

ATOMIC STRUCTURE AND BONDING

3AB CHEMISTRY 2010

Atomic structure and Periodic Table

- Explain the structure of the atom in terms of protons, neutrons and electrons
- Write the electron configuration using the shell model for the first twenty elements e.g. Na. 2, 8, 1
- Explain trends in ionisation energy, atomic radius and electronegativity across periods and down groups (for main group elements) in the Periodic Table
- Describe and explain the relationship between the number of valence electrons and an element's.
 - Bonding capacity
 - Position on Periodic Table
 - Physical and chemical properties_

Bonding

- Describe and apply the relationships between the physical properties and the structure of ionic, metallic, covalent network and covalent molecular substances
- Use the Valence Shell Electron Pair Repulsion (VSEPR) theory and Lewis structure diagrams to explain and predict and draw the shape of molecules and polyatomic ions (octet only)
- Explain polar and non-polar covalent bonds in terms of the electronegativity of the atoms involved in the bond formation
- Use the relationship between molecule shape and bond polarity to predict and explain the polarity of a molecule
- Explain the differences between intermolecular and intramolecular forces
- Describe and explain the origin and relative strength of the following intermolecular interactions for molecules of a similar size:
 - Dispersion forces
 - Dipole-dipole attractions
 - Hydrogen bonds
 - Ion-dipole interactions such as solvation of ions in aqueous solution
- Explain the relationships between physical properties such as melting and boiling point, and the types of intermolecular forces present in substances of similar size
- Apply an understanding of intermolecular interactions to explain the trends in melting and boiling points of hydrides of groups 15, 16 and 17 accounting for the anomalous behaviour of NH_3 , H_2O and HF
- Describe and explain the nature of the interaction between solute and solvent particles in a solution

Use the nature of the interactions, including the formation of ion-dipole and hydrogen bonds to explain water's ability to dissolve ionic, polar and non-polar solutes

Bonding

Electron Configuration and the Periodic Table

When elements are listed by atomic number similar outer level (valence level) configurations occur at regular intervals.

H				He
Li	Be	O	F	Ne
Na	Mg	S	Cl	Ar
K	Ca			

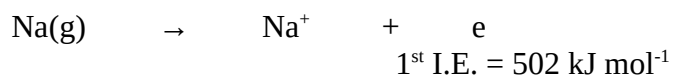
- **Groups** of elements have similar valence level configurations.
- They are chemically similar.
- Typical properties become more extreme ↓ metallic groups and ↑ non-metallic groups.
- The chemical properties of the elements are dependent on their valence level configurations.

Chemical Properties and Ion Formation

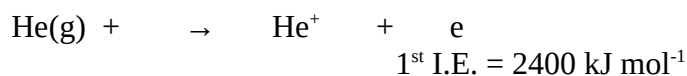
Metallic elements readily lose electrons to form +ive ions but some energy is needed to remove electrons from the influence of the nucleus. The amount of energy required to remove 1 mol of the most loosely bound electrons from the atoms of an element is called the **Ionisation Energy**.

The removal of the outermost electron from an atom is the 1st Ionisation Energy.

e.g. for sodium



For elements with complete octets of valence electrons the values are typically high.



Trends in First Ionisation Energy Values

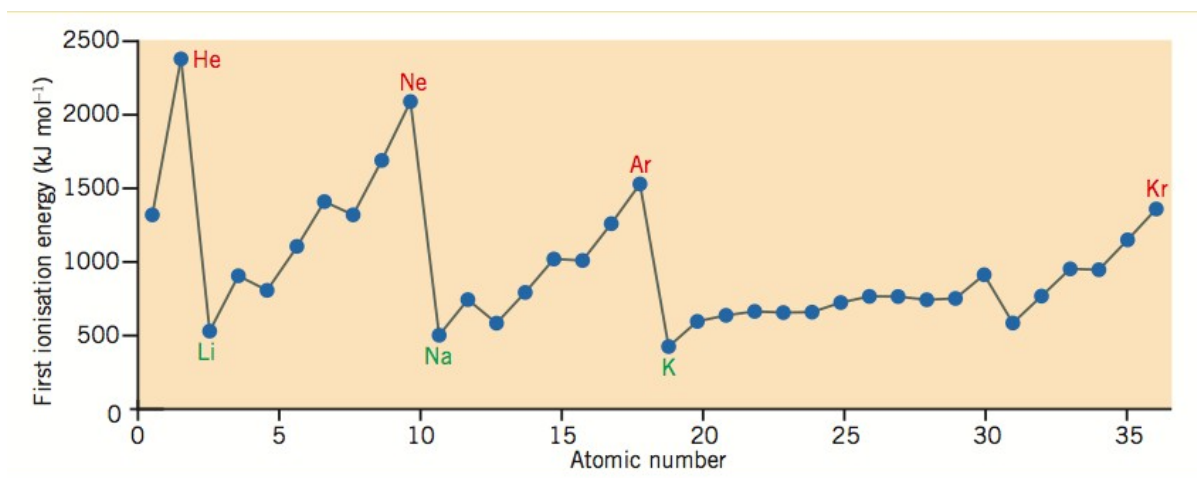


Figure 1.15 The first ionisation energies of elements from the first four periods of the periodic table

This trend shows:

	<u>Elements</u>	<u>1st I.E.</u>	(Obtain approx values from Fig 1.15 p20)
1	hydrogen		
3	lithium		
11	sodium		
19	potassium		

and elements

3	lithium
5	boron
8	neon

Describe the trends

-
-

Periodic Trends show a gradual decrease in 1st I.E. as Z increases down a Group and an increase across a period.

