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| **UNIT 5**  **ACIDS, BASES AND SALTS**  **Student Version**    www.s-cool.co.uk    **Contents**   1. Acids, Bases, Salts and Neutralisation 2. Physical Properties of Acids, Bases and Salts 3. Strong and Weak Acids and Bases 4. Further Reactions of Acids, Bases and Salts 5. Laboratory Preparation of Salts by Neutralisation 6. Qualitative Analysis 1a: Acid-Base Reactions 7. Indicators and the pH scale 8. Qualitative Analysis 1b: Indicators 9. Quantitative Analysis: Acid-Base Titrations   Key words: Arrhenius, Bronsted-Lowry, acid, base, salt, neutralisation, hydrogen ion, hydroxide ion, proton donor, proton acceptor, alkali, amphoteric, hygroscopic, deliquescent, efflorescent, water of crystallisation, strong acid, weak Acid, strong base, weak Base, amphoteric, salt hydrolysis, acidic, alkaline, neutral, pH, indicator, litmus, phenolphthalein, methyl orange, universal indicator, titration, equivalence point, initial volume, final volume, titre volume, concordancy  **Units which must be completed before this unit can be attempted:**  **Unit 1 – Atoms and the Periodic Table**  **Unit 2 – Particles, Bonding and Structures**  **Unit 3 – Amount of Substance and Measurement**  **Unit 4 – Introduction to Physical Chemistry**  **Estimated Teaching Time: 12 hours** |

**UNIT 5 SUMMARY AND SYLLABUS REFERENCE**

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| **Lesson** | **Title and Syllabus Reference** |
| **1** | **Introduction to Acids, Bases and Salts**  ***CA7a definitions of acids and bases*** *(Arrhenius concept of acids and bases, effects of acids on trioxocarbonate (IV) salts);* ***CA7b chemical properties of acids and bases*** *(balanced chemical equations of all reactions);* ***ISA7.1 simple definition of acids, bases, salts*** *(definition of acids and bases in terms of proton transfer (Bronsted-Lowry concept));* ***ISA7.2 chemical properties of acids and bases*** *(properties of acids and bases);* ***ISA7.3******examples of chemical substances classified as acids, bases or salts, ISA13.2 neutralisation*** *(equations representing neutralization reactions)* |
| **2** | **Physical Properties of Acids, Bases and Salts**  ***CA7b physical and chemical properties of acids and bases*** *(conductivities, taste etc);* ***CA7c acids, bases and salts as electrolytes*** *(electrolytes and non-electrolytes);* ***CA8b practical application of solubility*** *(generalisations of solubility of salts)* ***CC4 deliquescent, efflorescent and hygroscopic substances*** *(the use of deliquescent and hygroscopic substances as drying agents should be emphasized);* ***ISA7.2 physical properties of acids, bases, salts*** *(properties and uses of acids, bases and salts)* |
| **3** | **Strong and Weak Acids and Bases**  ***CA7c acids, bases and salts as electrolytes*** *(strong and weak electrolytes, evidence from conductivity and enthalpy of neutralisation);* ***CA7e weak acids and weak bases*** *(behaviour of acids and bases in water as example of equilibrium system; qualitative comparison of the conductances of molar solutions of strong and weak acids and bases);* ***ISA7.2 chemical properties of acids and bases*** *(properties of acids and bases)* |
| **4** | **Further Chemical Properties of Acids, Bases and Salts**  ***CA7b chemical properties of acids and bases*** *(concept of amphoterism);* ***CA7f hydrolysis*** *(qualitative explanation of hydrolysis, behaviour of some salts (e.g. NH4Cl, AlCl3, Na2CO3, CH3COONa) in water as examples of equilibrium systems));* ***CC7aiii binary compounds of oxygen: acidic oxides, basic oxides, amphoteric oxides****;* ***ISA7.2 chemical properties of acids and bases*** *(properties of acids, bases and salts)* |
| **5** | **Preparation of Salts by Neutralisation**  ***CA13aiii General Skills and Principles – filtration****;* ***CC4a methods of preparation of simple salts****;* ***ISA7.4 methods of preparation of salts*** *(preparation of salts using the following methods: neutralization, acid + salt)* |
| **6** | **Qualitative Analysis – Acid-Base Reactions**  ***CA5bii gases – properties of gases*** *(chemical properties of NH3, CO2, HCl and SO2);* ***CA7a definitions of acids and bases*** *(effects of acids on trioxocarbonate (IV) salts);* ***CA13ci Qualitative Analysis - characteristic test of NH4+******with dilute NaOH; confirmatory test for NH4+*** *(no formal scheme of analysis is required);* ***CA13cii Qualitative Analysis - characteristic reaction of dilute HCl on solids or aqueous solutions and conc. H2SO4 on solid samples of the following: Cl-, SO32-, CO32-, confirmatory tests for SO32-, CO32-*** *(no formal scheme of analysis is required);* ***CA13civ Qualitative Analysis - characteristic test for the following gases: NH3, CO2, HCl and SO2*** *(no formal scheme of analysis is required);* ***ISA7.2 chemical properties of acids and bases*** *(test for carbon dioxide and ammonia gases),* ***ISA7.3******examples of chemical substances classified as acids, bases or salts*** *(simple chemical tests to classify chemical substances as acids, bases, or salts)* |
| **7** | **The pH scale and simple indicators**  ***CA7a definitions of acids and bases*** *(effects of acids and bases on indicators);* ***CA7d pH*** *(knowledge of pH scale, pH as a measure of acidity and alkalinity);* ***CA7g acid-base indicators*** *(indicators as weak organic acids or bases (organic dyes); colour of indicator at any pH dependent on relative amounts of acid and base forms, working pH ranges of methyl orange and phenolphthalein),* ***ISA7.5 acid-base indicators*** *(description of the colours developed by phenolphthalein, litmus and methyl orange in dilute acids and dilute bases)* |

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| **Lesson** | **Title and Syllabus Reference** |
| **8** | **Universal Indicator and pH Colorimetry**  ***CA7a definitions of acids and bases*** *(effects of acids and bases on indicators);* ***CA13av General Skills and Principles - determination of pH value of various solutions by colorimetry****;* ***ISA7.3******examples of chemical substances classified as acids, bases or salts*** *(simple chemical tests to classify chemical substances as acids, bases, or salts);* ***ISA7.6 determination of the pH of given solutions*** *(the nature and use of the universal indicator and pH meter, determination of soil pH is required)* |
| **9** | **Qualitative Analysis – Indicators**  ***CA13ci Qualitative Analysis - confirmatory test for NH4+ with dilute NaOH; confirmatory tests for the anions and cations CA13ci Qualitative Analysis - characteristic test for the following gases: NH3; CO2; HCl and SO2*** *(no formal scheme of analysis is required);* ***ISA7.2 chemical properties of acids and bases*** *(test for ammonia gas);* ***ISA7.3 examples of chemical substances classified as acids, bases or salts*** *(simple chemical tests to classify chemical substances as acids, bases, or salts)* |
| **10** | **Acid-Base Titrations I**  ***CA7h acid-base titrations*** *(correct use of relevant apparatus, knowledge of how acid-base indicators work in titrations, titration involving weak acids versus strong bases, strong acids versus weak bases and strong acids versus strong bases using the appropriate indicators and their applications in quantitative determination; e.g. concentrations);* ***CA13ai General Skills and Principles - measurement of volume****;* ***CA13bi Quantitative Analysis - acid-base titrations; the use of standard solutions of acids and alkalis and the indicators methyl orange and phenolphthalein to determine the concentrations of acid and alkaline solutions*** |
| **11** | **Acid-Base Titrations II**  ***CA4cii standard solutions*** *(preparation of some primary standard solutions using Na2CO3, (COOH)2.2H2O);* ***CA7h acid-base titrations*** *(correct use of relevant apparatus, knowledge of how acid-base indicators work in titrations, titration involving weak acids versus strong bases, strong acids versus weak bases and strong acids versus strong bases using the appropriate indicators and their applications in quantitative determination; e.g. concentrations, purity, water of crystallisation and composition);* ***CA13ai General Skills and Principles - measurement of mass and volume****;* ***CA13aii General Skills and Principles - preparation and dilution of standard solutions; CA13bi Quantitative Analysis - acid-base titrations; the use of standard solutions of acids and alkalis and the indicators methyl orange and phenolphthalein to determine the concentrations of acid and alkaline solutions; CA13bii Quantitative Analysis - acid-base titrations; the use of standard solutions of acids and alkalis and the indicators methyl orange and phenolphthalein to determine the molar masses of acids and bases and water of crystallization; CA13biv Quantitative Analysis - acid-base titrations; the use of standard solutions of acids and alkalis and the indicators methyl orange and phenolphthalein to determine the percentage purity of acids and bases*** |
| **12** | **Unit 5 Revision and Summary** |

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***Lesson 1 – What are acids, bases and salts?***

