**ACIDS AND BASES**

**OBJECTIVES OF THE TOPIC**

At the end of the topic the learner should be able to:

* Define acids and bases according to Arrhenius
* Define acids and bases according to Lowry-Bronsted
* Distinguish between strong acids/bases and weak acids/bases.
* Distinguish between concentrated acids/bases and dilute acids/bases.
* Write down the reaction equations of aqueous solutions of acids and bases.
* Identify conjugate acid –base pairs for given compounds.
* Define an Amphiprotic substance.
* Write equations to show how an amphiprotic substance can act as acid or base.
* Write down neutralisation reactions of common laboratory acids and bases.
* Define hydrolysis.
* Determine the approximate pH (equal to, smaller than or larger than 7) of salts in salt hydrolysis.
* Define equivalence point.
* Define end point of a titration.
* Motivate the choice of a particular indicator in a titration.
* Perform stoichiometric calculations based on titrations.
* For a titration
	+ - List apparatus needed or identify apparatus from diagram.
		- Describe procedure to prepare a standard solution
		- Describe procedure to conduct titration.
		- Describe safety precautions.
		- Describe measures that need to be in place to ensure reliable results.
		- Interpret given results to determine the unknown concentration.
* Define Kw
* Explain the auto-ionisation of water.
* Explain the pH scale
* Calculate pH values of strong acids and strong bases using

pH = -log[H3O+].

* Interpret Ka values of acids to determine the relative strength of given acids.
* Interpret Kb values of bases to determine the relative strength of given bases.
* Compare strong and weak acids by looking at:
	+ - pH (monoprotic and diprotic acids)
		- Conductivity
		- Reaction rate

**KEY CONCEPTS:**

* Acid- base theories
* Reactions of acids and bases with water
* Conjugate acid-base pairs
* Amphiprotic substances
* Neutralisation reactions
* Hydrolysis
* Acid-base titrations
* Auto-ionisation of water
* pH Calculations
* Comparison of strong and weak acids

**TEACHING APPROACH**

* **Acid- base theories**

1.1 ***Arrhenius theory of acid and based***

 Acids are substances which produce hydrogen ions (H+) in the solution.

Bases are substances which produce hydrogen ions (OH-) in the solution.

1.2 ***Lowry-Bronsted theory of acid and bases***

Acids are proton (H+) donors

Bases are proton (H+) acceptors

* **Reactions of acid and bases with water**

2.1 ***Monoprotic and diprotic acids***

Monoprotic acid can donate only one proton in a reaction.

HCl(aq) + H2O(l) → H3O+(aq) + Cl- (aq)

1 mole of acid react with 1 mole of water to form 1 mole of H3O+ ions

Diprotic acid can donate two protons in a reaction.

H2SO4(aq) + 2H2O(l) → 2H3O+(aq) + SO42-(aq)

1 mole of acid react with 1 mole of water to form 2 moles of H3O+ ions

2.2 ***Strong and weak acids***

Strong acid

A strong acid ionises completely in water forming a high concentration of H3O+ ions.

The ionisation constants (Ka) of strong acid are very large

Weak acid

A weak acid ionise incompletely in water forming a low concentration of H3O+ ions.

All acids that ionise incompletely or only partially with Ka smaller than one (Ka<1) are referred to as weak acid.

2.3 ***Strong and weak base***

Strong base

A strong base dissociates completely in water producing a high concentration of OH- ions.

All bases that dissociates/ionise completely with Kb greater than one (Kb>1) are referred to as strong bases.

Weak base

A weak base dissociates incompletely in water producing a low concentration of OH- ions.

All bases that dissociates/ionise incompletely or partially with Kb less than one (Kb<1) are referred to as weak bases.

2.4 ***Concentrated and dilute acids / bases***

Concentrated acid/base: Contains a large amount of acid/base in proportion to the volume of water.

Dilute acid/base: Contains a small amount of acid/base in proportion to the volume of water.

