Achievement Standard 91164

Demonstrate understanding of bonding, structure, properties and energy changes

CHEMISTRY

2.4

Externally assessed 5 credits

Chemical bonding and properties of molecules

This chapter assumes knowledge of atomic structure.

Chemical bonding properties

Most elements do not consist of isolated atoms but are made up of atoms held together by **chemical bonds**. A bond is an **electrostatic force of attraction** between positively and negatively charged species. There are three general ways that atoms are bonded – ionic bonding, covalent bonding and metallic bonding.

Ionic bonding

When electrons are *transferred* between atoms, electrically charged particles called **ions** are formed. The force of attraction between positive ions (cations) and negative ions (anions) is called an **ionic bond**. Monatomic cations (having only one atom) form from metal atoms, and monatomic anions are formed from non-metal atoms.

Covalent bonding

When electrons are *shared* between two atoms, a **covalent bond** is formed. The bonding electrons 'belong' to both atoms at the same time. Covalent bonds form between the atoms of non-metals.

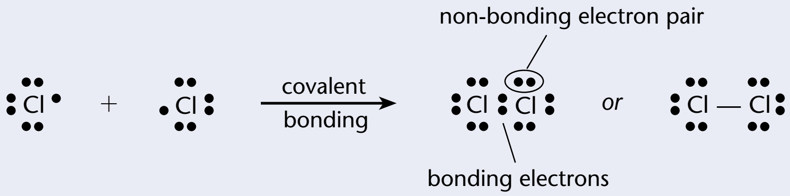
Metallic bonding

In metals, each atom is bonded to all neighbouring atoms by the electrostatic attraction between the loosely-held valence electrons and the positive metal ions (cations) that form when the valence electrons move away from the metal atom.

Lewis structures and shapes of molecules

Lewis structures (also called **electron dot diagrams**) can be used to represent the arrangement of electrons in species that contain covalent bonds.

A Lewis structure only shows the valence electrons, as these are the only ones involved in bonding. A Lewis structure must account for all the valence electrons of all the atoms in the molecule or ion.



The **shared** electrons are called **bonding** electrons. The electrons not involved in bonding are called **non-bonding** electrons or **lone pairs** of electrons. Shared electrons are usually represented by a line; e.g. chlorine is Cl–Cl

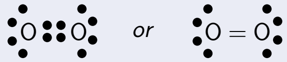
Example

A hydrogen atom and chlorine atom share electrons to form a hydrogen chloride molecule, HCl

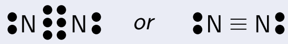
Atoms that share more than one pair of electrons, as double or triple bonds, form a **multiple covalent bond**.

Examples

1. The oxygen atoms in an oxygen molecule, O_2, are bonded together by a **double bond** (2 electron pairs).



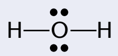
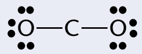
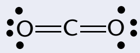
2. The nitrogen atoms in a nitrogen molecule, N_2, are bonded together by a **triple bond** (3 electron pairs).



Drawing Lewis structures

The following steps show how to draw Lewis structures.

Example

	Water, H_2O	Carbon dioxide, CO_2
Step 1 Count the total number of valence electrons present.	2 × H atoms (each with 1 valence e^-) $2e^-$ 1 × O atom $6e^-$ Total valence e^- = 8	2 × O atoms (each with 6 valence e^-) $12e^-$ 1 × C atom $4e^-$ Total valence e^- = 16
Step 2 Connect the atoms with single bonds. Each bond represents 2 electrons.	H—O—H	O—C—O
Step 3 Place the remaining electrons in pairs around the atoms, beginning with the outer atoms.		 Structure shows 16 electrons, but while the O atoms have 8 electrons, the C atom only has 4 electrons. To complete the 'octet' on C, move a pair of electrons from each O atom so that there are two double bonds.
Step 4 Check that the structure has the correct number of valence electrons and that the valence shells of all atoms are complete (octet rule). If not, move non-bonding pairs to form double or triple bonds.	The central O atom has 8 electrons – a full valence shell.	

