

3AB Chemistry

Acids & bases



<http://www.siraze.net/chemistry/sezennur/subjects/comics/comics24.htm>

Tyson

Name: _____

Acids and Bases in aqueous solutions

- Apply an understanding of the concept of an electrolyte to explain the self-ionisation of water
- Explain and apply the Arrhenius and Bronsted-Lowry models to describe acids and bases
- Apply the relationship between K_w and temperature to explain the pH value of a neutral solution at different temperatures
- Apply the relationship $\text{pH} = -\log [\text{H}^+(\text{aq})]$ to calculate the pH of:
 - Strong acid solutions
 - Strong base solutions
 - The resulting solution when strong acid-base solutions are mixed
- Apply the Bronsted-Lowry model to the hydrolysis of salts derived from monoprotic and polyprotic acids, and bases
- Describe and explain the conjugate nature of buffer solutions
 - Explain using Le Chatelier's Principle how buffers respond to the addition of H^+ and OH^-
 - Explain qualitatively the concept of buffering capacity

Chemical Reactions

Reactions, equations and stoichiometry

- Describe, write equations for and interpret observations for the following reaction types:
 - Neutralisation
 - Hydrolysis of salts of weak acids and weak bases
- Perform volumetric analysis using either acid-base or redox context, and:
 - Give a description of procedures used and methods for minimising experimental error
 - Describe and explain characteristics of primary standards and standard solutions
 - Demonstrate an understanding of end point and equivalence point to the selection of an appropriate indicator in an acid-base titration
 - Explain the choice of indicators (in acid-base only) or use self-indicators (redox)
- Perform calculations based on acid-base

Reactions

- Acid + base \rightarrow water + salt
- Acid + carbonate \rightarrow carbon dioxide + water + salt
- Acid + hydrogencarbonate \rightarrow carbon dioxide + water + salt
- Acid + active metal \rightarrow hydrogen gas + salt

Strong acids (completely ionise)

- Sulfuric acid H_2SO_4
- Nitric acid HNO_3
- Hydrochloric acid HCl
- Hydroiodic acid HI
- Hydrobromic acid HBr

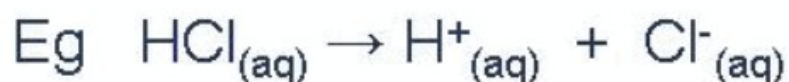
Weak acids (partially ionise)

Examples include

- Phosphoric acid H_3PO_4
- Carbonic acid H_2CO_3
- Acetic acid CH_3COOH
- Oxalic acid $\text{H}_2\text{C}_2\text{O}_4$

Arrhenius theory

- Arrhenius theory states that acids produce hydrogen ions in aqueous solution



H^+ ions often attach themselves to a water molecule to form a hydronium ion H_3O^+

Textbooks will often use H^+ and H_3O^+ interchangeably.

- Arrhenius theory states that bases are substances that produce hydroxide ions in solution.

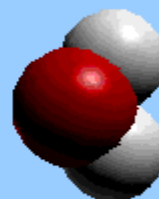
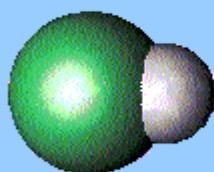
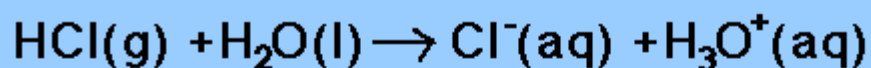


- Bases are metal hydroxides and metal oxides.

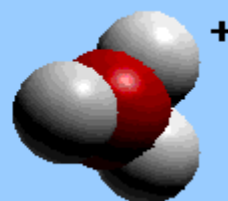
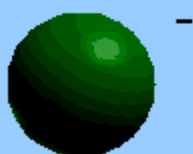
Bronsted_Lowry Theory

- Another name for a hydrogen ion is a proton (because H^+ consists of one single proton)
- BL theory states that acids are proton donors (they give away hydrogen ions)

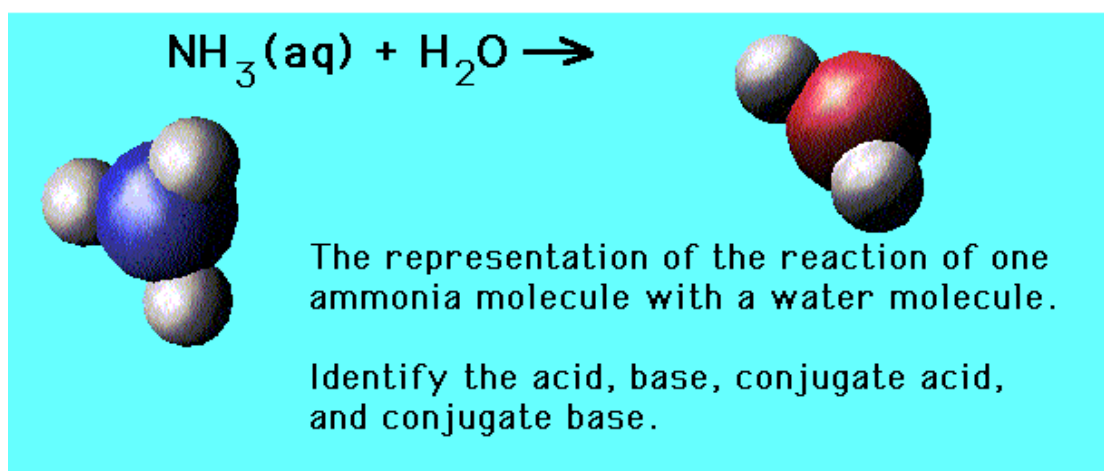
- How do acids give away H^+ ions?



Press 'continue' to start the animation.



- BL theory states that a base accepts protons (or hydrogen ions)

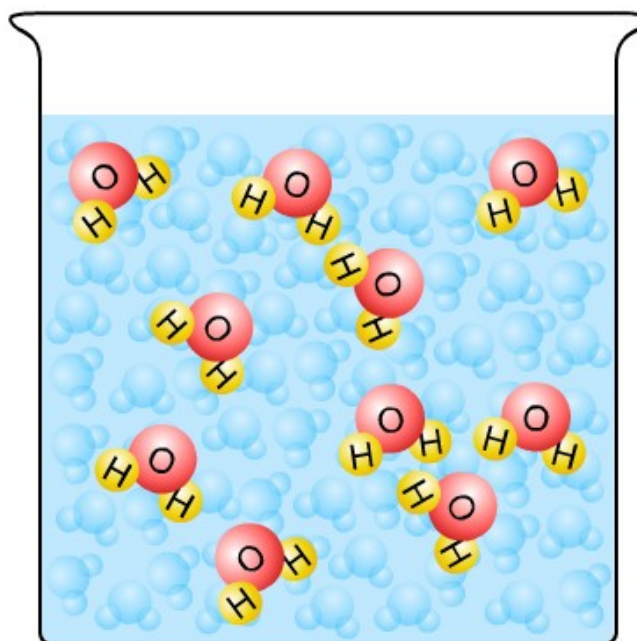
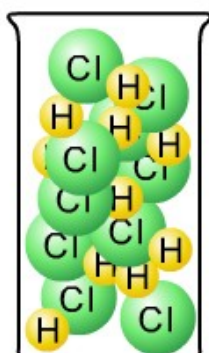


Draw the particles after the proton transfer has occurred in the space below

Strong versus weak acids

Strong Acid

Start



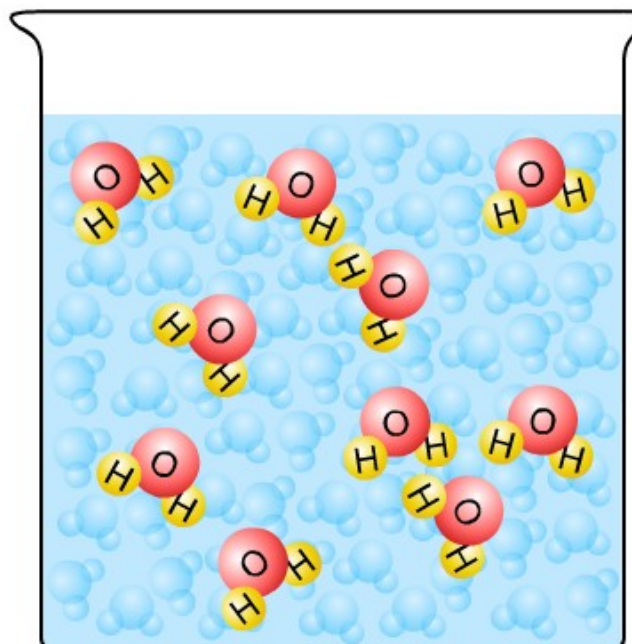
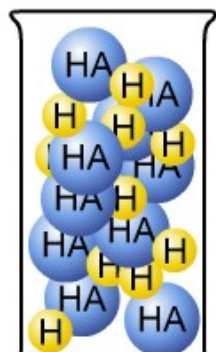
What happens when HCl is added to beaker of water?

Sketch a microscopic view of the contents of the beaker in the space below



Weak Acid

Start



Sketch a picture of beaker of water when the weak acid HAH is added



Electrolytes (textbook reference page 206)

Any substance that conducts electricity when dissolved in water is called an electrolyte

Strong acids are strong electrolytes because they ionise completely in solution

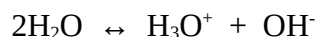
Weak acids are weak electrolytes because they only partially ionise in solution

Water is a weak electrolyte

Water partially ionises according to the following equation



Or can be written as



The value of the equilibrium constant for this reaction is very small (1×10^{-14} at 25°C). As a result of this the concentration of ions in water is very small and hence water is a poor conductor of electricity

Strength of acids and bases (ref page 210 textbook)

The strength of an acid or base is indicated by the position of equilibrium when the acid or base is added to water. If the equilibrium constant for the production of H_3O^+ or OH^- is large then the substance is described as a strong acid or base.

In general the more readily the acid donates a hydrogen ion, the less likely its conjugate base will accept a hydrogen ion. This means that a strong acid will have a very weak conjugate base.

Which will form the stronger conjugate base – hydrochloric acid or hydrofluoric acid?

Explain your choice

	Acid	Conjugate base	
Completely ionised in water	H_2SO_4 HCl HNO_3	HSO_4^- Cl^- NO_3^-	Negligible base strength
	\uparrow Acid strength increases	Base strength increases \downarrow	
	H_3O^+ HSO_4^- H_3PO_4 CH_3COOH H_2CO_3 NH_4^+ HCO_3^- H_2O OH^-	H_2O SO_4^{2-} H_2PO_4^- CH_3COO^- HCO_3^- NH_3 CO_3^{2-} OH^- O^{2-}	
Negligible acid strength			

Figure 7.6 Relative strengths of some common acid–base conjugate pairs

Multiprotic acids (textbook ref page)

HCl is referred to as a monoprotic acid as 1 molecule of HCl contains 1 acidic hydrogen.

H₃PO₄ is referred to as a triprotic acid

H₂SO₄ is a diprotic acid

The number of hydrogen ions an acid can donate does not relate to the strength of the acid.

Review exercise 7.1

1. Predict and explain the difference in the electrical conductivity of a solution of 1 mol L⁻¹ potassium nitrate and a solution of 1 mol L⁻¹ glucose, C₆H₁₂O₆.
2. The acidic nature of carbonic acid, H₂CO₃, can be explained by both the Arrhenius and Brønsted-Lowry theories. Using relevant equations, explain how carbonic acid can be classified as an acid according to both of these theories.
3. Hydrogen chloride, HCl, is a gas at room temperature while hydrogen nitrate, HNO₃, is a liquid. However, the two aqueous solutions of these different compounds have similar properties.
 - a) Explain the similarities between the two solutions
 - b) In terms of their intermolecular bonding, explain why hydrogen chloride is a gas at room temperature, but hydrogen nitrate, HNO₃, is a liquid. (can't do yet)
4. Write an equation for the following substances behaving as an:
 - a) Arrhenius base in water
 - i LiOH(s)
 - ii Ba(OH)₂(s)
 - b) Arrhenius acid in water
 - i H₃PO₄(l)
 - ii HNO₃(l)
5. Write an equation for the following substances behaving as a:
 - a) Brønsted-Lowry acid in water
 - i H₂PO₄⁻
 - ii H₂S(g)
 - b) Brønsted-Lowry base in water
 - i NH₃(g)
 - ii CH₃COO⁻
6. Write the formula for the conjugate acid of each of the following.
 - a) HCO₃⁻
 - b) H₂O
 - c) OH⁻
 - d) NH₃
7. Write the formula for the conjugate base of each of the following.
 - a) H₂S
 - b) CH₃COOH
 - c) NH₄⁺
 - d) HCO₃⁻

8. Boric acid, H_3BO_3 , is a weak triprotic acid. Write the equations to show its behaviour as a tri acid in water.
9. Hydrogen bromide, HBr , is a stronger acid than hydrogen sulfide, H_2S .

Identify the following statements as true or false, giving reasons for your answer.

- a) The equilibrium constant for the hydrolysis reaction of HBr will be larger than that for H_2S
- b) A solution of HBr will always have more H_3O^+ ions present than a solution of H_2S .
- c) $\text{Br}^-(\text{aq})$ is a stronger base than $\text{HS}^-(\text{aq})$.

Hydrolysis reactions of weak acids, bases and salts. (ref page 212 textbook)

When a substance donates to or accepts a hydrogen ion from water it is referred to as a hydrolysis reaction.

Acetic acid, CH_3COOH , is referred to as a weak acid because the equilibrium constant for the hydrolysis reaction is small



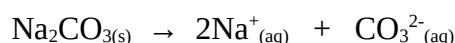
		Hydrolysis equation	Equilibrium constant
Acids	Phosphoric acid	$\text{H}_3\text{PO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{H}_2\text{PO}_4^-(\text{aq})$	$K_a = 7.5 \times 10^{-3}$
	Carbonic acid	$\text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$	$K_a = 4.3 \times 10^{-7}$
	Ammonium ion	$\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{NH}_3(\text{aq})$	$K_a = 5.6 \times 10^{-10}$
Bases	Carbonate ion	$\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$	$K_b = 1.8 \times 10^{-4}$
	Acetate ion	$\text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COOH}(\text{aq}) + \text{OH}^-(\text{aq})$	$K_b = 5.6 \times 10^{-10}$

Which is the stronger base – the acetate ion or the carbonate ion?
Explain your answer.

Solutions of salts in water – hydrolysis of ions

When a salt dissolves in water it dissociates completely into ions. Salts are strong electrolytes.

