



Aitchison College Academic Resources.

IGCSE CHEMISTRY (0620) RESOURCES

This document includes complete notes of each unit with unsolved and solved past paper questions. Don't forget to go through it and we hope this document will help you preparing for your exams.
Best of Luck!



Contents

Paper Pattern:.....	3
Specimen Papers:.....	3
States of Matter:.....	4
Atoms , Elements and Compounds.....	9
Stoichiometry:	16
Electrochemistry:	24
The Periodic Table:	34
Chemical Energetics	41
Chemical Bonding	47
Rate of Reaction	57
Redox Reactions.....	67
Chemistry of the Environment.....	70
Metals	80
Acids Bases And Salts	89
Organic Chemistry	100
Identification of Ions And Gasses	120
Chemical Equilibrium	125



Paper Pattern:

Extended candidates take Paper 2 and Paper 4. The questions are based on the Core and Supplement subject content:

Paper 2: Multiple Choice (Extended)

45 minutes
40 marks 30%
40 four-option multiple-choice questions
Externally assessed

Paper 4: Theory (Extended)

1 hour 15 minutes
80 marks 50%
Short-answer and structured questions
Externally assessed

Paper 6: Alternative to Practical

1 hour
40 marks 20%
Questions will be based on the experimental skills in Section 4
Externally assessed

Specimen Papers:

[Specimen Paper 2](#)

[Specimen Paper 4](#)

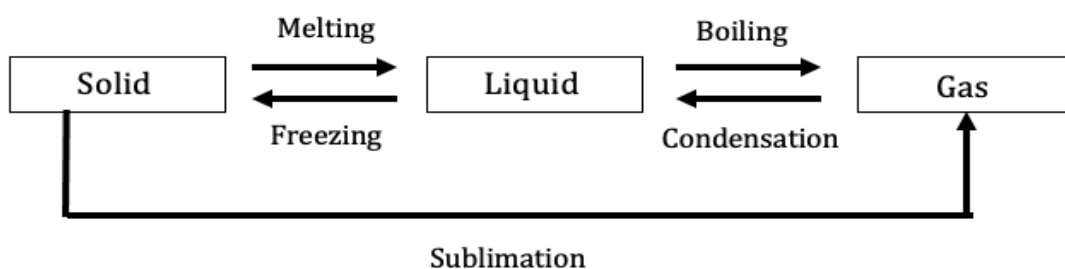
[Specimen Paper 6](#)



(i) States of Matter:

Solids, Liquids, Gases and their Changes of State

- There are 3 states of matter, Solid, Liquid, and Gas:



Process	Type of Reaction
Melting	Endothermic
Boiling/Evaporation	Endothermic
Condensation	Exothermic
Freezing	Exothermic
Sublimation	Endothermic

- Melting, Boiling/Evaporation, and Sublimation are all endothermic reactions as they are absorbing heat energy to break their bonds.
- Similarly, Freezing and Condensation are exothermic reactions as heat energy is released due to the formation of their bonds.
- Pure elements/compounds have fixed melting and boiling points, however impure elements/compounds have a lower melting point, and a higher boiling point, e.g.



	Melting Point	Boiling Point
Pure Water	0	100
Impure Water	-2	102

Diffusion and Evaporation

- Diffusion is when liquids or gasses mix due to their random movement.
- In diffusion, particles of the liquid/gas move from an area of higher concentration to an area of lower concentration.
- There are two factors that affect the rate of diffusion
 1. Temperature: The higher the temperature, the faster particles will diffuse and vice versa.
 2. Mass of particles: The larger a particle's Relative Atomic/Molecular mass is, the slower it will diffuse and vice versa.

For example, ammonia (NH_3) will diffuse faster than hydrogen chloride (HCl) as the Relative Molecular Mass (M_r) of is 17 and the M_r of HCl 36.5

- Evaporation is when a liquid changes its state from a gas, but at any given temperature. Unlike boiling, evaporation does not require a fixed temperature to take place.
- Evaporation only occurs at the surface of the liquid while boiling occurs throughout the liquid.



Physical and Chemical Changes

- A physical change involves the change of state of a substance and produces no new substance, e.g. melting of ice or boiling of water.
- A chemical change is when two or more substances react together to form a new compound with different properties. Chemical changes occur on an atomic level, e.g. combustion, decomposition, neutralization.

Kinetic Particle Theory

The Kinetic Particle theory of matter states that matter consists of microscopic particles , atoms at the smallest level , in constant random motion (brownian motion)

Solids :

Fixed volume

Fixed shape

Atoms vibrate in position

Regular, fixed arrangement

Particles are tightly packed together

Generally more dense than liquids and gases

Liquids :

Fixed Volume

Adapt to the shape of container

Slightly compressible

Particles move and slide past one another

Generally less dense than solids but more than gases

Gases :

No fixed volume

Take shape of container and fill it

Very low density



Easily compressible due to lots of space between particles

State Changes in terms of Kinetic Theory

When Substances are heated, they absorb it (thermal energy) and convert it into kinetic energy
For example, heating a solid causes it to vibrate more. As you increase the temperature, the particles vibrate so much that the structure breaks and the solid melts, turning into a liquid

When the particles gain enough energy to overcome the intermolecular forces of attraction, the structure breaks and the state of the substance changes

When a liquid is heated, the particles with the highest energy gain some more to overcome the intermolecular forces of attraction and evaporate .

When the substance reaches boiling point, all the particles have enough energy to overcome the forces of attraction and escape, called boiling

A heating curve shows the increase in temperature with respect to time and changing of state of the substance

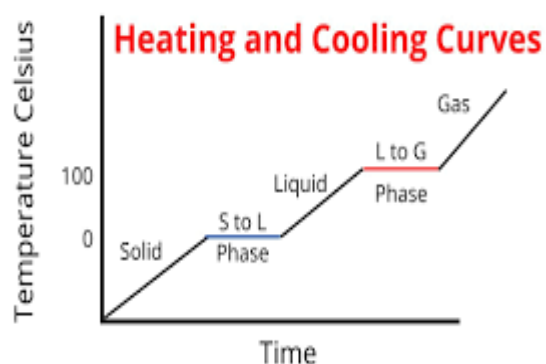
A cooling curve shows the decrease in temperature with respect to time and the state change of the substance

Important : when a substance is changing state, the temperature does NOT increase because :

For a heating curve, the particles are absorbing all that energy to overcome the forces of attraction

For a cooling curve, particles are releasing energy as the temperature decreases to maintain it

Heating curve





Effects of Temperature and Pressure On the Volume of Gases

Temperature:

As temperature increases, the kinetic energy of gas particles increases.

Particles move faster and spread further apart.

This increases collisions and the space the gas occupies (in an open system).

In a closed container, increasing temperature increases pressure (since volume can't expand).

More collisions with the container walls increase the force per unit area (pressure).

Pressure (Temperature constant):

Increasing pressure on a gas reduces its volume (compression).

Higher pressure forces particles closer together, decreasing the space they occupy.

Collisions become more frequent but with less space to move, lowering the volume.

This relationship follows Boyle's Law: pressure is inversely proportional to volume at constant temperature ($P \propto 1/V$).

Past Paper Questions P2 :

Which of the following processes cause the distance between particles to increase the most ?

Melting

Freezing

Vaporising

Condensing

A student left a bottle of Ammonia open in a corner of the laboratory. After a few minutes, the entire lab smells of Ammonia . Which process has occurred ?

Dispersion

Diffusion

Evaporation

Distillation

Which of the following statements is correct ?

A specific mass of gas has a fixed volume at RTP

A specific mass of liquid has a fixed volume at RTP

25cm³ of gas contains more particles than 25cm³ of liquid



Liquid particles placed inside sealed containers have fixed positions

In which substance are the particles close together and slowly moving past each other ?

Ice

Air

Water

Steam

Which of the following correctly describes the movement of particles in a solid according to the kinetic particle theory?

Particles move randomly and are far apart.

Particles vibrate in fixed positions

Particles slide past each other freely

Particles move rapidly in all directions.

P4 :

Q1 : Give the name of the process when a gas turns into a liquid

Q2 : Give the name of the process when a solid turns directly into a gas

Q3 : Why does the temperature not increase during state change even though the substance is being heated

Q4 : Compare the movement and arrangement of particles in solid nitrogen to gaseous nitrogen

Q5 : A sealed container contains nitrogen gas. Using the Kinetic theory, explain why the pressure inside the container increases when the temperature is increased

Chapter 2 : Atoms , Elements and Compounds

Atoms , Elements and Compounds

- All matter around us is made of tiny particles (atoms and molecules).



- An atom is the smallest particle of an element, while two or more atoms that are bonded together form a molecule.
- An element is a pure substance that cannot be broken down any further and is made up of only one type of atom. e.g. Nitrogen (N_2), Carbon (C).
- A mixture is when two or more elements are physically joined together, e.g. soil, cement and can be separated via physical means, for e.g distillation, evaporation, filtration .
- A compound is when two or more elements are chemically joined together, e.g. Sodium chloride (NaCl), Water (H_2O).

The subatomic particles inside an atom

- Inside every atom exist three types of subatomic particles:

1. Protons
2. Neutrons
3. Electrons

- Protons:

Protons (along with Neutrons) are found inside the nucleus of an atom. One proton has a relative mass of 1, and a positive charge of +1.

Each atom has the same number of protons as it does electrons.

- Neutrons:

Neutrons (along with Protons) are found inside the nucleus of an atom. One neutron has a relative mass of 1, and a neutral charge of 0.

- Electrons:

Electrons can be found orbiting the nucleus of an atom in electron shells. One electron has a relative mass of $1/1840$ where 1 is the mass of a proton/neutron, and a negative charge of -1.

Each atom has the same number of electrons as it does protons.

Each electron shell has a specific capacity of electrons it can hold. The first electron shell can only hold 2 electrons, while the rest can hold 8. e.g. a sodium atom () has 2 electrons in its first shell, 8 electrons in its second shell, and 1 electron in its first shell, giving it an electronic configuration of 2,8,1.

When the first electron shell is full , it is known as a duplet (2)

When any other electron shell is full , it is known as an octet (8)

The Electronic configuration of an element is written like this :

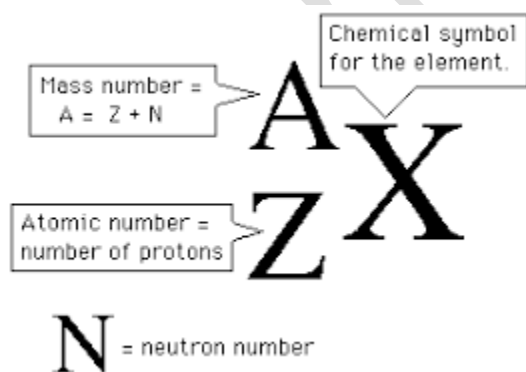


electrons in shell 1 . electrons in shell 2 . electrons in shell 3 etc until the valence shell

Subatomic Particle	Relative Mass	Charge
Proton	1	+1
Neutron	1	0
Electron	1/1840	-1

The Proton Number and Atomic Mass of an element

- The proton number of an element, also known as the atomic number of an element, is the number of protons one atom of an element contains.
- The atomic mass of an element, also known as the mass number or nucleon number of an element, is the number of protons and neutrons contained in the nucleus of one atom of the element.



- The relative atomic mass of an element is the average mass of an atom relative to the mass of a carbon-12 atom where a carbon 12 atom has a mass of exactly 12 units



- Between the proton number and atomic mass, the proton number is always the smaller number of the two, and the atomic mass is always the larger number of the two. e.g. in silicon (Si), its proton number is 14 and its atomic mass is 28.
- The periodic table is arranged in ascending order of the proton numbers of every element.

Ions

- Metal atoms can lose their valence shell electrons to form positive ions, known as cations, e.g. an Na atom loses its valence shell electron to form an ion.
- Similarly, non-metal atoms can gain electrons to complete their valence shell, which form negative ions, known as anions, e.g. an O atom gains 2 electrons to form an ion.
- The group numbers also signify the charge of their elements. Ions of elements in Groups I, II, and III, have a charge of 1+, 2+, and 3+ respectively. Atoms of elements in Group IV can both gain and lose electrons, while ions of elements in Groups V, VI, and VII have a charge of 3-, 2-, and 1- respectively. Ions of elements in Group VIII have no charge.

Isotopes

- Isotopes are atoms of the same element but with different nucleon numbers. They have the same number of protons and electrons, but they have a different number of neutrons, e.g. U-235 and U-239 .
- Isotopes also have the same chemical properties (as they have the same number of valence shell electrons), but they have different physical properties due to their different atomic masses.
- Often, questions will ask you to determine the relative atomic mass of an element by giving you two or more isotopes and their abundance.

Examples of Isotopes :

Carbon-12 (^{12}C) and Carbon-14 (^{14}C)

Carbon-12: Has 6 protons and 6 neutrons. It is the most common isotope of carbon.

Carbon-14: Has 6 protons and 8 neutrons. It is used in radiocarbon dating.

Hydrogen Isotopes

Protium (^1H): The most common form of hydrogen, with 1 proton and no neutrons.



Deuterium (^2H): Has 1 proton and 1 neutron. It's stable and used in some scientific applications.

Tritium (^3H): Has 1 proton and 2 neutrons. It's radioactive and used in nuclear fusion research.

Percentage Abundance and Relative atomic mass (calculation)

To calculate the relative atomic mass of an element with multiple isotopes, we multiply the relative atomic mass of one isotope by its percentage abundance and add this for all isotopes, then divide by 100

$$((\text{Isotopic mass 1} \times \% \text{ Isotope 1}) + (\text{Isotopic mass 2} \times \% \text{ Isotope 2})) / 100$$

Q. The table shows the relative abundance of the two naturally occurring atoms of lithium.

Atom	6	7
Relative abundance	10%	90%

Calculate the relative atomic mass of lithium to one decimal place.

$$\text{Ans : } (6 \times 10) + (7 \times 90) = 690 / 100 = 6.9$$

• Here is a list of important polyatomic ions and their charges which you must know.

Polyatomic Ion	Charge
Sulfate	-2
Carbonate	-2
Nitrate	-1
Hydroxide	-1
Phosphate	-3
Ammonium	+1



Past Paper Questions

Paper 2 :

The Atomic Number of element Q is 9 and its mass number is 19. In which Group of the Periodic table Should element Q be placed ?

Group VIII

Group I

Group VII

Group II

Which of the following Statements is Correct:

Elements contain different atoms

Compounds consist of elements that are not chemically combined

Mixtures consist of elements that are chemically combined

A mixture can contain elements and compounds

The X^+ ion has 16 electrons and a nucleon number of 37. What is the correct group number for element X

VII

VIII

III

I

Iridium has 2 Isotopes : ^{191}Ir and ^{193}Ir . The Percentage Abundances are as follows : Iridium-191 : 62.7% and Iridium- 193 : 37.3% . What is the relative atomic mass of iridium:

191.2

191.6

191.7

192.5

Element X contains 4 electrons in its outer shell and 125 neutrons in its nucleus. What is element X ?

Beryllium

Carbon



Tin

Lead

Paper 4 :

Q1 : Define the term Isotope.

Q2 : Magnesium , Calcium and Strontium are Group II metals. Write down the electronic configuration of Calcium

Q3 : What is meant by the term element?

Q4 : Carbon has 3 isotopes : Carbon-12 , Carbon-13 and Carbon-14. What are the similarities and differences between all 3 isotopes?

Q5 : How many protons, neutrons and electrons are there in an ion of Copper ⁶⁴



(v) Stoichiometry:

All about Formulae

- Atoms combine together in fixed ratios that give them a full outer shell of electrons. Chemical formulae of compounds tell you these fixed ratios. For example, the formula MgCl_2 shows us that 1 Magnesium atom reacts with 2 Chlorine atoms to form Magnesium Chloride, and consequently, a stable electronic configuration.
- Constructing chemical formulae: The 'Criss-Cross' Method
- Refer back to Topic (iii) – The Periodic Table. You will find that the number of electrons in the valence shell of an atom of an element depends on its group number. Let's solve an example with this formula constructing method

Q. Construct the formula of Aluminium Oxide

Step 1: Derive the elements in the compound.

We can derive that the elements contained in this compound will be Aluminium and Oxygen.

Step 2: Note their charges.

Aluminum is in Group III, therefore its charge is $3+$. Similarly, Oxygen is in Group VI, therefore its charge is $2-$

Step 3: Swap their charges.

- Empirical and Molecular Formulae
- Empirical Formulae are the simplest ratios of atoms in a compound, while Molecular Formulae are the actual ratios of atoms in a compound. For example, the molecular formula of ethanoic acid is $\text{C}_2\text{H}_4\text{O}_2$, thus the ratio of Carbon atoms : Hydrogen atoms : Oxygen atoms is $2:4:2$. However, the empirical formula of ethanoic acid is CH_2O . The previous ratio of $2:4:2$ has been simplified to $1:2:1$.

Q. Compound S has the empirical formula CH_2O and a relative molecular mass of 60.



Calculate the molecular formula of compound S.

Step 1: Calculate the E.F.M (Empirical Formula Mass).

$$= 12 + 2(1) + 16$$

$$= 28 + 2$$

$$= 30$$

Step 2: Divide its Molecular Mass by its E.F.M.

$$= 60/30$$

$$= 2$$

Step 3: Multiply the answer in Step 2 with its Empirical Formula.

$$= 2 (\text{CH}_2\text{O})$$

$$= \text{C}_2\text{H}_4\text{O}_2$$

• Therefore the molecular formula for this compound is $\text{C}_2\text{H}_4\text{O}_2$

• Relative molecular (formula) mass

• The symbol for the relative molecular mass is M_r and it refers to the total mass of the molecule.

• To calculate the M_r of a substance, you have to add up the relative atomic masses of all the atoms present in the formula.

The Mole & Avogadro's Constant

• Chemical amounts are measured in moles.

• The mole, symbol : mol, is the SI unit of amount of substance. It is often denoted by n.

• One mole of a substance contains the same number of the stated particles, atoms, molecules, or ions as one mole of any other substance

• One mole contains 6.02×10^{23} particles (atoms, ions or molecules). This number is known as the Avogadro Constant. For example,

• One mole of sodium (Na) contains 6.02×10^{23} atoms of sodium.

• One mole of hydrogen gas (H_2) contains 6.02×10^{23} molecules of hydrogen.

• One mole of sodium chloride (NaCl) contains 6.02×10^{23} formula units of sodium chloride.

• The mass of 1 mole of a substance is known as the molar mass.



For an element, it is the same as the relative atomic mass written in grams.

- For a compound it is the same as the relative formula mass or relative molecular mass in grams.

Moles – Formulae and Questions

- The following are the formulas concerning moles that you need to know:

$$(i) n = \frac{\text{mass (in grams)}}{A_r/M_r}$$

$$(ii) n = \frac{\text{volume (in dm}^3\text{)}}{24\text{dm}^3} \quad (\text{only for gases at r.t.p})$$

$$(iii) n = CV \quad (\text{where } C = \text{concentration in mol/dm}^3 \text{ and } V = \text{volume in dm}^3)$$

When sodium hydrogencarbonate is heated it forms three products.



Calculate the volume of carbon dioxide formed at room temperature and pressure when 12.6 g of NaHCO is heated using the following steps:

- Determine the molecular mass of one mole of NaHCO_3
 $= 23+1+12+3(16)$
 $= 36+48 = 84$
- Calculate the number of moles of NaHCO_3 used
- Determine the number of moles of carbon dioxide formed

(Look at the ratios in the balanced reaction given in the question)

$$= \text{NaHCO}_3 : \text{CO}_2$$

$$= 2 : 1$$

$$= 0.015 : x$$

$$2x = 0.015$$

$$x = 0.075 \text{ mol. of CO}_2$$

- Calculate the volume of carbon dioxide formed at room temperature and pressure.

$$n = \frac{\text{volume}}{24\text{dm}^3}$$

$$\text{volume} = n \times 24\text{dm}^3$$

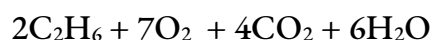
$$\text{volume} = 0.075 \times 24\text{dm}^3$$

$$\text{volume} = 1.8 \text{ dm}^3$$



- Limiting and Excess Reagents
- Often, questions may ask you to identify the limiting or excess reagent in a given equation.
- Limiting Reagent: The limiting reagent in a chemical reaction is the reactant that will be consumed completely. Once there is no more of that reactant, the reaction cannot proceed. Therefore it limits the reaction from continuing.
- Excess Reagent: The excess reagent is the reactant that could keep reacting if the other had not been consumed.
- Questions will often give you the mass (and maybe formulae) of the reactants, or will give you the concentration and volume of the reactants. You must then calculate their moles and then compare them with each other via their ratios in the balanced equation. If the number of moles you get from the ratio method is more than the number of moles of that specific reactant that you actually have, then that reactant is the limiting reagent.

Q. Use the following BALANCED equation.



If 15g of C_2H_6 reacts with 45g of O_2 , how many grams of water will be produced?

Step 1: Calculate the moles of the reactants.

$$n = \frac{\text{mass}}{M_r}$$

$$n = \frac{15}{2(12) + 6(1)}$$

$$n = 0.5 \text{ mol. of } \text{C}_2\text{H}_6$$

$$n = \frac{\text{mass}}{M_r}$$

$$n = \frac{45}{2(16)}$$

$$n = 1.4063 \text{ mol. of } \text{O}_2$$

Step 2: Compare the reactants using the ratios from the balanced equation

$$= 2 : 7$$

$$= \text{C}_2\text{H}_6 : \text{O}_2$$

$$= 2 : 7$$

$$= 0.5 : x$$

$$2x = 3.5$$

$$x = 1.75 \text{ mol. of } \text{O}_2$$



$$= 0.5 : x$$

$$2x = 3.5$$

$$x = 1.75$$

• This means that to completely react with 0.5 mol. of , there needs to be 1.75 . However, as we calculated previously, there is only 1.4063 present, which is less than the required amount to completely react. Thus, we can deduce is the limiting reagent in this question, and therefore is the excess reagent. This means that we must compare with to find the number of grams of water produced.

Step 3: Calculate the moles of water formed.

$$= \quad = \quad O_2 : H_2O$$

$$= 1.4 : 7 : 6$$

$$7x = 1.4063 : x$$

$$x = 1.4063 / 7 = 0.2009$$

Step 4: $x = 1.2054 \text{ mol. of } H_2O$

$$\text{mass} = n \times M_r$$

$$\text{mass} = 1.2054 \times 18$$

$$\text{mass} = 21.7 \text{ g. of } H_2O$$

Percentage Yield

- Yield is the term used to describe the amount of product you get from a reaction.
- In practice, you never get 100% yield in a chemical process for several reasons:
- Some reactants may be left behind in the equipment.
- The reaction may be reversible and in these reactions a high yield is never possible as the products are continually turning back into the reactants.
- Some products may also be lost during separation and purification stages such as filtration or distillation.
- There may be side reactions occurring where a substance reacts with a gas in the air or an impurity in one of the reactants.
- Products can also be lost during transfer from one container to another.

Actual & Theoretical Yield



The actual yield is the recorded amount of product obtained.

- The theoretical yield is the amount of product that would be obtained under perfect, practical, and chemical conditions .
- The Theoretical yield is calculated from the balanced equation and the reacting masses.
- The percentage yield compares the actual yield to the theoretical yield.
- For economic reasons, the objective of every chemical producing company is to have as high a percentage yield as possible to increase profits and reduce costs and waste.
- To calculate percentage yield, the following equation is used:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Percentage Purity

- A pure substance has nothing else mixed with it.
- Often, the product you are trying to obtain may become contaminated with unwanted substances such as unreacted reactants, catalysts and other impurities.
- To calculate percentage purity, the following equation is used:

$$\text{percentage purity} = \frac{\text{mass of pure substance}}{\text{total mass of substance}} \times 100\%$$

- The purity) are always given in the question.

Questions

P2 :

1. When 2 g of hydrogen gas reacts completely with oxygen, what mass of water is formed?
A) 9 g
B) 18 g
C) 36 g
D) 4.5 g



2. Which of the following is the correct ratio of moles when magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas?
- A) 1:1
 - B) 1:2
 - C) 2:1
 - D) 1:3
3. What is the number of moles of carbon dioxide produced when 44 g of carbon dioxide is formed from the complete combustion of carbon?
(Relative atomic mass: C = 12, O = 16)
- A) 0.5 mol
 - B) 1 mol
 - C) 2 mol
 - D) 0.25 mol
4. 50 cm³ of 0.5 mol/dm³ hydrochloric acid reacts completely with 25 cm³ of sodium hydroxide. What is the concentration of the sodium hydroxide solution?
- A) 0.5 mol/dm³
 - B) 1 mol/dm³
 - C) 2 mol/dm³
 - D) 0.25 mol/dm³
5. How many moles of sodium ions are present in 58.5 g of sodium chloride?
- A) 0.5 mol
 - B) 1 mol
 - C) 2 mol
 - D) 0.25 mol



P4:

1. Calcium carbonate reacts with hydrochloric acid to form calcium chloride, carbon dioxide, and water. Calculate the mass of calcium chloride produced when 10 g of calcium carbonate reacts completely with excess hydrochloric acid.
2. Magnesium burns in oxygen to form magnesium oxide. If 4.8 g of magnesium reacts completely, calculate the mass of magnesium oxide produced.
3. Sulfur reacts with oxygen to form sulfur dioxide. What volume of oxygen (at room temperature and pressure) is needed to react completely with 32 g of sulfur?
4. Sodium hydroxide reacts with sulfuric acid to form sodium sulfate and water. 25 cm³ of 1.0 mol/dm³ sodium hydroxide solution reacts completely with 50 cm³ of sulfuric acid. Calculate the concentration of the sulfuric acid solution.
5. Ethene (C₂H₄) reacts with oxygen to form carbon dioxide and water. A 0.05 mol sample of ethene reacts with excess oxygen. Calculate the mass of carbon dioxide produced in this reaction.



(iii) Electrochemistry:

Basic Principles of Electrochemistry

Basic Principles of Electrolysis

Electrolysis refers to the process in which a compound, typically an ionic substance, is decomposed by passing an electric current through it. The substance being decomposed can be either molten or in aqueous solution.

