Acids and Bases - Answers

1. It requires 37.0 mL of 0.0500 mol/L NaOH(aq) to neutralize 100.0 mL of gastric juice. We can assume that HCl(aq) is the only acid present in gastric juice.
   a) Calculate the concentration of HCl(aq) present in gastric juice.

   $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

   $$C_a = \frac{C_b V_b R_a}{V_a R_b} = \frac{(0.0500 \text{ mol/L})(0.0370 \text{ L})(1)}{(0.1000 \text{ L})(1)} = 0.0185 \text{ mol/L}$$

   b) Calculate the pH of gastric juice.

   $[H^+] = 0.0185$ mol/L $\quad \therefore$ pH = -log(0.0185) = 1.73

   c) Calculate the number of grams of HCl(aq) per 100 mL of gastric juice.

   n = CV = (0.0185 mol/L)(0.100L) = 0.00185 mol

   mass = (mol)(mol mass) = (0.00185 mol)(36.460 94 g/mol) = 0.0675 g

2. Ascorbic acid (vitamin C) is a weak monoprotic acid that has the formula $HC_6H_7O_6$ and Ka = 8.0×10^{-5}. What is the pH of a 0.035 mol/L solution of ascorbic acid?

	$HC_6H_7O_6 \leftrightharpoons$	H^+ +	$C_6H_7O_6^-$
[initial]	0.035 M	0	0
[Δ]	-x	+x	+x
[equilibrium]	0.035 – x	x	x

$$Ka = \frac{[H^+][C_6H_7O_6^-]}{[HC_6H_7O_6]}$$

$$8.0 \bullet 10^{-5} = \frac{(x)(x)}{(0.035 - x)}$$

$$\text{check}: \frac{[HA]}{Ka} = \frac{0.035}{8.0 \bullet 10^{-5}} = 437.5 < 500 \therefore \text{must keep } x \text{ in the denominator}$$

$$(8.0 \bullet 10^{-5})(0.035 - x) = x^2$$

$$2.8 \bullet 10^{-6} - 8.0 \bullet 10^{-5} x = x^2$$

$$0 = x^2 + 8.0 \bullet 10^{-5} x - 2.8 \bullet 10^{-6}$$

$$x = \frac{-8.0 \bullet 10^{-5} \pm \sqrt{(8.0 \bullet 10^{-5})^2 - (4)(1)(-2.8 \bullet 10^{-6})}}{(2)(1)}$$

$$x = \frac{-8.0 \bullet 10^{-5} \pm \sqrt{1.12 \bullet 10^{-5}}}{2}$$

$$x = \frac{-8.0 \bullet 10^{-5} \pm 0.00335}{2}$$

$$x = -0.0017 \quad \text{or} \quad 0.0016$$

$$[H^+] = 0.0016 \text{ mol/L}$$

$$pH = -\log(1.6 \bullet 10^{-3}) = 2.8$$

3. The pH of a 0.072 mol/L solution of benzoic acid ($HC_7H_5O_2$) is 2.68. Calculate the numerical value of the Ka for this acid.

$$HC_7H_5O_2 \leftrightarrows H^+ + C_7H_5O_2^-$$

[initial]	0.072 M	0	0
[Δ]	-x	+x	+x
[equilibrium]	0.072 – x	x	x

x = [H+]

[H+] = $10^{-2.68}$ = $2.09 \bullet 10^{-3}$ = x

$$K_a = \frac{[H^+][C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

$$K_a = \frac{(2.09 \bullet 10^{-3})(2.09 \bullet 10^{-3})}{(0.072 - 2.09 \bullet 10^{-3})} = 6.2 \bullet 10^{-5}$$

4. What is the pH of a solution formed by mixing 100.0 mL of a 0.150 mol/L HCl(aq) with 150.0 mL of 0.0900 mol/L NaOH(aq)?