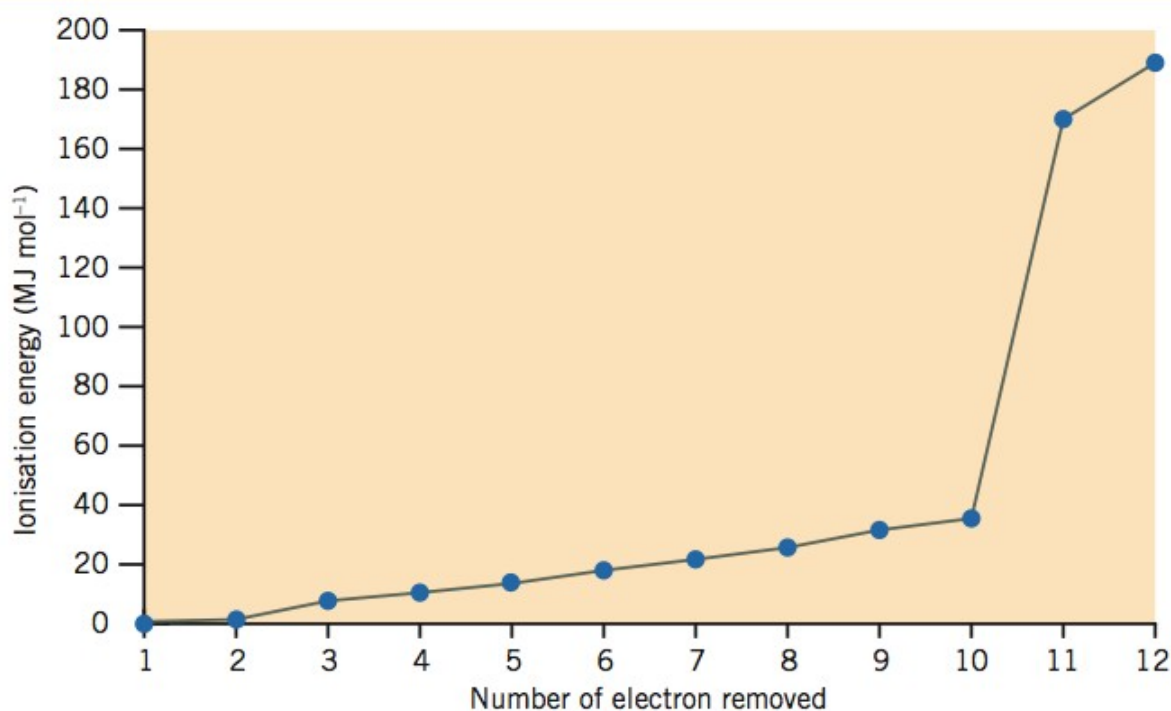
What are the reasons for this?

-
-

Successive Ionisation Energies (2nd, 3rd etc)

Successive ionisation energies provide us with evidence to support our theories about electronic configuration.

The removal of successive electrons from an atom requires increasing amounts of energy as the nuclear attraction for electrons increases with increasing positive charge.



- Consider Table 1.5, how does this data support the configuration 2,8,2 for magnesium?
- Explain how the data presented in Table 1.4 supports the same theory.
- Answer the questions in Ex 1.2
- Answer qu 1 to 4, 6, 9, 10 from Page 24

Electronegativity

An understanding of ionisation energy helps us explain how atoms can lose electrons to form positively charged ions but this does not explain why the opposite happens and negatively charged ions are formed. For this, we need to look at electronegativity.

TABLE 1.7 ELECTRONEGATIVITIES OF THE MAIN GROUP ELEMENTS

H 2.20						
Li 0.98	Be 1.57	B 2.04	C 2.55	N 3.04	O 3.44	F 3.98
Na 0.93	Mg 1.31	Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16
K 0.82	Ca 1.00	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96
Rb 0.82	Sr 0.95	In 1.78	Sn 2.96	Sb 2.05	Te 2.1	I 2.66
Cs 0.79	Ba 0.89	Tl 2.04	Pb 2.33	Bi 2.02	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

These give a measure of the electron attracting power of atoms within a molecule.

There is a Periodic trend here too.