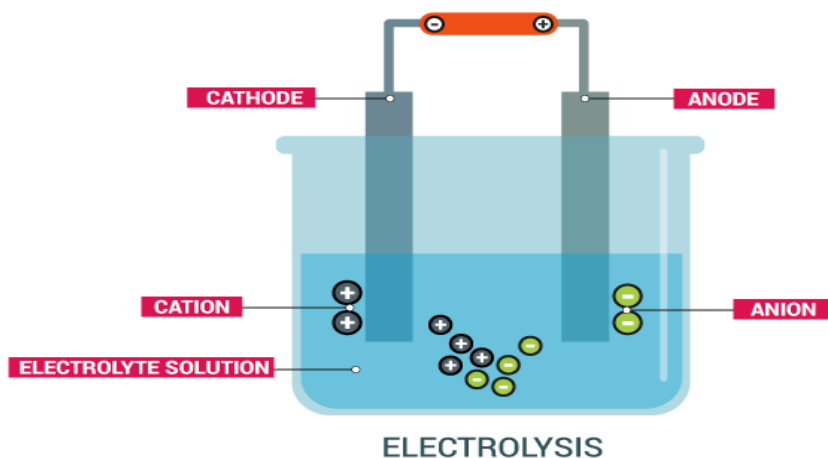
Electrolyte: The molten or aqueous ionic compound that undergoes electrolysis. It conducts electricity due to the presence of free-moving ions.

Electrodes: Conductors through which the electric current enters and leaves the electrolyte.

Anode (Positive Electrode): Electrons are lost here (oxidation occurs).

Cathode (Negative Electrode): Electrons are gained here (reduction occurs).

Inert Electrodes: Electrodes that do not react with the electrolyte. Graphite or platinum are commonly used.



Charge Transfer in Electrolysis:

At the Cathode (Negative Electrode):

Cations (positive ions) move towards the cathode and gain electrons, undergoing reduction.

The products may be metal deposition or hydrogen gas, depending on the reactivity of the ions.

At the Anode (Positive Electrode):

Anions (negative ions) move towards the anode and lose electrons, undergoing oxidation.

The products are typically non-metallic elements or oxygen gas.



Electrolysis of Molten Compounds

Ionic compounds decompose into their constituent elements when molten.

Example: Electrolysis of Lead(II) Bromide (PbBr_2):

Setup: Heat PbBr_2 until molten, then insert graphite electrodes.

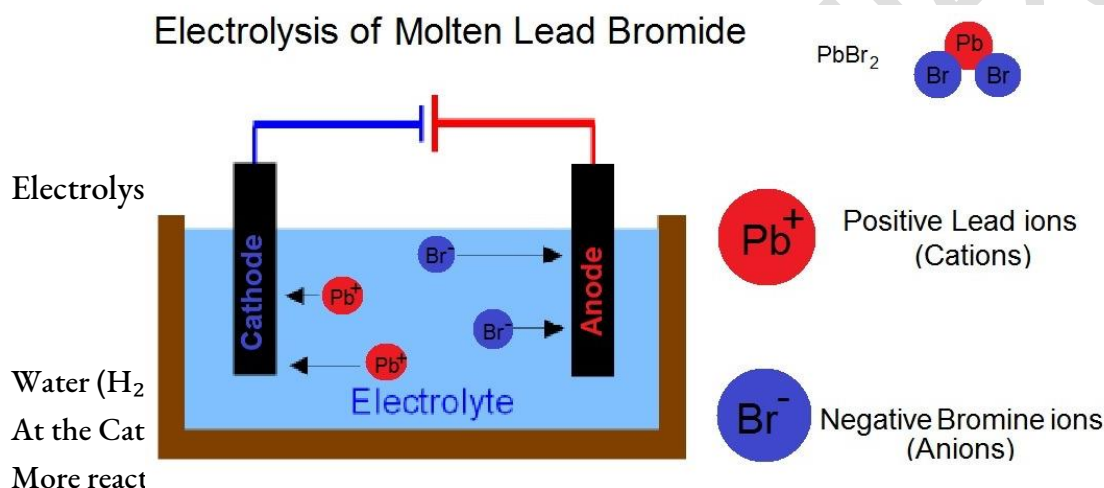
At Cathode: Lead (Pb^{2+}) ions gain electrons to form solid lead.

Half-equation: $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$

At Anode: Bromide (Br^-) ions lose electrons to form bromine gas.

Half-equation: $2\text{Br}^- \rightarrow \text{Br}_2(\text{g}) + 2\text{e}^-$

Overall Reaction: $\text{PbBr}_2(\text{l}) \rightarrow \text{Pb}(\text{s}) + \text{Br}_2(\text{g})$



Hydrogen gas (H_2) is produced if metal is more reactive than hydrogen.

Metal deposition occurs if metal is less reactive than hydrogen.

At the Anode:

Halide ions (Cl^- , Br^- , I^-) form halogen gas (Cl_2 , Br_2 , I_2).

Oxygen gas (O_2) is produced if halide ions are absent (from OH^- ions).

Electrolysis of Aqueous Sodium Chloride

Setup: Brine (NaCl solution) is electrolysed with inert electrodes (e.g., graphite or platinum).

At Cathode: Hydrogen ions (H^+) are reduced to form hydrogen gas.

Half-equation: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

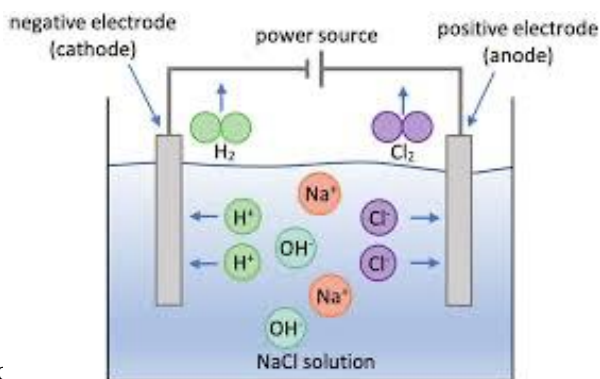
At Anode: Chloride ions (Cl^-) are oxidized to form chlorine gas.

Half-equation: $2\text{Cl}^- \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$

Overall Reaction: $2\text{NaCl}(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{Cl}_2(\text{g}) + \text{H}_2(\text{g}) + 2\text{NaOH}(\text{aq})$



Products: Chlorine gas, hydrogen gas, sodium hydroxide.



Electrolysis of the Aqueous

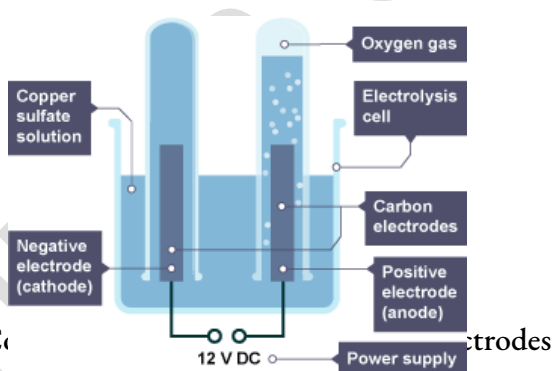
At Cathode: Copper ions (Cu^{2+}) are reduced to copper metal.

Half-equation: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu(s)}$

At Anode: Hydroxide ions (OH^-) are oxidized to form oxygen gas.

Half-equation: $4\text{OH}^- \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O} + 4\text{e}^-$

Products: Copper metal at cathode, oxygen gas at anode.



Electrolysis of the Aqueous Cu

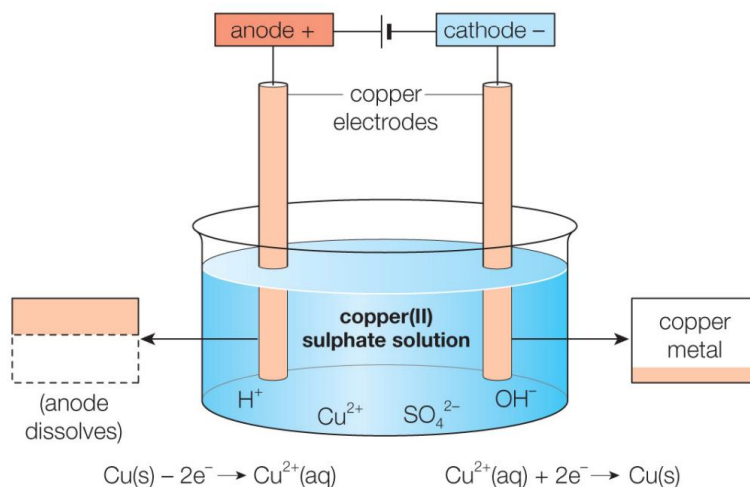
At the Anode: Copper metal is oxidized to copper ions (Cu^{2+}), which enter the solution.

Half-equation: $\text{Cu(s)} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$

At the Cathode: Copper ions (Cu^{2+}) are reduced to form copper metal on the cathode.

Half-equation: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu(s)}$

Outcome: The concentration of copper ions in the solution remains constant. Copper is transferred from the anode to the cathode.



Electrolysis of Dilute Sulfuric Acid

At Cathode: Hydrogen ions (H^+) are reduced to form hydrogen gas.

Half-equation: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

At Anode: Hydroxide ions (OH^-) are oxidized to form oxygen gas.

Half-equation: $4\text{OH}^- \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O} + 4\text{e}^-$

Products: Hydrogen gas at cathode, oxygen gas at anode.

Electrolysis of Water

Water is split into hydrogen and oxygen gases.

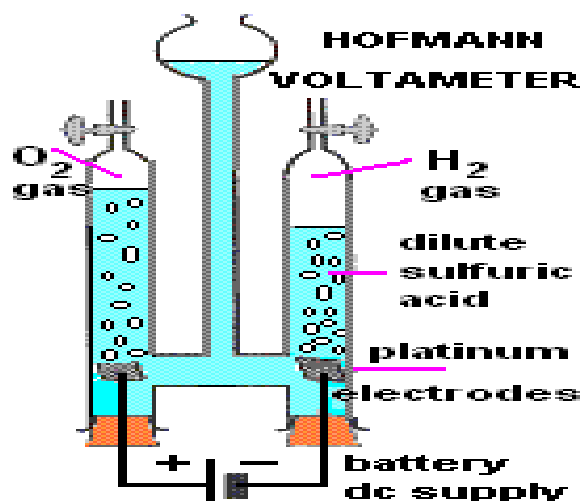
At Cathode: Hydrogen ions (H^+) are reduced to form hydrogen gas.

Half-equation: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

At Anode: Hydroxide ions (OH^-) are oxidized to form oxygen gas.

Half-equation: $4\text{OH}^- \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O} + 4\text{e}^-$

Overall Reaction: $2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$



Ionic Equations

Electrolysis is characterized by electron transfer.

Oxidation occurs at the anode (electron loss).

Reduction occurs at the cathode (electron gain).

Example: Electrolysis of water splits it into hydrogen and oxygen gases.

At the Cathode: $\text{H}^+ + \text{e}^- \rightarrow \text{H}_2(\text{g})$

At the Anode: $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+ + 4\text{e}^-$

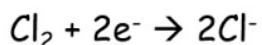
Metals:

All metallic ions are positively charged. If the metal is being produced during electrolysis, it will gain electrons to form the metal's atoms as the metal ions are being reduced. Number of electrons gained by a metal is determined by the charge value on the metal's ions. Following are a few examples:



Non-Metals:

All non-metallic ions are negatively charged. If a non-metal is produced during electrolysis, it will lose electrons to form non-metallic atoms as the non-metal ions are being reduced. Number of lost electrons corresponds to the charge on the non-metallic ion. Following is an example:



Hydrogen:

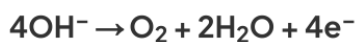
If hydrogen is being formed in the electrolysis, the following equation will be used:



Hydroxide:

In electrolysis, hydroxide ions are getting oxidized as they lose electrons to form oxygen and water.

For this, the following equation will be used for this:



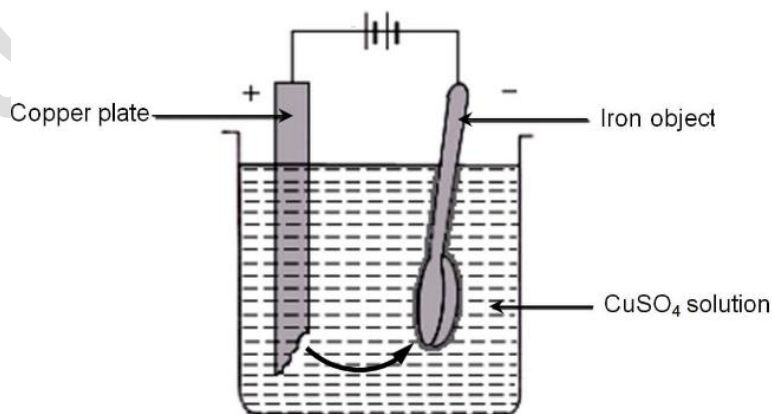
Electroplating

The process of electroplating involves applying a thin layer of one metal to the surface of another.

The object that has to be electroplated serves as the cathode in this operation (negative electrode).

The pure metal that will be deposited into the object's surface serves as the anode, or positive electrode. The electrolyte, which aids in the transfer of metal ions during electroplating, is a solution that contains a soluble salt of the metal utilized at the anode.

Example:



Copper atoms undergo oxidation at the anode, where they lose electrons to create copper ions that are released into the solution. At the same time, these copper ions reach the cathode and reduce to



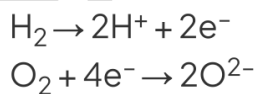
produce copper atoms by gaining electrons. The iron spoon is then covered with a layer of copper after these copper atoms are applied to its surface. Copper is successfully applied to the iron spoon by this procedure, giving it an effective protective layer.

Uses of Electroplating:

Metals can be made more resistant to corrosion and damage by electroplating them. For instance, galvanizing is applying a zinc coating to iron or steel to stop rust, and chromium and nickel plating is frequently used for protective purposes. Furthermore, electroplating enhances the look of metals. For example, silver electroplating gives flatware and jewellery a more appealing look.

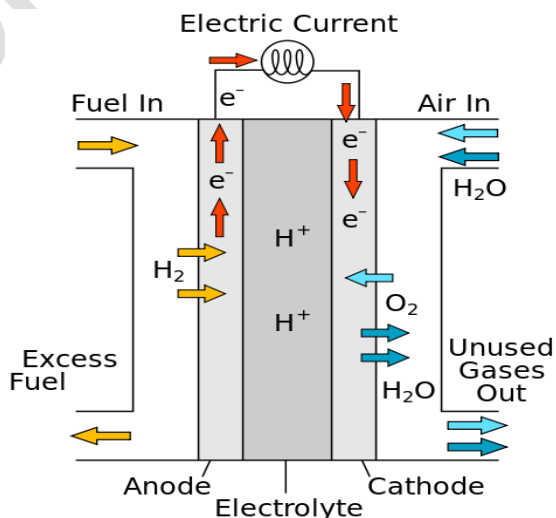
Hydrogen-Oxygen Fuel Cells

Any material that releases energy when burned is considered a fuel. In order to power specific vehicles, hydrogen is frequently employed as fuel in fuel cells and rocket engines. A fuel cell is an electrochemical device in which oxygen receives electrons at one electrode and a fuel, such as hydrogen, provides electrons at the other.



When hydrogen and oxygen are combined in a hydrogen-oxygen fuel cell, energy is released and water is produced as a byproduct. This process generates electricity. The hydrogen fuel cell's overall chemical reaction is:

Water = Hydrogen + Oxygen





In the above setup of a fuel cell, the oxygen is provided by the air that is entering and the fuel entering is the hydrogen. Water is the only waste product of the fuel cell. The movement of hydrogen, oxygen and electrons is shown in the above diagram.

Advantages and Disadvantages:

Advantages	Disadvantages
Process is renewable since the hydrogen can be obtained from the water.	Hydrogen is expensive and difficult to store as it is flammable and can easily explode.
More energy is released per kilogram as compared to petrol and diesel.	Hydrogen produced by combustion of fossil fuels releases carbon dioxide. Hydrogen produced by the electrolysis of water requires a very large amount of electricity.
It does not cause any pollution because water is the only product.	Fuel cells are badly affected by low temperatures which reduce their efficiency.
It does not contribute to noise pollution.	It is very expensive.
Since there are no moving parts, no power is lost in transmission.	Number of hydrogen filling stations across the countries is very low.

P2:

Q1. Concentrated aqueous sodium chloride and dilute sulfuric acid are both electrolysed using inert electrodes. Which row identifies the product at the cathode in each electrolysis?

	Aqueous Sodium Chloride	Dilute Sulfuric Acid
A	Hydrogen	Oxygen
B	Hydrogen	Hydrogen



C	Chlorine	Oxygen
D	Chlorine	Hydrogen

Q2. Which statement about hydrogen–oxygen fuel cells is correct?

Hydrogen is extracted from clean, dry air.

The only product is carbon dioxide.

The reaction is endothermic.

No toxic gasses are produced.

Q3. Which statement about electrolysis is correct?

Chemical energy is converted to electrical energy.

Electrons flow through the electrolyte.

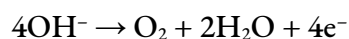
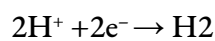
Ionic compounds are broken down.

Metals are formed at the positive electrode.

Q4. A metal object is electroplated with copper. One electrode is the metal object and the other electrode is copper. The electrolyte is aqueous copper(II) sulfate. Which row shows the ionic half-equation for the reaction at the anode and the observation of the electrolyte?

	Anode	Electrolyte
A	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	Blue colour fades
B	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	Blue colour does not change
C	$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$	Blue colour fades
D	$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$	Blue colour does not change

Q5. Aqueous copper(II) sulfate is electrolysed using copper electrodes. What is the half-equation for the reaction at the cathode?





P4:

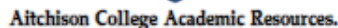
Q1. Write the ionic equation which represents a neutralization reaction between any acid and any alkali.

Q2. Name the process that is used to obtain water from aqueous sodium chloride.

Q3. State what is meant by the term electrolysis.

Q4. Dilute aqueous sodium fluoride undergoes electrolysis. Hydrogen is produced at the cathode. Write an ionic half-equation for the production of hydrogen at the cathode.

Q5. Name the process used to produce lead from molten lead(II) bromide.



The Arrangement of the Periodic Table

1																	2	3					
3 Li	4 Be																	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg																	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr						
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe						
55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn						
87 Fr	88 Ra																						

- 34



Each period number signifies the total number of electron shells an atom of the elements in that period have, e.g. all elements in Period 3 (Na, Mg, Al, Si, P, S, etc.) have a total of 3 electron shells. Similarly, all elements in Period 4 (K, Ca, Sc, Ti, V, Cr, etc.) have a total of 4 electron shells, and so on.

The Trends in Groups of the Periodic Table

- In the periodic table, you need to know about: Group I, Group II, Group VII, Group VIII, as well as the transition elements.

Group I (Alkali Metals):

The Group I elements include:

- Lithium (Li)
- Sodium (Na)
- Potassium (K)
- Rubidium (Rb)
- Caesium (Cs)
- Francium (Fr)

They are soft metals and have 1 electron in their outer-shell.

The Group I trends include:

- Melting/Boiling points decrease down the group.
- Density increases down the group.
- Reactivity increases down the group.

e.g. of reactions with water:

Observations of Group I metals reacting with water:

- The metal floats on the surface of the water and then dissolves.
- Effervescence/Bubbles of gas produced (produces popping sound with lighted splint).
- Colourless solution is formed which has a high pH value.

Group II (Alkaline Earth Metals):

The Group II elements include:

- Beryllium (Be)



- Magnesium (Mg)

Calcium (Ca)

- Strontium (Sr)
- Barium (Br)
- Radium (Ra)

They have 2 electrons in their outer shell.

The Group II trends include:

- Melting/Boiling points decrease down the group.
- Density increases down the group.
- Reactivity increases down the group.

e.g. of reactions with water:

(Mg is an exception, and reacts slow with water)

(Mg reacts fast with steam)

Group VII (Halogens):

The Group VII elements include:

- Fluorine (F)
- Chlorine (Cl)
- Bromine (Br)
- Iodine (I)
- Astatine (At)

They exist as diatomic gases in their free state and have 7 electrons in their outer-shell.

Appearance in Room Temperature and Pressure		
Element	Physical State	Colour
Fluorine (F)	Gas	Yellow
Chlorine (Cl)	Gas	Green



Bromine (Br)	Liquid	Orange/Red-Brown
Iodine (I)	Solid	Grey-Black
Astatine (At)	Solid	Black

The Group VII trends include:

- Melting/Boiling points increase down the group.
- Density increases down the group.
- Reactivity decreases down the group.
- Colour becomes darker down the group.

e.g. of displacement reactions:

(Only a more reactive element can displace a less reactive element, it cannot occur the other way around.)

Group VIII (Noble Gases/Group 0):

The Group VIII elements include:

- Helium (He)
- Neon (Ne)
- Argon (Ar)
- Krypton (Kr)
- Xenon (Xe)
- Radon (Rn)

They exist as monoatomic gases and are extremely unreactive as they have a stable electronic configuration.

The Transition Elements

The Transition Elements are the elements found in between Group II and Group III of the periodic table.

Some notable transition elements include:

- Iron (Fe)
- Copper (Cu)
- Cobalt (Co)
- Vanadium (V)



- Chromium (Cr)
- etc.

Their physical properties include:

- Having high melting/boiling points.
- Having high density.
- Being hard/strong.

Their chemical properties include:

- Showing variable oxidation states
- Forming coloured compounds.
- Being used as catalysts in reactions.

P2:

Q1. Which statements about the position of the elements in the periodic table are correct?

Elements in the same group have similar chemical properties.

Elements in the same period have similar chemical properties.

Elements in the same group have the same number of electron shells.

Elements in the same group have the same number of outer shell electrons.

1 and 3

1 and 4

2 and 3

2 and 4

Q2. Which statements about transition elements are correct?

They have a low density.

They form ions with variable oxidation numbers.

They have a high melting point.

They form only colourless compounds.

1 and 2

1 and 4

2 and 3



3 and 4

Q3. Which statement about Period 3 of the Periodic Table is correct?

Aluminium is a nonmetal in Group III.

Argon is in Group VIII and has eight electrons in its outer electron shell.

Magnesium is in Group II and has three electrons in its outer electron shell.

Sulfur is a metal in Group VI.

Q4. Atoms of Sodium, Rubidium and Element Q each have one outer shell electron. Some properties of these elements are shown.

Element	Melting Point in °C	Boiling Point in °C	Density
Sodium	98	883	0.97
Rubidium	39	688	1.53
Q	28	672	1.87

What is Q?

Hydrogen

Lithium

Potassium

Caesium

Q5. Which statement about elements in Group VIII of the periodic table are correct?

They all have a full outer shell of electrons.

They all react with Group I elements to form ionic compounds.

They are all diatomic molecules.

They are all liquids at room temperature and pressure.

P4:

Q1. The list shows properties of some elements.



- Have low densities
- Have low melting points
- Form acid or basic oxides
- Form coloured compounds
- Form positive or negative ions

Iron is a transition metal. Sodium is a Group I metal.

State which property from the list:

- (i). Is true for sodium but not iron.
- (ii). Is true for iron but not sodium.
- (iii). Is true for both sodium and iron.
- (iv). Is not true for both sodium and iron.

Q2. State one physical and one chemical property that is common in transition elements like Copper.

Q3. State the name given to Group VII elements and explain why they have similar chemical properties.

Q4. Describe 2 differences in the physical properties of Group I elements and transition elements.

Q5. Predict the physical state of astatine at room temperature and pressure.



(vi) Chemical energetics

Exothermic and Endothermic Reactions

- In this topic, we will be mainly focusing on two types of reactions: Exothermic reactions and Endothermic reactions.
- Remember the mnemonic “Bendo and Mexo” (more energy absorbed in breaking of bonds in endothermic reactions, and more energy released in making of bonds in exothermic reactions).
- The Enthalpy change (ΔH), also known as energy change, is the transfer of thermal energy during a reaction.
- The Activation energy is the minimum amount of energy required for a reaction to take place. It is always positive.

Exothermic Reactions

- Exothermic reactions are reactions in which high enthalpy reactants are converted into low enthalpy products.
- On the basis of bond energy, exothermic reactions are those reactions in which the amount of energy released during bond formation is greater than the amount of energy absorbed during the breaking of bonds.

Exothermic Reactions

- In exothermic reactions, heat is released into the surroundings and the temperature of the surroundings rises.

They are reactions where the energy released during bond making is greater than the energy consumed during bond breaking

The enthalpy change of an exothermic reaction is negative.

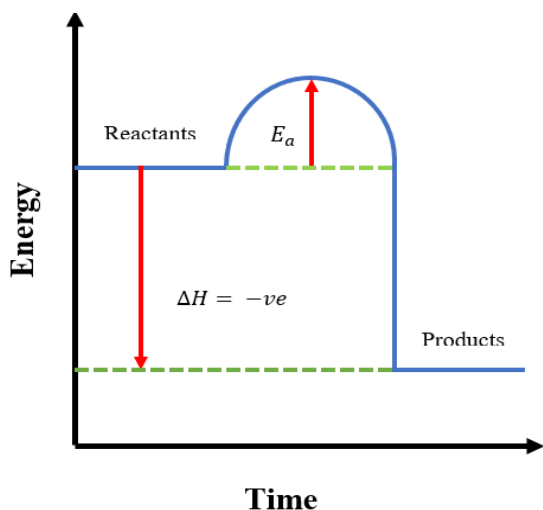


Examples of Exothermic Reactions:

- Respiration
- Combustion

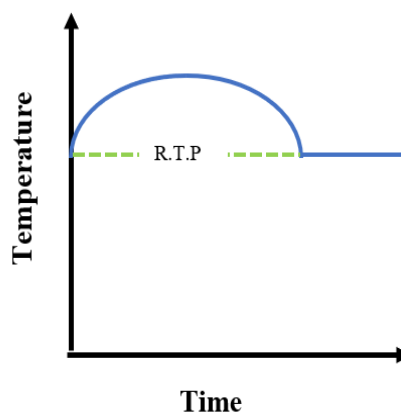
Neutralisation

Energy Profile Diagram:



Temperature-time graph

Temperature-Time Graph:



Endothermic Reactions

- Endothermic reactions are reactions in which low enthalpy reactants are converted into high enthalpy products.
- On the basis of bond energy, endothermic reactions are those reactions in which the amount of energy absorbed during bond breaking is greater than the amount of energy released during bond formation.
- In endothermic reactions, heat is absorbed and the temperature of the surroundings falls.

