Unit 2
ATOMIC STRUCTURE & BONDING **10 minutes**
Bonding

Test Yourself

1 An atom with an electron removed from its valence shell this is called an ______.
The energy needed to ionise a mole of gas atoms of an element is called its _______ energy.
For metals this energy is very (*small/large*), whereas for the ______ gases the value is extremely (*small/large*).
Going from top to bottom of a Group, the ionisation energy becomes (*larger/smaller*), Whereas going across a Period it becomes (*larger/smaller*).
To remove a second valence electron from an alkali metal will take (*more/less*) energy than the first.

2. The bonds between a metal and a non-metal are called _________ bonds. With these, the metal (*receives/donates*) valence electrons and the non-metal (*receives/donates*) them so that both species can achieve a ______ gas configuration. This is very (*stable/unstable*) e.g. if potassium gains an electron its structure becomes isoelectronic with the element called _________.

3. The ions in an ionic solid are arranged in a 3-D ________ structure which is strong and gives it a very (*high/low*) melting point. Ionic solids are (*good/bad*) electrical conductors unless they are changed into the _________ state by heating, where the ______ can then move around to transfer charge. When stressed, ionic solids snap with (*ease/ difficulty*) as they are brittle.

4. List: salt, graphite, diamond, silicon carbide, calcite, silicon dioxide, calcium carbide.
The 3 covalent molecular solids in the list above are: ___________, ____________ and _______________________. These compounds are held together in ________ dimensions by strong ___________ bonds. This makes them very (*soft/rigid*) and (*good/bad*) conductors, with (*high/low*) melting points.
The reason that graphite is slippery and conductive is because of the ___________ electrons from the carbon atoms, not used in covalent bonding. These form (*strong/weak*) bonds between the molecular sheets of graphite molecules.

5. Polyatomic ions contain several _________ held together by _________ bonds, with a net charge.
List: Permanganate, oxalate, ammonium, phosphate, cuprammonium ion.
From the list select one ion carrying the following net charge:

1- __________________ 2- ___________________ 3- _____________________

1+ __________________ 2+ __________________

The colours of the following ions are:

Fe^{2+} ________________ Co^{2+} _______________ Ag^{+} _____________________

Cu^{2+} ________________ Mn^{2+} _______________ Cr^{3+} ____________________

6. The attraction of an atom for an electron is called its _________-negativity. The value of this will (*increase/decrease*) along the Period and (*increase/decrease*) down the Group.
   Polar bonds are ones where a shared electron is (*closer to/further from*) the non-metal atom in the bond. This causes the metal atom to have a slight (*positive/negative*) charge and the non-metal to have a slight (*positive/negative*) charge on it. A net dipole is produced when the vector dipole from one polar bond in a molecule (*is/is not*) cancelled out by the other vectors. This kind of molecule is said to be a ________ molecule.

7. Where molecules of a compound are polar, this produces _______ forces between adjacent __________ and gives rise to relatively (*low/high*) melting and boiling points. The _________ bond is an extreme case of this and can only occur in molecules where the element ________ is bonded to either N, O or F atoms. Molecules with this kind of intermolecular bonds have unusually (*high/low*) melting and boiling points.

8. The least significant type of intermolecular forces are called _________ forces, produced because of the + and – charges in an atom or molecule are (*close to/far from)* each other. These ________ dipoles cause very (*large/small*) intermolecular forces to be exerted, pulling molecules together into a liquid at temperatures (*above/below*) the boiling point.

9. All metals are held together with forces called _______ bonds, formed due to ________ metal valence electrons. The attractive force between the electron cloud and the positive ________ causes metals to have (*low/high*) strength and malleability. The conductive properties of metals are due to the delocalised nature of their __________.

10. A polar solute will generally dissolve in _______ solvents due to the _______ between the + and – ends of its molecules and the solvent dipoles. Non-polar solvents will have (*weak/strong*) electrostatic forces between its molecules which results in them being able to ________ non-polar solutes.

Answers:

1. Ion; ionization; small; noble; large; smaller; larger; more.
2. Ionic; donates; receives; noble; stable; argon.
3. Lattice; high; bad; molten; ions; ease.
4. Diamond; silicon carbide; silicon dioxide; 3; covalent; rigid; bad; high; delocalised; weak.
5. Elements; covalent; 1- permanganate; 2- oxalate; 3- phosphate; 1+ ammonium; 2+ cuprammonium; Fe^{2+} green; Co^{2+} pink; Ag^{+} colourless; Cu^{2+} blue; Mn^{2+} pale pink; Cr^{3+} green.
6. Electro; increase; decrease; closer to; positive; negative; is not; polar.
7. Electrostatic; molecules; high; hydrogen; hydrogen; high.
8. Dispersion; far from; temporary; small; below.
9. Metallic; delocalised; ions; high; electrons.
10. Polar; attraction; weak; dissolve.