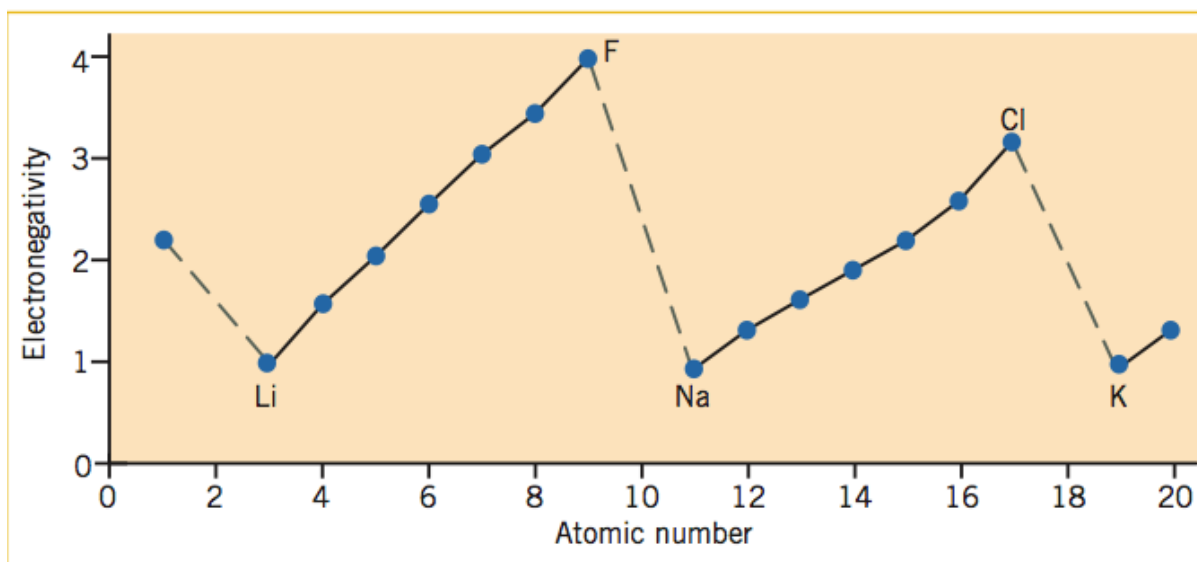


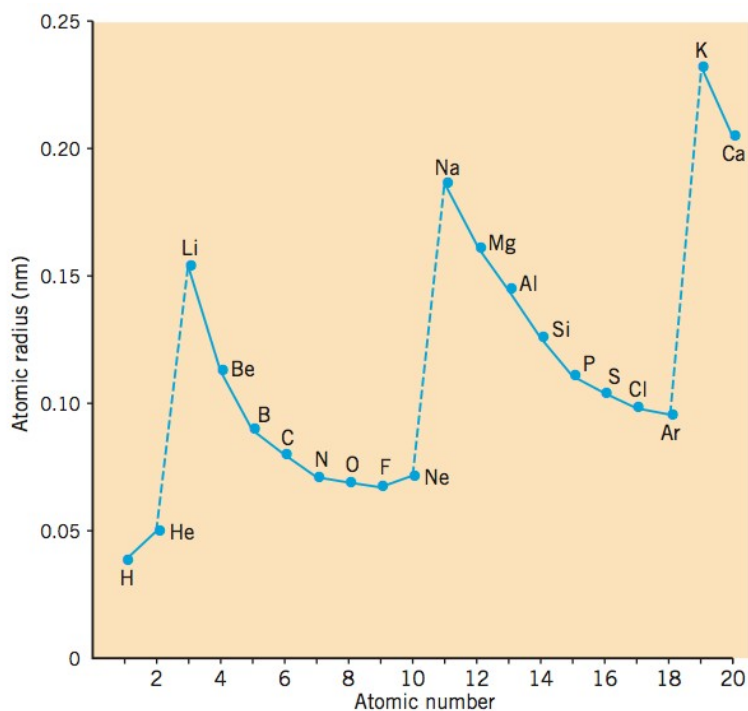
Figure 1.17 Electronegativity values of the first 20 elements in the periodic table

The values increase across the rows and decrease down the groups.

- Explain why this is so.

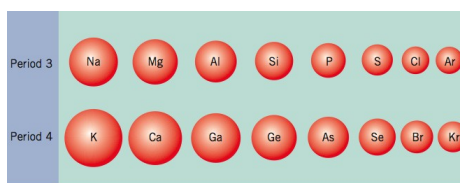
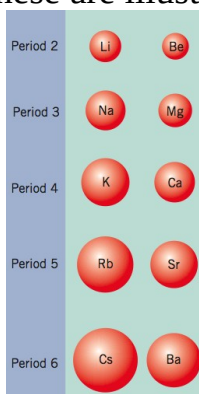
Ionisation energy, electronegativity and atomic radius

Just as the IE and electronegativity are related, so too is the atomic radius. Its size is determined by the number of protons in the nucleus, the number of electrons and the number of electron shells. The values for atomic radii show a periodic trend.



The trends seem to be:

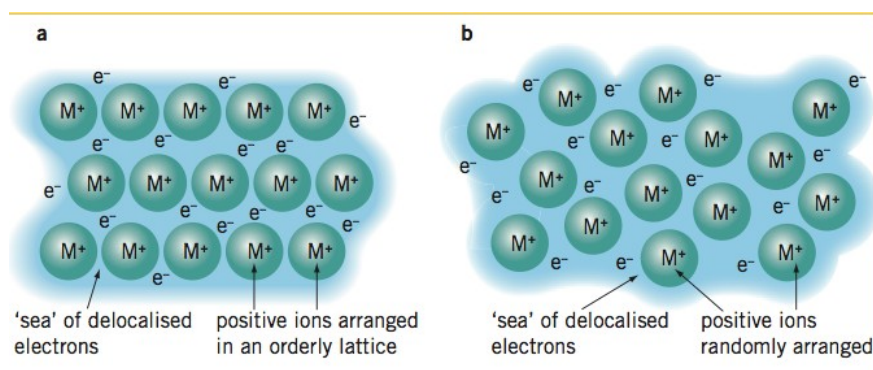
And these are illustrated by these diagrams



Review exercise 1.4 Qu 1, 3, 5

Metallic Bonding

We already know that the bonding in metals is quite different from that found in other substances.



The bonding found in metals is a consequence of the low values for the ionisation energies for the valence electrons.

Element	Metal or non-metal	First four ionisation energies (MJ mol^{-1})			
		1st	2nd	3rd	4th
Sodium	Metal	0.502	4.569	6.919	9.550
Calcium	Metal	0.596	1.152	4.918	6.480
Iron	Metal	0.766	1.567	2.964	5.29
Nitrogen	Non-metal	1.407	2.862	4.585	7.482
Chlorine	Non-metal	1.257	2.303	3.828	5.164

The 'sea of delocalised electrons becomes delocalised as ions attract electrons from adjacent atoms. This creates a series of non-directional bonds, which contribute to the physical properties of the metals. (See Table 2.2 p30)

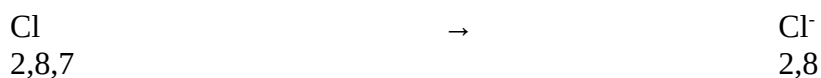
Review exercise 2.1 Qu 2, 3, 4, 5

Ionic and Covalent bonding

Ionic bonding principally occurs between a metallic and a non-metallic element.

e.g. sodium chloride

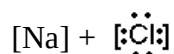
The metal atom donates its valence electrons to the non-metallic element to form charged particles called ions so that each element gains a noble gas electron configuration.



and the ions attract each other electrostatically.



