**IB Chemistry – HL**

**Topic 8 Questions**

**1.** The *K*a value for an acid is 1.0×10–2. What is the *K*b value for its conjugate base?

A. 1.0×10–2

B. 1.0×10–6

C. 1.0×10–10

D. 1.0×10–12

**2.** Separate 20.0 cm3 solutions of a weak acid and a strong acid of the same concentration are titrated with NaOH solution. Which will be the same for these two titrations?

I. Initial pH

II. pH at equivalence point

III. Volume of NaOH required to reach the equivalence point

A. I only

B. III only

C. I and II only

D. II and III only

**3.** When the following 1.0 mol dm–3 aqueous solutions are arranged in order of **increasing** pH, which is the correct order?

I. Ammonium chloride

II. Ammonium ethanoate

III. Sodium ethanoate

A. I, II, III

B. II, I, III

C. III, I, II

D. III, II, I

**4.** An acid-base indicator, HIn, dissociates according to the following equation.

HIn(aq)  H+(aq) + In–(aq)

colour A colour B

Which statement about this indicator is correct?

I. In a strongly acidic solution colour B would be seen.

II. In a neutral solution the concentrations of HIn(aq) and In–(aq) must be equal.

III. It is suitable for use in titrations involving weak acids and weak bases.

A. I only

B. II only

C. III only

D. None of the above

**5.** What is the concentration of OH– ions (in mol dm–3) in an aqueous solution in which  
[H+] = 2.0×10–3 mol dm–3? (*K*w = 1.0×10–14 mol2 dm–6)

A. 2.0×10–3

B. 4.0×10–6

C. 5.0×10–12

D. 2.0×10–17

**6.** What is the relationship between *K*a and p*K*a?

A. p*K*a = –log *K*a

B. p*K*a= 

C. p*K*a = log *K*a

D. p*K*a= 

**7.** Which curve is produced by the titration of a 0.1 mol dm–3 weak base with 0.1 mol dm–3 strong acid?



**8.** The acid dissociation constant of a weak acid HA has a value of 1.0×10–5 mol dm–3.   
What is the pH of a 0.10 mol dm–3 aqueous solution of HA?

A. 2

B. 3

C. 5

D. 6

**9.** Which mixture would produce a buffer solution when dissolved in 1.0 dm3 of water?

A. 0.50 mol of CH3COOH and 0.50 mol of NaOH

B. 0.50 mol of CH3COOH and 0.25 mol of NaOH

C. 0.50 mol of CH3COOH and 1.00 mol of NaOH

D. 0.50 mol of CH3COOH and 0.25 mol of Ba(OH)2

**10.** Which compound, when dissolved in aqueous solution, has the highest pH?

A. NaCl

B. Na2CO3

C. NH4Cl

D. NH4NO3

**11.** Which values are correct for a solution of NaOH of concentration 0.010 mol dm–3 at 298 K?

(*K*w= 1.0×10–14 mol2 dm–6 at 298 K)

A. [H+] = 1.0×10–2 mol dm–3 and pH = 2.00

B. [OH–] = 1.0×10–2 mol dm–3 and pH = 12.00

C. [H+] = 1.0×10–12 mol dm–3 and pOH = 12.00

D. [OH– ] = 1.0×10–12 mol dm–3 and pOH = 2.00

**12.** Which solution, of concentration 0.10 mol dm–3, has the highest pH value?

A. HCl(aq)

B. MgCl2(aq)

C. NaCl(aq)

D. AlCl3(aq)

**13.** Which statement about indicators is **always** correct?

A. The mid-point of an indicator’s colour change is at pH = 7.

B. The pH range is greater for indicators with higher p*K*4 values.

C. The colour red indicates an acidic solution.

D. The p*K*a value of the indicator is within its pH range.

**14.** Which compound will dissolve in water to give a solution with a pH greater than 7?

A. sodium chloride

B. potassium carbonate

C. ammonium nitrate

D. lithium sulfate

**15.** An aqueous solution has a pH of 10. Which concentrations are correct for the ions below?

[H+(aq)] mol dm–3 [OH–(aq)] mol dm–3

|  |  |  |
| --- | --- | --- |
| A. | 104 | 10–10 |
| B. | 10–4 | 10–10 |
| C. | 10–10 | 10–4 |
| D. | 10–10 | 10–4 |

**16.** Which graph shows how the pH changes when a weak base is added to a strong acid?



**17.** When the following acids are listed in decreasing order of acid strength (strongest first), what is the correct order?

