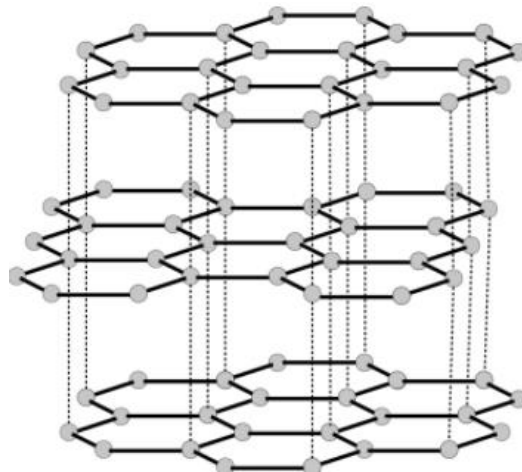


UNIT 2

PARTICLES, BONDING AND STRUCTURES

Teacher Version



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Contents

- Elements, Compounds and Mixtures
- Laws of Chemical Combination
- Simple Chemical Bonding
- Shapes of Molecules and Ions
- Intermolecular Forces
- The Kinetic Model of Matter
- The Gas Laws
- Chemical Structures
- Structure and Bonding in the Periodic Table

Key words: substance, material, element, compound, mixture, allotrope, law of constant composition, law of multiple proportions, valency, electronegativity, electropositive, electronegative, ionic bond, Lewis-dot structure, covalent bond, double bond, triple bond, dative covalent bond, molecule, intermolecular force, polyatomic ion, metallic bond, alloy, VSEPR theory, lone pairs, bonding pairs, linear, trigonal planar, tetrahedral, pyramidal, non-linear, temporary dipole, induced dipole, permanent dipole, Van der Waal's force, polar bond, polar molecule, hydrogen bond, kinetic model of matter, Celsius, Kelvin, absolute zero, solid, lattice, liquid, gas, melting, boiling, subliming, freezing, condensing, vapour pressure, evaporation, melting point, boiling point, state symbol, solution, solute, solvent, Brownian motion, diffusion, Graham's Law of Diffusion, pressure, Boyle's Law, Charles' Law, Gay-Lussac's Law, combined gas law, Dalton's Law of Partial Pressures structure, giant ionic, giant metallic, metal, non-metal, simple molecular, simple atomic, giant covalent, unit formula, molecular formula, allotrope, periodic

Units which must be completed before this unit can be attempted:

Unit 1 – Atoms and the Periodic Table

Estimated Teaching Time: 21 hours

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

UNIT 1 SUMMARY AND SYLLABUS REFERENCE

Lesson	Title and Syllabus Reference
1	Elements, Compounds and Mixtures; Laws of Chemical Combination CA4aiv laws of chemical combination (law of constant composition, law of multiple proportion); ISA4.2 Elements, compounds and mixtures (differences between elements, compounds and mixtures)
2	Electronegativity; Ionic Bonding CA2aii - Periodicity of the elements: trends in periodic properties - down a group and across a period (Periodic properties for the first 18 elements: electronegativity); CA3a inter-atomic bonding (Lewis dot structure for ionic compounds); CA3ai ionic bonding - factors influencing its formation (formation of stable compounds from ions, factors should include ionization energy, electron affinity and electronegativity); ISA4.3 ionic compounds (Ionic bond formation, IUPAC names of common ionic compounds)
3	Covalent Bonding, Dative Covalent Bonding, Molecules CA3a inter-atomic bonding (Lewis dot structure for covalent compounds); CA3aii covalent bonding - factors influencing covalent bond formation (co-ordinate bond as a type of covalent bond, factors should include electronegativity difference); Ca1biii molecules; ISA4.3 covalent compounds (covalent bond formation)
4	Metallic bonding; Summary of Interatomic bond Types CA3ai ionic bonding – factors affecting its formation (factors should include electronegativity difference); CA3aii covalent bonding – factors affecting covalent bond formation (factors should include electronegativity difference); CA3ci metallic bonding – factors affecting its formation (factors should include atomic radius, ionization potential, and number of valence electrons, type of specific packing not required); CA3e comparison of all bond types; CA12bii alloys
5	Shapes of Molecules and Ions CA3b simple models and their shapes (models should be used where applicable: (i) linear: CO ₂ ; (ii) non linear: H ₂ O; (iii) tetrahedral: CH ₄ ; (iv) pyramidal: NH ₃); CC1 shapes of molecules (the shapes of the following molecules should also be treated: H ₂ and O ₂)
6	Intermolecular Forces - Van der Waal's forces CA3di van der Waal's forces (Description of formation and nature should be treated. Dipole-dipole and induced dipole forces should be treated under van der Waal's forces)
7	Intermolecular Forces - Hydrogen Bonding; Kinetic Model of Matter CA1biii atoms, molecules and ions (definition of particles and treatment of particles as building blocks of matter); CA3dii hydrogen bonding; CA5ai postulates of the kinetic model of matter
8	States of Matter – Solids CA5aii use of the kinetic model to explain the nature of solids; use of the kinetic model to explain the changes of state of matter (changes of state of matter should be explained in terms of movement of particles, it should be emphasized that randomness decreases (and orderliness increases) from liquid state to solid state, illustrations of changes of state using the different forms of water, iodine, sulphur, naphthalene etc); CA5di solids (regular arrangement of ions, molecules and atoms in three dimensions in the solid state should be emphasized; knowledge of specific packing arrangements not required), CA13aiii General Skills and Principles – melting point determination
9	States of Matter – Liquids CA5aii use of the kinetic model to explain the nature of liquids; use of the kinetic model to explain the changes of state of matter (changes of state of matter should be explained in terms of movement of particles, it should be emphasized that randomness decreases and orderliness increases from gaseous state to liquid state); CA5c liquids, concept of vapour pressure (Liquids as an intermediate state between gases and solids in the kinetic-molecular sense should be emphasized, simple methods for determination of boiling points, standard boiling point); ISA9.1 physical and chemical properties of water (experiment to determine/demonstrate the boiling point of water)
10	States of Matter - Gases and Solutions CA4c solutions (concept of solution as made up of solvent and solute); CA5aii use of the kinetic model to explain the nature of gases; use of the kinetic model to explain the changes of state of matter (it should be emphasized that randomness decreases and orderliness increases from gaseous state to liquid state); CA5di solids (melting points and indicator of purity of solids)

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Lesson	Title and Syllabus Reference
11	<p>Evidence for the Kinetic Model - Brownian Motion and Diffusion CA5aii The use of the kinetic model to explain the nature of solids, liquids and gases (Brownian motion to be illustrated using any of the following experiments: (i) Pollen grains/powdered sulphur in water, viewed under a microscope, (ii) Smoke in a glass container illuminated by a strong light from the side, (iii) A dusty room being swept and viewed from outside under sunlight); CA5aii the use of the kinetic model to explain diffusion (demonstration could be given using the following: (i) Diffusion of bromine/iodine/NO₂ from a sealed tube into an empty tube, (ii) Spread of scent of ammonia in a room); CA5bi the gases – the gas laws (Graham's law; qualitative explanation using the kinetic model)</p>
12	<p>The Gas Laws CA5bi the gases, the gas laws (Charles', Boyle's, Dalton's law; qualitative explanation of each of the gas laws using the kinetic model, derivation of the general gas law $PV/T = k$, mathematical relations of the gas laws and calculations based on the laws will be required); CC2 The Gas Laws (statement of Gay Lussac's law and calculations based on the law)</p>
13	<p>Chemical Structures – Giant Ionic CA3ai ionic bonding - properties of ionic compounds (typical properties of ionic compounds using binary compounds which are largely ionic. e.g. melting points, boiling point); CA4ai chemical symbols; CA5di solids – types and structures (ionic solids – properties); ISA4.3 ionic compounds (characteristic properties of ionic compounds)</p>
14	<p>Chemical Structures - Giant Metallic and Simple Molecular CA3cii properties of metals (typical properties including conductivity; malleability, ductility); CA4ai chemical symbols; CA4aii molecular formulae; CA5di solids – types and structures (metallic solids – properties)</p>
15	<p>Chemical Structures – Simple Molecular CA3aii covalent bonding – properties of covalent compounds (typical properties compared with those of ionic compounds e.g. melting point, boiling point); CA3di van der Waal's forces (relative physical properties of polar and non-polar compounds, variation of the melting points and boiling points of noble gases and halogens explained in terms of van der Waal's forces); CA3dii hydrogen bonding (variation in the boiling points of H₂O, H₂S, H₂Se, H₂Te explained using hydrogen bonding); CA5di solids – types and structures (molecular solids – properties); ISA4.3 covalent compounds (characteristic properties of covalent compounds)</p>
16	<p>Chemical Structures - Giant Covalent CA3aii covalent bonding – properties of covalent compounds (typical properties compared with those of ionic compounds e.g. melting point, boiling point); CA4ai chemical symbols; CA5di solids – types and structures (covalent solids – properties); CA5diii structures, properties and uses of diamond and graphite (properties and uses – dependent on structures); ISA4.3 covalent compounds (characteristic properties of covalent compounds)</p>
17	<p>Chemical Structures – Solubility CA3ai ionic bonding – properties of ionic compounds (typical properties of ionic compounds using binary compounds which are largely ionic. e.g. solubility in various solvents); CA3aii covalent bonding - properties of covalent compounds (typical properties compared with those of ionic compounds e.g. solubility in various solvents like water, hexane, ether); CA5di solids – types and structures (ionic, metallic, covalent and molecular solids - comparison of their properties); ISA4.3 ionic and covalent compounds (characteristic properties of ionic and covalent compounds)</p>
18	<p>Chemical Structures - Comparison of Properties CA3e comparison of all bond types; CA5di solids – types and structures (ionic, metallic, covalent and molecular solids - comparison of their properties); ISA4.3 ionic and covalent compounds (characteristic properties of ionic and covalent compounds)</p>
19	<p>Periodic Trends in Properties of Elements CA2a Periodicity of the elements: trends in periodic properties - down a group and across a period (metallic to non-metallic character of elements); ISA10.1 classification of materials (into metals, semi-metals (metalloids), and non-metals)</p>
20	<p>Periodic Trends in Properties of Compounds CA2a Periodicity of the elements: trends in periodic properties - down a group and across a period (ionic to covalent bonding in compounds)</p>
21	<p>Revision and Summary</p>