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

mol H^+ = (0.150 mol/L)(0.1000 L) = 0.0150 mol H^+

mol OH^- = (0.0900 mol/L)(0.1500 L) = 0.0135 mol OH^-

excess mol H^+ = 0.0150 mol – 0.0135 mol = 0.0015 mol

$$C = \frac{n}{V} = \frac{0.0015\ \text{mol/L}}{0.250\ \text{L}} = 0.0060\ \text{mol/L} = [H^+]$$

pH = -log(0.0060) = 2.2

5. Acetylsalicylic acid (ASA) is the most widely used drug in the world. ASA has the chemical formula, $C_8H_7O_2COOH$, and its K_a equals 3.27 x 10^{-4}. Calculate the pH of a 0.250 mol/L ASA solution.

$$C_8H_7O_2COOH \leftrightarrows H^+ + C_8H_7O_2COO^-$$

[initial]	0.250 M	0	0
[Δ]	-x	+x	+x
[equilibrium]	0.250 – x	x	x

$$Ka = \frac{[H^+][C_8H_7O_2COO^-]}{[C_8H_7O_2COOH]}$$

$$3.27 \bullet 10^{-4} = \frac{(x)(x)}{(0.250 - x)}$$

$$3.27 \bullet 10^{-4} = \frac{(x)(x)}{(0.250)}$$

$$8.18 \bullet 10^{-5} = x^2$$

$$x = 9.04 \bullet 10^{-3}$$

$$pH = -\log\left(9.04 \bullet 10^{-3}\right) = 2.04$$

6. Calculate the $H_3O^+(aq)$ ion concentration of the resulting solution if 20.0 mL of 0.200 mol/L HCl(aq) is mixed with 30.0 mL of 0.350 mol/L NaOH(aq).

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

mol H^+ = (0.200 mol/L)(0.0200 L) = 0.00400 mol H^+
mol OH^- = (0.350 mol/L)(0.0300 L) = 0.0105 mol OH^-

excess mol OH^- = 0.0105 mol – 0.00400 mol = 0.0065 mol

$$C = \frac{n}{V} = \frac{0.0065\ \text{mol/L}}{0.0500\ \text{L}} = 0.130\ \text{mol/L} = [OH^-]$$

pOH = -log(0.130) = 0.88
pH = 14.0 – 0.88 = 13.1
$[H+] = 10^{-13.1} = 7.7 \bullet 10^{-14}$ mol/L

7. Calculate the $[H^+]$ in a solution of 0.10 mol/L hypochlorous acid, HOCl. The K_a value of HOCl is 2.9×10^{-8}.

	$HOCl$	$\leftrightarrows H^+$	$+\ OCl^-$
[initial]	0.10 M	0	0
[Δ]	-x	+x	+x
[equilibrium]	0.10 – x	x	x

$$Ka = \frac{[H^+][OCl^-]}{[HOCl]}$$

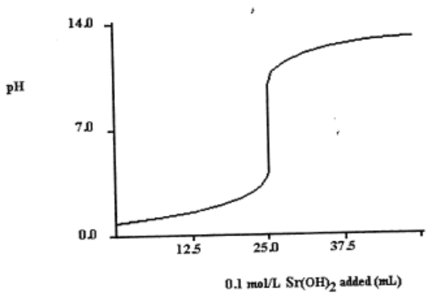
$$2.9 \bullet 10^{-8} = \frac{(x)(x)}{(0.10 - x)}$$

$$2.9 \bullet 10^{-8} = \frac{(x)(x)}{(0.10)}$$

$$2.9 \bullet 10^{-9} = x^2$$

$$x = 5.4 \bullet 10^{-5} = [H^+]$$

8. A student titrated 45.0 mL of a nitric acid solution of unknown concentration with a 0.10 mol/L strontium hydroxide solution. A pH meter was used to obtain the following titration curve:



a) Based on the results shown on the graph, calculate the concentration of the nitric acid solution.

$$2HNO_3(aq) + Sr(OH)_2(aq) \rightarrow Sr(NO_3)_2(aq) + 2H_2O(l)$$

$$C_a = \frac{C_b V_b R_a}{V_a R_b} = \frac{(0.10\ mol/L)(0.0250\ L)(2)}{(0.0450\ L)(1)} = 0.11\ mol/L$$

b) What would be an appropriate indicator for this titration and why?

litmus (4.5 – 8.3), bromothymol blue (6.0 – 7.6),
phenol red (6.6 – 8.0), phenolphthalein (8.2 – 10.0)

A suitable indicator must change colour during the vertical section of the graph.