Sodium carbonate dissociates in water according to the following equation



It is then possible that one or both of the ions produced in the dissociation reaction can undergo further reaction with water. (undergo hydrolysis)

In the case of the sodium carbonate solution only the carbonate ion undergoes hydrolysis

	neutral	basic	acidic
Anion	Derived from strong acids Cl ⁻ NO ₃ ⁻ Br ⁻ I ⁻ SO ₄ ²⁻	Derived from weak acids F ⁻ S ²⁻ ClO ⁻ CH ₃ COO ⁻ CO ₃ ²⁻ HCO ₃ ⁻ PO ₄ ³⁻ HPO ₄ ²⁻	Derived from polyprotic acids HSO ₄ ⁻ H ₂ PO ₄ ⁻
cation	Derived from strong bases Li ⁺ Mg ²⁺ Na ⁺ Ca ²⁺ K ⁺ Ba ²⁺	none	NH ₄ ⁺ Al ³⁺ Fe ³⁺

Review exercise 7.2

1.
 - a) A solution is labelled 0.100 mol L^{-1} hydrochloric acid. What is the concentration of hydronium ions in the solution?
 - b) Would you expect a 0.100 mol L^{-1} acetic acid solution to have the same concentration of hydronium ions in its solution? Explain your answer.
2. Aqueous solutions of acetic acid and ammonia are both poor conductors of electricity, although they are both better conductors than water. However, when acetic acid and ammonia solutions are mixed, the resulting solution is a much better conductor. Write equations to support an explanation of these observations.
3. Write equations for any hydrolysis reactions that take place after the dissociation of each of the following salts when dissolved in water.
 - a) Na_3PO_4
 - b) NH_4NO_3
 - c) FeCl_3
4. Predict whether the following salts when dissolved in water will produce neutral, acidic or basic solutions. Explain your reasoning and include relevant equations to support your explanation.
 - a) KHSO_4
 - b) NaHCO_3
 - c) $\text{Ca}(\text{NO}_3)_2$

pH (ref page 219 of textbook)

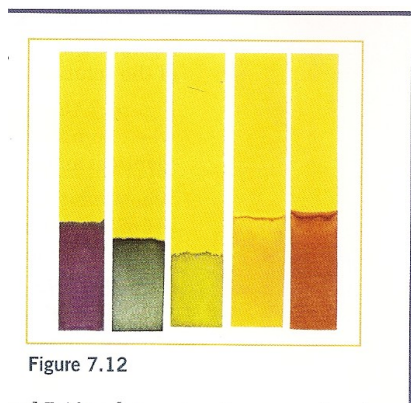
Review exercise 7.3

(For these problems assume a temperature of 25°C unless informed otherwise.)

1. Calculate the concentration of H_3O^+ in each of the following solutions.
 - a) $0.0125 \text{ mol L}^{-1} \text{ HCl}$
 - b) $0.0340 \text{ mol L}^{-1} \text{ LiOH}$
 - c) A solution containing 0.0550 mol of calcium hydroxide dissolved in 250 mL of solution
2. Determine the hydronium ion and hydroxide ion concentrations in a solution with a pH of 9.
3.
 - a) Calculate the pH solutions having the following concentrations of H_3O^+ .
 - i $5.39 \times 10^{-3} \text{ mol L}^{-1}$
 - ii $1.43 \times 10^{-9} \text{ mol L}^{-1}$
 - b) What is the hydronium ion concentration in solutions with the following pH values?
 - i 3.21
 - ii 7.50
 - iii 12.43
4. Calculate the hydronium ion concentration, hydroxide ion concentration and pH of the solutions formed when the following solutions are mixed.
 - a) 0.100 L of $0.300 \text{ mol L}^{-1} \text{ HNO}_3$ and 0.100 L of $0.300 \text{ mol L}^{-1} \text{ KOH}$
 - b) 0.500 L of $0.500 \text{ mol L}^{-1} \text{ HNO}_3$ and 0.500 L of $0.250 \text{ mol L}^{-1} \text{ KOH}$
 - c) 0.500 L of $0.250 \text{ mol L}^{-1} \text{ HNO}_3$ and 0.500 L of $0.250 \text{ mol L}^{-1} \text{ Ba(OH)}_2$
5. Calculate the pH of pure water at 10°C , given $K_w = 2.93 \times 10^{-15}$ at this temperature.
6. Explain why it is acceptable to use universal indicator to test the pH of a sample of garden soil, but a pH meter is preferable to measure the pH of the water in a river system.
7. Five colourless 0.1 mol L^{-1} solutions of the substances listed below were tested with universal indicator paper.

HCl
KF
 $\text{Mg(NO}_3)_2$
 NH_4NO_2
NaOH

The results obtained are shown in Figure 7.12; however, the solution tested for each strip of paper was not recorded.



- Use the universal indicator colours shown in Figures 7.9 and 7.10 to determine the approximate pH of the solution tested for each strip, and so determine whether each solution contains a strong acid, weak acid, strong base, weak base or neutral substance.
- Identify the solution tested by each strip of indicator paper. Explain your reasoning and give equations to support your answer.

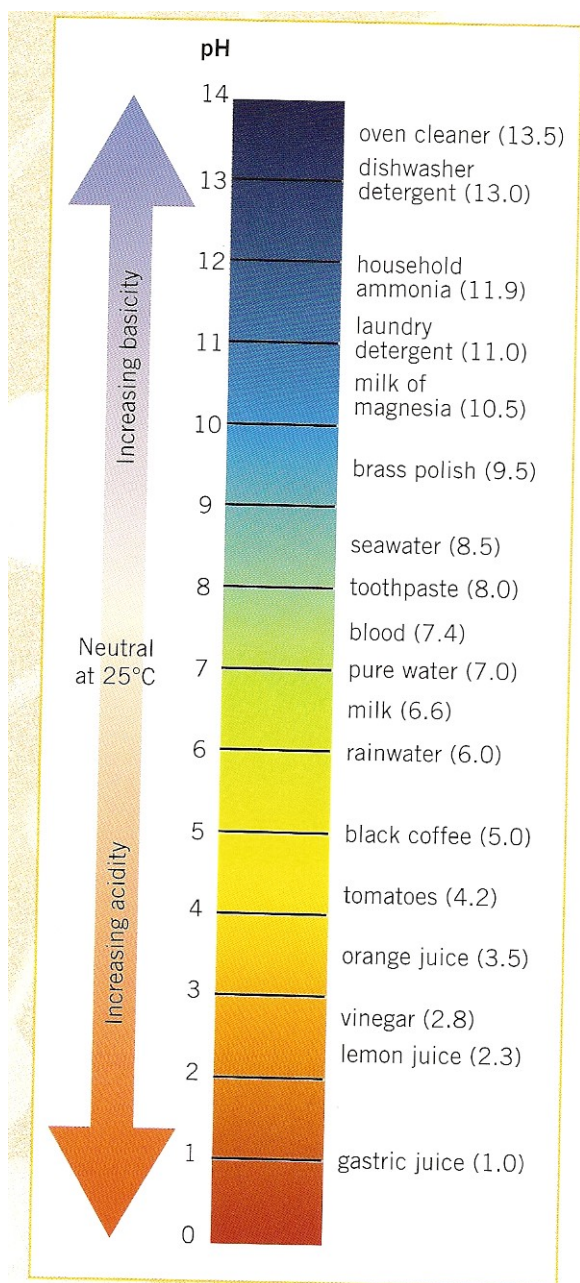
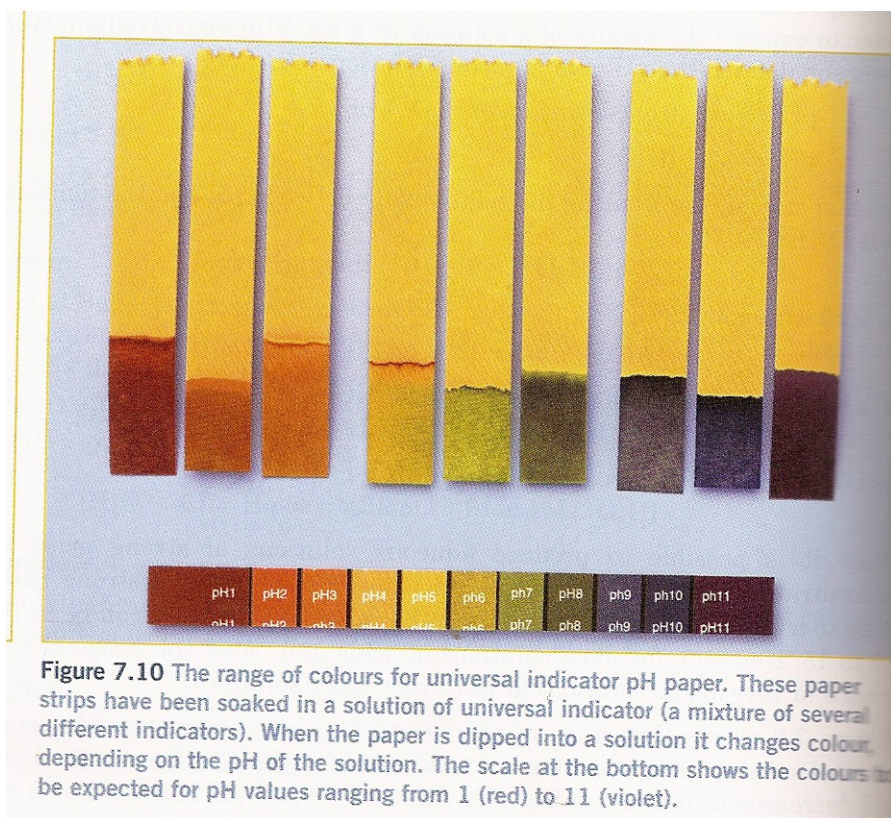


Figure 7.9 The pH values of some



Chemistry of buffer solutions (ref page 225 textbook)

A buffer is a solution that resists a change in its pH when small amounts of an acid or base are added.

A buffer typically consists of a weak acid and its conjugate base – often in equimolar amounts.

Buffer capacity is the extent to which a buffer can ‘absorb’ the extra acid or base added

The buffer capacity depends on two factors

- The relative concentrations of the weak acid and its conjugate base. The most effective buffer (the greatest buffer capacity) is one in which the concentrations of the acid and its conjugate base in the buffer are equal
- The concentrations of the weak acid and its conjugate base. The buffer capacity increases with an increase in the concentration of the conjugate pair. The greater the concentration of the conjugate pair, the more likely the buffer will be able to combine with the added acid or base by shifting the position of equilibrium either left or right

Blood as a buffer system

Review 7.4

1. Explain why a mixture of ammonium chloride and ammonia can act as a buffer, but a mixture of hydrochloric acid and sodium chloride solution cannot.
2. If a buffer solution is composed of a weak acid and its conjugate base in a 1:1 mole ratio, why does mixing 500 mL of 1.0 mol L⁻¹ acetic acid with 250 mL of 1.0 mol L⁻¹ sodium hydroxide produce a buffer solution?
3. Write equations to show what happens to a buffer solution containing equimolar amounts of H₂PO₄⁻ and HPO₄²⁻ when we add a small amount of:
 - a) OH⁻(aq)
 - b) H₃O⁺(aq)
4. Why would it be unwise to use tap water to calibrate a pH meter?

QUESTIONS

(Assume all the pH questions refer to solutions at 25°C, unless informed otherwise.)

- Using ammonium chloride and ammonia as example, explain the meaning of the terms 'strong electrolyte' and 'weak electrolyte'.
- Write an equation to show acetic acid acting as an acid according to the Arrhenius theory.
 - Write an equation to show sulfuric acid acting as an acid according to the Brønsted-Lowry theory.
 - Write an equation to show magnesium hydroxide acting as a base according to the Arrhenius theory.
 - Write an equation to show sodium carbonate acting as a base according to the Brønsted-Lowry theory.
- Write the formula for the conjugate acid of each of the following.
 - OH^-
 - HPO_4^{2-}
 - H_2SO_4
 - HS^-
 - CO_3^{2-}
 - Write the formula for the conjugate base of each of the following.
 - OH^-
 - HPO_4^{2-}
 - HCN
 - NH_4^+
 - HSO_4^-
- For each of the following equations identify the conjugate acid-base pairs, indicating which is the acid and which the base in each pair.
 - $\text{HCl}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
 - $\text{HSO}_4^-(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{SO}_4^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{PH}_3(\text{g}) + \text{HI}(\text{g}) \rightleftharpoons \text{PH}_4^+(\text{s}) + \text{I}^-(\text{s})$
 - $\text{CH}_3\text{NH}_2(\text{l}) + \text{CH}_3\text{CH}_2\text{OH}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{l}) + \text{CH}_3\text{CH}_2\text{O}^-(\text{l})$
 - $\text{O}^{2-}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2\text{OH}^-(\text{aq})$
- Using equations, explain why CH_3COOH is classified as a monoprotic acid, but H_3PO_4 is classified as a triprotic acid.
- Phosphoric acid is a triprotic acid but is not as strong an acid as monoprotic hydrochloric acid. Why is this the case?

7. Identify the following statements as true or false.

Comparing 1 L of 0.1 mol L⁻¹ acetic acid and 1 L of 0.1 mol L⁻¹ nitric acid:

- a) Both solutions have the same pH.
- b) The acetic acid would contain fewer hydroxide ions than the nitric acid.
- c) Both solutions are prepared by adding 0.1 mol of the original pure acid to water and making up to 1 L.
- d) The pH of the acetic acid is higher than 7 and that of the nitric acid is lower than 7.
- e) Both solutions contain the same mass of original acid.
- f) The equilibrium constant for the hydrolysis of acetic acid is smaller than that for the hydrolysis of nitric acid.
- g) Both solutions would require the same volume of 0.1 mol L⁻¹ sodium hydroxide for complete reaction.

8. Identify the following statements as true or false.

Comparing 1 L of 0.1 mol L⁻¹ potassium hydroxide and 1 L of 0.1 mol L⁻¹ ammonia:

- a) The pH of the potassium hydroxide solution would be higher than the pH of the ammonia solution.
- b) The ammonia solution would contain more hydronium ions than the potassium hydroxide solution.
- c) The value of K_w in the ammonia solution is less than the value of K_w in the potassium hydroxide solution.
- d) Both solutions would require the same volume of 0.1 mol L⁻¹ hydrochloric acid for complete reaction (to equivalence).

9. You have been supplied with 0.001 mol L⁻¹ solutions of four monoprotic acids and asked to measure their pH values.

The values you obtain are recorded below.

Solution M	pH = 5.7
Solution N	pH = 3.0
Solution O	pH = 4.5
Solution P	pH = 6.2

- a) Arrange these four solutions in order of increasing acid strength (weakest to strongest)
- b) Is any one of these acids a strong acid? Justify your answer.
- c) Another student in the class has been given 0.001 mol L⁻¹ solutions of some bases and records the following pH values.

Solution X	pH = 11
Solution Y	pH = 9
Solution Z	pH = 13

Comment on the likely acid-base nature of these three solutions and the measurements that this student has made.

10. Some solutions of sodium bromide, sodium hydrogensulfate and sodium hydrogencarbonate are found to have the pH values shown below.

Sodium bromide pH = 7

Sodium hydrogensulfate pH = 5.5

Sodium hydrogencarbonate pH = 8.0

Explain these results in terms of hydrolysis reactions that take place when the salts are dissolved in water.