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg**Thinkabout Activity 1.1: What do you know about acids and alkalis?**   * How many common acids can you name? * Which foods and drinks contain acids? * What do acids feel and taste like? * How many common alkalis can you name? * Which foods and drinks contain alkalis? * What do alkalis feel and taste like? |

1. **Definition of Acids, Bases, Salts and Neutralisation**

**(i) Acids**

* The **Arrhenius** definition of an acid is a substance which releases hydrogen (H+) ions when dissolved in water; common examples are:
* hydrogen chloride (HCl) dissolves in water to produce hydrogen (H+) and chloride (Cl-) ions:

HCl(g) 🡪 H+(aq) + Cl-(aq); the resulting solution is known as **hydrochloric acid**

* hydrogen nitrate (HNO3) dissolves in water to produce hydrogen (H+) and nitrate (NO3-) ions:

HNO3(l) 🡪 H+(aq) + NO3-(aq); the resulting solution is known as **nitric acid**

* hydrogen sulphate (H2SO4) dissolves in water to produce hydrogen (H+) and nitrate (SO42-) ions:

H2SO4(l) 🡪 2H+(aq) + SO42-(aq); the resulting solution is known as **sulphuric acid**

* Acids therefore mostly exist in aqueous solution (aq)
* the **Bronsted-Lowry** definition of an acid is a proton (H+) donor; the acid reacts with other ions and molecules in solution, losing its H+ in the process
* HCl, HNO3 and H2SO4 are all acids because they lose their H+ ions (protons) when in solution
* the Bronsted-Lowry definition of an acid includes all substances which are classified as Arrhenius acids; it also includes some substances which are not classified as Arrhenius acids (usually because they do not dissolve in water)

**(ii) Bases and Alkalis**

* The **Arrhenius** definition of a base is a substance which releases hydroxide (OH-) ions when dissolved in water; common examples are:
* some **metal oxides** such as CaO which react with water to produce hydroxide (OH-) ions:

CaO(s) + H2O(l) 🡪 Ca2+(aq) + 2OH-(aq)

* some **metal hydroxides** such as NaOH which dissolve in water to produce hydroxide (OH-) ions: NaOH(s) 🡪 Na+(aq) + OH-(aq)
* some **metal carbonates** such as Na2CO3 which react with water to produce hydroxide (OH-) ions: Na2CO3(s) + H2O(l) 🡪 2Na+(aq) + HCO3-(aq) + OH-(aq)
* **ammonia** (NH3), which reacts with water to produce hydroxide (OH-) ions:

NH3(g) + H2O(l) 🡪 NH4+(aq) + OH-(aq)

* The **Bronsted-Lowry** definition of a base is a proton (H+) acceptor; the base accepts H+ ions (protons) from other species in the solution (which are therefore behaving as acids); the four most common types of base are:
* OH- (hydroxide) ions, which accept H+ ions to form water: H+ + OH- 🡪 H2O
* CO32- (carbonate) ions, which accept H+ ions to form carbon dioxide and water: CO32- + 2H+ 🡪 CO2 + H2O
* O2- (oxide) ions, which accept H+ ions to form water: H+ + O2- 🡪 H2O
* NH3, which accepts H+ ions to become NH4+ (ammonium): NH3 + H+ 🡪 NH4+
* The Bronsted-Lowry definition of a base includes all substances which are classified as Arrhenius bases; it also includes some substances which are not classified as Arrhenius bases, usually because they do not dissolve in water
* A Bronsted-Lowry base which dissolves in water to produce a solution containing OH- ions is called an **alkali**; an alkali is therefore a soluble base; an alkali is therefore the same as an Arrhenius base
* According to Bronsted-Lowry, any reaction which involves the transfer of protons (ie from an acid to a base) can be classified as an acid-base reaction

**(iii) Salts and Neutralisation**

* When an acid reacts with a base, it loses its H+ ion; the H+ ion is replaced either with the metal ion from the base or with an ammonium ion; the resulting compound is called a **salt;** a salt is the product formed when the H+ ion from an acid is replaced by a metal ion or an ammonium ion
* A reaction between an acid and a base to make a salt is called a **neutralisation** reaction
* There are four main types of base, so there are four main types of neutralisation reaction in which a salt can be made from an acid and a base:
* from metal hydroxides: acid + metal hydroxide 🡪 metal salt + water

Eg HCl + NaOH 🡪 **NaCl** + H2O (the salt produced is sodium chloride)

* from metal oxides: acid + metal oxide 🡪 metal salt + water

Eg 2HNO3 + CaO 🡪 **Ca(NO3)2** + H2O (the salt produced is calcium nitrate)

* from metal carbonates: acid + metal carbonate 🡪 metal salt + carbon dioxide + water

Eg H2SO4 + K2CO3 🡪 **K2SO4** + CO2 + H2O (the salt produced is potassium sulphate)

* from ammonia: acid + ammonia 🡪 ammonium salt

Eg HCl + NH3 🡪 **NH4Cl** (the salt produced is ammonium chloride)

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| **https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpgTest your knowledge 1.2: Recognising Acids, Bases and Salts**  Deduce the formulae of the following substances and indicate whether they are acids, bases or salts: | |
| 1. Sodium oxide 2. Calcium hydroxide 3. Ammonium nitrate 4. Potassium carbonate 5. Strontium sulphate 6. Ammonium sulphate 7. Hydrochloric acid 8. Rubidium hydroxide 9. Magnesium carbonate | 1. Calcium nitrate 2. Sulphuric acid 3. Ammonium chloride 4. Nitric acid 5. Potassium sulphate 6. Magnesium oxide 7. Caesium bromide 8. Barium sulphate 9. Strontium nitrate |

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| **https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpgTest your knowledge 1.3: Understanding Neutralisation Reactions**  Write balanced symbol equations for the following neutralisation reactions: | |
| 1. Nitric acid with potassium hydroxide 2. Sulphuric acid with sodium hydroxide 3. Hydrochloric acid with calcium hydroxide 4. Nitric acid with calcium oxide 5. Hydrochloric acid with barium oxide 6. Sulphuric acid with magnesium oxide | 1. Nitric acid with calcium carbonate 2. Hydrochloric acid with barium carbonate 3. Sulphuric acid with sodium carbonate 4. Nitric acid with ammonia 5. Sulphuric acid with ammonia 6. Hydrochloric acid with ammonia |

###### *Lesson 2 – What are the physical properties of acids, bases and salts?*

1. **Physical Properties of Acids, Bases and Salts**

**(i) Solubility**

* Most acids are soluble in water; hydrochloric acid, sulphuric acid and nitric acid are all highly soluble in water; a small number of organic acids do not dissolve in water, but all acids will dissolve in alkalis such as NaOH
* The solubility of bases in water varies:
* **metal hydroxides** are usually insoluble in water; they are therefore bases but not alkalis

the exceptions are the Group I hydroxides (LiOH, NaOH, KOH etc) and strontium and barium hydroxide (Sr(OH)2 and Ba(OH)2), which are highly soluble in water, and calcium hydroxide (Ca(OH)2), which is slightly soluble in water; these hydroxides are alkalis as well as bases

* **metal oxides** are usually insoluble in water; they are therefore bases but not alkalis

the exceptions are the Group I oxides (Li2O, Na2O, K2O etc), and strontium and barium oxide (SrO and BaO), which are highly soluble in water; and calcium oxide (CaO), which is slightly soluble in water; when these oxides dissolve in water, they react with water to form their respective hydroxides:

Eg CaO + H2O 🡪 Ca(OH)2

* **metal carbonates** are usually insoluble in water; they are therefore bases but not alkalis

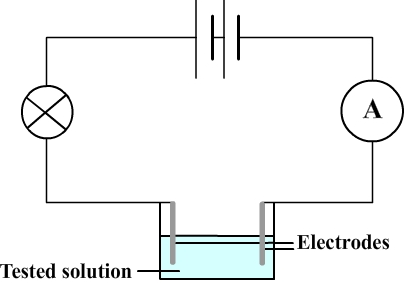
**t**he exceptions are the Group I carbonates (Li2CO3, Na2CO3, K2CO3 etc) which are highly soluble in water; these hydroxides are alkalis as well as bases

* ammonia is highly soluble in water; it is therefore an alkali as well as a base
* all bases will dissolve in acids such as HCl, even if insoluble in water
* Most common salts are soluble, with some exceptions:
* all nitrate salts are soluble
* most common chloride salts are soluble (except AgCl and PbCl2)
* most common sulphate salts are soluble, except barium sulphate and strontium sulphate (which are completely insoluble) and calcium sulphate (which is only slightly soluble) and a few other less common exceptions
* all Group I salts and all ammonium salts are soluble