Answers:

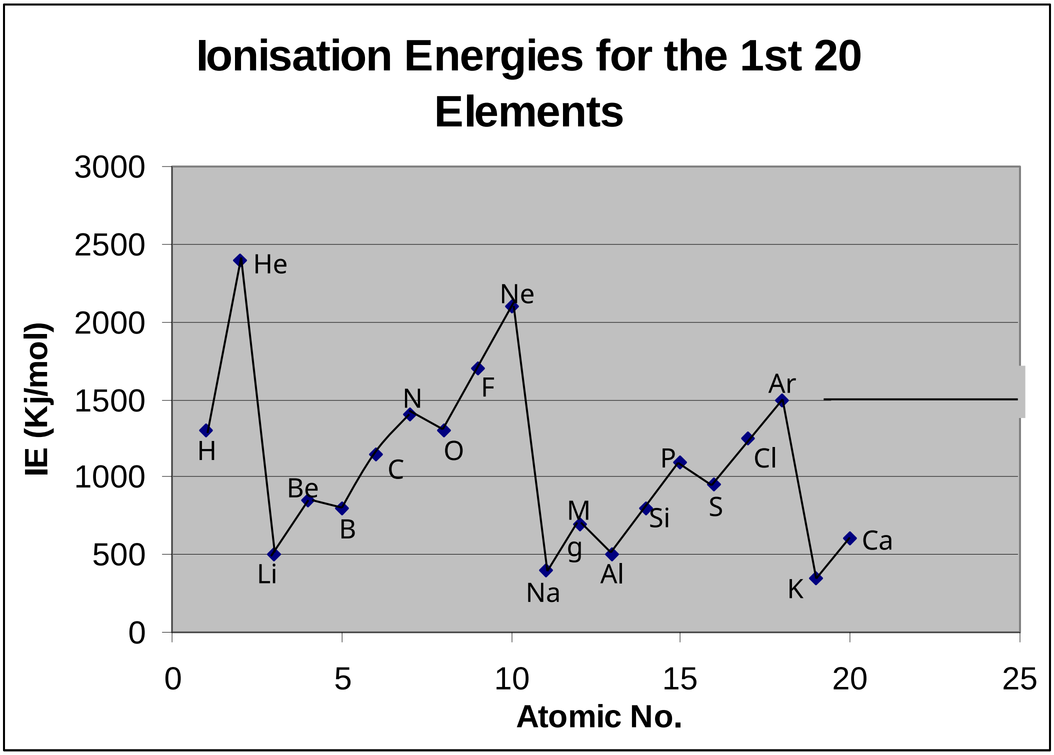
1. Spectra; model; nucleus; electrons; orbitals; atomic number; z; protons; neutrons; nucleus.
2. Electrons, other, groups; Periodic; levels.
3. Electrons; full; particular; pair; 5; 7.
4. Electrons; type; fluorine.
5. Element; shells; energy; group; noble gases.
6. 3 = Li; 8 = O; 12 = Mg; 18 = Ar; 25 = Mn; 30 = Zn.
7. Valence; reactions; electron; full; energy; left; lose; non; gain.
8. Ionic; loses; -; attracted; bond.
9. Covalent; shared; noble; donating; bond; co-ordinate.

Unit 2
ATOMIC STRUCTURE & BONDING **20 minutes**
Bonding

Key Points

1 The **ionisation energy** (IE) for an element is a measure of the energy needed to **extract** one electron from its outer orbital of its vapour M (g) + IE $\rightarrow$ M^+ (g) + e^-
IE is measured in kilojoules per mole of atoms ($kJmol^{-1}$).

- IE is **low** for atoms with only **one valence electron** (alkali metals) rising to very **high for the noble gases** where the attraction is great.



- Going **down a group** (e.g. Li $\rightarrow$ Na) the IE decreases because there are more electron shells beneath to repel the outer valence electron (e.g. Li has $2e^-$ repelling, Na has $8e^-$ repelling). Hence the trend is for **less energy** to be required to extract electrons from **heavier** elements.
- Going **across a period** the IE increases as there is a greater attractive force from a greater number of protons in the nucleus.
- As **successive electrons** are removed from the same atom the IE rises because the attractive force becomes greater because there are less electrons now available for repulsion.

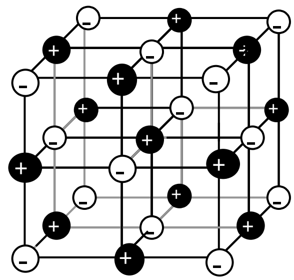
(1st) Na $\rightarrow$ Na^+ (IE = 502 $kJmol^{-1}$); (2nd) Na^+ $\rightarrow$ Na^{2+} (IE = 4569 $kJmol^{-1}$)

(3rd) $Na^{2+} \rightarrow Na^{3+}$ (IE = 6919kJmol^{-1}); (4th) $Na^{3+} \rightarrow Na^{4+}$ (IE = 9550 kJmol^{-1})