Such compounds may be represented by *dot diagrams* as follows



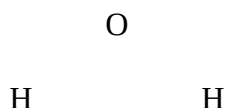
Draw dot diagrams to represent the ionic bonding present in the compounds calcium iodide, potassium sulfide, magnesium oxide, aluminium bromide, iron (III) oxide.

Covalent compounds are usually formed between non-metallic elements.

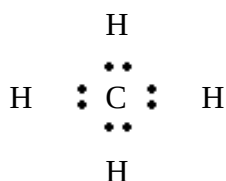
e.g. in water.

The valence configurations of each element involved in the bonding attains that of a noble gas by the sharing of electrons to form bonding pairs that are the covalent bonds. Such sharing may form a single bond (one shared electron pair) or a multiple bond (more than 1 shared pair).

In water the covalent bonds are single bonds.



Similarly with methane CH_4



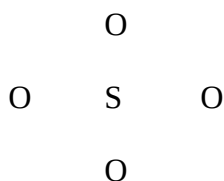
Draw electron dot diagrams to represent the covalent bonding present in ammonia, methanol (CH_3OH), carbon dioxide, sulfur trioxide.

For polyatomic ions and molecules the dot diagram illustrates which electrons are involved in bonding and which are non-bonding or lone pairs of electrons. To construct such diagrams we may follow a sequence of general guidelines. (Using the sulfate ion as an example).

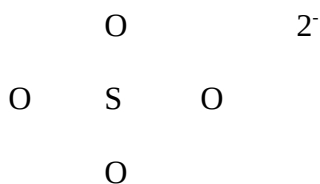
1. Decide which of the atoms in the ion or compound is to be central atom.
(in SO_4^{2-} S will be the central atom).
2. The charge on the ion determines the number of electrons to be placed around the central atom. (As the charge on sulfate is 2- we place sulfur's 6 electrons + 2 extras around the S atom).

S

3. Arrange the other atoms around the central atom such that all atoms have a complete octet of electrons around them. This may necessitate 'rotating' atoms to form multiple bonds.



4. The charge on the ion is distributed around the whole structure by placing the diagram in square brackets and defining the charge outside the bracket.



Draw electron dot diagrams to show the bonding present in the following substances

ammonia, ammonium ion, carbonate ion, phosphate ion, chlorate ion (ClO_3^-)

sodium carbonate, potassium cyanide, ammonium hydrogen carbonate.

Covalent Network Substances

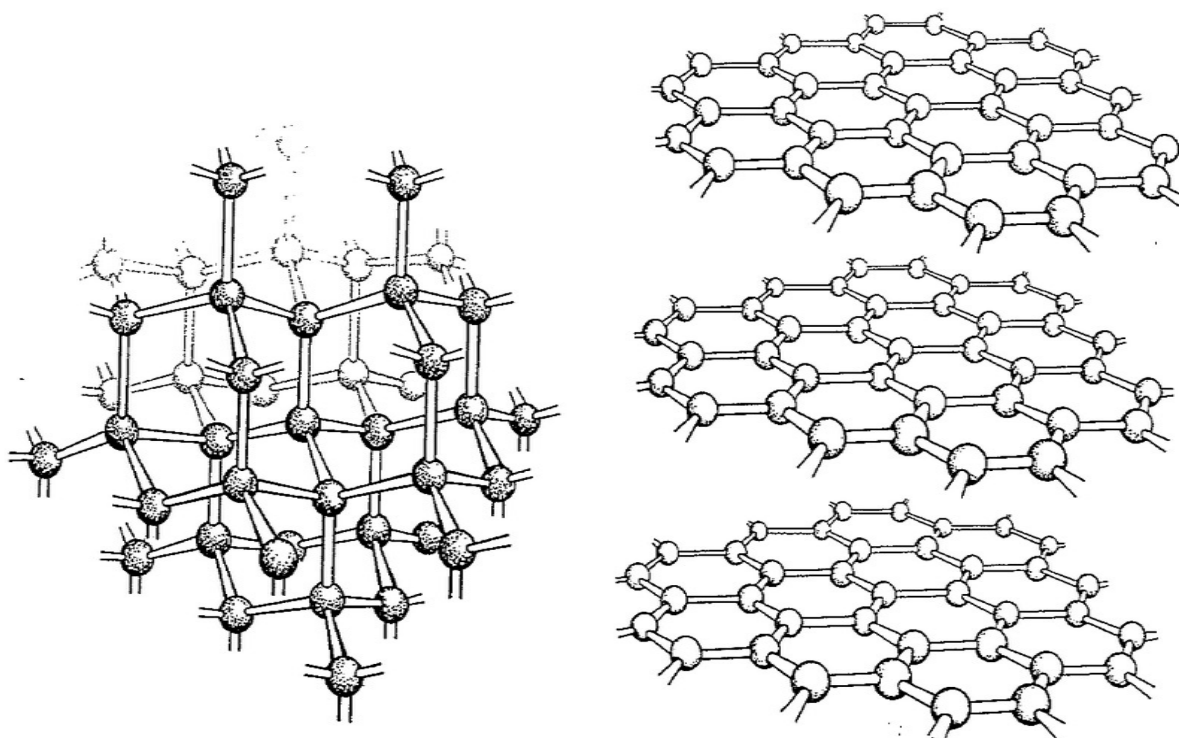
(continuous lattices, infinite arrays, giant structures)

Atoms occupy regular lattice positions.

Atoms are bound covalently.

Each atom is bound to several others.

Each crystal is a macromolecule.