Examples of Endothermic Reactions:

- Photosynthesis
- Thermal Decomposition
- Electrolysis

Energy Profile Diagram:

Energy Profile Diagram:



E_a = Activation Energy (positive)
 ΔH = Enthalpy Change (positive)

Temperature-Time Graph:

Temperature-Time Graph:





Bond Energy Calculations

- During bond energy calculations, always remember:

$\Delta H = \text{Energy absorbed in Bond breaking} - \text{Energy released in Bond formation}$

OR

$\Delta H = \text{Reactants} - \text{Products}$

Q. The chemical equation for this reaction can be represented as shown.

The energy change for the reaction is -180 kJ/mol .

Use the bond energies in the table to calculate the bond energy of a C-Cl bond, in kJ/mol .

bond	C-H	C=C	Cl-Cl	C-C
bond energy in kJ/mol	410	610	240	350

Use the following steps:

Step 1 Calculate the energy needed to break bonds.

$$= 4(\text{C-H}) + (\text{C=C}) + (\text{Cl-Cl})$$

$$= 4(410) + (610) + (240)$$

$$= 1640 + 850$$

$$= +2490 \text{ kJ}$$

Step 2 Use your answer in Step 1 and the energy change for the reaction to determine the energy released when bonds are formed.

$$2490 - x = -180$$

$$x = 2490 + 180$$

$$x = 2670 \text{ kJ}$$

Step 3 Use your answer in Step 2 and bond energy values to determine the energy of a C-Cl bond.

$$2(\text{C-Cl}) + (\text{C-C}) + 4(\text{C-H}) = 2670$$

$$2(\text{C-Cl}) + (350) + 4(410) = 2670$$



$$2(\text{C-Cl}) + 1990 = 2670$$

$$2(\text{C-Cl}) = 680$$

$$(\text{C-Cl}) = 240 \text{ kJ/mol.}$$

Past paper questions

P2 :

Which of the following processes is endothermic ?

Sodium reacting with water

Magnesium burning in oxygen

Distilling crude oil

Use of petrol in an engine

Which statement is correct ?

Bond breaking requires energy and is an exothermic process

Bond breaking requires energy and is an endothermic process

Bond breaking releases energy and is an exothermic process

Bond breaking releases energy and is an endothermic process

Endothermic reactions are used in emergency sports treatment packs .Which statement is true about them :

Absorb thermal energy from surrounding and transfer to injury

Can be reused

Usually contain ammonium nitrate and water in separate compartments

Cause temperature of surroundings to increase



P4:

Q1:

(b) Bond breaking is an endothermic change and bond forming is an exothermic change.

Bond energy is the amount of energy in kJ/mol needed to break one mole of the specified bond



Use the following bond energies to determine whether this reaction is exothermic or endothermic. You must show your reasoning.

bond	bond energies in kJ/mol
C-Cl	338
C-H	412
Cl-Cl	242
H-Cl	431
C-C	348

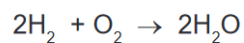
.....
.....
.....
..... [3]

Q2 : Explain the difference between an endothermic and exothermic reaction in terms of bond breaking and formation



Q3 :

Hydrogen and oxygen react to form water.



(i) Give an example of bond breaking in the above reaction.

..... [1]

(ii) Give an example of bond forming in the above reaction.

..... [1]

(iii) Is the change given in (i) exothermic or endothermic?

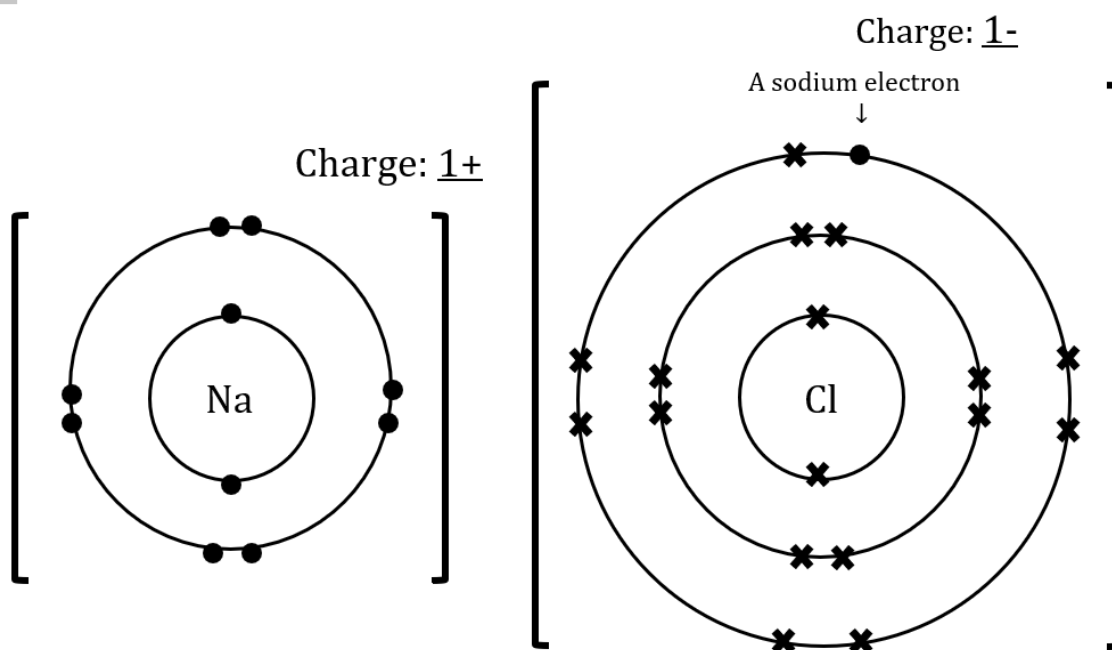
..... [1]



Chemical Bonding

Ionic Bonding

- The type of bonding between metal and non-metal atom(s) is known as ionic bonding
- The metal atom(s) lose their valence shell electrons to form positive ions. They give their electrons to the non-metal atom(s) which then form negative ions
- An ionic bond is the strong electrostatic forces of attraction between the metal cations and the non-metal anions
- Ionic bonds are often graphically represented as dot-and-cross diagrams, e.g. NaCl



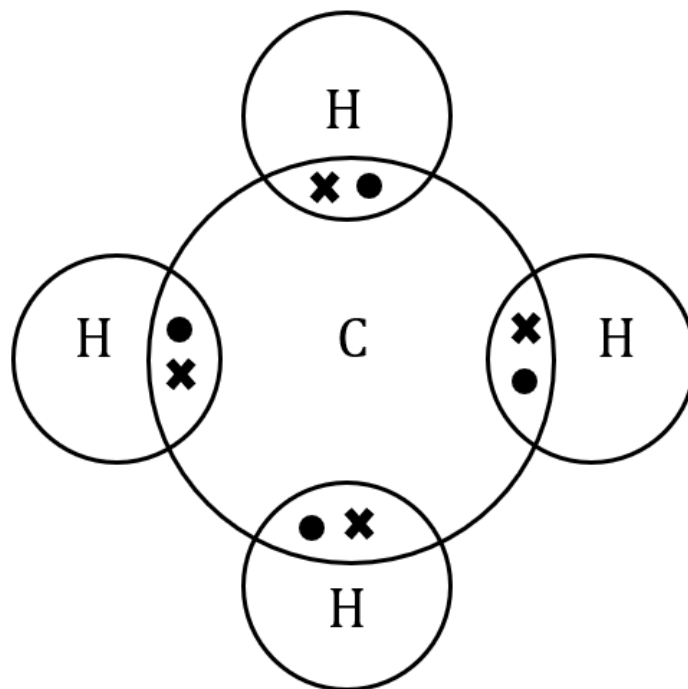


Na has 1 electron in its valence shell while Cl has 7 electrons in its valence shell. One Na atom loses 1 electron while 1 Cl atom gains 1 electron. This forms NaCl, an ionic compound, with a stable electronic configuration.

- Properties of Ionic Compounds:
- Have high melting/boiling points (as the strong electrostatic forces in the giant ionic lattice require a large amount of energy to break).
- Conduct electricity in molten or aqueous state, not as solids (as their ions can move freely).
- Are soluble in water.

Covalent Bonding

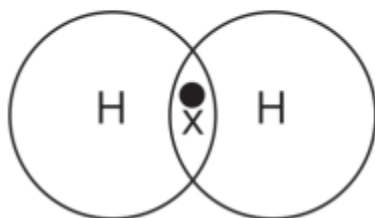
- The type of bonding between two or more non-metal atoms is known as covalent bonding.
- Non-metal atoms mutually share electrons amongst themselves to form a stable electronic configuration
- One electron pair shared = Single covalent bond.
- Two electron pairs shared = Double covalent bond.
- Three electron pairs shared = Triple covalent bond.
- Covalent bonds are often graphically represented as dot-and-cross diagrams, but these are different from the diagrams of Ionic bonding, e.g. C



- C has 4 electrons in its outer-shell and each H atom has 1 electron in their outer shell. C shares its 4 electrons with four separate H atoms, thus allowing C to complete its octet shell (shell with 8 electrons), and each H atom its duplet shell (shell with 2 electrons). This forms CH_4 , with a stable electronic configuration.

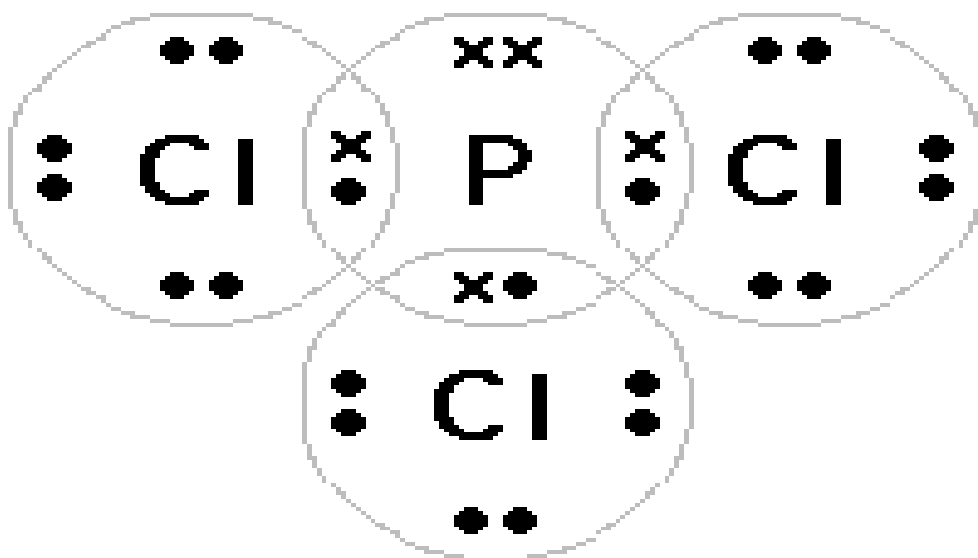
Examples of covalent bonds

'dot-and-cross' diagram



structural formula





PCl_3

- Simple Covalent molecules and Macromolecules:
- Covalent bonding leads to the formation of two types of molecules: Simple Covalent molecules and Giant Covalent Structures (Macromolecules).
- Simple Covalent molecules contain a fixed number of atoms joined by strong covalent bonds and are held together by weak intermolecular forces of attraction. These intermolecular forces require a small amount of energy to break.
- Simple Covalent molecules usually exist as gases/liquids while Macromolecules exist as solids

Properties of Simple Covalent Molecules:

- Have low melting/boiling points (due to weak intermolecular forces.).
- Don't conduct electricity
- Are (mostly) insoluble in water.
- Macromolecules have thousands of strong covalent bonds between atoms, which require a tremendous amount of energy to break.
- There are four main macromolecules that you need to know about:

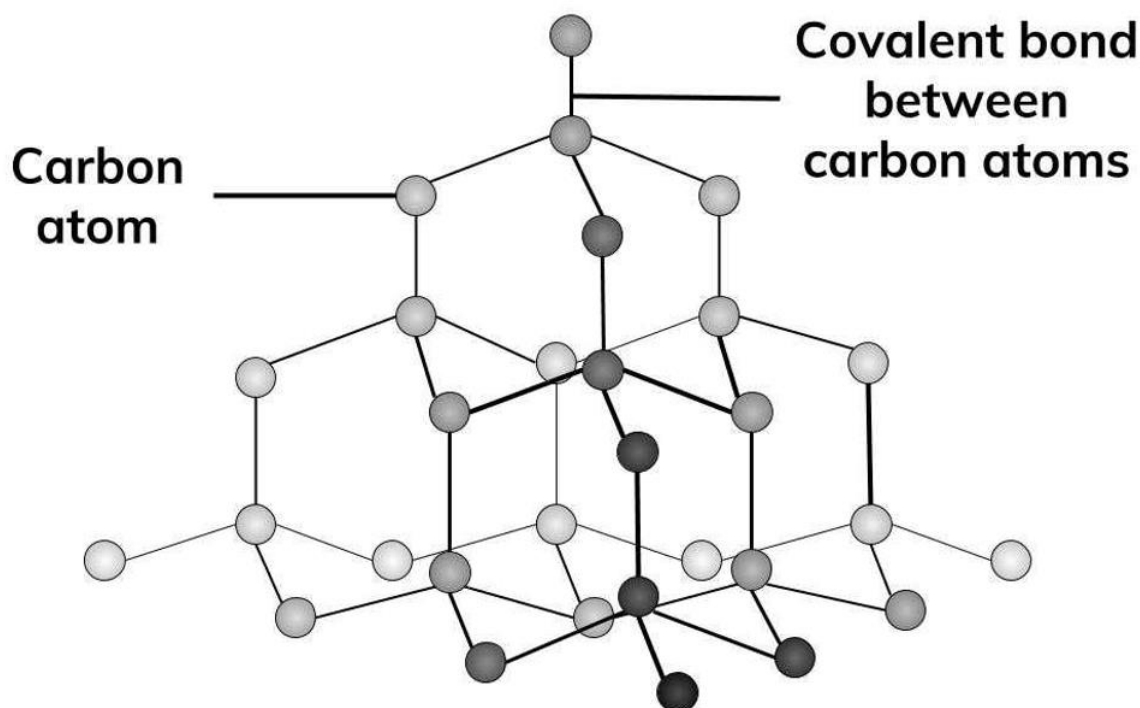


Diamond

- Graphite
- Silicon (Structure is similar to Diamond)
- Silicon(IV) Oxide / Silicon Dioxide / Silica

Diamond:

- Diamond consists of carbon atoms. Each carbon atom is bonded with four other carbon atoms.



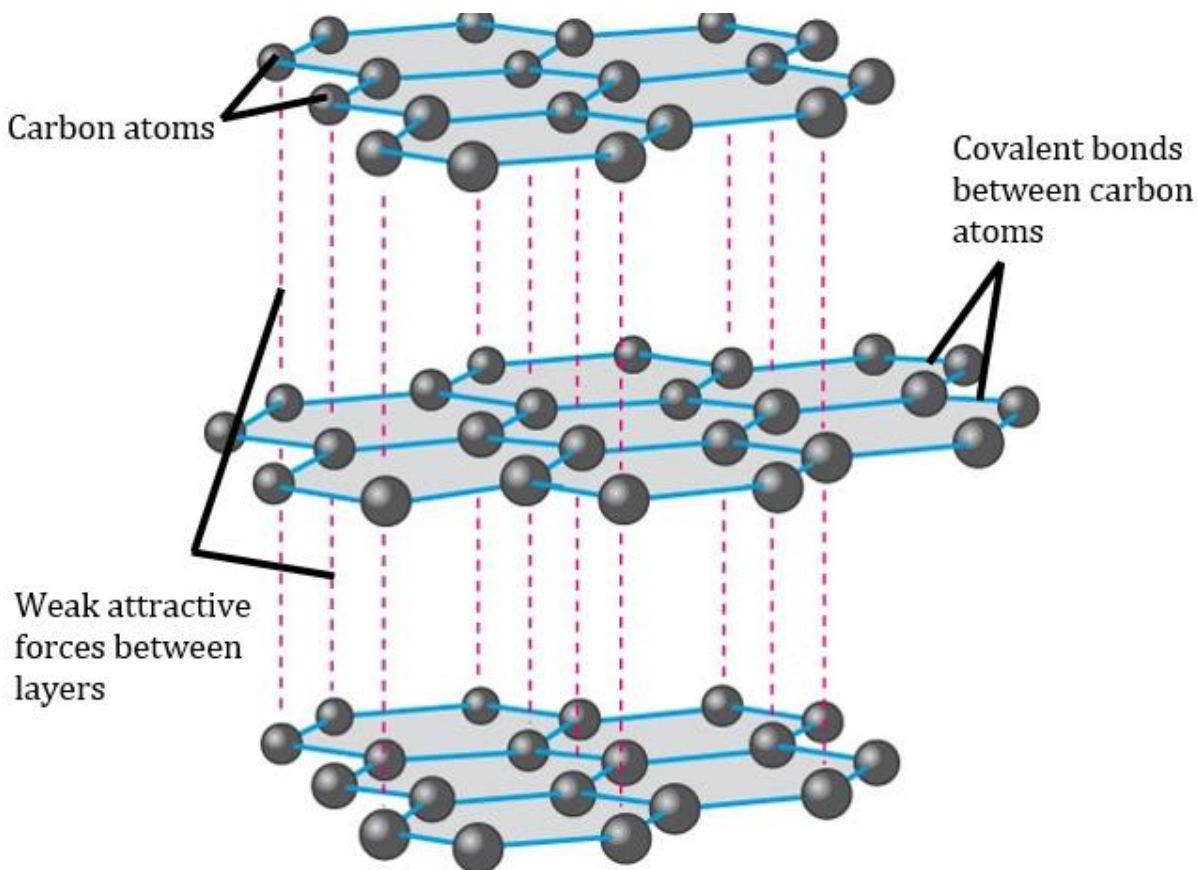
Diamond:

- Has a tetrahedral structure.
- Has a high melting/boiling point.
- Is extremely hard/strong.
- Does not conduct electricity (no freely moving electrons).
- Is used in making cutting tools and ornaments.



Graphite:

- Graphite also consists of carbon atoms. Each carbon atom is bonded with three other carbon atoms.



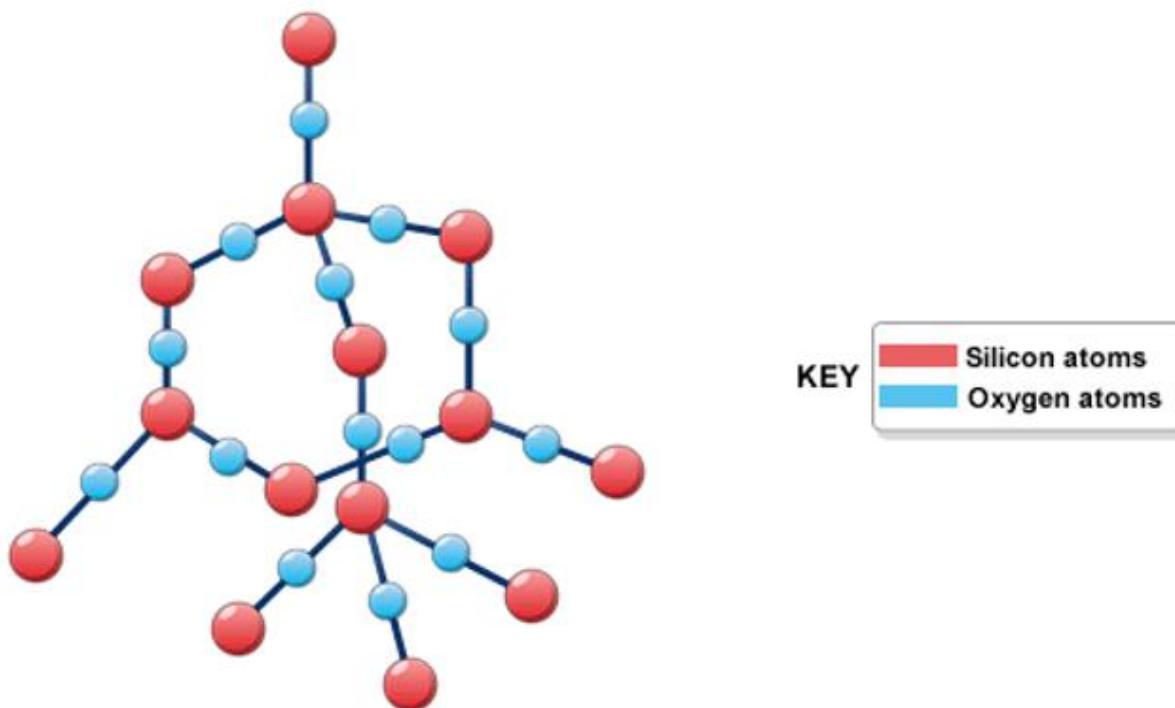
Graphite:

- Has a hexagonal structure with layers.
- Has a high melting/boiling point.
- Is soft.
- Conducts electricity (as it has freely moving electrons)
- Is used as a lubricant, in making pencils, and as an electrode due to conducting electricity and being inert.

Silica:



- Silicon(IV) Oxide/Silica consists of silicon and oxygen atoms. Each silicon atom is bonded with four oxygen atoms and each oxygen atom is bonded with two silicon atoms.



Silica:

- Has a tetrahedral structure.
- Has a high melting/boiling point.
- Is hard/strong.
- Does not conduct electricity (no freely moving electrons).
- Is used in making sandpaper, and with lining the inside of furnaces.

Metallic Bonding

- The type of bonding between metal atoms in a giant metallic lattice is known as metallic bonding.



- Within the metallic lattice, the atoms lose their outer shell electrons and become positively charged ions. The outer electrons no longer belong to a particular metal atom and are said to be delocalized. They move freely between the positive metal ions like a sea of electrons.
- A metallic bond is the strong electrostatic forces of attraction between the metal cations and the sea of delocalized electrons
- Metallic bonds are often graphically represented by a diagram of the metallic lattice.

Properties of Metals:

- Have high melting/boiling points (as the strong electrostatic forces in the metallic lattice require a large amount of energy to break).
- Conduct electricity in all states,
- Are malleable and ductile (as, when force is applied, layers of cations can slide over each other).

Past Paper Questions :

P2 :

Here are 5 IGCSE Chemistry multiple-choice questions on Chemical Bonding in increasing difficulty:

1. What type of bonding occurs between sodium and chlorine in sodium chloride (NaCl)?

- A) Covalent
- B) Ionic



- C) Metallic
- D) Hydrogen

Answer: B) Ionic

2. Which of the following pairs of elements will form a covalent bond?

- A) Sodium and chlorine
- B) Magnesium and oxygen
- C) Carbon and oxygen
- D) Potassium and bromine

Answer: C) Carbon and oxygen

3. Which property is characteristic of substances with giant covalent structures?

- A) High melting and boiling points
- B) Conduct electricity when solid
- C) Low melting and boiling points
- D) Soluble in water

Answer: A) High melting and boiling points

4. Which of the following is true about the bonding in methane (CH_4)?



- A) It involves a transfer of electrons.
- B) It consists of a lattice of positive ions.
- C) It involves shared pairs of electrons.
- D) It has ionic bonds between hydrogen and carbon.

Answer: C) It involves shared pairs of electrons.

5. Which statement explains why graphite conducts electricity?

- A) It has a giant ionic structure.
- B) It contains free-moving electrons between its layers.
- C) It has delocalized protons in its structure.
- D) It has covalent bonds that allow electron movement.

Answer: B) It contains free-moving electrons between its layers.

P4 :

Describe the type of bonding that occurs in metals and explain why metals are good conductors of electricity.

Explain why ionic compounds tend to have high melting points.

Describe how a covalent bond is formed between two atoms of hydrogen.

Explain why graphite can conduct electricity but diamond cannot, even though both are forms of carbon.

Why do substances with simple molecular structures, such as chlorine (Cl_2), tend to have low boiling points compared to substances with giant covalent structures like silicon dioxide (SiO_2)?



Rate of Reaction

Factors Affecting the Rate of Reaction

- The rate of reaction is the fastest at the start of the reaction. As the reaction progresses, the rate slows down and finally becomes zero when all the reactants have been used up.

Factors affecting the rate of reaction:

- Temperature
- Pressure (for gases only)
- Concentration
- Particle Size
- Catalysts

(i) Temperature

- By increasing the temperature, the rate of reaction increases, and vice versa.

This is because, when the temperature increases:

- The Kinetic Energy of the particles increases.
- The number of successful collisions between particles increases.
- More particles have energy greater than the activation energy.

(ii) Pressure (for gases only)

- By increasing the pressure the rate of reaction increases, and vice versa. This is because, when the pressure increases:

- The particles become tightly packed together and the separation between them decreases.
- The frequency of collisions increases.

(iii) Concentration



- By increasing the concentration the rate of reaction increases, and vice versa.

This is because, when the concentration increases:

- Particles become more crowded together.
- The frequency of collisions increases.

(iv) Particle Size

- Particles in powdered form have a higher surface area, therefore:
- The frequency of collisions increases.
- The rate of reaction becomes faster.
- Particles in the form of lumps have a smaller surface area, therefore:
- The frequency of collisions decreases.
- The rate of reaction becomes slower.

(v) Catalysts

- A catalyst is a substance which speeds up a chemical reaction without being used up.

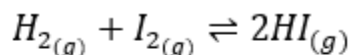
By adding a catalyst the rate of reaction increases. This is because, when a catalyst is added:

- The activation energy of the reaction is lowered, providing an alternate pathway for the reaction.

Reversible Reactions and Equilibrium

Reversible and Irreversible Reactions.

- Reversible Reactions are reactions which can proceed in both forward and backward directions, for example,



(The $\rightleftharpoons$ arrow means the reaction is reversible)



- Irreversible Reactions are reactions which can proceed in one direction only.
- In this topic, we will be focusing on reversible reactions.

Dynamic Equilibrium.

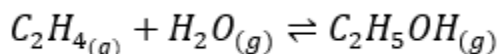
- Dynamic Equilibrium is a state in a reversible reaction at which:
- The rate of the forward reaction is equal to the rate of the backward reaction.
- The concentrations of the reactants and products are constant/fixed.
- Dynamic Equilibrium occurs in a closed container/system.
- We often use the term “position of equilibrium” when talking about equilibrium. The position of equilibrium can either shift to the left-hand side (reactant side), or the right-hand side (product side).
- Le Chatelier’s Principle: this principle states that when a change is made to the conditions of a system at equilibrium, the system automatically moves to oppose the change. This allows us to predict the changes to the position of equilibrium.