Bond polarity

Non-polar covalent bonds

A non-polar covalent bond forms when a shared electron pair is attracted to the nuclei of both atoms *equally*. This kind of bond is formed between non-metal atoms that have the same **electronegativity**. Electronegativity describes how strongly the bonding electrons are attracted to the nucleus of a bonded atom. Elements on the right-hand side of the periodic table are the most electronegative.

Questions
Chemical bonding and molecular properties

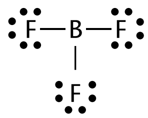
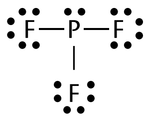
Question One

Year 2013
Ans. p. 95

a. Draw the Lewis structure for each of the following molecules.

Molecule	CH_4	H_2O	N_2
Lewis structure			

b. Boron and phosphorus both bond with three fluorine atoms to form BF_3 and PF_3. However, the molecules have different shapes and bond angles. The following table shows the Lewis structures for the molecules BF_3 and PF_3.

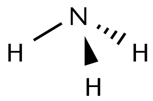
Molecule	BF_3	PF_3
Lewis structure		

Explain why these molecules have different shapes and bond angles. In your answer include:

- the shapes of BF_3 and PF_3
- factors that determine the shape of each molecule
- the approximate bond angle in BF_3 and PF_3
- justification of your chosen bond angles for each molecule.

2.4

c. i. The 3-dimensional diagram of NH_3 is shown.



Circle the word that describes the **polarity** of the molecule NH_3.

polar **non-polar**

Justify your choice.

ii. Elements M and X form a compound MX_2. Atoms of element X have a higher electronegativity value than atoms of element M, therefore the M–X bonds are polar. Depending on what elements M and X are, molecules of the compound formed will be **polar** or **non-polar**. State the most likely shape(s) of the molecule if it is:

Polar: **Non-polar:**

Justify your answer and draw diagrams of the possible molecules with dipoles labelled.

You do not need to identify what elements M and X are.

Achievement Standard 91165

Demonstrate understanding of the properties of selected organic compounds

CHEMISTRY

2.5

Externally assessed 4 credits

Characteristics and physical properties of hydrocarbons

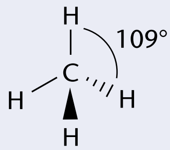
Hydrocarbons are organic compounds that contain only carbon and hydrogen. Alkanes, alkenes and alkynes are all hydrocarbons. They each form a **homologous** series (a 'family' where members of the series are represented by the same general formula, have the same functional group, have similar methods of preparation, and show similar chemical properties).

	Alkanes	Alkenes	Alkynes
General formula	C_nH_{2n+2}	C_nH_{2n}	C_nH_{2n-2}
Saturated?	yes	no	no
Functional group	–	double bond	triple bond

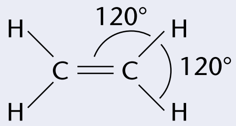
Saturated hydrocarbons contain single bonds. **Unsaturated** hydrocarbons contain a double or triple bond. A **functional group** is an atom or group of atoms within the organic molecule that determines the characteristic chemical properties of the particular family of organic compounds.

Examples of hydrocarbons

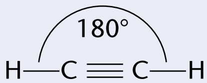
Methane is the simplest alkane and has the formula CH_4
A methane molecule has a tetrahedral shape.



Ethene is the simplest alkene and has the formula C_2H_4
An ethene molecule has two carbon atoms joined by a double bond. The four atoms around this double bond lie in a plane, with angles of 120° between the bonds.



Ethyne is an alkyne and has the formula C_2H_2
The ethyne molecule has two carbon atoms joined by a triple bond. It has a linear shape.