The Periodic Table of the Elements

	1	2	3	4	5	6	7	0											
	(1)	(2)	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	(13)	(14)	(15)	(16)	(17)	(18)	
	6.9 Li lithium 3	9.0 Be beryllium 4	23.0 Na sodium 11	39.1 K potassium 19	40.1 Ca calcium 20	44.9 Ti titanium 22	47.9 Ti titanium 22	54.9 Mn manganese 25	55.8 Fe iron 26	58.9 Co cobalt 27	58.7 Ni nickel 28	63.5 Cu copper 29	65.4 Zn zinc 30	10.8 B boron 5	12.0 C carbon 6	14.0 N nitrogen 7	16.0 O oxygen 8	19.0 F fluorine 9	4.0 He helium 2
	23.0 Na sodium 11	24.3 Mg magnesium 12	39.1 K potassium 19	40.1 Ca calcium 20	50.9 V vanadium 23	52.0 Cr chromium 24	54.9 Mn manganese 25	55.8 Fe iron 26	58.9 Co cobalt 27	58.7 Ni nickel 28	63.5 Cu copper 29	65.4 Zn zinc 30	69.7 Ga gallium 31	10.8 B boron 5	12.0 C carbon 6	14.0 N nitrogen 7	16.0 O oxygen 8	19.0 F fluorine 9	39.9 Ar argon 18
	85.5 Rb rubidium 37	87.6 Sr strontium 38	88.9 Y yttrium 39	91.2 Zr zirconium 40	92.9 Nb niobium 41	96.0 Mo molybdenum 42	98 Tc technetium	101.1 Ru ruthenium 44	102.9 Rh rhodium 45	106.4 Pd palladium 46	107.9 Ag silver 47	112.4 Cd cadmium 48	114.8 In indium 49	27.0 Al aluminium 13	28.1 Si silicon 14	31.0 P phosphorus 15	32.1 S sulfur 16	35.5 Cl chlorine 17	39.9 Ar argon 18
	132.9 Cs caesium 55	137.3 Ba barium 56	138.9 La * lanthanum 57	178.5 Hf hafnium 72	180.9 Ta tantalum 73	183.8 W tungsten 74	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	195.1 Pt platinum 78	197.0 Au gold 79	200.6 Hg mercury 80	204.4 Tl thallium 81	207.2 Pb lead 82	209.0 Bi bismuth 83	209 Po polonium 84	210 At astatine 85	222 Rn radon 86	
	[223] Fr francium 87	[226] Ra radium 88	[227] Ac † actinium 89	[267] Rf rutherfordium 104	[268] Db dubnium 105	[271] Sg seaborgium 106	[272] Bh bohrium 107	[270] Hs hassium 108	[276] Mt meitnerium 109	[281] Ds darmstadtium 110	[280] Rg roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated							
	140.1 Ce cerium 58	140.9 Pr praseodymium 59	144.2 Nd neodymium 60	145 Pm promethium 61	150.4 Sm samarium 62	152.0 Eu europium 63	157.3 Gd gadolinium 64	158.9 Tb terbium 65	162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68	168.9 Tm thulium 69	173.1 Yb ytterbium 70	175.0 Lu lutetium 71					
	232.0 Th thorium 90	231.0 Pa protactinium 91	238.0 U uranium 92	237 Np neptunium 93	244 Pu plutonium 94	243 Am americium 95	247 Cm curium 96	247 Bk berkelium 97	251 Cf californium 98	252 Es einsteinium 99	257 Fm fermium 100	258 Md mendelevium 101	259 No nobelium 102	262 Lr lawrencium 103					

* 58 – 71 Lanthanides

† 90 – 103 Actinides

Lesson 1 – What happens when atoms join together?

a) Elements, Compounds and Mixtures

- Atoms very rarely exist on their own; they are almost always found combined together with other atoms to form a very wide variety of chemical substances
- A **substance** is a collection of atoms with fixed composition and distinct chemical properties
- All objects in the universe are made of materials; a material can either be made from a single substance or from a mixture of different substances
- Materials can be classified in three different ways – as **elements**, **compounds** or **mixtures** - depending on their composition:

(i) Elements

- A material made up of only one type of atom is called an **element**; the name of the element is always the same as the name of the atom, and each atom therefore gives rise to one element (although the atoms can combine together in different ways, giving rise to different forms of the element called allotropes)
- the Periodic Table is a list of known atoms and therefore a list of known elements
- As elements only contain one type of atom, they have a fixed composition; each allotrope of the element is considered a single, pure substance
- Examples of known elements are hydrogen, oxygen and carbon

(ii) Compounds

- A substance made up of two or more different elements bonded together in a fixed proportion is called a **compound**; there are a large number of different compounds
- Compounds are pure substances with a specific chemical formula showing the relative composition of the different atoms in the compound
- Compounds can be broken down into their constituent elements, but not easily – usually a chemical reaction or electrolytic process is required
- Examples of compounds include sodium chloride (NaCl), water (H₂O), methane (CH₄), ethane (C₂H₆), silicon dioxide (SiO₂)

(iii) Mixtures

- A material made up of two or more elements or compounds but not in fixed proportion is called a **mixture**
- Mixtures are not pure substances, have no fixed composition and have no single chemical formula; mixtures consists of at least two different substances
- Mixtures can take almost any form; sea water, air, milk, crude oil and most types of earth are all examples of mixtures

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- In most cases the different components of the mixture are held together only by intermolecular forces
- Mixtures can be separated into the pure single substances from which they are made (elements and/or compounds) using simple physical separation techniques such as filtration, distillation (simple or fractional), evaporation or chromatography



Test your knowledge 1.1: Classifying materials

Classify the following materials as elements, compounds or mixtures; if they are compounds; give a chemical formula for the compound to show its chemical composition:

- | | | | |
|-------------|--------------------|------------|---------------|
| (a) bromine | (b) ammonia | (c) milk | (d) lemonade |
| (e) water | (f) carbon dioxide | (g) bronze | (h) crude oil |
| (i) diamond | (j) sand | (k) lead | (l) methane |

Answers: (a) element; (b) compound (NH_3); (c) mixture; (d) mixture; (e) compound (H_2O); (f) compound (CO_2); (g) mixture; (h) mixture; (i) element (an allotrope of carbon); (j) compound (SiO_2); (k) element; (l) compound (CH_4)

b) Laws of Chemical Combination

- Compounds have a fixed formula; this means that compounds obey the **Law of Constant Composition**: “a given chemical compound always contains its component elements in fixed ratio (by mass), and this fixed ratio does not depend on its source and method of preparation”
- Copper oxide can be used to illustrate the law of constant composition:
 - copper oxide contains 80% copper and 20% oxygen by mass;
 - there are three common methods for preparing copper oxide but all three methods will give copper oxide containing 80% copper and 20% oxygen
 - samples of copper oxide prepared by each of the three methods can be reduced to copper by passing hydrogen gas over the copper oxide
 - in all three cases, the mass of the solid will drop by exactly 20% as the oxygen is removed
- In many cases, different elements can combine to form more than one compound; when different elements combine they will obey the **Law of Multiple Proportions**: “the ratios of the masses of the second element which combine with a fixed mass of the first element will be ratios of small whole numbers”
- Carbon and oxygen can be used to illustrate the Law of Multiple Proportions:
 - when 1.2 g of carbon combines with oxygen, it will either combine with 1.6 g of oxygen (to form carbon monoxide, CO), or with 3.2 g of oxygen (to form carbon dioxide, CO_2)
 - the ratio of the two masses of oxygen = $3.2:1.6 = 2:1$



Test your knowledge 1.2 – Demonstrating the laws of chemical combination

- 10 g of methane was found to contain 7.5 g of carbon and 2.5 g of hydrogen; 40 g of methane was found to contain 30 g of carbon and 10 g of hydrogen; show that this is consistent with the law of constant composition
- In one experiment, 50 g of tin reacted with 6.8 g of oxygen; in another experiment, 50 g of tin reacted with 13.6 g of oxygen; show that this is consistent with the law of multiple proportions

- (a) $7.5/2.5 = 3$; $30/10 = 3$ so consistent (ie composition is always the same)
(b) $13.6/6.8 = 2$ (a small whole number ratio, so consistent)

Lesson 2 – What is ionic bonding?