9. a) A chemistry laboratory has solutions of four indicators: methyl orange, phenol red, phenolphthalein, and alizarin yellow R. Use the information given in the Chemistry Data Booklet to choose an appropriate indicator for each of the following titrations, and give your reasons in each case. (The concentrations of all acid and base solutions are in the 0.10 to 0.30 mol/L range.)

i) $CH_3COOH(aq)$ with $NaOH(aq)$

weak acid & strong base ➔ endpoint pH > 7
phenolphthalein (8.2 – 10.0), or alizarin yellow R (10.1 – 12)

ii) $HCl(aq)$ with $NH_3(aq)$

strong acid & weak base ➔ endpoint pH < 7
methyl orange (3.2 – 4.4)

iii) $HCl(aq)$ with $NaOH(aq)$

strong acid & strong base ➔ endpoint pH = 7

phenol red (6.6 – 8) or phenolphthalein (8.2 – 10.0)

b) For which titration would you have more than one choice among the above indicators? Which indicators could be used in this titration?

i) $CH_3COOH(aq)$ with $NaOH(aq)$ ➔ phenolphthalein or alizarin yellow R

iii) $HCl(aq)$ with $NaOH(aq)$ ➔ phenol red or phenolphthalein

10. Originally, Arrhenius believed that there must be the same number of positive and negative ions in a solution. However, the modern theory of electrolytes states that the total positive ionic charge must equal the total negative ionic charge. Is there a difference between these two statements? If so, why did Arrhenius' original theory have to be modified?

Yes, there is a difference between the two statements. The Arrhenius theory would have been correct if all positive ions and negative ions had equal but opposite charges. However, different positive ions can have charges of 1+, 2+, 3+, or 4+, and different negative ions can have charges of 1-, 2-, or 3-. Thus, in the case of a solution of a salt like calcium chloride ($CaCl_2$), there would be twice as many chloride ions as calcium ions: however, the total positive ionic charge does equal the total negative ionic charge.

11. Butanoic acid, C_3H_7COOH, is found in rancid butter and has a well-deserved, malodorous reputation (it stinks). In a well-ventilated room, 0.20 moles of the acid are dissolved in water to give 500 mL of a solution whose pH is found to be 2.60. Assuming that butyric acid is monoprotic (it has only one ionizable proton), calculate K_a for the acid.

	C_3H_7COOH	$\leftrightarrows$	H^+	+	$C_3H_7COO^-$
[initial]	0.040 M		0		0
[Δ]	-x		+x		+x
[equilibrium]	0.40 – x		x		x
	$0.40 - 2.51 \bullet 10^{-3}$		$2.51 \bullet 10^{-3}$		$2.51 \bullet 10^{-3}$

$$[C_3H_7COOH] = \frac{n}{V} = \frac{0.20\text{ mol}}{0.500\text{ L}} = 0.40\text{ mol/L}$$

x = [H+]

$[H+] = 10^{-2.60} = 2.51 \bullet 10^{-3} = x$

$$K_a = \frac{[H^+][C_3H_7COO^-]}{[C_3H_7COOH]}$$

$$K_a = \frac{(2.51 \bullet 10^{-3})(2.51 \bullet 10^{-3})}{(0.40 - 2.51 \bullet 10^{-3})} = 1.6 \bullet 10^{-5}$$

12. The K_w of water at 25°C is 1.0×10^{-14}, and the thermochemical equation for the reaction is:

$H_2O(l) \leftrightarrows H^+(aq) + OH^-(aq) \qquad \Delta H = +56\text{ kJ}$

Predict, giving your reason, how the Kw will change if the temperature is increases.

