Topic 18 Practice Problems

# **18.1 Lewis Acids and Bases**

1. AlCl3 reacts with chloride ions to form the complex ion AlCl4-.
	1. Draw Lewis structures to illustrate the interaction between AlCl3 and the chloride ion to form the complex ion AlCl4-.
	2. Explain how the behavior of AlCl3 is consistent with the definition of a Lewis acid.
	3. What is the name for the special type of covalent bond that exists between AlCl3 and the chloride ion?
2. For each of the following species, state whether it is most likely to act as a Lewis acid or a Lewis base:
	* 1. PH3
		2. BCl3
		3. H2S
		4. Cu2+
3. Which equation represents an acid-base reaction according to the Lewis theory **but** not the Brønsted-Lowry theory?

A. NH3 + HCl  NH4Cl

B. 2H2O  H3O+ + OH–

C. NaOH + HCl  NaCl + H2O

D. CrCl3 + 6NH3  [Cr(NH3)6]3+ + 3Cl–

1. Which statement explains why ammonia can act as a Lewis base?

A. Ammonia can donate a lone pair of electrons.

B. Ammonia can accept a lone pair of electrons.

C. Ammonia can donate a proton.

D. Ammonia can accept a proton.

1. Define an acid in terms of the Lewis theory. Deduce, giving a reason, whether NF3 is able to function as a Lewis acid or as a Lewis base.

# **18.2 Calculations Involving Acids and Bases**

1. For each of the following, determine [H+], [OH-] , pH and pOH
	1. 1.5 M HF (Ka = 7.0 X 10 –4)
	2. 1.0 M CH3COOH ( Ka =1.8 X 10 –5)
2. The pH of a 0.1 mol dm-3 solution of cyanic acid, HCN, is 3.
	1. Write the equation for the ionization of cyanic acid.
	2. Write the Ka expression.
	3. Calculate Ka.
3. Determine the pH of a 1.2 mol dm-3 solution of HF (Ka = 7.2 x 10-4).
4. Determine the pOH and the pH of a 1.8 mol dm-3 solution of KF that has a pKa 3.17.
5. The pKb for a base is 5. What is the pH of a 0.1 mol dm-3 solution of the base?
6. A 0.75 mol dm-3 solution of a weak acid, HX, has a pH of 1.35. Determine Ka for the acid.
7. The *K*a value for an acid is 1.0 × 10–2. What is the *K*b value for its conjugate base?
	1. 1.0 × 10-2
	2. 1.0 × 10-6
	3. 1.0 × 10-10
	4. 1.0 × 10-12
8. Four aqueous solutions, I, II, III and IV, are listed below. What is the correct order of **increasing** pH of these solutions?
	* 1. 0.100 mol dm-3 HCl
		2. 0.010 mol dm-3 HCl
		3. 0.100 mol dm-3 NaOH
		4. 0.010 mol dm-3 NaOH
9. I, II, III, IV
10. I, II, IV, III
11. II, I, III, IV
12. II, I, IV, III
13. Propanoic acid, CH3CH2COOH is a weak acid.

(a) Give the equation for the ionization of propanoic acid in water and deduce the expression for the ionization constant, *K*a, of propanoic acid.

 (b) Calculate the *K*a value of propanoic acid using the p*K*a value in the Data Booklet.

 (c) Use your answer from (b) to calculate the [H+] in an aqueous solution of propanoic acid of concentration 0.0500 mol dm–3, and hence the pH of this solution.

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

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

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

1. Determine the pH of the solution resulting when 100 cm3 of 0.50 mol dm–3 HCl(aq) is mixed with 200 cm3 of 0.10 mol dm–3 NaOH(aq).

# **18.3 pH Curves**

1. When the following 1.0 mol dm-3 solutions are arranged in order of increasing pH, which is the correct order?
	* 1. Ammonium chloride
		2. Ammonium ethanoate
		3. Sodium ethanoate
2. I, II, III
3. II, I, III
4. III, I, II
5. III, II, I
6. Which mixture would produce a buffer solution when dissolved in 1.0 dm3 of water?
	1. 0.50 mol of CH3COOH and 0.50 mol of NaOH
	2. 0.50 mol of CH3COOH and 0.25 mol of NaOH
	3. 0.50 mol of CH3COOH and 1.00 mol of NaOH
	4. 0.50 mol of CH3COOH and 0.25 mol of Ba(OH)2
7. 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.

a) Define the term *buffer solution.*

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

c) Calculate the pH of the resulting buffer solution by using information from Table 21 of the Data Booklet.

