Gases: Preparation & Tests

Describe how the following gases are prepared:

Carbon Dioxide

Hydrogen

Oxygen

Give equations.

Describe tests for the following gases:

Oxygen

Ammonia

Carbon Dioxide

Chlorine

Hydrogen

Sulfur Dioxide

Tests for Water

State two tests for water. Give equations.

Gas Preparation:

Carbon Dioxide:

React calcium carbonate (marble chips) with dilute hydrochloric acid:

$$CaCO_{3\,(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2\,(aq)} + CO_{2\,(g)} + H_2O_{(l)}$$

Hydrogen:

React zinc with dilute hydrochloric acid:

$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$$

Oxygen:

Mix hydrogen peroxide with manganese(IV) oxide (which acts as a catalyst):

$$2H_2O_{2\,(l)} \xrightarrow{\text{Manganese(IV) Oxide}} O_{2\,(g)} + 2H_2O_{(l)}$$

Gas Tests:

Oxygen: relights a glowing splint

Ammonia: turns damp red litmus paper **blue**

Carbon Dioxide: turns limewater milky

Chlorine: bleaches (i.e. turns white) damp litmus paper

Hydrogen: burns with a squeaky 'pop' when lit with a lighted splint

Sulfur Dioxide: decolourises filter paper which has been dipped into purple acidified potassium manganate(VII)

Tests for Water:

Cobalt Chloride Paper: Turns from blue to pink when water is added:

$$CoCl_{2\,(s)} + 6H_2O_{(l)} \underset{\text{Heat}}{\overset{\text{Water}}{\rightleftharpoons}} CoCl_2 \cdot 6H_2O_{(s)}$$

Blue — Pink

Anhydrous Copper(II) Sulfate: Turns from white to blue when water is added:

$$CuSO_{4\,(s)} + 5H_2O_{(l)} \underset{\text{Heat}}{\overset{\text{Water}}{\rightleftharpoons}} CuSO_4 \cdot 5H_2O_{(s)}$$

White — Blue

Chemical Tests: Anions

Describe how to test for the following ions:

Chloride

Bromide

Iodide

Carbonate

Sulfate

Nitrate

Sulfite

Give equations for the first five tests.

Halide ions:

- Acidify with dilute nitric acid
- Add silver nitrate solution

Positive results:

Chloride ions, Cl^-	White precipitate	$Ag^+_{(aq)} + Cl^-_{(aq)} \longrightarrow AgCl_{(s)}$
Bromide ions, Br^-	Cream precipitate	$Ag^+_{(aq)} + Br^-_{(aq)} \longrightarrow AgBr_{(s)}$
Iodide ions, I^-	Yellow precipitate	$Ag^+_{(aq)} + I^-_{(aq)} \longrightarrow AgI_{(s)}$

Carbonate ions (CO_3^{2-}):

- Add dilute hydrochloric acid

Positive result:
Carbon dioxide produced (effervescence). CO_2 turns limewater cloudy.

$$2H^+_{(aq)} + CO^{2-}_{3\,(aq)} \longrightarrow CO_{2\,(g)} + H_2O_{(l)}$$

Sulfate ions (SO_4^{2-}):

- Acidify with dilute hydrochloric acid
- Add barium nitrate solution

Positive result:
White precipitate of (insoluble) barium sulfate produced:

$$Ba^{2+}_{(aq)} + SO^{2-}_{4\,(aq)} \longrightarrow BaSO_{4\,(s)}$$

Nitrate ions (NO_3^-):

- Add sodium hydroxide solution followed by small pieces of aluminium foil
- Warm gently

Positive result: Ammonia gas is given off (this turns damp red litmus paper **blue**)

Sulfite ions (SO_3^{2-}):

- Acidify with dilute hydrochloric acid
- Warm gently

Positive result: Sulfur dioxide gas (SO_2) produced. SO_2 decolourises filter paper which has been dipped into purple acidified potassium manganate(VII).

Chemical Tests: Cations

Describe how to do flame tests.
State the resulting flame colours for the following:

Lithium ions	*Sodium ions*
Potassium ions	*Copper(II) ions*

Describe how to test for the following ions:

Copper(II)
Iron(II)
Iron(III)
Chromium(III)
Calcium
Zinc
Aluminium

Describe how to test for *Ammonium* ions.