|  |  |
| --- | --- |
|  | *K*a |
| benzoic | 6.31×10–5 |
| chloroethanoic | 1.38×10–3 |
| ethanoic | 1.74×10–5 |

A. chloroethanoic  benzoic  ethanoic

B. benzoic  ethanoic  chloroethanoic

C. chloroethanoic  ethanoic  benzoic

D. ethanoic  benzoic  chloroethanoic

**18.** The strengths of organic acids can be compared using *K*a and p*K*a values. Which acid is the strongest?

|  |  |  |
| --- | --- | --- |
| A. | Acid A | p*K*a = 6 |
| B. | Acid B | p*K*a = 3 |
| C. | Acid C | *K*a = 1×10–5 |
| D. | Acid D | *K*a = 1×10–4 |

**19.** Which is the correct statement about the pH and pOH values of an aqueous solution at 25°C?

A. pH + pOH =14.0

B. pH + pOH =1.0 ×10–14

C. pH × pOH =14.0

D. pH × pOH =1.0 ×10–14

**20.** Which salt, when dissolved in water to form a 1.0 mol dm–3 solution, produces the lowest pH value?

A. Ammonium chloride

B. Ammonium ethanoate

C. Sodium ethanoate

D. Sodium chloride

**21.** Which solution has the lowest pH value?

A. Aluminium sulfate

B. Sodium nitrate

C. Potassium chloride

D. Sodium ethanoate

**22.** Which neutralization reaction could use phenolphthalein (p*K*a = 9.3) and not methyl orange (p*K*a = 3.7) as an indicator?

A. NaOH(aq) and HNO3(aq)

B. NH3(aq) and CH3COOH(aq)

C. NaOH(aq) and CH3COOH(aq)

D. NH3(aq) and HNO3(aq)

**23.** Water dissociates according to the equation

H2O(l)  H+(aq) + OH–(aq) *H* = +56 kJ

At 25C water has a pH of 7. Which of the following occurs when water is heated to 30C?

A. It remains neutral and its pH decreases.

B. It becomes acidic and its pH decreases.

C. It remains neutral and its pH increases.

D. It becomes acidic and its pH increases.

**24.** Which mixture would produce a buffer solution when dissolved in 1.0 dm3 of water?

A. 0.30 mol of NH3(aq) and 0.30 mol of HCl(aq)

B. 0.30 mol of NH3(aq) and 0.15 mol of HCl(aq)

C. 0.30 mol of NH3(aq) and 0.60 mol of HCl(aq)

D. 0.30 mol of NH3(aq) and 0.15 mol of H2SO4(aq)

**25.** Ammonia (NH3) is a weak base in aqueous solution with an ionization constant *K*b. What expression is equal to the ionization constant for the following reaction?

NH4+(aq) + H2O(l)  NH3(aq) + H3O+(aq)

A. 

B. 

C. 

D. 

**26.** The p*K*a values of four acids are as follows.

W 4.87

X 4.82

Y 4.86

Z 4.85

What is the correct order when these acids are arranged in order of **increasing** acid strength?

A. X, Z, Y, W

B. X, Y, Z, W

C. W, Z, Y, X

D. W, Y, Z, X

**27.** 10 cm3 of 0.01 mol dm–3 nitric acid (HNO3) is diluted with 90 cm3 of water. What is the pH of the resulting solution?

A. 1

B. 2

C. 3

D. 4

**28.** A base of concentration 0.10 mol dm–3 is titrated with 25 cm3 of an acid of concentration 0.10 mol dm–3. Which base-acid pair would have the highest pH at the equivalence point?

A. NaOH(aq) and CH3COOH(aq)

B. NaOH(aq) and HNO3(aq)

C. NH3(aq) and HNO3(aq)

D. NH3(aq) and CH3COOH(aq)

**29.** What is the value of [H+] in a buffer solution in which [CH3COOH] = 2.0 mol dm–3 and [CH3COO–] 1.0 mol dm–3? For CH3COOH, *K*a = 1.8×10–5 mol dm–3.

