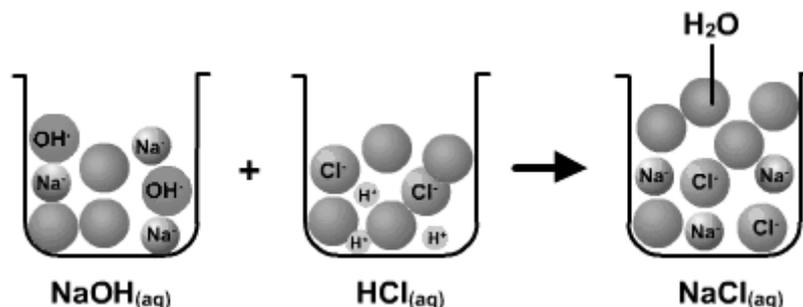


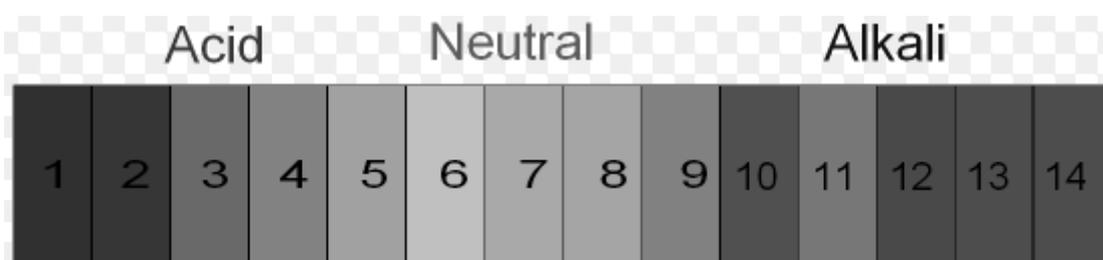
# UNIT 5

## ACIDS, BASES AND SALTS

### Teacher Version



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### Contents

- Acids, Bases, Salts and Neutralisation
- Physical Properties of Acids, Bases and Salts
- Strong and Weak Acids and Bases
- Further Reactions of Acids, Bases and Salts
- Laboratory Preparation of Salts by Neutralisation
- Qualitative Analysis 1a: Acid-Base Reactions
- Indicators and the pH scale
- Qualitative Analysis 1b: Indicators
- 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 – ACIDS, BASES AND SALTS

UNIT 5 SUMMARY AND SYLLABUS REFERENCE

Lesson	Title and Syllabus Reference
1	<p><b>Introduction to Acids, Bases and Salts</b>  <b>CA7a definitions of acids and bases</b> (Arrhenius concept of acids and bases, effects of acids on trioxocarbonate (IV) salts); <b>CA7b chemical properties of acids and bases</b> (balanced chemical equations of all reactions); <b>ISA7.1 simple definition of acids, bases, salts</b> (definition of acids and bases in terms of proton transfer (Bronsted-Lowry concept)); <b>ISA7.2 chemical properties of acids and bases</b> (properties of acids and bases); <b>ISA7.3 examples of chemical substances classified as acids, bases or salts, ISA13.2 neutralisation</b> (equations representing neutralization reactions)</p>
2	<p><b>Physical Properties of Acids, Bases and Salts</b>  <b>CA7b physical and chemical properties of acids and bases</b> (conductivities, taste etc); <b>CA7c acids, bases and salts as electrolytes</b> (electrolytes and non-electrolytes); <b>CA8b practical application of solubility</b> (generalisations of solubility of salts) <b>CC4 deliquescent, efflorescent and hygroscopic substances</b> (the use of deliquescent and hygroscopic substances as drying agents should be emphasized); <b>ISA7.2 physical properties of acids, bases, salts</b> (properties and uses of acids, bases and salts)</p>
3	<p><b>Strong and Weak Acids and Bases</b>  <b>CA7c acids, bases and salts as electrolytes</b> (strong and weak electrolytes, evidence from conductivity and enthalpy of neutralisation); <b>CA7e weak acids and weak bases</b> (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); <b>ISA7.2 chemical properties of acids and bases</b> (properties of acids and bases)</p>
4	<p><b>Further Chemical Properties of Acids, Bases and Salts</b>  <b>CA7b chemical properties of acids and bases</b> (concept of amphoterism); <b>CA7f hydrolysis</b> (qualitative explanation of hydrolysis, behaviour of some salts (e.g. <math>\text{NH}_4\text{Cl}</math>, <math>\text{AlCl}_3</math>, <math>\text{Na}_2\text{CO}_3</math>, <math>\text{CH}_3\text{COONa}</math>) in water as examples of equilibrium systems)); <b>CC7aiii binary compounds of oxygen: acidic oxides, basic oxides, amphoteric oxides</b>; <b>ISA7.2 chemical properties of acids and bases</b> (properties of acids, bases and salts)</p>
5	<p><b>Preparation of Salts by Neutralisation</b>  <b>CA13aiii General Skills and Principles – filtration</b>; <b>CC4a methods of preparation of simple salts</b>; <b>ISA7.4 methods of preparation of salts</b> (preparation of salts using the following methods: neutralization, acid + salt)</p>
6	<p><b>Qualitative Analysis – Acid-Base Reactions</b>  <b>CA5bii gases – properties of gases</b> (chemical properties of <math>\text{NH}_3</math>, <math>\text{CO}_2</math>, <math>\text{HCl}</math> and <math>\text{SO}_2</math>); <b>CA7a definitions of acids and bases</b> (effects of acids on trioxocarbonate (IV) salts); <b>CA13ci Qualitative Analysis - characteristic test of <math>\text{NH}_4^+</math> with dilute NaOH; confirmatory test for <math>\text{NH}_4^+</math></b> (no formal scheme of analysis is required); <b>CA13cii Qualitative Analysis - characteristic reaction of dilute HCl on solids or aqueous solutions and conc. <math>\text{H}_2\text{SO}_4</math> on solid samples of the following: <math>\text{Cl}^-</math>, <math>\text{SO}_3^{2-}</math>, <math>\text{CO}_3^{2-}</math>, confirmatory tests for <math>\text{SO}_3^{2-}</math>, <math>\text{CO}_3^{2-}</math></b> (no formal scheme of analysis is required); <b>CA13civ Qualitative Analysis - characteristic test for the following gases: <math>\text{NH}_3</math>, <math>\text{CO}_2</math>, <math>\text{HCl}</math> and <math>\text{SO}_2</math></b> (no formal scheme of analysis is required); <b>ISA7.2 chemical properties of acids and bases</b> (test for carbon dioxide and ammonia gases), <b>ISA7.3 examples of chemical substances classified as acids, bases or salts</b> (simple chemical tests to classify chemical substances as acids, bases, or salts)</p>
7	<p><b>The pH scale and simple indicators</b>  <b>CA7a definitions of acids and bases</b> (effects of acids and bases on indicators); <b>CA7d pH</b> (knowledge of pH scale, pH as a measure of acidity and alkalinity); <b>CA7g acid-base indicators</b> (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), <b>ISA7.5 acid-base indicators</b> (description of the colours developed by phenolphthalein, litmus and methyl orange in dilute acids and dilute bases)</p>

**UNIT 5 – ACIDS, BASES AND SALTS**

Lesson	Title and Syllabus Reference
8	<p><b>Universal Indicator and pH Colorimetry</b>  <b>CA7a definitions of acids and bases</b> (effects of acids and bases on indicators); <b>CA13av General Skills and Principles - determination of pH value of various solutions by colorimetry</b>; <b>ISA7.3 examples of chemical substances classified as acids, bases or salts</b> (simple chemical tests to classify chemical substances as acids, bases, or salts); <b>ISA7.6 determination of the pH of given solutions</b> (the nature and use of the universal indicator and pH meter, determination of soil pH is required)</p>
9	<p><b>Qualitative Analysis – Indicators</b>  <b>CA13ci Qualitative Analysis - confirmatory test for NH<sub>4</sub><sup>+</sup> with dilute NaOH; confirmatory tests for the anions and cations CA13ci Qualitative Analysis - characteristic test for the following gases: NH<sub>3</sub>; CO<sub>2</sub>; HCl and SO<sub>2</sub></b> (no formal scheme of analysis is required); <b>ISA7.2 chemical properties of acids and bases</b> (test for ammonia gas); <b>ISA7.3 examples of chemical substances classified as acids, bases or salts</b> (simple chemical tests to classify chemical substances as acids, bases, or salts)</p>
10	<p><b>Acid-Base Titrations I</b>  <b>CA7h acid-base titrations</b> (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); <b>CA13ai General Skills and Principles - measurement of volume</b>; <b>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</b></p>
11	<p><b>Acid-Base Titrations II</b>  <b>CA4cii standard solutions</b> (preparation of some primary standard solutions using Na<sub>2</sub>CO<sub>3</sub>, (COOH)<sub>2</sub>.2H<sub>2</sub>O); <b>CA7h acid-base titrations</b> (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); <b>CA13ai General Skills and Principles - measurement of mass and volume</b>; <b>CA13aii General Skills and Principles - preparation and dilution of standard solutions</b>; <b>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</b>; <b>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</b>; <b>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</b></p>
12	<p><b>Unit 5 Revision and Summary</b></p>