Physical properties of hydrocarbons

Physical property	Alkanes, alkenes and alkynes
Colour	colourless
Polarity	non-polar
Solubility in water	insoluble (ethyne is very slightly soluble)
Electrical conductivity	do not conduct
Conductivity of heat	poor
Density	less dense than water
Melting points and boiling points	• low due to weak forces between the molecules • gradually increase as molar mass increases

Naming alkanes

Straight-chain alkanes

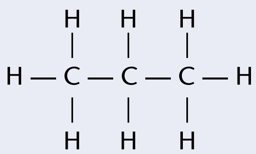
Ending (suffix) is *ane*.

Beginning of name (prefix) depends on the number of carbons.

Number of carbons	Prefix	Alkane name
1	*meth*	methane
2	*eth*	ethane
3	*prop*	propane
4	*but*	butane
5	*pent*	pentane
6	*hex*	hexane
7	*hept*	heptane
8	*oct*	octane

The **molecular formula** shows the identity and number of atoms in the molecule (e.g. molecular formula for propane is C_3H_8).

The **structural formula** (also known as **constitutional formula**) represents how the atoms in a molecule are arranged, e.g. structural formula for propane (shown alongside).



The 'stick' between the carbon and the hydrogen represents a shared pair of electrons (i.e. a single covalent bond).

For convenience, a **condensed structural formula** is mostly used, e.g. $CH_3CH_2CH_3$

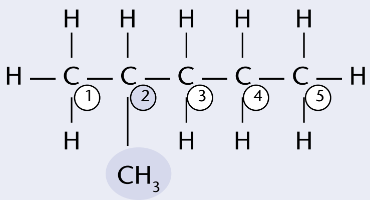
Branched-chain alkanes

Alkanes with more than 3 carbon atoms can also have a branched-chain structure.

Branched chains are called **alkyl** groups, e.g. methyl, CH_3– ; ethyl, CH_3CH_2–. Other alkyl groups are named accordingly, e.g. propyl, butyl, etc.

When naming branched-chain molecules, a number is used to indicate the position of the branch on the longest carbon chain in the molecule. The parent chain is numbered to give the lowest possible number for the branch.

Example



2-methylpentane

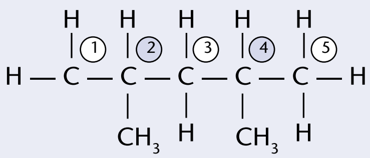
Longest chain of carbon atoms is 5, so pentane. The methyl group (methyl) is attached to carbon number 2 (2-methyl) of the 5-C chain (pentane).

For identical branches:

- *di* if there are 2 branches the same
- *tri* if there are 3 branches the same
- *tetra* if there are 4 branches the same.

A comma is placed between numbers.

Example



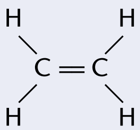
2,4-dimethylpentane

There are two (di) methyl groups attached at positions 2 and 4 (2,4-dimethyl) of the 5-C chain (pentane).

Naming alkenes

Ending (suffix) is *ene*.

The beginning of the name (prefix) depends on the number of carbons (as for alkanes); e.g.



is ethene

c. Molecule **D** can exist as geometric (*cis*- and *trans*-) isomers.

i. Draw the geometric (*cis*- and *trans*-) isomers for molecule **D** in the boxes below.

cis-isomer	*trans*-isomer

ii. Justify why molecule **D** can exist as geometric (*cis*- and *trans*-) isomers. Your answer should include:

- an explanation of the requirements for *cis*- and *trans*- isomers
- reference to the structure of molecule **D**.

d. Complete the following table to show the structural formula and IUPAC (systematic) name for each compound.

Structural formula	**IUPAC (systematic) name**
i.	pentanoic acid
ii.	3-methylbut-1-ene
$CH_3CH_2CH_2NH_2$	iii.
CH_3CHCH_2OH $\quad\;\;|$ $\quad\;\;Cl$	iv.
$CH_2CHCH_2CH_2CH_3$ $|\quad\;\;|$ CH_3CH_3	v.