A. 6.0×10–3

B. 3.6×10–5

C. 1.8×10–5

D. 9.1×10–6

**30.** Which salt forms the most acidic solution when added to water?

A. NaCl

B. MgSO4

C. Al(NO3)3

D. KHCO3

**31.** An acid-base indicator has a p*K*a value of 4.0. At what pH will this indicator change colour?

A. 2.0

B. 4.0

C. 8.0

D. 12.0

**32.** Which values are correct for a 0.010 mol dm–3 solution of NaOH(aq) at 298 K?  
(*K*w = 1.0×10–14 mol2 dm–6 at 298 K)

A. [H+] = 1.0×10–12 mol dm–3 and pH = 12.00

B. [OH–] = 1.0×10–12 mol dm–3 and pH = 12.00

C. [H+] = 1.0×10–12 mol dm–3 and pOH = 12.00

D. [OH–] = 1.0×10–12 mol dm–3 and pOH = 12.00

**33.** At 25°C, *K*a for an acid is 1.0×10–2. What is the value of *K*b for its conjugate base?

A. 1.0×102

B. 1.0×10–2

C. 1.0×1012

D. 1.0×10–12

**34.** Which statement about indicators is **always** correct?

A. The mid-point of the pH range of an indicator is 7.

B. The pH range is greater for indicators with higher p*K*a values.

C. The colour red indicates an acidic solution.

D. The p*K*a value of the indicator is within its pH range.

**35.** (a) (i) Calculate the *K*avalue of methanoic acid, HCOOH, using table 16 in the Data Booklet.

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(1)

(ii) Based on its *K*avalue, state and explain whether methanoic acid is a strong or weak acid.

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(2)

(iii) Calculate the hydrogen ion concentration and the pH of a 0.010 mol dm–3 methanoic acid solution. State **one** assumption made in arriving at your answer.

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(b) Explain how you would prepare a buffer solution of pH 3.75 starting with methanoic acid.

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(3)

(Total 10 marks)

**36.** The indicator bromophenol blue, HIn(aq), has a form that is yellow and an In–(aq) form that is blue.

(a) Write an equation to show how bromophenol blue acts as an indicator.

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(1)

(b) State and explain the colour of bromophenol blue

(i) on the addition of a strong acid.

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(ii) at the equivalence point of a titration.

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(3)

(Total 4 marks)

**37.** (a) The dissociation of water takes place as follows:

H2O(l)  H+(aq) + OH–(aq)

(i) State the expression for the ionic product constant of water, *K*w.

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(1)

(ii) The value of *K*w is 2.410–14 mol2 dm–6 at 310 K. Calculate the [H+] at 310 K.

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(1)

(b) Lactic acid CH3CH(OH)COOH is a weak monoprotic acid

(p*K*a = 3.85 and *K*a = 1.4×10–4 mol dm–3).

(i) Write an equation for the reaction of lactic acid with water.

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(1)

(ii) State the ionization constant expression, *K*a, for lactic acid.

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(1)

(iii) Calculate the pH of a 0.20 mol dm–3 solution of lactic acid.

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(2)

(iv) Determine the pH of a solution containing 0.10 mol dm–3 of lactic acid and 0.10 mol dm–3 of sodium lactate.

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(1)

(Total 7 marks)

**38.** (a) (i) Write the equation for the reaction of ammonia with water.

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(ii) Derive the expression for *K*b for this reaction.

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(b) Using information from Table 16 in the Data Booklet, determine the pOH of a 0.20 mol dm–3 solution of ammonia.

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(3)

(Total 5 marks)

**39.** Benzoic acid, C6H5COOH, is a weak acid.

(a) Deduce the equation for the ionization of benzoic acid in water.

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(1)

(b) Use information from Table 16 in the Data Booklet to calculate a value for the dissociation constant, *K*a, for benzoic acid.

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(c) Derive the ionization constant expression for benzoic acid and use it to determine the pH of a 0.20 mol dm–3 aqueous solution of benzoic acid.

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(3)

(Total 5 marks)

**40.** The hydrogen ion concentration in pure water varies with temperature. At a particular temperature [H+] =1.7×10–7 mol dm–3.

(a) State the expression for the ionic product constant of water, *K*w, and calculate the value of *K*w at this temperature.

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(b) Calculate the pH of water at this temperature.

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(1)

(c) State and explain whether water at this temperature is acidic, neutral or alkaline.

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(2)

(Total 5 marks)

**41.** Predict whether each of the following solutions would be acidic, alkaline or neutral. In each case explain your reasoning.

(i) 0.1 mol dm–3 FeCl3(aq)

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(ii) 0.1 mol dm–3 NaNO3(aq)

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(iii) 0.1 mol dm–3 Na2CO3(aq)

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(Total 6 marks)

**42.** The following graph shows how the pH changes during the titration of 10 cm3 of a solution of a weak acid (HA) with 0.10 mol dm–3 NaOH.



(i) State the pH at the equivalence point and explain why the pH changes rapidly in this region.

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(ii) Calculate the initial concentration of the acid (HA).