UNIT 5 – ACIDS, BASES AND SALTS

The Periodic Table of the Elements

	1	2	3	4	5	6	7	0	
(1)	6.9 <b>Li</b> lithium 3	9.0 <b>Be</b> beryllium 4	(13) 10.8 <b>B</b> boron 5	(14) 12.0 <b>C</b> carbon 6	(15) 14.0 <b>N</b> nitrogen 7	(16) 16.0 <b>O</b> oxygen 8	(17) 19.0 <b>F</b> fluorine 9	(18) 20.2 <b>Ne</b> neon 10	
	23.0 <b>Na</b> sodium 11	24.3 <b>Mg</b> magnesium 12	(13) 27.0 <b>Al</b> aluminium 13	(14) 28.1 <b>Si</b> silicon 14	(15) 31.0 <b>P</b> phosphorus 15	(16) 32.1 <b>S</b> sulfur 16	(17) 35.5 <b>Cl</b> chlorine 17	(18) 39.9 <b>Ar</b> argon 18	
	39.1 <b>K</b> potassium 19	40.1 <b>Ca</b> calcium 20	(12) 69.7 <b>Ga</b> gallium 31	(14) 72.6 <b>Ge</b> germanium 32	(15) 74.9 <b>As</b> arsenic 33	(16) 79.0 <b>Se</b> selenium 34	(17) 79.9 <b>Br</b> bromine 35	(18) 83.8 <b>Kr</b> krypton 36	
	85.5 <b>Rb</b> rubidium 37	87.6 <b>Sr</b> strontium 38	(12) 112.4 <b>Cd</b> cadmium 48	(14) 114.8 <b>In</b> indium 49	(15) 121.8 <b>Sb</b> antimony 51	(16) 127.6 <b>Te</b> tellurium 52	(17) 126.9 <b>I</b> iodine 53	(18) 131.3 <b>Xe</b> xenon 54	
	132.9 <b>Cs</b> caesium 55	137.3 <b>Ba</b> barium 56	(12) 200.6 <b>Hg</b> mercury 80	(14) 204.4 <b>Tl</b> thallium 81	(15) 209.0 <b>Po</b> polonium 84	(16) 207.2 <b>Pb</b> lead 82	(17) [210] <b>At</b> astatine 85	(18) [222] <b>Rn</b> radon 86	
	[223] <b>Fr</b> francium 87	[226] <b>Ra</b> radium 88	Elements with atomic numbers 112-116 have been reported but not fully authenticated						
			(11) 63.5 <b>Cu</b> copper 29	(10) 58.7 <b>Ni</b> nickel 28	(9) 58.9 <b>Co</b> cobalt 27	(8) 55.8 <b>Fe</b> iron 26	(7) 54.9 <b>Mn</b> manganese 25	(6) 52.0 <b>Cr</b> chromium 24	
			(11) 107.9 <b>Ag</b> silver 47	(10) 106.4 <b>Pd</b> palladium 46	(9) 102.9 <b>Rh</b> rhodium 45	(8) 101.1 <b>Ru</b> ruthenium 44	(7) [98] <b>Tc</b> technetium 43	(6) 96.0 <b>Mo</b> molybdenum 42	
			(11) 197.0 <b>Au</b> gold 79	(10) 195.1 <b>Pt</b> platinum 78	(9) 192.2 <b>Ir</b> iridium 77	(8) 190.2 <b>Os</b> osmium 76	(7) 186.2 <b>Re</b> rhenium 75	(6) 183.8 <b>W</b> tungsten 74	
			(11) [280] <b>Rg</b> roentgenium 111	(10) [281] <b>Ds</b> darmstadtium 110	(9) [276] <b>Mt</b> meitnerium 109	(8) [270] <b>Hs</b> hassium 108	(7) [272] <b>Bh</b> bohrium 107	(6) [271] <b>Sg</b> seaborgium 106	
			(11) 157.3 <b>Gd</b> gadolinium 64	(10) 152.0 <b>Eu</b> europium 63	(9) 150.4 <b>Sm</b> samarium 62	(8) [145] <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 162.5 <b>Dy</b> dysprosium 66	(10) 164.9 <b>Ho</b> holmium 67	(9) 167.3 <b>Er</b> erbium 68	(8) 168.9 <b>Tm</b> thulium 69	(7) 173.1 <b>Yb</b> ytterbium 70	(6) 175.0 <b>Lu</b> lutetium 71	
			(11) [251] <b>Cf</b> californium 98	(10) [252] <b>Es</b> einsteinium 99	(9) [257] <b>Fm</b> fermium 100	(8) [258] <b>Md</b> mendelevium 101	(7) [259] <b>No</b> nobelium 102	(6) [262] <b>Lr</b> lawrencium 103	
			(11) [247] <b>Bk</b> berkelium 97	(10) [247] <b>Cm</b> curium 96	(9) [244] <b>Pu</b> plutonium 94	(8) [237] <b>Np</b> neptunium 93	(7) 238.0 <b>U</b> uranium 92	(6) 231.0 <b>Pa</b> protactinium 91	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
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			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
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			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
			(11) 140.1 <b>Ce</b> cerium 58	(10) 140.9 <b>Pr</b> praseodymium 59	(9) 144.2 <b>Nd</b> neodymium 60	(8) 145 <b>Pm</b> promethium 61	(7) 144.2 <b>Nd</b> neodymium 60	(6) 140.9 <b>Pr</b> praseodymium 59	
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## Lesson 1 – What are acids, bases and salts?

**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?

Note: students should be encouraged to identify some common acids and alkalis and give some of their features

- acids: hydrochloric acid, sulphuric acid, nitric acid, lactic acid, citric acid
- vinegar (ethanoic acid), orange/lemon juice (citric acid); many fizzy sweet drinks (coke, sprite) also contain acids
- acids taste sour and bitter and they can sting if in contact with eyes or broken skin
- Alkalis: sodium hydroxide, calcium hydroxide, ammonia, soaps, detergents, bleach
- foods tend not to be strongly alkaline as they would be harmful
- alkalis are generally not good to eat; they feel and taste soapy

## a) 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:  
 $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$ ; the resulting solution is known as **hydrochloric acid**
  - hydrogen nitrate ( $HNO_3$ ) dissolves in water to produce hydrogen ( $H^+$ ) and nitrate ( $NO_3^-$ ) ions:  
 $HNO_3(l) \rightarrow H^+(aq) + NO_3^-(aq)$ ; the resulting solution is known as **nitric acid**
  - hydrogen sulphate ( $H_2SO_4$ ) dissolves in water to produce hydrogen ( $H^+$ ) and nitrate ( $SO_4^{2-}$ ) ions:  
 $H_2SO_4(l) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$ ; 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,  $HNO_3$  and  $H_2SO_4$  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)

## UNIT 5 – ACIDS, BASES AND SALTS

### (ii) Bases and Alkalis

- The **Arrhenius** definition of a base is a substance which releases hydroxide ( $\text{OH}^-$ ) ions when dissolved in water; common examples are:
  - some **metal oxides** such as  $\text{CaO}$  which react with water to produce hydroxide ( $\text{OH}^-$ ) ions:  
 $\text{CaO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$
  - some **metal hydroxides** such as  $\text{NaOH}$  which dissolve in water to produce hydroxide ( $\text{OH}^-$ ) ions:  $\text{NaOH(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
  - some **metal carbonates** such as  $\text{Na}_2\text{CO}_3$  which react with water to produce hydroxide ( $\text{OH}^-$ ) ions:  
 $\text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O(l)} \rightarrow 2\text{Na}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
  - **ammonia** ( $\text{NH}_3$ ), which reacts with water to produce hydroxide ( $\text{OH}^-$ ) ions:  
 $\text{NH}_3(\text{g}) + \text{H}_2\text{O(l)} \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
- The **Bronsted-Lowry** definition of a base is a proton ( $\text{H}^+$ ) acceptor; the base accepts  $\text{H}^+$  ions (protons) from other species in the solution (which are therefore behaving as acids); the four most common types of base are:
  - $\text{OH}^-$  (hydroxide) ions, which accept  $\text{H}^+$  ions to form water:  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
  - $\text{CO}_3^{2-}$  (carbonate) ions, which accept  $\text{H}^+$  ions to form carbon dioxide and water:  $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
  - $\text{O}^{2-}$  (oxide) ions, which accept  $\text{H}^+$  ions to form water:  $\text{H}^+ + \text{O}^{2-} \rightarrow \text{H}_2\text{O}$
  - $\text{NH}_3$ , which accepts  $\text{H}^+$  ions to become  $\text{NH}_4^+$  (ammonium):  $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$
- 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  $\text{OH}^-$  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  $\text{H}^+$  ion; the  $\text{H}^+$  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  $\text{H}^+$  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  $\rightarrow$  metal salt + water  
Eg  $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$  (the salt produced is sodium chloride)
  - from metal oxides: acid + metal oxide  $\rightarrow$  metal salt + water  
Eg  $2\text{HNO}_3 + \text{CaO} \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$  (the salt produced is calcium nitrate)
  - from metal carbonates: acid + metal carbonate  $\rightarrow$  metal salt + carbon dioxide + water  
Eg  $\text{H}_2\text{SO}_4 + \text{K}_2\text{CO}_3 \rightarrow \text{K}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$  (the salt produced is potassium sulphate)
  - from ammonia: acid + ammonia  $\rightarrow$  ammonium salt  
Eg  $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$  (the salt produced is ammonium chloride)

## UNIT 5 – ACIDS, BASES AND SALTS



### Test 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:

- |  |  |
|--|--|
| <ul style="list-style-type: none"> <li>a) Sodium oxide</li> <li>b) Calcium hydroxide</li> <li>c) Ammonium nitrate</li> <li>d) Potassium carbonate</li> <li>e) Strontium sulphate</li> <li>f) Ammonium sulphate</li> <li>g) Hydrochloric acid</li> <li>h) Rubidium hydroxide</li> <li>i) Magnesium carbonate</li> </ul> | <ul style="list-style-type: none"> <li>j) Calcium nitrate</li> <li>k) Sulphuric acid</li> <li>l) Ammonium chloride</li> <li>m) Nitric acid</li> <li>n) Potassium sulphate</li> <li>o) Magnesium oxide</li> <li>p) Caesium bromide</li> <li>q) Barium sulphate</li> <li>r) Strontium nitrate</li> </ul> |
|--|--|

- |   |  |
|---|--|
| <ul style="list-style-type: none"> <li>a) Na<sub>2</sub>O (base)</li> <li>b) Ca(OH)<sub>2</sub> (base)</li> <li>c) NH<sub>4</sub>NO<sub>3</sub> (salt)</li> <li>d) K<sub>2</sub>CO<sub>3</sub> (base)</li> <li>e) SrSO<sub>4</sub> (salt)</li> <li>f) (NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub> (salt)</li> <li>g) HCl (acid)</li> <li>h) RbOH (base)</li> <li>i) MgCO<sub>3</sub> (base)</li> </ul> | <ul style="list-style-type: none"> <li>j) Ca(NO<sub>3</sub>)<sub>2</sub> (salt)</li> <li>k) H<sub>2</sub>SO<sub>4</sub> (acid)</li> <li>l) NH<sub>4</sub>Cl (salt)</li> <li>m) HNO<sub>3</sub> (acid)</li> <li>n) K<sub>2</sub>SO<sub>4</sub> (salt)</li> <li>o) MgO (base)</li> <li>p) CsBr (salt)</li> <li>q) BaSO<sub>4</sub> (salt)</li> <li>r) Sr(NO<sub>3</sub>)<sub>2</sub> (salt)</li> </ul> |
|---|--|



### Test your knowledge 1.3: Understanding Neutralisation Reactions

Write balanced symbol equations for the following neutralisation reactions:

- |  |  |
|--|--|
| <ul style="list-style-type: none"> <li>a) Nitric acid with potassium hydroxide</li> <li>b) Sulphuric acid with sodium hydroxide</li> <li>c) Hydrochloric acid with calcium hydroxide</li> <li>d) Nitric acid with calcium oxide</li> <li>e) Hydrochloric acid with barium oxide</li> <li>f) Sulphuric acid with magnesium oxide</li> </ul> | <ul style="list-style-type: none"> <li>g) Nitric acid with calcium carbonate</li> <li>h) Hydrochloric acid with barium carbonate</li> <li>i) Sulphuric acid with sodium carbonate</li> <li>j) Nitric acid with ammonia</li> <li>k) Sulphuric acid with ammonia</li> <li>l) Hydrochloric acid with ammonia</li> </ul> |
|--|--|
- 
- |   |   |
|---|---|
| <ul style="list-style-type: none"> <li>a) <math>\text{HNO}_3 + \text{KOH} \rightarrow \text{KNO}_3 + \text{H}_2\text{O}</math></li> <li>b) <math>\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}</math></li> <li>c) <math>2\text{HCl} + \text{Ca}(\text{OH})_2 \rightarrow \text{CaCl}_2 + 2\text{H}_2\text{O}</math></li> <li>d) <math>2\text{HNO}_3 + \text{CaO} \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}</math></li> <li>e) <math>2\text{HCl} + \text{BaO} \rightarrow \text{BaCl}_2 + \text{H}_2\text{O}</math></li> <li>f) <math>\text{H}_2\text{SO}_4 + \text{MgO} \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}</math></li> </ul> | <ul style="list-style-type: none"> <li>g) <math>\text{HNO}_3 + \text{CaCO}_3 \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O}</math></li> <li>h) <math>2\text{HCl} + \text{BaCO}_3 \rightarrow \text{BaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}</math></li> <li>i) <math>\text{H}_2\text{SO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \text{Na}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}</math></li> <li>j) <math>\text{HNO}_3 + \text{NH}_3 \rightarrow \text{NH}_4\text{NO}_3</math></li> <li>k) <math>\text{H}_2\text{SO}_4 + 2\text{NH}_3 \rightarrow (\text{NH}_4)_2\text{SO}_4</math></li> <li>l) <math>\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}</math></li> </ul> |
|---|---|

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

**b) 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)<sub>2</sub> and Ba(OH)<sub>2</sub>), which are highly soluble in water, and calcium hydroxide (Ca(OH)<sub>2</sub>), 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 (Li<sub>2</sub>O, Na<sub>2</sub>O, K<sub>2</sub>O 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  $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
  - **metal carbonates** are usually insoluble in water; they are therefore bases but not alkalis the exceptions are the Group I carbonates (Li<sub>2</sub>CO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>, K<sub>2</sub>CO<sub>3</sub> 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 PbCl<sub>2</sub>)
  - 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

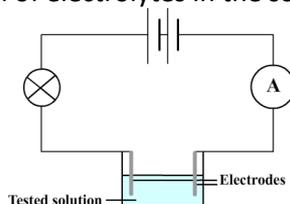
The solubility of salts will be studied in more detail later in the course

**(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<sup>+</sup>, the ionic compound is an **acid**
  - If the anion is O<sup>2-</sup>, OH<sup>-</sup>, CO<sub>3</sub><sup>2-</sup> or HCO<sub>3</sub><sup>-</sup>, 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)

## UNIT 5 – ACIDS, BASES AND SALTS

- 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 **deliquescent** 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:

Examples of hydrated solids	Unit formula
Hydrated copper sulphate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Hydrated sodium carbonate	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Hydrated ethanedioic acid	$\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$

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,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ , is  $106 + 10 \times 18 = 286$

- 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

## UNIT 5 – ACIDS, BASES AND SALTS

- 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 ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ); when plaster is left to dry, it loses its water and hardens



### Test your knowledge 2.1: Describing properties of acids, bases and salts

- Name two substances which are soluble bases (alkalis)
- Name two substances which are insoluble bases
- Name two soluble salts
- Name two insoluble salts
- Explain why soluble acids, alkalis and soluble salts are all electrolytes
- Describe the typical taste of acids, bases and salts and give the ion responsible for the taste
- Explain what is meant by the term hygroscopic and name one hygroscopic acid, one hygroscopic base and one hygroscopic salt
- Explain what is meant by the term deliquescent and name one deliquescent salt
- What is meant by the term “hydrated salt”?
- A sample of hydrated calcium nitrate is found to have a molar mass of 236.1. Deduce its dot formula.
- What is effluorescence? Give an example of an effluorescent salt.

- eg potassium hydroxide, sodium carbonate, ammonia etc
- eg magnesium oxide, calcium carbonate, aluminium hydroxide etc
- eg sodium chloride, magnesium nitrate, ammonium sulphate etc
- eg silver chloride, lead chloride, barium sulphate etc
- they contain ions and so conduct electricity when in solution
- acids are sour due to  $\text{H}^+$ ; bases are soapy due to  $\text{OH}^-$ ; salts are salty, often due to  $\text{Na}^+$
- absorbs water from the atmosphere; eg concentrated sulphuric acid, solid sodium hydroxide, solid calcium chloride
- absorbs water from the atmosphere and then dissolves in the water it has absorbed; eg calcium chloride
- a salt which has water locked into its crystal structure
- mr of  $\text{Ca}(\text{NO}_3)_2 = 164.1$ ;  $x\text{H}_2\text{O} = 236.1 - 164.1 = 72$ , so  $x = 72/18 = 4$ , so  $\text{Ca}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$
- the loss of water from a crystal structure; eg  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

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

## c) Strong and Weak Acids and Bases

### (i) strong and weak acids

- Acids react with water to form solutions containing  $\text{H}^+$  ions:

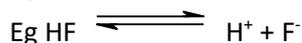


In hydrochloric acid and nitric acid, the molecule dissociates completely into the  $\text{H}^+$  ion and the anion; there is no molecular HCl or  $\text{HNO}_3$  remaining

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  $\text{H}^+$  ions**; HCl,  $\text{HNO}_3$  and  $\text{H}_2\text{SO}_4$  are strong acids

## UNIT 5 – ACIDS, BASES AND SALTS

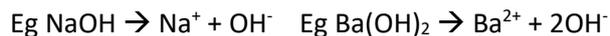
- Most acids, however, only dissociate slightly in water; the  $\text{H}^+$  ions, anions and undissociated molecules form a dynamic equilibrium:



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  $\text{H}^+$  ions**; all carboxylic acids, such as methanoic acid and ethanoic acid, are weak acids

### (ii) Strong and weak bases

- Bases react with water to form solutions containing  $\text{OH}^-$  ions:



In sodium hydroxide and barium hydroxide, the solid dissolves completely into the  $\text{OH}^-$  ion and the cation; there is no undissolved or undissociated  $\text{NaOH}$  or  $\text{Ba}(\text{OH})_2$ ; a base which completely dissociates in water to form  $\text{OH}^-$  ions is called a **strong base**; a **strong base is a substance which completely dissociates in water to form  $\text{OH}^-$  ions**;  $\text{NaOH}$ ,  $\text{KOH}$  and  $\text{BaO}$  are strong bases

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



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  $\text{OH}^-$  ions**; ammonia and sodium carbonate are weak bases.

- There are also some bases, such as  $\text{Ca}(\text{OH})_2$ , which are only slightly soluble in water but which fully dissociate into  $\text{OH}^-$  ions once they are dissolved; these are also classified as weak bases because they do not produce solutions containing a lot of  $\text{OH}^-$  ions due to their low solubility

### (iii) Properties of strong acids and bases compared to weak acids and bases

#### (α) Electrical conductivity:

- Strong acids and bases are completely dissociated into their ions:
  - a  $1 \text{ mol dm}^{-3}$  solution of  $\text{HCl}$  contains  $1 \text{ mol dm}^{-3} \text{ H}^+$  and  $1 \text{ mol dm}^{-3} \text{ Cl}^-$ ; it does not contain any undissociated  $\text{HCl}$
  - a  $1 \text{ mol dm}^{-3}$  solution of  $\text{NaOH}$  contains  $1 \text{ mol dm}^{-3} \text{ OH}^-$  and  $1 \text{ mol dm}^{-3} \text{ Na}^+$  it does not contain any undissociated  $\text{NaOH}$
  - 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:
  - $1 \text{ mol dm}^{-3}$  solution of  $\text{CH}_3\text{COOH}$  contains mostly undissociated  $\text{CH}_3\text{COOH}$ ; it contains a very small quantity of  $\text{CH}_3\text{COO}^-$  and  $\text{H}^+$  ions (around  $0.01 \text{ mol dm}^{-3}$ )
  - a  $1 \text{ mol dm}^{-3}$  solution of  $\text{NH}_3$  contains mostly undissociated  $\text{NH}_3$ . It contains a very small quantity of  $\text{NH}_4^+$  and  $\text{OH}^-$  ions (around  $0.01 \text{ mol dm}^{-3}$ )

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**

## UNIT 5 – ACIDS, BASES AND SALTS

### (β) Enthalpy of neutralisation:

- Solutions of strong acids contain fully dissociated  $\text{H}^+$  ions; solutions of strong bases contain fully dissociated  $\text{OH}^-$  ions:
  - when a solution of a strong acid reacts with a solution of a strong base, the  $\text{H}^+$  ions react with the  $\text{OH}^-$  ions to make  $\text{H}_2\text{O}$ :  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
  - 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 ( $-58 \text{ kJmol}^{-1}$ )
  - eg  $\text{HNO}_3 + \text{NaOH} \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}$  ( $\Delta H = -58 \text{ kJmol}^{-1}$ );  $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$  ( $\Delta H = -58 \text{ kJmol}^{-1}$ )
- Solutions of weak acids only contain small quantities of  $\text{H}^+$  ions
  - when weak acids are neutralised by strong bases, the  $\text{OH}^-$  ions react with the  $\text{H}^+$  ions which quickly run out
  - as a result, the weak acid dissociates more, producing more  $\text{H}^+$  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  $\text{H}^+$ :  $\text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COO}^- + \text{H}^+$
    - Step 2: the resulting  $\text{H}^+$  reacts with the  $\text{OH}^-$ :  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
  - the first reaction, the dissociation, is endothermic by around  $2 \text{ kJmol}^{-1}$ ; the second reaction, the neutralisation, has a molar enthalpy change of  $-58 \text{ kJmol}^{-1}$ ; the overall molar enthalpy change of neutralisation is therefore  $-56 \text{ kJmol}^{-1}$
  - eg  $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$   $\Delta H = -56 \text{ kJmol}^{-1}$
  - weak acids have a less exothermic enthalpy of neutralisation than strong acids, because the  $\text{H}^+$  ion has to be removed from the acid first, which absorbs some of the energy
- Solutions of weak bases only contain small quantities of  $\text{OH}^-$  ions:
  - When weak bases are neutralised by strong acids, the  $\text{H}^+$  ions react with the  $\text{OH}^-$  ions which quickly run out
  - as a result, the weak base dissociates more, producing more  $\text{OH}^-$  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  $\text{OH}^-$ :  $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$
    - Step 2: the resulting  $\text{OH}^-$  reacts with the  $\text{H}^+$ :  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
  - the first reaction, the dissociation, is endothermic by around  $5 \text{ kJmol}^{-1}$ ; the second reaction, the neutralisation, has a molar enthalpy change of  $-58 \text{ kJmol}^{-1}$ ; the overall molar enthalpy change of neutralisation is therefore  $-53 \text{ kJmol}^{-1}$
  - eg  $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$   $\Delta H = -53 \text{ kJmol}^{-1}$
  - Weak bases have a less exothermic enthalpy of neutralisation than strong bases, because the  $\text{OH}^-$  ion has to be produced from the base first, which absorbs some of the energy