c) Simple Chemical Bonding

**Summary Activity 2.1: Which atoms attract electrons most strongly?**

- As you cross a Period from left to right, how does the tendency of atoms to attract electrons change? Why?
- As you descend a Group, how does the tendency of atoms to attract electrons change? Why?
- As you cross a Period proton number increases but shielding stays the same, so electrons in the outer shell are more strongly attracted to the nucleus
- As you descend a Group the number of shells increases so shielding increases, so electrons in the outer shell are less strongly attracted to the nucleus

(i) Introduction to chemical bonding

- Atoms can form bonds with each other by gaining, losing or sharing electrons, depending on the properties of the atoms involved in the bond; this can lead to a number of different types of bond
- The number of electrons which an atom gains, loses or shares when it bonds with other atoms is called the **valency** of the atom; generally, atoms in the same group tend to have the same valency

(ii) Electronegativity

- Atoms on the right-hand side (especially at the top) of the Periodic Table attract other electrons more strongly than atoms on the left-hand side of the Periodic Table, due to their higher electron affinities; they also hold on to their own electrons more strongly than atoms on the left-hand side of the Periodic Table, due to their higher ionisation energies; such atoms are said to be **electronegative** atoms; fluorine is the most electronegative atom, followed by oxygen; nitrogen and chlorine are also electronegative; all atoms in Groups 5, 6 and 7 are considered fairly electronegative, especially near the top of the Periodic Table
- Atoms which hold their own electrons weakly (ie have low ionisation energies) and attract other electrons weakly (ie have low electron affinities) are said to be **electropositive** atoms; francium is the most electropositive atom, but all atoms in Group 1 are very electropositive; atoms in Group 2 are also electropositive, as are some atoms in Group 3
- The type of bond atoms will form with each other depends on how electropositive or electronegative the atoms are; **electronegativity is the ability of an atom to attract electrons in a bond**
- Electronegativity increases across a Period; electronegativity decreases down a Group

(iii) Ionic Bonding

- If an electropositive atom (like Na) and an electronegative atom (like Cl) come into contact with each other, the electronegative atom will take electrons from the electropositive atom until either the outer shell of the electronegative atom is full or the outer shell of the electropositive atom is empty
- Electropositive atoms lose electrons to become cations, and electronegative atoms gain electrons to become anions

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- Atoms in Groups I, II and III of the Periodic Table tend to form positively charged ions (cations) by losing all of the electrons in their outer shell; Na loses one electron to form Na^+ , Mg loses two electrons to form Mg^{2+} and Al loses three electrons to form Al^{3+} ; sodium therefore has a valency of 1, magnesium has a valency of 2 and aluminium has a valency of 3
- Simple positive ions have the same name as the atom they were formed from:
 - Na^+ is a sodium ion
 - Mg^{2+} is a magnesium ion
- Atoms in Groups V, VI and VII of the Periodic Table tend to form negatively charged ions (anions) by gaining electrons until their outer shell is full; Cl gains one electron to form Cl^- , O gains two electrons to form O^{2-} and N gains three electrons to form N^{3-} ; chlorine therefore has a valency of 1, oxygen has a valency of 2 and nitrogen has a valency of 3
- Simple negative ions are named by combining the first syllable of the atom they were formed from and then adding the ending “ide”:
 - Cl^- is a chloride ion
 - O^{2-} is an oxide ion

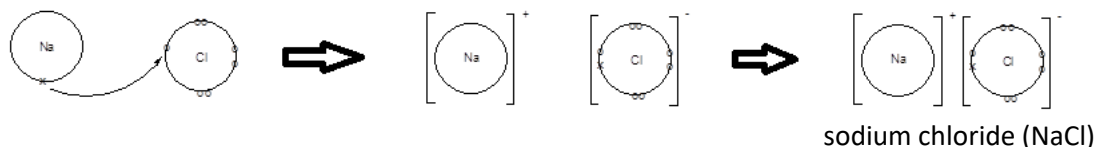


Test your knowledge 2.2: Forming simple ions

- Consider the atoms Li, K, Mg, Ca and Al; give the formula and name of the ion formed when these atoms come into contact with an electronegative atom
- Consider the atoms O, N, S, Br and P; give the formula and name of the ion formed when these atoms come into contact with an electropositive atom

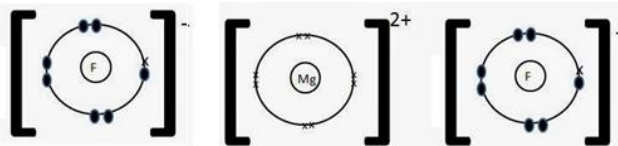
- (a) lithium ion Li^+ , potassium ion K^+ , magnesium ion Mg^{2+} , calcium ion Ca^{2+} , aluminium ion Al^{3+}
(b) oxide ion O^{2-} , nitride ion N^{3-} , sulphide ion S^{2-} , bromide ion Br^- , phosphide ion P^{3-}

- Positive and negative ions have opposite charges, so they attract each other; this attraction is called an ionic bond; **an ionic bond is an attraction between oppositely charged ions**
- Ionic bonding is most likely to exist when one atom has a high ionisation energy and electron affinity (and hence a high electronegativity) and one atom has a low ionisation energy and electron affinity (and hence a low electronegativity); there is always a significant electronegativity difference between the two atoms
- Ionic bonds are formed when electrons from an electropositive atom are transferred to an electronegative atom and the resulting ions attract each other
 - when sodium atoms and chlorine atoms interact, each sodium atom transfers an electron to a chlorine atom to form a sodium ion and a chloride anion
 - these two ions attract each other to form an ionic bond; the resulting substance is known as sodium chloride (NaCl); it is an example of an **ionic compound**



UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- The representation of ionic bonds as shown above, showing the outer electrons and charges on the ions, is an example of a **Lewis dot structure** (or dot-cross diagram)
 - When magnesium and fluorine atoms interact, the Mg loses two electrons and gives one electron each to two different fluorine atoms
 - the fluorine atoms accept the electrons and become fluoride ions
 - the oppositely charged ions attract each other to form a substance called magnesium fluoride (MgF_2):



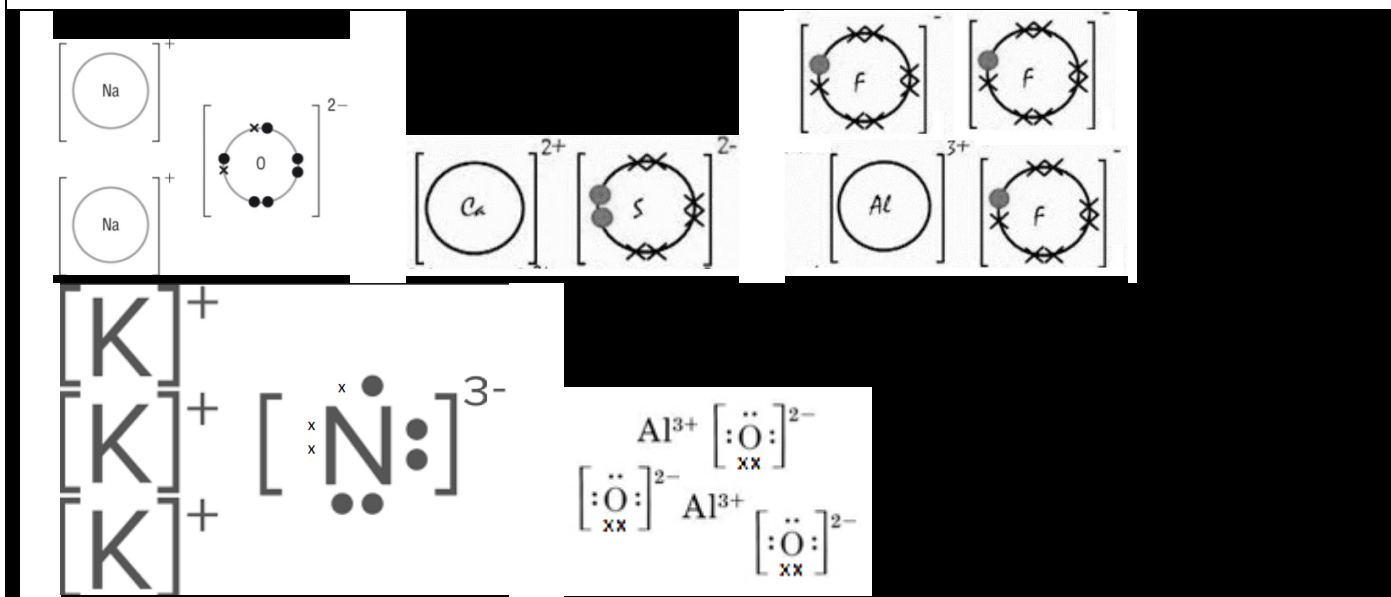
- Ionic compounds are named by naming the cation followed by the anion



Test your knowledge 2.3: Forming simple ionic compounds

Deduce the formula, and draw the Lewis dot structure, of the following ionic compounds:

- sodium oxide
- calcium sulphide
- aluminium fluoride
- potassium nitride
- aluminium oxide



Lesson 3 - What is covalent bonding?