The solubility of salts will be studied in more detail later in the course

**(ii) Electrical conductivity**

* Every ionic compound can be classified as an acid, base or salt, based on the cations and anions it contains:
* If the cation is H+, the ionic compound is an **acid**
* If the anion is O2-, OH-, CO32- or HCO3-, the ionic compound is a **base**
* Otherwise the ionic compound is a **salt**
* Most ionic compounds dissolve in water to produce free ions (note the exceptions above); as a result, aqueous solutions of ionic compounds (acids, bases or salts) behave as electrical conductors; molten ionic compounds also contain ions which can move; as a result, molten acids, bases and salts also behave as electrical conductors; any substance which can conduct electricity in the molten or aqueous states is called an **electrolyte;** molten and aqueous solutions of acids, bases and salts are **electrolytes**; any solution which shows electrical conductivity must therefore contain an acid, a base or a salt (or a mixture of them
* The electrical conductivity of a solution can easily be measured using a simple circuit; a cell, bulb and ammeter are all connected in series; the circuit is completed by two electrodes dipped into the solution but otherwise not in direct contact; if the bulb lights, the solution contains electrolytes; the greater the ammeter reading, the greater the concentration of electrolytes in the solution



**(iii) Taste**

* Acids have a sour taste, caused by the hydrogen ions; alkalis have a soapy taste, caused by the hydroxide ions; many salts have a salty taste, mainly caused by sodium ions but also caused by some other cations

**(iv) Deliquescence, Efflorescence, Hygroscopy and Water of Crystallisation**

* Many acids, bases and salts in the solid and aqueous states are hygroscopic, deliquescent, or efflorescent
* A **hygroscopic** substance is one which can absorb water from the atmosphere; this happens because the ions in the solid are strongly attracted to water; most concentrated solutions of acids and alkalis are hygroscopic and so are most solid salts if they are soluble in water; examples of hygroscopic substances are concentrated sulphuric acid, solid sodium chloride and solid sodium hydroxide; hygroscopic substances need to be stored in dry conditions to stop them from absorbing water
* Solids which are very hygroscopic are also deliquescent; a d**eliquescent** solid is one which can absorb so much water from the atmosphere that it will eventually dissolve in the water it has absorbed; sodium hydroxide and calcium chloride are both deliquescent solids
* Hygroscopic and deliquescent solids are often used as **drying agents** because of their ability to absorb water from the atmosphere
* Many bases and salts, and some acids, in the solid state contain water molecules locked in between the ions in the crystal structure; these are known as **hydrated** solids; they often have a fixed chemical formula; the water is shown after the unit formula, separated by a dot:

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| Examples of hydrated solids | Unit formula |
| Hydrated copper sulphate | CuSO4.5H2O |
| Hydrated sodium carbonate | Na2CO3.10H2O |
| Hydrated ethanedioic acid | H2C2O4.2H2O |

The water locked into the solid in these compounds is known as **water of crystallisation**; the water of crystallisation should be included in the molar mass of the solid:

* Eg the molar mass of hydrated sodium carbonate, Na2CO3.10H2O, is 106 + 10 x 18 = 286
* Some of these compounds lose this water on exposure to air over a period of time; this is known as efflorescence; an **efflorescent** solid is one which contains water within its crystal structure but which will gradually release the water on exposure to the atmosphere; efflorescence is the opposite of hygroscopy
* In many cases, there is an observable change in the solid as a result of effluorescence:
* hydrated copper sulphate is blue; when exposed to the atmosphere, it gradually loses its water and becomes anhydrous copper sulphate, which is white
* wet plaster contains hydrated calcium sulphate (CaSO4.2H2O); when plaster is left to dry, it loses its water and hardens

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| **Image result for test iconTest your knowledge 2.1: Describing properties of acids, bases and salts**   1. Name two substances which are soluble bases (alkalis) 2. Name two substances which are insoluble bases 3. Name two soluble salts 4. Name two insoluble salts 5. Explain why soluble acids, alkalis and soluble salts are all electrolytes 6. Describe the typical taste of acids, bases and salts and give the ion responsible for the taste 7. Explain what is meant by the term hygroscopic and name one hygroscopic acid, one hygroscopic base and one hygroscopic salt 8. Explain what is meant by the term deliquescent and name one deliquescent salt 9. What is meant by the term “hydrated salt”? 10. A sample of hydrated calcium nitrate is found to have a molar mass of 236.1. Deduce its dot formula. 11. What is effluoresence? Give an example of an effluorescent salt. |

###### *Lesson 3 – What is the difference between strong and weak acids?*

1. **Strong and Weak Acids and Bases**

**(i) strong and weak acids**

* Acids react with water to form solutions containing H+ ions:

Eg HCl 🡪 H+ + Cl- Eg HNO3 🡪 H+ + NO3-

In hydrochloric acid and nitric acid, the molecule dissociates completely into the H+ ion and the anion; there is no molecular HCl or HNO3 remaining

An acid which completely dissociates in water to form hydrogen ions is called a **strong** acid; a **strong acid is a substance which completely dissociates in water to form H+ ions;** HCl, HNO3 and H2SO4 are strong acids

* Most acids, however, only dissociate slightly in water; the H+ ions, anions and undissociated molecules form a dynamic equilibrium:

Eg CH3COOHH+ + CH3COO-

Eg HFH+ + F-

If the equilibrium position lies mainly on the left-hand side (ie only a small number of the molecules have dissociated), then the acid is called a **weak** acid; a **weak acid is a substance which only slightly dissociates in water to form H+ ions; a**ll carboxylic acids, such as methanoic acid and ethanoic acid, are weak acids

1. **Strong and weak bases**

* Bases react with water to form solutions containing OH- ions:

Eg NaOH 🡪 Na+ + OH- Eg Ba(OH)2 🡪 Ba2+ + 2OH-

In sodium hydroxide and barium hydroxide, the solid dissolves completely into the OH- ion and the cation; there is no undissolved or undissociated NaOH or Ba(OH)2; a base which completely dissociates in water to form OH- ions is called a **strong** base; a **strong base is a substance which completely dissociates in water to form OH- ions;** NaOH, KOH and BaO are strong bases

* Some bases are highly soluble in water, but do not fully react with water to form hydroxide ions; the OH- ions, cations and undissociated molecules form a dynamic equilibrium:

Eg NH3 + H2ONH4+ + OH-

Eg CO32- + H2O HCO3- + OH-

If the equilibrium position lies mainly on the left-hand side (ie only a small number of the molecules have reacted), then the bases is called a **weak** base; a **weak base is a substance which only slightly dissociates in water to form OH- ions;** ammonia and sodium carbonate are weak bases.

* There are also some bases, such as Ca(OH)2, which are only slightly soluble in water but which fully dissociate into OH- ions once they are dissolved; these are also classified as weak bases because they do not produce solutions containing a lot of OH- ions due to their low solubility

1. **Properties of strong acids and bases compared to weak acids and bases**

**(α) Electrical conductivity:**

* **S**trong acids and bases are completely dissociated into their ions:
* a 1 moldm-3 solution of HCl contains 1 moldm-3 H+ and 1 moldm-3 Cl-; it does not contain any undissociated HCl
* a 1 moldm-3 solution of NaOH contains 1 moldm-3 OH- and 1 moldm-3 Na+’ it does not contain any undissociated NaOH
* these solutions both contain the same number of ions and so will both show the approximately the same electrical conductivity in aqueous solution
* Weak acids and bases are only slightly dissociated into their ions:
* 1 moldm-3 solution of CH3COOH contains mostly undissociated CH3COOH; it contains a very small quantity of CH3COO- and H+ ions (around 0.01 moldm-3)
* a 1 moldm-3 solution of NH3 contains mostly undissociated NH3. It contains a very small quantity of NH4+ and OH- ions (around 0.01 moldm-3)

Solutions of weak acids and bases contain far fewer ions than solutions of strong acids and alkalis of the same concentration; they therefore have a much lower electrical conductivity than **solutions of strong acids and alkalis of the same concentration**

**(β) Enthalpy of neutralisation:**

* Solutions of strong acids contain fully dissociated H+ ions; solutions of strong bases contain fully dissociated OH- ions:
* when a solution of a strong acid reacts with a solution of a strong base, the H+ ions react with the OH- ions to make H2O: **H+ + OH- 🡪 H2O**
* the remaining ions are not involved in the reaction; this is true of a neutralisation reaction between any strong acid and any strong base – in all cases, an identical chemical reaction is taking place
* as a result, all reactions between strong acids and strong bases in solution have the same enthalpy of neutralisation (-58 kJmol-1)
* eg HNO3 + NaOH 🡪 NaNO3 + H2O (ΔH = -58 kJmol-1); HCl + KOH 🡪 KCl + H2O (ΔH = -58 kJmol-1)
* Solutions of weak acids only contain small quantities of H+ ions
* when weak acids are neutralised by strong bases, the OH- ions react with the H+ ions which quickly run out
* as a result, the weak acid dissociates more, producing more H+ ions (according to Le Chatelier’s principle); this process continues until the weak acid has fully dissociated; so during the neutralisation of weak acids, two reactions are taking place:

Step 1: the acid dissociates to form more H+: CH3COOH 🡪 CH3COO- + H+

Step 2: the resulting H+ reacts with the OH-: H+ + OH- 🡪 H2O

* the first reaction, the dissociation, is endothermic by around 2 kJmol-1; the second reaction, the neutralisation, has a molar enthalpy change of -58 kJmol-1; the overall molar enthalpy change of neutralisation is therefore -56 kJmol-1
* eg CH3COOH + NaOH 🡪 CH3COONa + H2O ΔH = -56 kJmol-1
* weak acids have a less exothermic enthalpy of neutralisation than strong acids, because the H+ ion has to be removed from the acid first, which absorbs some of the energy
* Solutions of weak bases only contain small quantities of OH- ions:
* When weak bases are neutralised by strong acids, the H+ ions react with the OH- ions which quickly run out
* as a result, the weak base dissociates more, producing more OH- ions (according to Le Chatelier’s principle); this process continues until the weak base has fully dissociated; so during the neutralisation of weak bases, two reactions are taking place:

Step 1: the base dissociates to form more OH-: NH3 + H2O 🡪 NH4+ + OH-

Step 2: the resulting OH- reacts with the H+: H+ + OH- 🡪 H2O

* the first reaction, the dissociation, is endothermic by around 5 kJmol-1; the second reaction, the neutralisation, has a molar enthalpy change of -58 kJmol-1; the overall molar enthalpy change of neutralisation is therefore -53 kJmol-1
* eg NH3 + HCl 🡪 NH4Cl ΔH = -53 kJmol-1
* Weak bases have a less exothermic enthalpy of neutralisation than strong bases, because the OH- ion has to be produced from the base first, which absorbs some of the energy

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 3.1: Compare the enthalpy of neutralisation of a strong base and a weak base by the same acid**   1. Place a clean, dry polystyrene cup inside a glass beaker. 2. Pour 25 cm3 of 2 moldm-3 HCl into the cup using a measuring cylinder. Record the initial temperature of the solution. 3. Measure out 25 cm3 of 2 moldm-3 NaOH using a measuring cylinder and record the initial temperature of the solution. Hence deduce the average initial temperature of the solutions. 4. Add the NaOH solution to the HCl solution and stir continuously, until the temperature either remains constant for over a minute or starts to change in the opposite direction. Record the final temperature reached. 5. Deduce the maximum temperature change. 6. Repeat steps 1 – 5, but this time use 25 cm3 of 2 moldm-3 NH3 instead of NaOH. 7. Which reaction gives you a greater temperature change? Can you explain why? |

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| **Image result for test iconTest your knowledge 3.2: Distinguishing between Strong and Weak Acids and Bases**   1. What is the difference between a strong acid and a weak acid? Give one example of each. 2. Give one example of a strong base, one example of a base which is weak due to low solubility and one example of a base which is weak due to small degree of ionisation 3. Why do strong acids conduct electricity better than weak acids? 4. Why is the enthalpy of neutralisation of weak acids less exothermic than the enthalpy of neutralisation of weak acids? |

***Lesson 4 - What are the other important reactions of acids, bases and salts?***

1. **Further Reactions of Acids, Bases and Salts**

**(i) Amphoterism**

* Proton donors (acids) will react with proton acceptors (bases), usually to form salts; as well as the common acids (HCl, HNO3 and H2SO4), which contain H+ ions, all oxides and hydroxides of non-metals (ie all covalently bonded oxides and hydroxides) can behave as acids; some dissolve in water to form solutions containing H+ ions; some do not dissolve in water but react with alkalis to form salts; examples of acidic oxides are CO2, NO2, P4O6, P4O10, SO2, SO2, Cl2O
* Eg SO3 + H2O 🡪 H2SO4 and SO3 + 2NaOH 🡪 Na2SO4 + H2O (salt = sodium sulphate)
* Eg SiO2 + 2NaOH 🡪 Na2SiO3 + H2O (salt = sodium silicate)
* Proton acceptors (bases) will react with proton donors (acids), usually to form salts; most oxides and hydroxides of metals (ie all ionically bonded oxides and hydroxides) can behave as bases; some dissolve in water to form solutions containing OH- ions; ome do not dissolve in water but react with acids to form salts:
* Eg Na2O + HO 🡪 2NaOH and Na2O + HCl 🡪 2NaCl + H2O
* Eg Mg(OH)2 + 2HCl 🡪 MgCl2 + 2H2O
* There are a small number of oxides and hydroxides of metals which can form salts by reacting with both acids and alkalis:
* Eg Al(OH)3 + 3HCl 🡪 AlCl3 + 3H2O and Al(OH)3 + NaOH 🡪 NaAl(OH)4
* Eg ZnO + 2HCl 🡪 ZnCl2 + H2O and ZnO + 2NaOH 🡪 Na2ZnO2 + H2O

These substances can react with both acids and bases; they are proton donors and proton acceptors and are said to be **AMPHOTERIC**; the bonding in these oxides and hydroxides tends to be intermediate between ionic and covalent

(The acid-base properties of oxides and hydroxides will be studied in more detail later in the course)

* Other substances can show amphoteric properties by behaving as both proton donors and proton acceptors; one of the most common examples is the hydrogencarbonate ion (HCO3-)

NaHCO3 + HCl 🡪 NaCl + CO2 + H2O and NaHCO3 + NaOH 🡪 Na2CO3 + H2O

**(ii) Acid-Salt Reactions**

* If a strong acid reacts with a salt of a weaker acid, the stronger acid will displace the weaker acid from its salt, by forcing the salt of the other acid to accept the hydrogen ion
* For example, sulphuric acid is a stronger acid than hydrochloric acid; as a result, if concentrated sulphuric acid is reacted with any salt of hydrochloric acid (eg potassium chloride), the sulphuric acid will give its hydrogen ion to the chloride ion, forming hydrochloric acid and a sulphate salt (eg potassium sulphate); H2SO4 + 2KCl 🡪 K2SO4 + 2HCl
* In some cases, a weak acid will even react with the salt of a stronger acid; for example, phosphoric acid can react with chloride salts to make phosphate salts, even though phosphoric acid is a weaker acid than hydrochloric acid; the reaction is possible because the HCl escapes as a gas, preventing the reverse reaction and pushing the position of equilibrium to the right according to Le Chatelier’s principle: H3PO4(l) + 3KCl(s) 🡪 K3PO4(s) + 3HCl(g)
* In acid-salt reactions, the anion in the salt is accepting a proton from the acid; it is therefore acting as a base; acid-salt reactions are just a special type of acid-base reaction

**(iii) Salt Hydrolysis**

* Salts can be classified into three main types:
* salts formed from strong acids and strong bases (eg NaNO3, K2SO4 and BaCl2)
* salts formed from weak acids and strong bases (eg CH3COONa, KCN)
* salts formed from strong acids and weak bases (eg NH4Cl, NH4NO3)

Salts can also be formed from weak acids and weak bases, but these tend not to be stable

* Anions produced from weak acids can react partially with the water in the solution to reform the original acid: eg CH3COO- + H2OCH3COO + OH-
* the resulting solution, therefore, contains OH- ions as well as the salt
* this type of reaction occurs in salts made from weak acids; if the acid is reluctant to give up its H+ in the first place, it is more likely that the anion will pick up an H+ to reform the acid
* salts containing anions which do this are known as basic salts; this is an example of salt hydrolysis
* examples of basic salts are CH3COONa and KCN; they are made from weak acids and strong bases
* Cations produced from weak bases in the salt can react partially with the water in the solution to reform the original base: eg NH4+NH3 + H+
* the resulting solution, therefore, contains H+ ions as well as the salt
* this type of reaction occurs in salts made from weak bases; if the base is reluctant to accept an H+ ion in the first place, its cation is more likely release an H+ and reform the original base
* salts which do this are known as **acidic salts**; this is another example of salt hydrolysis
* examples of acidic salts are NH4Cl and NH4NO3; they are made from strong acids and weak bases
* Some highly charged cations also react with water to form H+ ions, in order to reduce their charge; they are accepting OH- ions from the water and partly reforming the original hydroxide:

eg Al3+ + H2OAl(OH)2+ + H+

* the resulting solution contains H+ ions as well as the salt
* salts which do this are also known as acidic salts; this is another example of salt hydrolysis
* an example of this type of acidic salt is AlCl3 or CrCl3
* Anions made from strong acids and cations made from strong bases do not undergo salt hydrolysis; salts made from strong acids and strong bases are therefore known as neutral salts
* The tendency of a salt to itself react with water and show acidic or basic properties is known as **salt hydrolysis**

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| **Image result for test iconTest your knowledge 4.1: Understanding Further Reactions of Acids, Bases and Salts**   1. What is meant by the term “amphoteric”; give two examples of amphoteric substances 2. Write equations for the following reactions:   (i) sulphuric acid and potassium chloride  (ii) nitric acid and sodium fluoride  (iii) phosphoric acid (H3PO4) and potassium bromide   1. Predict whether the following salts will undergo salt hydrolysis; if they do, write an equation to show how one of the ions in the salt reacts with water to produce H+ or OH- ions   (i) sodium nitrate  (ii) ammonium nitrate  (iii) sodium ethanoate  (iv) potassium sulphate  (v) aluminium sulphate  (vi) potassium cyanide |