Associated properties

Poor electrical conductors

Very high melting points

Hard

Brittle

- **Graphite is an exceptional substance. Use the diagrams above to help you to clearly describe and explain why.**

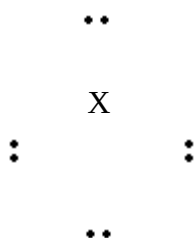
Electron pair repulsion and the shapes of molecules

The shape of a molecule is determined by the electron pair arrangement around its central atom.

The valence electrons of an atom involved in covalent bonding form pairs. These pairs may be
bonding pairs
lone pairs (non-bonding)

The theory that we are going to examine (VSEPR) says that these pairs are arranged so that they maximise the distance between them and so minimise the repulsions that occur between them.

For a complete octet the optimum position of the electron pairs describes a tetrahedron around the central atom.



The theory allows us to predict the shapes of molecules provided we can draw their electron dot diagrams.

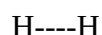
Molecules with only one bonding electron pair (two atoms)

Take the simplest form of molecule – diatomic hydrogen H_2 .



With only one bonding pair in the molecule it must be LINEAR

The shape would be drawn



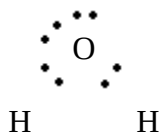
Other diatomic molecules (O_2 , HCl , HBr and so on) must have this linear arrangement of their atoms. Draw their dot diagrams and shapes here.

Molecules with two bonding electron pairs (three atoms)

There are two possible shapes, determined by the presence or absence of lone pairs.

LINEAR $\text{F} \text{ : } \text{Be} \text{ : } \text{F}$

BENT



Draw dot diagrams and shapes for these molecular compounds.

CS_2

H_2S

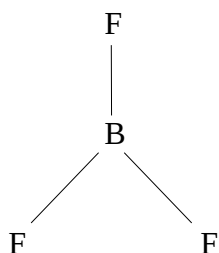
CO_2

SO_2

Molecules with three bonding electron pairs (four atoms)

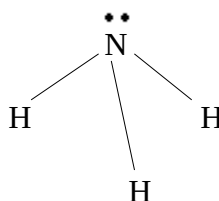
Again there are two possible shapes.

TRIGONAL PLANAR



(This is a non-octet molecule)

PYRAMIDAL



Draw electron dot diagrams and shapes for

SO_3

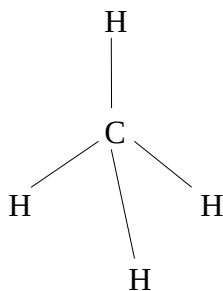
CO_3^{2-}

SO_3^{2-}

PCl_3

Molecule with four bonding electron pairs (five atoms)

These molecules must be TETRAHEDRAL.



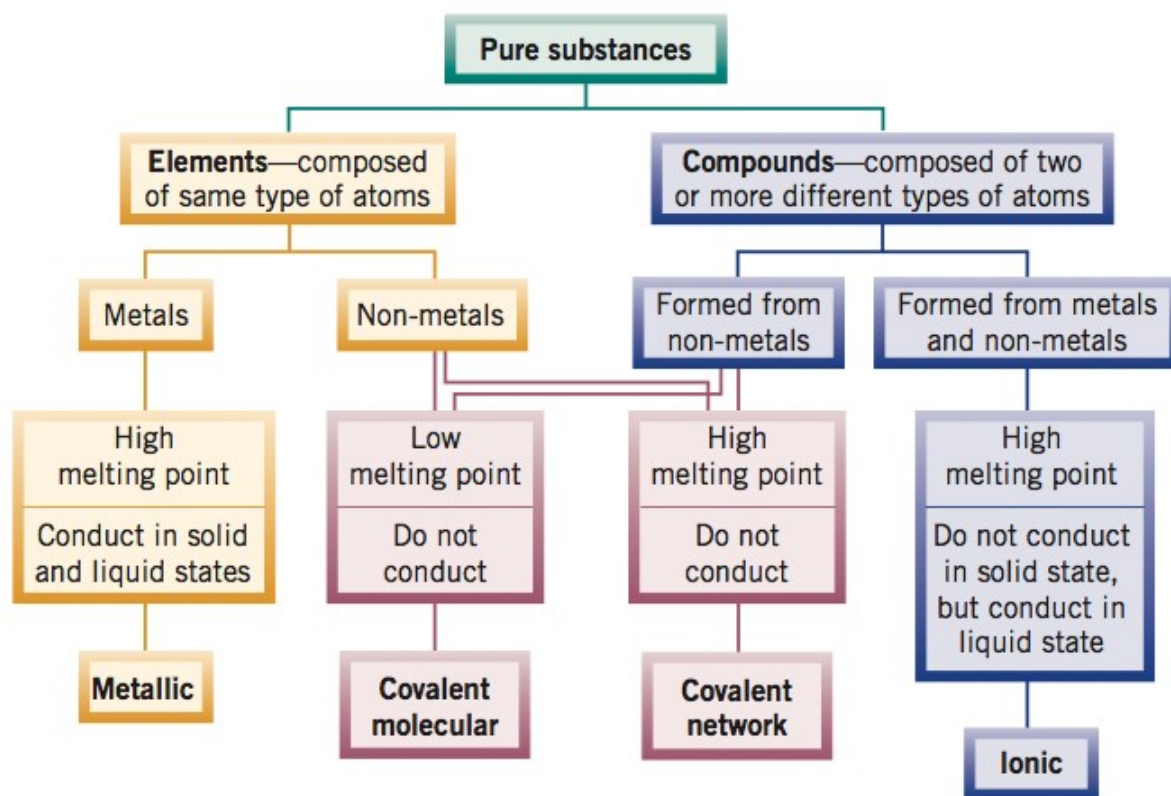
Draw dot diagrams and shapes for



Complete Review Ex2.3 Qu 1, 2, 3, 4, 5, 6, 7

Ex 2.4 Qu 1 to 4

Page 51 qu 1, 3, 5, 6, 7, 12, 15



Electronegativity differences and bond polarity

Electronegativity is a measure of an atom's ability to attract electrons towards its nucleus. Electronegativity values show periodicity. They tend to increase from left to right across the rows of the Periodic Table and decrease from top to bottom of each group. So, the non-metals show greater attraction for electrons than do the metals in the same period.