11. Explain the difference between the terms 'ionisation', 'dissociation' and 'hydrolysis'.
12. a) Water is a weaker acid than the hydrogenselenate ion, HSeO_4^- . Write a balanced ionic equation for the hydrolysis of the hydrogenselenate ion.
b) Water is a weaker base than the cyanide ion, CN^- . Write a balanced ionic equation for the hydrolysis of the cyanide ion.
13. The labels have fallen off three bottles, each of which contains white crystals. The bottles are known to contain ammonium nitrate, sodium nitrate and sodium carbonate. Using only pH paper, explain how you would determine the contents of each bottle.
14. A dilute solution of ammonia has its pH recorded as 9.5. Determine the hydronium and hydroxide ion concentrations in the solution.
15. A 0.100 mol L^{-1} solution of acetic acid is found to be 1.35% ionised. Determine the pH of this solution.
16. Explain the meaning of the following statement with particular reference to the terms 'dilute', 'concentrated', 'strong' and 'weak'.
A 10 mol L^{-1} solution of acetic acid is a concentrated solution of a weak acid, while a 0.001 mol L^{-1} solution of sulfuric acid is a dilute solution of a strong acid.
17. Calculate the pH of the following solutions.
- a) A solution in which the concentration of hydroxide ions is $5.38 \times 10^{-3} \text{ mol L}^{-1}$
b) 45 mL of a $0.00735 \text{ mol L}^{-1}$ solution of the strong acid HBr.
c) 3.25 mol L^{-1} sodium hydroxide solution
d) A solution prepared by dissolving 4.48 L of hydrogen chloride gas as STP in water to form 0.250 mol L^{-1} nitric acid with 0.0532 g of barium hydroxide
18. Solution A has a concentration of hydronium ions that is 10 000 times greater than that of solution B. By what factor is the pH of solution A different from the pH of solution B?

19. Copy and complete the following table.

	$[\text{H}_3\text{O}^+]$ (mol L ⁻¹)	$[\text{OH}^-]$ (mol L ⁻¹)	pH	Acidic or basic?
Solution 1		3.45×10^{-4}		
Solution 2			7.59	
Solution 3	2.96×10^{-6}			

20. In January 2008 researchers in the USA developed a buffer system that resists pH changes when the temperature changes. In their investigations with biologically active molecules that are very sensitive to pH they discovered that in some buffer systems the pH decreased when the temperature dropped while in other buffer systems the pH increased when the temperature dropped.
- Why does the pH of water change when the temperature is changed?
 - What can you conclude about the buffer systems that caused a decrease in pH when the temperature dropped?
 - What can you conclude about the buffer systems that caused an increase in pH when the temperature dropped?
- (And the answer to the problem of pH varying with temperature faced by the researchers was to simply mix two of the different types of buffer solution until there was no significant change in pH during the cooling process.)
21.
 - A 0.1 mol L⁻¹ solution of acetic acid has a pH of 2.93. This changes to a pH of 4.74 when 0.1 mol of sodium acetate is added. Explain why the pH changes in this way.
 - The mixture of acetic acid and sodium acetate solutions produced in part a) has 0.5 mL of 0.1 mol L⁻¹ HCl added to it. Explain why the pH of the solution does not change significantly.
22.
 - Using only ammonium chloride and ammonia solution, describe how you could produce a buffer solution.
 - Explain how this buffer solution would 'work'.
 - If you wanted to decrease the pH of the buffer solution produced in part a) and you only have access to ammonium chloride and ammonia solution, what would you do? Explain your answer.
 - If you wanted to produce a solution with a greater buffer capacity than that produced in part a) describe the changes you would make to the production of the buffer.

23. Strenuous athletic activity can lead to the formation of lactic acid in muscles from the oxidation of glucose in a limited oxygen supply. The presence of the excess acid in the muscle tissue is claimed to interfere with muscular activity and it has been claimed that athletic performance is reduced.

Some athletes involved in short-distance events hyperventilate just before the race starts. Hyperventilation involves rapid, deep breathing. It is argued that hyperventilation increases oxygen flow to the bloodstream and expels carbon dioxide.

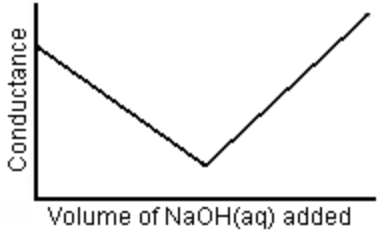
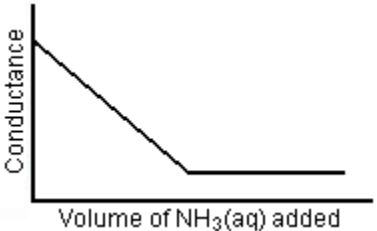
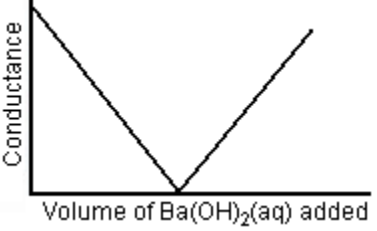
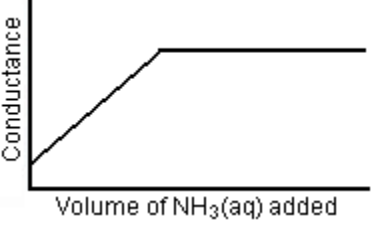
- Use the equations involving the two equilibria in the carbonic acid/hydrogencarbonate/ CO_2 buffer system to show what would be the effect of this rapid breathing on the position of equilibrium.
 - What would happen to the pH of blood as a result of hyperventilation?
 - Lactic acid has the formula $\text{C}_3\text{H}_6\text{O}_2$ and is a weak monoprotic acid. Write an equation to show what will happen to lactic acid when it dissolves in an aqueous medium such as blood.
 - How does hyperventilation help prevent the build up of excess H_3O^+ from lactic acid in the muscle tissue?
 - If the race starts later than the athlete anticipated, why might hyperventilation be dangerous?
24. One method that can be used to monitor the progress of a reaction or determine the equivalence point of a titration involves measuring the electrical conductivity of the reaction mixture.

In a solution, the electric current is carried by the anions and cations. The conductance of a solution is therefore a measure of the concentration and type of cations and anions present. Most ions have similar conductances; however, hydrogen ions and hydroxide ions have particularly high conductances, as shown in the table below.

Cation	Relative conductance	Anion	Relative conductance
H^+	349.8	OH^-	198.3
Na^+	50.1	Cl^-	76.3
K^+	73.5	NO_3^-	71.5
NH_4^+	73.5	SO_4^{2-}	159.6
Ba^{2+}	127.3	CH_3CO_2^-	40.9

When one solution is added to another, the conductivity of the reaction mixture will change as the type and proportion of ions present varies as the reaction proceeds. The changes in conductivity during four titrations are shown in the graphs on the next page. In each titration, the addition of the solutions from the burette was continued past the equivalence point.

For each graph (on the next page), explain in terms of the ions present before, during and after the reaction, why the conductance changes in the way shown. Also explain how you could use the graphs to determine when the equivalence point is reached for titration 1 and titration 2.

Titration number	Solution being titrated, i.e. solution present in the flask at the beginning of the titration	Solution added from the burette	Conductance curve of the reaction mixture
1	20 mL of approximately 0.1 mol L ⁻¹ HCl	0.1 mol L ⁻¹ NaOH	
2	20 mL of approximately 0.1 mol L ⁻¹ HCl	0.1 mol L ⁻¹ NH ₃	
3	20 mL of approximately 0.1 mol L ⁻¹ H ₂ SO ₄	0.1 mol L ⁻¹ Ba(OH) ₂	
4	20 mL of approximately 0.1 mol L ⁻¹ CH ₃ COOH	0.1 mol L ⁻¹ NH ₃	

Review exercise 8.1

1. Write balanced ionic equations, and give observations, for each of the following chemical reactions.
 - a) Sulfuric acid is added to a solution of lithium hydroxide.
 - b) Nitric acid is added to a solid copper(II) oxide.
 - c) Iron(II) hydroxide is reacted with a solution of hydrochloric acid.
 - d) Solid potassium oxide is added to a solution of sulfuric acid.
 - e) Nitric acid is added to calcium carbonate.
 - f) A solution of potassium hydrogencarbonate is added to hydrochloric acid.
 - g) Nitric acid is added to solid ammonium sulfide.(not in syllabus)
 - h) A solution of rubidium sulfite is reacted with sulfuric acid.(not in syllabus)
2. Write balanced ionic equations for the following oxidation and reduction reactions.
 - a) Zinc metal is added to dilute sulfuric acid solution.
 - b) Copper metal is added to a solution of silver nitrate.
 - c) Very dilute nitric acid is poured onto a piece of calcium.
 - d) Acidified potassium permanganate solution is added to a solution of iron(II) sulfate and the colour of the purple permanganate solution is observed to disappear.
 - e) Acidified potassium dichromate solution is added to a solution of sodium oxalate and the orange colour of the dichromate solution is observed to change to a deep green colour.
3. The halogens in group 17 form a displacement series. All the halogens are oxidants, but the oxidant strength decreases down the group. Fluorine is the strongest oxidant while astatine is the weakest. Use this information to predict whether or not the following combinations will react. If there would be no reaction, write 'No reaction'. Write an ionic equation for any predicted reactions.
 - a) Chlorine gas is bubbled through a solution of potassium iodide.
 - b) A solution of bromine in water is added to a solution of sodium fluoride.
 - c) A solution of bromine in water is added to a solution of ammonium iodide.
4. Aluminium hydroxide and magnesium hydroxide are both white solids, sparingly soluble in water. What chemical reaction can be used to distinguish between these two metal hydroxides?

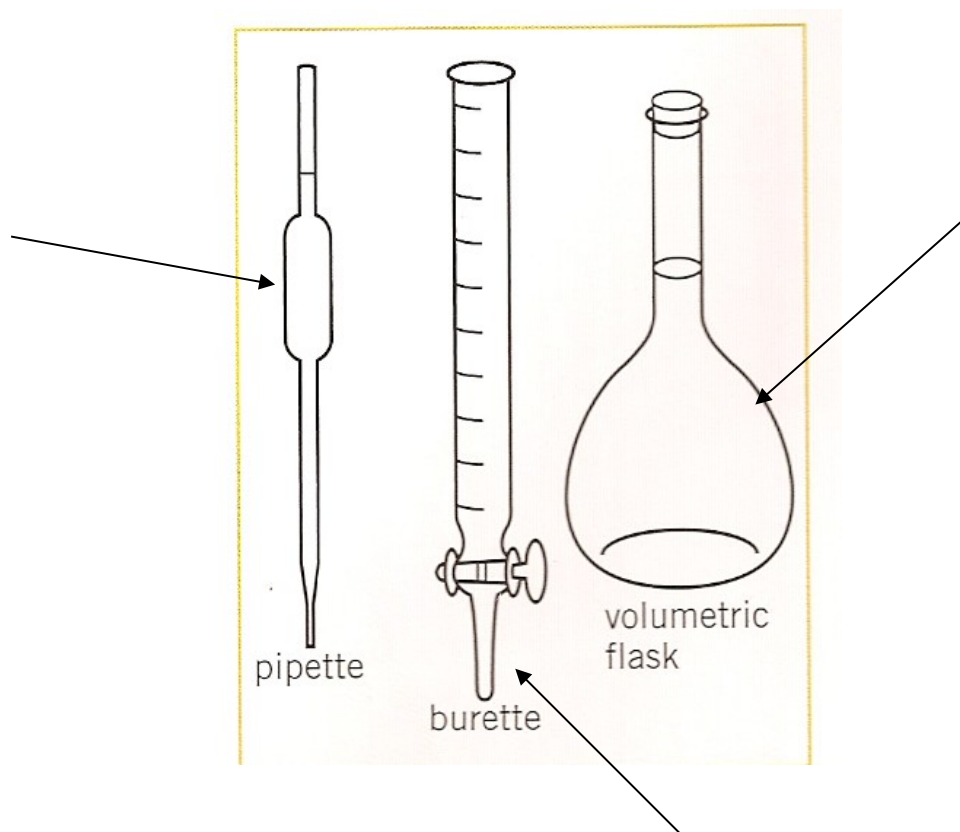
Review exercise 8.2

1. For each of the following state whether a gravimetric or volumetric analysis would be the most appropriate.
 - a) Determining the amount of water of crystallisation in a sample of hydrated barium chloride
 - b) Calculating the concentration of lead ions in a sample of water using the equation:
$$\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$$
 - c) Determining the percentage of ammonia in a household cleaner by reacting a solution of the cleaner with hydrochloric acid
 - d) Calculating the amount of sulfur dioxide dissolved in a sample of wine by oxidising it with a solution of potassium permanganate
2. A sample of lawn food is known to contain iron as a soluble iron(II) salt. A 0.5479 g sample of the lawn food was dissolved in dilute hydrochloric acid and then a little nitric acid was added. The mixture was boiled and turned an orange-brown colour. An excess of ammonia solution was added and the precipitate of iron(III) oxide, Fe_2O_3 , was collected and dried to constant mass. The mass of iron(III) oxide collected was 0.1832 g. Calculate the percentage by mass of iron in the lawn food.
3. Using your knowledge of chemical reactions, reactions rates and principles of equilibrium outline the particular characteristics of a chemical reaction that make it suitable for use in a titration.

Equipment

TABLE 8.2 EQUIPMENT USED IN VOLUMETRIC ANALYSIS

Pipette	<ul style="list-style-type: none">• Used to accurately deliver a specific, measured volume of solution, e.g. 20.00 mL.• The volume of solution delivered from the pipette is called an aliquot.• Prior to being used in an analysis, the pipette should be rinsed with a small amount of the solution being delivered from it in the titration.
Burette	<ul style="list-style-type: none">• Used to accurately deliver a variable volume of solution.• The volume of solution delivered from the burette is called the titre.• Prior to being used in an analysis, the burette should be rinsed with a small amount of the solution that is being delivered for the titre.
Volumetric flask	<ul style="list-style-type: none">• Holds on accurately known volume of a solution.• Used to make up a standard solution.• Should be rinsed with distilled or de-ionised water prior to use.
Analytical balance	<ul style="list-style-type: none">• Used to measure accurate masses of solids that are used to make up standard solutions.
Conical flask	<ul style="list-style-type: none">• Used to hold the reaction mixture during the titration.• Should be rinsed with distilled water before the aliquot is put into it.



Standard solution (ref page 244 of textbook)

A standard solution has an accurately known concentration.