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(iii) Calculate the [H+] of the acid before any sodium hydroxide is added. Use this value to determine the *Ka* value and the *pKa* value of the acid.

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(5)

(Total 10 marks)

**43.** A buffer solution can be made by dissolving 0.25 g of sodium ethanoate in 200 cm3 of 0.10 mol dm–3 ethanoic acid. Assume that the change in volume is negligible.

(i) Define the term *buffer solution.*

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(ii) Calculate the concentration of the sodium ethanoate in mol dm–3.

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(iii) Calculate the pH of the resulting buffer solution by using information from Table 16 of the Data Booklet.

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(3)

(Total 8 marks)

**44.** An experiment was carried out to determine the concentration of an aqueous solution of ammonia by titrating it with a solution of sulfuric acid of concentration 0.150 mol dm–3. It was found that 25.0 cm3 of the ammonia solution required 20.1 cm3 of the sulfuric acid solution for neutralization.

(a) Write the equation for the reaction and calculate the concentration, in mol dm–3, of the ammonia solution.

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(b) Several acid-base indicators are listed in Table 17 of the Data Booklet. State and explain which one of the following indicators should be used for this experiment: bromocresol green, phenol red or phenolphthalein.

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(c) Determine the pOH of a solution with an ammonia concentration of 0.121 mol dm–3.   
(p*K*bof ammonia is 4.75.)

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(Total 11 marks)

**45.** (i) State what is meant by the term *buffer solution*, and describe the composition of an acid buffer solution in general terms.

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(3)

(ii) Calculate the pH of a mixture of 50 cm3 of ammonia solution of concentration 0.10 mol dm–3 and 50 cm3 of hydrochloric acid solution of concentration 0.050 mol dm–3.

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(4)

(Total 7 marks)

**46.** The pH of a solution is 4.8. Using information from Table 17 of the Data Booklet, deduce and explain the colours of the indicators bromophenol blue and phenol red in this solution.

(Total 3 marks)

**47.** Calculate the pH of a buffer solution containing 0.0500 mol dm–3 of ethanoic acid  
(*K*a = 1.7410–5) and 0.100 mol–3 of sodium ethanoate.

(Total 3 marks)

**48.** Describe the composition and behaviour of a buffer solution.

(Total 3 marks)

**49.** (i) Define the term pH.

(1)

(ii) A 25.0 cm3 sample of 0.100 mol dm–3 hydrochloric acid was placed in a conical flask, and 0.100 mol dm–3 sodium hydroxide is added until a total of 50.0 cm3 had been added. Sketch a graph of pH against volume of NaOH(aq) added, clearly showing the volume of NaOH(aq) needed for complete reaction and the pH values at the start, the equivalence point and finish.

(4)

(iii) The experiment in (ii) was repeated, but with a 25.0 cm3 sample of 0.100 mol dm–3 ethanoic acid in the conical flask instead of the hydrochloric acid. Use information from Table 16 of the Data Booklet to calculate the pH at the start of the experiment. State the approximate pH value at the equivalence point.

(5)

(Total 10 marks)

**50.** (i) Describe how an indicator, HIn, works.

(3)

(ii) Name a suitable indicator for the reaction between ethanoic acid and sodium hydroxide. Use information from Table 17 in the Data Booklet to explain your choice.

(2)

(Total 5 marks)

**51.** (i) Identify **two** substances that can be added to water to form a basic buffer solution.

(1)

(ii) Describe what happens when a small amount of acid solution is added to the buffer solution prepared in (i). Use an equation to support your explanation.

(2)

(Total 3 marks)

**52.** Predict and explain whether an aqueous solution of 0.10 mol dm–3 AlCl3 will be acidic, alkaline or neutral.

(Total 2 marks)

**53.** A titration was carried out to determine the concentration of 25.0 cm3 of an aqueous solution of nitric acid. The pH value of the liquid in the flask was measured as 0.100 mol dm–3, aqueous sodium hydroxide was added. The results are shown on the graph below.



(i) Use the graph to determine the value of [H+] of the nitric acid solution.

(1)

(ii) Determine the pH value when the value of [H+] has decreased to 1×10–3 mol dm–3.

(1)

(iii) Use the graph to determine the volume of 0.100 mol dm–3 aqueous sodium hydroxide solution needed to exactly neutralize the nitric acid.

(1)

(iv) Calculate the concentration, in mol dm–3, of the nitric acid.

(2)

(Total 5 marks)

**54.** In aqueous solution at 298 K, ammonia is a weak base with a p*K*b value of 4.75 and a *K*b value of 1.7×10–5 mol dm–3.