Concentration of acid/base can be calculated using the following formulae:

C=n/v or C=m/MV

* **Conjugate acid- base pairs**

CH3COOH(aq) + H2O(l)  H3O+ + CH3COO-(aq)

acid 1 base 2 acid 2 base1

 Conjugate pair

 Conjugate pair

The conjugate base of a weaker acid is a stronger base and the conjugate acid of a weaker base is stronger acid

* **Amphiprotic substances**

Some substances can act as either and acid or a base depending on the other reactant. Such substances are called amphiprotic or ampholytes.

**Task1: Homework/Classwork**

***Question 1: Multiple choice questions***

1.1 Consider the equilibrium H2SO4 + HSO3-  HSO4- + H2SO3

 The Lowry- Bronsted bases are

 A HSO4- and H2SO4 B HSO3- and HSO3

 C HSO3- and HSO4- D HSO4- and H2SO3

1.2 Water can act as either an acid or a base. Which equation represents water reacting as an acid?

A H2O(l) + NH3(g) OH- + NH4+(aq)

B H2O(l) + HCl(aq)  H3O+(aq) +Cl-(aq)

C H2O(l) H2(g) + ½O2(g)

D H2O(l) +C(s )  CO(g) + H2(g)

1.3 Which one of the following best describes the difference between a base and its conjugate acid?

A The base has an additional OH- ion

B The base has an additional OH- ion

C The conjugate acid has an additional OH-ion

D The conjugate acid has an additional H+ ion.

***Contextual questions***

***Question 2***

2.1 For each of the following, write down the formula of the conjugate acid

2.1.1 NH3 2.1.2 NH2- 2.1.3 C10H14N2 2.1.4 H2O

2.2 For each of the following write down the formula of the conjugate base

2.2.1 HCl 2.2.2 H2CO3 2.2.3 H2O 2.2.4 HPO42-

* **Neutralisation reactions**

A neutralisation reaction takes place when an acid react with a base in aqueous solution to form a salt and water

5.1 Reaction of strong acid and a strong base

When a strong acid reacts with a strong base, the salt that forms will be neutral

5.2 Reaction of a weak acid with a strong base

When a weak acid react with a strong base, the salt solution will be basic (an alkaline solution)

5.3 Reaction of a strong acid with a weak base

When a strong acid reacts with a weak base, the salt that forms will be acidic.

* **Hydrolysis**

Hydrolysis is the reaction of a salt with water

* Split the salt into its ions and decide from which acid and which base the ions are coming.
* Also determine whether each of the acid and base is strong or weak.
* If an ion comes from either strong base or a strong acid. It will not undergo hydrolysis.
* The ion from the weak base or weak acid will undergo hydrolysis
* The pH of the salt solution is finally determined by the hydrolysis reaction that will take place. If OH- is formed during hydrolysis, the salt is basic. If H3O+ is formed during hydrolysis, the salt is acidic.

***Example 1:***

Will a solution of NH4Cl be basic, acidic or neutral? Use hydrolysis to fully explain the answer.

Step 1: The two ions present in this salt are NH4+and Cl-

NH4+ comes from a weak base, NH3. Cl- comes from a strong acid HCl

Step 2: Cl- will not undergo hydrolysis because it is the conjugate base of a strong acid

NH4+ is the conjugate acid of a weak base and will undergo hydrolysis

Step 3: NH4+ undergo hydrolysis according to the following equation

NH4+ + H2O  NH3  + H3O+

Step 4: The salt is acidic because H3O+ ions are formed during hydrolysis. The pH of the salt will be less than 7.

***Example 2***

Will a solution of NaNO3 be basic, acidic or neutral? Use hydrolysis to fully explain the answer.

Step 1: The two ions present in this salt are Na+ and NO3-

Na+ comes from a strong base, NaOH. NO3- comes from a strong acid HNO3

Step 2: NO3- will not undergo hydrolysis because it is the conjugate base of a strong acid.

Na+ will not undergo hydrolysis because it comes from a strong base

Step 3: None of the two ions will undergo hydrolysis

Step 4: No H3O+ and OH- ions are formed; therefore, the salt is neutral and the pH=7

***Example 3***

Will a solution of NaHCO3 be basic, acidic or neutral? Use hydrolysis to fully explain the answer.