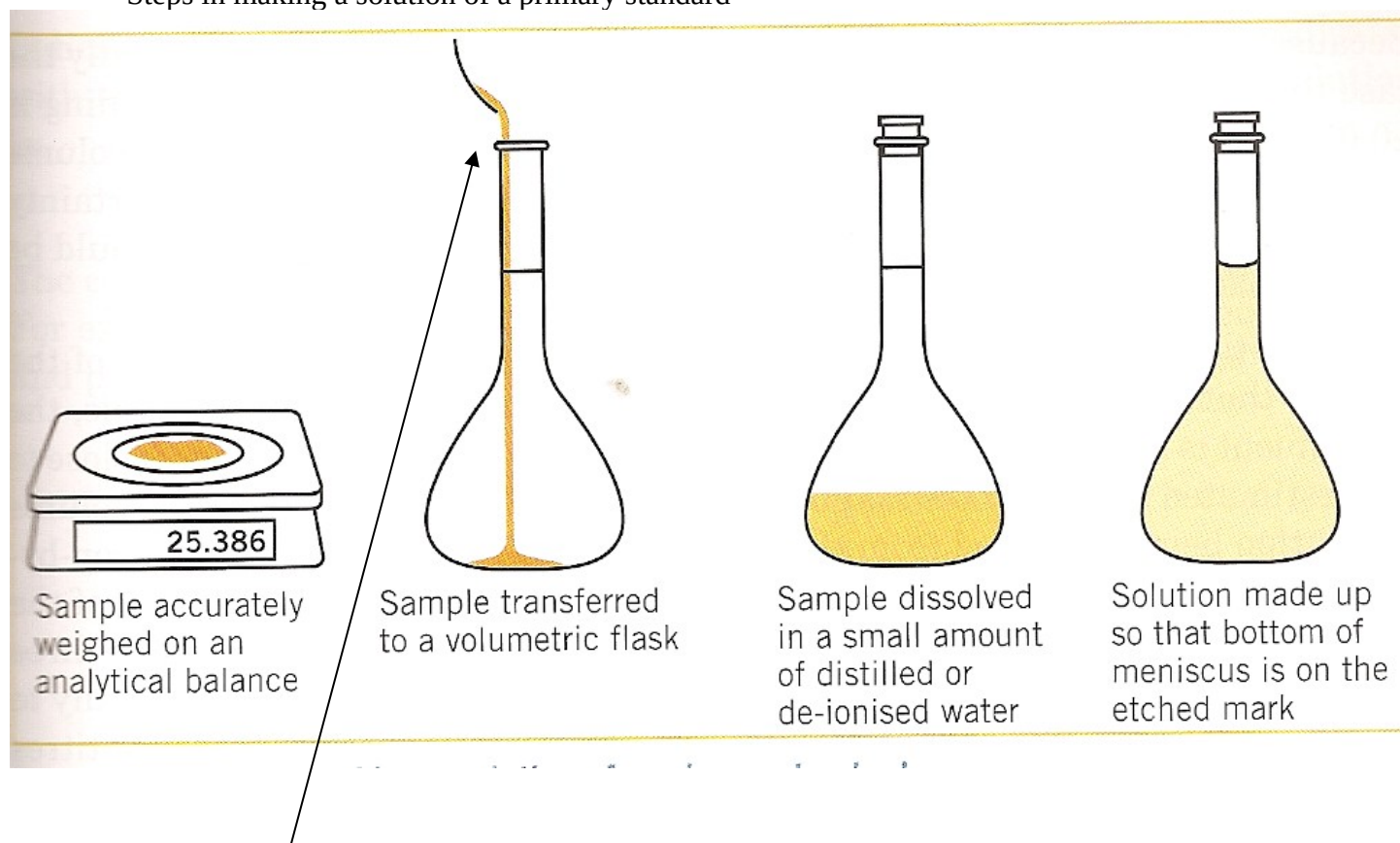
There are two types of standards that are used in titrations: **primary standards** and **secondary standards**.

A primary standard is made up using the mass of the solid to determine the number of moles dissolved.

Characteristics of good primary standard

- Obtained with a high degree of purity and has a known formula
- Undergoes reactions according to known chemical equations
- Must be stable eg not absorb water from atmosphere
- Relatively high formula mass to reduce weighing errors

Steps in making a solution of a primary standard



Why don't we transfer solid directly into volumetric flask?

Titration

Using a burette

- Rinse with the solution you are going to fill it with
- Do NOT have to fill to exactly 0 mark
- You take 2 readings – initial reading and final reading
- Read to 2 decimal places – the second decimal place is 0 or 5
- Do NOT leave funnel in the top during the titration
- Ensure burette is vertical



In titrations, it is essential to know when the reaction is complete. This point is termed the **equivalence point** (when stoichiometric amounts have reacted) and should be signalled by some visible change such as a colour change. In some cases, such as titrations involving permanganate ion this change is provided by the reactant itself, but in most cases it is necessary to use an extra reagent to indicate the equivalence point.

Such reagents are called indicators and must be carefully chosen so that the visible change the **end point** of the titration coincides as nearly as possible with the equivalence point of the reaction.

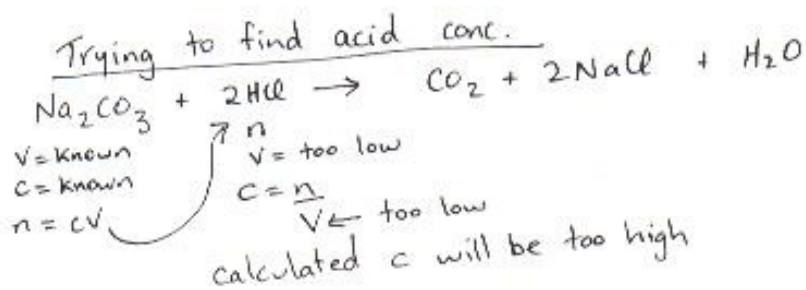
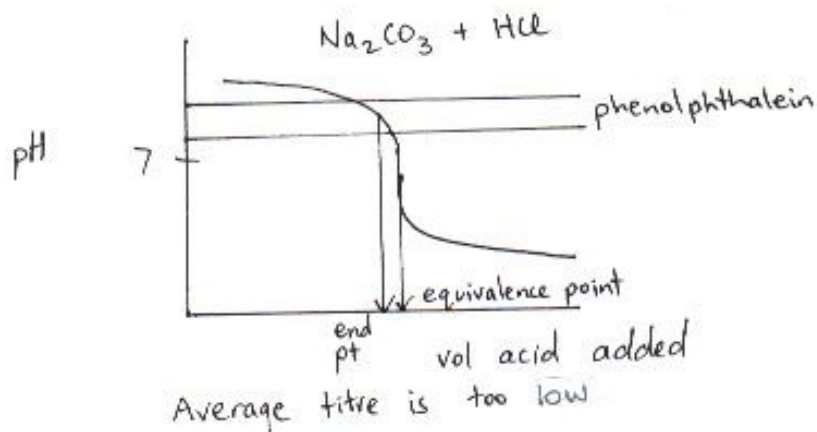
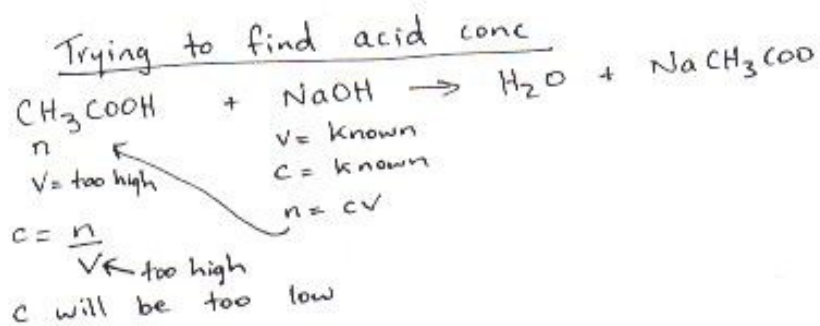
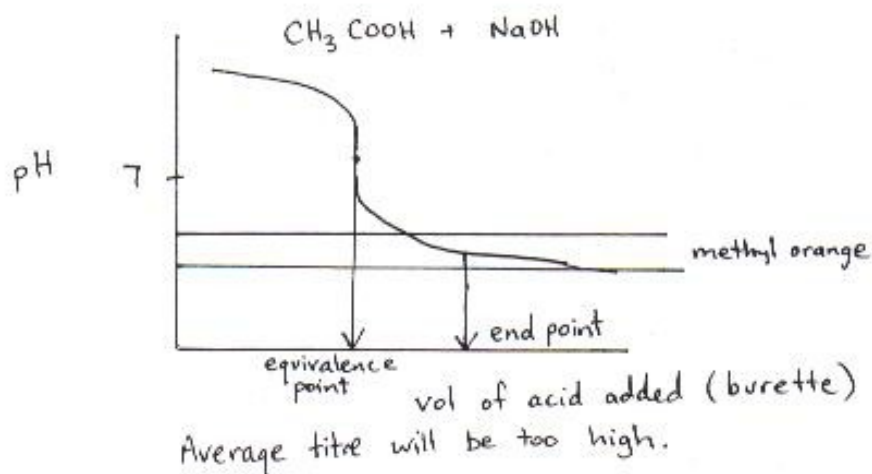
Acid-base indicators are substances which change colour according to the hydrogen ion concentration of the solution to which they are added. The indicators are themselves weak acids. One, or both, of the acid and its conjugate base are coloured.

INDICATOR	IN ACID	IN BASE
Litmus	red (below pH 4.7)	blue (above pH 8.1)
Phenolphthalein	colourless (below pH 8.4)	red (above pH 9.9).
Methyl Orange	red (below pH 3.1)	yellow (above pH 4.4)

These ranges may seem large, but near the **equivalence point**, the point at which the materials are equal, there is a large change in pH. The equivalence point may not occur at pH 7, neutral pH, so the appropriate pH indicator must be chosen for the type of acid and base being titrated.

The aim of the titration is to get the indicator end point as close as possible to the reaction equivalence point. This then requires that the indicator chosen is suitable for the reaction is the indicator endpoint is appropriate for the reaction equivalence point.

12 Chem Volumetric Analysis PowerPoint handout



Review exercise 8.3

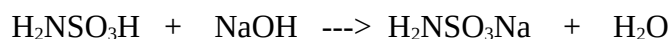
1. What is a standard solution?
2. Anhydrous sodium carbonate is a primary standard. What are the characteristics of this compound that allow it to be used as a primary standard?
3.
 - a) Two students are arguing over part of an experimental procedure. One claims that the volumetric flask into which the primary standard is to be placed must be dry before use. The other claims that as long as the flask has been thoroughly rinsed with distilled water, it doesn't need to be dry. How would you resolve this argument?
 - b) Later, these same two students argue over whether the conical flask into which the aliquot of a solution of known concentration is added needs to be dry. One states that it must be because any water in the flask would alter the concentration of the material dissolved in the aliquot. The other claims that the presence of water won't make a difference to the final titre. Both students are correct so how do you resolve this argument?
4. In preparing a standard solution of oxalic acid, $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$, in a 250 mL volumetric flask, you record the following measurements.
Mass of watch glass + oxalic acid = 52.37 g
Mass of watch glass after transfer to the volumetric flask = 48.98 g
 - a) Why is the weighing procedure described superior to simply recording the mass of the empty watch glass and the mass of watch glass + oxalic acid?
 - b) Calculate the concentration of the oxalic acid solution produced.
5. What is the difference between the equivalence point and the end point in a titration?
6. In a titration between acetic acid and potassium hydroxide, a student records the following in his laboratory notebook.
 - (i) The burette was rinsed with 5 mL of the acetic acid solution of unknown concentration.
This was discarded before the burette was then filled with the acetic acid solution.
 - (ii) The pipette was rinsed with water and then used to transfer 20.00 mL of the standard potassium hydroxide solution to a conical flask. Three drops of methyl orange were added to the flask to produce a yellow coloured solution.
 - (iii) The acetic acid was added to the flask from the burette until the colour of the solution changed from yellow to red. The volume on the burette was recorded and the titre calculated.

The experimental report then goes on to detail the steps in the calculation of the concentration of the acid from this result.

This student has made several mistakes in this procedure. Identify these mistakes.

ACID-BASE TITRATION CALCULATIONS

1. 23.45 mL of 0.275 M sodium hydroxide was used to titrate against mL of acetic acid. What was the concentration of acetic acid?
2. 17.05 mL of 0.247 M barium hydroxide was used to titrate against 10 mL of nitric acid. What was the concentration of nitric acid?
3. 35.79 mL of 0.275 M sodium hydroxide was used to titrate against 15 mL of sulfuric acid. What was the concentration of sulfuric acid in g L^{-1} ?
4. The active ingredient of a drain cleaning powder is sodium hydroxide. A sample of this powder weighing 11.4 g is dissolved in water and made up to 250.0 mL in a volumetric flask. A 20.0 mL sample of the resulting solution required 18.75 mL of 0.320 mol L^{-1} sulfuric acid for complete neutralisation. You may assume that none of the other ingredients of the drain cleaning powder react with the sulfuric acid.
 - a) How many moles of sulfuric acid were used ?
 - b) What is the molar concentration of the sodium hydroxide solution?
 - c) How many grams of sodium hydroxide are contained in 1.00 kg of the drain cleaning powder?
5. Sulfamic acid is a strong acid. It is a good primary standard substance for standardising alkalis. The reaction of sulfamic acid with sodium hydroxide is as follows:



2.50 g of sulfamic acid were dissolved in distilled water and diluted in a 250.0 mL volumetric flask. It was found that 20.5 mL of sulfamic acid solution neutralized 20.0 mL of sodium hydroxide solution. Calculate the molarity of the sodium hydroxide solution.

(0.106 mol L^{-1})

6. 5.10 g of potassium hydrogen phthalate were dissolved in distilled water and diluted to 250.0 mL in a volumetric flask. 20.0 mL of this solution were titrated with 0.105 mol L^{-1} sodium hydroxide solution, using phenolphthalein as indicator. The volume of sodium hydroxide required to reach end point was 19.05 mL. Calculate the formula mass of potassium hydrogen phthalate, given that 1 mole of the substance produced 1 mole of H^+ ions.

(204)

7. 1.10 g of a carbonate of formula, MCO_3 , was dissolved in 25.0 mL of 1.0 mol L^{-1} hydrochloric acid. The resultant solution was titrated with 0.10 mol L^{-1} sodium hydroxide to determine the excess acid. A volume of 30.0 mL of the hydroxide was needed for neutralisation.
 - a) Calculate
 - i) the formula mass of MCO_3 . (100)
 - ii) the relative atomic mass of M. (40)
 - b) What element is M? (Ca)

8. 1.50 g of a sample of sodium carbonate decahydrate was dissolved in 25.0 mL (an excess) of 0.50 mol L⁻¹ hydrochloric acid. The resultant solution was titrated with 0.10 mol L⁻¹ sodium hydroxide solution. 25.0 mL of the hydroxide solution were needed for neutralisation. Calculate the percentage purity of the hydrated salt.

(95.3%)

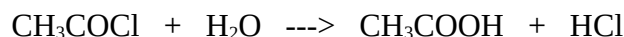
9. 6.0 mL of a laboratory sample of glacial acetic acid were transferred to a 500.0 mL volumetric flask and were then made up to volume with distilled water. 20.0 mL of the diluted solution were titrated with 0.20 mol L⁻¹ potassium hydroxide solution using phenolphthalein as indicator. It was found that 20.9 mL of the hydroxide were needed for neutralization. Calculate

- a) the mass of acetic acid in 1 litre of glacial acetic acid. (1045 g L⁻¹)
 b) the molarity of glacial acetic acid. (17.4 mol L⁻¹)

10. 3.00 g of sodium oxide, Na₂O, and 2.00g of potassium oxide, K₂O, were dissolved in exactly 250.0 mL of distilled water. 25.0 mL aliquots of the solution were titrated with 0.50 mol L⁻¹ hydrochloric acid using methyl red as indicator. Calculate the volume of hydrochloric acid required for neutralisation.

(27.8 mL)

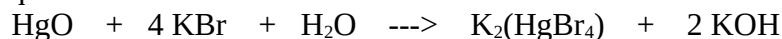
11. Acetyl chloride (also known as ethanoyl chloride) is completely hydrolysed in water to form ethanoic acid and hydrochloric acid:



A certain mass of acetyl chloride was dissolved in distilled water and then diluted to 250.0 mL in a volumetric flask. 20.0 mL of this solution required 21.50 mL of 0.105 mol L⁻¹ sodium hydroxide solution for neutralisation. Calculate the mass of acetyl chloride dissolved.

(1.11 g)

12. Mercury (II) oxide reacts quantitatively with potassium bromide according to the equation:

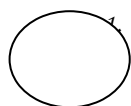


The potassium hydroxide generated by the reaction can be titrated with a solution of hydrochloric acid. A 2.00 g sample of impure mercury (II) oxide was dissolved in a solution of potassium bromide and was then diluted to 250.0 mL in a volumetric flask. 20.0 mL of the solution, when titrated with 0.100 mol L⁻¹ hydrochloric acid, required 14.0 mL of hydrochloric acid for neutralisation. Calculate the percentage purity of mercury (II) oxide. (Assume that the impurity had no reaction with HCl or NaOH).