***Lesson 5 – How can we use acid-base reactions to prepare salts in the laboratory?***

1. **Laboratory Preparation of Salts by Neutralisation**

* The most suitable method of preparing a salt by a neutralisation reaction depends on the solubility of the salt and the solubility of the acid and the base used to make it; insoluble salts are generally not prepared by neutralisation reactions, but soluble salts frequently are; there are two common types of preparation:
* Soluble salts can be prepared from **an acid and an insoluble base**:
* if the salt being prepared is soluble, the easiest way to make it is from an acid and an insoluble base
* an excess of the insoluble base must be added to ensure that the acid reacts completely
* the salt solution can be separated from the excess base by simple filtration
* the salt can be extracted from the salt solution by gradually evaporating off the water
* copper sulphate is an example of a salt which can be prepared by this method

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 5.1: Prepare a salt by neutralising an acid with an insoluble base**   1. Measure out 20 cm3 of 0.5 moldm-3 sulphuric acid into a 100 cm3 beaker. 2. Use a Bunsen burner to warm the beaker gently on a tripod until the temperature reaches 50 oC. 3. Add 1 g of copper oxide gradually to the beaker over a period of 2 minutes, stirring gently. 4. Heat gently for a few minutes until it is almost boiling. 5. Allow the mixture to cool. 6. Place a folded piece of filter paper inside a filter funnel, and then place the funnel into the neck of a 100 cm3 conical flask. 7. Pour the warm mixture into the filter funnel and allow the solution to filter through. A clear blue solution should collect in the conical flask. 8. Rinse the beaker and then pour the clear blue solution back into it. Label the beaker with your name and leave it for a week. 9. Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.  * How many moles of sulphuric acid (H2SO4) are used in the reaction? * How many moles of copper oxide (CuO) are used in the reaction? * Why is it important that the CuO is in excess? |

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| Cannot do this experiment? Watch this video: [www.youtube.com/watch?v=qIOMlwBoe\_4](http://www.youtube.com/watch?v=qIOMlwBoe_4) |

* Group I and ammonium salts cannot be prepared by neutralisation of insoluble bases as ammonia and all Group I bases are soluble; these salts can only be prepared by reacting acids and alkalis
* the acid and alkali must be added in the exact amounts required, as there is no way to separate the salt produced from any remaining acid or alkali; the final mixture should therefore only contain the salt and water
* the salt can be extracted from the salt solution by gradually evaporating off the water
* ammonium sulphate is an example of a salt which can only be prepared by this method

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 5.2: Prepare a salt by neutralising an acid with a soluble base**   1. Measure out 25 cm3 of 2 moldm-3 ammonia solution into a 100 cm3 beaker. 2. Add 25 cm3 of 1 moldm-3 sulphuric acid solution to the same beaker. Stir the mixture gently. 3. Pour the resulting mixture into an evaporating dish and heat over a Bunsen burner until most (but not all) of the water has evaporated. 4. Leave the remaining solution for a week. 5. Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.  * Write an equation for the reaction between sulphuric acid and ammonia. * How many moles of sulphuric acid are used in the reaction? * How many moles of ammonia are used in the reaction? * Why is it important to use these exact quantities of the reactants? * Why is it easier to make a pure salt if the base is insoluble rather than soluble? |

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| Cannot do this experiment? Watch this video: <https://www.youtube.com/watch?v=Yh7z2ThG32o> |

* Salts can also be prepared from acid-salt reactions; the salts most commonly prepared in this way are phosphate salts, by reacting concentrated H3PO4 with chloride salts:

Eg 3CaCl2(s) + 2H3PO4 (l) 🡪 Ca3(PO4)2(s) + 6HCl(g)

***Lesson 6 – How are acid-base reactions useful in qualitative analysis?***

1. **Qualitative Analysis Part 1a: Acid-Base Reactions**

* Qualitative analysis is the experimental identification of a substance or of a particular species present in a substance; this can include cations, anions, gases or organic molecules
* Some cations, anions and gases can be identified using simple acid-base reactions:

**(i) H+ ions with CO32- and SO32- ions**

* H+ ions react with CO32- and HCO3- ions to give carbon dioxide and water:

2H+ + CO32- 🡪 CO2 + H2O; H+ + HCO32- 🡪 CO2 + H2O

* H+ ions react with SO32- and HSO3- ions to give sulphur dioxide and water:

2H+ + SO32- 🡪 SO2 + H2O; H+ + HSO32- 🡪 SO2 + H2O

* if a solid sample of a carbonate (eg CaCO3) is added to a solution and bubbles/fizzing (also known as effervescence) is evolved, it shows that the solution contains H+ ions
* if an acid such as HCl is added to a solid sample and effervescence is observed, then the gas should be tested; if the gas is shown to be CO2 then CO32- or HCO3- ions are present in the solid; if the gas is shown to be SO2 then SO32- or HSO3- ions are present in the solid

**(ii) CO2 and SO2 with limewater**

* CO2 reacts with Ca(OH)2 (limewater) to produce insoluble CaCO3, which turns the limewater a milky colour: CO2(g) + Ca(OH)2(aq) 🡪 CaCO3(s) + H2O(l)
* SO2 reacts with Ca(OH)2 (limewater) to produce insoluble CaSO3, which also turns the limewater a milky colour: SO2(g) + Ca(OH)2(aq) 🡪 CaSO3(s) + H2O(l)
* if either gas is bubbled through the limewater for a longer period of time, a second reaction will happen and the milkiness will disappear:

CaCO3(s) + CO2(g) + H2O(l) 🡪 Ca(HCO3)2(aq)

CaSO3(s) + SO2(g) + H2O(l) 🡪 Ca(HSO3)2(aq)

* CO2 and SO2 are the only common gases which have this effect on limewater
* Note: it is easy to distinguish between CO2 and SO2; CO2 is odourless but SO2 smells like a burning match

**(iii) NH4+ ions with OH- ions**

* NH4+ ions react with OH- ions to produce NH3; NH4+ + OH- 🡪 NH3 + H2O
* when warmed gently, the NH3 is given off as a pungent-smelling gas
* if NaOH is added to a solution and the mixture is warmed, and as a result a pungent smell is observed, then the solution contains NH4+ ions
* if NH4Cl is added to a solution and the mixture is warmed, and as a result a pungent smell is observed, then the solution contains OH- ions

1. **NH3 with HCl**

* NH3 (an alkaline gas) and HCl (an acidic gas) react together to give a salt (NH4Cl):

NH3(g) + HCl(g) 🡪 NH4Cl(s)

* this salt is clearly visible as a white smoke which appears when the gases come into contact with each other
* If a piece of filter paper is soaked in concentrated HCl and placed close to a source of NH3, a white smoke can be seen, confirming the presence of NH3
* If a piece of filter paper is soaked in concentrated NH3 and placed close to a source of HCl, a white smoke can be seen, confirming the presence of HCl

1. **Concentrated H2SO4 and Cl- ions**

* If concentrated sulphuric acid is added to a solid sample containing Cl- ions, an acid-salt reaction will take place and HCl (g) will be produced: H2SO4 + Cl- 🡪 HSO4- + HCl
* If a piece of filter paper is soaked in concentrated NH3 and placed close to a source of HCl, a white smoke can be seen, confirming the presence of HCl

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 6.1: Use acid-base reactions to identify cations, anions and gases**  You are given four solid samples to test, labelled A, B, C and D; each solid contains a different one of the following ions: NH4+, OH-, CO32-, SO32-   1. Place a small spatula of sample A onto a watch glass; add a few drops of HCl and observe whether bubbles are formed; if bubbles are formed, observe whether or not the gas smells like a burning match; repeat for samples B, C, and D 2. If bubbles are observed, half-fill a test-tube with limewater; add a spatula measure of the same sample into another test-tube; add 3 cm3 of HCl to the sample, attach a bung with a delivery tube to the test tube and insert the delivery tube into the limewater in the other test tube; wait for a few minutes until there is no further change in the limewater 3. Place a small spatula of sample A onto a watch glass; add a few drops of NaOH; observe whether or not a pungent smell is given off; repeat for samples B, C and D 4. Place a small spatula of sample A onto a watch glass; add a few drops of NH4Cl; observe whether or not a pungent smell is given off; repeat for samples B, C and D 5. Place a very small spatula of sample A into a test tube and add 2 – 3 cm3 of water so that the solid is mostly dissolved; then add a small spatula measure of powdered CaCO3; observe whether bubbles are formed; repeat for samples B, C and D 6. Use your observations to identify which ion is present in which sample; present your answers in a table similar to the one below:      1. Write equations for the reaction occurring with each sample which results in a positive identification |