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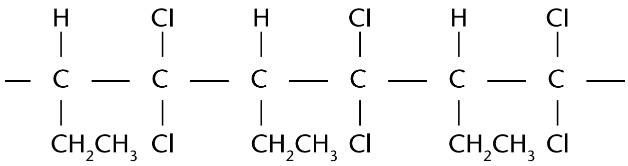
Question Two

a. i. The molecule tetrafluoroethene, shown alongside, is the monomer for the polymer commonly known as Teflon. $CF_2{=}CF_2$

Draw TWO repeating units for the Teflon polymer in the box.

ii. The diagram shows three repeating sections of another polymer.



Draw the structural formula of the monomer molecule used to make this polymer.

b. Five separate colourless organic liquids are known to be:

pentan-1-ol   ethanol   pent-1-ene   pentane   ethanamine.

Write a valid method to show how each of these liquids can be identified using **only** water, litmus paper, bromine water, $Br_2(aq)$, and acidified potassium dichromate, $H^+/Cr_2O_7^{2-}(aq)$. Your method should allow another student to identify these liquids, and should include:

- the reagent used
- any observations made.

You do not need to include equations in your answer.

Achievement Standard 91166

Demonstrate understanding of chemical reactivity

CHEMISTRY

2.6

Externally assessed 4 credits

Essential exam notes

Rates of reactions

Rate of reaction is 'the speed at which a reaction occurs'.

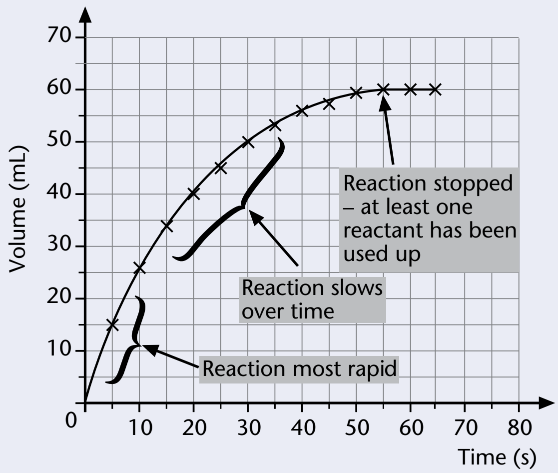
Some reactions occur quickly (e.g. adding zinc to hydrochloric acid), others occur slowly (e.g. iron rusting in a car).

Measuring rates of reactions

The rate at which a reaction occurs is measured by how quickly reactants are consumed or how quickly products form (e.g. a gas product can be measured).

The graph shows how the volume of hydrogen gas given off in the reaction between zinc and hydrochloric acid changes over time.

Volume of hydrogen gas formed against time



Collision theory

For a reaction to occur, particles must collide with sufficient energy for bonds to break and new bonds to form to make new substances.

Factors that control rates of reactions are:

- the frequency of collisions – the more collisions in a given time, the faster the reaction
- the effectiveness of the collisions – this is in terms of *orientation* (reactant particles must be positioned so that bonds break and bonds form) and *sufficient energy* (particles must collide with sufficient energy to break existing bonds and make new bonds).

Particles must collide with enough **kinetic** ('moving') energy. Total kinetic energy must be greater than the **activation energy**. Activation energy, E_A, is the minimum amount of energy required for a reaction to proceed between colliding particles.

Factors affecting rates of reactions

The rate of a reaction depends on the concentration, surface area, temperature of reactants, and the presence of catalysts.

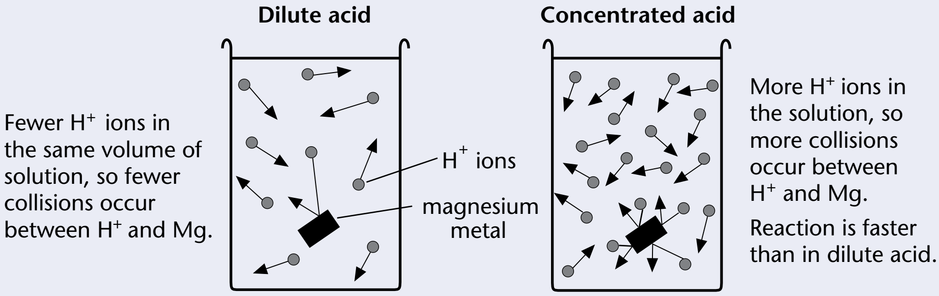
Concentration

Concentration affects reactions in which at least one of the reactants is in solution or the reactants are gases.