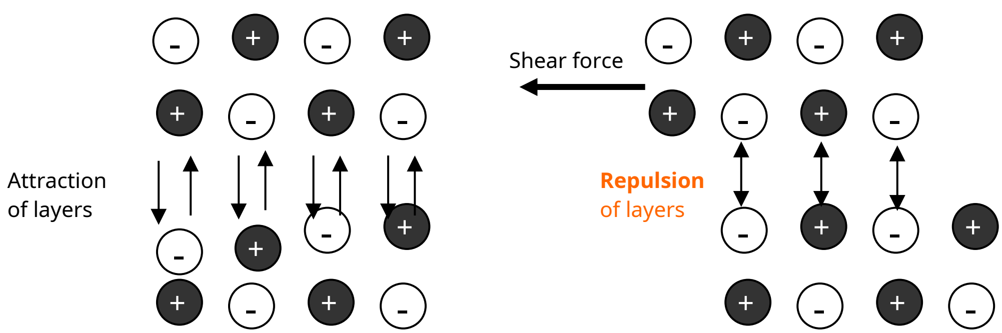
2 Where ionic bonding takes place (metal to non-metal) electrons will always move from an element with a low IE to another with higher IE (e.g. Na + F $\rightarrow$ Na^+F^-). Here IE_{Na} = 502 kJmol^{-1}, IE_F = 1700 kJmol^{-1} hence about 1200 kJmol^{-1} of energy can be released by the reaction.

- The electron configuration of an ion will be **isoelectronic** with the noble gases as ionisation occurs to achieve the lowest energy state (isoelectronic means the same electron arrangement).
  e.g. B^{3+} ion: B atom is $1s^22s^22p^1$ but B^{3+} is $\mathbf{1s^2}$ (isoelectronic with He)
  e.g. O^{2-} ion: O atom is $1s^22s^22p^4$ but O^{2-} is $\mathbf{1s^22s^22p^6}$ (isoelectronic with Ne)

1 In ionic solids the + and – ions exist in a **3-D lattice**, held together by electrostatic attraction:

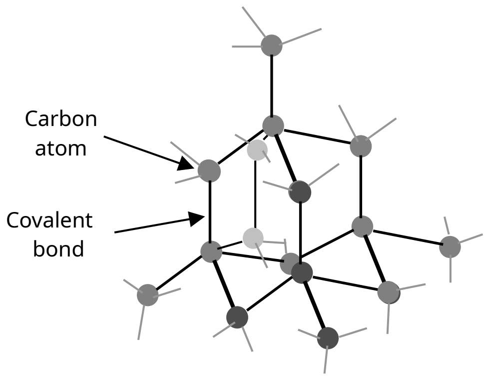


- Due to the strong attractive forces, substances with ionic lattices have a **high melting point** and boiling point because a large amount of heat energy is required to separate the ions.
- Ionic lattices are **brittle** (easily shattered) due to **repulsion of adjacent layers** when they slip under shear forces



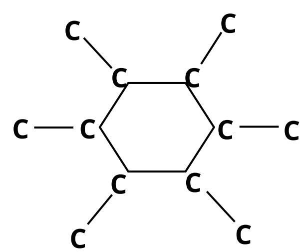
- In molten (fused) ionic solids **mobile ions** are freely available for conduction as they break away from the lattice, making them **good conductors** when molten.

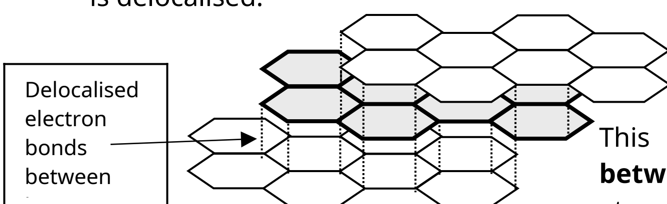
4. **Covalent network solids** are really one giant 3-d network of atoms bonded together in all directions. This makes them extremely **rigid and hard**. Covalent network solids thus have **high melting points** (mp) and boiling points (bp) and are very **poor electrical conductors** because they have no available free electrons.
Some examples of covalent network substances are **diamond**, silicon dioxide and silicon carbide.



Diamond Network structure

- An allotrope of carbon called **graphite** is not 3-D bonded as only 3 of carbon's 4 available electrons are used to form **three** C-C bonds whilst the other electron is delocalised.



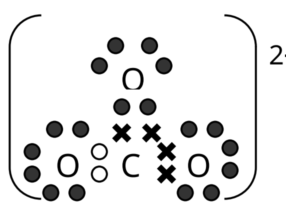


This electron forms very weak bonds **between** the layers of bonded carbon atoms. This means that each layer can be easily moved sideways by shear forces

This gives graphite its lubricating properties (slippery) that makes it useful in graphite grease. The freely-moving electrons between layers make graphite quite a good conductor, even though it is a non-metal.