(iv) Covalent Bonding

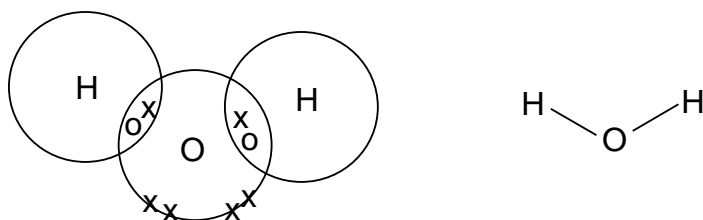
- If two electronegative atoms come into contact with each other, both will attract each other's electrons but neither will lose their electrons; as a result they share electrons; the shared electrons are attracted to the nuclei of both atoms at the same time; this is known as a covalent bond; **a covalent bond is a pair of electrons shared between two atoms**
- Covalent bonding generally exists when both atoms have relatively high ionization energies and electron affinities, and hence high (and often similar) electronegativities)

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

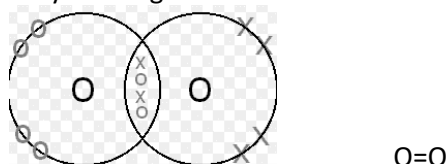
- A covalent bond is formed when two atomic orbitals on different atoms overlap:



- in most cases, each orbital contains one electron; the resulting covalent bond therefore contains one electron from each atom
 - in some cases, one of the orbitals contains two electrons and the other contains no electrons; both of the electrons in the resulting covalent bond therefore come from the same atom; this is known as a **dative covalent bond**
- Atoms will generally accept as many shared electrons as they can accommodate in their outer shell:
 - hydrogen can only accommodate one extra electron and so each hydrogen atom can only form one covalent bond (ie H has a valency of 1)
 - atoms in Group 7 (eg Cl) can only accommodate one extra electron and so each atom can only form one covalent bond (ie atoms in group 7 have a valency of 1)
 - oxygen (and usually S) can accommodate two extra electrons and so each atom will form two covalent bonds (O and usually S have a valency of 2)
 - nitrogen (and usually P) can accommodate three extra electrons so each atom will form three covalent bonds (N and usually P have a valency of 3)
 - carbon and silicon can accommodate four extra electrons so each atom will form four covalent bonds (C and Si have a valency of 4)
- A covalent bond is represented by a short straight line between the two atoms; it can also be represented in a Lewis-dot structure by showing dots and crosses inside overlapping circles:
 - Eg two hydrogen atoms each form one covalent bonds with an oxygen atom, which forms one covalent bond with both hydrogen atoms:



- It is possible for two atoms to share more than one pair of electrons; the sharing of two pairs of electrons between two atoms is called a **double bond**; the sharing of three pairs of electrons between two atoms is called a **triple bond**:
 - Eg two oxygen atoms can join by forming a double covalent bond with each other:



- A small group of two or more atoms held together by covalent bonds is called a **molecule**; covalent bonding usually results in the formation of molecules
- These molecules act as single particles and there are no covalent bonds holding the different molecules together; there are, however, weaker forces of attraction between these molecules which are known as **intermolecular forces**

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

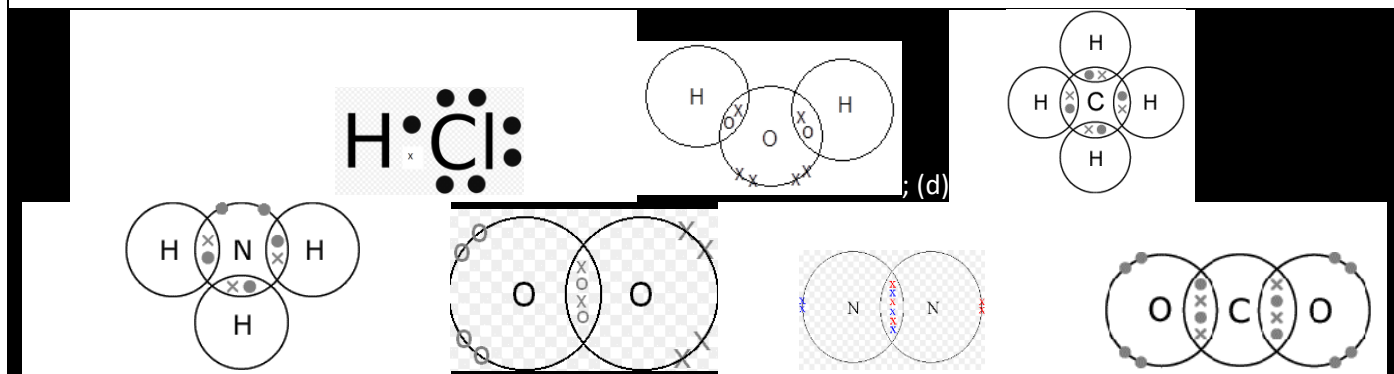
- Individual molecules can be elements (eg Cl₂) or compounds (eg HCl); many materials are mixtures of different molecules (eg crude oil, air)



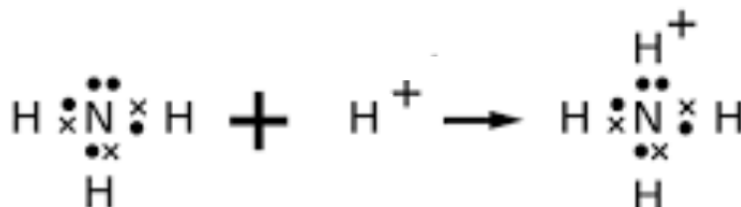
Test your knowledge 3.1: Forming simple molecules

Deduce the formula and draw the structure of the molecules formed when the following atoms combine to form covalent bonds:

(a) H and H; (b) H and Cl; (c) H and O; (d) C and H; (e) N and H; (f) O and O; (g) N and N; (h) C and O



- It is possible for atoms and ions to form covalent bonds with each other; the resulting particles, which are charged, are called **polyatomic ions**; they are ions containing covalent bonds: eg NH₄⁺, SO₄²⁻, CO₃²⁻, OH⁻
- Dative covalent bonds** are represented by a short arrow from the electron providing both electrons to the electron providing neither, eg N→H⁺, or N→B
- Polyatomic ions often result from dative covalent bonds:
 - Eg NH₄⁺, is formed when a molecule of NH₃, which contains two electrons in a single orbital on the N atom, forms a dative covalent bond with a hydrogen ion, which has no electrons:



The NH₄⁺ ion can be drawn as follows:

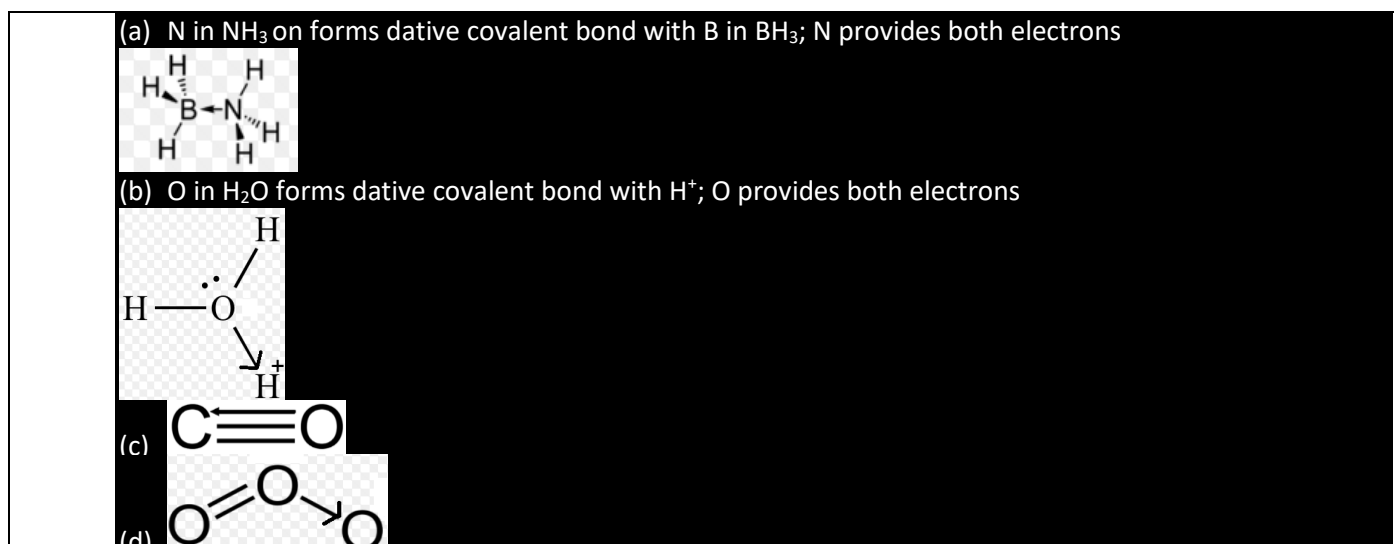
- Dative bonding can also result in unusual molecules such as CO and O₃ which could not be formed using normal covalent bonds



Test your knowledge 3.2: Understanding dative covalent bonding

- Explain how BH₃ forms a covalent bond with NH₃ to form BH₃NH₃; draw the structure of BH₃NH₃
- Explain how H₂O forms a covalent bond with H⁺ to form H₃O⁺; draw the structure of H₃O⁺
- Use dative covalent bonding to draw the structure of carbon monoxide (CO)
- Use dative covalent bonding to draw the structure of ozone (O₃)

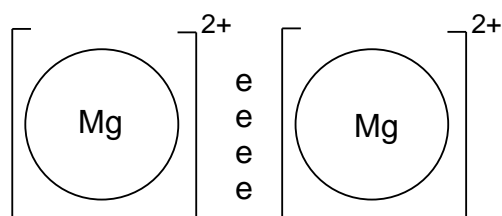
UNIT 2 – PARTICLES, BONDING AND STRUCTURES



Lesson 4 - What is a metallic bond?