Factors affecting the position of equilibrium:

- Temperature
- Pressure (for gases only)
- Concentration.

(i) Temperature

- An increase in temperature favours the endothermic reaction, while a decrease in temperature favours the exothermic reaction. For example,



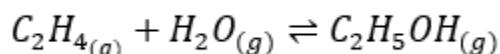
- The forward reaction is exothermic, thus the backward reaction must be endothermic.



- Increasing the temperature would favour the endothermic reaction. In this case, this would cause the position of equilibrium to shift to the left-hand side, thereby decreasing the yield of ethanol .
- Decreasing the temperature would favour the exothermic reaction. In this case, this would cause the position of equilibrium to shift to the right-hand side, thereby increasing the yield of ethanol .

(ii) Pressure (for gases only)

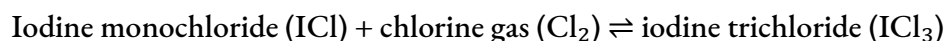
- An increase in pressure favours the side with fewer moles of gases, while a decrease in pressure favours the side with more moles of gases. Let us take the same example as before,



- The left-hand side has 1 mole of ethene gas (C_2H_4) and 1 mole of water vapor (H_2O), therefore it has 2 moles in total. The right-hand side has only 1 mole of ethanol vapor (C_2H_5OH).
- Increasing the pressure would favor the side with fewer moles of gases. In this case, this would cause the position of equilibrium to shift to the right-hand side, thereby increasing the yield of ethanol vapor (C_2H_5OH).
- Decreasing the pressure would favor the side with more moles of gases. In this case, this would cause the position of equilibrium to shift to the left-hand side, thereby decreasing the yield of ethanol vapor (C_2H_5OH).

(iii) Concentration

- An increase in concentration favors the opposite side to where the concentration was increased, and a decrease in concentration favors the opposite side to where the concentration was decreased. For example:





- Increasing the concentration of either iodine monochloride (ICl) or chlorine gas (Cl_2) will cause the position of equilibrium to shift to the right-hand side, thereby increasing the yield of iodine trichloride (ICl_3).
- Decreasing the concentration of either iodine monochloride (ICl) or chlorine gas (Cl_2) will cause the position of equilibrium to shift to the left-hand side, thereby decreasing the yield of iodine trichloride (ICl_3).

Equilibrium and Rate of Reaction

- Do not confuse the effects of temperature, pressure, and concentration on the rate of reaction with their effects on the position of equilibrium.
- While an increase in temperature may favor the endothermic reaction and not the exothermic reaction, it will still increase the rate of reaction for both the exothermic and endothermic reactions.
- While an increase in pressure may favor the side with fewer moles and not the side with more moles, it will still increase the rate of reaction for both the forward and backward reactions.
- While an increase in concentration may favor the opposite side to where the concentration was increased and not the other side, it will still increase the rate of reaction for both the forward and backward reactions.
- Adding a catalyst will only increase the speed/rate of both the forward and backward reactions. It has no effect on the position of equilibrium.
- Industries often use an optimum/compromised temperature and pressure. While increasing the temperature may decrease the yield of the product industries are trying to produce, it also increases the rate of reaction at which this product forms. Similarly, pressure may have the same effect depending on the product.
- If industries were to use too low of a temperature (similarly with pressure) just to increase the yield of the product, the reaction would be extremely slow. If they were to use too high of a temperature (similarly with pressure) just to increase the speed of the reaction, the yield of the product would be extremely low.
- Thus, there is an optimum temperature and pressure for most chemical reactions. These are the temperatures and pressures at which the yield of the product and the speed of the reaction are balanced.

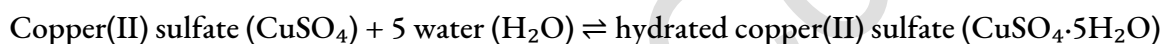


Hydrated and Anhydrous Salts (Test for Water)

- Hydrated salts are salts that contain water of crystallization, which affects their molecular shape and color.
- Water of crystallization is the water that is stoichiometrically included in the structure of some salts during the crystallization process.
- Water of crystallization is indicated with a dot written between the salt molecule and the surrounding water molecules.
- Anhydrous salts are those that have lost their water of crystallization, usually by heating.
- In our syllabus, there are two examples of hydrated salts that we must know about:

(i) Hydrated and Anhydrous Copper(II) Sulfate

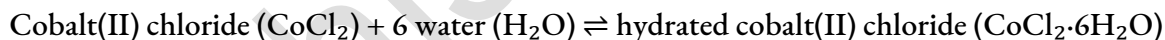
- Water can be added to anhydrous copper(II) sulfate to form hydrated copper(II) sulfate. This is a reversible reaction:



- Anhydrous copper(II) sulfate is a white powder, while hydrated copper(II) sulfate is a blue crystal.
- In this case, the forward reaction is exothermic, and the backward reaction is endothermic.

(ii) Hydrated and Anhydrous Cobalt(II) Chloride

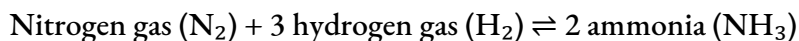
- Water can be added to anhydrous cobalt(II) chloride to form hydrated cobalt(II) chloride. This is a reversible reaction:



- Anhydrous cobalt(II) chloride is a blue powder, while hydrated cobalt(II) chloride is a pink crystal.
- In this case, the forward reaction is exothermic, and the backward reaction is endothermic.

The Haber Process

- The Haber Process is the manufacturing of ammonia at an industrial scale. There are certain conditions regarding the Haber Process that you must know:



- In this reaction:
 - The temperature is 450°C
 - The pressure is 200 atm / 20,000 kPa
 - The catalyst used is iron (Fe)

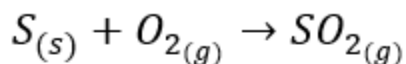


- In the Haber Process, nitrogen is sourced from the air via fractional distillation, and hydrogen is obtained through the cracking of hydrocarbons or electrolysis of water (H_2O).

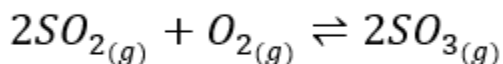
The Contact Process

- The Contact Process is the manufacturing of sulfuric acid at an industrial scale. This process consists of 4 stages, and there are certain conditions regarding the Contact Process that you must know.

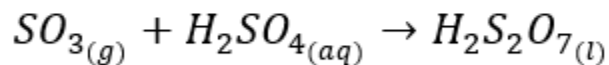
Stage 1:



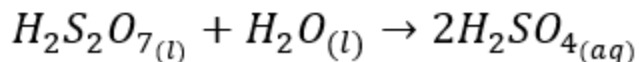
Stage 2 (Reversible Reaction):



Stage 3:



Stage 4:





- In this reaction (Stage 2):
- The temperature is 450°C
- The pressure is 2 atm / 200 kPa
- The catalyst used is vanadium(V) oxide (V_2O_5)

- Sulfur dioxide is formed in stage 1, then sulfur trioxide is formed in stage 2. In stage 3, sulfur trioxide is first acidified with sulfuric acid to form a liquid known as oleum ($\text{H}_2\text{S}_2\text{O}_7$), and finally, in stage 4, oleum is diluted with water to form more sulfuric acid.

Past Paper Questions

P2 :

1. Which of the following factors increases the rate of a reaction?

- A) Increasing the particle size of a solid reactant
- B) Lowering the concentration of reactants
- C) Adding a catalyst
- D) Decreasing the temperature

Answer: C) Adding a catalyst

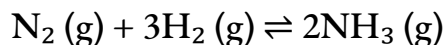
2. Which of the following will increase the rate of reaction between magnesium and dilute hydrochloric acid?

- A) Decreasing the temperature
- B) Using larger pieces of magnesium
- C) Using concentrated hydrochloric acid
- D) Removing the catalyst

Answer: C) Using concentrated hydrochloric acid



3. In the reversible reaction between nitrogen and hydrogen to form ammonia, which condition would shift the position of equilibrium to the right?



- A) Increasing the pressure
- B) Decreasing the concentration of nitrogen
- C) Increasing the temperature
- D) Adding a catalyst

Answer: A) Increasing the pressure

4. In a reaction where magnesium reacts with hydrochloric acid, which of the following changes would cause the rate of reaction to decrease?

- A) Increasing the surface area of magnesium
- B) Using dilute hydrochloric acid instead of concentrated acid
- C) Heating the reaction mixture
- D) Increasing the pressure of the system

Answer: B) Using dilute hydrochloric acid instead of concentrated acid

5. Which of the following describes how a catalyst affects a reaction?

- A) It increases the energy of reactant particles
- B) It increases the activation energy
- C) It provides an alternative reaction pathway with a lower activation energy
- D) It is used up in the reaction

Answer: C) It provides an alternative reaction pathway with a lower activation energy

6. The decomposition of ammonium chloride is a reversible reaction:



Which of the following changes would result in more ammonia being produced?

- A) Decreasing the temperature
- B) Increasing the pressure



C) Increasing the concentration of ammonium chloride

D) Increasing the temperature

Answer: D) Increasing the temperature

P4 :

Define what is meant by the rate of a chemical reaction.

State two ways in which you can increase the rate of a reaction between a solid and a gas.

Explain how increasing the temperature affects both the rate of reaction and the position of equilibrium in a reversible reaction.

A catalyst is used in a reaction. Explain how it increases the rate of the reaction without affecting the position of equilibrium.

In the reaction between nitrogen and hydrogen to form ammonia, how would decreasing the temperature affect the yield of ammonia?



REDOX Reactions

A redox reaction involves the transfer of electrons between substances.

When a substance gains electrons, it is reduced.

When a substance loses electrons, it is oxidised.

There are also multiple other definitions for oxidation and reduction, for example:

Reduction is the loss of oxygen atoms.

Oxidation is the gain of oxygen atoms.

Notes: BOTH of these are used in exams.

Oxidation Numbers (Oxidation States)

An oxidation number (or oxidation state) is a numerical value assigned to an element in a compound that reflects its degree of oxidation (loss of electrons) or reduction (gain of electrons).

Rules for Assigning Oxidation Numbers:

Elements in their Standard State:

The oxidation number of an element in its pure form (e.g., O_2 , N_2 , Cl_2 , Fe, etc.) is 0.

Example: O_2 has an oxidation number of 0.

Monatomic Ions:

The oxidation number is equal to the charge of the ion.

Example: For Na^+ , the oxidation number is +1; for Cl^- , it is -1.

Oxygen:

Typically has an oxidation number of -2 in compounds, except in peroxides (e.g., H_2O_2) where it is -1, and in OF_2 where it is +2.

Hydrogen:

Typically has an oxidation number of +1 when bonded to non-metals and -1 when bonded to metals.

Alkali Metals (Group 1):

Always have an oxidation number of +1 in compounds.



Alkaline Earth Metals (Group 2):

Always have an oxidation number of +2 in compounds.

Halogens (Group 17):

Generally have an oxidation number of -1.

Sum of Oxidation Numbers:

In a neutral compound, the sum of the oxidation numbers of all atoms must equal 0.

In a polyatomic ion, the sum of the oxidation numbers must equal the charge of the ion.

Examples of Assigning Oxidation Numbers:

Water (H₂O):

H: +1 (2 hydrogen atoms) $\rightarrow 2(+1) = +2$

O: -2 $\rightarrow$ Total: $+2 + (-2) = 0$

Carbon Dioxide (CO₂):

C: +4 (since O is -2, total -4 from 2 O's, C must be +4)

O: -2 $\rightarrow$ Total: $+4 + 2(-2) = 0$

Oxidizing and Reducing Agents

Oxidizing Agent:

An oxidizing agent (oxidant) is a substance that gains electrons (is reduced) in a redox reaction, causing the oxidation of another substance.

Characteristics:

Typically contains elements with high electronegativity (e.g., oxygen, halogens).

Increases its oxidation number.

Common Oxidizing Agents:

Potassium Permanganate (KMnO₄): Often used in redox titrations.

Hydrogen Peroxide (H₂O₂): Commonly used as a bleaching agent and disinfectant.

Chlorine Gas (Cl₂): Reacts with various compounds to oxidize them.

Example Reactions:

Chlorine Oxidizing Iron: $2\text{Fe} + \text{Cl}_2 \rightarrow 2\text{FeCl}_2$



Color Change: The iron changes from a shiny gray to orange-brown (iron (III) chloride color).
Fe is oxidized from 0 to +2.

Reducing Agent:

A reducing agent (reductant) is a substance that loses electrons (is oxidized) in a redox reaction, causing the reduction of another substance.

Characteristics:

Usually contains elements with lower electronegativity (e.g., metals).

Decreases its oxidation number.

Common Reducing Agents:

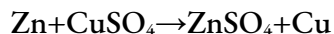
Sodium Borohydride (NaBH_4): Used in organic chemistry for reductions.

Zinc (Zn): Commonly used in metal displacement reactions.

Carbon (C): Reduces metal oxides in metallurgy.

Example Reactions:

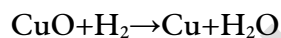
Zinc Displacing Copper:



Color Change: The blue solution of copper sulfate (CuSO_4) turns colorless, and copper metal is deposited.

Zn is oxidized from 0 to +2.

Hydrogen as a Reducing Agent:



Color Change: The black copper oxide (CuO) turns into red copper metal.

H_2 is oxidized, and CuO is reduced.



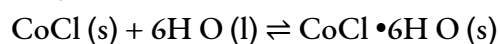
Chemistry of the Environment

Chemical Tests for Water

Cobalt (II) Chloride:

Anhydrous Cobalt (II) Chloride is blue in colour. Cobalt (II) Chloride paper is used to test the presence of water. In the presence of water, it changes from blue to pink.

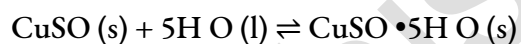
anhydrous cobalt(II) chloride + water $\rightleftharpoons$ hydrated cobalt(II) chloride



Copper (II) Sulfate:

Anhydrous Copper (II) Sulfate is white in colour and can be used to test the presence of water. In the presence of water, it changes from white to blue.

anhydrous copper(II) sulfate + water $\rightleftharpoons$ hydrated copper(II) sulfate



Tests for the Purity of Water

All pure substances have a fixed melting and boiling point. The melting point of pure water is 0°C and the boiling point of pure water is 100°C . When water is mixed with impurities, the boiling point tends to increase while the melting point tends to decrease.

Distilled Water

Distilled water is created by heating water until it turns into vapour and then condensing it back into a liquid. This process removes most impurities, making distilled water highly pure. It is often used in practical chemistry for this reason, while tap water, which contains more impurities, is generally avoided as it can interfere with chemical reactions.



Uses of Water

Agricultural Uses: irrigation and as a drink for animals

Industrial Uses: as a coolant, as a solvent and as steam for electricity

Domestic Uses: drinking, cooking, sanitation, gardening

Impurities in Water

Some of these impurities are naturally occurring while some are the results of human actions.

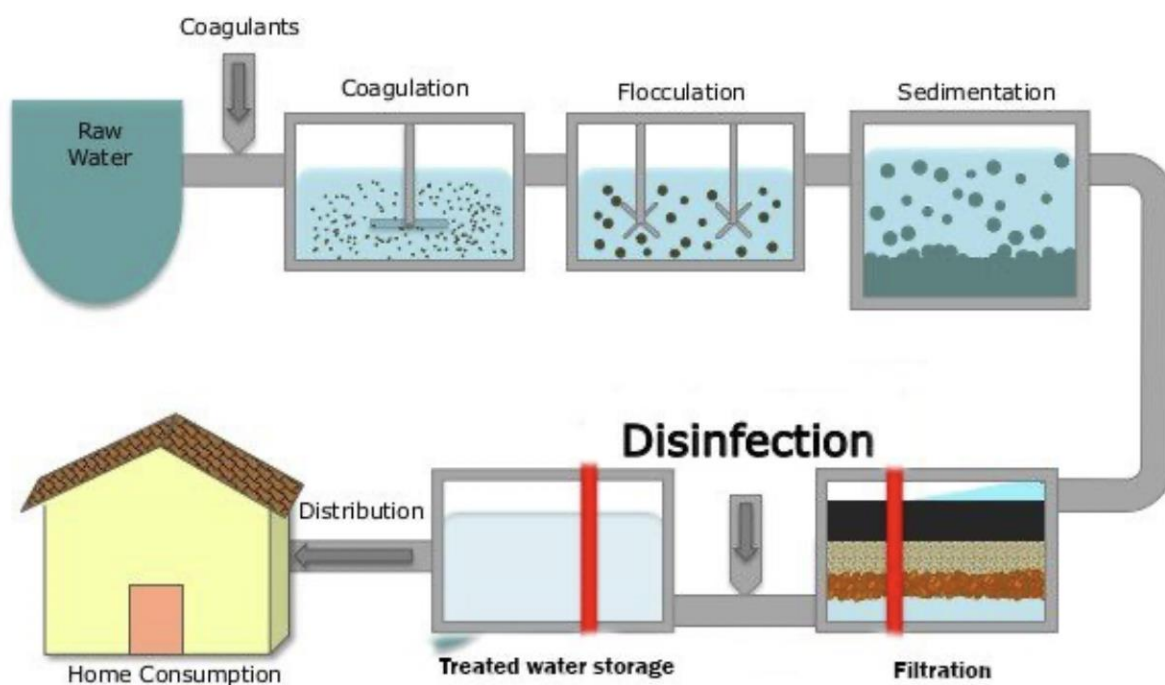
Impurity	Beneficial or Harmful	Reason
Dissolved Oxygen	Beneficial	It is essential for the aquatic life
Metal Ions (e.g. Ca, K, Na, Mg, Pb, Al)	Some are beneficial while some are harmful	Some are essential for life like Ca, K, Na and Mg. Others are toxic like Pb and Al
Plastics	Harmful	These may be harmful to aquatic life in many ways
Harmful Microbes	Harmful	These are added in the water through sewage and can lead to several diseases
Nitrates and Phosphates from Fertilisers	Harmful	It can lead to Eutrophication.

Treatment of Water

Untreated water contains both soluble and insoluble impurities. Insoluble impurities include soil, plant debris, and other organic matter, while soluble impurities consist of dissolved calcium, metallic compounds, and inorganic pollutants. The first step in water treatment is sedimentation or filtration. Water is pumped into sedimentation tanks and left to sit for several hours, allowing mud, sand, and other particles to settle at the bottom due to gravity, forming a sediment layer. Afterward, the water is filtered through sand and gravel to eliminate smaller particles. The second step involves filtration or treatment with carbon (charcoal), which removes unpleasant tastes and



odors. The final step is chlorination. Since bacteria and other microorganisms are too small to be captured by filters, chlorine is added to the water supply to kill them. Diseases like cholera and typhoid can result from consuming untreated water containing harmful bacteria.



Fertilizers

NPK fertilizers are crucial for enhancing plant growth and health by supplying key nutrients:

Nitrogen (N): Essential for producing chlorophyll and proteins, it helps plants grow healthy, green leaves. Nitrogen is provided through ammonium ions (NH_4^+) and nitrate ions (NO_3^-).

Phosphorus (P): Aids in the development of strong roots, improving the plant's stability and nutrient absorption. It is available as phosphate ions (PO_4^{3-}).

Potassium (K): Encourages the growth of robust fruit, flowers, and overall plant health. It dissolves in water as potassium ions (K^+).

Common fertilizer compounds include:

Ammonium nitrate (NH_4NO_3)

Ammonium phosphate ($(\text{NH}_4)_3\text{PO}_4$)

Potassium nitrate (KNO_3)

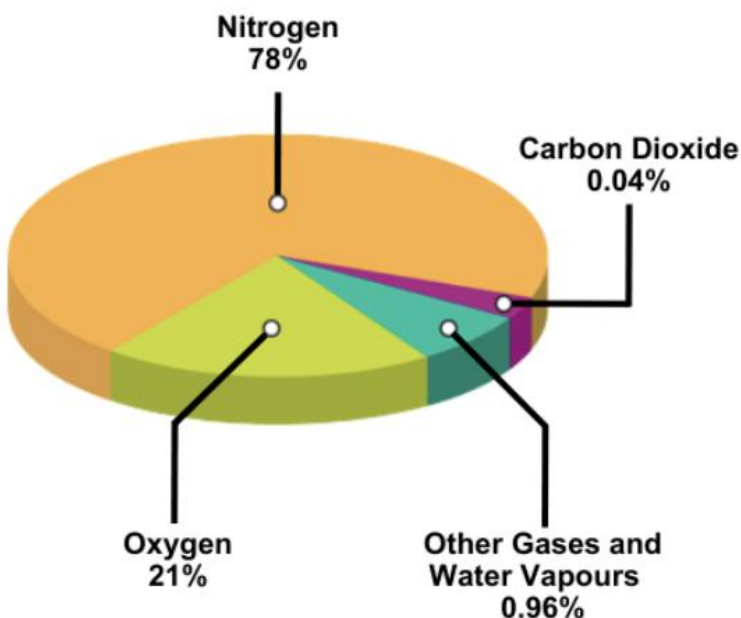


Different fertilisers offer varying amounts of nitrogen, phosphorus, and potassium, depending on their specific formulation.

Air

Composition of Air

Air is a mixture of gases. The two main gases found in the air are nitrogen and oxygen. Other gases include carbon dioxide, water vapours and argon etc.



Air Pollution

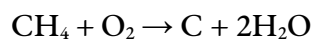
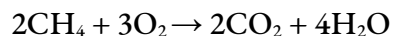
While there are naturally occurring gases found in the air, other gases that are the result of human actions are also found and that are grouped as air pollutants. Following are some of these air pollutants:

Carbon Dioxide: Carbon Dioxide is produced by the complete combustion of carbon-containing fuels, e.g. $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$. It can lead to severe global warming which ultimately results in climate change.

Carbon Monoxide: Carbon Monoxide is a toxic gas formed by the incomplete combustion of carbon-containing fuels, e.g. $\text{C}_8\text{H}_{18} + 9\text{O}_2 \rightarrow 5\text{CO} + 2\text{CO}_2 + 9\text{H}_2\text{O}$. Carbon Monoxide has the ability to combine with a human's hemoglobin and stop it from carrying oxygen throughout the body.



Air Particulates: Air particulates can lead to the production of carbon particles, sometimes referred as soot. For example, the incomplete combustion of Methane can produce carbon monoxide as well as carbon.



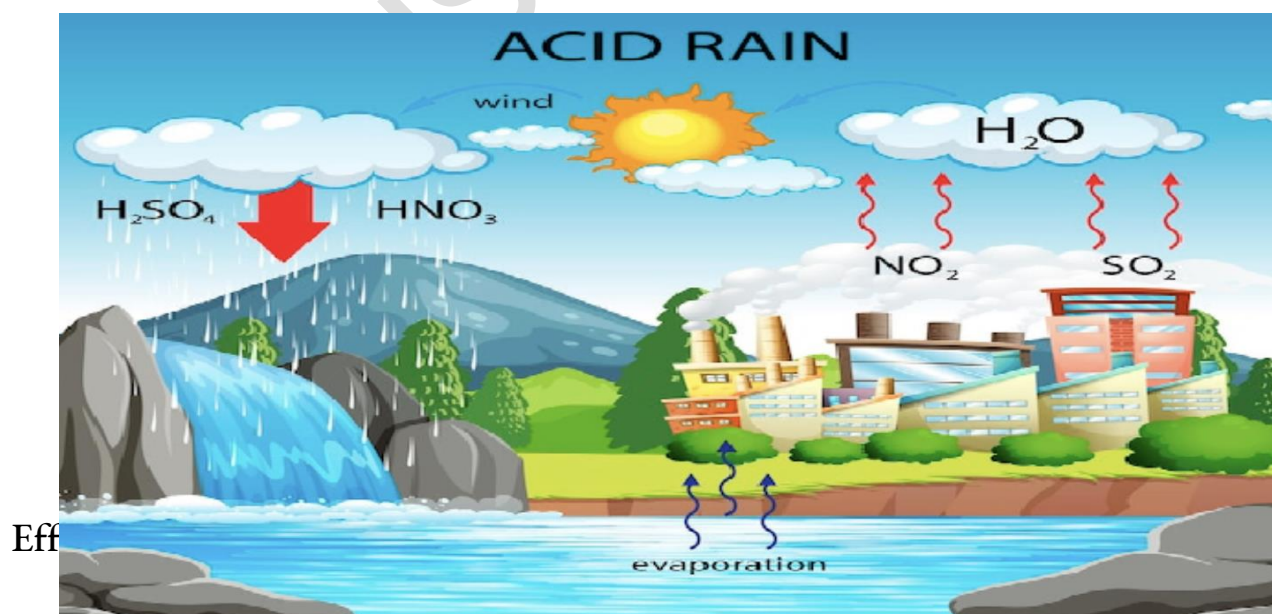
The air particulates can lead to respiratory problems or cancer.

Methane: There are several sources of methane including the decomposition of vegetation, waste gases from the digestive processes of animals and the bacterial action in swamps, rice paddies and

lakes. Methane can lead to a substantial increase in global warming which increases climate change.