Try to explain why.

Electronegativity accounts for covalent bond formation.

The attraction of a nucleus for the valence electrons of a nearby atom within a pair of atoms causes the electron to be shared. Can you represent this diagrammatically?

In diatomic molecules of elements this sharing is equal and the shared electron pair spends as much time under the influence of one atom as the other.

This is PURE COVALENT BONDING.

Where the atoms sharing the electron pair have different electronegativity values the more electronegative has more influence over the bonding pair than the less electronegative.

This uneven sharing is POLAR COVALENT BONDING

e.g.	HCl	H-----Cl
	Electronegativity	2.2 3.2

The bonding pair is influenced more by the chlorine atom than by the hydrogen atom.

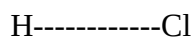
The Cl end of the molecule will be negative compared with the H end.

delta positive	$\delta+$	$\delta-$	delta negative
	H-----Cl		

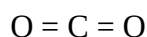
Polar covalency represents the transition between ionic and pure covalent bonding.

Bond polarity and molecular polarity

Where covalent molecules contain polar bonds the whole molecule may or may not be polar in nature.



polar bond?
show the charge distribution.
polar molecule?



polar bonds?
show the charge distribution.
polar molecule?

Draw electron dot diagrams for methane, fluoromethane, boron trifluoride (non-octet), water and ammonia.

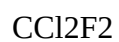
Draw shapes for the molecules listed above and include the bond polarities.
State whether each molecule is polar or non-polar.

In summary, which molecules are polar and which non-polar? (Is there a general rule that we can look to?)

Review exercise 3.4 Qu 1, 3

For each of the following molecules/ions draw a dot diagram, name the shape and determine whether it is polar or non-polar.

<u>Molecule</u>	<u>Dot diagram</u>	<u>Shape</u>	<u>Polar/Non-polar</u>
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- Review Ex 3.1 Qu 1 and 2
- Ex 3.2 Qu 2, 3, 4, 5

The Physical Properties of Covalent Molecular Substances

In metallic, ionic and network solids the properties are determined largely by forces existing within the structures,

Melting and boiling points are influenced by the amount of energy needed to disrupt metallic, ionic or covalent bonding respectively.

In **covalent molecular** solids these properties are ***not*** dependent on breaking covalent bonds. Rather, the mp, bp and solubility of molecular solids are determined by forces that exist between molecules (inter-molecular forces).

What evidence do we have that inter-molecular forces actually exist?

There are a number of types of inter-molecular force that operate in molecular substances and the nature of these forces is determined by the shape and molecular polarity of the molecules involved.

The types of inter-molecular force are:

- Dispersion forces (temporary induced dipoles)
- Dipole to dipole attractions
- Hydrogen bonds

These attractions are collectively described as Van der Waal's Forces.

Before we consider the type(s) of inter-molecular force found within a molecular substance we need to revise our ideas on polarity.

What molecular shapes do we know about?

Complete the table below. Where bonds are polar covalent in nature indicate the bond polarity. (+ or -)

Name of shape	Drawing of shape	Example of a polar covalent compound	Example of a non-polar covalent compound
Linear			
Bent			
Triangular planar			
Pyramidal			
Tetrahedral			

Inter-Molecular forces (Van der Waal's Forces)

Dispersion forces

Dispersion forces are temporary induced dipoles. They are the main force acting between non-polar molecules though it must be noted that they contribute to the forces attracting all molecules.

The capacity of monatomic gases to be liquefied and solidified is evidence to suggest that some kind of force exists between the atoms.

Look at fig 3.16, fig 3.17 and 3.18

Draw the diagrams here and explain what is happening to create an attractive force between the atoms.

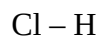
Dispersion forces increase in magnitude with increasing molar mass. This is reflected in melting point data.

e.g.	M_r	M Point ($^{\circ}\text{C}$)
He	4	-270
Ne	20	-249
Ar	40	-189
Kr	84	-157
F_2	38	-220
Cl_2	71	-101
Br_2	160	-7
I_2	254	114

This occurs because the strength of the dispersion force depends on the size of the electron charge cloud creating it.

Dipole to dipole attraction

Neighbouring polar molecules orientate themselves such that unlike – charged ends are adjacent.



A weak dipole to dipole attraction occurs which influences the physical properties of the substance. These dipole to dipole attractions must be disrupted to melt or boil it.

The melting and boiling points of polar substances would be expected to be greater than those of non-polar substance of similar molar mass.

	Substance	Molar mass	Boiling point °C
e.g.	Oxygen	32	-183
	Nitrogen monoxide	30	-152
	Krypton	84	-152
	Hydrogen bromide	81	-67
	Bromine	160	59
	Iodine chloride	162	97