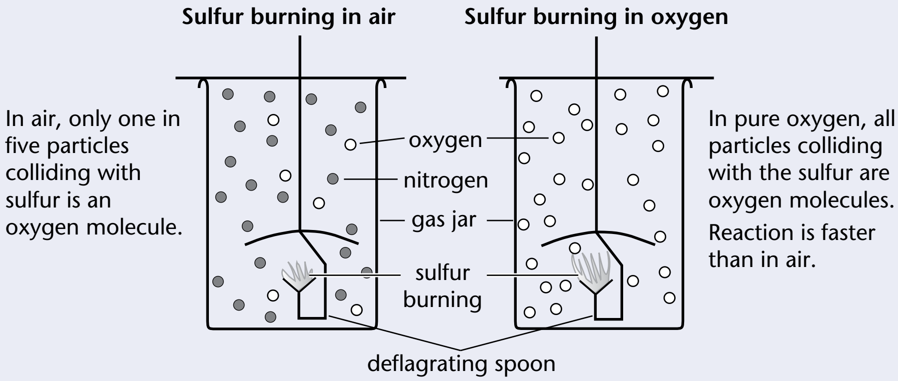
With increased concentration, particles are closer together so there is an increased chance of colliding. As a result, the *frequency of collisions increases* and hence the rate of reaction increases.

The rate of a reaction will decrease as the reaction continues because the concentration of the reactants is decreasing.

For solutions, using a more concentrated solution will increase the rate of reaction.

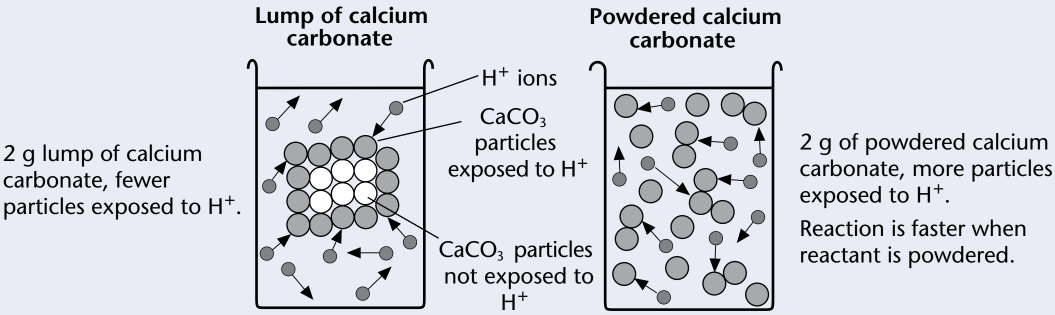


For gases, increasing the pressure or introducing more of the reactant gas into the container will increase the concentration of gases.



Surface area

Increasing the surface area means exposing a greater part of one reactant to the other reactant, e.g. using a powder instead of a solid. Because there is an increase in the surface area exposed, the frequency of collisions increases and hence the rate of reaction increases. Stirring and shaking can expose the surface of reactants. This *increases the frequency of collisions*.

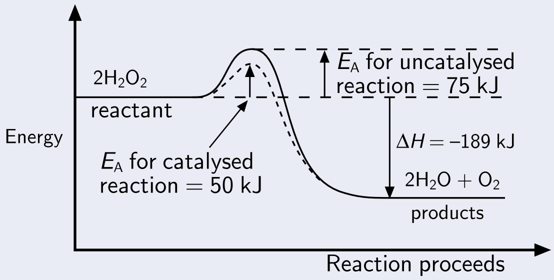


Temperature

Temperature is a measure of the average kinetic energy of particles. At higher temperatures, the *collisions are more energetic* (*effective*) and since more reacting particles will have greater energy than the activation energy, the reaction is faster. If temperature is increased, particles also move faster and there is more chance of collisions. As a result, the *frequency of collisions increases* and hence the rate of reaction increases.