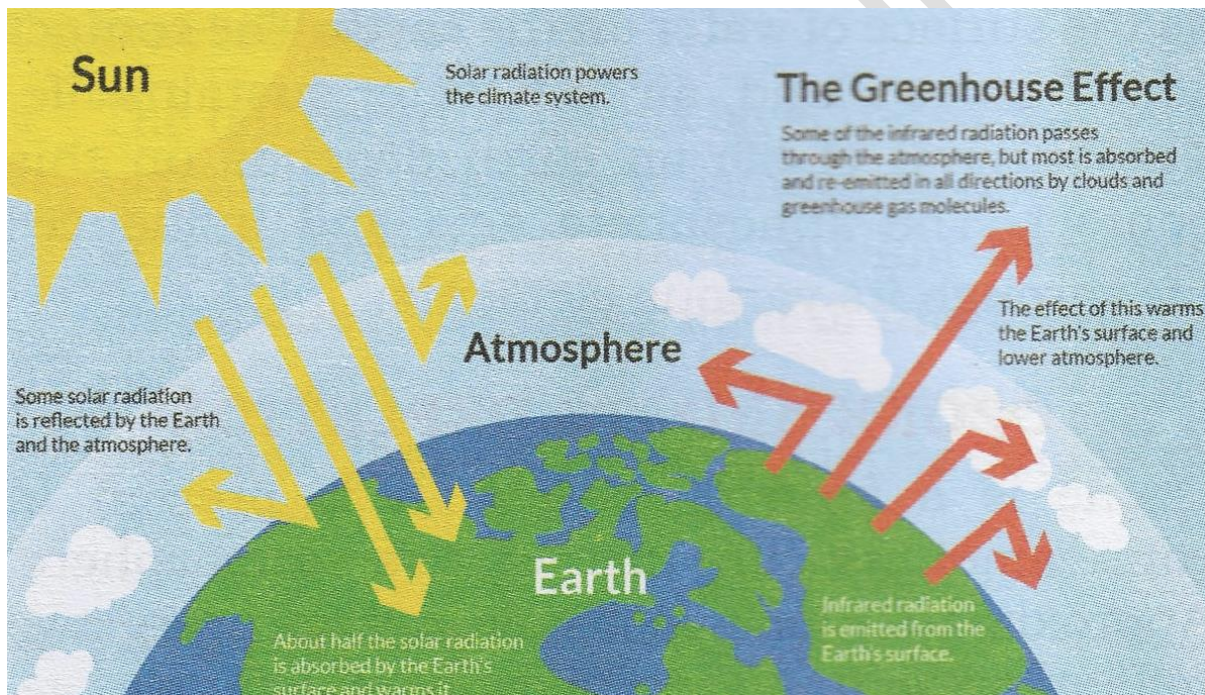
Oxides of Nitrogen: Nitrogen reacts with oxygen at high temperatures, such as in car engines, furnaces, or during lightning strikes, and can also form as a result of bacterial activity in soil. This reaction produces nitrogen oxides, which contribute to photochemical smog and dissolve in rainwater, leading to acid rain. Acid rain damages metal structures, carbonate-based buildings and statues, harms aquatic life, and pollutes crops and water supplies. Additionally, nitrogen oxides can irritate the lungs, throat, and eyes, causing respiratory problems.

Sulfur Dioxide: Sulfur Dioxide is produced by the combustion of sulfur-containing fossil fuels, e.g. $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$. Another major source of Sulfur Dioxide is the power stations. When it dissolves in rain, acid rain is formed with similar effects as the oxides of nitrogen.





The Sun emits energy in the form of radiation, which enters the Earth's atmosphere. While some of this thermal energy is reflected off the Earth's surface, most is absorbed and then re-emitted. As this energy moves through the atmosphere, part of it escapes into space, but some is absorbed by greenhouse gases like carbon dioxide and methane. These gases re-emit the thermal energy in all directions, trapping heat within the Earth's atmosphere and helping to maintain a warm climate—a process known as the greenhouse effect. However, as human activity increases the concentration of greenhouse gases, more thermal energy is trapped, leading to a rise in the Earth's average temperature. This is known as the enhanced greenhouse effect and contributes to global warming.



Consequences of Global Warming

The temperature rising on Earth is causing a change in the climate. Warm temperatures will lead glaciers to melt, raising water levels and creating flooding in nations with low lying areas. Another immediate consequence is the extinction of species as a result of habitat loss. It can also result in the migration of species as they will relocate to more livable (drought-free) places. There is also a risk for the diseases to be spread due to a much warmer climate. Global Warming also leads to a loss of habitat in areas of low-lying countries and areas surrounding glaciers.



Minimizing the Effects of the Environmental Issues

Reducing Carbon Dioxide Emissions:

Measures to reduce atmospheric carbon emissions and combat climate change include becoming more mindful consumers of energy. This involves using cleaner energy sources, like hydrogen and renewables (solar or wind), instead of fossil fuels. Opting for fuel-efficient vehicles, such as electric and hybrid cars, and reducing the number of vehicles on the road through public transport and car-sharing can further reduce emissions. Recycling and reusing products derived from crude oil, lowering household energy consumption (e.g., turning off lights and using efficient appliances),

and addressing deforestation by planting more trees or reforestation are also impactful steps.

Increased tree cover can absorb more atmospheric carbon dioxide through photosynthesis, helping to balance greenhouse gas levels.

Reducing Methane Emissions:

Decrease livestock farming, as methane is released during the animals' digestion process.

Reducing Acid Rain:

Acid rain results from nitrogen oxides and sulfur dioxide emissions. Its effects can be minimised by reducing the production of these pollutants. Catalytic converters in vehicles can help by removing nitrogen oxides from exhaust gases.

Reducing Sulfur Dioxide Emissions:

Sulfur dioxide emissions can be lowered by using fuels with minimal sulfur content. Another effective method is flue gas desulfurization, the primary technique for reducing sulfur dioxide. In this process, waste gases from coal-fired power plants are directed into a scrubbing chamber where they are sprayed with a wet slurry of calcium oxide and calcium carbonate. These calcium compounds react with sulfur dioxide, forming calcium sulfate as a byproduct.



Oxides of Nitrogen and Catalytic Converters

Nitrogen oxides (NO and NO₂) are produced when nitrogen and oxygen react under the high pressure and temperature conditions of internal combustion engines and blast furnaces. Exhaust gases also contain unburned hydrocarbons and carbon monoxide. To neutralize these harmful emissions, cars are equipped with catalytic converters as part of their exhaust systems.

Catalytic converters contain catalysts made from transition metals, such as platinum and rhodium, arranged in a honeycomb structure to maximise the surface area for reactions. Through a series of redox reactions, these catalysts neutralize pollutant gases. For instance, carbon monoxide is oxidized to form carbon dioxide: $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$.

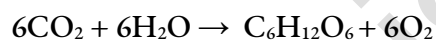
Nitrogen oxides are reduced to nitrogen gas: $2\text{NO} \rightarrow \text{N}_2 + \text{O}_2$ and $2\text{NO}_2 \rightarrow \text{N}_2 + 2\text{O}_2$.

A single reaction that represents the interaction of nitrogen monoxide and carbon monoxide within a catalytic converter is: $2\text{NO} + 2\text{CO} \rightarrow \text{N}_2 + 2\text{CO}_2$.

Photosynthesis

Photosynthesis is an endothermic reaction in which energy is absorbed from the environment by chloroplasts containing chlorophyll. In this process, carbon dioxide and water act as reactants and are converted into glucose and oxygen. Chlorophyll and light energy are essential for this reaction to take place. Following is the equation for the process:

Carbon Dioxide + Water → Glucose + Oxygen



P2:

Q1. How do carbon dioxide and methane cause global warming?

They emit the thermal energy they have absorbed back to the Earth.

They absorb the radiation directly from the Sun.

They increase thermal energy loss to space.

They reduce reflection of thermal energy from the Earth's surface.



Q2. Four statements about photosynthesis are listed.

Chlorophyll is required for photosynthesis.

The equation for photosynthesis is $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$.

Photosynthesis requires energy from light.

Photosynthesis releases carbon dioxide, which can lead to climate change.

Which statements are correct?

1 and 3

1 and 4

2 and 3

2 and 4

Q3. Several processes are used to treat domestic water. Which row identifies a reason for the given process?

	Process	Reason
A	Chlorination	Removes impurities
B	Filtration	Removes insoluble solids
C	Sedimentation	Removes soluble solids
D	Use of Carbon	Kills bacteria

4. Which compounds both contribute to acid rain?

Carbon monoxide and carbon dioxide

Carbon monoxide and oxides of nitrogen

Oxides of nitrogen and sulfur dioxide

Sulfur dioxide and carbon dioxide

Q5. Which statements about water are correct?

Tap water has fewer impurities than distilled water.

Tap water will turn anhydrous cobalt(II) chloride pink.

The domestic water supply is treated with carbon to kill microbes.

Phosphates from fertilisers can cause deoxygenation of water.



- 1 and 2
- 1 and 3
- 2 and 4
- 3 and 4

P4:

Q1. State which 2 oxides are responsible for acid rain.

Q2. Give the symbol of the element that is used to kill the microbes in water.

Q3. Name one gas which is responsible for both photochemical smog and acid rain.

Q4. Nitrogen dioxide and carbon monoxide are removed from a car exhaust by a catalytic converter. Write the symbol equation for this reaction.

Q5. Name a Halide compound that can be used to test the presence of water.



Metals

Physical Properties:

High melting and boiling points due to strong metallic bonds within a lattice structure.

Excellent conductors of heat and electricity because of free-moving delocalized electrons in the metallic structure.

Malleability and Ductility: Metals can be bent or drawn into wires because metal layers can slide over each other without breaking bonds.

Shiny (lustrous) appearance due to electron interaction with light.

Chemical Properties:

Formation of Positive Ions (Cations): Metals lose electrons easily, forming cations, due to low ionization energy.

Oxidation: Metals react with oxygen to form metal oxides, which are mostly basic, meaning they neutralize acids to form salt and water.

Reactivity: Metals differ in reactivity, affecting their reactions with water, acids, and oxygen, and the methods used for their extraction.

Reactivity Series

The reactivity series is an arrangement of metals in order of their reactivity, primarily with water, acids, and oxygen. This reactivity series determines:

Whether metals will displace others in compounds.

The methods used to extract the metal from its ore.

Suitability for various practical applications (e.g., more reactive metals are not ideal for construction as they corrode quickly).

Order of Reactivity (Most to Least):

Most Reactive: Potassium (K), Sodium (Na), Calcium (Ca)

Moderately Reactive: Magnesium (Mg), Aluminum (Al), Zinc (Zn), Iron (Fe)

Least Reactive: Copper (Cu), Silver (Ag), Gold (Au), Platinum (Pt)

A helpful way to remember it is this mnemonic

Please Send Cats, Monkeys And Zebras Into Hot Countries, Signed Gordon

Reactivity and Reactions with Water, Acids, and Oxygen:



Cold Water: Very reactive metals like Potassium, Sodium, and Calcium react vigorously with cold water.

Steam: Moderately reactive metals like Magnesium, Zinc, and Iron do not react with cold water but do react with steam to form metal oxides and release hydrogen gas.

Dilute Acids: Metals above hydrogen in the series react with dilute acids, producing hydrogen gas.

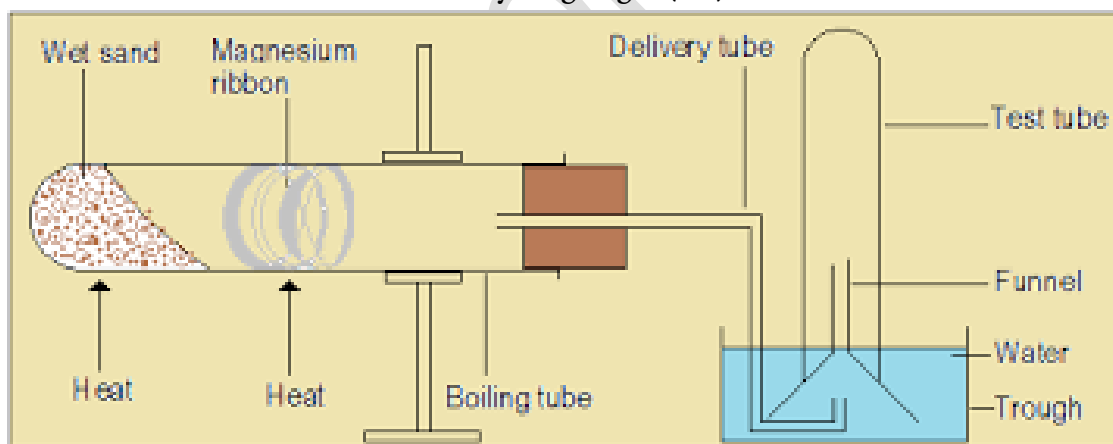
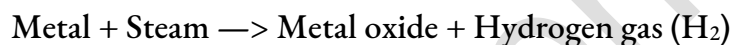
Oxygen: All metals, except for gold and platinum, react with oxygen, though the intensity and product type vary.

3. Detailed Reactions of Metals with Different Substances

Reaction with cold water : Reactive metals form hydroxides and hydrogen gas. (potassium , Sodium , Calcium)

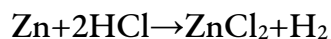


Moderately reactive metals like Magnesium, Zinc, and Iron react with steam, forming metal oxides and hydrogen gas.



Metals above hydrogen in the reactivity series react with dilute hydrochloric or sulfuric acid, releasing hydrogen gas.

Zinc with Hydrochloric Acid:



Effervescence (bubbles) observed; reaction is exothermic.

Metals like iron and aluminum require initial oxide layer removal before reacting.



Most metals react with oxygen to form an oxide except for gold, silver and platinum as they are unreactive .

Metal + Oxygen $\rightarrow$ Metal Oxide (often basic or amphoteric)

Magnesium: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

a bright white flame and basic oxide formed.

Alloys

It is A mixture of a metal with other elements to enhance certain properties. They are made to Improve hardness, strength, corrosion resistance, and other physical properties.

Examples:

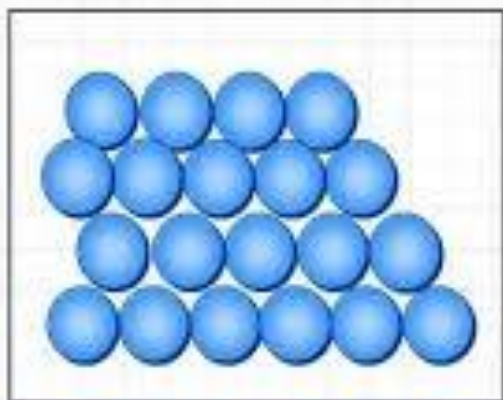
Brass (Copper + Zinc): Stronger and more durable than copper; used in musical instruments, plumbing.

Stainless Steel (Iron + Chromium + Nickel): Highly resistant to corrosion; used in cutlery, surgical tools.

Solder (Lead + Tin): Low melting point for electronic connections.

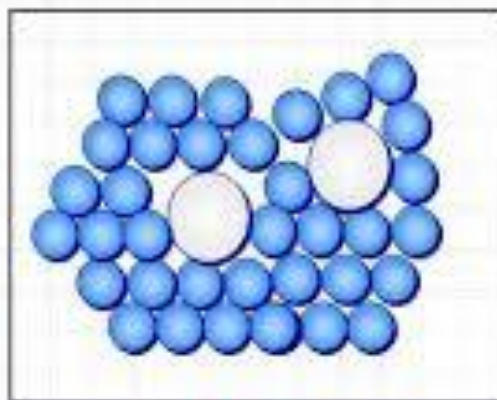
Structure and Properties of Alloys:

Alloys disrupt the uniform lattice structure of pure metals, since now there are 2 or more different types of atoms in the lattice . Since they are of different sizes , the layers cannot now easily slide over one another , increasing the strength and hardness of the new alloy



Pure metal

Orderly arrangement
becos of same size atoms



Alloy

Disorderly arrangement
becos of different size
atoms



Extraction of metals

In IGCSE , we learn about the extraction of 2 metals , Iron and Aluminium

There are multiple reasons for different extraction methods ,mainly accredited to the reactivity series . Since Aluminium is above Carbon and hydrogen in the Reactivity series , it cannot be extracted using a blast furnace since Carbon cannot displace it . Hence we use electrolysis to extract Aluminium while we use a blast furnace to extract iron , since it is below carbon in the reactivity series .

Extraction of Iron

Process: Extracted in a Blast Furnace from iron ore (mainly hematite, Fe_2O_3).

Raw Materials and Their Functions:

Iron Ore (Hematite): The source of iron.

Coke: A form of carbon; acts as a reducing agent.

Limestone (Calcium Carbonate): Removes impurities as slag.

Steps in the Blast Furnace:

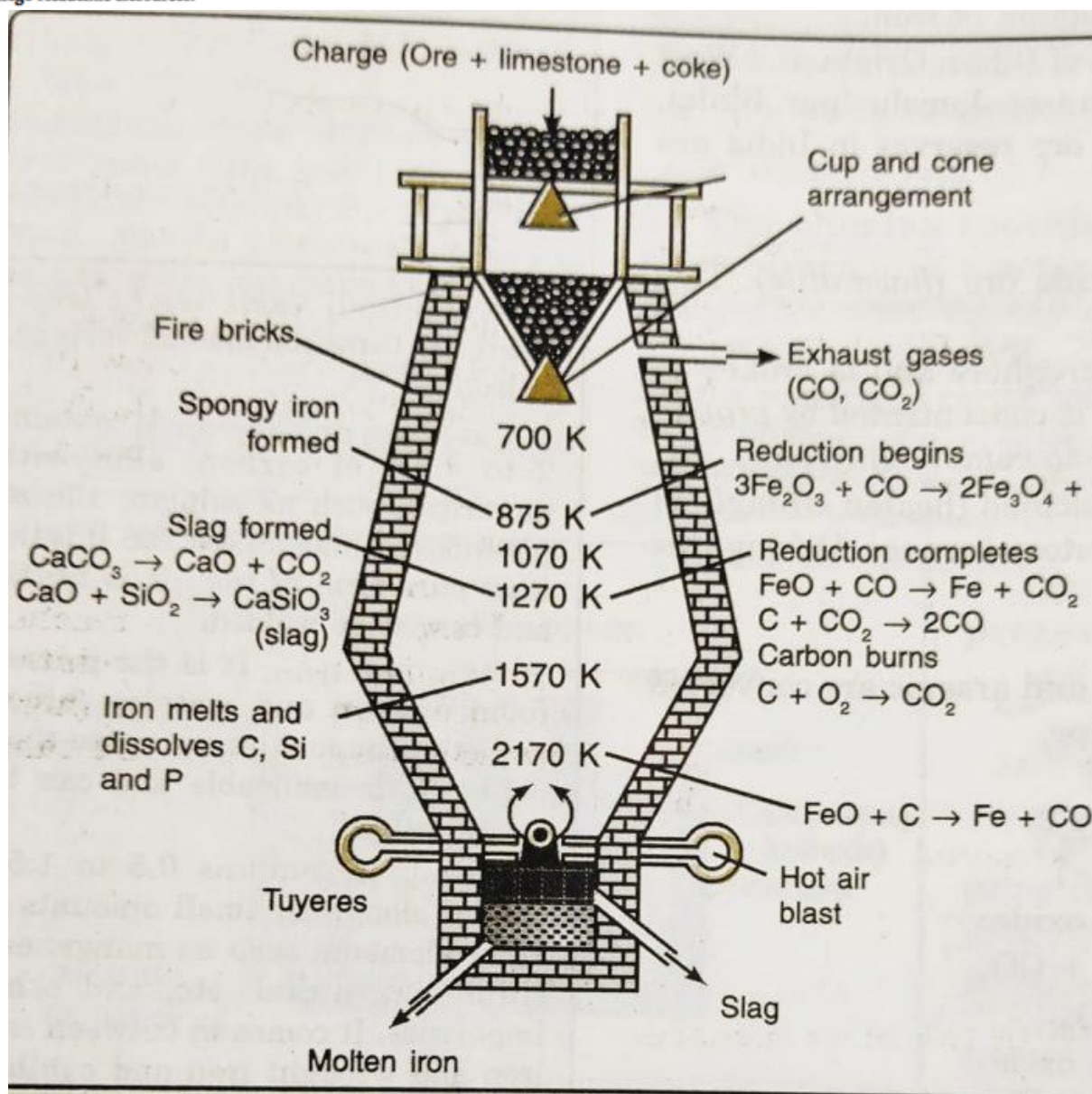
Combustion of Coke: $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ (releases heat so an exothermic reaction)

Reduction of CO_2 to CO: $\text{CO}_2 + \text{C} \rightarrow 2\text{CO}$ (endothermic reaction).

Reduction of Iron Ore by Carbon Monoxide: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$

Formation of Slag: Limestone decomposes into calcium oxide, which reacts with impurities like silica: $\text{CaO} + \text{SiO}_2 \rightarrow \text{CaSiO}_3$

Purpose of Slag: Acts as a waste product, floating on molten iron for easy removal.





Extraction of Aluminum

Process: Electrolysis of aluminum oxide (bauxite) dissolved in molten cryolite.

Aluminium Oxide has a high melting point so requires more heat; cryolite reduces this, making electrolysis feasible.

Another use of Cryolite is that it acts as a solvent for Aluminium

Electrolytic Cell Design:

Cathode (Reduction): Aluminum ions are reduced to molten aluminum.

Reaction: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$

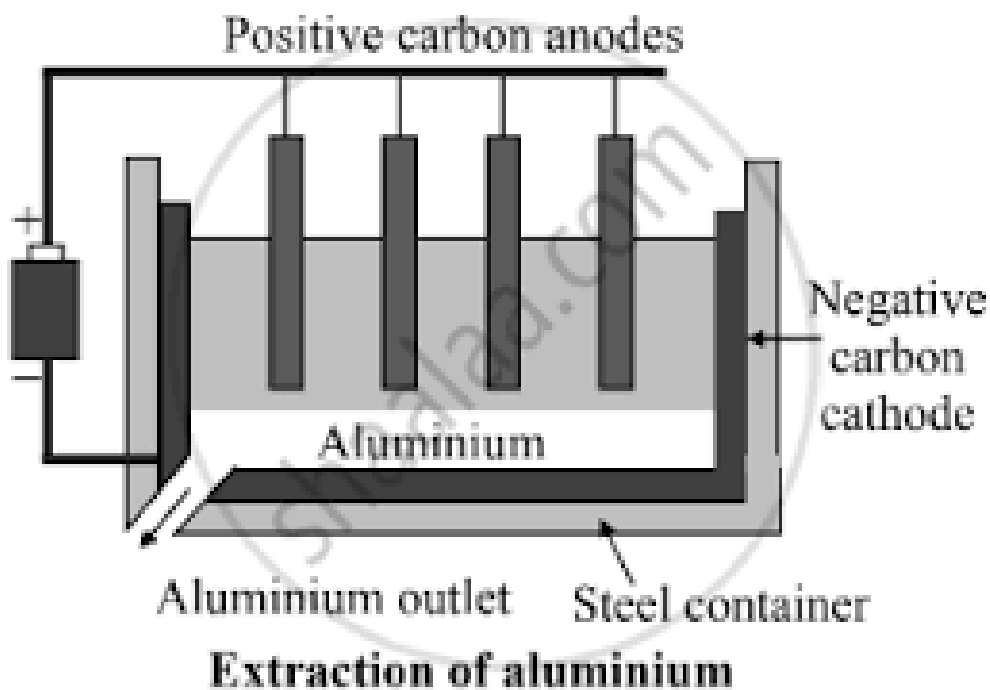
Anode (Oxidation): Oxide ions are oxidized to oxygen gas, which then reacts with carbon anodes, producing CO_2 .

Reaction: $4\text{Al}^{3+} + 12\text{e}^- \rightarrow 4\text{Al}$

Anode Consumption: Anodes made of graphite must be replaced periodically due to oxidation.

Energy and Cost Considerations:

The process is energy-intensive and costly, making aluminum relatively expensive.





The carbon anodes (made of graphite) need to be replaced after some time because the oxygen at the anode reacts with the carbon in graphite , forming carbon dioxide .

7. Corrosion and Rusting

Corrosion: Gradual degradation of metals, primarily due to reactions with water, oxygen, or acids.

Iron Corrosion (Rusting): Rusting specifically refers to the formation of iron oxide (rust) on iron.

Conditions Needed for Rusting:

Both water and oxygen are essential for rust to form on iron.

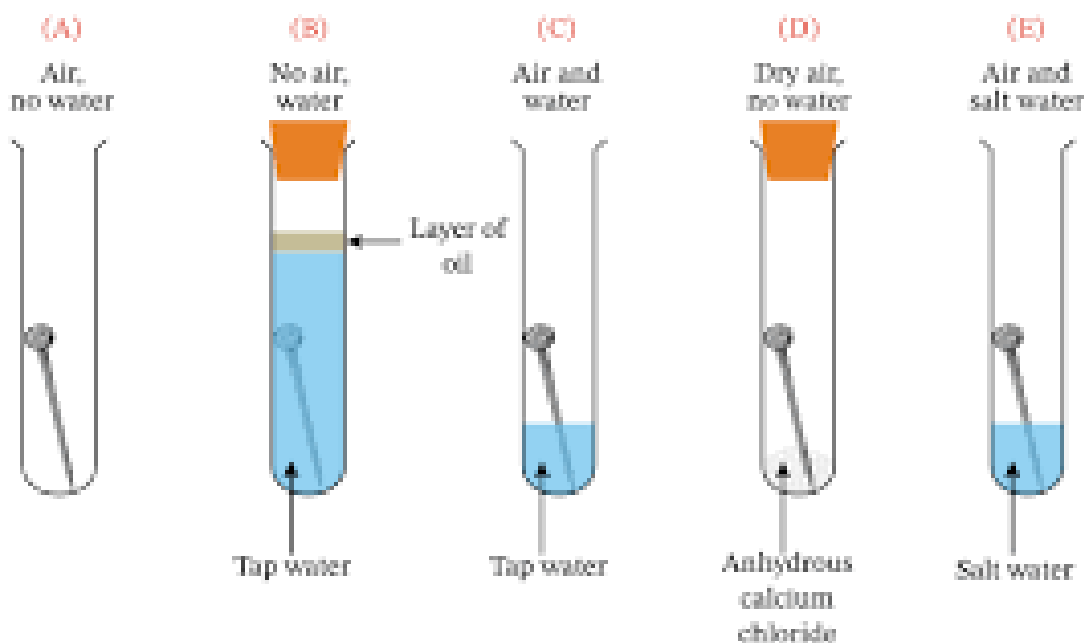
Rusting Reaction: $4\text{Fe} + 3\text{O}_2 + 6\text{H}_2\text{O} \rightarrow 2\text{Fe}_2\text{O}_3 \cdot 6\text{H}_2\text{O}$

Rust Prevention Methods:

Barrier Methods: Painting, oiling, or plastic coating to prevent oxygen and water contact.

Galvanization: Coating iron with a layer of zinc; zinc reacts first due to its position in the reactivity series.

Sacrificial Protection: Using a more reactive metal (e.g., magnesium) that corrodes in place of iron.





Past Paper Questions

P2

Which of the following metals is extracted from its ore using electrolysis?

- A) Iron
- B) Copper
- C) Aluminum
- D) Zinc

Answer: C) Aluminum

Which metal is commonly used for making electrical wires?

- A) platinum
- B) Steel
- C) Bronze
- D) Copper

Answer: D) Copper

What is the primary purpose of adding carbon to iron when making steel?

- A) To improve ductility
- B) To increase strength
- C) To reduce cost
- D) To prevent rusting

Answer: B) To increase strength

Which of the following statements about the reactivity of metals is correct?

- A) Gold is more reactive than iron.
- B) Sodium can displace copper from copper sulfate solution.
- C) Magnesium is less reactive than zinc.
- D) Potassium is less reactive than calcium.