(a) Write an equation for the reaction of ammonia with water.

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(1)

(b) State the ionization constant expression, *K*b, for ammonia.

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(1)

(c) Calculate the pH of a 0.25 mol dm–3 solution of ammonia.

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(Total 5 marks)

**55.** Nitric acid and ammonia may be used to make a buffer solution.

(i) Describe the behaviour of a buffer solution.

(2)

(ii) Describe how you could prepare a buffer solution using 0.100 mol dm–3 solutions of nitric acid and ammonia.

(3)

(Total 5 marks)

**56.** With reference to Table 16 in the Data Booklet, determine the pH of a 0.100 mol dm–3 solution of propanoic acid.

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(Total 3 marks)

**57.** 0.100 mol dm–3 hydrochloric acid solution is added to 25.0 cm3 0.100 mol dm–3 ammonia solution and the pH is recorded until a total of 35.0 cm3 hydrochloric acid has been added.

(i) Sketch a graph to show how the pH changes as hydrochloric acid is added to the ammonia solution. Use a pH scale of 0–14, and an acid volume scale of 0–35 cm3. Explain the shape of the curve.

(6)

(ii) Use table 17 of the Data Booklet to suggest an indicator that could be used in the titration, explaining your choice.

(2)

(Total 8 marks)

**58.** (i) State the composition of an acidic buffer solution.

(1)

(ii) Suggest the identity of an acid and its amount that could be added to a solution containing 0.10 mol ammonia in order to prepare a buffer.

(2)

(iii) Explain how the solution you prepare in (ii) can act as a buffer solution when a strong acid and a strong base are added to separate portions of it. Write an equation to illustrate the buffer action in **each** case.

(4)

(iv) Write an equation for the reaction of ammonia with water, and write its *K*b expression.

(2)

(Total 9 marks)

**59.** (a) Predict and explain, using equations where appropriate, whether the following solutions are acidic, alkaline or neutral.

(i) 0.1 mol dm–3 FeCl3(aq)

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(1)

(ii) 0.1 mol dm–3 NaNO3(aq)

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(1)

(iii) 0.1 mol dm–3 Na2CO3(aq)

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(1)

(b) Acidic gases can be released into the atmosphere that have an environmental impact when they are deposited as acid rain. State **two** elements that form the acidic gases and describe **two** impacts they have on the natural environment.

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(3)

(Total 6 marks)

**60.** An experiment was carried out to determine the concentration of aqueous ammonia by titrating it with a 0.150 mol dm–3 sulfuric acid solution. It was found that 25.0 cm3 of the aqueous ammonia required 20.1 cm3 of the sulfuric acid solution for neutralization.

(a) Write the equation for the reaction and calculate the concentration, in mol dm–3, of the aqueous ammonia.

(4)

(b) Several acid-base indicators are listed in Table 16 of the Data Booklet. Identify **one** indicator that could be used for this experiment. Explain your answer.

(3)

(c) (i) Determine the pOH of 0.121 mol dm–3 aqueous ammonia (p*K*b= 4.75).

(4)

(ii) State what is meant by the term *buffer solution*, and describe the composition of an acid buffer solution in general terms.

(3)

(iii) Calculate the pH of a mixture of 50.0 cm3 of 0.100 mol dm–3 aqueous ammonia and 50.0 cm3 of 0.0500 mol dm–3 hydrochloric acid solution.

(4)

(Total 18 marks)

**IB Chemistry – HL**

**Topic 8 Answers**

**1.** D

**2.** B

**3.** A

**4.** D

**5.** C

**6.** A

**7.** C

**8.** B

**9.** B

**10.** B

**11.** B

**12.** C

**13.** D

**14.** B

**15.** C

**16.** C

**17.** A

**18.** B

**19.** A

**20.** A

**21.** A

**22.** C

**23.** A

**24.** B

**25.** C

**26.** D

**27.** C

**28.** A

**29.** B

**30.** C

**31.** B

**32.** A

**33.** D

**34.** D

**35.** (a) (i) p*K*a = 3.75, therefore *K*a = 1.78×10–4 (accept 1.8×10–4) 1

No units required.