(v) Metallic bonding

- If two electropositive atoms come into contact with each other, both atoms will give up their electrons but neither will accept electrons from the other; as a result, both atoms become cations and the electrons become free, or “delocalised”; a large number of delocalised electrons is also known as a “sea of electrons”; **a metallic bond is an attraction between cations and a sea of electrons**
- Metallic bonds are formed when two or more atoms all lose electrons and the resulting electrons are attracted to all the resulting cations:
 - Magnesium atoms lose two electrons each, and the resulting electrons are attracted to all the cations

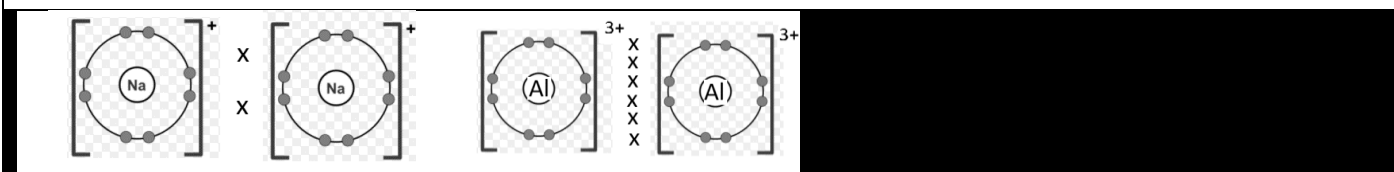


- sodium atoms lose one electron each
- aluminium atoms lose three electrons each
- Metallic bonding happens because the delocalised electrons are attracted to more than one nucleus and hence more stable
- Metallic bonding is most likely to be found when atoms have a relatively large size and a low ionization energy; for this reason they tend to involve atoms on the left of the Periodic Table, which have relatively few valence electrons (1, 2 or 3)
- Metallic bonding between atoms of the same type results in the formation of elements; metallic bonding between atoms of different types results in the formation of mixtures known as **alloys**; alloys are not compounds as they do not have a fixed composition


Test your knowledge 4.1: Forming metallic bonds

Draw diagrams to show the bonding between:

- Two sodium atoms
- Two aluminium atoms

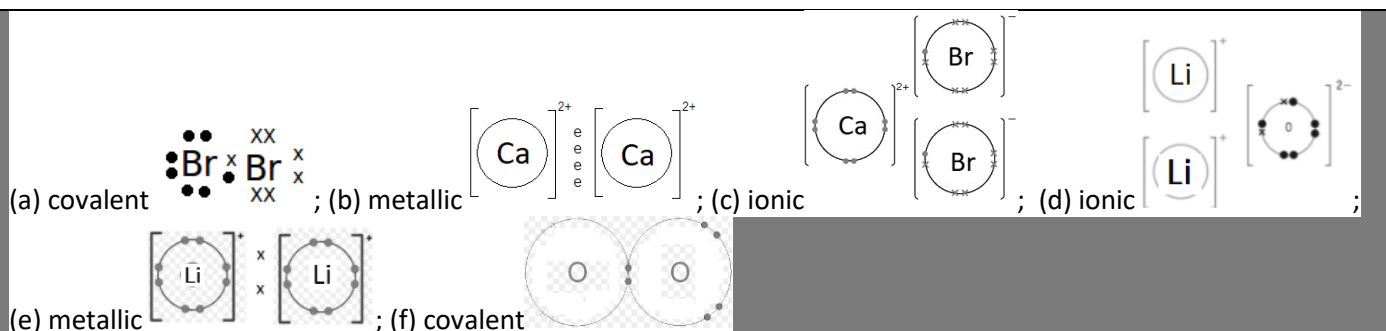

(vi) Summary of different types of bond

- Ionic bonds are usually formed when an electropositive atom bonds with an electronegative atom – ie when there is a large electronegativity difference between the two atoms; the electropositive atom gives electrons to the electronegative atom and the resulting ions attract
- Covalent bonds are usually formed when two electronegative atoms bond together by sharing electrons; they usually result in the formation of small particles called molecules
- Metallic bonds are usually formed when two electropositive atoms bond together by donating their electrons into a delocalised electron sea; the resulting cations are attracted to the electron sea
- Covalent bonds can also form between atoms and ions to form polyatomic ions (eg NH_4^+ , SO_4^{2-}); these polyatomic ions will form bonds with other ions
- Molecules do not form strong bonds with other molecules but can form weak bonds called intermolecular forces

Test your knowledge 4.2: Understanding ionic, covalent and metallic bonding

Predict the type of bond which will be formed between the following pairs of atoms and draw the Lewis dot structure to show each bond:

- Br and Br
- Ca and Ca
- Ca and Br
- Li and O
- Li and Li
- O and O



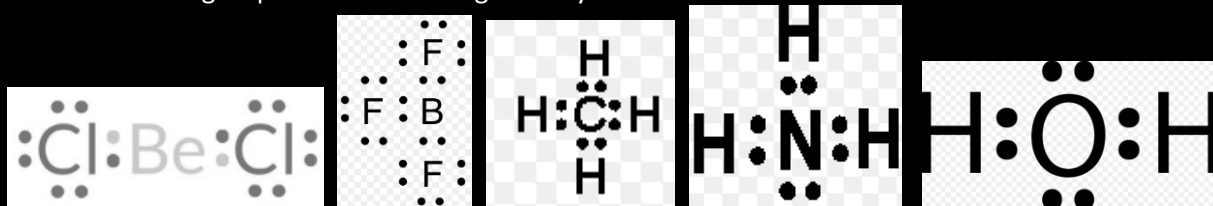
Lesson 5 – What shapes do different molecules have?

d) Shapes of Molecules and Polyatomic Ions

**Summary Activity 5.1: Lewis dot structures for different molecules**

- Explain the meaning of the terms “covalent bond” and “molecule”
- Draw Lewis-dot structures for the following molecules: BeCl_2 , BF_3 , CH_4 , NH_3 and H_2O

- Covalent bond: pair of electrons shared between two atoms
- Molecule: small group of atoms held together by covalent bonds



- When an atom forms more than one covalent bond, the two covalent bonds arrange themselves at a very specific angle; the angles between the covalent bonds in a molecule or polyatomic ion result in the species having a characteristic shape
- When an atom forms a covalent bond with two other atoms at the same time, the electrons in the different bonds and the non-bonding electrons in the outer shells of the central atom repel each other; in order to minimise this repulsion, all the outer shell electrons spread out as far apart in space as possible; the resulting shapes, and the angles between the covalent bonds, can be predicted by the VSEPR theory (VSEPR = valence shell electron pair repulsion)
- VSEPR theory consists of two basic rules:
 - All bonded electron pairs and all lone pairs arrange themselves as far apart in space as is possible; double bonds are counted as one pair of electrons for the purpose of determining shapes
 - Lone pairs repel more strongly than bonding pairs and result in the angle between the bonds becoming slightly smaller
- These two rules can be used to predict the shape of any covalent molecule or ion, and the angles between the bonds.

**Activity 5.2: Investigate the shapes balloons form when tied together**

Take two balloons, blow them up and tie them together. What angle do they naturally adopt between them? Now repeat with three and four balloons. What angle and shape will they naturally adopt?

Equipment needed: 4 balloons

Two inflated balloons should naturally sit 180° apart when tied; three inflated balloons 120° apart and four inflated balloons 109.5° apart, in the shape of a tetrahedron when tied

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

(i) two electron pairs

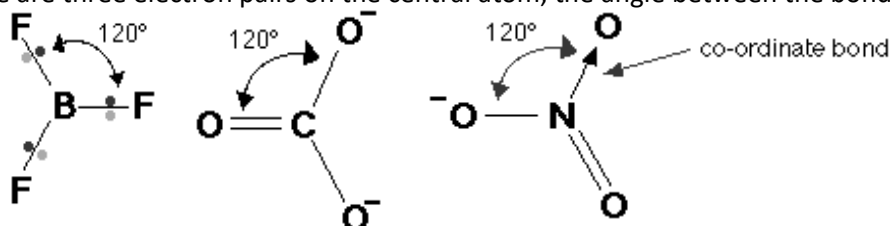
- If there are two electron pairs on the central atom, the angle between the bonds is 180°



- Molecules which adopt this shape are said to be LINEAR
- E.g. BeCl_2 , CO_2 (remember that double bonds count as one electron pair)

(ii) three electron pairs

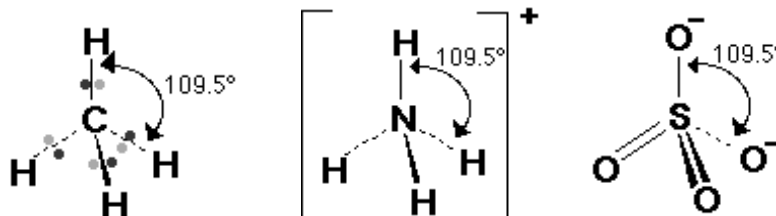
- If there are three electron pairs on the central atom, the angle between the bonds is 120°