5. Polyatomic ions exhibit covalent bonding
e.g. carbonate ion:

○ = Extra ionic electrons



- Polyatomic ions that you need to be familiar with are:

1-	1-	2-	3-	1+	2+
OH^-	ClO_4^-	CO_3^{2-}	PO_4^{3+}	NH_4^+	Hg_2^{2+}
NO_3^-	ClO_3^-	$Cr_2O_7^{2-}$			$[Cu(NH_3)_4]^{2+}$
HCO_3^-	ClO_2^-	SO_4^{2-}			$[Zn(NH_3)_4]^{2+}$
HSO_4^-	ClO^-	SO_3^{2-}			
CH_3COO^-	$[Al(OH)_4]^-$	O_2^{2-}			
MnO_4^-	$[Au(CN)_2]^-$	CrO_4^{2-}			
CN^-	$[Cr(OH)_4]^-$	HPO_4^{2-}			
$H_2PO_4^-$		$C_2O_4^{2-}$			
		$[Zn(OH)_4]^{2-}$			

- **Colours** of monatomic ions in solution to be familiar with are:

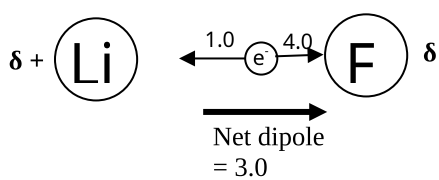
Cr^{3+}	Mn^{2+}	Fe^{2+}	Fe^{3+}	Co^{2+}	Ni^{2+}
GREEN	PALE PINK	GREEN	BROWN	PINK	GREEN
Cu^{2+}	Cd^{2+}	Zn^{2+}	Ag^+	Hg^{2+}	Hg_2^{2+}
BLUE	C/LESS	C/LESS	C/LESS	C/LESS	C/LESS

6. The **electronegativity** of an element is a measure of an atom's relative **attraction** for an electron. **Metals will lose electrons** to achieve the noble gas structure and so they have a **low** value of electronegativity. This value will **increase across the period** moving left to right e.g. Values: Li = 1.0, F = 4.0

Down a group the electronegativity value will **decrease**, as the attractive force of the nucleus decreases when the atomic radius gets larger.

Polar bonds exist where the shared electrons of a bond sit **closer to one** atom than another to give it a slight negative charge (δ -).

- E.g. in LiF the attractive force of the fluorine atom is 4x larger than that of the lithium atom and so the electron sits closer to the F atom, giving it a slight negative charge (δ -).

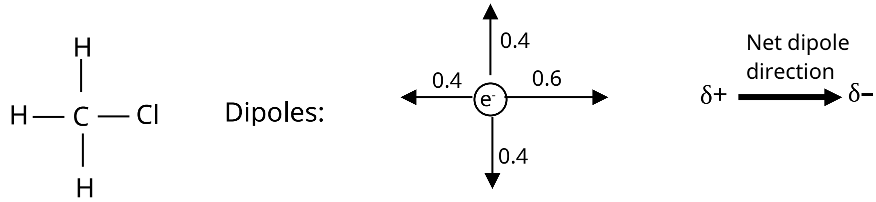


In this extreme case of sharing, the bond is **highly polar** and called an **Ionic Bond**.

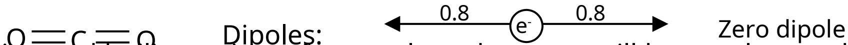
An O-F bond would be far less polar O $\xleftarrow{3.4}$ e^- $\xrightarrow{4.0}$ F but the electron still sits closer to the F atom, making it a polar bond (net dipole 0.6) with a δ – on the F atom. The only pure non-polar bonds exist between molecular elements e.g. O=O, Cl-Cl, H-H, etc.

- Bond dipoles can add or subtract (like vectors) to give an overall **dipole moment** and a **polar molecule**, or subtract to give a **zero dipole moment** and a **non-polar molecule.** The melting and boiling points of polar substances are relatively high because of the large intermolecular forces.

e.g. 1 **CH_3Cl** is a polar molecule with polar bonds

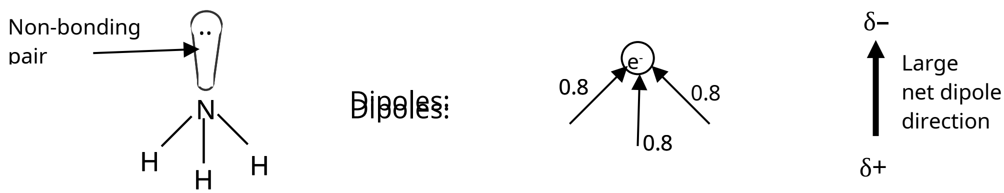


e.g. 2 **CO_2** is a non-polar molecule with polar bonds

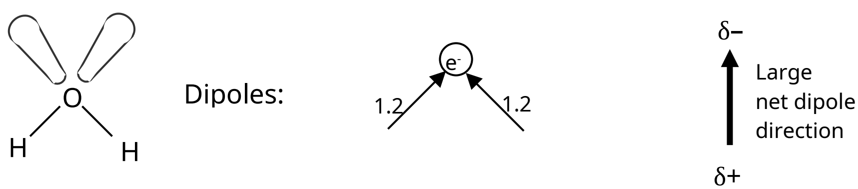


The melting and boiling points of non-polar substances will be very low as there are no dipole-dipole forces acting. Only dispersion forces hold the CO_2 molecules together.