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| Image result for test icon**Test your knowledge 6.2: Using acid-base reactions to identify certain cations and anions**  Describe a simple test to show how you would confirm the presence of the following ions:   1. CO32- in a solid sample 2. SO32- in a solid sample 3. Cl- in a solid sample 4. H+ in solution 5. OH- in solution 6. NH4+ in solution 7. NH3 as a gas 8. HCl as a gas   In each case, give the equations for the reactions occurring |

***Lesson 7 - What is the pH scale?***

1. **Indicators and the pH Scale**

**(i) The pH Scale**

* Water dissociates very slightly to produce H+ and OH- ions: H2OH+ + OH-; as a result, all aqueous solutions contain both H+ and OH- ions
* Solutions which contain acids contain many more H+ ions than pure water; as a result, these solutions also contain fewer OH- ions than pure water, because the extra H+ ions push the above equilibrium to the left; the higher the concentration of H+ ions, the lower the concentration of OH- ions; any solution in which the concentration of H+ ions is greater than the concentration of OH- ions is said to be an **acidic** solution
* Solutions which contain alkalis contain many more OH- ions than pure water; as a result, these solutions also contain fewer H+ ions than pure water, because the extra OH- ions push the above equilibrium to the left; the higher the concentration of OH- ions, the lower the concentration of H+ ions; any solution in which the concentration of OH- ions is greater than the concentration of H+ ions is said to be an **alkaline** solution
* A solution in which the concentration of H+ ions is equal to the concentration of OH- ions, as is the case in pure water, is said to be a neutral solution
* The level of acidity or alkalinity of a solution (ie the relative concentrations of H+ and OH- ions is measured on a scale called the **pH scale**:
* the lower the pH, the more acidic the solution (the higher the concentration of H+ ions and the lower the concentration of OH- ions); the most acidic solutions possible have a pH of around -1
* the higher the pH, the more alkaline the solution (the higher the concentration of OH- ions and the lower the concentration of H+ ions). The most alkaline solutions possible have a pH of around 15
* Neutral solutions have a pH of 7
* The relationship between pH, acidity and alkalinity is summarised in the table below:

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| pH | -1 | 1 | 3 | 5 | 7 | 9 | 11 | 13 | 15 |
| Acidity | strongly acidic | | slightly acidic | | Neutral | slightly alkaline | | Strongly alkaline | |
| [H+] | very high | | quite high | | Normal | quite low | | very low | |
| [OH-] | very low | | quite low | | Normal | quite high | | very high | |

* Examples of the pH of common solutions are:
* 1 moldm-3 HCl has a pH of 0
* lemon juice and vinegar have a pH of approx. 2
* orange juice has a pH of approx. 3
* rainwater has a pH of approx. 6
* pure water has a pH of 7
* household bleach has a pH of approx. 12
* 1 moldm-3 NaOH has a pH of 14
* The pH of a solution can be measured directly using a **pH meter** (although these can be difficult to use)

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| Image result for test icon**Test your knowledge 7.1: Understanding the pH scale**  Classify the following solutions as acidic, alkaline or neutral based on the information provided: | |
| 1. it has a pH of 5 2. it is a solution of ammonia 3. [H+] = [OH-] 4. it is lemon juice 5. it has a pH of 11 6. [H+] > [OH-] 7. It is a solution of sodium hydroxide 8. It is a solution of sodium chloride | 1. it is a solution of sodium carbonate 2. it has a pH of 7 3. it has a pH of -1 4. [OH-] > [H+] 5. it is rainwater 6. it is pure water 7. it is a solution of ammonium chloride 8. it is a solution of potassium ethanoate |

1. **Indicators**

* An indicator is a substance which appears as one colour in acidic solutions and as a different colour in alkaline solutions; indicators can therefore be used to determine whether a solution is acidic or alkaline; most indicators are soluble organic acids or bases
* Most indicators are weak acids in which the acid and its anion have different colours:

HA  H+ + A-

Colour 1 Colour 2

Some indicators are weak bases in which the base and its cation have different colours:

BH+  H+ + B

Colour 1 Colour 2

* at low pH, the H+ ions shift the position of equilibrium to the left and the indicator will appear as colour 1
* at high pH, the [H+] is low so the equilibrium shifts to the right to produce more H+ and the indicator will appear as colour 2
* **hence indicators therefore appear as one colour in low pH, and a different colour in high pH**
* The pH range over which an indicator changes colour varies from indicator to indicator:

|  |  |  |  |
| --- | --- | --- | --- |
| Indicator | Colour 1 (low pH) | Colour 2 (high pH) | pH of colour change |
| Methyl orange | Pink | Yellow | 3 – 5 |
| Litmus | Red | Blue | 5 – 8 |
| Phenolphthalein | Colourless | Pink | 8 – 10 |

* The effect of acids and bases on indicators depends on the exact pH of the solution, and the pH at which the indicator changes colour:
* Strongly acidic solutions with a pH of 3 or below will cause all three indicators to appear as Colour 1
* Weakly acidic solutions with a pH of 5 – 6 will cause methyl orange to appear as Colour 2 but will cause litmus and phenolphthalein to appear as colour 1
* Neutral solutions (pH = 7) will cause methyl orange to appear as Colour 2, litmus to appear as an intermediate colour (or whatever colour it was before) but will cause phenolphthalein to appear as colour 1
* Alkaline solutions with a pH of 10 or more will cause all three indicators to appear as Colour 2
* Single indicators, therefore, cannot give you the exact pH of a solution; they will simply tell you whether the pH is above or below a certain number

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| cid:ii_jepnvfe00_1621fa54a497745d**Practical 7.2: Investigate the effect of solutions of different pH values on different indicators**  You have access to five different solutions, with pH values of 3, 5, 7, 9 and 11 and labelled accordingly.  You have access to three different indicators: litmus, methyl orange and phenolphthalein   1. Pour 1 cm3 of the solution of pH 3 into each of three different test tubes 2. Add two drops of methyl orange to the first test tube and note the colour; add two drops of litmus to the second test tube and note the colour; add two drops of phenolphthalein to the third test tube and note the colour 3. Repeat steps (a) and (b) for the solutions with pH values of 5, 7, 9 and 11 4. Record your results in a table similar to the one below:      1. Deduce the approximate pH at which each of the three indicators changes colour |

***Lesson 8 - What is universal indicator and why is it useful?***

1. **Universal Indicator**

* Universal indicator is a mixture of several different indicators, each of which changes colour at a different pH; as a result, universal indicator appears at a different colour for very pH unit between 3 and 11:
* at pH values of 3 and below, universal indicator appears red
* between the pH values of 3 and 5, universal indicator appears orange
* between the pH values of 5 and 6, universal indicator appears yellow
* between the pH values of 6 and 8, universal indicator appears green
* between the pH values of 8 and 9, universal indicator appears blue
* between the pH values of 9 and 11, universal indicator appears indigo
* at pH values of 11 and above, universal indicator appears violet
* Universal indicator is most useful for finding out the approximate pH of aqueous solutions; it can be used in paper form or in solution form; in most cases, a colour chart is also provided and the colour of the indicator can be compared to the chart to deduce the pH of the solution; it is not an accurate method of determining pH, but is very simple and easy to use
* Using universal indicator to measure pH is an example of **colorimetry** – the use of colour to make measurements
* pH meters measure pH more accurately, but they are not as simple or convenient to use

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg**Activity 8.1: Universal Indicator and pH**  The following chart shows the different colours shown by universal indicator in solutions with different pH values, but it is not in colour:    Copy the table above, but use coloured pencils, pens or crayons to make sure that each box is shaded in a colour which matches that shown by universal indicator at that pH. |
| https://image.freepik.com/free-icon/plus-sign_318-54005.jpg**Extension 8.2: Universal Indicator and pH**  Completed Activity 20.1? Underneath your chart, give an example of a common substance with a pH similar to that shown. |

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| cid:ii_jepnvfe00_1621fa54a497745d**Practical 8.3: Determine the pH value of various solutions by colorimetry**  You have access to five different solutions, labelled A, B, C, D and E. Each solution has a different pH.   1. Using a dropping pipette, place around 1 cm3 of solution A into a test tube 2. Add three drops of universal indicator and note the colour 3. Repeat Steps 1 – 3 using solutions B, C, D and E 4. Record your results in the table below; state the pH of each solution and suggest what the solution might be. |

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| cid:ii_jepnvfe00_1621fa54a497745d**Practical 8.4: Determine the pH of different soil samples**  You are given access to three different soild samples, taken from somewhere close to where you live.   1. Take 20 g of the soil to be tested and place it in a 100 cm3 beaker. 2. Add 40 cm3 of water to the beaker using a measuring cylinder. 3. Leave the beaker for around 30 minutes, stirring every few minutes. 4. Fold and insert a piece of filter paper into a funnel and pour the mixture through the funnel into a boiling tube. When around 10 cm3 of filtrate has been collected, add a few drops of UI solution and note the colour. 5. Repeat for any other soil samples you need to test. 6. Record your results in a suitable table. |