- Molecules which adopt this shape are said to be TRIGONAL PLANAR
- E.g. BF_3 , AlCl_3 , CO_3^{2-} , NO_3^-

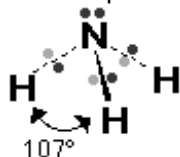
(iii) four electron pairs

- If there are four bonded pairs on the central atom, the angle between the bonds is 109.5°



- Molecules which adopt this shape are said to be TETRAHEDRAL
- E.g. CH_4 , SiCl_4 , NH_4^+ , SO_4^{2-}

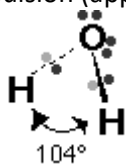
- If one of the electron pairs is a lone pair, the bond angle is slightly less than 109° , due to the extra lone pair repulsion which pushes the bonds closer together (approx 107°)



- Molecules which adopt this shape are said to be PYRAMIDAL
- E.g. NH_3 , PCl_3

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- If two of the electron pairs are lone pairs, the bond angle is also slightly less than 109° , due to the extra lone pair repulsion (approx 104°)



- Molecules which adopt this shape are said to be NON-LINEAR
- E.g. H_2O , OF_2

(iv) Diatomic molecules

- Molecules such as H_2 and O_2 , which only contain two atoms and one bond, are generally described as **linear**

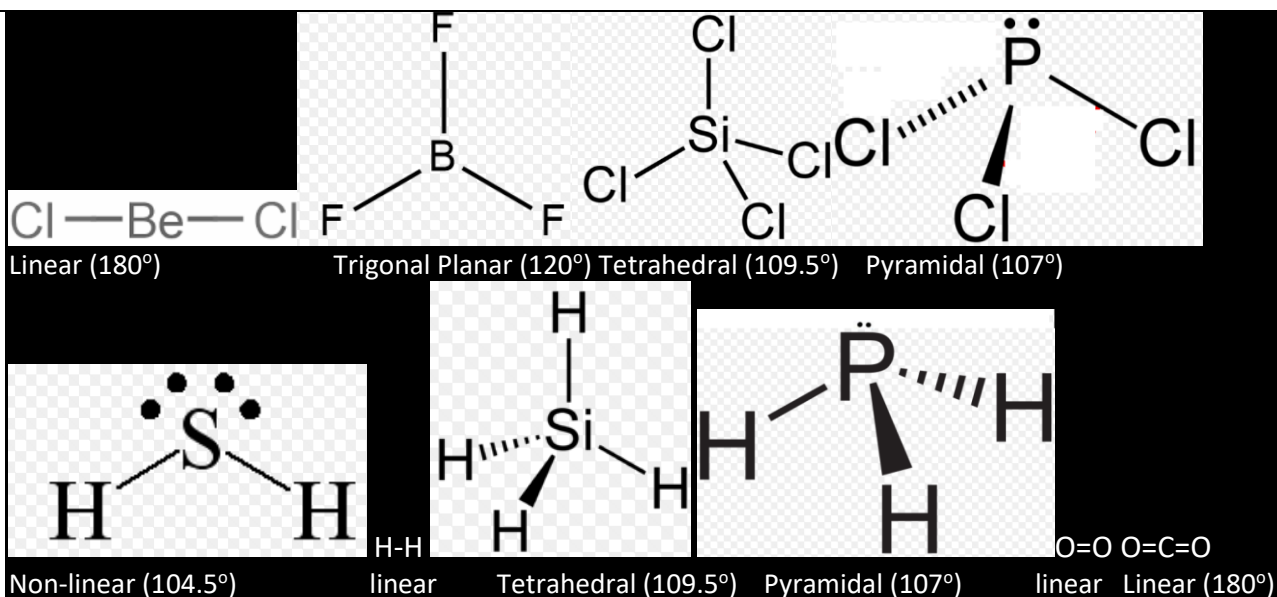
(v) Summary of molecular shapes

Valence shell electron pairs around central atom	Bonding Pairs around central atom	Lone Pairs around central atom	Shape	Bond Angle ($^\circ$)
2	2	0	LINEAR 	180
3	3	0	TRIGONAL PLANAR 	120
4	4	0	TETRAHEDRAL 	109.5
4	3	1	PYRAMIDAL 	107
4	2	2	NON-LINEAR 	104.5


Test your knowledge 5.3: Understanding Molecular Shapes

Draw and name the shapes of following molecules, indicating the bond angles:

- (a)
- BeCl_2
- (b)
- BF_3
- (c)
- SiCl_4
- (d)
- PCl_3
- (e)
- H_2S
- (f)
- H_2
- (g)
- SiH_4
- (h)
- PH_3
- (i)
- O_2
- (j)
- CO_2


Extension 5.4: Understanding molecular shapes

- (k) Completed Test Your Knowledge 5.3? Think of any other molecules you know and try to work out their shape and the angle between the bonds

Free choice question so no answers available

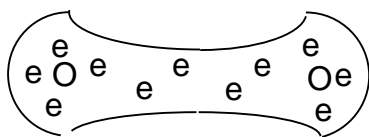
Lesson 6 – What are Intermolecular Forces?
e) Intermolecular forces

- When atoms form covalent bonds with other atoms to form molecules, these molecules act as discrete particles and there are no strong bonds (covalent, ionic or metallic) holding the different molecules together; there are, however, weaker forces of attraction between these molecules and these forces are known as **intermolecular forces**; there are two main types of intermolecular force:

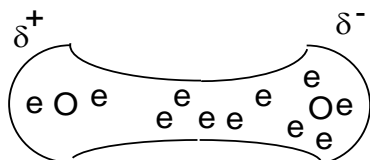
UNIT 2 – PARTICLES, BONDING AND STRUCTURES

(i) Van der Waal's forces

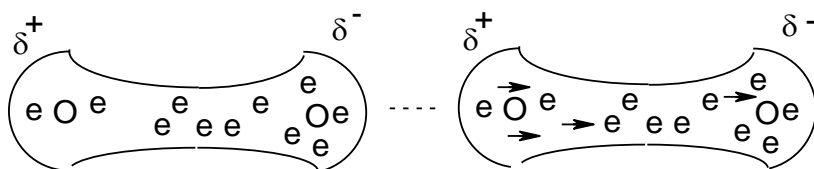
- Consider a molecule of oxygen, O₂:



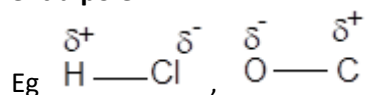
- the electrons in this molecule are not static; they are in a state of constant motion; it is therefore likely that at any given time the distribution of electrons will not be exactly symmetrical - there is likely to be a slight surplus of electrons on one of the atoms



- this is known as a **temporary dipole**; it lasts for a very short time as the electrons are constantly moving; temporary dipoles are constantly appearing and disappearing
- Consider now an adjacent molecule; the electrons on this molecule are repelled by the negative part of the dipole and attracted to the positive part and move accordingly:

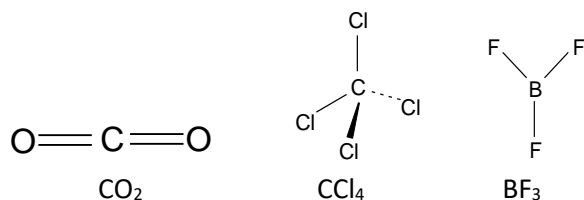


- this is known as an **induced dipole**
- there is a resulting attraction between the two molecules, and this known as a **Van der Waal's force**; Van der Waal's forces exist between all molecules
- In most covalent bonds there is a difference in electronegativity between the atoms; this causes an uneven distribution of charge and the covalent bond is said to be **polar**; the separation of charges is known as a **permanent dipole**

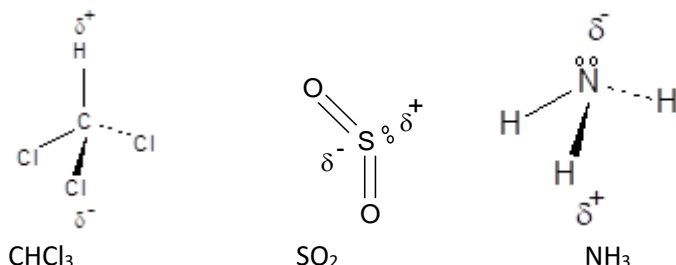


UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- In many cases, however, the presence of polar bonds (dipoles) does not result in the overall molecule having a permanent dipole on the molecule, as there are other polar bonds (dipoles) in the same molecule which have the effect of cancelling each other out; this effect can be seen in a number of linear, planar and tetrahedral substances:



- in all the above cases, there are dipoles resulting from polar bonds but the vector sum of these dipoles is zero; i.e. the dipoles cancel each other out
 - the molecule thus has no overall permanent dipole and is said to be **non-polar**; the Van der Waal's forces in these molecules will result from temporary dipoles and induced dipoles only
 - in other words, although there are permanent dipoles on the individual covalent bonds, there is no overall permanent dipole on the molecule because the individual permanent dipoles cancel each other out
- Some molecules, however, contain covalent bonds with dipoles which do not cancel each other out:



- In all the above cases, there are dipoles resulting from polar bonds whose vector sum is not zero; i.e. the dipoles do not cancel each other out; the molecule thus has a **permanent dipole** and is said to be **polar**
- The attraction between two molecules with permanent dipoles is another type of Van der Waal's force; molecules with permanent dipoles attract each other slightly more strongly than similar molecules which only have temporary and induced dipoles



