

chemistry spm form 4 short notes chapter 7 acids and bases

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Introduction

In the South African school system, the SPM (School Certificate) is the culmination of secondary education, and for many students, Chemistry is a compulsory subject. Chapter 7 of the Form 4 Chemistry curriculum—Acids and Bases—lays the foundation for understanding everyday chemical reactions, industrial processes, and biological systems. This article provides a concise, yet comprehensive, set of short notes that will help you grasp the core concepts, solve exam questions, and ace your SPM Chemistry exams.

Overview of Chapter /7: Acids and Bases

Chapter 7 introduces the basic building blocks of chemical reactions that involve the transfer of protons or electrons. It covers the definitions, classifications, properties, and practical applications of acids and bases, along with the tools used to measure acidity and basicity. By mastering these concepts, you'll be able to predict reaction outcomes, calculate concentrations, and design experiments—all essential skills for the SPM Chemistry exam.

Definition of Acids and Bases

Acids and bases are defined by different theories. The most widely taught definitions for Form /4 students are:

- Arrhenius definition: Acids produce H^+ ions in aqueous solution; bases produce OH^- ions.
- Brønsted–Lowry definition: Acids donate protons (H^+); bases accept protons.
- Lewis definition: Acids accept electron pairs; bases donate electron pairs.

Understanding these perspectives allows students to classify substances accurately and predict reaction behavior.

Historical Context

The study of acids and bases dates back to the 17th century when Robert Boyle first described the “acidic” taste of vinegar. In the 19th century, Svante Arrhenius introduced the ionisation concept, which laid the groundwork for modern chemistry. Later, Brønsted and Lowry expanded the theory to include proton transfer, and Lewis further generalized the concept to electron pair interactions. These milestones illustrate how scientific thought evolves and why a multi-theoretical approach is useful for students.

Classification of Acids

Arrhenius Acids

These acids release hydrogen ions when dissolved in water. Common examples include:

- Hydrochloric acid (HCl)
- Sulphuric acid (H_2SO_4)
- Hydrogen sulphide (H_2S)

Brønsted–Lowry Acids

Any substance that can donate a proton qualifies. For instance:

- Acetic acid (CH_3COOH)
- Hydrogen fluoride (HF)
- Hydrogen peroxide (H_2O_2)

Lewis Acids

Lewis acids accept electron pairs. Typical examples include:

- Aluminium chloride (AlCl_3)
- Iron(III) chloride (FeCl_3)
- BF_3 (boron trifluoride)

Classification of Bases

Arrhenius Bases

These substances produce hydroxide ions in aqueous solution:

- Sodium hydroxide (NaOH)
- Calcium hydroxide (Ca(OH)_2)

Brønsted–Lowry Bases

Any proton acceptor qualifies. Common examples:

- Ammonia (NH_3)
- Ammonium hydroxide (NH_4OH)
- Water acting as a base ($\text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+$)

Lewis Bases

Lewis bases donate electron pairs:

- Ammonia (NH_3)
- Water (H_2O)
- Phosphorus pentachloride (PCl_5 ...)

Properties of Acids and Bases

Taste, Texture, and Reaction with Metals

Acids often have a sour taste, while bases feel slippery. When acids react with metals, they produce hydrogen gas:

1. $2\text{H}^+ + \text{Fe} \rightarrow \text{Fe}^{2+} + \text{H}_2$
2. $2\text{H}^+ + \text{Zn} \rightarrow \text{Zn}^{2+} + \text{H}_2$

Indicators and Colour Changes

Indicators are weak acids or bases that change colour depending on the pH:

- Litmus: blue in basic solutions, red in acidic solutions.
- Methyl orange: red in acidic, yellow in basic.
- Phenolphthalein: colourless in acidic, pink in basic.

pH Scale and Measurement

pH Scale Overview

The pH scale ranges from 0 to 14. A pH of 7 is neutral, values below 7 indicate acidity, and values above 7 indicate

basicity. The pH is defined as:

$$\text{pH} = -\log \cdot \epsilon[\text{H}^+]$$

pH Meter and Litmus Paper

pH meters provide an electronic reading of hydrogen ion concentration, while litmus paper offers a quick visual indicator. For accurate results, calibrate the meter with standard buffers before measuring unknown solutions.

Calculating pH of Common Solutions

For a 0.01 M hydrochloric acid solution:

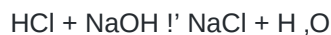
1. $[\text{H}^+] = 0.01 \text{ M}$
2. $\text{pH} = -\log \cdot \epsilon(0.01) = 2$

Similarly, a 0.01 M sodium hydroxide solution has $[\text{OH}^-] = 0.01 \text{ M}$. Using $K_w = 1.0 \times 10^{-14}$, calculate $[\text{H}^+] = K_w / [\text{OH}^-] = 1.0 \times 10^{-12}$, so $\text{pH} = 12$.

Acid–Base Reactions

Neutralisation Reactions

When an acid reacts with a base, the result is typically a salt and water:



Neutralisation is exothermic, releasing heat that can be measured with a calorimeter.

Acid–Base Titration

Titration is a quantitative technique to determine the concentration of an unknown acid or base. The general procedure involves:

1. Preparing a standard solution of known concentration.
2. Adding the standard solution dropwise to the unknown until the endpoint is reached, indicated by a colour change.
3. Calculating the concentration using the stoichiometry of the reaction.

Buffer Solutions

Buffers resist changes in pH when small amounts of acid or base are added. A typical buffer consists of a weak acid and its conjugate base (e.g., acetic acid and sodium acetate). The Henderson–Hasselbalch equation describes the buffer capacity:

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

Important Equations and Formulas

Ionisation Constant (K_a)

K_a measures the strength of an acid:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Base Dissociation Constant (K_b)

K_b quantifies the strength of a base:

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Relationship Between K_a and K_b

For a conjugate acid–base pair:

$$K_a \times K_b = K_w = 1.0 \times 10^{-14}$$

Common Acids and Bases in the Classroom

Household Acids

- Vinegar (acetic acid)
- Lemon juice (citric acid)
- Carbonated drinks (carbonic acid)

Household Bases

- Baking soda (sodium bicarbonate)
- Ammonia solution (household cleaner)
- Dish soap (sodium lauryl sulfate)

Practical Experiments for Form 4 Students

Titration with Hydrochloric Acid and Sodium Hydroxide

Materials: 0.1 M HCl, 0.1 M NaOH, phenolphthalein indicator, burette, flask.

1. Fill the burette with NaOH.
2. Add a measured volume of HCl to the flask.
3. Stir and add phenolphthalein until a faint pink colour appears.
4. Record the volume of NaOH used.

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