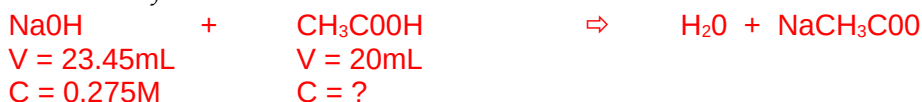
(94.8%)

13. When dissolved in water, one mole of a hydrated salt, $B \cdot xH_2O$, produces two moles of hydroxide ions. 5.350g of the hydrated salt were dissolved in water and then diluted to 250.0 mL in a volumetric flask. Using methyl orange as indicator, 20.0 mL aliquots of the solution required an average of 22.0 mL of 0.102 mol L^{-1} hydrochloric acid for neutralization.
- Calculate the molarity of the solution. $(5.60 \times 10^{-2} \text{ mol L}^{-1})$
 - What is the formula mass of the hydrated salt. (381 g)
 - If the anhydrous salt has a formula mass of 202, calculate the value of x . (10)
14. A student analysed a brand of vinegar bought from a supermarket in the following way:
20.12 g of the vinegar were diluted to 250.0 mL in a volumetric flask. Using phenolphthalein as indicator, 20.0 mL of the solution required 11.50 mL of 0.105 mol L^{-1} sodium hydroxide solution for neutralisation.
- Calculate the number of moles of acetic acid in 20.0 mL of the solution. $(1.21 \times 10^{-3} \text{ mol L}^{-1})$
 - What is the molarity of the diluted vinegar solution? $(6.05 \times 10^{-2} \text{ mol L}^{-1})$
 - Calculate the mass of acetic acid in 250 mL of the diluted solution. (0.980 g)
 - What is the percentage by mass of acetic acid in the original vinegar solution? (4.51%)

ACID-BASE TITRATION CALCULATIONS



23.45 mL of 0.275 M sodium hydroxide was used to titrate against mL of acetic acid. What was the concentration of acetic acid?

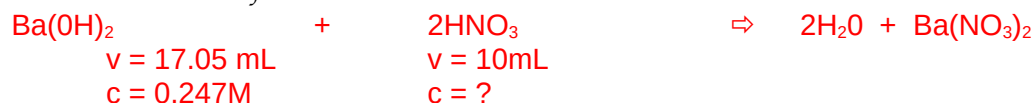


$$n = CV = (0.275)(0.02345) = 6.45 \times 10^{-3}$$

$$n_{\text{CH}_3\text{COOH}} = 1 \times \text{moles NaOH} = 6.45 \times 10^{-3}$$

$$C_{\text{CH}_3\text{COOH}} = \frac{n}{V} = \frac{6.45 \times 10^{-3}}{0.02} = 0.322\text{M (3sf)}$$

2. 17.05 mL of 0.247 M barium hydroxide was used to titrate against 10 mL of nitric acid. What was the concentration of nitric acid?

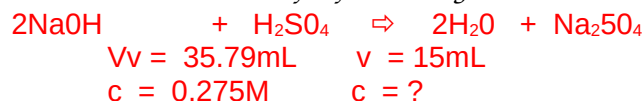


$$n = cv = (0.247)(0.01705) = 4.211 \times 10^{-3}$$

$$\text{moles HNO}_3 = 2 \times \text{moles Ba(OH)}_2 = 2 \times 4.211 \times 10^{-3} = 8.423 \times 10^{-3}$$

$$C_{\text{HNO}_3} = \frac{n}{v} = \frac{8.423 \times 10^{-3}}{(0.01)} = 0.842\text{M}$$

3. 35.79 mL of 0.275 M sodium hydroxide was used to titrate against 15 mL of sulfuric acid. What was the concentration of sulfuric acid in gL⁻¹?



$$n = CV = (0.275)(0.03579) = 9.84 \times 10^{-3}$$

$$\text{moles H}_2\text{SO}_4 = \frac{1}{2} \times \text{moles NaOH} = \frac{1}{2} \times 9.84 \times 10^{-3} = 4.92 \times 10^{-3}$$

$$C_{\text{H}_2\text{SO}_4} = \frac{4.92 \times 10^{-3}}{(0.015)} = 0.328\text{M}$$

$$= 0.328 \times (2 \times 1.008 + 32.06 + 4 \times 16)$$

$$= 0.328 \times 98.076$$

$$= 32.2\text{ gL}^{-1}$$

4. The active ingredient of a drain cleaning powder is sodium hydroxide. A sample of this powder weighing 11.4 g is dissolved in water and made up to 250.0mL in a volumetric flask. A 20.0 mL sample of the resulting solution required 18.75 mL of 0.320 mol L⁻¹ sulfuric acid for complete neutralisation. You may assume that none of the other ingredients of the drain cleaning powder react with the sulfuric acid.

- How many moles of sulfuric acid were used ?
- What is the molar concentration of the sodium hydroxide solution?
- How many grams of sodium hydroxide are contained in 1.00 kg of the drain cleaning powder?

$$V = 250\text{mL}$$

$$\text{mass} = 11.4\text{g (not all NaOH)}$$



a) $V = 20\text{mL}$ $V = 18.75\text{mL}$
 $c = 0.320\text{M}$
 $n = cv$
 $= (0.320)(0.01875)$
 $= 6 \times 10^{-3} \text{ mol of H}_2\text{SO}_4$

b) $\text{moles NaOH} = \frac{2}{1} \times \text{moles H}_2\text{SO}_4 = \frac{2}{1} \times 6 \times 10^{-3} = 0.012$

$C = \frac{n}{V} = \frac{0.12}{0.02} = 0.6\text{M}$

c) $\text{moles NaOH in 250mL of solution}$
 $n = CV = (0.6)(0.25) = 0.15\text{mol}$
 $\text{mass NaOH} = 0.15 \times (22.99 + 16 + 1.008)$
 $= 0.15 \times 39.998 = 6.00\text{g}$

$\% \text{ NaOH in drain cleaner} = \frac{6.00}{11.4} \times 100$
 $= 52.6\%$

grams of NaOH in 1kg of drain cleaner

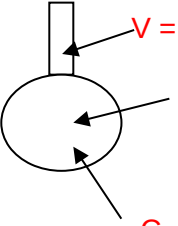
$\frac{52.6}{100} \times 1000 = 526\text{g}$

(remember there are 1000g in 1kg)

5. Sulfamic acid is a strong acid. It is a good primary standard substance for standardising alkalis. The reaction of sulfamic acid with sodium hydroxide is as follows:

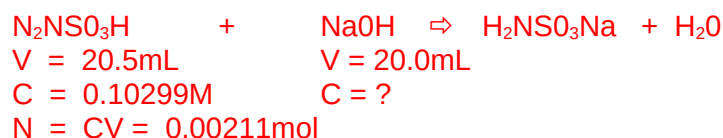
$$\text{H}_2\text{NSO}_3\text{H} + \text{NaOH} \rightarrow \text{H}_2\text{NSO}_3\text{Na} + \text{H}_2\text{O}$$

2.50 g of sulfamic acid were dissolved in distilled water and diluted in a 250.0 mL volumetric flask. It was found that 20.5 mL of sulfamic acid solution neutralized 20.0 mL of sodium hydroxide solution. Calculate the molarity of the sodium hydroxide solution. (0.106 mol L⁻¹)

5. 

$V = 250\text{mL}$
 $\text{mass} = 2.50\text{g}$
 $n_{\text{H}_2\text{NSO}_3\text{H}} = \frac{2.50}{(2 \times 1.008) + (14.01) + (32) + (3 \times 16) + (1.008)}$
 $= 0.0257\text{mol}$

$C = \frac{n}{v} = \frac{0.0257}{0.25} = 0.10299\text{M}$

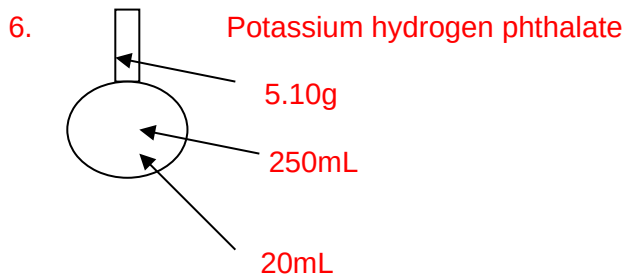


$$n_{\text{NaOH}} = \frac{1}{1} \times n_{\text{H}_2\text{NSO}_3\text{H}} = \frac{1}{1} \times 0.00211$$

$$c_{\text{NaOH}} = \frac{n}{V} = \frac{0.00211}{0.02} = 0.1055\text{M or } 0.1055 \text{ mol L}^{-1}$$

$$= \underline{0.106\text{M}} \text{ (to 3sig fig)}$$

6. 5.10 g of potassium hydrogen phthalate were dissolved in distilled water and diluted to 250.0 mL in a volumetric flask. 20.0 mL of this solution were titrated with 0.105 mol L⁻¹ sodium hydroxide solution, using phenolphthalein as indicator. The volume of sodium hydroxide required to reach end point was 19.05 mL. Calculate the formula mass of potassium hydrogen phthalate, given that 1 mole of the substance produced 1 mole of H⁺ ions.



$$\begin{aligned} V &= 20\text{mL} & V &= 19.05\text{mL} \\ C &= 0.105\text{M} \\ N &= CV \\ &= (0.105)(0.01905) \\ &= 2.000 \times 10^{-3} \end{aligned}$$

$$\begin{aligned} \text{moles pot hydrogen phthalate} &= \frac{1}{1} \times \text{moles NaOH} \\ \text{(in 20mL sample)} &= 2.000 \times 10^{-3} \end{aligned}$$

$$\begin{aligned} \text{moles potassium hydrogen phthalate in 250ml} \\ &= 2.000 \times 10^{-3} \times \frac{250}{20} \end{aligned}$$

$$= 0.025\text{mol}$$

$$n = \frac{\text{mass}}{\text{mass of 1 mole (formula mass)}}$$

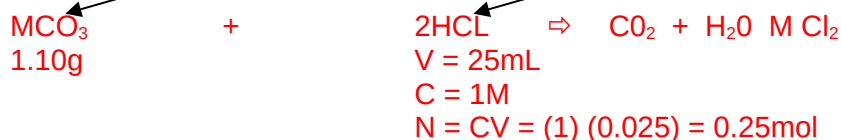
$$0.025 = \frac{5.10}{\text{formula mass}}$$

$$\Rightarrow \text{formula mass} = \underline{204}$$

7. 1.10 g of a carbonate of formula, MCO_3 , was dissolved in 25.0 mL of 1.0 mol L^{-1} hydrochloric acid. The resultant solution was titrated with 0.10 mol L^{-1} sodium hydroxide to determine the excess acid. A volume of 30.0 mL of the hydroxide was needed for neutralisation.

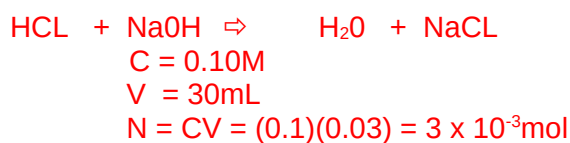
- a) Calculate i) the formula mass of MCO_3 .
 ii) the relative atomic mass of M.
 b) What element is M?

7. Equ 1 all of this reacts not all of this reacts (in excess)



The excess HCL is then reacted as follows

Equ 2



$$\text{Moles HCL} = 1 \times \text{moles NaOH} = 3 \times 10^{-3}$$

the amount of HCL that reacted in Equ 1
 $= 0.025 - 3 \times 10^{-3} = 0.022 \text{ mol}$

$$\begin{array}{l} \text{moles MCO}_3 = \frac{1}{2} \text{ moles HCL} \\ = \frac{1}{2} \times 0.022 = 0.011 \end{array}$$

- a) i) formula mass of MCO_3

$$\begin{array}{l} 0.011 = \frac{1.10}{\text{formula mass}} \\ \Rightarrow \text{formula mass} = 100 \end{array}$$

- ii) atomic mass of M

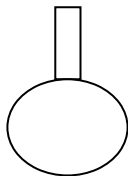
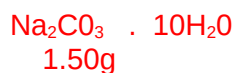
$$\begin{array}{l} \text{MCO}_3 = 100 = M + 12 + (3 \times 16) \\ \Rightarrow M = 40 \end{array}$$

- b) element M is Ca

8. 1.50 g of a sample of sodium carbonate decahydrate was dissolved in 25.0 mL (an excess) of 0.50 mol L⁻¹ hydrochloric acid. The resultant solution was titrated with 0.10 mol L⁻¹ sodium hydroxide solution. 25.0 mL of the hydroxide solution were needed for neutralisation. Calculate the percentage purity of the hydrated salt.

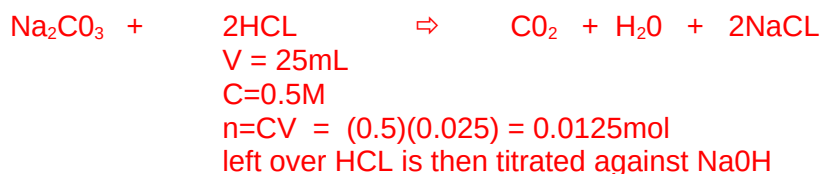
8. note decahydrate means 10 lots of water of crystallisation

hydrated salt means that there is water of crystallisation in the solid

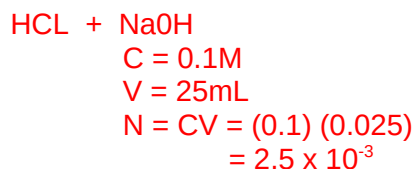


V = 25mL of HCL

Equ 1



Equ 2



$$\text{moles HCL} = \frac{1}{1} \times \text{moles NaOH} = 2.5 \times 10^{-3}$$

moles HCL that reacted with Na₂CO₃

$$= 0.0125 - 2.5 \times 10^{-3} = 0.01 \text{ mol}$$

$$\begin{aligned} \text{moles Na}_2\text{CO}_3 &= \frac{1}{2} \times \text{moles HCL} \\ &= \frac{1}{2} \times 0.01 = 5 \times 10^{-3} \\ \text{moles Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O} &= 5 \times 10^{-3} \end{aligned}$$

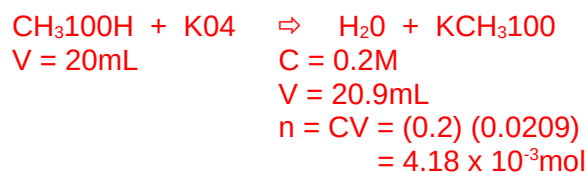
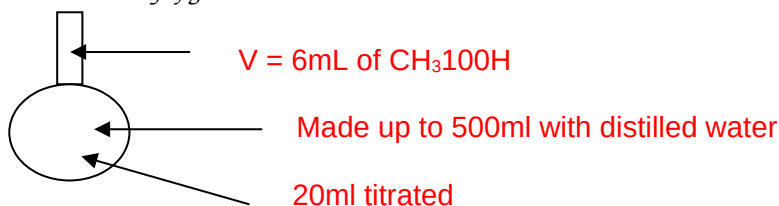
$$\begin{aligned} \text{mass Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} &= 5 \times 10^{-3} \times (2 \times 22.99 + 12.01 + 3 \times 16 + 20 \times 1.008 + 10 \times 16) \\ &= 5 \times 10^{-3} \times 286.15 \\ &= 1.43\text{g} \end{aligned}$$

$$\% \text{ purity} = \frac{\text{mass pure}}{\text{mass impure}} \times 100$$

$$= \frac{1.43}{1.50} \times 100 = 95.3\%$$

9. 6.0 mL of a laboratory sample of glacial acetic acid were transferred to a 500.0 mL volumetric flask and were then made up to volume with distilled water. 20.0 mL of the diluted solution were titrated with 0.20 mol L⁻¹ potassium hydroxide solution using phenolphthalein as indicator. It was found that 20.9 mL of the hydroxide were needed for neutralization. Calculate
- a) the mass of acetic acid in 1 litre of glacial acetic acid. (1045 g L⁻¹) b) the molarity of glacial acetic acid.