Test your knowledge 6.1: Distinguishing between temporary and permanent dipoles

Using your molecular shapes from Exercise 5.3, deduce which of the molecules will have a permanent dipole:

(a) BeCl_2 (b) BF_3 (c) SiCl_4 (d) PCl_3 (e) H_2S (f) H_2 (g) SiH_4 (h) PH_3 (i) O_2 (j) CO_2 (k) HCl

- (a) BeCl_2 no because dipoles on bonds cancel
 (b) BF_3 no because dipoles on bonds cancel
 (c) SiCl_4 no because dipoles on bonds cancel
 (d) PCl_3 yes because dipoles on bonds don't cancel
 (e) H_2S yes because dipoles on bonds don't cancel
 (f) H_2 no because the bond is not polar
 (g) SiH_4 no because the dipoles on bonds cancel
 (h) PH_3 yes because the dipoles on bonds don't cancel
 (i) O_2 no because the bond is not polar
 (j) CO_2 no because dipoles on bonds cancel
 (k) HCl yes because there is only one bond and it is polar



Extension 6.2: Distinguishing between temporary and permanent dipoles

Completed Test Your Knowledge 6.1? Think of any other molecules you know and try to work out whether or not they have a permanent dipole

Free choice question so no answers available

- Van der Waal's forces are present between all molecules, although they can be very weak and are often not strong enough to hold the molecules together
- The strength of the Van der Waal's forces in between molecules depends on two factors:
 - **the number of electrons in the molecules**
The greater the number of electrons in a molecule, the greater the likelihood of a distortion and thus the greater the frequency and magnitude of the temporary dipoles; thus the Van der Waal's forces between the molecules are stronger
Example: the intermolecular forces between O₂ molecules (16 electrons) are stronger than those between N₂ molecules (14 electrons)
 - **the surface area of the molecules**
The larger the surface area of a molecule, the more contact it will have with adjacent molecules; thus the greater its ability to induce a dipole in an adjacent molecule and the greater the Van der Waal's forces.
Example: the intermolecular forces between CH₄ molecules (10 electrons) are stronger than those between F₂ molecules (18 electrons)
 - **the existence of permanent dipoles on the molecules**
If the molecule has a permanent dipole, there is a slightly stronger attraction between the molecules than there would have been if there were only temporary and induced dipoles; thus the Van der Waal's forces between the molecules are stronger
Example: the intermolecular forces between HCl molecules (18 electrons) are stronger than those between F₂ molecules (18 electrons)



Test your knowledge 6.3: Understanding factors affecting the strength of Van der Waal's forces

State, with a reason, which of the substances below will have stronger Van der Waal's forces between the molecules:

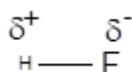
- (a) Ar and Ne
- (b) CH₄ and C₂H₆
- (c) Cl₂ and Br₂
- (d) CH₃CH₂CH₂CH₃ and CH₃CH(CH₃)CH₃
- (e) CO₂ and SO₂

- (a) Ar, because there are more electrons in each atom, and so stronger Van der Waal's forces
- (b) C₂H₆, because there are more electrons in each molecule, so stronger Van der Waal's forces
- (c) Br₂, because there are more electrons in each molecule, so stronger Van der Waal's forces
- (d) CH₃CH₂CH₂CH₃, because the surface area of the molecule is larger, so stronger Van der Waal's forces
- (e) SO₂, because there are more electrons in each molecule, so stronger Van der Waal's forces, and because it has a permanent dipole

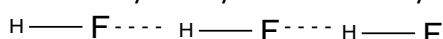
Lesson 7 – What is hydrogen bonding?

(ii) Hydrogen Bonding

- In most cases, the presence of permanent dipoles only makes a slight difference to the magnitude of the intermolecular forces; there is one exceptional case, however, where the permanent dipole makes a huge difference to the strength of the bonding between the molecules:
 - consider a molecule of hydrogen fluoride, HF; this molecule clearly has a permanent dipole as there is a large difference in electronegativity between H (2.1) and F (4.0); the electrons in this bond are on average much closer to the F than the H:

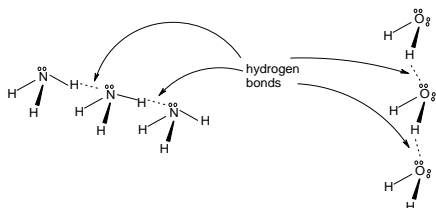


- as a result, the H atom has on almost no electron density around its nucleus at all and is therefore very small; this H atom is therefore able to approach the lone pairs of electrons on electronegative atoms on adjacent molecules very closely and form a very strong intermolecular dipole-dipole bond:



- this is known a **hydrogen bond**; it can only exist when a hydrogen atom is bonded to a very electronegative element; i.e. N, O or F; a **hydrogen bond is an attraction between an electropositive hydrogen atom (ie covalently bonded to N, O or F) and the lone pair of an electronegative atom on an adjacent molecule**

- Examples of molecules which are held together by hydrogen bonds are HF, H₂O and NH₃; hydrogen bonding also exists between many organic molecules such as alcohols, carboxylic acids, amines and amino acids



- Molecules which can form hydrogen bonds with each other tend to attract each other more strongly than molecules of similar size which cannot form hydrogen bonds with each other; the intermolecular forces between molecules of H₂O and NH₃, for example, are much stronger than those between larger molecules such as C₄H₁₀ or Cl₂, even though these larger molecules have stronger Van der Waal's forces


Test your knowledge 7.1: Recognising hydrogen bonding

- (a) State, with a reason, which of the substances below can form hydrogen bonds:
 (i) BeH₂; (ii) BH₃; (iii) CH₄; (iv) PH₃; (v) H₂O; (vi) H₂; (vii) HCl; (viii) NH₃; (ix) HF; (x) CO₂; (xi) H₂S
- (b) Explain why the intermolecular forces between water molecules are stronger than the intermolecular forces between carbon dioxide molecules

- (a) (i) BeH₂ no - no electropositive H; (ii) BH₃ no - no electropositive H; (iii) CH₄ no - no sufficiently electropositive H; (iv) PH₃ no - no sufficiently electropositive H; (v) H₂O yes - electropositive H because bonded to O; (vi) H₂ no - no electropositive H; (vii) HCl no - no sufficiently electropositive H; (viii) NH₃ yes - electropositive H because bonded to N; (ix) HF; yes - electropositive H because bonded to F; (x) CO₂ no - no H; (xi) H₂S no - no sufficiently electropositive H
- (b) Water molecules can form strong hydrogen bonds with other water molecules; but CO₂ cannot; CO₂ will have stronger Van der Waal's forces between molecules but the effect of hydrogen bonding is more significant

f) The Kinetic Model of Matter

(i) Particles and forces between them

- All matter is made from particles; Dalton believed that these particles were all indivisible atoms, but since then, we have a better understanding of what particles are and how they are held together; there are three types of particle:
 - **atoms**
 - **molecules** (small groups of atoms held together by covalent bonds)
 - **ions** (atoms or molecules with a charge)
- The way in which individual atoms, ions and molecules (particles) are attracted to each other is called **bonding**:
 - oppositely charged ions are held together by **ionic** bonding
 - some electropositive atoms are held together by **metallic** bonding
 - most other atoms are held together by **covalent** bonding
 - molecules (and atoms in Group 0) are held together by **intermolecular** forces (Van der Waal's forces and/or hydrogen bonds)
- All matter is made from particles (atoms, molecules or ions) joined together by bonds; particles can thus be described as the building blocks of matter
- In between the particles, there is empty space
- All particles have kinetic energy, which means that all particles are always moving; this idea is known as the **kinetic model of matter**

(ii) The meaning of temperature

- The amount of kinetic energy particles have depends on the temperature; the temperature is a direct measure of the average kinetic energy of the particles; the higher the temperature, the more energy the particles have and the faster they are moving
- Temperature is most commonly measured using the **Celsius scale (°C)**:
 - the temperature of the air on a normal day in West Africa is around 35 °C
 - the temperature of the water in the sea around West Africa is around 20 °C
 - at 0 °C, water freezes and turns to ice
 - at 100 °C, water boils and turns into steam
- It is possible to cool materials down to below 0 °C; as you keep cooling materials down, their kinetic energy decreases and they move more and more slowly; eventually, if you keep cooling, the particles will have no kinetic energy at all and will be completely still; this happens at a temperature of -273 °C and this temperature is known as **absolute zero**; this is the lowest temperature possible; it means zero kinetic energy
- For this reason, scientists tend to use a different scale for measuring temperature, called the **Kelvin scale (K)**; the Kelvin scale starts at absolute zero (0 K) and the unit is the same as the Celsius unit, so 0 °C is 273 K and 100 °C is 373 K; the two scales can be interconverted as follows:

$$T (\text{K}) = T (^\circ\text{C}) + 273$$

$$T (^\circ\text{C}) = T (\text{K}) - 273$$



Test your knowledge 7.2: Using the Kelvin temperature scale

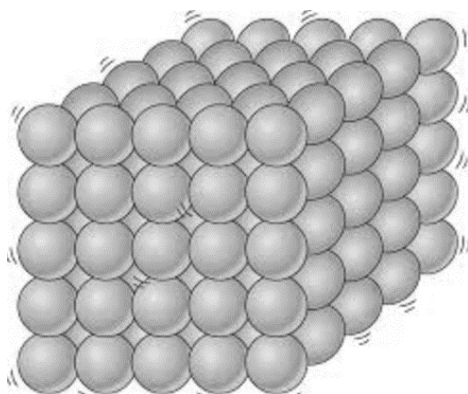
- (a) Convert the following temperatures from °C into K:
 (i) 25 °C (ii) 35 °C (iii) 300 °C (iv) -50 °C (v) -200 °C
- (b) Convert the following temperatures from K into °C:
 (i) 50 K (ii) 293 K (iii) 200 K (iv) 308 K (v) 600 K

- (a) (i) 298 K; (ii) 308 K; (iii) 573 K; (iv) 223 K, (v) 73 K
 (b) (i) -223 °C; (ii) 20 °C; (iii) -73 °C; (iv) 35 °C; (v) 327 °C

Lesson 8 – What are solids and what are the properties of solids?