**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 cm<sup>3</sup> of 2 mol dm<sup>-3</sup> HCl into the cup using a measuring cylinder. Record the initial temperature of the solution.
- 3) Measure out 25 cm<sup>3</sup> of 2 mol dm<sup>-3</sup> 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 cm<sup>3</sup> of 2 mol dm<sup>-3</sup> NH<sub>3</sub> instead of NaOH.
- 7) Which reaction gives you a greater temperature change? Can you explain why?

**Equipment needed per group: two measuring cylinders (25 cm<sup>3</sup>), polystyrene cup (250 cm<sup>3</sup>), beaker (250 cm<sup>3</sup>), thermometer, access to 2.0 mol dm<sup>-3</sup> HCl (50 cm<sup>3</sup> per group), 2 mol dm<sup>-3</sup> NH<sub>3</sub> (25 cm<sup>3</sup> each per group) and 2 mol dm<sup>-3</sup> NaOH (25 cm<sup>3</sup> each per group)**

**Caution:** the alkalis are corrosive at this molarity and should be handled with great care

- **The temperature increase should be around 7 °C with NaOH and around 5 °C with NH<sub>3</sub>**
- **The temperature increase is higher with NaOH because the OH<sup>-</sup> ions are already in the solution; NH<sub>3</sub> dissociates during the neutralisation to give OH<sup>-</sup> ions; this process is endothermic so the overall neutralisation process is less exothermic**



**Test your knowledge 3.2: Distinguishing between Strong and Weak Acids and Bases**

- (a) What is the difference between a strong acid and a weak acid? Give one example of each.
- (b) 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
- (c) Why do strong acids conduct electricity better than weak acids?
- (d) Why is the enthalpy of neutralisation of weak acids less exothermic than the enthalpy of neutralisation of strong acids?

- (a) Strong acid fully dissociates in water to give H<sup>+</sup> ions (eg H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub>, HCl); weak acid only partially dissociates in water to give H<sup>+</sup> ions (eg ethanoic acid etc)
- (b) Strong base: NaOH; weak base due to low solubility: Ca(OH)<sub>2</sub>, weak base due to limited dissociation: NH<sub>3</sub>
- (c) Strong acids are fully dissociated into ions, so the concentration of ions is greater, and the conductivity is due to the presence of ions
- (d) Weak acids must dissociate into ions during neutralisation; this process is endothermic, so the overall reaction is less exothermic

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

## d) 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, HNO<sub>3</sub> and H<sub>2</sub>SO<sub>4</sub>), which contain H<sup>+</sup> 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<sup>+</sup> ions; some do not dissolve in water but react with alkalis to form salts; examples of acidic oxides are CO<sub>2</sub>, NO<sub>2</sub>, P<sub>4</sub>O<sub>6</sub>, P<sub>4</sub>O<sub>10</sub>, SO<sub>2</sub>, SO<sub>3</sub>, Cl<sub>2</sub>O
  - Eg SO<sub>3</sub> + H<sub>2</sub>O → H<sub>2</sub>SO<sub>4</sub> and SO<sub>3</sub> + 2NaOH → Na<sub>2</sub>SO<sub>4</sub> + H<sub>2</sub>O (salt = sodium sulphate)
  - Eg SiO<sub>2</sub> + 2NaOH → Na<sub>2</sub>SiO<sub>3</sub> + H<sub>2</sub>O (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<sup>-</sup> ions; some do not dissolve in water but react with acids to form salts:
  - Eg Na<sub>2</sub>O + H<sub>2</sub>O → 2NaOH and Na<sub>2</sub>O + HCl → 2NaCl + H<sub>2</sub>O
  - Eg Mg(OH)<sub>2</sub> + 2HCl → MgCl<sub>2</sub> + 2H<sub>2</sub>O
- 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)<sub>3</sub> + 3HCl → AlCl<sub>3</sub> + 3H<sub>2</sub>O and Al(OH)<sub>3</sub> + NaOH → NaAl(OH)<sub>4</sub>
  - Eg ZnO + 2HCl → ZnCl<sub>2</sub> + H<sub>2</sub>O and ZnO + 2NaOH → Na<sub>2</sub>ZnO<sub>2</sub> + H<sub>2</sub>O
 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 (HCO<sub>3</sub><sup>-</sup>)
  - NaHCO<sub>3</sub> + HCl → NaCl + CO<sub>2</sub> + H<sub>2</sub>O and NaHCO<sub>3</sub> + NaOH → Na<sub>2</sub>CO<sub>3</sub> + H<sub>2</sub>O

## (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);
 
$$\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$$
- 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:
 
$$\text{H}_3\text{PO}_4(\text{l}) + 3\text{KCl}(\text{s}) \rightarrow \text{K}_3\text{PO}_4(\text{s}) + 3\text{HCl}(\text{g})$$
- 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

## UNIT 5 – ACIDS, BASES AND SALTS

### (iii) Salt Hydrolysis

- Salts can be classified into three main types:
  - salts formed from strong acids and strong bases (eg  $\text{NaNO}_3$ ,  $\text{K}_2\text{SO}_4$  and  $\text{BaCl}_2$ )
  - salts formed from weak acids and strong bases (eg  $\text{CH}_3\text{COONa}$ ,  $\text{KCN}$ )
  - salts formed from strong acids and weak bases (eg  $\text{NH}_4\text{Cl}$ ,  $\text{NH}_4\text{NO}_3$ )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  $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$ 
  - the resulting solution, therefore, contains  $\text{OH}^-$  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  $\text{H}^+$  in the first place, it is more likely that the anion will pick up an  $\text{H}^+$  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  $\text{CH}_3\text{COONa}$  and  $\text{KCN}$ ; 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  $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$ 
  - the resulting solution, therefore, contains  $\text{H}^+$  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  $\text{H}^+$  ion in the first place, its cation is more likely release an  $\text{H}^+$  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  $\text{NH}_4\text{Cl}$  and  $\text{NH}_4\text{NO}_3$ ; they are made from strong acids and weak bases
- Some highly charged cations also react with water to form  $\text{H}^+$  ions, in order to reduce their charge; they are accepting  $\text{OH}^-$  ions from the water and partly reforming the original hydroxide:
  - eg  $\text{Al}^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{Al}(\text{OH})^{2+} + \text{H}^+$
  - the resulting solution contains  $\text{H}^+$  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  $\text{AlCl}_3$  or  $\text{CrCl}_3$
- 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**


**Test your knowledge 4.1: Understanding Further Reactions of Acids, Bases and Salts**

- (a) What is meant by the term “amphoteric”; give two examples of amphoteric substances
- (b) Write equations for the following reactions:
- sulphuric acid and potassium chloride
  - nitric acid and sodium fluoride
  - phosphoric acid ( $\text{H}_3\text{PO}_4$ ) and potassium bromide
- (c) 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  $\text{H}^+$  or  $\text{OH}^-$  ions
- sodium nitrate
  - ammonium nitrate
  - sodium ethanoate
  - potassium sulphate
  - aluminium sulphate
  - potassium cyanide

- (a) reacts with acids and bases/can behave as an acid or a base, eg  $\text{NaHCO}_3$  or  $\text{Al}(\text{OH})_3$  or  $\text{ZnO}$  etc
- (b) (i)  $\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$ ; (ii)  $\text{HNO}_3 + \text{NaF} \rightarrow \text{NaNO}_3 + \text{HF}$ ; (iii)  $\text{H}_3\text{PO}_4 + 3\text{KBr} \rightarrow \text{K}_3\text{PO}_4 + 3\text{HBr}$
- (c) (i) no; (ii) yes  $\text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}^+$ ; (iii) yes  $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{COOH} + \text{OH}^-$ ; (iv) no; (v) yes  $\text{Al}^{3+} + \text{H}_2\text{O} \rightarrow \text{Al}(\text{OH})^{2+} + \text{H}^+$ ; (vi) yes  $\text{CN}^- + \text{H}_2\text{O} \rightarrow \text{HCN} + \text{OH}^-$

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

## e) 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



**Practical 5.1: Prepare a salt by neutralising an acid with an insoluble base**

- 1) Measure out 20 cm<sup>3</sup> of 0.5 moldm<sup>-3</sup> sulphuric acid into a 100 cm<sup>3</sup> beaker.
- 2) Use a Bunsen burner to warm the beaker gently on a tripod until the temperature reaches 50 °C.
- 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 cm<sup>3</sup> 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 (H<sub>2</sub>SO<sub>4</sub>) 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?

**Equipment needed per group: measuring cylinder (25 cm<sup>3</sup>), beaker (250 cm<sup>3</sup>), Bunsen burner, tripod, gauze, thermometer, weighing boat, spatula, stirrer, filter paper, funnel, conical flask (100 cm<sup>3</sup>), access to a mass balance, access to 0.5 moldm<sup>-3</sup> H<sub>2</sub>SO<sub>4</sub> (20 cm<sup>3</sup> per group), access to CuO (1 g per group)**

-  $0.02 \times 0.5 = 0.01$

-  $1/79.5 = 0.0126$

- **To ensure the H<sub>2</sub>SO<sub>4</sub> fully reacts; excess CuO can be removed by filtration; it would be much more difficult to remove excess H<sub>2</sub>SO<sub>4</sub>**

Cannot do this experiment? Watch this video: [www.youtube.com/watch?v=qlOMlwBoe\\_4](http://www.youtube.com/watch?v=qlOMlwBoe_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



**Practical 5.2: Prepare a salt by neutralising an acid with a soluble base**

- 1) Measure out 25 cm<sup>3</sup> of 2 moldm<sup>-3</sup> ammonia solution into a 100 cm<sup>3</sup> beaker.
- 2) Add 25 cm<sup>3</sup> of 1 moldm<sup>-3</sup> 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?