Flame Tests for Metal ions:

- Dip a loop of **nichrome wire** in hydrochloric acid
- Dip the loop of wire in the unknown compound
- Hold the loop of wire in a hot Bunsen burner flame

Positive results (flame colours):

Lithium ions, Li^+	Red
Sodium ions, Na^+	Orange/yellow
Potassium ions, K^+	Lilac
Copper(II) ions, Cu^{2+}	Blue-green

Precipitation Tests for Metal ions:

- Add a few drops of sodium hydroxide or ammonia solution
- Use colour of precipitate formed to identify metal ions present
- Add **excess** sodium hydroxide or ammonia – certain precipitates will redissolve

Positive results:

	with Sodium Hydroxide	**with Ammonia**
Copper(II) ions, Cu^{2+}	Light blue precipitate Insoluble in excess	Light blue precipitate Soluble in excess *(gives dark blue solution)*
Iron(II) ions, Fe^{2+}	Green precipitate Insoluble in excess	Green precipitate Insoluble in excess
Iron(III) ions, Fe^{3+}	Red-brown precipitate Insoluble in excess	Red-brown precipitate Insoluble in excess
Chromium(III) ions, Cr^{3+}	Green precipitate Soluble in excess	Grey-green precipitate Insoluble in excess
Calcium ions, Ca^{2+}	White precipitate Insoluble in excess	No precipitate, or very slight white precipitate
Zinc ions, Zn^{2+}	White precipitate Soluble in excess	White precipitate Soluble in excess
Aluminium ions, Al^{3+}	White precipitate Soluble in excess	White precipitate Insoluble in excess

Ammonium ions (NH_4^+):

- Add sodium hydroxide solution
- Warm gently

Positive result: Ammonia gas released, which turns damp red litmus paper **blue**

$$NH^+_{4\,(aq)} + OH^-_{(aq)} \longrightarrow NH_{3\,(g)} + H_2O_{(l)}$$

Acids and Bases

Define *Acid*, *Base* and *Alkali*.

Which ions are responsible for making a solution acidic or alkaline?

Name 4 common acids and 4 common bases.
Give their formulae.

Explain the difference between:
Strong Acids, Weak Acids, Strong Bases & *Weak Bases*

What are the hazards of working with acids and bases?

Definitions:

*Acid: A proton (H^+) donor

*Base: A proton (H^+) acceptor

Alkali: A soluble base

Ions:

Acidic solutions contain H^+ ions (hydrogen ions)

Alkaline solutions contain OH^- ions (hydroxide ions)

Common Acids:

Hydrochloric acid	HCl
Sulfuric acid	H_2SO_4
Nitric acid	HNO_3
Ethanoic acid	CH_3COOH

Common Bases:

Sodium Hydroxide	$NaOH$
Potassium Hydroxide	KOH
Calcium Hydroxide	$Ca(OH)_2$
Ammonia	NH_3

Strong Acids and Bases:

- **Strong acids dissociate completely** into ions when dissolved in water
  e.g. hydrochloric acid, sulfuric acid, nitric acid

$$HCl_{(aq)} \xrightarrow{100\%} H^+_{(aq)} + Cl^-_{(aq)}$$

- **Strong bases dissociate completely** into ions when dissolved in water
  e.g. sodium hydroxide, potassium hydroxide, calcium hydroxide

$$NaOH_{(aq)} \xrightarrow{100\%} Na^+_{(aq)} + OH^-_{(aq)}$$

Weak Acids and Bases:

- **Weak acids do not dissociate completely** into ions when dissolved in water
  e.g. ethanoic acid

$$CH_3COOH_{(aq)} \underset{}{\overset{\text{much less than 100\%}}{\rightleftharpoons}} H^+_{(aq)} + CH_3COO^-_{(aq)}$$

- **Weak bases do not dissociate completely** into ions when dissolved in water
  e.g. ammonia solution

$$NH_{3\,(aq)} + H_2O_{(l)} \overset{\text{much less than 100\%}}{\rightleftharpoons} NH^+_{4\,(aq)} + OH^-_{(aq)}$$

Hazards:

Strong acids and bases are corrosive (they can eat away metals, skin and cloth)

Reactions of Acids & Bases

Give equations for the reactions of acids with:

Bases

Metals

Carbonates

Give an equation for the reaction of a base with an ammonium salt.

Give an equation to show how an insoluble salt can be produced.

Which salts are soluble?