The reaction is endothermic. If the temperature is increased, the equilibrium shifts to the right, producing more $H^+(aq)$ and $OH^-(aq)$. Since $Kw = [H^+][OH^-]$, Kw will increase as the concentrations of $H^+(aq)$ and $OH^-(aq)$ increase.

13. The $HCO_3^-(aq)$ ion is capable of acting as either a Brønsted-Lowry acid or base. Illustrate this behaviour by writing balanced equations for its reactions with $CN^-(aq)$ ions and with $HNO_3(aq)$.

HCO_3^- as an acid: $HCO_3^-(aq) + CN^-(aq) \leftrightarrows CO_3^{2-}(aq) + HCN(aq)$

HCO_3^- as a base: $HCO_3^-(aq) + HNO_3(aq) \leftrightarrows H_2CO_3(aq) + NO_3^-(aq)$

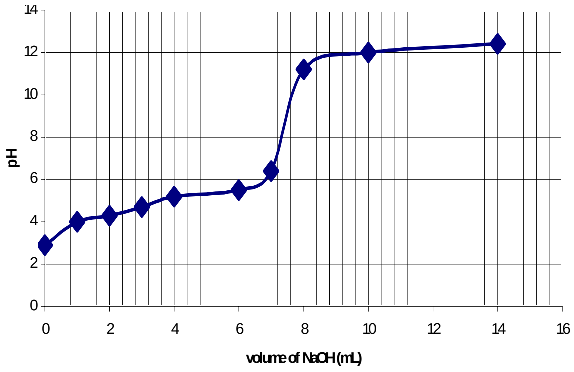
14. Three solutions are known to have pH values of 4.5, 6.0, and 8.0. Using two acid-base indicators (Chemistry Data Booklet, page 8), describe how you would distinguish among them.

The experiment design involves testing each solution with two indicators. One indicator must distinguish the solution that has pH 4.5 from the other two – use bromocresol green, litmus or methyl red. The other indicator must distinguish a solution that has pH = 8 from the other two – use bromothymol blue or phenol red. The third solution can be identified by elimination.

15. The following data show how the pH changes during a titration when 0.100 mol/L NaOH(aq) is added to 10.0 mL of aqueous acetic acid (CH_3COOH).

mL NaOH	0.0	1.0	2.0	3.0	4.0	6.0	7.0	8.0	10.0	14.0
pH	2.9	4.0	4.3	4.7	5.2	5.5	6.4	11.2	12.0	12.4

a) Using graph paper, plot the pH titration curve. Determine the value of the end-point pH.



b) Calculate the concentration of $CH_3COOH(aq)$

$CH_3COOH(aq) + NaOH(aq) \rightarrow Na\ CH_3COO(aq) + H_2O(l)$

$$C_a = \frac{C_b V_b R_a}{V_a R_b} = \frac{(0.100\ \text{mol/L})(0.0075\ \text{L})(1)}{(0.010\ \text{L})(1)} = 0.075\ \text{mol/L}$$

c) Use the information provided in this question and your answer in part (b) to calculate the numerical value of Ka for $CH_3COOH(aq)$

	CH_3COOH	$\leftrightarrows$	H^+	+	CH_3COO^-
[initial]	0.075 M		0		0
[Δ]	-x		+x		+x
[equilibrium]	0.075 – x		x		x
	$0.075 - 1.3 \bullet 10^{-3}$		$1.3 \bullet 10^{-3}$		$1.3 \bullet 10^{-3}$

x = [H+], $[H+] = 10^{-2.9} = 1.3 \bullet 10^{-3} = x$

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH3COOH]}$$

$$K_a = \frac{(1.3 \bullet 10^{-3})(1.3 \bullet 10^{-3})}{(0.075 - 1.3 \bullet 10^{-3})} = 2.1 \bullet 10^{-5}$$