Step 1: The two ions present in this salt are Na+ and HCO3-

Na+ comes from a strong base, NaOH. HCO3-comes from a weak acid H2CO3.

Step 2: Na+ will not undergo hydrolysis because it is the conjugate acid of a strong base.

HCO3- is the conjugate base of a weak acid and will undergo hydrolysis.

Step 3: HCO3-  undergo hydrolysis according to the following equation

HCO3- + H2O  H2CO3  + OH-

Step 4: The salt is basic because OH- ions are formed during hydrolysis. The pH of the salt will be greater than 7.

***SUMMARY***

* Hydrolysis of the salt of a weak acid and a strong base forms an alkaline solution, i.e. the pH>7. Examples of such salts are sodium ethanoate. Sodium oxalate and sodium carbonate
* Hydrolysis of the salt of a strong acid and a weak base forms acidic solution, i.e. the pH <7. an example of such salt is ammonium chloride
* The salt of a strong acid and a strong base does not undergo hydrolysis and the solution of the salt will be neutral, i.e. pH=7.

**Task 2: Homework/Classwork**

***Question 1 Multiple choice questions***

* 1. Consider the equation H+(aq) + OH-  H2O(l). Which type of reaction does this equation represent?

A Esterification B Decomposition

C Hydrolysis D Neutralisation

* 1. Consider the ionisation constants for four weak acids, I,II,III and IV

|  |  |
| --- | --- |
| Weak acid | Ka |
| I | 1x10-3 |
| II | 3x10-5 |
| III | 2.6x10-7 |
| IV | 4x10-9 |

The weakest conjugate base will be produced by acid…..

A I B II C III D IV

* 1. Which ONE of the following salts produces neutral solutions when dissolved in water?

A NaNO3 B CH3COONa C Na2CO3 D NH4Cl

1.4 When solid K2CO3 is added to water, the pH…….

A becomes less than 7 because of hydrolysis of K+

B becomes greater than 7 because of hydrolysis of K+

C becomes greater than 7 because of hydrolysis of CO32-

D becomes less than 7 because of hydrolysis of CO32-

***Contextual questions***

***QUESTION 2***

The salt sodium carbonate is dissolved in water.

2.1 Define the term hydrolysis

2.2 Write down the formula of the acid and the base that reacts to form salt.

2.3 Classify each of the acid and base in QUESTION 2.2 as either weak or strong.

2.4 From your answer to QUESTION 2.3, predict whether the pH of the salt solution will be EQUALL TO 7, GREATER THAN 7 OR LESS THAN7

2.5 Write down a balanced equation to explain the answer in QUESTION 2.4

***QUESTION 3***

The salt ammonium nitrate is dissolved in water.

3.1 Write down the formula of the acid and the base that reacts to form salt.

3.2 Classify each of the acid and base in QUESTION3.1 as either weak or strong.

3.3 From your answer to QUESTION 3.2, predict whether the pH of the salt solution will be EQUALL TO 7, GREATER THAN 7 OR LESS THAN7.

3.4 Write down a balanced equation to explain the answer in QUESTION 3.3

* **ACID - BASE titration**

An acid-base titration is a procedure for determining the amount of acid (or base) in a solution by determining the volume of the base/ acid of known concentration that will completely react with it.

The solution of known concentration is called a ***standard solution***.

***Acid-base indicators***

|  |  |  |  |
| --- | --- | --- | --- |
| **Indicator** | **pH range** | **Colour in acid** | **Colour in base** |
| Bromothymol blue | 6.0 - 7.6 | yellow | blue |
| Phenolphthalein | 8.3 - 10.0 | colourless | pink |
| Methyl orange | 2.5 - 4.4 | red | yellow |

Precautionary measures when doing a titration

* Rinse the pipette first with distilled water and then twice with the acid (or base) to the measured
* Before filling the burette, it should be rinsed at least twice with the base (or acid) with which it will be filled.
* When measuring the volume of solution in the burette, make sure that your eye is in level with the meniscus to prevent a parallax mistake.
* Wash any titrant (solution in burette) spilled against the sides of the Erlenmeyer flask down with distilled water before the endpoint is reached. This water will not change the number of moles of acid (or base) present in the flask and thus will not influence the results.