9.



$$\text{moles CH}_3\text{COOH} = \frac{1}{1} \times \text{moles KOH}$$

$$= 4.18 \times 10^{-3}$$

$$C = \frac{n}{V} = \frac{4.18 \times 10^{-3}}{0.02} = 0.209\text{M} \text{ this is the conc of the diluted glacial acetic acid}$$

- b) find conc of undiluted acid

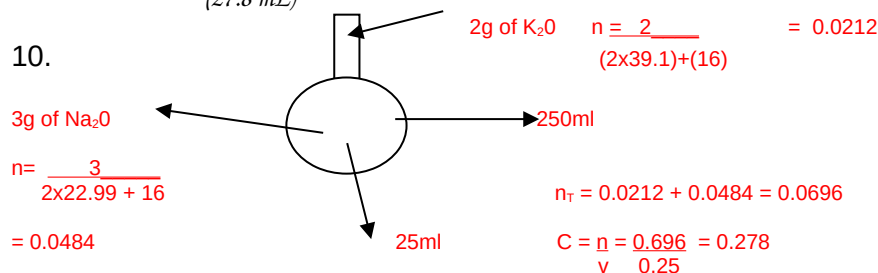
$$C_1 V_1 = C_2 V_2$$

$$C_1 (6) = (0.209)(500)$$

$$\Rightarrow C_1 = 17.4 \text{ mol L}^{-1}$$

- a) $n = CV = (17.4)(0.006) = 0.1044 \text{ mol}$
- $$\text{mass}_{\text{CH}_3\text{COOH}} = 0.1044 \times (12.01 + 3 \times 1.008 + 12.01 + 2 \times 16 + 1.008)$$
- $$= 0.1044 \times 60.052 = 6.27\text{g in 500ml}$$
- $$12.54\text{g in 1 litre}$$

10. 3.00 g of sodium oxide, Na_2O , and 2.00 g of potassium oxide, K_2O , were dissolved in exactly 250.0 mL of distilled water. 25.0 mL aliquots of the solution were titrated with 0.50 mol L^{-1} hydrochloric acid using methyl red as indicator. Calculate the volume of hydrochloric acid required for neutralisation.
(27.8 mL)



oxide solution where x represents Na or K



$$V = 25 \text{ mL} \quad V = ?$$

$$C = 0.278 \text{ M} \quad C = 0.50 \text{ M}$$

$$N = CV$$

$$= (0.278)(0.025)$$

$$= 6.96 \times 10^{-3}$$

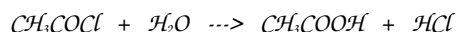
$$\text{moles HCl} = \frac{2}{1} \times \text{moles X}_2\text{O}$$

$$= \frac{2}{1} \times 6.96 \times 10^{-3}$$

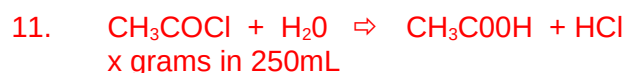
$$= 0.0139$$

$$V_{\text{HCl}} = \frac{n}{C} = \frac{0.0139}{0.5} = 27.8 \text{ mL}$$

11. Acetyl chloride (also known as ethanoyl chloride) is completely hydrolysed in water to form ethanoic acid and hydrochloric acid:



A certain mass of acetyl chloride was dissolved in distilled water and then diluted to 250.0 mL in a volumetric flask. 20.0 mL of this solution required 21.50 mL of 0.105 mol L^{-1} sodium hydroxide solution for neutralisation. Calculate the mass of acetyl chloride dissolved.



$$20 \text{ mL}$$

$$V = 21.50 \text{ mL}$$

$$C = 0.105 \text{ M}$$

$$N = CV = (0.105)(0.0215) = 2.257 \times 10^{-3} \text{ mol}$$

$$\text{Moles CH}_3\text{COCl (in 20 mL)} = \frac{1}{2} \times \text{moles NaOH}$$

$$= 1.1285 \times 10^{-3}$$

$$\text{Moles CH}_3\text{COCl (in 250 mL)} = 1.1285 \times 10^{-3} \times \frac{250}{20}$$

$$= 0.0141$$

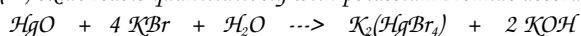
$$\text{mass CH}_3\text{COCl} = 0.0141 \times (12.01 + 3 \times 1.008 + 12.01 + 16 + 35.45)$$

$$= 0.0141 \times 78.494$$

$$= 1.107\text{g}$$

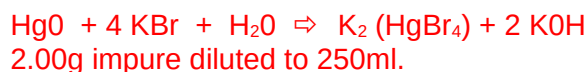
$$= 1.11\text{g (3 sf)}$$

12. Mercury (II) oxide reacts quantitatively with potassium bromide according to the equation:



The potassium hydroxide generated by the reaction can be titrated with a solution of hydrochloric acid. A 2.00 g sample of impure mercury (II) oxide was dissolved in a solution of potassium bromide and was then diluted to 250.0 mL in a volumetric flask. 20.0 mL of the solution, when titrated with 0.100 mol L⁻¹ hydrochloric acid, required 14.0 mL of hydrochloric acid for neutralisation. Calculate the percentage purity of mercury (II) oxide. (Assume that the impurity had no reaction with HCl or KOH).

12 Equ 1



Equ 2



$$v = 20\text{mL}$$

$$c = 0.100\text{M}$$

$$v = 14.0\text{mL}$$

$$n = CV = (0.1)(0.014) = 1.4 \times 10^{-3}$$

$$\text{Moles KOH} = \frac{1}{1} \times \text{moles HCl} = 1.4 \times 10^{-3}$$

$$c = \frac{n}{v} = \frac{1.4 \times 10^{-3}}{0.02} = 0.07\text{M}$$

find moles KOH in 250mL

$$n = cv = (0.07)(0.25) = 0.0175\text{mol}$$

From Equ 1

$$\text{Moles HgO} = \frac{1}{2} \times \text{moles KOH} = \frac{1}{2} \times 0.0175$$

$$= 8.75 \times 10^{-3}$$

$$\text{Mass HgO} = 8.75 \times 10^{-3} \times (200.6 + 16)$$

$$= 1.895\text{g}$$

$$\% \text{ purity} = \frac{\text{mass pure}}{\text{mass impure}} \times 100$$

$$= \frac{1.895}{2} \times 100$$

$$= 94.8\%$$

13. When dissolved in water, one mole of a hydrated salt, B.xH₂O, produces two moles of hydroxide ions. 5.350g of the hydrated salt were dissolved in water and then diluted to 250.0 mL in a volumetric flask. Using methyl orange as indicator, 20.0 mL aliquots of the solution required an average of 22.0 mL of 0.102 mol L⁻¹ hydrochloric acid for neutralization.

a) Calculate the molarity of the solution.

b) What is the formula mass of the hydrated salt. (381 g)

c) *If the anhydrous salt has a formula mass of 202, calculate the value of x , (10)*

13. a) B. \times H₂O \Rightarrow 2 OH⁻
5.350g
in 250mL

OH⁻ + HCL
V = 20mL V = 22.0mL
C = 0.102 mol L⁻¹
N = CV = (0.102) (0.022)
= 2.244 \times 10⁻³

$$\text{moles OH}^- = \frac{1}{1} \times \text{mole HCL} = 2.244 \times 10^{-3}$$

$$C = \frac{n}{V} = \frac{2.244 \times 10^{-3}}{0.02} = 0.1122\text{M}$$

$$\begin{array}{l} \text{Moles OH}^- \\ \text{In 250ml} \end{array} \quad n = CV = (0.1122) (0.25) = 0.02805$$

$$\begin{aligned} \text{Moles B. } \times \text{ H}_2\text{O} &= \frac{1}{2} \times \text{moles OH}^- \\ &= \frac{1}{2} \times 0.02805 = 0.0140 \end{aligned}$$

$$C_{\text{B} \times \text{H}_2\text{O}} = \frac{n}{V} = \frac{0.0140}{0.25} = 0.0561 \text{ M}$$

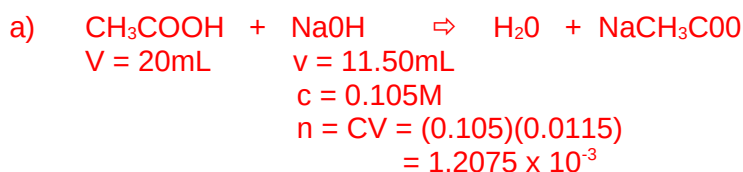
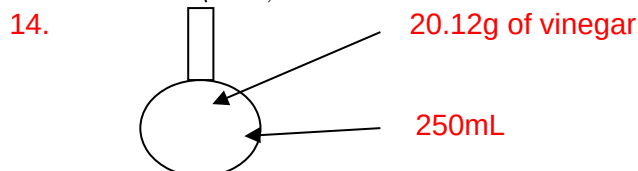
b) mass B. \times H₂O = 5.350g
 $n = 0.0140 = \frac{\text{mass}}{\text{formula mass}} = \frac{5.350}{\text{formula mass}}$

$$\Rightarrow \text{formula mass} = \frac{5.350}{0.0140} = 382\text{g}$$

13. c)

$$\begin{aligned} 382 &= B + x \cdot 18 \\ 382 &= 202 + 18 \\ 180 &= x \cdot 18 \Rightarrow x = 10 \end{aligned}$$

14. A student analysed a brand of vinegar bought from a supermarket in the following way:
 20.12 g of the vinegar were diluted to 250.0 mL in a volumetric flask. Using phenolphthalein as indicator, 20.0 mL of the solution required 11.50 mL of 0.105 mol L⁻¹ sodium hydroxide solution for neutralisation.
- Calculate the number of moles of acetic acid in 20.0 mL of the solution.
 - What is the molarity of the diluted vinegar solution?
 - Calculate the mass of acetic acid in 250 mL of the diluted solution.
 - What is the percentage by mass of acetic acid in the original vinegar solution? (4.51%)



$$\text{moles CH}_3\text{COOH} = \frac{1}{1} \times \text{mole NaOH} = 1.2075 \times 10^{-3}$$

b) $C = \frac{n}{v} = \frac{1.2075 \times 10^{-3}}{0.02} = 0.060375$

c) moles CH₃COOH in 250mL
 $n = CV = (0.060375)(0.25)$
 $= 0.0151$

$$\text{mass CH}_3\text{COOH} = 0.0151 \times (12.01 + 3 \times 1.008 + 12.01 + 2 \times 16 + 1.008)$$

$$= 0.0151 \times 60.052 = 0.906\text{g}$$

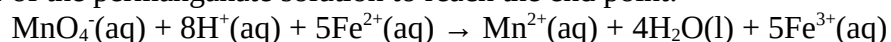
d) % by mass = $\frac{\text{mass CH}_3\text{COOH}}{\text{mass vinegar}} \times 100$

$$= \frac{0.906}{20.12} \times 100 = 4.77\%$$

Review exercise 8.4

1. Some photographic developing agents contain a base that provides hydroxide ions. A 10.00 mL sample of developer was diluted to 100.0 mL in a volumetric flask. 20.00 mL of this diluted solution was titrated with 0.0900 mol L⁻¹ sulfuric acid and required 21.40 mL to reach a methyl orange end point. Calculate the concentration, in mol L⁻¹, of the hydroxide ions in the original developer. (Hint: Sulfuric acid is diprotic.)
2. A patient suffering from indigestion provides a sample of gastric juices for analysis. A 40.00 mL sample of gastric juice was titrated with 0.1050 mol L⁻¹ sodium hydroxide to a phenolphthalein end point. The volume of sodium hydroxide solution required was 17.25 mL. Calculate the pH of the gastric juice in the patient.
3. A sample of iron ore having a mass of 4.60 g was treated in sulfuric acid to convert the iron in the ore to a mixture of soluble Fe²⁺ and Fe³⁺ ions. A solution of tin(II) chloride was carefully added to just convert all the Fe³⁺ ions to Fe²⁺. Excess Sn²⁺ was removed before the next step.

The resulting solution, assumed to be a solution containing Fe²⁺(aq), was diluted to 250.0 mL in a volumetric flask. A 25.00 mL sample of the diluted iron(II) solution was titrated with 0.0278 mol L⁻¹ potassium permanganate and required 35.6 mL of the permanganate solution to reach the end point:



Calculate the percentage of Fe₂O₃ present in the ore, assuming that this is the only iron-containing compound present.

4. A 4.674 g sample of limestone (impure CaCO₃) was treated with 375 mL of 0.261 mol L⁻¹ hydrochloric acid. After the reaction was over, the mixture was filtered and a 25.0 mL sample of the filtrate was titrated with 0.0517 mol L⁻¹ Ba(OH)₂ to the phenolphthalein end point. The volume of the barium hydroxide solution required was 17.62 mL.

Calculate the percentage, by mass, of CaCO₃ in the limestone sample.

QUESTIONS

1. Write ionic equations for the following reactions.
 - a) A solution of sulfuric acid is added to solid magnesium oxide.
 - b) Ammonia gas is bubbled through hydrochloric acid.
 - c) Aluminium oxide is added to a solution of nitric acid.
 - d) A solution of potassium sulfite is mixed with hydrochloric acid.(not in syllabus)
 - e) Solid lithium hydroxide is added to a solution of sulfuric acid.
 - f) A solution of nitric acid is added to solid iron(III) sulfide.(not in syllabus)
 - g) A solution of acetic acid is added to solid calcium carbonate.
 - h) Aluminium oxide is warmed with a solution of potassium hydroxide.
 - i) Acidified potassium permanganate solution is added to a solution of sodium oxalate.

2. For each of the following write observations and provide an ionic equation if relevant. For those combinations that do not result in a chemical reaction, write 'no reaction'.
 - a) A solution of ammonia is added to a solution of acetic acid.
 - b) A solution of calcium hydroxide is added to a solution of sodium hydroxide.
 - c) Zinc is added to a dilute sulfuric acid solution.
 - d) Nitric acid is added to solid copper(II) carbonate.
 - e) Copper is added to a solution of silver nitrate.
 - f) Solid chromium(III) oxide is added to hydrochloric acid.
 - g) Sulfuric acid is added to a solution of potassium hydrogencarbonate.

3. DDT (from its non-IUPAC name, dichlorodiphenyltrichloroethane) is a pesticide that was banned from general use several decades ago because of its effects on animals further up the food chain. A soil sample is known to be contaminated with DDT and was treated to convert all of the chlorine in the sample to chloride ions. These chloride ions can then be reacted with a solution of silver nitrate to produce solid silver chloride that can be determined gravimetrically.