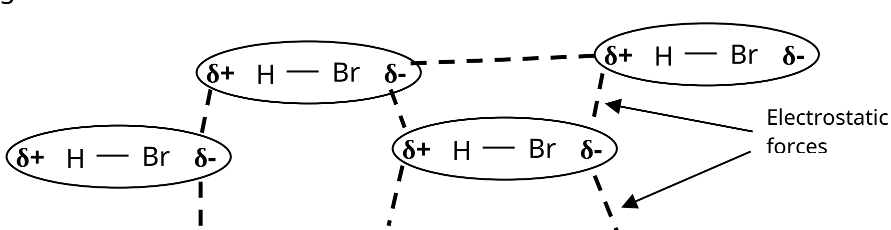
e.g. 3 **NH_3** has 3 equal polar bonds but the molecule does have a net dipole moment due to its shape. This shape is due to the electron repulsion by the unbonded pair. The non-bonding pair repels the H-bonds down, according to the Valence Shell Electron Pair Repulsion (VSEPR) theory.

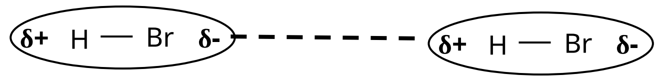


e.g. 4 **H_2O** is also a polar molecule due to the electron pair repulsion (VSEPR), forming a non-symmetrical shape.



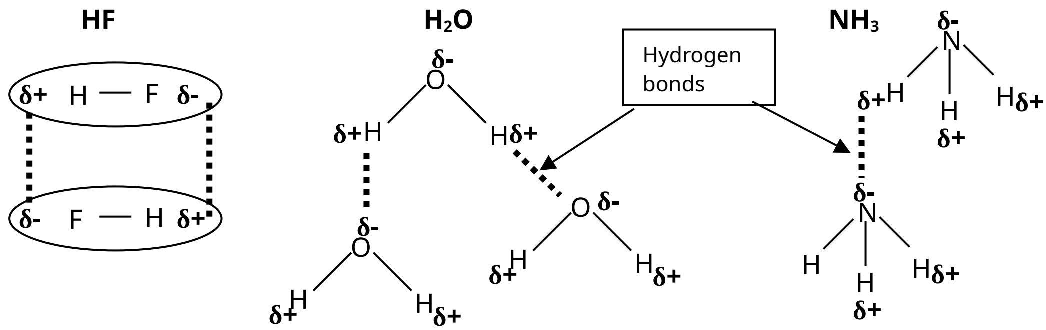
7. **Intermolecular attractive forces** act between molecules because of the dipoles present on the bonds. Strong intermolecular forces give rise to high melting points and boiling points. E.g.





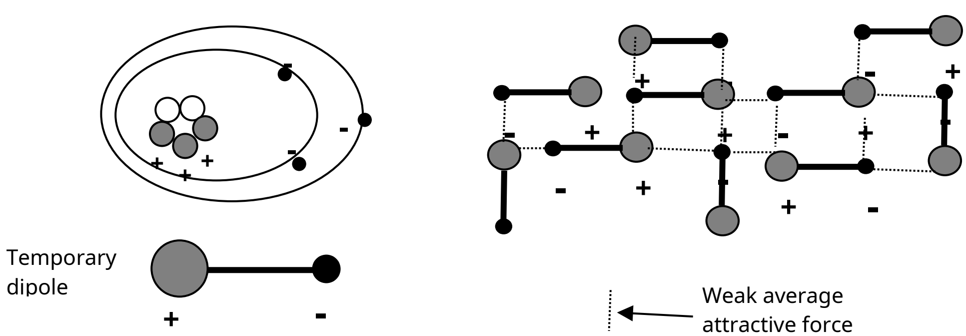
- **Hydrogen Bonding** is an extreme case of dipole-dipole attraction with molecules where hydrogen is attached to either a **N, O or F atom**. H-O, H-N and H-F bonds are extremely polar and yield abnormally high forces between their molecules. Attraction between dipoles containing these pairs of atoms is called Hydrogen Bonding.

  e.g.