***Lesson 9 – How can indicators be used in qualitative analysis?***

**h) Qualitative Analysis Part 1b – Using Indicators**

* Indicators can be used in qualitative analysis to identify the presence of H+ ions or OH- ions in a solution; the best indicator to use if you simply want to determine whether a solution is acidic, alkaline or neutral is litmus; this is because litmus changes colour at a pH of around 7; i is usually used soaked in paper, as **litmus paper**; litmus paper can be red or blue; red litmus paper contains litmus in Colour 1 form and blue litmus paper contains litmus in Colour 2 form
* acidic solutions (containing H+) will turn blue litmus paper red but will have no effect on red litmus paper
* alkaline solutions (containing CO32- or HCO3-) will turn red litmus paper blue but will have no effect on blue litmus paper
* neutral solutions will have no effect on either type of litmus paper
* Litmus paper can also be used to test for certain gases which have acidic or alkaline properties:
* HCl and SO2 are acidic gases which turn blue litmus paper red
* NH3 is an alkaline gas which turns red litmus paper blue

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| cid:ii_jepnvfe00_1621fa54a497745d**Practical 9.1: Investigate the effect of acidic, alkaline and neutral solutions on different indicators**  You have access to four different solutions: HCl(aq); NaOH(aq), Na2CO3(aq) and NaCl(aq)  You have access to three different indicators: litmus, methyl orange and phenolphthalein   1. Pour 5 cm3 of HCl(aq) into a beaker 2. Dip a strip of red litmus paper into the solution; note any colour change 3. Dip a strip of blue litmus paper into the solution; note any colour change 4. Rinse out the beaker and measuring cylinder and repeat steps (a) – (c) using the other solutions (NaOH, Na2CO3 and NaCl) 5. Pour 1 cm3 of each solution (HCl, NaOH, Na2CO3 and NaCl) separately into four test tubes 6. Add two drops of methyl orange indicator to each test tube; note any colour change 7. Rinse out the test tubes and repeat steps (e) and (f), but this time add two drops of phenolphthalein; note any colour changes 8. Record your results in a table similar to the one below: |

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| Image result for test icon**Test your knowledge 9.2: Summarising the Qualitative Analysis of Cations, Anions and Gases**   1. Describe a test for H+ ions which uses a simple chemical reaction 2. Describe a test for H+ ions which uses an indicator 3. Describe a test for OH- which uses a simple chemical reaction and a smell 4. Describe a test for OH- which uses an indicator 5. Name three gases which turn blue litmus paper red 6. Name one gas which turns red litmus paper blue 7. Name two gases which turn limewater milky and then clear again and suggest how the gases could be distinguished 8. Describe a test for HCl gas which does not use an indicator 9. Describe a test for NH3 gas which does not use an indicator 10. Describe a test for Cl- ions in a solid sample 11. Describe a test for NH4+ ions 12. Describe a test for CO32- ions in a solid sample 13. Describe a test for SO32- ions in a solid sample 14. Suggest how you could demonstrate that a solution contains a neutral salt |

***Lesson 10 – How I determine how much of an acid or a base is present in a sample (practical)?***

1. **Quantitative Analysis - Acid-Base Titrations**

* Quantitative analysis is the experimental determination of how much of a substance is present in a sample; it is possible to carry out the quantitative analysis of acids and bases using a technique called titration
* A titration is an experiment used to find out what exact volume of one solution is required to react completely with a fixed volume of another solution in a flask; it is the most common example of volumetric analysis (a type of quantitative analysis which measures volumes); from the titration result, you can deduce the concentration of the alkali if the concentration of the acid is known, or vice versa

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| **Summary Activity 10.1: What can you remember about the different instruments used to measure the volume of a solution?**   * What are the names of the different instruments used to measure the volume of solution? * What are the advantages and disadvantages of each instrument? * Which instruments would be most useful when carrying out a titration? |

* In most titrations, a fixed volume (usually 25 cm3) of alkali is prepared in a conical flask using a pipette and an indicator added to the alkali; the acid is then added gradually from the burette into the conical flask until the indicator changes colour (it is better to place the acid in the burette as alkalis can leave residue in the burette which makes it harder to use)
* Burettes are calibrated to measure the volume of solution delivered – the readings on the burette do not tell you the volume of solution actually in the burette; you must first record the burette reading before any solution has been delivered, and then record it again after the reaction is complete; the difference between these two volumes is the volume of solution delivered; this is also known as the titre volume: titre volume = final burette reading – initial burette reading
* It is important to choose the indicator correctly:
* the point at which the alkali has been completely neutralised by the acid is called the **equivalence point** of the titration, and the pH of the solution changes sharply at this point; the precise pH change at the end point depends on the type of titration being carried out
* different indicators change colour at different pH values; for example methyl orange changes colour at pH 3 -5 and phenolphthalein changes colour at pH 8 – 10; the pH range over which the indicator changes colour is called the **end-point** of the indicator

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| Indicator | Colour 1 (low pH) | Colour 2 (high pH) | End point pH range |
| Methyl orange | Red | Yellow | 3 – 5 |
| Phenolphthalein | Colourless | Pink | 8 – 10 |

* an indicator will only change colour at the equivalence point of a titration if the end-point pH range of the indicator lies within the pH change at the equivalence point of the titration

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| Type of titration | pH change at equivalence point | methyl orange suitable? | phenolphthalein suitable? |
| strong acid - strong base | 11 to 3 | Yes | yes |
| weak acid - strong base | 11 to 7 | No | Yes |
| strong acid - weak base | 7 to 3 | Yes | No |
| weak acid-weak base | No sharp change | No | No |

* either methyl orange or phenolphthalein can be used for strong acid-strong base titrations
* only phenolphthalein can be used for weak acid-strong base titrations
* only methyl orange can be used for strong acid-weak base titrations
* there is no suitable indicator for weak acid-weak base titrations

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 10.2: Determine the concentration of a solution of NaOH by titration against 1.0 moldm-3 HCl**   1. **Add 0.1 moldm-3 hydrochloric acid to a burette and record the initial volume**   **Set up the burette for the first time (you only have to do this before the first titration)**   * Ensure that the tap is closed * Rinse the burette with acid to remove any impurities: pour around 10 cm3 of acid into the burette using a funnel; turn the burette as far horizontally as you can without spilling the acid and then rotate the burette a few times so that the acid covers all parts of the inside of the glass; then empty the burette by holding the tip of the burette downwards over a sink and opening the tap until the acid has come out; then close the tap * Use a boss to attach a clamp to a stand and clamp a burette vertically into position; the tip of the burette should be pointing downwards and should be around 20 cm above the bench surface; make sure the tap is closed * Using a funnel, add the acid solution gradually to the burette until it is around half-full; open the tap of the burette and allow acid to escape until the tip of the burette is completely full (there should be no air bubbles); you may need to flick the tip a few times to remove air bubbles; once the tip is completely full, close the tap again   **Prepare the burette for a titration (you need to do this before every titration)**   * Add more acid into the burette, using the funnel, until the level of solution in the burette is in between the 1 cm3 and the 5 cm3 mark; record this level accurately, to the nearest 0.05 cm3, as the “initial volume” * Remove the funnel from the burette * Once you have reached this stage for the first time, stop and ask the teacher to check your progress  1. **Pipette 25.0 cm3 of NaOH(aq) into a conical flask and add a suitable indicator**  * Pour around 100 cm3 of the NaOH into a beaker * Use a pipette with a pipette filler to suck the alkali into the pipette until it is just above the graduation mark; remove the pipette filler and quickly place your thumb over the top of the pipette; release the pressure gradually until the base of the meniscus lies on the mark; then increase the thumb pressure to fix the solution at that level (it takes practice to get this right) * Still applying pressure with your thumb, move the pipette out of the alkali solution and over a clean, empty 250 cm3 conical flask; remove your thumb and allow the pipette to empty itself under gravity * Once the solution has stopped draining out of the pipette, dip the tip of the pipette into the solution in the conical flask for a few seconds and then remove the pipette completely from the conical flask * Place the empty pipette to one side * Add two drops of phenolphthalein indicator to the alkali; the indicator should turn pink * Once you have reached this stage for the first time, stop and ask the teacher to check your progress  1. **Add the HCl to the NaOH until the indicator changes colour**  * Place the conical flask containing 25.0 cm3 of the alkali and the indicator and place it below the burette; the rim of the conical flask should be just below the tip of the burette * open the tap of the burette and add the solution gradually into the conical flask, swirling the conical flask continuously until around 10 cm3 has been added * close the tip gradually until the solution is dripping slowly into the conical flask at the rate of one drop per second; keep swirling the mixture and continue until the indicator changes colour (into Colour 1); then close the tap * Record the new volume of solution accurately (to the nearest 0.05 cm3) as the “final volume” * Once you have reached this stage for the first time, stop and ask the teacher to check your progress  1. **Repeat the titration until concordancy is achieved**  * After rinsing out the concical flask, steps 1 to 3 (except the burette set-up phase) should be repeated; both titre volumes for both titrations should be calculated and compared * If the two titre volumes are within 0.10 cm3 of each other, the results can be described as concordant and you have completed the experiment * If the values differ by more than 0.10 cm3, the results are not concordant and you must repeat the experiment until you have obtained two concordant titre volumes * Record all your results in a table similar to the one below:      1. **Hence determine the concentration of the sodium hydroxide solution**  * Calculate the average of your concordant values * Hence calculate the number of moles of HCl used in the titration (n = C x V) * Hence calculate the number of moles of NaOH in the conical flask (same as the moles of HCl) * Hence calculate the molarity of the NaOH solution (C = n/V) |
| Need more assistance with this practical? Try these videos: [www.youtube.com/watch?v=Y-5QJIr7Xm4](http://www.youtube.com/watch?v=Y-5QJIr7Xm4) and [www.youtube.com/watch?v=VADhF7M8bJ8](http://www.youtube.com/watch?v=VADhF7M8bJ8) |

* Here is a simple worked example of how to use a titre volume to deduce the molarity of the unknown solution:

Question: In a titration 28.3 cm3 of a 0.10 moldm-3 solution of NaOH was required to react with 25 cm3 of a solution of H2SO4. What was the concentration of the H2SO4 solution?