Catalyst

A catalyst can speed up a reaction. A catalyst increases the rate of reaction by lowering the activation energy for the reaction. This means that *collisions are more effective as more particles have sufficient energy to react*. A catalyst is not used up by a reaction and does not change the products formed. It provides an alternative pathway for the reaction. Biological catalysts are proteins called **enzymes** (e.g. amylase in the mouth is involved in digestion, breaking starch down into glucose).



NCEA practice Questions
Rates of reactions

Question One

Year 2013
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a. Hydrochloric acid was reacted with calcium carbonate in the form of marble chips (lumps) and powder (crushed marble chips) in an experiment to investigate factors affecting the rate of a chemical reaction.

Marble chips (lumps)

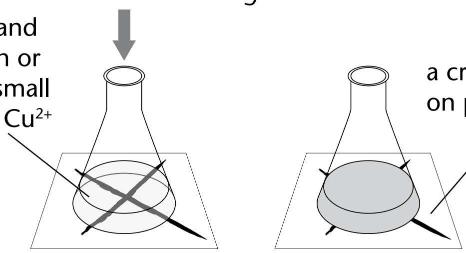
Crushed marble chips (powder)

i. Identify the factor being investigated.

ii. Explain why the hydrochloric acid would react faster with the powder.

b. A *clock reaction* involves mixing solution X and solution Y with starch present. When the reaction is complete, the solution turns blue-black in colour.

A student carried out this reaction between solution X and solution Y in a conical flask. Over time, the cross on the piece of paper under the flask disappeared when viewed from above.



Three experiments were carried out, and the times taken for the cross to disappear recorded.

Experiment		Temperature / °C	Time for cross to disappear / s
1	No Cu^{2+} present	25	42
2	No Cu^{2+} present	50	23
3	Cu^{2+} present	25	5

Elaborate on why the reactions in **Experiment 2** and **Experiment 3** occur faster than the reaction in **Experiment 1**. In your answer, include the following words or terms.

collisions **activation energy** **temperature** **effective** **catalyst**

2.6

Reason: There are three regions of electrons on the central (C) atom. To minimise repulsion, these are in a trigonal planar arrangement.

c. i. and ii. HCl – **polar bond** Explanation: The bonded atoms have different electronegativities, so the bonded electrons are more attracted to the Cl atom (the more electronegative atom), creating a charge separation (a bond dipole).

N_2 – **non-polar bond** Explanation: The bonded atoms are the same, so the bonded electrons are evenly shared and the bond is non-polar.

d. i. H_2S is a polar molecule

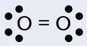
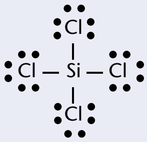
CO_2 is a non-polar molecule

ii. Both molecules have polar bonds since, for each bond, the bonding atoms have different electronegativities. H_2S is a V-shaped molecule with the polar bonds arranged asymmetrically around the central atom, so the polarities of the bonds will not cancel and the molecule is polar. CO_2 is a linear molecule so the polarities of the bonds cancel and the molecule is non-polar.

(*A* – some of correct Lewis structures or shapes or polarity of bonds or molecules; *M* – links shape to number of electron clouds or links bond polarity to electronegativity and polarity of molecules to their shape; *E* – full explanations for shape and polarity)

Question Four

p. 11

a. i.  ii.  iii. 

b. i. tetrahedral

ii. trigonal pyramid

iii. trigonal planar

c. NOCl: bond angle is approximately 120°; three electron clouds/regions of electron density/regions of negative charge around the central Atom. To minimise the repulsion between the electron clouds, they are arranged as far apart as possible. This means the electron clouds form a trigonal planar arrangement, which gives a bond angle of (approximately) 120°. There are two bonding clouds and one non-bonding cloud, so the observed structure will be bent.