A 25.0 kg sample of soil is treated to convert all of the chlorine in the DDT to chloride ions. A solution containing the dissolved chloride ions is separated from the soil mixture and reacted with excess silver nitrate solution. The precipitate was collected and dried. The mass obtained was 42.9 mg.

 - a) Write the ionic equation for the reaction of silver nitrate with the solution containing chloride ions.
 - b) What amount, in mole, of solid was obtained.
 - c) What amount, in mole, of chloride ions must have been present in the solution obtained from the soil sample?
 - d) From the non-IUPAC name for DDT, deduce how many chlorine atoms are present in one molecule of DDT and hence determine the amount, in mole, of DDT in 1 kg of soil.

- e) If the molar mass of DDT is $354.49 \text{ g mol}^{-1}$, calculate the amount of DDT in the soil in ppm (parts per million, or mg per kg).
4. How does a primary standard differ from a secondary standard?
5. There are a number of steps involved in the preparation of 250.0 mL of a 0.050 mol L^{-1} solution of sodium carbonate. From the list below choose those steps that would be included in the preparation of such a solution and list them in the correct order.
- Dry the inside of the volumetric flask.
 - Weigh out 3.577 g of $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$.
 - Wash a 250.0 mL volumetric flask with de-ionised water.
 - Weigh out 1.325 g of anhydrous sodium carbonate.
 - Transfer the sodium carbonate to a 250 mL conical flask.
 - Rinse the volumetric flask with the sodium carbonate solution.
 - Swirl the contents of the volumetric flask until the solid dissolves.
 - Fill the volumetric flask with de-ionised water until the top of the meniscus is level with the calibration mark.
 - Weigh out 1.325 g of hydrated sodium carbonate.
 - Shake the volumetric flask with your thumb over the end to prevent the solution from escaping.
 - Transfer all of the solid sodium carbonate to the 250.0 mL volumetric flask.
 - Put a stopper on the flask and shake gently to distribute the solution.
 - Add some de-ionised water to the flask.
 - Fill the volumetric flask with de-ionised water until the bottom of the meniscus is level with the calibration mark.
 - Add several drops of phenolphthalein indicator to the solution in the flask.
6. For each of the following descriptions of the preparation of a standard solution, explain why the procedure is not accurate, and state, with reasons, whether the actual concentration of the solution would be higher or lower than that calculated using the data given.

Solution 1: A clean, dry 250 mL volumetric flask is placed on an analytical balance and weighed. To this flask is added 2.50 g of a solution of concentrated ammonia labelled as 'ammonia – $25 \text{ g per } 100 \text{ mL}$ '. The ammonia is diluted with de-ionised water and made up to the calibration mark on the flask before the solution is stoppered and gently shaken.

Solution 2: A watch glass is placed on an analytical balance and 5.70 g of potassium hydroxide is transferred to a 500.0 mL volumetric flask and de-ionised water is added to make the solution up to the calibration mark.

Solution 3: Into an empty beaker on an analytical balance is weighed 2.905 g of hydrated sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$. The solid is transferred to a 1000 mL volumetric flask and is then made up to the calibration mark with de-ionised water.

7. a) Why is it unnecessary to have the solution in the burette adjusted to the 0.00 mL 'mark' on the scale at the beginning of any titration?
b) The reading of the meniscus on a 50 mL burette scale after a titration is 36.44. The previous titration required 18.58 mL. One of the members of a practical groups suggests that the burette needs to be 'topped up' with more solution before proceeding to the next titration. The other member of the group suggests that they add the remaining solution to the 50.00 mL mark and then top up the burette to complete the experiment. Whose suggestion is better and why?
8. A standard solution of sodium carbonate is prepared by adding an accurately weighed amount of anhydrous sodium carbonate to a 250.0 mL volumetric flask and making the solution up to the calibration mark with distilled water. measured aliquots of this standard solution are then titrated with a hydrochloric acid solution of unknown concentration.

What will be the effect on the calculated concentration of hydrochloric acid (compared to its actual concentration) of each of the following? (Assume the only factor being altered is the one specifically identified and all other parts of the experiment are performed correctly.)

- a) The solution in the volumetric flask has water added to a point just above the calibration mark.
b) The pipette is rinsed with water prior to measuring out an aliquot.
c) The chemist performing the titration is colour blind and consistently adds more acid from the burette than needed.
d) The insides of the conical flask are washed down with de-ionised water during titration.
e) The burette was rinsed with de-ionised water prior to being filled with the acid used for the titration.
9. An unknown acid has the general formula H_xA , where all of the hydrogen atoms in the formula are acidic. A 0.432 g sample of this acid was dissolved in water and made up to 250.0 mL in a volumetric flask. A 25.00 mL aliquot of the acid solution was placed in a conical flask together with a few drops of phenolphthalein, and the whole titrated with a solution of $0.0400 \text{ mol L}^{-1}$ potassium hydroxide. The end point of the titration occurred when 20.00 mL of base had been added.

If the molar mass of the acid is 108 g mol^{-1} , what is the value of x ?

10. The percentage of ammonia in a commercial cleaner was determined by volumetric analysis.

A 18.73 g sample of the cleaner was diluted to 250.0 mL in a volumetric flask and 25.00 mL aliquots of the diluted solution were titrated with a previously standardised solution of hydrochloric acid using methyl orange as the indicator. The concentration of the acid solution was 0.115 mol L^{-1} .

The following titres were obtained.

Volume of HCl added				
Final burette reading (mL)	46.75	37.35	36.85	43.50
Initial burette reading (mL)	27.80	18.30	16.95	24.35

- What would you use to rinse the following equipment just prior to the titration?
 - pipette
 - volumetric flask
 - burette
 - conical flask
 - Determine the percentage by mass of ammonia in the cleaner showing details of all steps in the calculation.
11. A titration of a solution of potassium hydroxide of unknown concentration is to be performed. The solution is going to be titrated with a standardised solution of hydrochloric acid, added from a burette. The following indicators have been provided.

Indicator	Colour		pH range of indicator
	Acid	Base	
Methyl violet	Yellow	Violet	0.0 – 1.6
Methyl orange	Red	Yellow	3.2 – 4.4
Litmus	Red	Blue	5.0 – 8.0
Bromothymol blue	Yellow	Blue	6.0 – 7.6
Phenol red	Yellow	Red	6.6 – 8.0
Phenolphthalein	Colourless	Pink	8.2 – 10.0
Alizarin yellow	Yellow	Red	11.1 – 12.0

- Which indicators would you not consider for this titration? Give reasons for your choice.
- For each indicator chosen in part a) explain what the effect of that choice would be on the calculated concentration of the potassium hydroxide.
- The potassium hydroxide solution, once standardised, is to be used to determine the concentration of a solution of citric acid, a weak acid found in citrus fruits. Which of the indicators would be suitable for this titration? Once again, justify your answer.

12. Compare the titration of a strong monoprotic acid with that of a weak monoprotic acid, using sodium hydroxide as the base, with respect to:
- pH at the beginning of the titration
 - pH after a slight excess of base has been added
 - amount of base required to reach the equivalence point of equal molar amounts of the two acids are compared
 - pH at the equivalence point
 - indicator that should be used for the titration
13. The graph in Figure 8.16 shows the titration curve for the reaction of two different bases with a standard solution of hydrochloric acid.

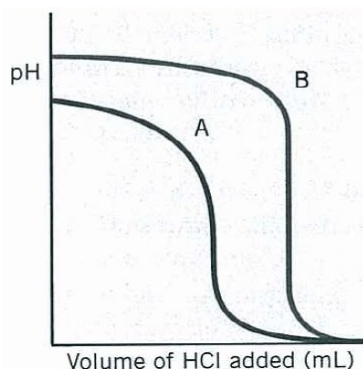


Figure 8.16 Titration curves for reaction of two different bases with HCl

- Which of these curves is with the strong base and which one is with the weak base?
 - What is the approximate pH at the equivalence point for each titration?
 - Why do the pH values of both solutions eventually reach approximately the same value when they start off at such different pH values?
 - If 20.0 mL of each of the bases A and B was present at the beginning of the titration, provide two possible reasons for the difference between the volume of acid required to reach the equivalence point in each titration.
14. A mixture of sodium chloride and anhydrous potassium carbonate with a mass of 0.260 g was made up to 25.00 mL with de-ionised water. This solution required 13.2 mL of 0.0995 mol L⁻¹ hydrochloric acid to reach the equivalence point.
- What volume of 0.0500 mol L⁻¹ AgNO₃ solution would be required to precipitate all of the chloride ions introduced with the hydrochloric acid) from the resulting solution?
15. A sample of pure iron weighing 0.835 g is dissolved in sulfuric acid, causing all of the iron to be oxidised to iron(II) ions. The resulting solution is titrated with a solution of potassium dichromate containing 12.5 g L⁻¹ of K₂Cr₂O₇. What volume of potassium dichromate solution is required?
16. A sample of lawn food is known to contain Fe²⁺ as iron(II) sulfate. A 2.000 g sample of lawn food is dissolved in 0.1 mol L⁻¹ sulfuric acid in a conical flask. The contents of the flask are quickly titrated with a solution of 0.0200 mol L⁻¹ potassium permanganate. The first permanent pink colour of the solution appears after 13.20 mL of permanganate has been added. Calculate the percentage by mass of iron in the lawn food sample.

17. Vitamin C tablets contain ascorbic acid, $\text{C}_6\text{H}_4\text{O}_2(\text{OH})_4$. A tablet with a mass of 320 mg is crushed and dissolved in 30 mL of water in a conical flask. The contents of the flask is titrated with a $0.0458 \text{ mol L}^{-1}$ solution of iodine.

Ascorbic acid reacts with iodine according to the equation:



Starch indicator is added to the titration mixture as it turns a deep blue colour when iodine is in excess. The starch indicator changes colour after 24.6 mL of iodine solution has been added.

Determine the mass of ascorbic acid in the tablet and hence the percentage by mass of ascorbic in the vitamin C tablet.

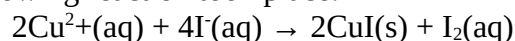
18. The major acid present in a sample of wine is tartartic acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$, which is a weak diprotic acid. A 25.00 mL aliquot of wine was added to 20.0 mL of 0.180 mol L^{-1} potassium hydroxide solution.

The resulting solution was basic due to excess hydroxide and required 19.85 mL of 0.100 mol L^{-1} nitric acid to reach a methyl orange end point.

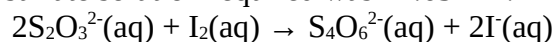
Calculate the mass of tartaric acid, in g L^{-1} , in the wine sample. What assumption has been made in this calculation?

19. It is possible to determine the percentage by mass of copper in a copper ore by converting the copper ions into a soluble form, and using them to oxidise iodide ions to iodine. The iodine produced in this reaction can then be titrated with a standard solution containing thiosulfate ions.

A concentrated sample of a copper ore having a mass of 0.450 g was treated with nitric acid to dissolve all of the copper ions as $\text{Cu}^{2+}(\text{aq})$. To the conical flask containing the copper ions was added 3.5 g of potassium iodide, KI. The following reaction took place:



The iodine produced was then titrated with 0.110 mol L^{-1} sodium thiosulfate solution using starch indicator to establish the end point. The volume of thiosulfate solution required was 21.65 mL:



- Determine the percentage of copper, by mass, in the ore sample.
- What assumptions have been made in this calculation?

20. 0.132 g of an ammonium salt was dissolved in water and enough potassium hydroxide was added to convert all of the ammonium ions to ammonia:

$$\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$$
The mixture was warmed and the ammonia gas, together with some water, was distilled into a flask containing 40.00 mL of 0.100 mol L⁻¹ hydrochloric acid:

$$\text{NH}_3(\text{g}) + \text{H}^+(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq})$$
The excess acid was titrated with 0.0500 mol L⁻¹ sodium hydroxide and required 40.00 mL for neutralisation.
- What amount, in mole, of acid was originally present in the flask?
 - How much acid was left over in the flask after the ammonia had reacted?
 - What amount, in mole, of ammonia was generated in the first reaction?
 - What was the percentage, by mass, of ammonium ions in the original salt?
 - If the negative ion in the ammonium salt has a charge of -2, what amount, in mole, of the salt was present in the 0.132 g and what is the formula mass of the negative ion? What ion might this be?
21. 1.010 g of a hydroxide of iron, Fe(OH)_x, was mixed with 20.00 mL of 2.000 mol L⁻¹ nitric acid and the mixture made up to 200.0 mL in a volumetric flask with distilled water. A 25.00 mL aliquot of this solution was placed in a conical flask and required 14.56 mL of 0.1000 mol L⁻¹ sodium hydroxide to reach the equivalence point. What is the value of x?
22. Lindane, C₆H₆Cl₆, is an insecticide that has been used in the treatment of head lice. On combustion, lindane produced, among other compounds, hydrogen chloride gas. An impure sample of lindane having a mass of 1.200 g was burnt in oxygen. The hydrogen chloride produced was bubbled through 50.00 mL of 0.946 mol L⁻¹ sulfuric acid and a titre of 26.55 mL was required. Calculate the percentage, by mass, of lindane in the impure sample.
23. Sulfur dioxide is used in wine as a preservative and antimicrobial agent. It is suggested that wine should contain no more than 125 ppm of free sulfur dioxide. An investigation into the sulfur dioxide content of a sample of imported white wine involves the following titration.
- 50.00 mL of wine was placed in a conical flask together with 5 mL of 25% sulfuric acid solution and 2-3 mL of starch indicator solution. The mixture was titrated with a standard solution of 0.0110 mol L⁻¹ iodine solution. The volume of iodine required to just turn starch indicator blue was 9.85 mL. Determine whether this sample of wine exceeded the 125 ppm limit.
- $$\text{SO}_2(\text{aq}) + \text{I}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{I}^-(\text{aq})$$
24. 24.92 g of a certain lead ore consisting of lead carbonate and lead sulfate was heated gently with 50.0 mL of 2.44 mol L⁻¹ nitric acid. The mixture was then filtered and the filtrate diluted to a volume of 500.0 mL. 25.00 mL aliquots of the diluted filtrate required 22.7 mL of 0.151 mol L⁻¹ sodium hydroxide to neutralise the unreacted nitric acid. Calculate the percentage, by mass, of lead carbonate in the ore.

25. 1.50 g of an acid was dissolved in water and its solution was made up to 250 mL. If 40.0 mL of this acid solution neutralised 25.0 mL of 0.120 mol L⁻¹ KOH solution, and if the acid is diprotic (that is, each acid molecule can donate two hydrogen ions to a base), find its relative formula mass.

YEAR 12 CHEMISTRY

ACIDS AND BASES ASSIGNMENT 2009

TIME ALLOWED: 45 minutes

MULTIPLE CHOICE ANSWER SHEET

1	A	B	C	D
2	A	B	C	D
3	A	B	C	D
4	A	B	C	D
5	A	B	C	D
6	A	B	C	D
7	A	B	C	D
8	A	B	C	D
9	A	B	C	D
10	A	B	C	D

	MARK	OUT OF	STRENGTH	WEAKNESS
PART 1		10		
PART 2		22		
PART 3		13		
TOTAL		45		

PART ONE: MULTIPLE CHOICE (10 questions 10 marks)

Please answer these questions on the separate multiple choice answer sheet provided.