(ii) weak acid;  
less [H+]/partial dissociation/more reactants/less products/  
*K*a << 1/small *K*a; 2

(iii) (HCOOH(aq)  H+(aq) + HCOO–(aq))

*K*a =;  
(*x*2 = 1.78×10–6)  
*x* = 1.33×10–3 mol dm–3 = [H+] (*no mark without units*);

ECF from (a)(i).  
No penalty for incorrect significant figures.

pH = 2.88/2.9 *(ECF)*;  
assume *x* << 0.010/25°C/negligible dissociation; 4

(b) add strong base/sodium hydroxide or other named alkali/salt of methanoic  
acid/HCOONa to methanoic acid;  
in equimolar amounts/quantities/so that [HCOOH] = [HCOO–];  
(from *K*a expression) pH = p*K*a (= 3.75); 3

[10]

**36.** (a) HIn(aq)  H+(aq) + In–(aq); 1

 needed for mark. State symbols not essential.

(b) (i) yellow as equilibrium shifts to left to remove (added) H+(aq); 1

Colour and explanation needed for the mark.

(ii) green/blue-yellow;  
both HIn(aq) and In–(aq) are present; 2

[4]

**37.** (a) (i) *K*w = [H+][OH]; 1

(ii) [H+] = 1.5×107(mol dm3); 1

Accept answer in range 1.5 to 1.55.

(b) (i) CH3CH(OH)COOH + H2O  CH3CH(OH)COO + H3O+; 1

Ignore state symbols even if incorrect.

The double arrow is necessary for the mark.

(ii)  1

Allow [H3O+] for [H+] in the expression.

(iii) 5.3×103 = [H+];

pH = 2.3; 2

Allow ECF pH based on wrong [H+] in the value, award **[1]**.

Award **[2]** for correct pH.

(iv) pH = 3.85; 1

Accept answer in range 3.8 to 3.9.

[7]

**38.** (a) (i) NH3 + H2O  NH4+ + OH; 1

Do not penalise 

Do not accept NH4OH

(ii)  1

(b) *Kb* = 104.75 = 1.78×105;



pOH = log[OH] = 2.72; 3

Accept answer in range 2.68 to 2.76.

Correct answer scores **[3]**.

Apply ECF throughout this part.

[5]

**39.** (a) C6H5COOH + H2O  C6H5COO + H3O+; 1

Ignore state symbols.

Accept C6H5COOH  C6H5COO + H+.

 needed for mark.

(b) *Ka* (= 104.20) = 6.31×105 (mol dm3); 1

Units not needed for mark, but penalize incorrect units.

(c) *Ka* 

[H+] =  *K*a[C6H5COOH]/3.55×103 (mol dm3);

pH = 2.45; 3

Apply ECF from (b) and from [H+] to pH.

Correct final answer scores [3].

[5]

**40.** (a) (*K*w = ) [H+][OH]/[H3O+] [OH];

= 2.89×1014 / 2.9×1014; 2

Units not needed.

(b) pH = 6.8; 1

Accept answer in range 6.7 to 6.8.

(c) neutral;

[H+] = [OH]/[H3O+] = [OH]/*OWTTE*; 2

[5]

**41.** (i) acidic;  
Fe(H2O)63+ is a weak acid/Fe3+ reacts with OH–/equation to show  
formation of HCl or H+;  
*“FeCl3 is acidic” is not acceptable.* 2

(ii) neutral;  
NaNO3 / sodium nitrate is formed from strong base and strong  
acid/ions do not hydrolyse; 2

(iii) alkaline;  
As CO32– is weak base/combines with H+/equation showing  
formation of OH–; 2

Acidic, neutral, alkali mark in each case is independent of reason.

[6]

**42.** (i) 8.7 ± 0.7;  
low [H+] thus small addition of OH– has great effect/OH– increases  
rapidly as NaOH is a strong base/logarithmic nature of pH; 2

(ii) volume of NaOH = 8.2 cm3 (*exact*);  
amount of NaOH =×0.1 = 0.00082 mol;  
[HA] == 0.082 mol dm–3/0.082 M; 3

Correct answer **[3]**, units needed for last mark.

(iii) correct pH reading from graph (2.9) *(allow 2.8 or 3.0)*;  
thus [H+] = 1.26×10–3 (mol dm–3)

*K*a =   
 = 1.9×10–5 (mol dm–3);  
p*K*a = 4.71 5

Accept 4.7 and allow ECF from (ii).

If pH given as 2.8, Ka = 3.06×10–5 and pKa = 4.51  
If pH given as 2.8, Ka = 1.22×10–5 and pKa = 4.91

If half equivalence method used:  
Volume = 4.1 cm3  
pKa = 4.75  
Award **[2]** out of last **[4]**.