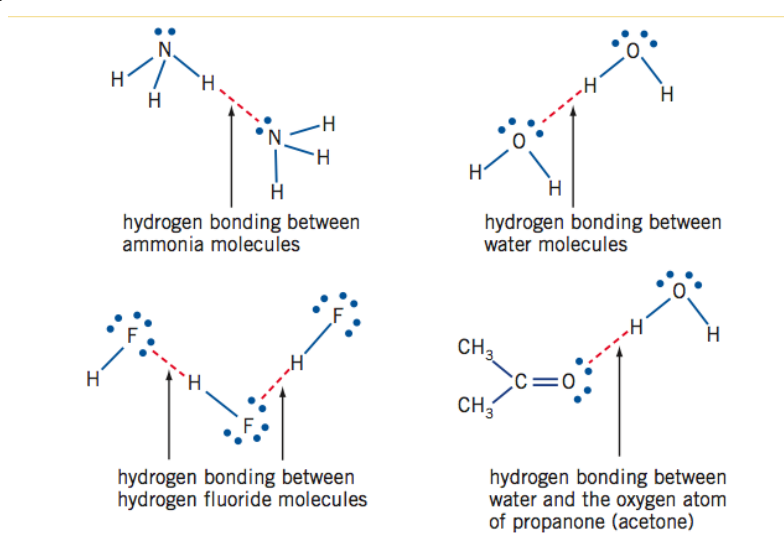
Review ex 3.6 Qu 1, 2, 3

Hydrogen Bonding

Hydrogen bonding occurs between molecules containing hydrogen bound to one of the three most electronegative elements – oxygen, fluorine or nitrogen.
e.g. water, ammonia, hydrogen fluoride, ethanol, sucrose

These forces are stronger than other inter-molecular attractions because

- The charge separation is high.
- The atoms of O, F and N have lone pairs of electrons attracted to the hydrogen atom of an adjacent molecule.
- Hydrogen atoms that are covalently bound do not enjoy any shielding of the electrons of adjacent molecules from the nucleus.
- Hydrogen atoms are very small and can approach the δ^- portion of adjacent molecules very closely.



The effect of hydrogen bonding on physical properties is described by fig 3.20

Draw a graph similar to fig 3.20 but include only the data for Group VI

The influence of hydrogen bonding on compounds is most clearly seen in those containing the hydroxyl group -OH (e.g. water, ethanol, acetic acid).

Hydrogen bonding and the properties of water

- Boiling point – exceptionally high in relation to other hydrides of Gp VI elements.
- Expansion on freezing – each water molecule hydrogen bonds to 4 others when freezing. This holds them at fixed distances apart.
- Specific heat capacity – amongst the highest values known.
- Surface tension – molecules at the surface of water are attracted sideways and downwards forcing the water to form spherical droplets.

Draw the structures of water, ethanol and acetic acid.

In each case circle the hydroxyl group.

For each substance draw diagrams to show how the hydroxyl groups are involved in hydrogen bonding between molecules.

Review exercise 3.7 Qu 1, 3, 4, 5, 6

Questions p 88 1, 3, 4, 5, 9,

Solubility in Liquids

A solution is a homogenous mixture of two or more substances.
Although solutions are common not all solutes will dissolve in all solvents.

- Oil spillages at sea are difficult to deal with
- Water cannot be sprayed onto an oil or petrol fire
- Dirty dishes cannot be cleaned effectively in water alone
- Our clothes need to be 'dry cleaned' rather than washed

Observations of the solubilities of molecular substances indicate the following

- Non-polar solutes dissolve in non-polar solvents
- Non-polar solutes do not dissolve in polar solvents
- Polar solutes dissolve in polar solvents
- Polar solutes do not dissolve in non-polar solvents

Explanation of the solubilities of molecular substances in molecular liquids

The strengths of intermolecular forces are the important factors determining the solubility of substances.

Two non-polar substances

e.g. benzene dissolves in hexane

- Benzene molecules separate from each other
- Hexane molecule separate from each other
- Hexane and benzene molecules are attracted to each other

This occurs because the intermolecular forces acting within each substance are similar – dispersion forces.

List some pairs of non-polar molecular substances that will readily mix in this way.

A non-polar substance with a polar substance

These interactions do not readily occur.

The strength of the intermolecular forces acting in the polar substance cannot be disrupted by the non-polar molecules.

List some pairs of molecular substances that will not dissolve in each other.

Two polar substances

e.g. alcohol in water

- Separation of alcohol molecules is resisted by the hydrogen bonds between them
- Separation of water molecules is resisted by the hydrogen bonds between them
- Alcohol and water dissolve because hydrogen bonds form between their molecules

List some pairs of polar substances that will dissolve because of hydrogen bonding.

List some pairs of polar substances that will dissolve because of dipole to dipole interactions

- *As the carbon chain in alcohol molecules gets longer the solubility in water decreases. Explain why.*
- *Propanol, C_3H_7OH is soluble in petrol (a non-polar hydrocarbon) but 1,2,3- propantriol is not. Explain why.*
- *Acetic acid is soluble in water but decanoic acid ($C_9H_{17}COOH$) is not. Explain why.*

The solubility of ionic solids in molecular liquids

When an ionic solid dissolves in water

- Ions are separated from each other
- Water molecules are separated from each other
- Ions surround the water molecules
- Ion – dipole attractions form

Using the Periodic Table as a guide to properties

- The Periodic Table is arranged in order of **increasing atomic number** (number of protons and hence number of electrons).
- **Metals** appear to the LHS of the table and **non-metals** to the RHS
- Similar elements occupy **Groups** in the table e.g.
 - Gp I the alkali metals
 - Gp II the alkaline earth metals
 - Gp VII the halogens
 - Gp VIII the noble gases
- Elements which differ slightly occupy **Series, Periods or Rows** e.g.
 - 2nd Row from Li to Ne
- Position in the table indicates the number of **valence electrons** and hence the most likely valency.
- The table is divided into **blocks s, p and d** and the elements within a block have similarities.
- The table enables us to predict the likely **bond type** for any combination of elements.
 - LHS with LHS – **metallic**
 - LHS with RHS – **ionic**
 - RHS with RHS – **covalent**

Trends that are evident in the table

Across the Rows left to right

- Increasing **1st Ionisation Energy**
- Decreasing **atomic radius**
- Increasing **electronegativity**
- Increasing **oxidising strength**
- Increasing **acidity** of oxides and hydroxides