Equipment needed per group: two measuring cylinders (25 cm<sup>3</sup>), beaker (100 cm<sup>3</sup>), stirring rod, Bunsen burner, tripod, gauze, evaporating dish, spatula, filter paper; access to 1.0 moldm<sup>-3</sup> H<sub>2</sub>SO<sub>4</sub> (20 cm<sup>3</sup> per group), and 2.0 moldm<sup>-3</sup> NH<sub>3</sub> (25 cm<sup>3</sup> each per group)



-  $0.025 \times 1 = 0.025$

-  $0.025 \times 2 = 0.05$

- to ensure that both reactants are fully used up and that there is no excess of either reactant in the solution at the end

- excess insoluble base can be removed easily by filtration; excess soluble base cannot be easily removed

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 H<sub>3</sub>PO<sub>4</sub> with chloride salts:  
Eg  $3\text{CaCl}_2(\text{s}) + 2\text{H}_3\text{PO}_4(\text{l}) \rightarrow \text{Ca}_3(\text{PO}_4)_2(\text{s}) + 6\text{HCl}(\text{g})$

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

## f) 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:
  - H<sup>+</sup> ions with CO<sub>3</sub><sup>2-</sup> and SO<sub>3</sub><sup>2-</sup> ions**
    - H<sup>+</sup> ions react with CO<sub>3</sub><sup>2-</sup> and HCO<sub>3</sub><sup>-</sup> ions to give carbon dioxide and water:  
 $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ ;  $\text{H}^+ + \text{HCO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
    - H<sup>+</sup> ions react with SO<sub>3</sub><sup>2-</sup> and HSO<sub>3</sub><sup>-</sup> ions to give sulphur dioxide and water:  
 $2\text{H}^+ + \text{SO}_3^{2-} \rightarrow \text{SO}_2 + \text{H}_2\text{O}$ ;  $\text{H}^+ + \text{HSO}_3^{2-} \rightarrow \text{SO}_2 + \text{H}_2\text{O}$
    - if a solid sample of a carbonate (eg CaCO<sub>3</sub>) is added to a solution and bubbles/fizzing (also known as effervescence) is evolved, it shows that the solution contains H<sup>+</sup> 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 CO<sub>2</sub> then CO<sub>3</sub><sup>2-</sup> or HCO<sub>3</sub><sup>-</sup> ions are present in the solid; if the gas is shown to be SO<sub>2</sub> then SO<sub>3</sub><sup>2-</sup> or HSO<sub>3</sub><sup>-</sup> ions are present in the solid
  - CO<sub>2</sub> and SO<sub>2</sub> with limewater**
    - CO<sub>2</sub> reacts with Ca(OH)<sub>2</sub> (limewater) to produce insoluble CaCO<sub>3</sub>, which turns the limewater a milky colour:  $\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
    - SO<sub>2</sub> reacts with Ca(OH)<sub>2</sub> (limewater) to produce insoluble CaSO<sub>3</sub>, which also turns the limewater a milky colour:  $\text{SO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaSO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
    - if either gas is bubbled through the limewater for a longer period of time, a second reaction will happen and the milkiness will disappear:  
 $\text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$   
 $\text{CaSO}_3(\text{s}) + \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HSO}_3)_2(\text{aq})$
    - CO<sub>2</sub> and SO<sub>2</sub> are the only common gases which have this effect on limewater
    - Note: it is easy to distinguish between CO<sub>2</sub> and SO<sub>2</sub>; CO<sub>2</sub> is odourless but SO<sub>2</sub> smells like a burning match

## UNIT 5 – ACIDS, BASES AND SALTS

### (iii) $\text{NH}_4^+$ ions with $\text{OH}^-$ ions

- $\text{NH}_4^+$  ions react with  $\text{OH}^-$  ions to produce  $\text{NH}_3$ ;  $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- when warmed gently, the  $\text{NH}_3$  is given off as a pungent-smelling gas
- if  $\text{NaOH}$  is added to a solution and the mixture is warmed, and as a result a pungent smell is observed, then the solution contains  $\text{NH}_4^+$  ions
- if  $\text{NH}_4\text{Cl}$  is added to a solution and the mixture is warmed, and as a result a pungent smell is observed, then the solution contains  $\text{OH}^-$  ions

### (iv) $\text{NH}_3$ with $\text{HCl}$

- $\text{NH}_3$  (an alkaline gas) and  $\text{HCl}$  (an acidic gas) react together to give a salt ( $\text{NH}_4\text{Cl}$ ):  
 $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$
- 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  $\text{HCl}$  and placed close to a source of  $\text{NH}_3$ , a white smoke can be seen, confirming the presence of  $\text{NH}_3$
- If a piece of filter paper is soaked in concentrated  $\text{NH}_3$  and placed close to a source of  $\text{HCl}$ , a white smoke can be seen, confirming the presence of  $\text{HCl}$

### (v) Concentrated $\text{H}_2\text{SO}_4$ and $\text{Cl}^-$ ions

- If concentrated sulphuric acid is added to a solid sample containing  $\text{Cl}^-$  ions, an acid-salt reaction will take place and  $\text{HCl}$  (g) will be produced:  $\text{H}_2\text{SO}_4 + \text{Cl}^- \rightarrow \text{HSO}_4^- + \text{HCl}$
- If a piece of filter paper is soaked in concentrated  $\text{NH}_3$  and placed close to a source of  $\text{HCl}$ , a white smoke can be seen, confirming the presence of  $\text{HCl}$

## UNIT 5 – ACIDS, BASES AND SALTS



### 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:  $\text{NH}_4^+$ ,  $\text{OH}^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{SO}_3^{2-}$

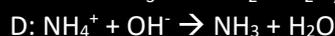
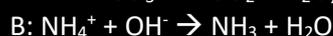
- Place a small spatula of sample A onto a watch glass; add a few drops of HCl and observe whether bubbles are formed; if bubbles are formed, observe whether or not the gas smells like a burning match; repeat for samples B, C, and D
- If bubbles are observed, half-fill a test-tube with limewater; add a spatula measure of the same sample into another test-tube; add 3 cm<sup>3</sup> 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
- 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
- Place a small spatula of sample A onto a watch glass; add a few drops of  $\text{NH}_4\text{Cl}$ ; observe whether or not a pungent smell is given off; repeat for samples B, C and D
- Place a very small spatula of sample A into a test tube and add 2 – 3 cm<sup>3</sup> of water so that the solid is mostly dissolved; then add a small spatula measure of powdered  $\text{CaCO}_3$ ; observe whether bubbles are formed; repeat for samples B, C and D
- Use your observations to identify which ion is present in which sample; present your answers in a table similar to the one below:

Sample	Bubbles with HCl? Y/N	Smell with HCl?	Effect of gas on limewater?	Smell with NaOH?	Smell with $\text{NH}_4\text{Cl}$ ?	Bubbles with $\text{CaCO}_3$ ? Y/N	Ion present
A							
B							
C							
D							

- Write equations for the reaction occurring with each sample which results in a positive identification

Equipment needed per group: three watch glasses, four test tubes, test tube rack, bung to fit test tube with delivery tube connected; access to solid samples of:  $\text{Na}_2\text{SO}_3$  (labelled A), KOH (labelled B),  $\text{Na}_2\text{CO}_3$  (labelled C),  $\text{NH}_4\text{Cl}$  (labelled D) (1 g per group) each with their own spatula; access to 1 mol dm<sup>-3</sup> HCl with dropping pipette (10 cm<sup>3</sup> per group), limewater (10 cm<sup>3</sup> per group), 1 mol dm<sup>-3</sup> NaOH (5 cm<sup>3</sup> per group) with dropping pipette, 1 mol dm<sup>-3</sup>  $\text{NH}_4\text{Cl}$  (5 cm<sup>3</sup> per group) with dropping pipette, powdered  $\text{CaCO}_3$  (2 g per group) with its own spatula

Sample	Bubbles with HCl? Y/N	Smell with HCl?	Effect of gas on limewater?	Smell with NaOH?	Smell with $\text{NH}_4\text{Cl}$ ?	Bubbles with $\text{CaCO}_3$ ? Y/N	Ion present
A	Y	Burning match	Milky then clear again	none	none	N	$\text{SO}_3^{2-}$
B	N (but it might get warm)	none	n/a	none	pungent	N	$\text{OH}^-$
C	N	none	n/a	none	none	Y	$\text{H}^+$
D	Y	none	Milky then clear again	none	none	N	$\text{CO}_3^{2-}$




**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:

- $\text{CO}_3^{2-}$  in a solid sample
- $\text{SO}_3^{2-}$  in a solid sample
- $\text{Cl}^-$  in a solid sample
- $\text{H}^+$  in solution
- $\text{OH}^-$  in solution
- $\text{NH}_4^+$  in solution
- $\text{NH}_3$  as a gas
- $\text{HCl}$  as a gas

In each case, give the equations for the reactions occurring

- add  $\text{HCl}$  (aq); odourless gas evolved which turns limewater milky and then clear again:  
 $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ ;  $\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$ ;  $\text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$
- add  $\text{HCl}$  (aq); gas evolved with burning-match smell which turns limewater milky and then clear again:  $2\text{H}^+ + \text{SO}_3^{2-} \rightarrow \text{SO}_2 + \text{H}_2\text{O}$ ;  $\text{SO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaSO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$ ;  $\text{CaSO}_3(\text{s}) + \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HSO}_3)_2(\text{aq})$
- add concentrated  $\text{H}_2\text{SO}_4$ ; if gas given off, test with filter paper soaked in concentrated  $\text{NH}_3$ ; a white smoke should be seen:  $\text{H}_2\text{SO}_4 + \text{Cl}^- \rightarrow \text{HSO}_4^- + \text{HCl}$ ;  $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$
- add  $\text{CaCO}_3(\text{s})$ ; gas evolved:  $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- add  $\text{NH}_4\text{Cl}(\text{aq})$  and warm; pungent smell given off:  $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- add  $\text{NaOH}(\text{aq})$  and warm; pungent smell given off:  $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- test with filter paper soaked in concentrated  $\text{HCl}$ ; a white smoke should be seen:  $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$
- test with filter paper soaked in concentrated  $\text{NH}_3$ ; a white smoke should be seen:  $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$