[10]

**43.** (i) a solution that resists pH change/maintains a (nearly) constant pH;  
when **small** amounts of acid or alkali are added; 2

(ii) *M*r of sodium ethanoate;  
moles of sodium ethanoate =  = (0.0030);

[CH3COO–] =  = 0.015 (mol dm–3) *2* *sig figs only*; 3

(iii) *K*a =  *(or with substituted values)*;  
*May be assumed from later work.*

[H+] = = (1.159×10–4);

pH = 3.9(4); 3

Allow ECF throughout (ii) and (iii).

[8]

**44.** (a) 2NH3 + H2SO4 → (NH4)2SO4

Accept correct equation with NH4OH instead of NH3.

. mol H2SO4 = 0.0201×0.150;  
2NH3 = H2SO4/mol NH3 = 6.03×10–3;  
[NH3] = 0.241 (mol dm–3); 4

Apply –1(SF) if appropriate.

Award **[3]** for the correct final answer for the concentration calculation.

(b) bromocresol green;  
reaction of weak base and strong acid/*OWTTE*;  
pH range of bromocresol green is 3.8 to 5.4 / occurs at pH < 7; 3

(c) *K*b = 10–4.75 = 1.78×10–5;

*K*b = /[OH–] = ;

[OH–] = ;

pOH = 2.83; 4

Award **[4]** for the correct final answer.

Allow ECF, for example any correct conversion of [OH–]

[11]

**45.** (i) a solution which resists change in pH/changes pH very slightly/  
keeps pH constant/*OWTTE*;  
when small amounts of acid or base are added;  
weak acid and its salt/weak acid and its conjugate base; 3

(ii) mol NH3 = 0.0050 and mol HCl = 0.0025;  
[NH4+] = [NH3];  
[OH–] = *K*b = 1.78×10–5  
(pOH = 4.75 so) pH = 9.25 *(allow 9.2 to 9.3)*; 4

Award **[4]** for correct final answer.

Accept other valid methods such as Henderson-Hasselbach equation.

[7]

**46.** bromophenol blue is blue **and** phenol red is yellow;

pH of 4.8 is above range of bromophenol blue/bromphenol blue shows  
its alkaline colour/*OWTTE*;

pH of 4.8 is below range of phenol red/phenol red shows its acidic  
colour/*OWTTE*; 3

[3]

**47.** /rearrangement for [H+];



pH ( = log[H+]) = 5.06; 3

**OR**

pH = 

pH = 4.76 + log

pH = 5.06;

Accept answer in range 5.0 to 5.1.

ECF from [H+].

Award **[3]** for correct final answer.

[3]

**48.** weak acid + salt of weak acid/weak acid + conjugate base.

Accept equivalent descriptions of a basic buffer.

the solution resists pH change;

Do not accept pH does not change.

when small amounts of acid or base are added;

Only award if previous answer correct.

[3]

**49.** (i) pH =  log[H+]; 1

(ii) curve should include the following:

starting pH = 1;

equivalence point: 25.0 cm3 of NaOH;

pH at equivalence point = 7;

pH to finish = 1213;



4

Penalise **[1]** if profile incorrect.

(iii) *Ka* = 104.76/1.74×105;

*Ka* = [H+]2÷[CH3COOH]/1.74×105 = 

[H+] = 1.32×103 (mol dm3);

starting pH = 2.88;

Accept 3 sig. fig.

Award **[4]** for correct pH.

Allow ECF.

pH at equivalence point: 89; 5

[10]

**50.** (i) HIn is a weak acid;  
HIn  H+ + In and two colours indicated;

In acid equilibrium moves left or vice versa; 3

(ii) phenolphthalein/phenol red/bromothymol blue;  
colour change of indicator occurs within the range of pH at equivalence  
point/on vertical part of graph; 2

[5]

**51.** (i) specific examples of weak base and its salt/specific strong acid and  
weak base; 1  
e.g. NH3 and NH4Cl.

(ii) pH changes very little/most acid neutralized by base;  
equation from (i); 2  
e.g. NH3 + H+  /NH4OH + H+   + H2O.

[3]

**52.** acidic;

[Al(H2O6)]3+ is (weak) acid due to the formation of H+/  
[Al(H2O)6]3+  [Al(H2O)5(OH)]2+ + H+; 2

[2]

**53.** (i) 0.1 (mol dm3); 1

(ii) 3; 1

(iii) 28(.0) (cm3 ); 1

(iv) nNaOH/HNO3 ( = 0.100×0.0280) = 2.80×103 (mol);

ECF from value in (iii).

[HNO3] (= 2.80×103÷0.025) = 0.112 (mol dm3); 2

ECF from n above.