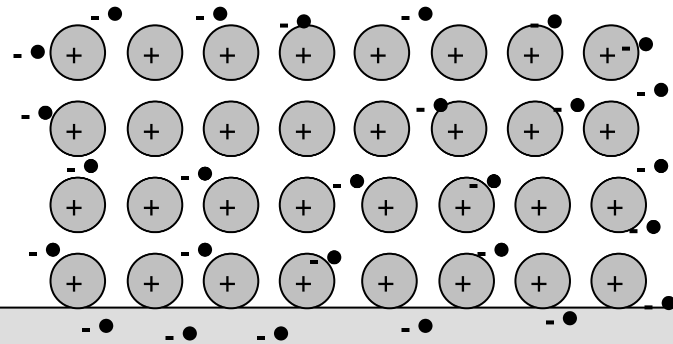
- Generally the trend is for the melting and boiling points of all molecules to rise with greater molar mass but HF has an unusually **high** m.p. and b.p. compared with HCl, although its molar mass is smaller. This is an indication of the strength of the hydrogen-bonding present in HF but absent in HCl.

8. **Dispersion forces** are present in all substances, but with non-polar molecules these are the only **attractive** forces, otherwise they could never exist as anything but gases. Dispersion forces are the **weakest forces** of all and are caused by an attraction between **temporary molecular dipoles** formed because of the separation (dispersion) of the protons and electrons in the atoms or molecules.



Dispersion forces **increase with atomic number** as the more electrons that are present, the greater the temporary dipole value and therefore the greater the melting and boiling point of the substance. E.g. melting point of Noble gases:
He = -272 °C, Ne = -249 °C, Ar = -189°C.

9. **Metallic bonding:** Due to the **low ionisation energy** of metals, at room temperature there is sufficient heat energy available to completely **ionise** all the valency electrons of metals. Consequently these electrons become **delocalised** and hence can move throughout the **lattice of positive metal ions**. This availability of free electrons means that all metals will be **good conductors** of hat and electricity.



The attractive force between the lattice of positive ions and the electron cloud in metals gives rise to **metallic bonding** and accounts for their high melting and boiling points and also their other physical properties, such as ductility and malleability.

10. Solvents that are composed of molecules with a net dipole are called **polar solvents**. Examples are H_2O, NH_3, CH_3Cl, etc. Polar solvents tend to dissolves polar solutes and non-polar solvents tend to dissolve non-polar solutes e.g. sugar/salt dissolves in water (all polar) and grease dissolves in petrol (both non-polar). [Also see Unit 6 on Common Organic Substances for a more detailed explanation of this].

Checklist – can you:

1. Define the Ionisation Energy (IE) of an element?
2. Describe the trends in IE for elements across Periods and down Groups?
3. Explain how ionic bonding allows elements to achieve a stable electronic configuration?
4. Draw and explain the formation of 3-D ionic lattices and the way this type of bonding affects physical properties?
5. Name some ionic compounds?
6. Draw and explain the formation of covalent molecular solids and the way their bonding affects physical properties?

7. Name some covalent molecular compounds?
8. Explain the 3-D bonding arrangement of graphite and the way this affects its physical properties?
9. Recall the names and covalent bonding structure of some common polyatomic ions?
10. Recall the colours of some common polyatomic ions?
11. Explain how the electronegativity of elements affects the polar nature of the bonds formed?
12. Describe how electronegativity changes across Periods and down Groups?
13. Explain how dipole moments can add or subtract to produce an overall polar or non-polar molecule?
14. Explain how VSEPR theory predicts the shapes and polar nature of compound such as H_2O and NH_3?
15. Explain how the polar nature of molecules affects physical properties, such as melting and boiling points?
16. Explain the phenomenon, and give examples of, substances containing hydrogen-bonds and their affect on physical properties, such as melting and boiling points?
17. Explain the phenomenon, and give examples of, substances having only dispersion forces and their affect on physical properties, such as melting and boiling points?
18. Explain the phenomenon of metallic bonding and how this accounts for the physical properties of metals e.g. melting and boiling points, malleability and conductivity?
19. Recall some common polar and non-polar solvents?
20. Explain the link between solubility of different solutes in various types of solvents and the polarity of the solvent?

Unit 2

ATOMIC STRUCTURE & BONDING **25 minutes**

Bonding

Key Exam Questions

1. a) Which would you expect to have the highest value for its first ionization energy, magnesium or calcium? Why? [3 marks]

b) Which would you expect to have the highest value for its first ionization energy, sodium or chlorine? Why? [3 marks]

c) An element X has ionization energies (in $kJmol^{-1}$) for successive electron removals as follows: 1^{st} 584, 2^{nd} 1823, 3^{rd} 2751, 4^{th} 11,584, 5^{th} 14,837. To what group in the periodic table does element X belong? Explain. [3 marks]

2. a) Draw an electron-dot diagram for calcium chloride and describe the type of bonding present. [3 marks]

b) In the above compound, what elements are the ions present isoelectronic with? [2 marks]

c) What is the electron configuration of the calcium ion in calcium chloride? [2 marks]

3. a) If a sample of rock-salt (NaCl) is struck with a hammer it breaks cleanly into cuboid pieces. Explain this in terms of the bonding present in the mineral. [2 marks]

b) A miner is convinced that he can obtain sodium metal by the electrolysis of a solution of rock-salt, but fails in his attempt. Explain why this is so and describe a successful method by which sodium could be obtained from this mineral.[3 marks]