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

## g) Indicators and the pH Scale

### (i) The pH Scale

- Water dissociates very slightly to produce  $\text{H}^+$  and  $\text{OH}^-$  ions:  $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$ ; as a result, all aqueous solutions contain both  $\text{H}^+$  and  $\text{OH}^-$  ions
- Solutions which contain acids contain many more  $\text{H}^+$  ions than pure water; as a result, these solutions also contain fewer  $\text{OH}^-$  ions than pure water, because the extra  $\text{H}^+$  ions push the above equilibrium to the left; the higher the concentration of  $\text{H}^+$  ions, the lower the concentration of  $\text{OH}^-$  ions; any solution in which the concentration of  $\text{H}^+$  ions is greater than the concentration of  $\text{OH}^-$  ions is said to be an **acidic** solution
- Solutions which contain alkalis contain many more  $\text{OH}^-$  ions than pure water; as a result, these solutions also contain fewer  $\text{H}^+$  ions than pure water, because the extra  $\text{OH}^-$  ions push the above equilibrium to the left; the higher the concentration of  $\text{OH}^-$  ions, the lower the concentration of  $\text{H}^+$  ions; any solution in which the concentration of  $\text{OH}^-$  ions is greater than the concentration of  $\text{H}^+$  ions is said to be an **alkaline** solution
- A solution in which the concentration of  $\text{H}^+$  ions is equal to the concentration of  $\text{OH}^-$  ions, as is the case in pure water, is said to be a neutral solution

## UNIT 5 – ACIDS, BASES AND SALTS

- 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  $\text{mol dm}^{-3}$  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  $\text{mol dm}^{-3}$  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)



### Test your knowledge 7.1: Understanding the pH scale

Classify the following solutions as acidic, alkaline or neutral based on the information provided:

(a) it has a pH of 5	(i) it is a solution of sodium carbonate
(b) it is a solution of ammonia	(j) it has a pH of 7
(c) $[H^+] = [OH^-]$	(k) it has a pH of -1
(d) it is lemon juice	(l) $[OH^-] > [H^+]$
(e) it has a pH of 11	(m) it is rainwater
(f) $[H^+] > [OH^-]$	(n) it is pure water
(g) It is a solution of sodium hydroxide	(o) it is a solution of ammonium chloride
(h) It is a solution of sodium chloride	(p) it is a solution of potassium ethanoate

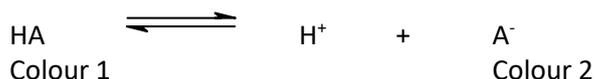
(a) acidic; (b) alkaline; (c) neutral; (d) acidic; (e) alkaline; (f) acidic; (g) alkaline; (h) neutral; (i) alkaline; (j) neutral; (k) (very) acidic; (l) alkaline; (m) acidic; (n) neutral; (o) acidic (due to salt hydrolysis); (p) alkaline (due to salt hydrolysis)

### (ii) Indicators

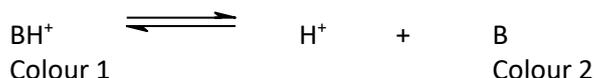
- An indicator is a substance which appears as one colour in acidic solutions and as a different colour in alkaline solutions; indicators can therefore be used to determine whether a solution is acidic or alkaline; most indicators are soluble organic acids or bases

## UNIT 5 – ACIDS, BASES AND SALTS

- Most indicators are weak acids in which the acid and its anion have different colours:



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



- at low pH, the  $\text{H}^+$  ions shift the position of equilibrium to the left and the indicator will appear as colour 1
- at high pH, the  $[\text{H}^+]$  is low so the equilibrium shifts to the right to produce more  $\text{H}^+$  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



### 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

- Pour 1 cm<sup>3</sup> of the solution of pH 3 into each of three different test tubes
- 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
- Repeat steps (a) and (b) for the solutions with pH values of 5, 7, 9 and 11
- Record your results in a table similar to the one below:

pH of solution	methyl orange colour	litmus colour	Phenolphthalein colour
3			
5			
7			
9			
11			

- Deduce the approximate pH at which each of the three indicators changes colour

## UNIT 5 – ACIDS, BASES AND SALTS

Equipment needed per group: five test tubes and one test tube rack; access to:  $0.1 \text{ mol dm}^{-3} \text{ CH}_3\text{COOH}$  (labelled pH 3); a solution containing  $0.1 \text{ mol dm}^{-3} \text{ CH}_3\text{COOH}$  and  $0.5 \text{ mol dm}^{-3} \text{ CHCOO}^-\text{Na}^+$  (labelled pH 5); distilled water (labelled pH = 7); a solution containing  $0.1 \text{ mol dm}^{-3} \text{ NH}_3$  and  $0.5 \text{ mol dm}^{-3} \text{ NH}_4\text{Cl}$  (labelled pH 9);  $0.1 \text{ mol dm}^{-3} \text{ NH}_3$  (labelled pH = 11);  $5 \text{ cm}^3$  per group for each solution; each bottle should have its own dropping pipette; access to solutions of phenolphthalein, methyl orange and litmus, each with their own dropping pipette ( $1 \text{ cm}^3$  per group)

pH of solution	methyl orange colour	litmus colour	Phenolphthalein colour
3	pink	red	colourless
5	yellow	red	colourless
7	yellow	red/blue	colourless
9	yellow	blue	pink/colourless
11	yellow	blue	pink

- MO changes colour at pH = 5; L changes colour at pH 7; PP changes colour at pH 9

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

#### (iii) 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



**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.

Students will need access to the different colours shown by universal indicator.

Answer: 1 – 3 red; 4 - 5 orange; 6 – yellow; 7 – green; 8 – blue; 9 – 11 indigo; 12 – 14 violet



**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.

Eg pH 1 stomach acid (HCl); pH 2 lemon juice/vinegar; pH 3 orange juice; pH 4 tomato juice; pH 5 black coffee; pH 6 milk; pH 7 water; pH 8 sea water; pH 9 a solution of baking soda; pH 10 milk of magnesia; pH 11 ammonia solution; pH 12 soapy water; pH 13 bleach



**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 cm<sup>3</sup> 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.

Solution	Colour it turns UI	pH	Possible identity
A			
B			
C			
D			
E			

## UNIT 5 – ACIDS, BASES AND SALTS

Equipment needed per group: five test tubes and one test tube rack; access to: rainwater (labelled A) (you can ensure a PH of 5 by bubbling CO<sub>2</sub> through it); distilled water (labelled B); a diluted solution of bleach (C); a solution of baking soda (D); diluted vinegar or lemon juice (E); the exact identity of the solution is not important but they should turn UI the following colours: A – orange/yellow ; B – green; C – violet; D – blue/indigo; E - red  
5 cm<sup>3</sup> per group for each solution; each bottle should have its own dropping pipette; access to universal indicator solution with dropping pipette (1 cm<sup>3</sup> per group)

Solution	Colour it turns UI	pH	Possible identity
A	orange/yellow	5	rainwater (or urine)
B	green	7	water
C	violet	11	bleach
D	blue/indigo	9	soda or magnesia
E	red	3	lemon juice or vinegar



### Practical 8.4: Determine the pH of different soil samples

You are given access to three different soil 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 cm<sup>3</sup> beaker.
- 2) Add 40 cm<sup>3</sup> 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 cm<sup>3</sup> 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.

Equipment needed per group: one beaker (100 cm<sup>3</sup>); one measuring cylinder (50 cm<sup>3</sup>); one stirring rod; one funnel; three pieces of filter paper; three boiling tubes; access to three different soil samples (20 g per group), each with its own spoon; access to a mass balance; access to universal indicator with its own dropping pipette (1 cm<sup>3</sup> per group)

Note: the soil samples should ideally cover a range of different types, ideally with a range of different acidities.

### 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<sup>+</sup> ions or OH<sup>-</sup> 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; it 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<sup>+</sup>) will turn blue litmus paper red but will have no effect on red litmus paper
  - alkaline solutions (containing CO<sub>3</sub><sup>2-</sup> or HCO<sub>3</sub><sup>-</sup>) 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 SO<sub>2</sub> are acidic gases which turn blue litmus paper red
  - NH<sub>3</sub> is an alkaline gas which turns red litmus paper blue

## UNIT 5 – ACIDS, BASES AND SALTS



### 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), Na<sub>2</sub>CO<sub>3</sub>(aq) and NaCl(aq)

You have access to three different indicators: litmus, methyl orange and phenolphthalein

- (a) Pour 5 cm<sup>3</sup> of HCl(aq) into a beaker
- (b) Dip a strip of red litmus paper into the solution; note any colour change
- (c) Dip a strip of blue litmus paper into the solution; note any colour change
- (d) Rinse out the beaker and measuring cylinder and repeat steps (a) – (c) using the other solutions (NaOH, Na<sub>2</sub>CO<sub>3</sub> and NaCl)
- (e) Pour 1 cm<sup>3</sup> of each solution (HCl, NaOH, Na<sub>2</sub>CO<sub>3</sub> and NaCl) separately into four test tubes
- (f) Add two drops of methyl orange indicator to each test tube; note any colour change
- (g) Rinse out the test tubes and repeat steps (e) and (f), but this time add two drops of phenolphthalein; note any colour changes
- (h) Record your results in a table similar to the one below:

Solution	Effect on red litmus paper	Effect on blue litmus paper	Effect on methyl orange	Effect on phenolphthalein
HCl				
NaOH				
Na <sub>2</sub> CO <sub>3</sub>				
NaCl				

Equipment needed per group: one small beaker (<50 cm<sup>3</sup>) and one small measuring cylinder (10 cm<sup>3</sup>); four strips of red litmus paper, four strips of blue litmus paper, access to solutions of 0.1 mol dm<sup>-3</sup> HCl, NaOH, Na<sub>2</sub>CO<sub>3</sub> and NaCl, (10 cm<sup>3</sup> per group) access to phenolphthalein and methyl orange indicator solutions (1 cm<sup>3</sup> per group), 8 test tubes, 1 test tube rack

Solution	Effect on red litmus paper	Effect on blue litmus paper	Effect on methyl orange	Effect on phenolphthalein
HCl	stays red	turns red	pink	colourless
NaOH	turns blue	stays blue	yellow	pink
Na <sub>2</sub> CO <sub>3</sub>	turns blue	stays blue	yellow	pink
NaCl	stays red	stays blue	Yellow	colourless