Correct final answer scores **[2]**.

[5]

**54.** (a) NH3(aq) + H2O(l)  N(aq) + OH(aq); 1

Ignore state symbols and accept .

(b)  1

(c) [OH] = 2.1×103

pOH = 2.7/[H+] = 4.8×1012

pH = 11.3; 3

Allow ECF for the value of pOH and pH.

[5]

**55.** (i) a solution which resists change in pH;  
when a small amount of strong acid or base is added to it; 2

(ii) react excess ammonia with nitric acid;  
stated volumes with about 50% more ammonia solution;  
gives a solution containing the weak base and its salt with the acid/  
NH4+ and NH3; 3

Accept suitable volumes from about 20 cm3 to about 500 cm3 for 2nd mark.

[5]

**56.** (p*K*a (propanoic) = 4.87)

*k*a

[H3O+]= 1.16×103 (mol dm3);

pH = 2.94; 3

Award **[3]** for correct answer.

[3]

**57.** (i)



graph starting at pH < 13;

Award **[0]** for pH=13.

equivalence point pH < 7;

Accept anything between 4 and 6

bottom end of graph: pH between 3 and 1;

NH3 is a weak base/partially dissociated/[OH]</<<0.10mol dm3

(therefore, pH < 13);

NH4+ formed is a weak acid/NH4+  NH3 + H+/NH4+ dissociates into  
a weak base and a strong acid (thus acidic at equivalence point);  
HCl is a strong acid, thus graph finishes close to pH = 1; 6

(ii) methyl orange/bromocresol green/bromophenol blue/methyl red;  
*pK*a of indicator centred around pH at equivalence/end point/indicator  
pH range falls where there is a sharp pH change/*OWTTE*; 2

[8]

**58.** (i) weak acid **and** salt of the weak acid/its conjugate base; 1

(ii) HCl/HNO3/H2SO4;

Amount < 0.10 mol for HCl/HNO3/< 0.05 mol for H2SO4; 2

(iii) (added) OH reacts with NH4+ present/acid of buffer;

(added) H+ reacts with NH3 present/base of buffer;

OH + NH4+  NH3 + H2O (strong base replaced by weak base);

H+ + NH3  NH4+ (strong acid replaced by weak acid); 4

(iv) NH3(aq) + H2O(l)  NH4+(aq) + OH(aq);

States not required for mark

*K*b = ; 2

[9]

**59.** (a) (i) acidic **and** [Fe(H2O)6]3+ is a weak acid  
[Fe(H2O)6]3+(aq)  [Fe(OH)(H2O)5]2+(aq) + H+(aq); 1

“FeCl3 is acidic” is not acceptable.

(ii) neutral **and** NaNO3/sodium nitrate is formed from strong base  
and strong acid/ions do not hydrolyze; 1

(iii) alkaline **and** CO32– is a weak base/  
CO32–(aq) + H2O(l)  HCO3–(aq) + OH–(aq); 1

Award **[1]** only for correct identification of solutions as acidic, neutral and alkaline only, without explanation.

(b) nitrogen **and** sulfur;  
kills/harms fish/aquatic life in lakes/rivers;  
leaching of soils damages plant life/trees;3

[6]

**60.** (a) 2NH3(aq) + H2SO4(aq)  (NH4)2SO4(aq); 4

Accept correct equation with NH4OH instead of NH3.  
n(H2SO4) = 0.0201×0.150 (mol);  
n(NH3) = 6.03×10–3 (mol);  
[NH3] = 0.241 (mol dm–3);

Award **[3]** for the correct final answer for the concentration calculation.

(b) bromocresol green;  
reaction of weak base and strong acid;  
pH range of bromocresol green is 3.8 to 5.4/occurs at pH < 7; 3

(c) (i) *K*b = 10–4.75 = 1.78×10–5;  
*K*b =   
[OH–] =   
pOH = 2.83; 4

Award **[4]** for the correct final answer.

Allow ECF, for example any correct conversion of [OH–] to pOH.

(ii) a solution which resists change in pH/changes pH very slightly;  
when small amounts of acid or base are added;  
weak acid and its salt/weak acid and its conjugate base; 3

(iii) n(NH3) = 0.00500 (mol) **and** n(HCl) = 0.00250 (mol);  
  
[OH–] = *K*b = 1.78×10–5;  
(pOH = 4.75 so) pH = 9.25 *(allow 9*.*2 to 9*.*3)*; 4

Award **[4]** for correct final answer.

[18]