$K_a = \frac{(0.003\ 98)(0.003\ 98)}{(0.0157 - 0.003\ 98)} = 0.001\ 35 \text{ or } 1.35 \times 10^{-3}$$

18. A 0.10 mol/L solution of lactic acid, found in sour mild, has a pH of 2.43. Calculate the percent ionization of lactic acid in water.

$HC_3H_5O_3 + H_2O \leftrightarrows H_3O^+ + C_3H_5O_3^-$

pH = 2.43 ∴ $[H^+]$ = antilog(-2.43) = 3.72×10^{-3} mol/L

$$\text{\% ionization} = \frac{\text{amount ionized}}{\text{original amount of acid}} \times 100$$

$$\text{\% ionization} = \frac{3.72 \times 10^{-3}\ \text{mol/L}}{0.10\ \text{mol/L}} \times 100 = 3.7\%$$

19. Ascorbic acid is the chemical ingredient of Vitamin C. A student prepares a 0.20 mol/L aqueous solution of ascorbic acid, measures its pH, and finds it to be 2.40. Based on this evidence, what is the K_a for ascorbic acid?

pH = 2.40 $\therefore$ $[H^+]$ = antilog(-2.40) = 3.98×10^{-3} mol/L

Let ascorbic acid = HA

	HA	+	H_2O	$\leftrightarrows$	H_3O^+	+	A^-
Initial	0.20				0		0
Δ	- x				+ x		+ x
equilibrium	0.20 - x				x		x
equilibrium	0.20 – 0.003 98				0.003 98		0.003 98

$$K_a = \frac{(0.003\ 98)(0.003\ 98)}{(0.20 - 0.003\ 98)} = 8.1 \times 10^{-5}$$

20. A solution was made up to be 0.100 M in chloroacetic acid ($HC_2H_2O_2Cl$) and also 0.002 00 M sodium chloroacetate ($NaC_2H_2O_2Cl$). The Ka for chloroacetic acid is 1.36×10^{-3}. Determine the [H^+] and the pH of the solution.

	$NaC_2H_2O_2Cl \rightarrow$	Na^+ +	$C_2H_3O_2Cl^-$
[initial]	0.0.00200	0	0
[final]	0	0.00200	0.00200

	$HC_2H_2O_2Cl$	$\leftrightarrows$	H^+ +	$C_2H_2O_2Cl^-$
[initial]	0.100		0	0.00200
Δ	– x		+ x	+ x
[equilibrium]	0.100 – x		x	0.00200 + x

$$Ka = \frac{[H^+][C_2H_2O_2Cl^-]}{[HC_2H_2O_2Cl]}$$

$$1.36 \cdot 10^{-3} = \frac{(x)(0.00200 + x)}{0.100 - x}$$

$$1.36 \cdot 10^{-3}(.100 - x) = 0.00200x + x^2$$

$$1.36 \cdot 10^{-4} - 1.36 \cdot 10^{-3}x = 0.00200x + x^2$$

$$0 = x^2 + 0.00336x - 1.36 \cdot 10^{-4}4$$

$$x = \frac{-0.00336 \pm \sqrt{(0.00336)^2 - (4)(1)(-1.36 \cdot 10^{-4})}}{(2)(1)}$$

$$x = \frac{-0.00336 \quad \pm \sqrt{0.0005552896}}{2}$$

$$x = \frac{-0.00336 \quad \pm 0.02335645836}{2}$$

$$x = -0.0135 \quad \text{or} \quad 0.0101$$

$$[H^+] = 0.0101 \ M$$

$$pH = -\log(0.0101) = 2.00$$

21. The hydroxide ion concentration in a 0.157 mol/L solution of sodium propanoate, $NaC_2H_5COO(aq)$, is found to be 1.1×10^{-5} mol/L. Calculate the base ionization constant for the propanoate ion.

$NaC_2H_5COO \leftrightarrows Na^+ + C_2H_5COO^-$ ($C_2H_5COO^-$ = propanoate ion)

	$C_2H_5COO^-$	+ H_2O $\leftrightarrows$	C_2H_5COOH	+ OH^-
Initial	0.157		0	0
Δ	- x		+ x	+ x
equilibrium	0.157 - x		x	x
equilibrium	$0.157 - 1.1 \times 10^{-5}$		1.1×10^{-5}	1.1×10^{-5}

$$K_b = \frac{(1.1 \times 10^{-5})(1.1 \times 10^{-5})}{(0.157 - 1.1 \times 10^{-5})} = 7.7 \times 10^{-10}$$

22. Aniline, $C_6H_5NH_2$, is closely related to ammonia and is also a weak base. If the pH of a 0.10 mol/L aniline solution was found to be 8.81, what is its K_b?

pH = 8.81 $\therefore$ pOH = 5.19 $\therefore$ [OH-] = antilog(-5.19) = 6.46×10^{-6} mol/L

	$C_6H_5NH_2$	+ H_2O $\leftrightarrows$	$C_6H_5NH_3^+$	+ OH^-
Initial	0.10		0	0
Δ	- x		+ x	+ x
equilibrium	0.10 - x		x	x
equilibrium	$0.10 - 6.46 \times 10^{-6}$		6.46×10^{-6}	6.46×10^{-6}

$$K_b = \frac{(6.46 \times 10^{-6})(6.46 \times 10^{-6})}{(0.10 - 6.46 \times 10^{-6})} = 4.2 \times 10^{-10}$$

23. Codeine has a K_b of 1.73×10^{-6}. Calculate the pH of a 0.020 mol/L codeine solution.

	Codeine	+ H_2O	$\leftrightarrows$	$CodeineH^+$	+	OH^-
Initial	0.020			0		0
Δ	- x			+ x		+ x
equilibrium	0.020 - x			x		x

$$1.73 \times 10^{-6} = \frac{(x)(x)}{(0.020 - x)} \qquad \therefore \qquad x = [OH^-] = 1.9 \times 10^{-4}$$

$pOH = -\log(1.9 \times 10^{-4}) = 3.7$

$pH = 14 - 3.7 = 10.3$