**Test your knowledge 9.2: Summarising the Qualitative Analysis of Cations, Anions and Gases**

- (a) Describe a test for  $H^+$  ions which uses a simple chemical reaction
- (b) Describe a test for  $H^+$  ions which uses an indicator
- (c) Describe a test for  $OH^-$  which uses a simple chemical reaction and a smell
- (d) Describe a test for  $OH^-$  which uses an indicator
- (e) Name three gases which turn blue litmus paper red
- (f) Name one gas which turns red litmus paper blue
- (g) Name two gases which turn limewater milky and then clear again and suggest how the gases could be distinguished
- (h) Describe a test for HCl gas which does not use an indicator
- (i) Describe a test for  $NH_3$  gas which does not use an indicator
- (j) Describe a test for  $Cl^-$  ions in a solid sample
- (k) Describe a test for  $NH_4^+$  ions
- (l) Describe a test for  $CO_3^{2-}$  ions in a solid sample
- (m) Describe a test for  $SO_3^{2-}$  ions in a solid sample
- (n) Suggest how you could demonstrate that a solution contains a neutral salt

(a) Add  $CaCO_3(s)$  - effervescence; (b) add blue litmus paper – it turns red; (c) add  $NH_4Cl$  and warm – a pungent smell; (d) add red litmus paper – it turns blue; (e) hydrogen chloride, sulphur dioxide and nitrogen dioxide; (f) ammonia; (g) carbon dioxide and sulphur dioxide,  $SO_2$  smells of burning matches and also turns blue litmus red; (h) hang filter paper soaked in concentrated  $NH_3$  close to the source – white smoke formed; (i) hang filter paper soaked in concentrated HCl close to the source – white smoke formed; (j) add concentrated  $H_2SO_4$  and then test the gas either with filter paper soaked in concentrated  $NH_3$  – white smoke formed, or with blue litmus paper – it turns red; (k) add NaOH and warm – pungent smell; (l) add HCl and bubble gas through limewater – odourless gas which turns limewater milky and then clear again; (m) add HCl and bubble gas through limewater – burning-match-smelling gas which turns limewater milky and then clear again and also turns blue litmus paper red; (n) conducts electricity which proves that ions are present; turns universal indicator green which shows that  $H^+$ ,  $OH^-$  and  $CO_3^{2-}$  are not present

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

## (i) 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



**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?

## UNIT 5 – ACIDS, BASES AND SALTS

- Pipette: very accurate but can only deliver one volume
- Volumetric flask: very accurate but can only store one volume
- Burette: slightly less accurate than a pipette but can deliver any volume up to 50 cm<sup>3</sup>
- Measuring cylinder: not accurate
- Pipettes and burettes are most useful for carrying out titrations

- In most titrations, a fixed volume (usually 25 cm<sup>3</sup>) 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

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

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


**Practical 10.2: Determine the concentration of a solution of NaOH by titration against 1.0 mol dm<sup>-3</sup> HCl**
**1) Add 0.1 mol dm<sup>-3</sup> 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 cm<sup>3</sup> 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 cm<sup>3</sup> and the 5 cm<sup>3</sup> mark; record this level accurately, to the nearest 0.05 cm<sup>3</sup>, 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

**2) Pipette 25.0 cm<sup>3</sup> of NaOH(aq) into a conical flask and add a suitable indicator**

- Pour around 100 cm<sup>3</sup> 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 cm<sup>3</sup> 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

**3) Add the HCl to the NaOH until the indicator changes colour**

- Place the conical flask containing 25.0 cm<sup>3</sup> 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 cm<sup>3</sup> has been added
- close the tap 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 cm<sup>3</sup>) as the “final volume”
- Once you have reached this stage for the first time, stop and ask the teacher to check your progress

## UNIT 5 – ACIDS, BASES AND SALTS

### 4) Repeat the titration until concordancy is achieved

- After rinsing out the conical 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  $0.10 \text{ cm}^3$  of each other, the results can be described as concordant and you have completed the experiment
- If the values differ by more than  $0.10 \text{ cm}^3$ , 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:

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Final volume ( $\text{cm}^3$ )					
Initial volume ( $\text{cm}^3$ )					
Titre volume ( $\text{cm}^3$ )					
Concordant? (Y or N)					

### 5) 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 ( $n = C \times V$ )
- 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-5QJlr7Xm4](http://www.youtube.com/watch?v=Y-5QJlr7Xm4) and [www.youtube.com/watch?v=VADhF7M8bJ8](http://www.youtube.com/watch?v=VADhF7M8bJ8)

Equipment needed per group: 1 burette, 1  $25 \text{ cm}^3$  pipette, 1 pipette filler, one conical flask, one funnel, two  $100 \text{ cm}^3$  beakers and the means to label them, clamp, stand, boss; access to  $0.1 \text{ mol dm}^{-3}$  HCl, a solution of NaOH of approximately  $0.08 \text{ mol dm}^{-3}$  but with the concentration not labelled) (approx  $100 \text{ cm}^3$  per group), phenolphthalein indicator and suitable dropping pipette

Note: this practical is the most important practical in SS Chemistry; it is also the most challenging in terms of the techniques required; it is recommended that when carrying out this practical for the first time, each group performs each step together so that the teacher can check that the student is performing the practical correctly. Particular things to check in each group are: no air bubbles in burette tip, initial reading is recorded to  $0.05 \text{ cm}^3$  and is correct, funnel is removed

- if the NaOH molarity is  $0.08 \text{ mol dm}^{-3}$  then a typical titre volume should be around  $15 \text{ cm}^3$
- a correctly completed table would look something like this: it may be an idea to share this with the students beforehand:

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Final volume ( $\text{cm}^3$ )	23.30	24.85	23.15		
Initial volume ( $\text{cm}^3$ )	1.25	23.30	1.70		
Titre volume ( $\text{cm}^3$ )	22.05	21.55	21.45		
Concordant? (Y or N)	NO	YES	YES		

- the average titre volume (using concordant results only) should be  $(21.55 + 21.45)/2 = 21.40 \text{ cm}^3$  (using the above results)
- moles of HCl =  $0.1 \times \text{titre volume} / 1000$
- moles of NaOH = moles of HCl
- molarity of NaOH = moles of NaOH /  $0.025$

## UNIT 5 – ACIDS, BASES AND SALTS

- 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 cm<sup>3</sup> of a 0.10 moldm<sup>-3</sup> solution of NaOH was required to react with 25 cm<sup>3</sup> of a solution of H<sub>2</sub>SO<sub>4</sub>. What was the concentration of the H<sub>2</sub>SO<sub>4</sub> solution?  
Worked answer:
  - Equation: H<sub>2</sub>SO<sub>4</sub> + 2NaOH → Na<sub>2</sub>SO<sub>4</sub> + 2H<sub>2</sub>O
  - Moles of NaOH used in titration = 28.3/1000 x 0.1 = 2.8 x 10<sup>-3</sup>
  - 2:1 ratio so moles of H<sub>2</sub>SO<sub>4</sub> present in conical flask = 2.8 x 10<sup>-3</sup>/2 = 1.4 x 10<sup>-3</sup>
  - so concentration of H<sub>2</sub>SO<sub>4</sub> = 1.4 x 10<sup>-3</sup>/25 x 1000 = 0.056 moldm<sup>-3</sup>

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



#### 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 cm<sup>3</sup> of a 0.1 moldm<sup>-3</sup> solution of NaOH, starting with solid NaOH?
  - How would you prepare 250 cm<sup>3</sup> of a 0.2 moldm<sup>-3</sup> solution of H<sub>2</sub>O<sub>2</sub>, starting with 2.0 moldm<sup>-3</sup> H<sub>2</sub>O<sub>2</sub>?
- Moles of NaOH required = 0.25 x 0.1 = 0.025 so mass of NaOH required = 0.025 x 40 = 1.0 g  
Weigh out exactly 1.0 g of NaOH on a mass balance using a weighing boat  
Pour the NaOH into a beaker  
Rinse the weighing boat with distilled water and ensure all the rinsings run into the beaker  
Dissolve the NaOH in a small quantity of distilled water  
Pour the solution into a 250 cm<sup>3</sup> volumetric flask using a funnel  
Rinse the beaker with distilled water and pour the rinsings into the volumetric flask  
Add distilled water until the meniscus rests exactly on the graduation mark, shaking continuously
- Dilution factor = 2/0.2 = 10 so volume needed = 250/10 = 25 cm<sup>3</sup>  
Pipette 25 cm<sup>3</sup> of 2.0 moldm<sup>-3</sup> H<sub>2</sub>O<sub>2</sub> into a 250 cm<sup>3</sup> volumetric flask  
Add distilled water until the meniscus rests exactly on the graduation mark, shaking continuously

- 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



**Practical 11.2: Determine the relative formula mass, and hence water of crystallisation, of hydrated sodium carbonate,  $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ , by titration ( $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$ )**

- 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 cm<sup>3</sup> of distilled water.
- 3) Transfer this solution into a 250 cm<sup>3</sup> volumetric flask, and make the volume in the flask up to 250 cm<sup>3</sup> by using washings from the beaker; shake continuously to ensure that the contents are thoroughly mixed
- 4) Pipette a 25.0 cm<sup>3</sup> 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<sup>-3</sup> 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  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

Equipment needed per group: 1 burette, 1 25 cm<sup>3</sup> pipette, 1 pipette filler, clamp, boss, stand, one conical flask, two funnels, two 100 cm<sup>3</sup> beakers with the means to label them, one 250 cm<sup>3</sup> volumetric flask, one weighing boat; access to mass balance, access to 0.1 moldm<sup>-3</sup> HCl, a sample of hydrated sodium carbonate with the formula not labelled, with spatula, methyl orange indicator and suitable dropping pipette

- If the student uses 3.5 g of solid the titre volume should be around 24 cm<sup>3</sup>
- Using 24 cm<sup>3</sup>, moles of HCl = 0.1 x 0.0224 = 0.00224 so moles of  $\text{Na}_2\text{CO}_3$  in conical flask is 0.00224/2 = 0.00112
- Moles of  $\text{Na}_2\text{CO}_3$  in volumetric flask = 0.00112 x 10 = 0.012
- Molar mass of  $\text{Na}_2\text{CO}_3 = 3.5/0.012 = 292$
- $106 + 18x = 292$  so  $x = 10$



**Practical 11.3: Determine the percentage purity of a sample of vinegar ( $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$ )**