16. A set of six solutions was prepared for a demonstration. The solutions were:

0.05 mol/L $Ba(OH)_2(aq)$	0.10 mol/L $NaOH(aq)$	0.05 mol/L $H_2SO_4(aq)$
0.10 mol/L $HCl(aq)$	0.10 mol/L $CH_3COOH(aq)$	0.10 mol/L $NH_3(aq)$

Before the solutions were labeled, they became scrambled, and the technician was forced to label the bottles A to F. The technician tested all six solutions and four mixtures (prepared by mixing equal volumes of two solutions) first in a conductivity apparatus and then with bromothymol blue. The technician knew that of the possible mixtures of acids and bases only the following would produce an insoluble salt:

$$Ba(OH)_2(aq) + H_2SO_4(aq) \leftrightharpoons BaSO_4(s) + H_2O(l)$$

Solution(s)	Ability to Conduct Electricity	Colour of Bromothymol Blue
A only	good	yellow
B only	poor	blue
C only	poor	yellow
D only	good	blue
E only	good	blue
F only	good	yellow
A & D mixed	good	green
A & E mixed	good	green
D & F mixed	good	green
E & F mixed	poor	green

a) Classify the solutions A to F as: strong acids, weak acids, strong bases, or weak bases.

Strong Acids: A & F	Weak Acids: C
Strong Bases: D & E	Weak Bases: B

b) Explain the differences in the tests results of mixture E & F compared with the other mixtures, and identify solutions E and F.

The mixture of E and F was the only mixture of an acid and a base that was a poor conductor of electricity. This must have been a mixture of aqueous solutions of barium hydroxide (strong base, E) and sulfuric acid (strong acid, F). The product of this mixture would be barium sulfate, and it is an insoluble salt that is a poor conductor of electricity.

c) Match the letters A to D with the appropriate solutions.

A = hydrochloric acid (HCl)
C = acetic acid (CH_3COOH)
B = ammonia (NH_3)
D = sodium hydroxide (NaOH)

17. a) Give one similarity and one difference between the Arrhenius and the Brønsted-Lowry definitions of acids and bases.

Both the Arrhenius and the Brønsted-Lowry theories predict that substances that have ionisable hydrogen atoms are acidic and that ionic hydroxide compounds are basic. However, the Brønsted-Lowry definition is a broader one in which acids and bases are defined in terms of how they react and not merely whether they produce hydrogen ions or hydroxide ions in aqueous solution.

b) Explain how well CH_3COOH fits into each definition.

CH_3COOH fits well in both definitions. The hydrogen atom that is bonded to the oxygen atom does ionize in aqueous solutions, and CH_3COOH is a proton donor.

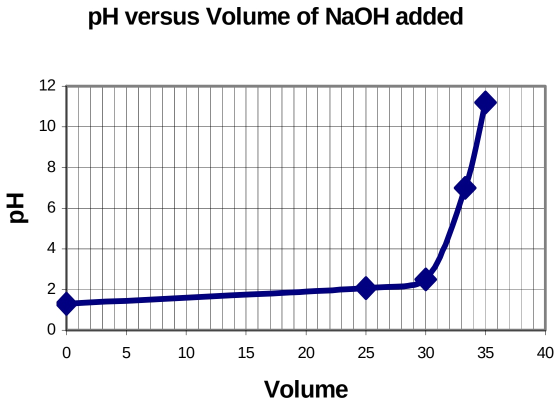
$$CH_3COOH + H_2O \leftrightarrows CH_3COO^- + H_3O^+$$

c) Explain how well NH_3 fits into each definition.

NH_3 fits well into the Brønsted-Lowry definition since it can accept a proton:

$$NH_3 + H_2O \leftrightarrows NH_4^+ + OH^-$$

However, it does not fit well into the Arrhenius definition since it is not an ionic hydroxide and does not dissociate in aqueous solution to produce hydroxide ions.