Answer: B) Sodium can displace copper from copper sulfate solution.

In the extraction of iron from its ore in a blast furnace, which gas is produced by the reaction of coke with oxygen?

- A) Carbon dioxide
- B) Carbon monoxide
- C) Methane
- D) Hydrogen



Answer: B) Carbon monoxide

P4

1. Define what an alloy is and give two examples of common alloys.
2. Describe the main steps involved in the extraction of iron from its ore using a blast furnace. Include the role of coke and limestone in the process.
3. Give two uses of Cryolite in the extraction of aluminium .
4. Discuss the factors that influence the choice of method for extracting a metal from its ore. Include considerations such as reactivity, availability of resources, and economic factors.
5. Evaluate the environmental impact of metal extraction and processing. Discuss how these processes can affect ecosystems and suggest ways to mitigate these effects.



ACIDS BASES AND SALTS

Acids :

Acids are substances with PH values less than 7

The lower the pH of an acid, the stronger it is

The higher the pH of an acid, the weaker it is

When Acids are added to water, they form positively charged Hydrogen Ions H^+ ions also known as hydronium ions

These Ions make a solution acidic.

Bases :

Bases are substances with pH values greater than 7

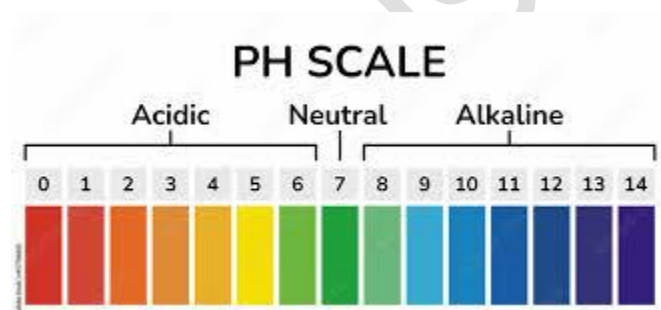
The greater the value of pH of a base, the stronger it is

The lesser the value of pH of a base, the weaker it is

Soluble bases are called alkalis

When alkalis are added to water, they form negatively charged hydroxide ions

These ions make solutions basic / alkaline



Characteristic Reactions

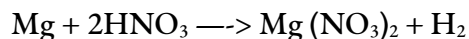
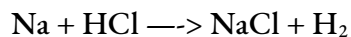
Acids and Metals :

Acid + metal $\rightarrow$ Salt + Hydrogen

The cation in the salt depends upon the metal

The Anion in the salt depends upon the acid

Examples :



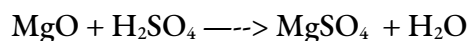
Acids and Bases :

Acid + Base $\longrightarrow$ Salt + Water

Metal Oxides and Hydroxides can act as bases

The Reaction between an acid and a base is called a neutralization reaction

Example :



Acids can also react with metal carbonates to form a salt , carbon dioxide and water

Indicators

Indicators are substances that change color when added to acidic or basic solutions, helping to determine the pH of a substance by showing whether it is acidic, basic, or neutral. Here's a breakdown of commonly included indicators for IGCSE Chemistry and their color changes in different solutions:

Litmus

Acid: Red

Neutral: Purple

Base: Blue

Methyl Orange

Acid: Red

Neutral: Yellow-Orange

Base: Yellow

Thymolphthalein

Acid: Colorless

Neutral: Colorless

Base: blue

Universal Indicator (covers a broad pH range and shows a gradient of colors)

Acid: Red to Yellow (pH 0-6)

Neutral: Green (pH 7)

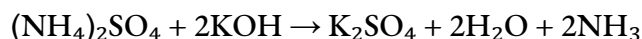
Base: Blue to Purple (pH 8-14)



Ammonium Salts :

Ammonium salts decompose when gently heated with an alkali. Despite ammonia being a weak base, its high volatility allows it to be readily displaced from the ammonium salt by another alkali, resulting in the formation of a salt, water, and ammonia gas.

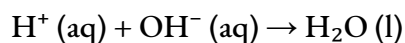
For example:



This reaction is commonly used to test for ammonium ions (NH_4^+). To perform the test, alkali is added to the sample, followed by gentle heating. Ammonia gas, if released, can be detected by holding damp red litmus paper near the reaction; the paper will turn blue, indicating the presence of ammonia gas.

Neutralization Reactions

A neutralization reaction occurs when an acid and a base react together. Acids release hydrogen ions (H^+) in solution, while bases (or alkalis) release hydroxide ions (OH^-). When combined, these ions react to form water, which has a pH of 7, resulting in a neutral solution. This is the general ionic equation for any acid-base neutralization:



Non-Neutralization Reactions with Acids:

Not every reaction involving an acid is a neutralization. For instance, when a metal reacts with an acid, a salt is produced without forming water. Such reactions do not meet the strict definition of neutralization, as they don't involve H^+ and OH^- ions producing water.

The pH Scale and Hydrogen Ion Concentration

Understanding the pH Scale:

The pH scale, ranging from 1 to 14, indicates whether a solution is acidic, alkaline, or neutral. Acids have pH values below 7, with stronger acids near 0, while alkalis have pH values above 7, with strong alkalis near 14. The strength of an acid or alkali can be classified as follows:

Strong acids: pH 0-2

Weak acids: pH 3-6

Neutral solutions: pH 7

Weak alkalis: pH 8-11

Strong alkalis: pH 12-14

Logarithmic Nature of the pH Scale:

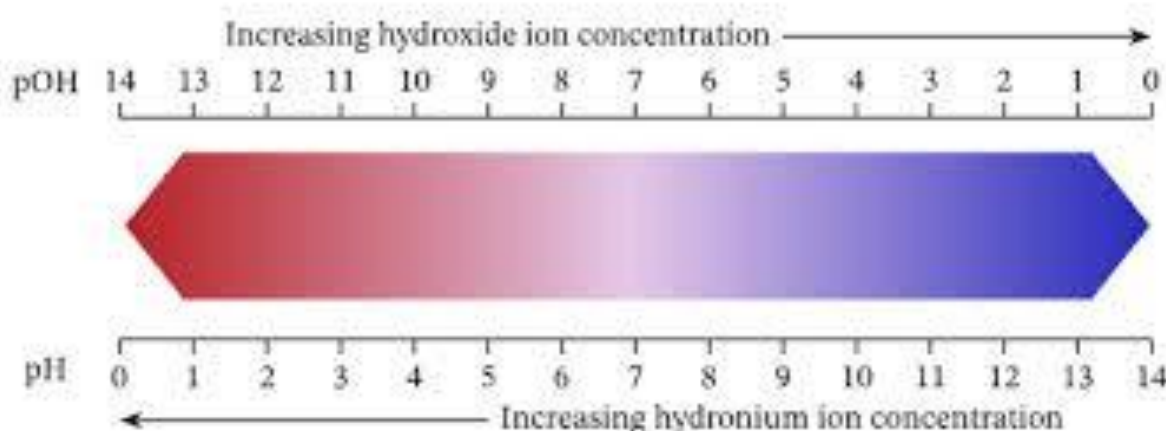
The pH scale is logarithmic, so each change of 1 in pH represents a tenfold change in H^+



concentration. For example, an acid with a pH of 2 has 100 times the H^+ concentration of an acid with a pH of 4.

Hydrogen and Hydroxide Ion Concentration:

Acidity is determined by the concentration of hydrogen ions in solution: the higher the H^+ concentration, the lower the pH. Conversely, a higher concentration of hydroxide ions increases the pH, making the solution more alkaline.



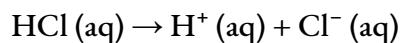
Proton Transfer in Acids and Bases

Acids are classified as proton donors because they release hydrogen ions (H^+) in solution, which contributes to the acidic nature of the solution. Conversely, bases act as proton acceptors, taking in the H^+ ions that acids release. For instance, when hydrochloric acid reacts with water, water acts as a base by accepting a proton from the acid.

Strong and Weak Acids

Definition of a Strong Acid:

A strong acid is one that fully dissociates in water, meaning it completely releases all available H^+ ions, creating a solution with a low pH. Some common examples of strong acids include hydrochloric acid (HCl) and sulfuric acid (H_2SO_4). For example, in solution:

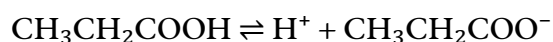


This reaction shows hydrochloric acid breaking down entirely into its ions, leading to a very acidic solution.

Definition of a Weak Acid:

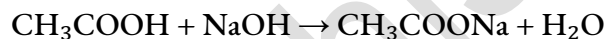
Weak acids only partially dissociate in water, meaning that only some of the acid molecules release H^+ ions, while others remain intact. This partial dissociation leads to an equilibrium between dissociated ions and undissociated acid molecules. The pH of weak acids is still below 7 but closer to neutral.

For example, with propanoic acid ($\text{CH}_3\text{CH}_2\text{COOH}$), equilibrium is established as follows:



The equilibrium favors the intact acid molecules, resulting in a lower concentration of H^+ ions in the solution.

Another example of a weak acid is ethanoic acid, which can react with an alkali like sodium hydroxide to form ethanoate salts:



Concentration vs. Strength of Acids

Understanding Acid Concentration:

The concentration of an acid refers to how many acid molecules are present per cubic decimeter (dm^3) of solution. A concentrated acid solution contains more acid molecules, while a dilute solution has fewer.

Strength vs. Concentration:

It's important to note that concentration and strength are not the same. Strength refers to how completely an acid dissociates in water, while concentration refers to the number of acid molecules



in a given volume. For example, a dilute solution of hydrochloric acid (HCl) is often more acidic than a concentrated solution of ethanoic acid (CH₃COOH), because even in lower amounts, HCl fully dissociates, whereas CH₃COOH does not.

Oxides

Definition of Oxides

Oxides are compounds consisting of oxygen atoms bonded with atoms of another element. Some common oxides include calcium oxide (CaO), nitrogen dioxide (NO₂), phosphorus pentoxide (P₂O₅), sulfur trioxide (SO₃), and water (H₂O). Oxides are generally classified according to their acidic or basic nature.

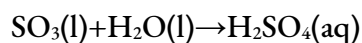
Types of Oxides

Acidic Oxides

Acidic oxides are produced when non-metal elements react with oxygen. They generally react with bases to form salt and water. When dissolved in water, acidic oxides create acidic solutions, lowering the pH. Examples of acidic oxides include sulfur trioxide (SO₃), nitrogen dioxide (NO₂), and phosphorus pentoxide (P₂O₅).

Example: Sulfur Trioxide (SO₃)

Reaction:

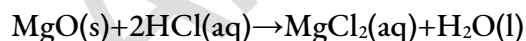


Basic Oxides

Basic oxides form when metals combine with oxygen, often producing a high pH solution when dissolved in water. They react with acids to produce salt and water. Examples of basic oxides are potassium oxide (K₂O) and magnesium oxide (MgO).

Example: Magnesium Oxide (MgO)

Reaction:

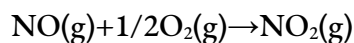


Neutral Oxides

Neutral oxides neither react with acids nor bases and thus are classified as neutral. They do not alter the pH of water significantly. Common examples of neutral oxides include nitrous oxide (N₂O), nitric oxide (NO), and carbon monoxide (CO).

Example: Nitric Oxide (NO)

Reaction:

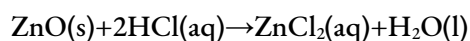


Amphoteric Oxides

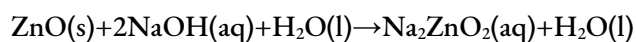
Amphoteric oxides are unique oxides that can react with both acids and bases, forming salt and water in each case. This dual behavior makes them versatile in chemical reactions. Zinc oxide (ZnO) and lead(II) oxide (PbO) are common amphoteric oxides.

Example: Zinc Oxide (ZnO)

Reaction with Acid:

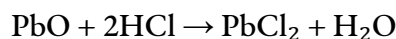


Reaction with Base:

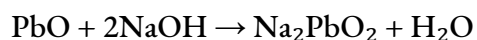


Example Reactions of Lead(II) Oxide (PbO):

As a base with an acid



As an acid with a base



Salts

A salt is a compound formed when the hydrogen ion (H^+) in an acid is replaced by a metal ion. For example, if we replace the hydrogen in hydrochloric acid (HCl) with a sodium atom, the salt sodium chloride (NaCl) is produced.

Salts play an essential role in various applications, including fertilizers, batteries, cleaning agents, healthcare, and fungicides.

Naming Salts

Salts are named based on the reactants involved:

1. First Part – Derived from the metal, metal oxide, or metal carbonate that reacts.
2. Second Part – Based on the acid used in the reaction.

Examples:

- Sodium hydroxide + hydrochloric acid $\rightarrow$ Sodium chloride
- Zinc oxide + sulfuric acid $\rightarrow$ Zinc sulfate

Acid naming conventions:



- Hydrochloric acid produces salts ending in chloride (Cl^-).
- Sulfuric acid produces salts ending in sulfate (SO_4^{2-}).
- Nitric acid produces salts ending in nitrate (NO_3^-)

Salt Solubility Table

Soluble salts		Insoluble salts	
All sodium salts			
All potassium salts			
All ammonium salts			
All nitrates			
All chlorides		Silver chloride Lead(II) chloride	
All sulphates		Barium sulphate Lead(II) sulphate Calcium sulphate (sparingly soluble)	
Sodium carbonate Potassium carbonate Ammonium carbonate		All carbonates	

Preparing Soluble Salts

There are two main methods for preparing soluble salts, depending on the nature of the reactants:

Method 1: Adding Acid to an Insoluble Substance (metal, base, or carbonate)

Procedure:

1. Heat dilute acid in a beaker.
2. Gradually add an insoluble base, carbonate, or metal until no more dissolves and the excess is visible in the beaker.
3. Filter the mixture to remove undissolved solids.
4. Heat the filtered solution to evaporate water and achieve saturation.
5. Allow the solution to cool for crystallization.
6. Collect the crystals, decant excess solution, and dry with filter paper.

Example: Preparing Copper(II) Sulfate Crystals

- Combine sulfuric acid with copper(II) oxide (CuO).
- $\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$

Method 2: Titration of a Dilute Acid and Alkali (soluble base)

Procedure:



1. Measure the alkali into a conical flask and add a few drops of indicator (e.g., methyl orange or thymolphthalein).
2. Fill a burette with acid and record the initial volume.
3. Slowly add acid from the burette to the flask until the indicator changes color.
4. Record the final burette reading to find the volume of acid used.
5. Repeat the procedure without the indicator, adding the same volume of acid to the alkali.
6. Heat the solution to partially evaporate water, forming a saturated solution.
7. Let the solution cool and form crystals, then decant excess liquid and dry the crystals.

Preparing Insoluble Salts

Insoluble salts can be made through a precipitation reaction. In this method, an insoluble salt is formed as a solid precipitate, so it's essential that the salt product is insoluble in water, while the reactants are soluble.

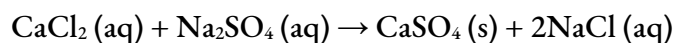
Using Two Soluble Reactants

Method:

1. Dissolve each of the soluble salts in water separately in two beakers.
2. Combine both solutions in a single beaker and stir. An insoluble salt will form as a precipitate.
3. Filter the mixture to separate the solid precipitate from the solution.
4. Rinse the residue with distilled water to eliminate any leftover traces of other solutions.
5. Dry the precipitate by placing it in an oven, leaving behind the purified insoluble salt.

Example: Preparing Pure, Dry Calcium Sulfate Crystals

1. Dissolve calcium chloride (CaCl_2) in water in one beaker.
2. Dissolve sodium sulfate (Na_2SO_4) in water in another beaker.
3. Pour the two solutions into a single beaker, stirring to create the precipitate, calcium sulfate (CaSO_4).
4. Filter the mixture to separate the calcium sulfate precipitate.
5. Rinse the precipitate with distilled water to wash away any sodium chloride (NaCl) solution.
6. Dry the calcium sulfate by leaving it in an oven.





Practice Questions (Solved)

1. Explain the difference between an alkali and a base, and give examples of each.

Answer:

A base neutralizes acids and is insoluble , while an alkali is a base that dissolves in water to release OH^- ions. For example, NaOH is an alkali, and CaCO_3 is a base.

2. Describe how you would carry out a titration to find the concentration of a sulfuric acid solution using a standard solution of sodium hydroxide and an appropriate indicator.

Answer:

Measure NaOH with a pipette, add indicator, and titrate with H_2SO_4 from a burette until the endpoint. Record the volume to calculate H_2SO_4 concentration.

3. Explain the amphoteric nature of aluminum oxide and describe its reactions with both acids and bases.

Answer:

Al_2O_3 is amphoteric; it reacts with acids to form salts and with bases to form aluminate.

Practice Questions (Unsolved)

P2

1. Which of the following substances will turn red litmus paper blue?

- A) Hydrochloric acid
- B) Sodium hydroxide
- C) Water
- D) Ethanoic acid

2. Which of these reactions is a neutralization reaction?

- A) $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- B) $\text{HCl} + \text{Mg} \rightarrow \text{MgCl}_2 + \text{H}_2$
- C) $\text{HCl} + \text{ZnCO}_3 \rightarrow \text{ZnCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$
- D) $\text{H}_2\text{SO}_4 + \text{Cu} \rightarrow \text{CuSO}_4 + \text{H}_2$

3. What type of oxide is formed when a non-metal reacts with oxygen?

- A) Basic oxide
- B) Acidic oxide



- C) Amphoteric oxide
- D) Neutral oxide

4. What is the name of the salt formed when sulfuric acid reacts with potassium hydroxide?

- A) Potassium sulfate
- B) Potassium nitrate
- C) Potassium chloride
- D) Potassium carbonate

5. Which of the following indicators would show a color change in an acidic solution?

- A) Phenolphthalein turning pink
- B) Universal indicator turning green
- C) Litmus turning red
- D) Methyl orange turning yellow

P4

1. Describe how you would prepare a pure sample of an insoluble salt such as barium sulfate in the laboratory, including the names of reagents used and the steps involved.
2. Explain the concept of strong and weak acids in terms of ionization, and provide examples of each. Describe how you could differentiate between them experimentally.
3. Describe what happens when ammonia gas dissolves in water. Write an equation for this reaction and explain the role of ammonia in terms of proton transfer.
4. Explain the purpose of using indicators in titration experiments. Describe how you would carry out a titration experiment to determine the concentration of a hydrochloric acid solution using sodium hydroxide and an appropriate indicator.



Organic Chemistry

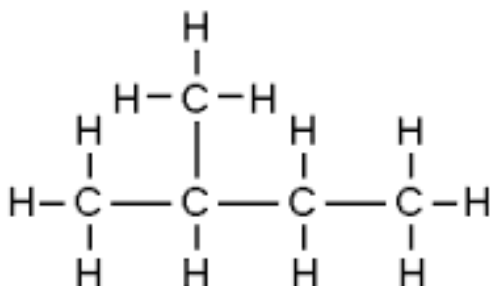
Organic Chemistry studies the structure, properties, and reactions of organic compounds (those containing carbon).

Exceptions: Metal carbonates, CO_2 , and CO are not considered organic compounds.

Hydrocarbons

Compounds containing only hydrogen and carbon atoms.

Displayed Formulae - Shows spatial arrangement of all atoms and bonds.



General Formulae - Shows composition for a homologous series.

Structural Formulae - Shows key bonds but omits most covalent bonds.

Example: Ethene : $\text{CH}_2 = \text{CH}_2$

Structural Isomers

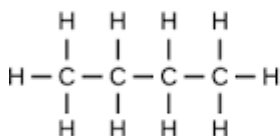
Compounds with the same molecular formula but different atom arrangements.

Example of Structural Isomers

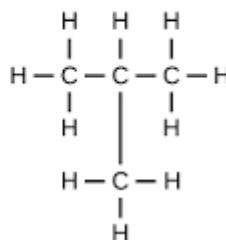


Structural Isomers

Butane



Isobutane



Homologous Series

A family of organic compounds with similar features due to the same functional group.

Functional group is responsible for characteristic reactions.

Functional Groups and Endings:

Family	Functional group	Name
Alkane	$\begin{array}{c} \quad \\ -\text{C}-\text{C}- \\ \quad \end{array}$	-ane
Alkene	$\begin{array}{c} \quad \\ \text{C}=\text{C} \\ \quad \end{array}$	-ene
Alcohol	-OH	-anol
Carboxylic acid	$\begin{array}{c} \text{O} \\ \\ -\text{C}-\text{OH} \end{array}$	-anoic acid
Amine	-NH ₂	-amine
Ester	$\begin{array}{c} \text{O} \\ \\ -\text{C}-\text{O}- \end{array}$	-yl-anoate



General Formulae of Common Homologous Series

- Alkane: $C_n H_{2n+2}$
- Alkene: $C_n H_{2n}$
- Alcohol: $C_n H_{2n+1}OH$
- Carboxylic Acid: $C_n H_{2n+1}COOH$

Saturated vs. Unsaturated Compounds

Saturated Compounds: Contain only C-C single bonds (e.g., alkanes, $C_n H_{2n+2}$).

Unsaturated Compounds: Contain one or more C=C double bonds (e.g., alkenes, $C_n H_{2n}$).

Naming Organic Compounds

Name Parts:

Stem - Based on the longest carbon chain.

Suffix - Based on functional group.

Stem Based on Chain Length:

Prefixes

No. of carbon atoms	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-



Suffix Based on Functional Group:

Functional Group Name	Suffix Ending	Functional Group Structure
Alkane	-ane	C-H atoms
Alcohol	-ol	--OH
Alkene	-ene	C=C

Structure	Classification	Suffix Name	Substituent Name
higher priority ▲ $\text{R}-\text{C}(=\text{O})\text{OH}$	Carboxylic acid	-oic acid	carboxy-

Additional Naming Notes

Functional Group Location:

Number the chain from the end that gives the smallest number for the functional group.

Example: Propan-1-ol has -OH on the first carbon.

Common Organic Families

Alkanes

Contain only C–C single bonds.

Example: Methane, Ethane.

Alkenes

Alcohols

Contain -OH group, suffix “-anol.”



Example: Ethanol, Propan-1-ol.

Carboxylic Acids

Contain -COOH group, suffix “-anoic acid.”

Example: Ethanoic acid, Butanoic acid.

Esters

Formed from alcohol and carboxylic acid, suffix “-yl-anoate.”

Example: Ethyl ethanoate.

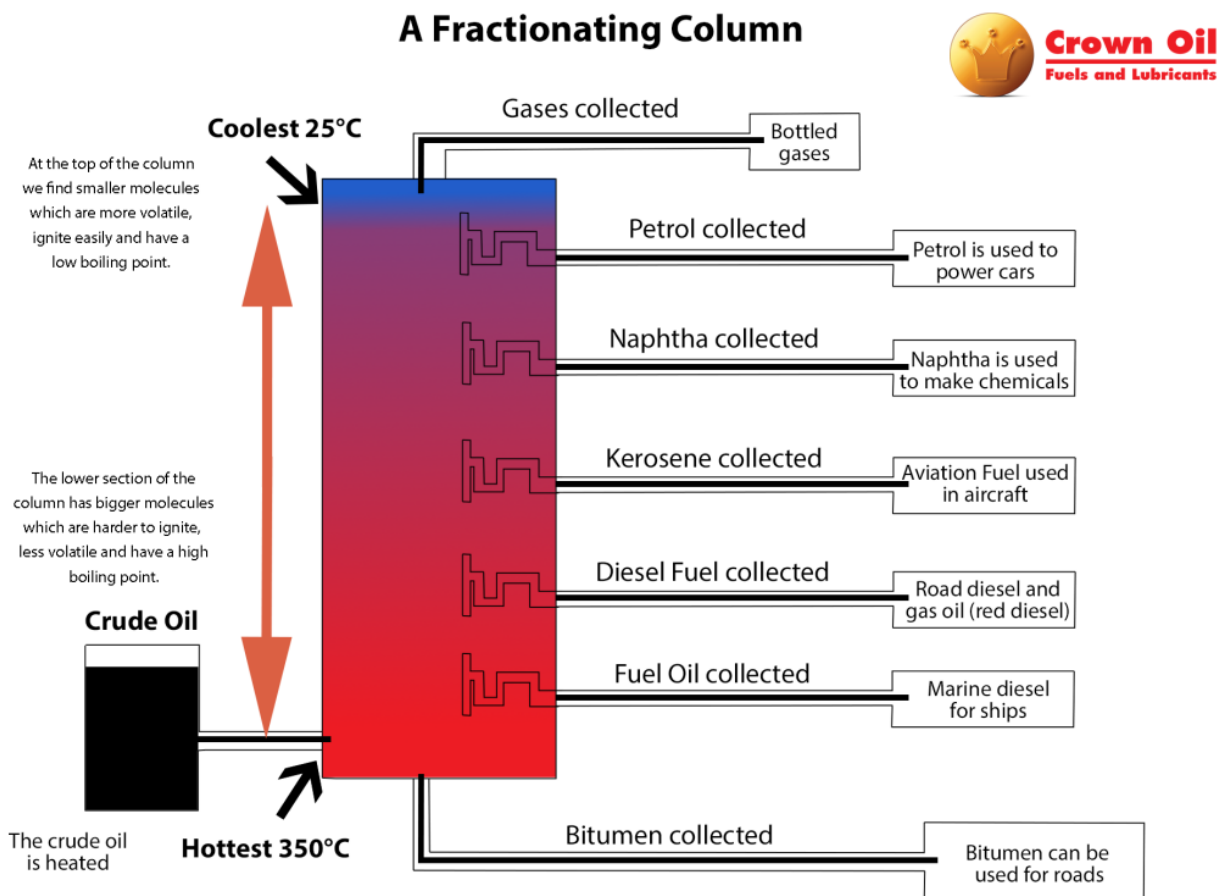
Petroleum and Fractional Distillation

Petroleum (crude oil) is a mixture of hydrocarbons, separated by fractional distillation.

- Fractional Distillation Process: [fractional distillation of crude oil diagram]



Properties and Uses of Fractions



Alkanes: Properties & Bonding

- Alkanes are saturated hydrocarbons, meaning they contain only single carbon-carbon (C-C) bonds.
- Alkanes have covalent bonds.
- General Formula: $C_n H_{2n+2}$.