Safety precautions when dong acid-base titrations

* Always add acid to water and not water to acid.
* Wear protective equipment, gloves, goggles, lab gown etc.

**7.1 Titration of a strong acid and a strong base**

The best choice of indicator will be bromothymol blue, pH 6.0-7.6

**7.2 Titration of a weak acid and a strong base**

The best choice of indicator will be phenolphthalein, pH 8.3-10.0

**7.3 Titration of a strong acid and a weak base**

The best choice of indicator will be methyl orange, pH 2.5-4.4

**Task3: Homework/Classwork**

***Question 1 Multiple choice question***

* 1. Consider the following indicator equilibrium:

HIn + H2O  H3O+ + In-

Colourless blue

What is the effect of adding HCl to a blue sample of this indicator?

|  |  |  |
| --- | --- | --- |
|  | Shift of equilibrium position | Colour change |
| A | To the right | More blue |
| B | To the left | Less blue |
| C | To the left | More blue |
| D | To the right | Less blue |

***Contextual questions***

***Question 2***

A learner used a standard solution of sodium hydrogen carbonate to determine the concentration of a sulphuric acid solution

* 1. Write down the definition of a standard solution.
	2. Why H2SO4 regarded as a strong acid?
	3. Write the balanced equation for the reaction of sulphuric acid with water.

In a titration, the learner finds that 20 cm3 of a 0.2moldm-3 solution of sodium hydrogen carbonate neutralises 12 cm3 of the sulphuric acid solution. The balanced equation for the reaction is

2NaHCO3 + H2SO4 → Na2SO4 + 2CO2 + 2H2O

2.4 How many moles of NaHCO3 are present in 20cm3 of the 0.2 mol.dm-3 NaHCO3 solution?

2.5 Determine the number of moles of H2SO4 that are neutralised by 20 cm3 of a 0.2moldm-3 NaHCO3 solution.

2.6 Calculate the concentration of H2SO4 solution?

2.7 From the table below, select the most suitable indicator for use in this titration.

|  |  |
| --- | --- |
| indicator | pH range |
| Bromothymol blue | 6.0-7.6 |
| Phenolphthalein | 8.3-10.0 |
| Methyl orange | 2.5-4.4 |

* **Auto-ionisation of water**

H2O(l) + H2O (l)  H3O+(aq) + OH-(aq)

Kw – the equilibrium constant for the ionisation of water.

* **pH calculations**

 pH= -log[H3O+]

 pOH= -log[OH-]

 Kw= 1.0x10-14=[H3O+][OH-]

 pH +pOH=14

* **Comparison of strong and weak acids**

|  |  |  |
| --- | --- | --- |
|  | **Strong acid**  | **Weak acid** |
| ionisation | Completely ionised | Partially ionised |
| Ka value | Ka>1 | Ka<1 |
| [H3O+] for acids of same concentration | Equal to the original acid concentration | Much less than the original acid concentration |
| pH of acids of same concentration | Lower pH | Higher pH |
| conductivity | Strong electrolyte | Weaker electrolyte |
| Reaction rate | Faster rate | Slower rate |

***Task4: Homework/Classwork***

***QUESTION 1***

Calculate the:

* 1. pH of a 0.25 mol.dm-3 KOH solution.
	2. [H3O+] of a KOH solution that has a pH of 13.48.
	3. pH of a 0.5 moldm-3 H2SO4 solution.

1.4 Hydroxide ion concentration in 0.01mol.dm-3 solution of ammonia with pH of 10.6.

***QUESTION 2***

A learner adds 0.06 mole of 0.5 moldm-3 NaOH to 1 dm3 of a 0.5moldm-3 HCl solution

2.1 Write down the balanced equation for the reaction takes place.

2.2 Calculate the initial number of moles of HCl present in the solution.

2.3 Determine the number of moles of NaOH needed to react with the acid.

2.4 Which one of the two substances is excess?

2.5 Calculate the pH of the final solution.