18. Construct a graph of pH versus volume of base added for the titration of 50.0 mL of 0.0500 M HI with 0.0750 M NaOH, based on the calculations below.

a) How many mL of base are required to neutralize the HI?

$$NaOH + HI \rightarrow NaI + H_2O$$

$$V_b = \frac{C_a V_a R_b}{C_b R_a} = \frac{(0.0500\ mol/L)(0.0500\ L)(1)}{(0.0750\ mol/L)(1)} = 0.0333\ L \text{ or } 33.3\ mL$$

b) What is the pH before any base is added?

pH = -log(0.0500) = 1.30

c) What is the pH after the addition of 25.0 mL of NaOH?

mol H^+ = (0.0500 mol/L)(0.0500 L) = 0.00250 mol H^+
mol OH^- = (0.0750 mol/L)(0.0250 L) = 0.001875 mol OH^-

excess mol H^+ = 0.00250 mol – 0.001875 mol = 0.000625 mol

$$C = \frac{n}{V} = \frac{0.000625\ mol/L}{0.0750\ L} = 0.00833\ mol/L = [H^+]$$

pH = -log(0.00833) = 2.08

d) What is the pH after 30.0 mL of NaOH has been added?

mol H^+ = (0.0500 mol/L)(0.0500 L) = 0.00250 mol H^+
mol OH^- = (0.0750 mol/L)(0.0300 L) = 0.00225 mol OH^-

excess mol H^+ = 0.00250 mol – 0.00225 mol = 0.00025 mol

$$C = \frac{n}{V} = \frac{0.00025\ mol/L}{0.0800\ L} = 0.0031\ mol/L = [H^+]$$

pH = -log(0.0031) = 2.5

e) What is the pH at the equivalence point?

Titration of a strong acid with a strong base, therefore the pH is seven at the equivalence point.

f) What is the pH after 35.0 mL of NaOH have been added?

mol H^+ = (0.0500 mol/L)(0.0500 L) = 0.00250 mol H^+
mol OH^- = (0.0750 mol/L)(0.0350 L) = 0.002625 mol OH^-

excess mol OH^- = 0.00250 mol – 0.002625 mol = 0.000125 mol

$$C = \frac{n}{V} = \frac{0.000125\ \text{mol/L}}{0.0850\ \text{L}} = 0.00147\ \text{mol/L} = [OH^-]$$

pOH = -log(0.00147) = 2.8
pH = 14.0 – 2.8 = 11.2

19. Using the K_a table in your Data Booklet, determine the $[OH^-]$ of 0.068 M of HCN.

	$HCN \leftrightarrows$	H^+ +	CN^-
[initial]	0.068 M	0	0
[Δ]	-x	+x	+x
[equilibrium]	0.068 – x	x	x

$$Ka = \frac{[H^+][CN^-]}{[HCN]}$$

$$6.2 \bullet 10^{-10} = \frac{(x)(x)}{(0.068 - x)}$$

$$6.2 \bullet 10^{-10} = \frac{(x)(x)}{(0.068)}$$

$$4.2 \bullet 10^{-11} = x^2$$

$$x = 6.5 \bullet 10^{-6} = [H^+]$$

pH = -log($6.5 \bullet 10^{-6}$) = 5.2
pOH = 14.0 – 5.2 = 8.8
[OH-] = $10^{-8.8}$ = $1.5 \bullet 10^{-9}$ mol/L

20. a) Write equations for the reactions of the amphoteric substance, HS^-, with HCl and with KOH.

HS^- with HCl: $HS^-(aq) + HCl(aq) \leftrightarrows H_2S(aq) + Cl^-(aq)$

or $HS^-(aq) + H^+(aq) \leftrightarrows H_2S(aq)$

HS^- with KOH: $HS^-(aq) + KOH(aq) \leftrightarrows S^{2-}(aq) + K^+(aq) + H_2O(l)$

or $HS^-(aq) + OH^-(aq) \leftrightarrows S^{2-}(aq) + H_2O(l)$

b) Write the formula for the conjugate base of HSO_4^-.

SO_4^{2-}

c) Write the formula for the conjugate acid of HSO_4^-.