Down the Groups top to bottom

- **1st IE** decreases
- **Electronegativity** decreases
- **Atomic radius** increases
- **Metallic** nature increases
- **Melting point and boiling point** of metals decreases
- **Melting point and boiling point** of non-metals increases

Metallic	Property	Explanation
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substances	Good conductors of electricity and heat	Delocalised electrons transfer charge and thermal energy
	High melting point and boiling point	Strong metallic bonding extending throughout the lattice
	Dense	Ions tightly packed in the lattice
	Malleable and ductile	Distortion does not disrupt the metallic bonding
	Lustrous	Delocalised electrons cause reflection of light

Ionic substances	Property	Explanation
	Non-conductors of electricity when solid	Ions firmly bound in the lattice, so no mobile charged particles
	Conductors of electricity when molten or in aqueous solution	Ions free to move
	High melting point and boiling point	Strong ionic bonding extending throughout the lattice
	Hard	Ions strongly bound in the lattice
	Brittle	Distortion causes repulsion between ions of like charge

Covalent molecular substances	Property	Explanation
	Non-conductors of electricity when solid, molten or dissolve in water, unless a reaction occurs with water to produce ions	The molecules are uncharged and electrons are localised in covalent bonds or on the atoms
	Low melting point and boiling point	Weak forces between molecules
	Soft solids, liquids or gases	Weak forces between molecules

Covalent network substances	Property	Explanation
	Non-conductors when solid or molten	Electrons localised in covalent bonds or on the atoms
	Hard	Atoms strongly bound in the lattice
	Brittle	Distortion breaks covalent bonds

- Write the electron configurations for the following
 - silicon atom
 - argon atom
 - potassium atom
 - sulfide ion
 - magnesium ion
 - oxide ion
- Why is the first ionisation energy of Na greater than that of Cs?
- Why is the first ionisation energy of S greater than that of Mg?
- An unknown element X has successive ionisation energies of 500, 1010, 5200, 7100, 9200 (in kJ per mole)
 - To which group of the Periodic table does X belong? How did you know?
 - Write formulae to represent the compounds formed between X and”

chlorine, oxygen, bromine
- Sketch a rough graph (actual values not needed) to show the first 4 ionisation energies of Na plotted against number of electrons removed.
- List THREE important differences between ionic and covalent bonding.
- Draw neat electron dot diagrams to show the bonding in

carbon dioxide, ammonia, hydrogen sulfide, ethane, phosphorus trifluoride,
nitrogen gas, potassium oxide
- Draw electron dot diagrams to represent the bonding in the following polyatomic ions:

SO_4^{2-} , NO_3^- , HCO_3^- , ClO_2^-
- Which of the following compounds contains BOTH ionic and covalent bonding? Justify your answer.

hydrogen chloride, potassium chloride, trichloromethane, sodium phosphate,
water, ammonium chloride, ammonia
- What is the essential characteristic of the valence electrons in a metal? LIST 4 significant properties of metals which results from this characteristic.
- Describe 3 important differences in the properties of covalent network solids and covalent molecular solids. Clearly explain WHY these differences exist.

12. Elements X and Y react with oxygen to form oxides having the formulae XO_2 and Y_2O respectively. Some of the properties of these oxides are tabulated below:

OXIDE	BP	CONDUCTIVITY when MOLTEN	HARDENS
XO_2	132° C	nil	v soft
Y_2O	822 ° C	high	hard

To which class of solid does each oxide belong? Give a reason to support your answer.

13. The number of valence electrons of four elements are tabulated below:

ELEMENT	No of valence electrons
E	one
J	three
Q	five
Z	seven

For compounds between the following pairs of elements

- say whether the bonding is ionic or covalent
- write the chemical formula for that compound

(i) E and Z (ii) J and Z (iii) Q and Z

14. NAME one element or compound which fits each of the following descriptions:

- diatomic gas at room temperature
- a compound that is classified as a network solid
- a Group V metal
- a metal with a Melting Point $< 0^\circ\text{C}$
- an ionic solid in which the ratio of cations to anions is 2 : 1
- an element that exists as a covalent molecular solid at room temperature
- an element that reacts with oxygen gas to form another gas which has acidic properties
- a network solid which will conduct electricity to some extent
- the most reactive metal
- the most reactive non-metal

15. The first 5 ionisation energies of an element X are listed below:
- 793, 1585, 3238, 4366, 16 098 (all in kJ mol^{-1})
- Why is $I_3 > I_2$?
 - Why is I_5 substantially greater than I_4 ?
 - How many valence electrons does X have? To which group of the PT does X belong?
16. Clearly account for the following differences in properties:
- Copper metal is very malleable but NaCl is brittle.
 - Molten KBr is a good electrical conductor whereas molten candle wax does not conduct at all.
 - SiO_2 melts at 1845°C whereas CO_2 melts at -78°C .
17. Draw electron dot diagrams for the following ions:
- carbonate, sulfite, chlorate, nitrite
18. For each of the following molecules draw an electron dot diagram, name and draw the shape of the molecule:
- HCl
 - NCl_3
 - H_2O
 - CCl_4
 - SF_2
 - BBr_3
 - SO_3
 - PF_3
 - CH_2Cl_2
 - CH_3F
19. What are the shapes of these molecules? Name and draw them.
- SbH_3 H_2Te BeBr_2 SiF_4 CS_2
20. Draw diagrams to show how each of these solutions forms
- methane in ethane
 - ammonium nitrate in water
 - acetic acid in water
21. Why is the arrangement of water molecules around dissolved bromide ions different from the arrangement of water molecules around dissolved aluminium ions?
22. In which of the following solvents would you expect iodine to be most soluble?
methanol, ethanol or propanol
23. Butanol ($\text{C}_4\text{H}_9\text{OH}$) is soluble in both water and tetrachloromethane. Explain which features are responsible for this.