Acid + Base:

Acid + Base ⟶ Salt + Water

e.g. $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

e.g. $H_2SO_{4\,(aq)} + CuO_{(s)} \longrightarrow CuSO_{4\,(aq)} + H_2O_{(l)}$

Acid + Metal:

Acid + Metal ⟶ Salt + Hydrogen

e.g. $Mg_{(s)} + H_2SO_{4\,(aq)} \longrightarrow MgSO_{4\,(aq)} + H_{2\,(g)}$

Metals which form insoluble salts will not usually react with acids because a layer of the salt forms on the surface of the metal, preventing any more metal from reacting

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Acid + Carbonate:

Acid + Carbonate ⟶ Salt + Water + Carbon Dioxide

e.g. $CaCO_{3\,(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$

Base + Ammonium Salt:

Base + Ammonium Salt $\xrightarrow{\text{Heat}}$ Salt + Water + Ammonia

e.g. $Ca(OH)_{2\,(s)} + 2NH_4Cl_{(s)} \xrightarrow{\text{Heat}} CaCl_{2\,(s)} + 2H_2O_{(l)} + 2NH_{3\,(g)}$

This reaction is used to make ammonia in the lab

Making Insoluble Salts:

Soluble Salt + Soluble Salt $\xrightarrow{\text{Precipitation}}$ Insoluble Salt + Soluble Salt

e.g. $BaCl_{2\,(aq)} + MgSO_{4\,(aq)} \longrightarrow BaSO_{4\,(s)} + MgCl_{2\,(aq)}$

Soluble Salts

- All sodium, potassium & ammonium salts
- All nitrates
- Chlorides ***except*** silver & lead chlorides
- Sulfates ***except*** calcium, barium & lead sulfates

All other salts are insoluble

pH & Neutralisation

Explain the pH Scale.
List 3 indicators with their colours in acidic, neutral and alkaline solutions.

What is ***Neutralisation***? Give an ionic equation.

Give an agricultural use for neutralisation.
Which bases could you use for this?

Oxides

Explain what is meant by
Acidic Oxides
Basic Oxides
Amphoteric Oxides
Neutral Oxides.
Give examples of each.

pH Scale:

A measure of acidity or alkalinity on a scale of 1 to 14:

1 = most acidic 7 = neutral 14 = most alkaline

pH values can be determined using **Universal Indicator** *(UI)*

Indicators:

Indicator	Acid	Neutral	Alkali
Universal Indicator	Red – Orange – Yellow	Green	Blue – Purple
Litmus	Red	Purple	Blue
Methyl Orange	Red	Orange	Yellow

Neutralisation:

Neutralisation occurs when acids and bases react together:

$$\text{Hydrogen ions} + \text{Hydroxide ions} \longrightarrow \text{Water}$$

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

Reducing Soil Acidity:

Acidic soil can be neutralised with a base, e.g.

- crushed limestone (calcium carbonate, $CaCO_3$)
- lime (calcium oxide, CaO)
- slaked lime (calcium hydroxide, $Ca(OH)_2$)

Crops do not grow well if soil is too acidic

Oxides:

Type of Oxide	Properties	Examples
Acidic Oxides	React with bases (dissolve in water to produce acids)	Most **non-metal oxides**, e.g. • Carbon Dioxide, CO_2 • Sulfur Dioxide, SO_2 • Phosphorus Pentoxide, P_2O_5
Basic Oxides	React with acids	**Metal oxides**, e.g. • Magnesium Oxide, MgO • Iron(III) Oxide, Fe_2O_3 • Copper(II) Oxide, CuO
Amphoteric Oxides	React with both acids and bases	• Aluminium Oxide, Al_2O_3
Neutral Oxides	Do not react with acids or bases	• Carbon Monoxide, CO • Dinitrogen Oxide, N_2O

Limestone

Give the chemical name and formula of *Limestone*.

Give equations for the *Thermal Decomposition of Limestone* and the *Slaking of Lime*.

Give uses for *Limestone*, *Lime* and *Slaked Lime*.

Name: Calcium carbonate

Formula: $CaCO_3$

Thermal Decomposition of Limestone:

$$\text{Calcium Carbonate} \xrightarrow{\text{Heat}} \text{Calcium Oxide} + \text{Carbon Dioxide}$$

$$CaCO_{3\,(s)} \xrightarrow{\text{Heat}} CaO_{(s)} + CO_{2\,(g)}$$

Calcium Oxide, CaO, is referred to as ***Lime***

Slaking Lime:

$$\text{Calcium Oxide} + \text{Water} \longrightarrow \text{Calcium Hydroxide}$$

$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2\,(s)}$$

Only a small amount of water is added, so the calcium hydroxide does not dissolve.

Solid Calcium Hydroxide is referred to as ***Slaked Lime****.*

Uses:

Limestone:

- Making iron (extracting it from iron ore)
- Making cement

Lime /Slaked Lime:

- Neutralising acidic soil (to make it better for growing plants)
- Neutralising acidic industrial waste products (e.g. flue gas desulfurisation)