Hint 1

4. List: CO_2, K_2SO_4, SiO_2, CCl_4, $Ca(OH)_2$, NH_3.

a) From the above list, classify the chemicals into groups of: (i) Molecular compounds, (ii) ionic compounds, (iii) covalent network solids. [6 marks]

b) Explain, in terms of the intermolecular forces present, which one of the above would have the highest boiling point and which one would have the lowest boiling point. *Hint 2* [4 marks]

c) Draw electron-dot diagrams for SiO_2, CCl_4 and NH_3. [6 marks]

5. a) The gas SO_2 contains a "co-ordinate covalent bond". Explain the characteristics of this type of bond and use its electron-dot diagram to illustrate this. [3 marks]

b) Draw the electron-dot diagrams for the nitrate and sulfate polyatomic ions

Hint 3 [4 marks]

6. List: F_2O, CF_4, $BeCl_2$, N_2, PCl_3, CO_2, BBr_3.

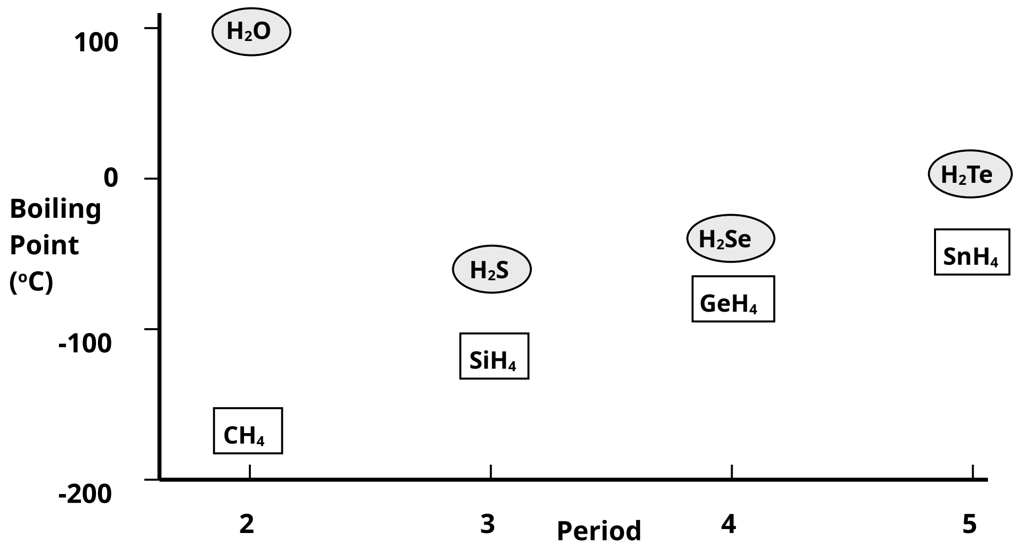
a) Place each of the compounds in the list into the correct classification boxes below which correctly describes it. [7 marks]

A polar compound with polar bonds	A polar compound with non-polar bonds	A non-polar compound with polar bonds	A non-polar compound with non-polar bonds

b) Explain how the shape of a covalent compound can affect its overall polarity

[4 marks]

7. The graph below compares boiling points of the hydrides of Group IV elements with the hydrides of Group VI elements.



Explain the following trends in the Periodic Table with reference to the intermolecular forces present:

a) There is a gradual rise in the boiling points of the Group IV hydrides from carbon to tin. [2 marks]

b) Group VI hydrides have higher boiling points than Group IV hydrides. [2 marks]

c) The Group VI hydride with the lowest molar mass has the highest boiling point.

[2 marks]

8. a) Explain why the noble gases are monatomic and have low boiling points[4 marks]

b) Describe the type of bonding present in copper and how this explains its physical properties. [3 marks]

9. Nappy rash in small babies is caused through blistering of the skin when ammonia comes into contact with it. The compound urea, from the baby's urine, breaks down into ammonia under bacterial action. Nappy rash can be reduced by applying Vaseline to the

baby's skin before putting on a new nappy. Vaseline (or petroleum jelly) is a member of the organic alkane series.

Explain, in terms of the polarity of bonds, why this treatment of nappy rash is particularly effective. *Hint 4* [3 marks]

Hint 1: Think about reactive products
Hint 2: Think about intermolecular forces
Hint 3: Some bonds could be co-ordinate
Hint 4: Think about solubility

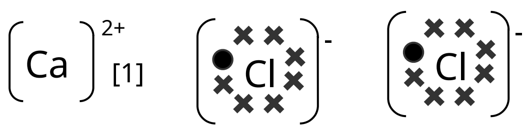
Unit 2
ATOMIC STRUCTURE & BONDING
Bonding

Key Exam Questions - Answers

1. a) Magnesium would have the higher IE because it is in the same Group as calcium, but with lower atomic weight [1] Calcium has another shell of electrons to repel the outer valence electron, [1] making it easier to extract. [1]

   b) Chlorine would have the higher IE than sodium [1] as, although they are in the same period, [1] it has more protons attracting the outer electron. [1]

   c) Group 3. [1] There is a large increase in IE between the 3rd and 4th electron extraction, showing that the 4th electron is in a full shell [1] (more energy required to extract it). So 3 electrons must be in the valence shell. [1]