Worked answer:

* Equation: H2SO4 + 2NaOH 🡪 Na2SO4 + 2H2O
* Moles of NaOH used in titration = 28.3/1000 x 0.1 = 2.8 x 10-3
* 2:1 ratio so moles of H2SO4 present in conical flask = 2.8 x 10-3/2 = 1.4 x 10-3
* so concentration of H2SO4 = 1.4 x 10-3/25 x 1000 = 0.056 moldm-3

***Lesson 11 – How I determine how much of an acid or a base is present in a sample II?***

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| **Summary Activity 11.1: How do you prepare a standard solution?**   * In SS1 you prepared some standard solutions, either by dissolving a solid or by diluting a concentrated solution * How would you prepare 250 cm3 of a 0.1 moldm-3 solution of NaOH, starting with solid NaOH? * How would you prepare 250 cm3 of a 0.2 moldm-3 solution of H2O2, starting with 2.0 moldm-3 H2O2? |

* Once the titre volume has been determined, other desired quantities can be determined using techniques from the amount of substance topic; the simplest quantity to determine is the concentration of the unknown solution (acid or base); other quantities which can be determined from titration results include:
* the molar masses of acids and bases and water of crystallization
* the solubility of acids and bases
* the percentage purity of acids and bases
* In many cases, the solution in the pipette needs to be prepared by dissolving a solid sample or by diluting a more concentrated solution; this requires a standard solution of the substance to be prepared first

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| cid:ii_jepnvfe00_1621fa54a497745d**Practical 11.2: Determine the relative formula mass, and hence water of crystallisation, of hydrated sodium carbonate, Na2CO3.xH2O, by titration (Na2CO3 + 2HCl 🡪 2NaCl + CO2 + H2O)**   1. Weigh out approximately 3.5 g of hydrated sodium carbonate into a weighing boat. Record the exact mass of the solid. 2. Dissolve the solid in approximately 100 cm3 of distilled water. 3. Transfer this solution into a 250 cm3 volumetric flask, and make the volume in the flask up to 250 cm3 by using washings from the beaker; shake continuously to ensure that the contents are thoroughly mixed 4. Pipette a 25.0 cm3 sample of this solution into a conical flask and add a few drops of methyl orange indicator (the solution should turn yellow) 5. Fill a burette with 0.1 moldm-3 HCl and record the initial volume 6. Add the HCl solution to the sodium carbonate solution until the indicator changes colour (from yellow to pink); record the final volume and deduce the titre volume 7. Repeat the titration until concordancy is achieved 8. Hence determine the moles of sodium carbonate present in the conical flask, and in the original volumetric flask 9. Hence determine the molar mass of the hydrated sodium carbonate sample 10. Hence calculate the value of x, given that the formula of sodium carbonate is Na2CO3.10H2O |

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| cid:ii_jepnvfe00_1621fa54a497745d**Practical 11.3: Determine the percentage purity of a sample of vinegar (CH3COOH + NaOH 🡪 CH3COONa + H2O)**   1. A solution of vinegar is labelled as 62.3 gdm-3. Pipette 25 cm3 of this vinegar sample into a 250 cm3 volumetric flask; then rinse out the pipette 2. Make the volume in the flask up to the mark, shaking continuously to ensure that the contents are thoroughly mixed 3. Transfer some of this diluted acid into a burette and record the initial volume 4. Pipette a 25.0 cm3 sample of the 0.05 moldm-3 sodium hydroxide solution into a conical flask and add a few drops of phenolphthalein indicator (the solution should turn pink) 5. Add the diluted vinegar solution to the NaOH solution until the indicator changes colour (from pink to colourless); record the final volume and deduce the titre volume 6. Repeat the titration until concordancy is achieved 7. Hence determine the concentration of the diluted vinegar solution 8. Hence determine the concentration of the original vinegar solution 9. Hence determine the percentage purity of the original vinegar solution 10. Explain why phenolphthalein was used as the indicator for this titration |

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| Image result for test icon**Test your knowledge 11.4: Volumetric Analysis – Titrations**   1. Succinic acid has the formula (CH2)n(COOH)2 and reacts with dilute sodium hydroxide as follows:   (CH2)n(COOH)2 + 2NaOH 🡪 (CH2)n(COONa)2 + 2H2O  2.0 g of succinic acid were dissolved in water and the solution made up to 250 cm3. This solution was placed in a burette and 18.4 cm3 was required to neutralise 25 cm3 of 0.1 moldm-3 NaOH. Deduce the molecular formula of the acid and hence the value of n.   1. Sodium carbonate exists in hydrated form, Na2CO3.xH2O, in the solid state. 3.5 g of a sodium carbonate sample was dissolved in water and the volume made up to 250 cm3. 25.0 cm3 of this solution was titrated against 0.1 moldm-3 HCl and 24.5 cm3 of the acid were required. Calculate the value of x given the equation: Na2CO3 + 2HCl 🡪 2NaCl + CO2 + H2O 2. 25 cm3 of a sample of vinegar (CH3COOH) was pipetted into a volumetric flask and the volume was made up to 250 cm3. This solution was placed in a burette and 13.9 cm3 were required to neutralise 25 cm3 of 0.1 moldm-3 NaOH. Calculate the molarity of the original vinegar solution and its concentration in gdm-3, given that it reacts with NaOH in a 1:1 ratio. 3. 2.5 g of a sample of impure ethanedioic acid, H2C2O4.2H2O, was dissolved in water and the solution made up to 250 cm3. This solution was placed in a burette and 21.3 cm3 were required to neutralise 25 cm3 of 0.1 moldm-3 NaOH. Given that ethanedioic acid reacts with NaOH in a 1:2 ratio, calculate the percentage purity of the sample. 4. A toilet cleaner containing sodium hydrogensulphate, NaHSO4 is believed to have been contaminated. 5.678 g of the sample were dissolved in water and the solution was made up to 250 cm3. This solution was placed in a burette and 23.1 cm3 of it were required to neutralise 25 cm3 of 0.1 moldm-3 sodium hydroxide.   Calculate the percentage purity of the sample given that the species react in a 1:1 ratio. |

***Lesson 12 – What have you understood about Acids, Bases and Salts?***

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| 12.1 END-OF-TOPIC QUIZ  UNIT 5 – ACIDS, BASES AND SALTS  Image result for test icon   1. Write symbol equations for the following neutralisation reactions: 2. hydrochloric acid and sodium hydroxide 3. sulphuric acid and magnesium oxide 4. nitric acid and calcium carbonate 5. Explain the meaning of the term deliquescent, hygroscopic and efflorescent and give one example of an acid, base or salt with each property. 6. State the difference between a strong acid and a weak acid. Explain how 1.0 moldm-3 HCl and 1.0 moldm-3 CH3COOH would differ in terms of their: 7. pH 8. enthalpy of neutralisation by NaOH 9. rate of reaction with calcium carbonate 10. electrical conductivity 11. Suggest whether the following salts undergo salt hydrolysis, and if so, whether they are acidic or basic salts: 12. sodium chloride 13. ammonium chloride 14. sodium ethanoate 15. Describe a chemical test which would identify the following: 16. H+ ions in solution 17. OH- ions in solution 18. NH4+ ions in solution 19. Cl- ions from a solid sample 20. SO32- ions from a solid sample 21. NH3 gas 22. Describe how: 23. universal indicator could distinguish between samples of water, lemon juice and bleach 24. litmus paper could distinguish between HCl and NH3 25. Deduce the identity of the Group 1 metal M based on the following experiment:   2.5 g of the metal carbonate M2CO3 was dissolved in water and the volume made up to 250 cm3. A 25 cm3 sample of this solution was titrated against 0.2 moldm-3 HCl and 10.8 cm3 of the acid was required to neutralise the sample. |