H_2SO_4

21. a) Calculate the pH of a 2.5×10^{-3} mol/L NaOH solution.

$[OH^-] = 2.5 \times 10^{-3}$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-3}} = 4.0 \times 10^{-12}$$

$$pH = -\log[H_3O^+] = -\log(4.0 \times 10^{-12}) = \underline{\underline{11.4}}$$

or $pOH = -\log[OH^-] = -\log(2.5 \times 10^{-3}) = 2.6$

$pH = 14.00 - pOH = 14.00 - 2.6 = \underline{\underline{11.4}}$

b) Calculate the pH of a solution of HCl(aq) made by diluting 3.50 mL of 12.0 mol/L HCl to a final volume of 500.0 mL.

$C_1V_1 = C_2V_2$ $\quad [H_3O^+] = \frac{(12.0\ \text{mol/L})(0.003\ 50\ \text{L})}{0.5000\ \text{L}} = 0.0840\ \text{mol/L}$

$pH = -\log(0.0840) = 1.08$

22. a) Write equations for the reactions of the amphoteric substance, HPO_4^{2-}, with

i) with HCl

$HCl + HPO_4^{2-} \leftrightarrows Cl- + H_2PO_4^-$

or: $H_3O^+ + HPO_4^{2-} \leftrightarrows Cl- + H_2PO_4^-$

ii) with KOH

$KOH + HPO_4^{2-} \leftrightarrows K^+ + H_2O + PO_4^{3-}$

or: $OH- + HPO_4^{2-} \leftrightarrows K^+ + H_2O + PO_4^{3-}$

b) Write the formula for the conjugate base of each of these species:

i) H_2SO_4 HSO_4^-

ii) PH_4^+ PH_3

c) Write the formula for the conjugate acid of each of these species:

i) HPO_4^{2-} $H_2PO_4^-$

ii) NO_3^- HNO_3

23. a) What is the pH of a 0.155 mol/L HCl solution?

HCl is a strong acid, therefore, [H+] = 0.155 mol/L

pH = -log(0.155) = $\underline{\underline{0.810}}$

b) Calculate the volume of 0.120 mol/L NaOH needed to neutralize 25.0 mL of a 0.155 mol/L HCl

$$HCl + NaOH \rightarrow NaCl + H_2O$$

$$V_b = \frac{C_aV_aR_b}{C_bR_a} \qquad C_b = \frac{(0.155\ \text{mol/L})(0.0250\ \text{L})(1)}{(0.120\ \text{mol/L})(1)} = \underline{\underline{0.0323\ \text{L or } 32.3\ \text{mL}}}$$

c) Explain why the endpoint of an acid-base titration does not always coincide with the equivalence point.

The equivalence point represents the stoichiometric quantity of titrant required by the balanced chemical equation. The endpoint of a titration is indicated by a sudden change in some property such as the colour change of an acid-base indicator. If a suitable indicator is chosen and the titration is performed carefully, the endpoint should be the equivalence point.

If an unsuitable indicator is chosen (one which does not complete its colour change within the vertical region of the pH curve), or sloppy titration techniques are used, the endpoint will differ from the equivalence point.

24. The data given in the table has been compiled for 0.10 mol/L aqueous solutions of these compounds: CH_3COOH, HBr, HCN, H_2SO_4, Na_2HBO_3, $NaHCO_3$, and NaH_2PO_4. Complete the COMPOUND column of the table. In the table, n/a means not applicable. (Pages 8 and 9 of the Chemistry Data Booklet may be useful.)

Compound	pH	Orange IV Colour	Reaction with zinc	mL needed to neutralize 20 mL 0.10 mol/L NaOH	mL needed to neutralize 20 mL 0.10 mol/L HCl
H_2SO_4		red	fast	10	n/a
$NaHCO_3$		yellow		20	20
NaH_2PO_4		yellow		10	20
HCN	5.1		slow		n/a
Na_2HBO_3		yellow		20	10
HBr		red	fast		n/a
CH_3COOH	2.9		slow		n/a