H_2S: bond angle approximately 109°; four electron clouds/electron pairs/regions of negative charge around the central atom. To minimise the repulsion between the electron clouds, they form a tetrahedral arrangement, which gives a bond angle of (approximately) 109°. There are two bonding clouds and two non-bonding clouds, so the observed structure will be bent.

d. i. polar

ii. non-polar

iii. polar

e. For a molecule to be polar, it needs to have polar bonds that are arranged asymmetrically around the central atom so that the polarity of the bonds is not cancelled. This causes an uneven spread of charge across the molecule.

CCl_4 is a non-polar molecule. It has polar bonds because of the electronegativity difference between the C and Cl atoms. However, the symmetrical tetrahedral arrangement of the C–Cl bonds cancels out the polarity of the individual bonds, making the molecule non-polar.

CH_3Cl is a polar molecule because the polarity of the C–H bonds is less than that of the C–Cl bond. This means that there is a higher electron density on the Cl atom than on the H atoms, giving one end of the CH_3Cl molecule a small negative charge and the other end a small positive charge, i.e. charges are not spread evenly around the molecule.

NH_3 is a polar molecule because the polar N–H bonds are not arranged symmetrically around the N atom; also, there is a lone pair of electrons on the N atom.

(**A** – some of Lewis structures, shapes of molecules, polarity of bonds or molecules correct; **M** – links shape to number of electron clouds or links bond polarity to electronegativity and polarity of molecules to their shape; **E** – full explanations for shape and polarity)

2.4 Types of solid

Question One

p. 17

a.

Substance	Type of substance	Type of particle	Attractive forces between particles
C(*s*) (graphite)	(Covalent) Network	Atoms	Covalent bonds
Cl_2(*s*) (chlorine)	Molecular	Molecules	Weak intermolecular forces
$CuCl_2$(*s*) (copper chloride)	Ionic	Ions	Ionic bonds
Cu(*s*) (copper)	Metallic	Cations and electrons	Metallic bonds

b. i. Cl_2(*s*) consists of molecules with weak forces of attraction between them. These are easily broken at low temperatures (little energy needed), so at room temperature there is sufficient energy to separate the molecules from one another so Cl_2 exists as a gas. $CuCl_2$ consists of oppositely charged ions (Cu^{2+} and Cl^-) that are strongly attracted to one another forming an ionic bond. A large amount of energy is needed to break this bond, and hence at room temperature the bond is unbroken and the compound is in the solid state.

ii. In graphite, each carbon atom is covalently bonded to 3 other carbon atoms, forming a (hexagonal) planar array of atoms. The 4th valence electron is delocalised between the layers. It is free to move, and hence graphite is able to conduct electricity. The layers of atoms are weakly bonded to each other through the delocalised (4th) valence electron, and hence when force is applied the layers will separate from each other, and hence graphite cannot be drawn into a wire.

In copper metal, the valence electrons are delocalised and bond to all the atoms that are packed into a regular 3-D array. The delocalisation of the valence electrons means that they are free to move, and hence the metal is able to conduct electricity. When pressure is applied to this array, the atoms are able to slide over each other without disrupting the bonding, so it is possible to draw out copper into a wire and still maintain the strong solid array.

(**A** – majority of: identifies type of solid / particles / forces between particles, recognises relationship between strength of bond and melting point / state at room temperature / energy for state change, recognises factors affecting conductance/ductility; **M** – links state of substance/conductance/ductility to structure and bonding; **E** – full complete explanations that contrast the properties of solids linked to structure and bonding)

Question Two

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a.

Substance	Type of particle	Attractive forces between particles
Ammonia, NH_3	Molecule	Intermolecular forces
Zinc, Zn	Atom / cations and electrons	Metallic bonds
Silicon dioxide, SiO_2	Atoms	Covalent bonds