2. a) 

[1] [1]

   Ionic bonding is present, where the 2 valence electrons from calcium are donated, one to each chlorine atom. [1]

   b) Ca^{2+} is isoelectronic with argon. [1]
   Cl^- is also isoelectronic with argon [1]

   c) Ca^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6$ [2]

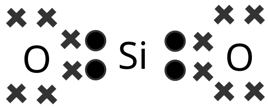
3. a) NaCl is bonded as Na^+ and Cl^- ions in an ionic network lattice. [1] It is brittle because, when the ions are moved slightly out of line with a hammer blow, repulsion of + ions occurs, causing the lattice to split along the crystal plane. [1]

   b) Electrolysing a solution of NaCl would only produce hydrogen and oxygen, as these elements are preferentially deposited and any sodium produced would immediately react with the water. [1]
   If NaCl is heated to melting point [1] and then electrolysed (without water) sodium would be produced at the cathode. [1]

4. a) (i) Molecular compounds are CO_2, CCl_4, NH_3 [3]
   (ii) Ionic compounds are K_2SO_4, $Ca(OH)_2$ [2]
   (iii) Covalent molecular solid is SiO_2 [1]

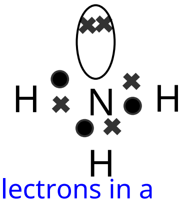
   b) SiO_2 would have the highest boiling point [1] as covalent network solids have very strong 3-D bonds which need to be overcome by heat energy. [1]
   CO_2 and NH_3 are both gases but hydrogen-bonding would give NH_3 a higher boiling point than CO_2. [1] CO_2 would boil at the lowest temperature as it has only dispersion forces holding the molecules together. [1]

c)



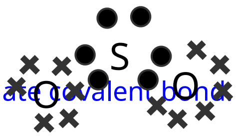
[2]

[2]



[2]

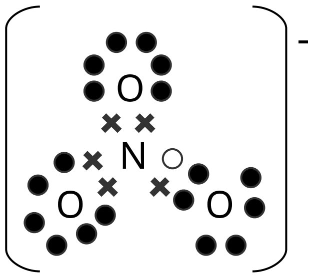
5. a) A co-ordinate covalent bond is one where one atom donates **both** electrons in a pair to form a covalent bond. [1]



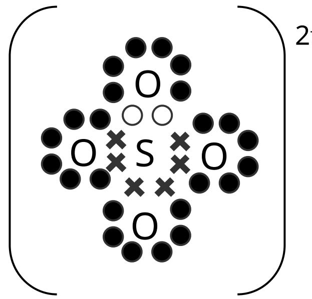
[2]

Both electrons in the right hand S-O bond come from the sulfur atom. This is a co-ordinate covalent bond.

b)



[2]



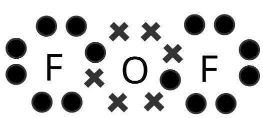
[2]

6. a)

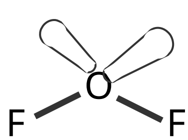
A polar compound with polar bonds	A polar compound with non-polar bonds	A non-polar compound with polar bonds	A non-polar compound with non-polar bonds
F_2O	none	CF_4	N_2
PCl_3		$BeCl_2$	
		CO_2	
		BBr_3	

[1 mark each]

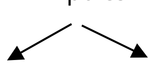
b) Equal polar dipoles can cancel out if a molecule is flat but will add up vectorally if a molecule is bent e.g. F_2O [2]



This appears flat on paper but the 2 non-bonding pairs will repel the 2 fluorine atoms downwards producing a net dipole moment. [2]



Dipoles



Net dipole down

7. a) Down Group IV, as the molar mass increases, the attractive dispersion forces also increase, raising the boiling point. [2]
   b) Group VI hydrides have a higher molar mass than Group IV. As the molar mass increases the attractive dispersion forces also increase, raising the boiling point. [2]
   c) The lowest molar mass Group VI hydride (H_2O) has strong hydrogen-bonds between its molecules which give it an unusually high boiling point. [2]

8. a) Noble gases have full shells [1] and so do not need to bond into pairs like hydrogen, oxygen, chlorine, etc. to achieve stability. [1]
   The only forces present between noble gas atoms are dispersion forces [1] which are very weak, allowing atoms to separate and boil at low energy (temperature) values. [1]
   b) Copper has metallic bonding holding its atoms together [1]. This makes it quite strong [1] with delocalised valence electrons allowing it to conduct heat and electricity easily [1]

9. Ammonia is a polar molecule [1] and Vaseline, being an alkane, will be non-polar. [1] This means that the ammonia will not dissolve in the Vaseline and penetrate the baby's skin. [1]