- 1) A solution of vinegar is labelled as 62.3 gdm<sup>-3</sup>. Pipette 25 cm<sup>3</sup> of this vinegar sample into a 250 cm<sup>3</sup> 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 cm<sup>3</sup> sample of the 0.05 moldm<sup>-3</sup> 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

## UNIT 5 – ACIDS, BASES AND SALTS

Equipment needed per group: 1 burette, 1 25 cm<sup>3</sup> pipette, 1 pipette filler, one conical flask, two funnels, two 100 cm<sup>3</sup> beakers with the means to label them, one 250 cm<sup>3</sup> volumetric flask; access to 0.1 mol dm<sup>-3</sup> NaOH, a sample of ethanoic acid (approx 1 mol dm<sup>-3</sup>) labelled “vinegar 62.3 g dm<sup>-3</sup>”, phenolphthalein indicator and suitable dropping pipette

- If the ethanoic acid is around 1 mol dm<sup>-3</sup> the titre volume should be around 13 cm<sup>3</sup>
- moles of NaOH =  $0.05 \times 25/1000 = 0.00125$  so moles of CH<sub>3</sub>COOH in titration = 0.00125
- molarity of = CH<sub>3</sub>COOH / (titre volume/1000); using 13 cm<sup>3</sup>,  $C = 0.00125/0.013 = 0.096 \text{ mol dm}^{-3}$
- so molarity before dilution =  $0.096 \times 250/25 = 0.96 \text{ mol dm}^{-3}$
- mass concentration =  $0.96 \times 60 = 57.7 \text{ g dm}^{-3}$
- percentage purity =  $57.7/62.3 \times 100 = 92.6 \%$
- phenolphthalein necessary because acid is weak so methyl orange will not work


**Test your knowledge 11.4: Volumetric Analysis – Titrations**

- (a) Succinic acid has the formula  $(\text{CH}_2)_n(\text{COOH})_2$  and reacts with dilute sodium hydroxide as follows:  
 $(\text{CH}_2)_n(\text{COOH})_2 + 2\text{NaOH} \rightarrow (\text{CH}_2)_n(\text{COONa})_2 + 2\text{H}_2\text{O}$   
 2.0 g of succinic acid were dissolved in water and the solution made up to  $250 \text{ cm}^3$ . This solution was placed in a burette and  $18.4 \text{ cm}^3$  was required to neutralise  $25 \text{ cm}^3$  of  $0.1 \text{ mol dm}^{-3}$  NaOH. Deduce the molecular formula of the acid and hence the value of  $n$ .
- (b) Sodium carbonate exists in hydrated form,  $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ , in the solid state. 3.5 g of a sodium carbonate sample was dissolved in water and the volume made up to  $250 \text{ cm}^3$ .  $25.0 \text{ cm}^3$  of this solution was titrated against  $0.1 \text{ mol dm}^{-3}$  HCl and  $24.5 \text{ cm}^3$  of the acid were required. Calculate the value of  $x$  given the equation:  $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$
- (c)  $25 \text{ cm}^3$  of a sample of vinegar ( $\text{CH}_3\text{COOH}$ ) was pipetted into a volumetric flask and the volume was made up to  $250 \text{ cm}^3$ . This solution was placed in a burette and  $13.9 \text{ cm}^3$  were required to neutralise  $25 \text{ cm}^3$  of  $0.1 \text{ mol dm}^{-3}$  NaOH. Calculate the molarity of the original vinegar solution and its concentration in  $\text{g dm}^{-3}$ , given that it reacts with NaOH in a 1:1 ratio.
- (d) 2.5 g of a sample of impure ethanedioic acid,  $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ , was dissolved in water and the solution made up to  $250 \text{ cm}^3$ . This solution was placed in a burette and  $21.3 \text{ cm}^3$  were required to neutralise  $25 \text{ cm}^3$  of  $0.1 \text{ mol dm}^{-3}$  NaOH. Given that ethanedioic acid reacts with NaOH in a 1:2 ratio, calculate the percentage purity of the sample.
- (e) A toilet cleaner containing sodium hydrogensulphate,  $\text{NaHSO}_4$  is believed to have been contaminated. 5.678 g of the sample were dissolved in water and the solution was made up to  $250 \text{ cm}^3$ . This solution was placed in a burette and  $23.1 \text{ cm}^3$  of it were required to neutralise  $25 \text{ cm}^3$  of  $0.1 \text{ mol dm}^{-3}$  sodium hydroxide.  
 Calculate the percentage purity of the sample given that the species react in a 1:1 ratio.

- (a) Moles of NaOH =  $0.0025$ ; moles of SA used =  $0.0025/2 = 0.00125$ ; moles of SA in volumetric flask =  $0.00125 \times 250/18.4 = 0.016984$ ,  $\text{mr of SA} = 2/0.016984 = 118$ ;  $(\text{CH}_2)_n = 118 - 90 = 24$  so  $n = 28/14 = 2$
- (b) Moles of HCl =  $0.0245 \times 0.1 = 0.00245$ ; moles of  $\text{Na}_2\text{CO}_3$  used =  $0.00245/2 = 0.001225$ ; moles of in volumetric flask =  $0.001225 \times 10 = 0.01225$ , molar mass of  $\text{Na}_2\text{CO}_3 = 3.5/0.001225 = 286$ ;  $x\text{H}_2\text{O} = 280 - 106 = 180$ ;  $x = 180/18 = 10$
- (c) Moles of NaOH =  $0.025 \times 0.1 = 0.0025$ ; moles of  $\text{CH}_3\text{COOH}$  used =  $0.0025$ ; molarity of  $\text{CH}_3\text{COOH}$  used =  $0.0025/(13.9/1000) = 0.180$ ; molarity before dilution =  $0.180 \times 250/25 = 1.80 \text{ mol dm}^{-3}$ ; mass concentration =  $1.8 \times 60 = 108 \text{ g dm}^{-3}$
- (d) Moles of NaOH =  $0.0025$ ; moles of acid used =  $0.0025/2 = 0.00125$ ; moles of acid in volumetric flask =  $0.00125 \times 250/21.3 = 0.0147$ ; mass of pure acid =  $0.0147 \times 126 = 1.85 \text{ g}$ ; % purity =  $1.85/2.5 \times 100 = 73.9 \%$
- (e) Moles of NaOH =  $0.0025$ ; moles of acid used =  $0.0025$ ; moles of acid in volumetric flask =  $0.0025 \times 250/21.3 = 0.0271$ ; mass of acid =  $0.0271 \times 120.1 = 3.25$ ; % purity =  $57.2 \%$

## UNIT 5 – ACIDS, BASES AND SALTS

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

#### 12.1 END-OF-TOPIC QUIZ

#### UNIT 5 – ACIDS, BASES AND SALTS



- Write symbol equations for the following neutralisation reactions:
  - hydrochloric acid and sodium hydroxide
  - sulphuric acid and magnesium oxide
  - nitric acid and calcium carbonate
- Explain the meaning of the term deliquescent, hygroscopic and efflorescent and give one example of an acid, base or salt with each property.
- State the difference between a strong acid and a weak acid. Explain how  $1.0 \text{ mol dm}^{-3}$  HCl and  $1.0 \text{ mol dm}^{-3}$   $\text{CH}_3\text{COOH}$  would differ in terms of their:
  - pH
  - enthalpy of neutralisation by NaOH
  - rate of reaction with calcium carbonate
  - electrical conductivity
- Suggest whether the following salts undergo salt hydrolysis, and if so, whether they are acidic or basic salts:
  - sodium chloride
  - ammonium chloride
  - sodium ethanoate
- Describe a chemical test which would identify the following:
  - $\text{H}^+$  ions in solution
  - $\text{OH}^-$  ions in solution
  - $\text{NH}_4^+$  ions in solution
  - $\text{Cl}^-$  ions from a solid sample
  - $\text{SO}_3^{2-}$  ions from a solid sample
  - $\text{NH}_3$  gas
- Describe how:
  - universal indicator could distinguish between samples of water, lemon juice and bleach
  - litmus paper could distinguish between HCl and  $\text{NH}_3$
- Deduce the identity of the Group 1 metal M based on the following experiment:

2.5 g of the metal carbonate  $\text{M}_2\text{CO}_3$  was dissolved in water and the volume made up to  $250 \text{ cm}^3$ . A  $25 \text{ cm}^3$  sample of this solution was titrated against  $0.2 \text{ mol dm}^{-3}$  HCl and  $10.8 \text{ cm}^3$  of the acid was required to neutralise the sample.

## UNIT 5 – ACIDS, BASES AND SALTS

- $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
  - $\text{H}_2\text{SO}_4 + \text{MgO} \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}$
  - $\text{CaCO}_3 + 2\text{HNO}_3 \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O}$
- Deliquescent – absorbs water from atmosphere and dissolves in it (eg NaOH or CaCl<sub>2</sub>)  
Hygroscopic – absorbs water from atmosphere (eg conc. H<sub>2</sub>SO<sub>4</sub> or any deliquescent substance)  
Efflorescent – contains water which it releases (eg CaSO<sub>4</sub>·2H<sub>2</sub>O)
- strong acid – lower pH
  - strong acid – more exothermic enthalpy of neutralisation
  - strong acid – faster reaction with calcium carbonate
  - strong acid – greater electrical conductivity
- no - neutral (salt of strong acid and strong base)
  - yes - acidic (salt of weak base)
  - yes – basic (salt of weak acid)
- Add CaCO<sub>3</sub>; observe fizzing
  - Add NH<sub>4</sub>Cl and warm – pungent smell
  - Add NaOH and warm – pungent smell
  - Add concentrated H<sub>2</sub>SO<sub>4</sub> – white fumes which turn blue litmus red and give white smoke in presence of filter paper soaked in NH<sub>3</sub>
  - Add acid; gas given off which smells like burning matches
  - Gives white smoke in presence of filter paper soaked in concentrated HCl
- water – green; lemon juice – red; bleach – violet
  - HCl turns blue litmus red; NH<sub>3</sub> turns red litmus blue
- Moles of HCl =  $0.2 \times 10.8/1000 = 0.00216$ ;  $\text{M}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{MCl} + \text{CO}_2 + \text{H}_2\text{O}$  so moles of M<sub>2</sub>CO<sub>3</sub> used =  $0.00216/2 = 0.00108$ ; moles of M<sub>2</sub>CO<sub>3</sub> in volumetric flask =  $0.00108 \times 10 = 0.0108$ ; molar mass =  $2.5/0.0108 = 231$ ;  $2M = 231 - 60 = 171$ ;  $M = 171/2 = 85.5$ ;  $M = \text{Rb}$