1. Which of the following statements about acetic (ethanoic) acid are true?
- I the pH of 0.1 mol L⁻¹ solution of ethanoic acid is 1
 - II the ethanoate ion is the conjugate base of ethanoic acid
 - III a solution of ethanoic acid consists mainly of molecules
 - IV ethanoic acid is a strong acid
- a) I and II only
 - b) II and III only
 - c) I, II and IV only
 - d) all of I, II, III and IV
2. Which one of the following represents the conjugate acid of HAsO_4^{2-} ?
- a) AsO_4^{3-}
 - b) H_3AsO_4
 - c) H_2AsO_4^-
 - d) H_3O^+
3. Which one of the following acids would have the weakest conjugate base?
- a) HNO_3
 - b) H_2SO_3
 - c) HClO
 - d) HCN
4. What is the best description of a neutral solution?
- a) A neutral solution is one where the concentration of hydrogen ions is negligible
 - b) A neutral solution is one with a pH of exactly 7.0
 - c) A neutral solution is one resulting from the reaction between an acid and a base
 - d) A neutral solution is one where the concentrations of hydrogen ions and hydroxide ions are equal
5. Which of the following oxides would react with a solution of a strong base?
- | | | |
|-----|-------------------------|--|
| I | MgO | This question beyond new syllabus |
| II | Al_2O_3 | |
| III | SO_2 | |
| IV | ZnO | |
| V | CO_2 | |
- a) II and IV only
 - b) I, II and IV only
 - c) III and V only
 - d) II, III, IV and V only

6. Which one of the following statements about sulfuric acid is NOT true?
- V_2O_5 is used as a catalyst in the preparation of H_2SO_4
 - The pH of a 0.01 mol L^{-1} solution of H_2SO_4 is 1.7
 - It is used as the electrolyte in car batteries
 - It reacts with barium to form hydrogen gas and a white precipitate
7. Which one of the following lists contains solutions all with a pH greater than 7?
- NH_3 , $Ca(OH)_2$, Na_2CO_3 , KH_2PO_4
 - CH_3COOH , HNO_3 , H_2S , NH_4Br
 - $MgCl_2$, NH_3 , KOH , $Al(NO_3)_3$
 - $NaHCO_3$, $Ba(HSO_4)_2$, $LiOH$, Na_3PO_4
8. The volume of water required to dilute 100 mL of 0.400 mol L^{-1} nitric acid to $0.0800 \text{ mol L}^{-1}$ is
- 500 mL
 - 100 mL
 - 400 mL
 - 200 mL
9. What are the possible products of the following Bronsted-Lowry reaction?
- $$H_2O + NH_2^-(aq) \rightarrow$$
- $OH^- + NH_3$
 - $H_3O^+ + NH_3$
 - $OH^- + NH_4^+$
 - $OH^- + NH_2^-$

Note that Q10 below is NOT IN NEW CHEMISTRY COURSE OF STUDY

10. Which one of the following reactions would result in a white precipitate forming which then dissolves on the addition of further sodium hydroxide solution?
- sodium hydroxide solution added to silver nitrate solution
 - sodium hydroxide solution added to chromium hydroxide
 - sodium hydroxide solution added to zinc metal
 - sodium hydroxide solution added to aluminium chloride solution

PART TWO: SHORT ANSWER

3 questions 22 marks

1. Write balanced IONIC equations for any reactions that occur in the following procedures.

In each case describe in full what you would observe, including any

- Colours
- Odours
- Precipitates (give the colour)
- Gases evolved (give the colour or describe as colourless)

If a reaction occurs but no change is visible you should state this.

- a) Zinc oxide is stirred with potassium hydroxide solution

Equation THIS REACTION IS NOT IN NEW COURSE

Observation

- b) Solutions of barium hydrogen carbonate and sulfuric acid are mixed.

Equation

Observation

- c) Dilute NaOH solution is added to a solution of aluminium nitrate

Equation

Observation

- d) Ammonia gas is bubbled through hydrochloric acid

Equation

Observation

(12 marks)

2. You are given samples of white solids known to be ammonium nitrate, NH_4NO_3 , and potassium nitrate, KNO_3 . Describe, with the aid of equations, how you could identify the solids using only water and Universal Indicator.

(4 marks)

3. a) Explain with the aid of an equation, why pure water conducts electricity, even though to only a very small extent.

(2 marks)

- b) Explain, with the aid of an equation, why the conductivity of 1 L of pure water increases when 1 mole of hydrogen chloride gas is bubbled through it. (Assume no volume change)

(2 marks)

- c) Calculate the concentration of hydroxide ions in the solution in b).

(2 marks)

PART THREE**CALCULATIONS****2 questions 13 marks*****Express all answers to three significant figures***

1. 111.0 mL of hydrochloric acid was required to dissolve (react with) 3.16 g of iron(III) hydroxide.

a) Write the equation of the reaction

(1 mark)

b) Calculate the concentration of the hydrochloric acid.

(4 marks)

- c) If 150.0 mL and not 111.0 mL of acid was added to the iron(III) hydroxide what would be the pH of the final solution?

(4 marks)

2. 0.0200 mol of Na_2SO_4 was added to 125 mL of $0.0100 \text{ mol L}^{-1} \text{Na}_3\text{PO}_4$ solution.
- Calculate the final sodium ion concentration.

2009 acid-base assign answers

Multiple choice

1. b 2. c 3. a 4. d 5. d (not in new syllabus) 6. b
7. a 8. c 9. 10. (not in new syllabus)



Year 12 Chemistry
Test : Acids and Bases 2009

Part One: Multiple Choice

(15 questions, 15 marks)

Please answer these questions on the separate multiple choice answer sheet provided.

1. Which one of the following substances is the most suitable as an indicator for strong acid-weak base titrations?

- (a) Universal indicator
- (b) Litmus
- (c) Methyl orange
- (d) Phenolphthalein

2. Bromocresol green produces the following colours when added to solutions of known pH.

pH	Colour
1	Yellow
2	Yellow
3	Blue
4	Blue

Bromocresol green is added to 0.1 mol L^{-1} solutions of acetic acid, carbonic acid and nitric acid. Which one of the following correctly identifies the colours the indicator will produce in each solution?

	Acetic acid	Carbonic acid	Nitric acid
(a)	Yellow	Yellow	Yellow
(b)	Yellow	Yellow	Blue
(c)	Blue	Blue	Blue
(d)	Blue	Blue	Yellow

3. 20.0 mL of a 0.030 mol L^{-1} solution of KOH is added to 20.0 mL of 0.010 mol L^{-1} solution of HCl . What is the pH of the resulting solution?

- (a) 1
- (b) 2.0
- (c) 12
- (d) 13

4. The water in a seriously neglected dam is tested and found to have a pH of 9. By what factor must the hydrogen ion concentration be changed to decrease the water pH to 7?

- (a) decrease the hydrogen ion concentration by 100 times
- (b) double the hydrogen ion concentration
- (c) halve the hydrogen ion concentration
- (d) increase the hydrogen ion concentration by 100 times

5. Which one of the following cannot act as both a Bronsted-Lowry acid and base?

- (a) HPO_4^{2-}
- (b) HCO_3^-
- (c) HSO_4^-
- (d) SO_4^{2-}

6. What is the pH of a 0.05 mol L^{-1} solution of barium hydroxide?

- (a) 1.3
- (b) 7.0
- (c) 12.7
- (d) 13.0

7. Oxalic acid is a weaker acid than sulfuric acid. Which of the following best explains this?

- (a) Oxalic acid is found in plant material and sulfuric acid isn't.
- (b) Sulfuric acid has two hydrogen ions available for ionisation per mole and oxalic acid has one.
- (c) Oxalic acid is less soluble in water than sulfuric acid.
- (d) Oxalic acid ionises to a lesser extent than sulfuric acid.

8. Which of the following correctly identifies the acidity of the listed salts when dissolved in water?

	Sodium chloride	Sodium acetate	Ammonium acetate	Potassium hydrogencarbonate
(a)	Neutral	Acidic	Acidic	Basic
(b)	Acidic	Acidic	Acidic	Acidic
(c)	Neutral	Neutral	Neutral	Basic
(d)	Neutral	Basic	Neutral	Basic

9. In which of the following reactions is water behaving as a base?
- $\text{H}_2\text{O} + \text{C} \rightarrow \text{H}_2 + \text{CO}$
 - $\text{H}_2\text{O} + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{OH}^-$
 - $\text{H}_2\text{O} + \text{HCl} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$
 - $\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3 + \text{OH}^-$
10. What volume of sodium hydroxide of pH 12 would be needed to neutralise 10 mL of hydrochloric acid of pH 1?
- 100 mL
 - 10 mL
 - 1 mL
 - 10×10^{12} mL
11. Which of the following is not a property of sulfuric acid?
- The dilute acid will react with copper to produce hydrogen gas.
 - The concentrated acid is an oxidising agent.
 - The dilute acid will react with magnesium.
 - The dilute acid absorbs water and may be used as a drying agent.
12. Which of the following oxides dissolve in water to give an acidic solution?
- Na_2O
 - Al_2O_3
 - P_4O_{10}
 - SO_2
- I and III only
 - II and III only
 - II, III and IV only
 - III and IV only
13. An acetic acid (ethanoic acid) solution is titrated with sodium hydroxide solution (in the burette), using methyl orange as the indicator. Which one of the following statements about this titration is **true**?

- (a) The end point and equivalence point occur at the same time.
 - (b) The end point occurs after the equivalence point.
 - (c) The end point occurs before the equivalence point.
 - (d) Whether the end point or equivalence point occurs first depends on the relative concentrations of acid and base.
14. The value of K_w at 35°C is 2.09×10^{-14} . This means that
- (a) the pH of water at 35°C is 13.7 and it is basic.
 - (b) the pH of water at 35°C is 6.3 and it is acidic.
 - (c) the pH of water at 35°C is 6.3 and it is neutral.
 - (d) the pH of water at 35°C is 13.7 and it is neutral.
15. Which of the following statements about aqueous solutions of weak acids is true?
- (a) A weak acid is a concentrated acid that has been diluted.
 - (b) A 1.00 mol L⁻¹ solution of a weak acid contains more molecules of acid than ions.
 - (c) Less than 1.0 mol of sodium hydroxide is needed to react completely with 1.0 mol of a monoprotic weak acid.
 - (d) The salt produced through the neutralisation of a weak acid by a strong base is slightly acidic.



Year 12 Chemistry
Test : Acids and Bases 2009

Name _____

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
a	a	a	a	a	a	a	a	a	a	a	a	a	a	a
b	b	b	b	b	b	b	b	b	b	b	b	b	b	b
c	c	c	c	c	c	c	c	c	c	c	c	c	c	c
d	d	d	d	d	d	d	d	d	d	d	d	d	d	d

Please answer these questions in the spaces provided.

1. Calculate the pH of a solution containing 25.0 mL of 2.0 mol L⁻¹ sodium hydroxide and 49.5 mL of 1.0 mol L⁻¹ hydrochloric acid.

(4 marks)

2. Write equations for any reactions that occur in the following procedures. If no reaction occurs write 'no reaction'.
In each case describe what you would observe, including any: colours, gases evolved, odours and precipitates.
If no change is observed you must state this as the observation.

(a) A solution of phosphoric acid is added to a solution of sodium hydroxide.

Equation

Observation

(b) Potassium hydroxide solution is added to solid zinc hydroxide.

Equation **NOT IN NEW
COURSE**

Observation

(c) Dilute hydrochloric acid solution is added to chromium (III) sulfide solid.

Equation NOT IN NEW COURSE

Observation

(d) Ammonium chloride and sodium hydroxide solutions are mixed and gently warmed.

Equation

Observation

(12 marks)

3. A student was required to standardise a solution of sulfuric acid by titrating it with a sodium carbonate solution made up from an accurately weighed mass of pure sodium carbonate. The sulfuric acid was placed in the burette. The sodium carbonate solution was pipetted into the conical flask with a few drops of methyl orange indicator added.

(a) Explain why sodium carbonate was chosen for this experiment.

(2 marks)

(b) Explain why methyl orange was chosen as the indicator.

(2 marks)

4. Answer the following questions about terms used in volumetric analysis.

a) What does the word 'standardise' mean?

b) Why is a primary standard often needed for volumetric analysis?

c) What is meant by the term 'end point'?

d) What is meant by the term 'equivalence point'?

(4 marks)

5. 3.12 g of washing soda, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, was dissolved in water and the solution made up to a volume of 200.0 mL in a volumetric flask. 18.18 mL of $0.0600 \text{ mol L}^{-1} \text{ H}_2\text{SO}_4$ neutralised 20.0 mL of the solution. Calculate the percentage, by mass, of Na_2CO_3 in the washing soda.

(6 marks)

6. Using your knowledge of Bronsted-Lowry theory, the properties of water and the pH scale, discuss the pH values in the table below. All solutions are 0.1 mol L^{-1} .

Solution 0.1 mol L^{-1}	pH
HNO_3	1.0
HClO	4.2
NH_4NO_3	5.1
KNO_3	7.0

(8 marks)



Year 12 Chemistry
Test : Acids and Bases 2009

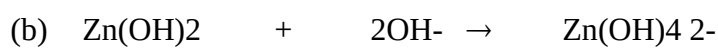
Answers

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
C	D	C	D	D	D	D	D	C	A	A	D	C	C	C

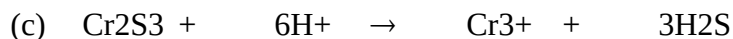
1. $n(\text{OH}) = 0.025 \times 2.0 = 0.050 \text{ mol}$
 $n(\text{H}) = 0.0495 \times 1.0 = 0.0495 \text{ mol}$
 $n(\text{XS OH}) = 0.005 \text{ mol}$
 $[\text{OH}] = 0.005 / 0.0745 = 6.71 \times 10^{-2} \text{ mol L}^{-1}$
 $\text{pOH} = -\log [\text{OH}] = 1.17$
 $\text{pH} = 14 - 1.17 = 12.83$



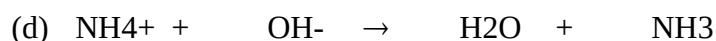
no visible change



white solid dissolves in a colourless solution to form a colourless solution.



green solid dissolves in a colourless solution to form a green solution and a colourless gas with an unpleasant smell.



two colourless solutions form a colourless solution and a colourless gas with a pungent smell.

- 3(a) sodium carbonate solution is basic. It is neither deliquescent nor efflorescent and does not absorb gases from the air. The solid has a high molar mass and may be obtained in pure form.

- (b) methyl orange has an end point colour change in the same pH range as the equivalence point of the reaction between sodium carbonate and sulfuric acid.

- 4.(a) Standardise means, determine the exact concentration of.
- (b) Many substances may only be obtained in an impure form, hence the concentration of their solutions may only be approximately known.
- (c) The end point of a titration occurs when the indicator changes colour.
- (d) The equivalence point of a titration occurs when there are chemically equivalent amounts of both reactants present in the conical flask.



$$n = cV = 0.001091 \text{ mol}$$

$n = 0.001091$ in 20 mL therefore 0.001091×10 in 200 mL.

$$m = nM = 1.156 \text{ g}$$

$$\% = 1.156 / 3.12 = 37.1\%$$

6. 2 marks per compound, if pH explained fully.