1. Which substances could be added to a solution of ethanoic acid to prepare an acidic buffer solution?

I. Hydrochloric acid

II. Sodium ethanoate

III. Sodium hydroxide

A. I and II only

B. I and III only

C. II and III only

D. I, II and III

1. Calculate the pH of a buffer solution made by reacting 20 cm3 0.10 mol dm-3 HCl(aq) with 40 cm3 0.10 mol dm-3 NH3(aq) at 298K. The pKb for ammonia is 4.75.
2. (a) (i) Calculate the *K*avalue of methanoic acid, HCOOH, using table 21

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

(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.

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

1. (i) 25.0 cm3 of 1.00 × 10–2 mol dm–3 hydrochloric acid solution is added to 50.0 cm3 of
1.00 × 10–2 mol dm–3 aqueous ammonia solution. Calculate the concentrations of both ammonia and ammonium ions in the resulting solution and hence determine the pH of the solution.

(ii) State what is meant by a buffer solution and explain how the solution in (i), which contains ammonium chloride dissolved in aqueous ammonia, can function as a buffer solution.

1. For each of the following salts, tell if the aqueous solution is acidic, basic, or neutral. When acidic or basic, show the hydrolysis equation. KF, KCl, NH4NO3, NaClO4, Mg(NO3)2
2. Which compound will dissolve in water to give a solution with a pH greater than 7?
	1. Sodium chloride
	2. Potassium carbonate
	3. Ammonium nitrate
	4. Lithium sulfate
3. Deduce whether the pH of the resulting salt solution will be greater than, less than or equal to 7 when the following solutions exactly neutralize each other.
	1. H2SO4(aq) + NH3(aq)
	2. H3PO4(aq) + KOH(aq)
	3. HNO3(aq) + Ba(OH)2(aq)
4. 50 mL of a base with unknown concentration was titrated with 1.5M HCl. Determine the concentration of the base if it required 25 mL of acid to neutralize the base. (Think how a dilution calc is set up)
5. Sketch titration curves for:
6. The titration of a strong acid and a weak base (base added to acid)
7. The titration of a strong acid and a strong base (base added to acid)
8. The titration of a weak acid and a strong base (base added to acid)
9. The titration of a weak acid and a weak base (base added to acid)
10. Hydrochloric acid (in the flask) is to be titrated with aqueous sodium carbonate (in the burette).
11. Would you choose methyl orange or phenolphthalein for this titration? Explain. (use data booklet)
12. What color change would you expect to see at the end point?
13. Explain why adding too much indicator could lead to an inaccurate titration result.
14. The laboratory has run out of both methyl orange and phenolphthalein. Below are listed some indicators that are available. Which would you use to replace your original choice? Explain.

# Indicator pKa Color Change

bromophenol blue 4.0 yellow to blue

bromothymol blue 7.0 yellow to blue

thymol blue 8.9 yellow to blue

1. 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

1. Which curve is produced by the titration of a 0.1 mol dm−3 weak base with 0.1 mol dm−3 strong acid?
2. (a) Explain why a 1.0 mol dm–3 solution of sodium hydroxide has a pH of 14 whereas 1.0 mol dm–3 ammonia solution has a pH of about 12. Use equations in your answer.

(b) 20.0 cm3 of a known concentration of sodium hydroxide is titrated with a solution of nitric acid. The graph for this titration is given below.



(i) State an equation for the reaction between sodium hydroxide and nitric acid.

(ii) Calculate the concentration of the sodium hydroxide solution before the titration.

(iii) From the graph determine the volume of nitric acid required to neutralize the sodium hydroxide and calculate the concentration of the nitric acid.

(iv) Predict the volume of ethanoic acid of the same concentration as the nitric acid in (b) (iii), required to neutralize 20.0 cm3 of this sodium hydroxide solution.

1. Which statement about indicators is always correct?
	1. The mid-point of an indicator’s color change is at pH=7
	2. The pH range is greater for indicators with higher pKa values
	3. The color red indicates an acidic solution
	4. The pKa values of an indicator is within its pH range
2. Bromocresol green has a pH range of 3.8-5.4 and changes color from yellow to blue as the pH increases
	1. Of the four types of titrations shown, state in which two of these this indicator could be used.
	2. Suggest a value for the pKa of this indicator
	3. What color will the indicator be at pH 3.6?