Molecular Name	Formula	Melting Point (°C)	Boiling Point (°C)	Density (20°C) [†]	Physical State (at 20°C)
methane	CH ₄	-182	-164	0.668 g/L	gas
ethane	C ₂ H ₆	-183	-89	1.285 g/L	gas
propane	C ₃ H ₈	-190	-42	1.867 g/L	gas
butane	C ₄ H ₁₀	-138	-1	2.493 g/L	gas
pentane	C ₅ H ₁₂	-130	36	0.626 g/mL	liquid
hexane	C ₆ H ₁₄	-95	69	0.659 g/mL	liquid
octane	C ₈ H ₁₈	-57	125	0.703 g/mL	liquid
decane	C ₁₀ H ₂₂	-30	174	0.730 g/mL	liquid

Physical Properties

- Colorless compounds with increasing boiling and melting points as the carbon chain length increases.

Chemical Properties

Combustion: Alkanes react with oxygen to release energy.

Catalytic Cracking

Substitution with halogens under UV light

(Complete) Combustion of Alkanes

Alkane + Oxygen → Carbon Dioxide + Water

Examples:

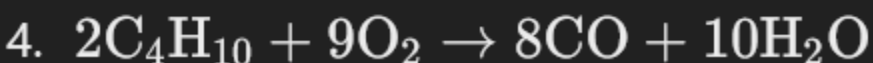
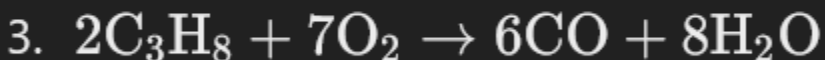
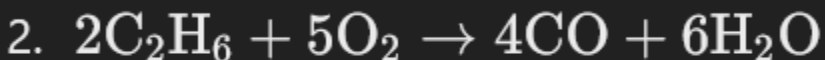
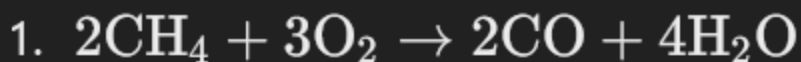
1. Methane: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
2. Ethane: $\text{C}_2\text{H}_6 + 3.5\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$
3. Propane: $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$
4. Butane: $\text{C}_4\text{H}_{10} + 6.5\text{O}_2 \rightarrow 4\text{CO}_2 + 5\text{H}_2\text{O}$

(Incomplete) Combustion of Alkanes

Alkane + Oxygen → Carbon Monoxide + Water



Examples :



Substitution Reactions with Halogens

Alkanes react with halogens in the presence of ultraviolet light.

The Halogen atom replaces a hydrogen atom bonded with a carbon in the chain

Given that the halogen is in excess, one by one, all the hydrogens can be replaced

So Methane would, after 4 reactions, form Carbon Tetrachloride

Example: $\text{CH}_4 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{HCl}$ (Methane reacts with chlorine to form chloromethane and hydrogen chloride)

Cracking of Alkanes

Cracking is the process of breaking down long-chain alkanes into shorter alkanes and alkenes. This is done to increase the supply of shorter, more useful hydrocarbons and produce alkenes, which are valuable for making plastics.

Types of Cracking:

Catalytic Cracking: Lower temperatures (450°C) and slight pressure with a catalyst (zeolite).

Produces mainly branched alkanes and aromatic hydrocarbons.

Example Reaction:

Hexane (C_6H_{14}) $\rightarrow$ Butane (C_4H_{10}) + Ethene (C_2H_4)



Purpose: Cracking produces alkenes for polymers and provides fuels for gasoline, which burns more efficiently in car engines.



Alkenes

Properties & Bonding

- Alkenes are unsaturated hydrocarbons containing at least one C=C double bond.
- General Formula: $C_n H_{2n}$

No. of C atoms	Name of alkene	Molecular formula
2	Ethene	C_2H_4
3	Propene	C_3H_6
4	Butene	C_4H_8
5	Pentene	C_5H_{10}
6	Hexene	C_6H_{12}
7	Heptene	C_7H_{14}
8	Octene	C_8H_{16}

Reactivity

- Alkenes are reactive and undergo addition reactions.

Distinguishing Between Alkanes and Alkenes

- Bromine Water Test:

- Alkanes: Solution remains orange.
- Alkenes: Decolorizes bromine water (orange to colorless).

Addition Reactions of Alkenes

Hydrogenation (Addition of Hydrogen)



Reaction: Alkenes react with hydrogen to form alkanes.

Conditions: Nickel catalyst at 150°C.

Example: $\text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6$

Use: Used to hydrogenate vegetable oils, converting unsaturated fats to saturated fats.

Addition of Halogens (e.g., Bromination)

Reaction: Alkenes react with halogens (e.g., bromine, chlorine) to form dihaloalkanes.

Observation: Bromine water decolorizes (orange-brown to colorless), indicating the presence of alkenes.

Example: $\text{C}_2\text{H}_4 + \text{Br}_2 \rightarrow \text{C}_2\text{H}_4\text{Br}_2$

Hydration (Addition of Steam)

Reaction: Alkenes react with steam to form alcohols.

Conditions: Requires high temperature (300°C), high pressure (60-70 atm), and phosphoric acid catalyst.

Example: $\text{C}_2\text{H}_4 + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_5\text{OH}$

Use: Industrially used to produce ethanol from ethene.

Alcohols

- Functional Group: -OH

- General Formula: $\text{C}_n \text{H}_{2n+1}\text{OH}$

Naming Examples:

- Methanol: CH_3OH

- Ethanol: $\text{C}_2\text{H}_5\text{OH}$

- Propanol: $\text{C}_3\text{H}_7\text{OH}$

Key Alcohol: Ethanol

- Combustion: $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$

Formation of Carboxylic Acids from Alcohols

Oxidation Using Oxidizing Agents

Reaction: Alcohols react with oxidizing agents (e.g., potassium dichromate in acidic conditions) to produce carboxylic acids.

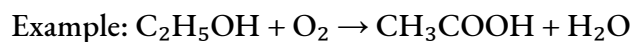
Conditions: Heat with acidified potassium dichromate (H_2SO_4).

Example: $\text{C}_2\text{H}_5\text{OH} + [\text{O}] \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$

Microbial Oxidation



Reaction: Alcohols, such as ethanol, can be oxidized by microbial action (e.g., acetic acid bacteria) in the presence of oxygen.



Use: Commonly used in vinegar production.

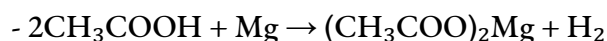
Carboxylic Acids

General Properties

- Contain the -COOH functional group.
- React with metals, carbonates, and bases.

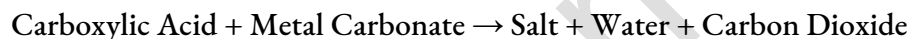
Reaction Example

- Ethanoic acid + Magnesium $\rightarrow$ Magnesium ethanoate + Hydrogen



Reactions of Carboxylic Acids with Carbonates

Carboxylic acids react with metal carbonates to form salt, water, and carbon dioxide. This reaction is used to test for carboxylic acids.



Example Reaction with Ethanoic Acid:



Observation: Effervescence (bubbling) occurs due to the release of CO_2 gas, which can be confirmed by bubbling it through limewater, causing it to turn cloudy.

Esters

1. Definition and Structure

Esters are organic compounds formed by the reaction of a carboxylic acid and an alcohol. They typically have pleasant, fruity smells and are used in perfumes and food flavorings.

Functional Group: -COO- .

General Structure: R-COO-R' , where "R" represents an alkyl group from the acid and "R'" represents an alkyl group from the alcohol.

2. Formation of Esters (Esterification Reaction)

Esterification Reaction: The process of forming an ester is known as esterification. In this reaction, a carboxylic acid reacts with an alcohol in the presence of a catalyst (typically concentrated sulfuric acid).



General Equation: Carboxylic Acid+Alcohol→Ester+Water

Example Reaction:

Ethanoic Acid + Ethanol → Ethyl Ethanoate + Water

$\text{CH}_3\text{COOH} + \text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O}$

Catalyst: Concentrated sulfuric acid (H_2SO_4) is used to speed up the reaction and shift the equilibrium towards ester formation.

3. Properties of Esters

Physical Properties:

Esters are generally colorless, volatile liquids with distinctive fruity or floral smells.

They are typically insoluble in water due to their lack of hydrogen bonding capability but are soluble in organic solvents.

Boiling Points: Esters have lower boiling points than the corresponding acids due to the absence of hydrogen bonds.

4. Uses of Esters

Esters are used for their pleasant odors and flavors in perfumes, foods, and drinks.

Some esters (like ethyl ethanoate) are used as solvents in products like nail polish remover and paints.

Esters are used as plasticizers to make plastics like PVC flexible and durable.

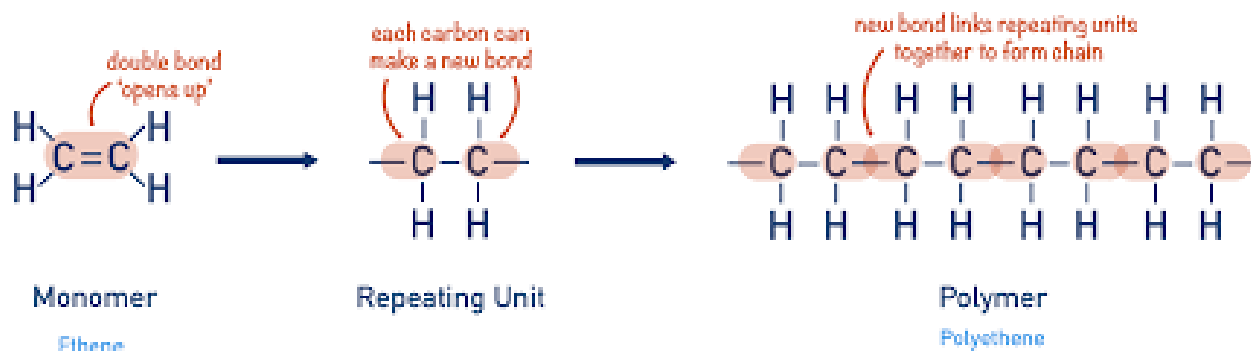
: Esters are found in some drugs due to their ability to be hydrolyzed by enzymes in the body.

Polymers

- Polymers are large molecules formed by linking 50 or more monomers.
- Each repeat unit is connected by covalent bonds.
- Examples of polymers include PVC and nylon.

Formation of Poly(ethene) from Ethene

- Poly(ethene) is formed by the addition polymerisation of ethene monomers.
- Addition polymerisation involves the joining of multiple monomers to form a long-chain polymer.
- In ethene, the $\text{C}=\text{C}$ double bond breaks, allowing the monomers to join and form poly(ethene).



Repeat Unit of Poly(ethene)

- Poly(ethene) has a repeat unit formed by addition polymerisation from ethene monomers.

Addition Polymers

Key Characteristics

- Addition polymers are formed from alkene monomers with $C=C$ bonds.
- One bond in the $C=C$ breaks, allowing bonding with adjacent monomers.
- The polymer structure consists of single bonds between carbon atoms.
- Many addition polymers are derived from alkenes with additional groups like chlorine (as in PVC) or hydroxyl (OH).

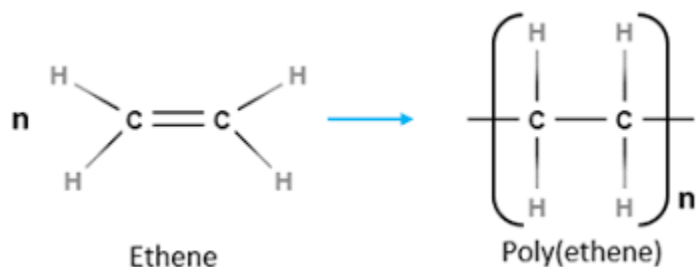
Naming Addition Polymers

- Named by adding “poly-” before the monomer’s name.
- Example: Poly(propene) from propene.

Deducing the Polymer from the Monomer

To draw a repeat unit:

1. Change the double bond in the monomer to a single bond in the repeat unit.
2. Add continuation bonds extending outside the brackets on each end.
3. Add a small n at the bottom right to indicate many repeat units.
4. Place groups in the same order around the bond as in the monomer.



Deducing the Monomer from the Polymer

To identify the monomer:

1. Find the repeating unit in the polymer.
2. Change the single bond in the repeat unit back to a double bond for the monomer.
3. Remove the continuation bonds.

Plastics & Their Disposal

- Plastics are made from polymers and are non-biodegradable, meaning they don't break down easily, leading to environmental issues.

Incineration

- Burning plastics releases heat and carbon dioxide (CO_2), a greenhouse gas.
- Toxic fumes may be released by some polymers:
 - Poly(vinyl chloride) (PVC) releases hydrogen chloride gas when burned.
- Incomplete combustion can release carbon monoxide (CO), a toxic gas.

Environmental Impact

1. Polluting Oceans: Plastic waste disrupts marine ecosystems.
2. Landfills: Non-biodegradable polymers accumulate, taking up valuable land as they do not decompose.

Disposal of Polymers

PET Re-Polymerisation



- PET (polyethylene terephthalate) is a condensation polymer used for plastic bottles.
- Recycling PET is challenging as high temperatures and pressures degrade the monomers.
- PET can be depolymerised back into monomers using enzymes or chemical catalysts at mild conditions.
- Recovered monomers are re-polymerised, saving resources and energy.



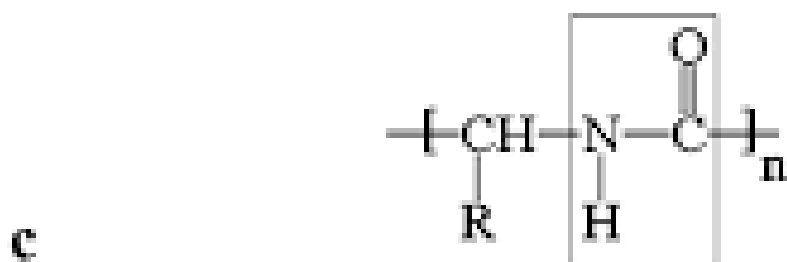
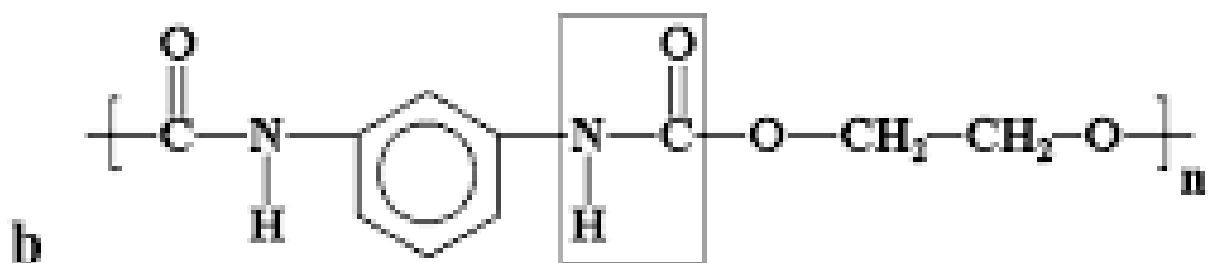
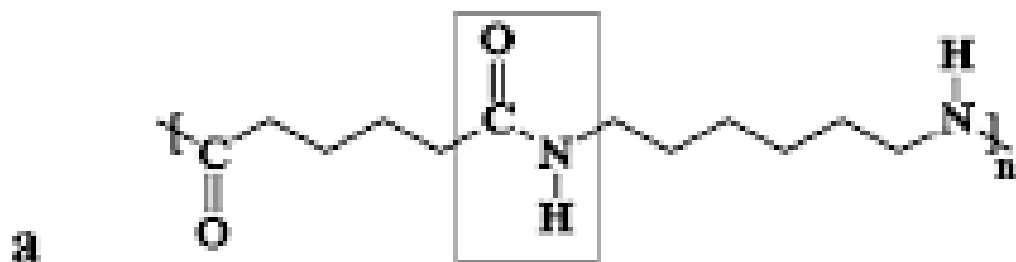
Condensation Polymers

Condensation polymers form when two different monomers link, releasing a small molecule, usually water.

Key Difference: Unlike addition polymers, condensation polymerisation produces a polymer and water.

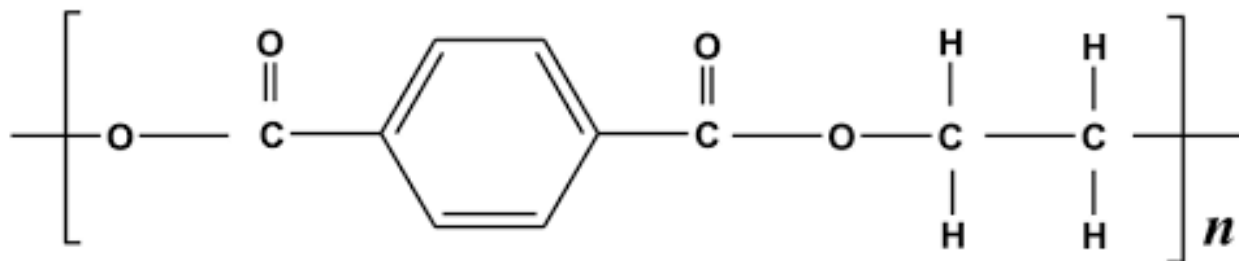
Forming Nylon

- Nylon is a polyamide formed from dicarboxylic acid (with two -COOH groups) and diamines (with two -NH_2 groups).
- Each -COOH group reacts with an -NH_2 group on another monomer, forming an amide linkage and releasing water.



Forming Polyesters (Example: PET)

- PET (polyethylene terephthalate) is formed from dicarboxylic acids (-COOH groups) and diols (-OH groups).
- Ester linkages form with the loss of water.
- PET is widely used in synthetic fibers and sold under trade names like terylene.



Proteins

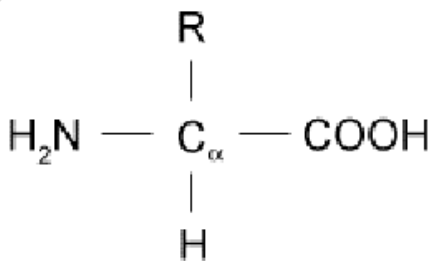
Condensation polymers

Formed from monomers of amino acids

Joined via amide links .

Amino acids are much smaller , containing amide (NH₂) and Carboxylic acid (COOH) groups .

Structure of an amino acid (General)



Proteins are formed from tens to hundreds of amino acid molecules

One water molecule is released when the Carboxylic acid functional group of one amino acid and the NH₂ (amide) group of the other amino acid combine



Practice questions (solved)

1. What is a homologous series, and what is a key characteristic that all members share?

A homologous series is a group of organic compounds that have the same functional group and similar chemical properties. A key characteristic they share is that each member differs from the next by a CH_2 unit.

2. How can you distinguish between alkanes and alkenes using a chemical test?

You can distinguish between alkanes and alkenes by using bromine water. Alkenes will decolorize bromine water (orange to colorless), while alkanes will not cause any change in the color of the bromine water.

3. What is the general formula for carboxylic acids, and how do they differ from alcohols?

The general formula for carboxylic acids is $\text{C}_n \text{H}_{2n+1} \text{COOH}$. They differ from alcohols, which have the general formula $\text{C}_n \text{H}_{2n+1} \text{OH}$, by the presence of the carboxyl group ($-\text{COOH}$) instead of the hydroxyl group ($-\text{OH}$).

4. Describe one common use of esters in everyday life.

Esters are commonly used as flavoring agents and fragrances in food and cosmetic products, providing characteristic fruity or floral scents.

5. What effect does increasing chain length have on the boiling point of alkanes?

Increasing the chain length of alkanes typically results in a higher boiling point due to increased van der Waals forces (intermolecular forces) between the longer hydrocarbon chains.

Practice Questions (Unsolved)

P2 :

1. Which of the following compounds is an alkane?

A) C_2H_4

B) $\text{C}_3\text{H}_6\text{O}$

C) C_4H_{10}

D) $\text{C}_6\text{H}_{12}\text{O}_6$



2. What is the functional group present in alcohols?

- A) -COOH
- B) -NH_2
- C) -OH
- D) -C=O

3. Which of the following reactions is an example of an addition reaction?

- A) Ethene + Water $\rightarrow$ Ethanol
- B) Ethanol + Acid $\rightarrow$ Ester + Water
- C) Propane + Chlorine $\rightarrow$ Chloropropane + Hydrogen Chloride
- D) Butanoic Acid + Sodium Bicarbonate $\rightarrow$ Sodium Butanoate + Water + Carbon Dioxide

4. Which of the following statements is true about alkenes?

- A) Alkenes are saturated hydrocarbons.
- B) Alkenes do not react with bromine water.
- C) Alkenes can undergo hydrogenation to form alkanes.
- D) Alkenes have higher boiling points than their corresponding alkanes.

5. What is the product of the reaction between a carboxylic acid and an alcohol in the presence of an acid catalyst?

- A) Amide
- B) Ester
- C) Alkane
- D) Ether

P4 :

1. Explain the difference between saturated and unsaturated hydrocarbons, providing examples of each.

2. Describe the process of esterification, including the reactants, conditions, and the type of bond formed in the product.

3. Outline the steps involved in the complete combustion of an alkane and the general products formed.



4. Discuss the methods of producing alcohols from alkenes, including the specific conditions and catalysts required.
5. Compare the properties and uses of alkanes and alkenes, highlighting at least two key differences between the two classes of hydrocarbons.

Aitchison College



Identification of Ions And Gasses

*It is preferable to memorize these tables

For the identification of cations and ions, we react them with different substances under different conditions and either base it on our observations of, for example the precipitate, or test the gas(es) released or both.

The Following tables show the tests for cations, anions, gases and Flame tests for metals



Flame test (Metals)

Metal Ion solution	Solution colour	Observations
Na^+	colourless	Flame turns orange
K^+	colourless	Flame turns purple
Ca^{2+}	colourless	Flame turns reddish orange
Sr^{2+}	colourless	Flame turns red
Ba^{2+}	colourless	Flame turns yellowish green
Cu^{2+} (sulfate solution)	deep blue	Flame turns green
Cu^{2+} (MAC solution)	dark red	Flame turns green
Sn^{2+}	clear/yellow	No change
$\text{Na}^+ / \text{Ba}^{2+}$	colourless	Flame turns orange
$\text{K}^+ / \text{Sr}^{2+}$	colourless	Flame turns red
$\text{Ca}^{2+} / \text{Cu}^{2+}$	deep blue	Flame turns green
$\text{Cu}^{2+} / \text{Sn}^{2+}$	deep blue	Flame turns green



Cation Tests

METAL CATION	EFFECT OF ADDING NaOH	EFFECT OF ADDING AMMONIA SOLUTION
ALUMINIUM (Al ³⁺)	WHITE PRECIPITATE, DISSOLVES IN EXCESS NaOH TO FORM A COLOURLESS SOLUTION	WHITE PRECIPITATE, INSOLUBLE IN EXCESS AMMONIA, WHITE PRECIPITATE REMAINS
AMMONIUM (NH ⁴⁺)	AMMONIA PRODUCED IF WARMED	–
CALCIUM (Ca ²⁺)	WHITE PRECIPITATE, INSOLUBLE SO REMAINS IN EXCESS NaOH	VERY FAINTLY VISIBLE WHITE PRECIPITATE
CHROMIUM (III) (Cr ³⁺)	GREEN PRECIPITATE WHICH FORMS A GREEN SOLUTION IN EXCESS	GREY – GREEN PRECIPITATE, INSOLUBLE IN EXCESS
COPPER (II) (Cu ²⁺)	LIGHT BLUE PRECIPITATE, INSOLUBLE IN EXCESS	LIGHT BLUE PRECIPITATE, SOLUBLE IN EXCESS TO FORM DARK BLUE COLOUR
IRON (II) (Fe ²⁺)	GREEN PRECIPITATE, INSOLUBLE IN EXCESS	GREEN PRECIPITATE, INSOLUBLE IN EXCESS
IRON (III) (Fe ³⁺)	RED – BROWN PRECIPITATE, INSOLUBLE IN EXCESS	RED – BROWN PRECIPITATE, INSOLUBLE IN EXCESS
ZINC (Zn ²⁺)	WHITE PRECIPITATE, DISSOLVES IN EXCESS TO FORM COLOURLESS SOLUTION	WHITE PRECIPITATE, DISSOLVES IN EXCESS TO FORM COLOURLESS SOLUTION



Anion tests

Carbonate (CO_3^{2-})	Add dilute acid and test the gas released	Effervescence, gas produced is CO_2 which turns limewater milky
Chloride (Cl^-)	Acidify with dilute nitric acid and add aqueous silver nitrate	White precipitate formed
Bromide (Br^-)	Acidify with dilute nitric acid and add aqueous silver nitrate	Cream precipitate formed
Iodide (I^-)	Acidify with dilute nitric acid and add aqueous silver nitrate	Yellow precipitate formed
Nitrate (NO_3)	Add aqueous NaOH and aluminium foil, warm gently and test the gas released	Gas given off is ammonia, has a pungent smell and turns moist red litmus paper blue
Sulfate (SO_4^{2-})	Acidify with dilute nitric acid and add aqueous barium nitrate	White precipitate formed
Sulfite (SO_3^{2-})	Add dilute acid, warm gently and test the gas released	Gas decolourises purple acidified aqueous potassium manganate(VII) solution



Gas identification

Tests for Gases

Gases	Colour and Smell	Test	Test Result
Ammonia	Colourless, pungent	Hold damp red litmus paper in gas	Paper turns blue
Carbon dioxide	Colourless, odourless	Bubble gas through limewater	White ppt
Chlorine	Pale green, choking smell	Hold damp litmus paper in gas	Paper is bleached white
Hydrogen	Colourless, odourless	Hold a lighted splint in gas	Burns with a squeaky pop
Oxygen	Colourless, odourless	Hold a glowing splint in gas	The splint re-lights
Sulfur Dioxide	Colourless, choking smell	Bubble gas through acidified aqueous potassium manganate(VII)	Potassium manganate (VII) change from purple to colourless



CHEMICAL EQUILIBRIUM

Reversible reactions are reactions in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously.

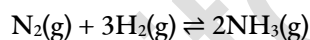
They consist of backwards and forwards reactions, whose respective rates are affected by a variety of factors, namely temperature, pressure and concentration. They are represented by the symbol $\rightleftharpoons$

A reversible reaction is at equilibrium when:

- (a) the rate of the forward reaction is equal to the rate of the reverse reaction, so products are being formed at the same rate as reactants
- (b) the concentrations of reactants and products are no longer changing, since they are both being formed at the same rate

Factors Affecting the Equilibrium:

A general principle that might make understanding the following easier is that the equilibrium always moves to minimise change. If the temperature rises, there is a tendency to automatically lower it, if the concentration of a certain substance drops, there is a tendency to raise it back up. Below is the reaction to form ammonia, which will be used as an example throughout and is included in the syllabus.



Temperature: Higher temperature makes the equilibrium move in the direction of the endothermic reaction, while lower temperature makes it shift in the direction of the exothermic reaction. For example, in the reaction to form ammonia, the forward reaction is exothermic, creating heat. Therefore, a lower temperature will shift the equilibrium forward, in the direction of the exothermic reaction. A higher temperature would shift it backwards, in the direction of the reverse endothermic reaction.

Concentration: Higher concentration of the reactants makes the equilibrium shift forward, while higher concentration of the products makes it shift backwards.

Pressure: This factor only applies if the reaction involves gases. In that case, increasing the pressure will cause the equilibrium to shift in the direction with less moles of gas and decreasing the pressure will cause it to shift in the direction with more moles of gas. For instance, in the ammonia reaction, there are four moles of gas on the



reactants' side and two moles on the products' side. Therefore, increasing the pressure will shift the equilibrium forward and decreasing will shift it backwards

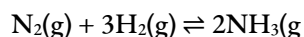
Catalyst: This has no effect whatsoever on the equilibrium's position. It will increase the rate of reaction, but it will not affect the position.

In addition to these, there are other factors to consider as well for the reactions involving copper (II) sulfate and cobalt (II) chloride in your syllabus. In their reactions, the hydrated substance reacts in the forward reaction to form its anhydrous form and water, with the opposite happening in the backwards reaction.

Adding water: This shifts the equilibrium backwards, to form more of the hydrated compound.

Heat: This shifts the equilibrium forward, as both hydrated substances decay in the presence of heat

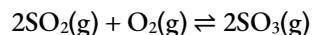
The Haber Process (The Formation of Ammonia)



Sources of Reactants: Hydrogen is produced from methane and nitrogen from distillation of liquid air.

Conditions: 450°C, 20000kPa / 200 atmospheres and an iron catalyst (remember that the catalyst has no effect on the position of the equilibrium.)

The Contact Process (Conversion of Sulfur Dioxide to Sulfur Trioxide):



Sources of Reactants: Sulfur dioxide is produced by burning sulfur or roasting sulfide ores. Oxygen is produced by the distillation of liquid air.

Conditions: 450°C, 200kPa / 2 atmospheres and a vanadium oxide catalyst (remember that the catalyst has no effect on the position of the equilibrium.)

Why are the conditions for both processes like this?

In both cases, the forward reaction is exothermic, so increasing the temperature leads to a decrease in yield by shifting the equilibrium backwards. However, decreasing the temperature would lead to a decrease in the rate, so these temperatures have been chosen to keep both factors in a desirable range.

In both cases, the products have less gaseous moles, so increasing the pressure increases the yield by shifting the equilibrium forward as well as the rate of reaction. However, high pressures can require a lot of effort and equipment to maintain, so they are not infinitely high.



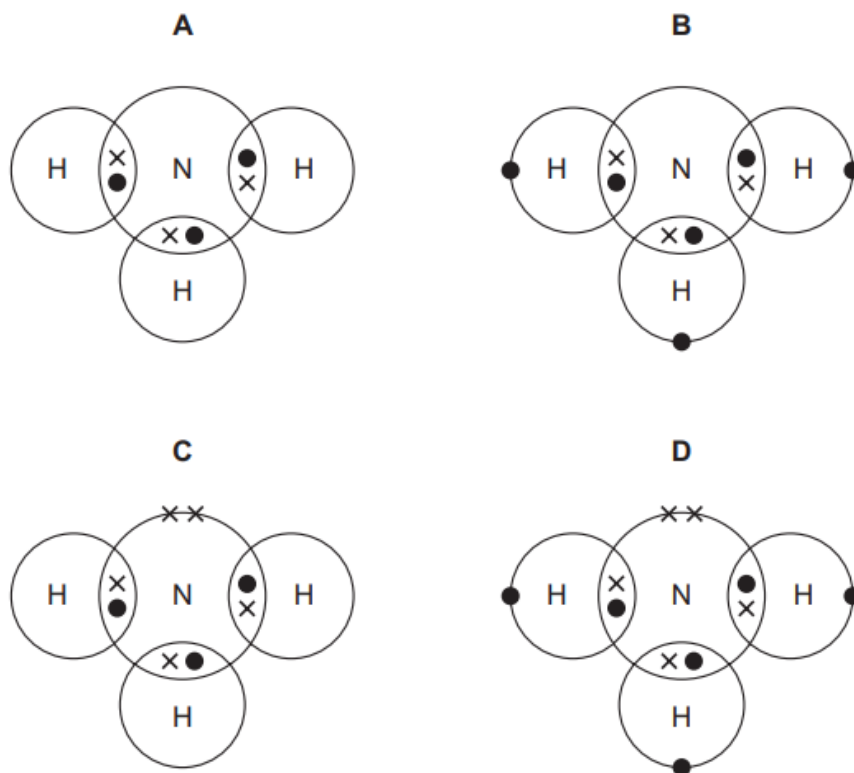
MCQS

Bonding:

Q1. 0620_s23_qp_22

Ammonia, NH_3 , is a covalent molecule.

Which diagram shows the outer-shell electron arrangement in a molecule of ammonia?



ANS: The correct answer is C, as in this example Nitrogen's octet is being completed according to the octet rule, and Hydrogen has the correct number of electrons in its outermost shell. A is not the correct answer, because Nitrogen does not have a full octet, and thus is not stable. B is incorrect because the Hydrogen atoms have three electrons in their outermost shell, alongside Nitrogen being unstable. D is incorrect because Hydrogen has three electrons in its outermost shell, whereas it should only have two in the case of a stable ion.

Q2. 0620_s23_qp_22



Which structure does silicon(IV) oxide most closely resemble?

- A** carbon dioxide
- B** diamond
- C** graphite
- D** sodium chloride

ANS: The correct answer is B, as Diamond is a giant covalent structure, like silicon(IV) oxide. Diamond also has a central atom bonded to four other atoms, which is a characteristic that silicon(IV) oxide shares. So it is the most similar to silicon(IV) oxide.

Q3. 0620_s23_qp_22

Substance P conducts electricity when solid.

Which particles move in solid P so that it can conduct electricity?

- 1 anions
- 2 cations
- 3 electrons

- A** 1 and 2
- B** 1 only
- C** 2 and 3
- D** 3 only

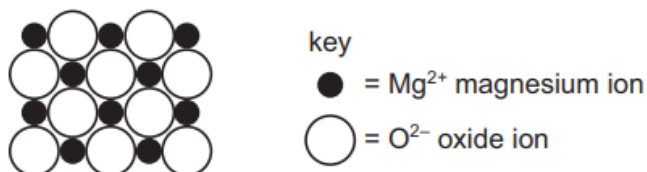
ANS: The correct answer is D, as in the solid state, individual ions (cations & anions) cannot move due to the lattice structure of solids. So, electrons are the only charge carriers able to freely move through the material.



Q4. 0620_s23_qp_23

Magnesium oxide is a white solid at room temperature and pressure.

Part of the structure of solid magnesium oxide is shown.



Three statements are listed.

- 1 Magnesium ions are smaller than oxide ions because they contain fewer electrons.
- 2 Magnesium oxide has good electrical conductivity when molten because the ions are mobile.
- 3 Magnesium oxide has a high melting point because of the strong electrostatic attraction between the ions and delocalised electrons in the giant lattice.

Which statements are correct?

- A** 1 and 2 **B** 1 and 3 **C** 2 and 3 **D** 2 only

ANS: The correct answer is D, because Magnesium oxide is an ionic compound that conducts electricity when molten, due to the free movement of ions which act as carriers of charge (option 2). Option 1 is not correct because their size depends on the charge on the nucleus. Option 3 is incorrect because the forces of attraction are between oppositely charged ions, not ions and delocalized electrons.

Q5. 0620_s23_qp_23

In which molecule are all the outer-shell electrons involved in covalent bonding?

- A** Cl_2 **B** CH_4 **C** HCl **D** NH_3

ANS: The correct answer is B, because In CH_4 , the central carbon atom containing four electrons, is bonded with four hydrogen atoms, so all four of the carbon's outer shell electrons are bonded to the four hydrogen atoms covalently.



Q6. 0620_s23_qp_23

Which row describes the properties of silicon(IV) oxide?

	giant covalent structure	melting point
A	no	high
B	no	low
C	yes	high
D	yes	low

ANS: The correct answer is C, because Silicon(IV) oxide is a compound that exists as a giant covalent structure, and due to the high number of covalent bonds keeping the structure together, also has a high melting point.

Q7. 0620_s23_qp_23

Which row describes the structure of a solid metal and explains the property?

	structure of solid metal	property of solid metal
A	lattice of negative ions in a sea of electrons	conducts electricity because the electrons are free to move
B	lattice of negative ions in a sea of electrons	is malleable because the layers of ions can slide over each other
C	lattice of positive ions in a sea of electrons	conducts electricity because the ions are free to move
D	lattice of positive ions in a sea of electrons	is malleable because the layers of ions can slide over each other

ANS: The correct answer is D, because Solid metals are made up of positive metal ions in a sea of delocalized electrons, and are malleable due to this property that allows them to slide over each other. A is not correct because the metal ions that make up the lattice are positive, not negative. B is incorrect due to the same reason. C is not correct because in a solid metal, only the delocalized electrons can move freely through the solid, while the metal ions cannot.



Q8. 0620_w23_qp_23

Magnesium reacts with oxygen to form magnesium oxide.

What happens to magnesium atoms and oxygen atoms during this reaction?

- A** Magnesium and oxygen share two electrons.
- B** Magnesium gains two electrons and oxygen loses two electrons.
- C** Magnesium loses one electron and oxygen gains one electron.
- D** Magnesium loses two electrons and oxygen gains two electrons.

ANS: The correct answer is D, because Magnesium and oxygen combine to make the compound MgO. In MgO, the ionic bonding means that it is made up of a positive metal ion (magnesium), and a negative non-metal ion (oxygen). In MgO, magnesium ions have a +2 charge, as they lose two electrons, while oxygen ions have a -2 charge, meaning they gain two electrons.

Q9. 0620_w23_qp_23

Which row about the properties of both diamond and silicon(IV) oxide is correct?

	conductor of electricity	type of molecule
A	yes	giant covalent
B	yes	simple covalent
C	no	giant covalent
D	no	simple covalent

ANS: The correct answer is C, because Both diamond and silicon(IV) oxide are giant covalent structures, and do not have any free outer shell electrons in their central atoms, making them unable to conduct electricity.



Q10. 0620_w23_qp_22

Potassium reacts with iodine to form potassium iodide.

Which statement about potassium iodide is correct?

- A** Each potassium atom shares a pair of electrons with an iodine atom.
- B** In potassium iodide, the particles of potassium have more protons than electrons.
- C** Potassium iodide has a high melting point because it is a covalent compound.
- D** Potassium iodide has a low melting point because it is an ionic compound.

ANS: The correct answer is B, because The potassium ions in KI have a +1 charge, meaning they lose an electron and thus have more protons than electrons. A is incorrect because electron sharing only occurs in covalent bonds, whereas potassium iodide has ionic bonding (between a metal and a non-metal). C is not correct because potassium iodide is an ionic compound, not a covalent one. D is not correct because all ionic compounds have relatively high melting points.

Q11. 0620_w23_qp_22

Which substance has the lowest melting point?

- A** graphite
- B** methanol
- C** silicon(IV) oxide
- D** sodium chloride

ANS: The answer is B, because Methanol is a simple covalent compound, which have the weakest bonds out of all the types. A and C are incorrect because they are giant covalent structures and have more bonds and thus are bonded more strongly together. D is incorrect because sodium chloride is an ionic compound that has a high melting and boiling point.

Q12. 0620_w22_qp_22

Which statement about solid magnesium oxide is correct?

- A** It is a giant structure made up of magnesium and oxygen atoms bonded covalently.
- B** It is an electrical conductor with mobile magnesium ions and oxygen ions.
- C** Magnesium loses electrons and these electrons move freely through a lattice.
- D** Oxygen ions and magnesium ions are attracted to each other in a giant lattice.



ANS: The answer is D, because Ionic compounds such as MgO always exist as ionic lattices where cations and anions are bonded together. A is incorrect because magnesium oxide has ionic bonding, not covalent. B is incorrect because in a solid, individual ions cannot move freely. C is incorrect because electrons cannot move freely through an ionic lattice, as they are donated by metal atoms to non-metal atoms to form cations and anions.

Q13. 0620_w22_qp_22

Which molecule contains only three shared pairs of electrons?

A CH_3OH

B Cl_2

C H_2O

D N_2

ANS: The answer is D, because Each Nitrogen atom needs three electrons to complete its octet. So, each of the two nitrogen atoms share three pairs of bonded electrons with each other.

Q14. 0620_w22_qp_22

Which particles are present in the structure of metals?

1 positive ions

2 negative ions

3 shared pairs of electrons

4 mobile electrons

A 1 and 2

B 1 and 4

C 2 and 3

D 2 and 4

ANS: The answer is B, because Metals consist of positive metal ions (option 1) in a sea of delocalized electrons (option 4). Option 2 is incorrect because only positive ions exist in a metal lattice. Option 3 is incorrect because electrons are shared in covalent compounds, not metal lattices.

Chemical equilibrium:

1. Anhydrous cobalt(II) chloride is blue and turns pink when water is added. How is this reaction reversed?

A Adding dilute acid

B Filtering

C Heating

D Cooling



The answer is C, because heating the pink hydrated cobalt (II) chloride will cause it to revert to its previous blue anhydrous form. A will not work because an acid cannot directly do so, B will not work because the water is bonded to the cobalt (II) chloride molecules, so it is impossible to physically separate them, and D will not work either for the same reasons.

2. Which statement about hydrated cobalt(II) chloride is correct?

- A It turns blue when it is heated.
- B It turns blue when water is added to it.
- C It turns pink when water is added to it.
- D It turns white when it is heated.

The correct answer is A, because heating hydrated cobalt(II) chloride would yield anhydrous cobalt(II) chloride, which is blue. Adding water to it would have no effect, as it is already hydrated, so B and C are incorrect. Heating it cannot turn it white, as that would be the case for copper(II) sulfate.

3 Which reaction is not a reversible reaction?

- A combustion of alkanes
- B hydration of anhydrous copper(II) sulfate
- C melting lead(II) bromide
- D thermal decomposition of hydrated cobalt(II) chloride

The correct answer is A, combustion of alkanes, because alkanes cannot be formed again by the combination of their products, carbon dioxide and water. B and D cannot be the answers because the hydration and thermal decomposition of both copper (II) sulfate and cobalt (II) chloride is a reversible reaction. C is also not the answer because lead (II) bromide can be cooled to receive once more its solid state, making this too a reversible reaction.

4 The equation shows a reaction that is reversed by changing the conditions



How can the forward reaction be reversed? .

	By adding water	By adding heat
A	✓	✓
B	✓	✗
C	✗	✓



D	X	X
---	---	---

The correct answer is B, because the forward reaction is that of dehydration, so to reverse it we would need to add more water but not heat it. A and C are incorrect because adding heat would cause it to remain as a forward reaction, while D is incorrect because we need to add water to shift the equilibrium.

5. Substance Y is a pink solid. When substance Y is heated gently it becomes a blue solid. When the blue solid is cooled down it remains blue. When water is added to the blue solid it becomes pink.

What is substance Y?

A anhydrous cobalt(II) chloride

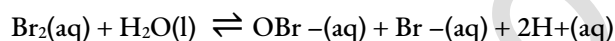
B anhydrous copper(II) sulfate

C hydrated cobalt(II) chloride

D hydrated copper(II) sulfate

The answer is C, because hydrated cobalt (II) chloride is pink in colour and turns to blue anhydrous cobalt (II) chloride when heated. It cannot be A or B, because the substance reacts to heat to produce its anhydrous form, meaning that it must be hydrated, and it cannot be D because it is pink, not blue.

6. Aqueous bromine is an equilibrium mixture.



Aqueous bromine is orange in colour. The species on the right-hand side of the equation are colourless. Changes are made to three separate portions of the equilibrium mixture. Which row shows how the colour of the mixture changes when a small amount of each substance is added?

	Adding sulfuric acid	Adding solid sodium bromide	Adding water
A	Darker orange	Darker orange	Darker orange
B	Darker orange	Darker orange	Paler orange
C	Darker orange	Paler orange	Darker orange
D	Paler orange	Darker orange	Paler orange

The answer is B, because an increase in the products, i.e sulfuric acid and sodium bromide, will cause the equilibrium to shift to the left and thus turn darker, while adding a reactant, i.e water, will cause the equilibrium to shift to the right and thus the solution will turn lighter. A and C are incorrect because adding water would not



cause the equilibrium to shift to the left and turn darker, while D is incorrect because adding sulfuric acid would not cause it to shift to the right and turn lighter.

7. Which of the following reactions cannot be reversed?

- A. Melting ice
- B. Thermal decomposition of hydrated cobalt(II) chloride
- C. Combustion of propane
- D. Hydration of anhydrous copper(II) sulfate

The answer is C, combustion of propane, since its products, carbon dioxide and water, will not react together to produce propane once more. A cannot be the answer because water can be frozen to once more provide ice, while B and D cannot be the answers because both these substances can be hydrated and reduced again to their anhydrous forms respectively.

8. Reversible reactions occur when products formed in the forward reaction behave as reactants and proceed in the reverse direction to produce the original reactants again.

Which of the following statements is correct for a reversible reaction?

- 1 If the forward reaction is exothermic, then the reverse reaction must be endothermic
- 2 If the forward reaction is endothermic, then the reverse reaction must be exothermic
- 3 Both directions are exothermic
- 4 Both directions are endothermic

-
- A. 1 and 2
 - B. 3 and 4
 - C. 1 only
 - D. 3 only

The answer is A, 1 and 2, because in a reversible reaction, one of the two reactions must each be exothermic and endothermic. B and D cannot be correct because both reactions cannot be of the same type, and C is incorrect because both the forward and backward reactions can be exothermic or endothermic.

9. Which row correctly describes the conditions used in the Haber process?

	temperature/ $^{\circ}\text{C}$	pressure/atm	catalyst
--	---------------------------------	--------------	----------



A	400	300	vanadium(V) oxide
B	450	200	iron
C	200	300	vanadium(V) oxide
D	300	200	iron

The answer is B, because a temperature of 450 degrees centigrade is used in the Haber process to create ammonia, as all the other answers are incorrect due to the incorrect temperatures given

10. Nitrogen and hydrogen are the raw ingredients used in the production of ammonia, NH_3 Which row correctly describes the source of these reactants and the catalyst used?

	nitrogen	hydrogen	catalyst
A	air	methane	Vanadium pentoxide
B	methane	water	iron
C	methane	water	Vanadium pentoxide
D	air	methane	iron

The answer is D, because nitrogen is extracted from air, hydrogen from methane and the catalyst used for the Haber process is iron. The answer cannot be A or C because the catalyst used is not vanadium pentoxide, and the answer cannot be B because nitrogen cannot be extracted from methane.

STRUCTURED QUESTIONS:

Bonding:

Q1. 0620_s23_qp_41



3 This question is about ionic and covalent compounds.

- (a) (i) Sodium reacts with oxygen to form the ionic compound sodium oxide.
The electronic configurations of an atom of sodium and an atom of oxygen are shown in Fig. 3.1.



Fig. 3.1

Ions are formed by the transfer of electrons from sodium atoms to oxygen atoms.

Complete the dot-and-cross diagrams in Fig. 3.2 to show the electronic configuration of **one** sodium ion and **one** oxide ion. Show the charges on the ions.

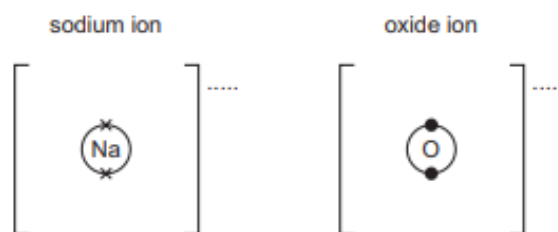


Fig. 3.2

[3]

- (ii) Write the formula of sodium oxide.

..... [1]

- (b) Carbon dioxide, CO_2 , is a covalent compound.

Complete the dot-and-cross diagram in Fig. 3.3 to show the electronic configuration in a molecule of carbon dioxide. Show outer shell electrons only.

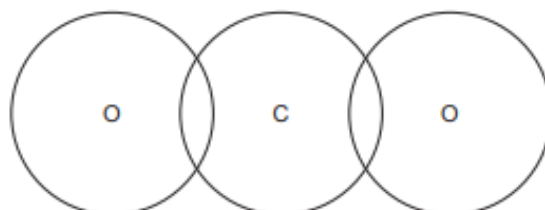


Fig. 3.3

[2]



(c) The melting points of sodium oxide and carbon dioxide are shown in Table 3.1.

Table 3.1

	melting point/ $^{\circ}\text{C}$
sodium oxide	1275
carbon dioxide	-78

(i) Explain, in terms of bonding, why sodium oxide has a high melting point.

.....

.....

.....

..... [2]

(ii) Carbon dioxide has a low melting point.

State the general term for the weak forces that cause carbon dioxide to have a low melting point.

..... [1]

[Total: 9]

Q2. 0620_w22_qp_42

Aitchison



- 1 Diamond and graphite are different solid forms of carbon. The carbon atoms in diamond and graphite are arranged in different ways.

(a) State the number of covalent bonds each carbon atom has in diamond.

..... [1]

(b) State the term used to describe the structure of diamond.

..... [1]

(c) Name an oxide that has a similar structure to diamond.

..... [1]

(d) Describe the arrangement of atoms in graphite.

.....
..... [2]

(e) Explain how graphite conducts electricity.

..... [1]

(f) Buckminsterfullerene is a simple molecular form of carbon.

The relative molecular mass of Buckminsterfullerene is 720.

Determine the number of carbon atoms in one molecule of Buckminsterfullerene.

..... [1]

(g) All forms of carbon burn to produce carbon dioxide.

Name the substance used to test for carbon dioxide.

..... [1]

[Total: 8]



Q3. 0620_w23_qp_43

5 This question is about iron.

(a) (i) Describe the bonding in a metallic element such as iron.

You may include a labelled diagram as part of your answer.

.....
.....
..... [3]

(ii) Explain why iron conducts electricity when it is solid.

..... [1]

Q4. 0620_w22_qp_43

Aitchi



3

- 2 The table shows the melting points, boiling points and electrical conductivities of six substances, **D**, **E**, **F**, **G**, **H** and **I**.

substance	melting point /°C	boiling point /°C	conducts electricity when solid	conducts electricity when liquid
D	1083	2567	yes	yes
E	-117	79	no	no
F	3550	4827	no	no
G	119	445	no	no
H	-210	-196	no	no
I	801	1413	no	yes

- (a) Identify the substance, **D**, **E**, **F**, **G**, **H** or **I**, which is:

- (i) a liquid at 25°C [1]
 (ii) a gas at 25°C [1]
 (iii) a solid consisting of simple molecules at 25°C. [1]

- (b) Identify the substance, **D**, **E**, **F**, **G**, **H** or **I**, which is a metal. Give a reason for your choice.

substance
 reason [2]

- (c) Identify the substance, **D**, **E**, **F**, **G**, **H** or **I**, which has a macromolecular structure. Give **two** reasons for your choice.

substance
 reason 1
 reason 2 [3]

- (d) Identify the substance, **D**, **E**, **F**, **G**, **H** or **I**, which is an ionic solid. Give a reason for your choice.

substance
 reason
 [2]

[Total: 10]



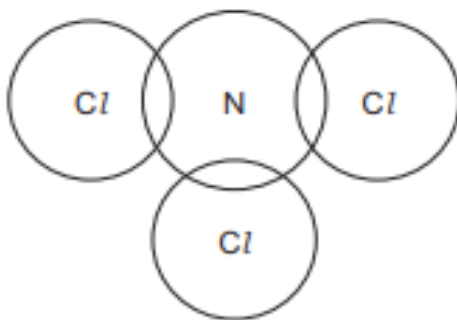
Q5. 0620_s22_qp_42

molecular formula =

1. Nitrogen trichloride, NCl_3 , is a covalent compound.

Complete the dot-and-cross diagram to show the electron arrangement in a molecule of

Show outer electrons only.



Chemical equilibrium:

1. The equation for the Haber process is shown. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ $\Delta H = -92 \text{ kJ/mol}$ The reaction is reversible. The forward reaction is exothermic.

(i) State the typical conditions and name the catalyst used in the Haber process. temperature
°C pressure, kPa, catalyst [3]



450, 200000, iron

(ii) What would be the effect of the following on the yield of ammonia? [2]

Increasing pressure (Increased yield)

Increasing temperature (Decreased yield)

Increasing the concentration of ammonia (Decreased yield)

Increasing the catalyst (No change)

2. Hydrated copper(II) sulfate is a coloured compound. It exists as hydrated crystals which contain water molecules.

(i) State the term given to water molecules present in hydrated crystals.

..... [1]

Water of crystallisation

(ii) State the colour of hydrated copper(II) sulfate crystals.

..... [1]

Blue

(iii) Write the formula of hydrated copper(II) sulfate.

..... [2]

$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

3. Sulphuric acid is made by the Contact process in the following sequence of reactions.

sulphur sulphur dioxide sulphur trioxide sulphuric acid

(a)

(i) How is sulphur dioxide made from sulphur? [1]

By burning sulphur in air

(b) The equation for a stage of the Contact process is $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$. The percentage of sulphur trioxide in the equilibrium mixture varies with:

(i) How does the percentage of sulphur trioxide in the equilibrium mixture vary as the temperature increases? Circle the correct answer.



Increases stays the same decreases [1] Decreases

(ii) Is the forward reaction in the equilibrium $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$ exothermic or endothermic? Give a reason for your choice. [2]

It is exothermic, because increasing the temperature shifts the equilibrium backwards, so the backward reaction is endothermic, therefore the forward one must be exothermic.

(iii) Explain, mentioning both rate and percentage yield, why the temperature used in the Contact process is 450°C . [2]

It is 450°C because that is the ideal temperature that is low enough for a good yield and high enough for a favourable economic rate.