(iii) States of Matter

- The way in which particles interact with each other depends on:
 - the strength of the forces (bonds) between the particles
 - how much energy the particles have
- Particles will arrange themselves into three different physical states depending on the balance between two factors; these three states are known as **solids**, **liquids** and **gases**
- In **solids**:
 - all particles are close together and arranged in a regular pattern (or lattice) in three dimensions; the particles are fixed in place by the adjacent particles and so move by vibrating about their fixed positions



- solids have a fixed shape and cannot flow, because the particles are held in fixed positions by the bonds between them
- solids also have a fixed volume; they cannot be compressed or squashed because the particles are already close together and have no space to move in to
- a solid is a therefore state with high order and low randomness

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- As solids are heated, the particles gain more energy and vibrate more:
 - eventually, the particles have sufficient energy to break some of the bonds holding the particles together
 - when the bonds start to break, the particles are able to move past each other
 - when this happens change of state from **solid** to **liquid** has taken place
 - this change of state is known as **melting**
 - the reverse is known as **freezing** (ie liquid to solid)
 - in some cases, all of the bonds holding the particles together break at the same time, and the solid changes directly into a gas; this change of state is known as **subliming**
 - the temperature at which a solid turns into a liquid is known as the **melting point** of a solid
 - the melting point and other physical properties of a solid depend on the type of particles present and the type and strength of the forces between them



Demonstration 8.1: Warm ice, sulphur and iodine

- Gently warm a block of ice in a crucible
- Gently warm some sulphur in a crucible
- Gently warm some iodine in a crucible

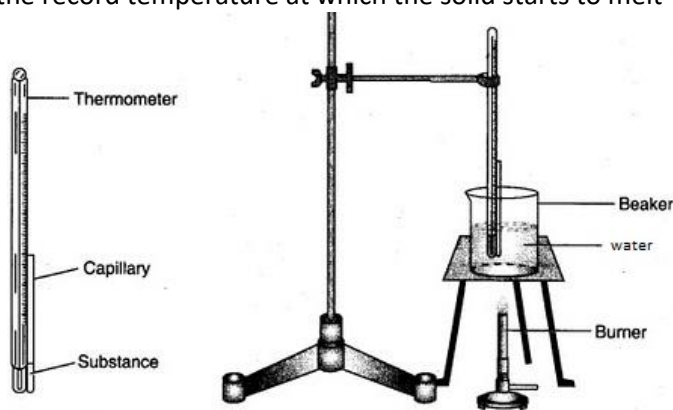
Describe the changes which you see in the solids as they are heated.
Why are these changes taking place?

Equipment needed: clay pipe triangle, three crucibles, tripod, Bunsen burner and gas, tongs, a block of ice, a spatula of sulphur powder, a spatula of iodine crystals

If possible, warm the sulphur and the iodine in a fume cupboard; if not, use an open space and keep the students a safe distance from the crucible

Place a clay pipe triangle on a tripod over a Bunsen burner connected to gas; take the ice block and place it into a crucible; then place the crucible into the clay pipe triangle; light the Bunsen burner and heat the crucible gently (medium flame) until it starts to melt; then remove the crucible, replace it with a crucible containing one spatula of sulphur, and repeat the warming process; then remove the crucible, replace it with a crucible containing one spatula of iodine, and repeat the warming process

- The melting point of a solid can be determined as follows:
 - packing a small quantity of the solid into a capillary tube until it is approximately 1 cm full
 - strap the capillary tube to the stem of a thermometer
 - place the thermometer into a heating bath (usually a beaker of water)
 - heat the bath slowly and record the temperature at which the solid starts to melt



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Demonstration 8.2: Determine the melting point of naphthalene

- (a) Observe your teacher melt a sample of naphthalene using the procedure described above

What is the melting point of naphthalene?

What is the maximum melting point which can be measured using this method?

Cannot see this experiment? Watch it here: www.youtube.com/watch?v=t8vFW56ZrRI

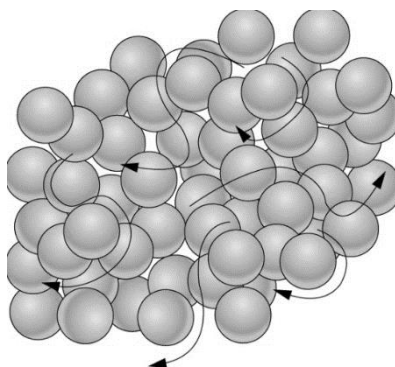
Equipment needed: thermometer with range up to 100 °C, capillary tube sealed at one end, a small quantity (less than 1 g) of finely powdered naphthalene, a spatula, a clamp, stand and boss, a 250 cm³ beaker half-filled with water, a tripod, gauze, Bunsen burner and access to gas

Fill the capillary tube with 1 cm of naphthalene and strap it to a thermometer using an elastic band; place a gauze onto a tripod over a Bunsen burner connected to a gas supply; half-fill the 250 cm³ beaker with water; use the boss to attach the clamp to the stand; place the thermometer into the beaker at a depth so that the naphthalene is submerged in the water but the top of the capillary tube is not; fix the thermometer in place with the clamp; light the Bunsen burner and heat the water on a medium flame until the naphthalene starts to melt

The melting point of naphthalene is 80 °C

Lesson 9 – What are liquids and what are their properties?

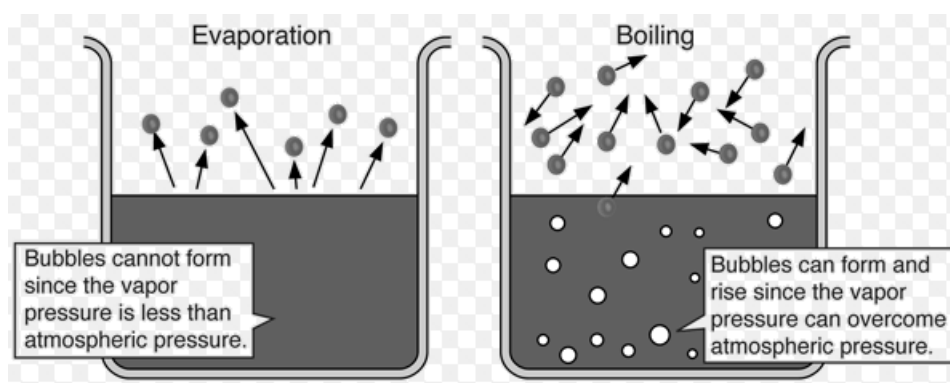
- In liquids:
 - all the particles are close together but able to move past each other because some of the bonds between the particles have been broken
 - the particles are therefore vibrating as well as moving around
 - most of the bonds holding the particles together have not been broken



- liquids do not have a fixed shape and can flow, because the particles can move past each other
 - liquids have a fixed volume; liquids cannot be compressed or squashed because the particles are already close together and have no space to move in to
 - a liquid is a state with lower order (more disorder) and more randomness than a solid
- Some particles at the surface of the liquid, if they have a lot of energy, can escape the surface of the liquid, temporarily becoming **gas** particles; this process is known as **evaporation**; as a result, all liquids contain a small amount of vapour immediately above it; the pressure exerted by this vapour is known as the **vapour pressure** of the liquid; this is why some liquids have a smell (some of the liquid particles become gas particles and enter your nose)

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- As a liquid is heated, the particles gain more energy and both move around and vibrate more:
 - the vapour pressure of the liquid therefore increases, as more of the particles at the liquid surface have enough energy to escape
 - when most of the particles have enough energy to overcome the break the bonds between them, the particles move apart and are no longer attracted to each other
 - at this temperature, the vapour pressure of the liquid is equal to the atmospheric pressure
 - when this happens change of state from liquid to gas is taking place
 - this change of state is known as **boiling**
 - the reverse of boiling and evaporating is known as **condensing**
 - the temperature at which the vapour pressure is equal to the atmospheric pressure is known as the **boiling point** of the liquid
 - the **standard boiling point** of a liquid is the temperature at which the vapour pressure of the liquid is equal to 1 atm pressure (100,000 Pa – this is normal atmospheric pressure)
 - the boiling point and other physical properties of a liquid depend on the type of particles present and the type and strength of the forces between them
- The difference between evaporation and boiling can be shown as follows:



Source: www.brainly.in



Demonstration 9.1: Heat water

- (a) Gently heat some water in a beaker until it boils

Describe the changes which you see in the water as it is heated.

How can you tell that the water is boiling?

Why do you see some bubbles before the water boils?

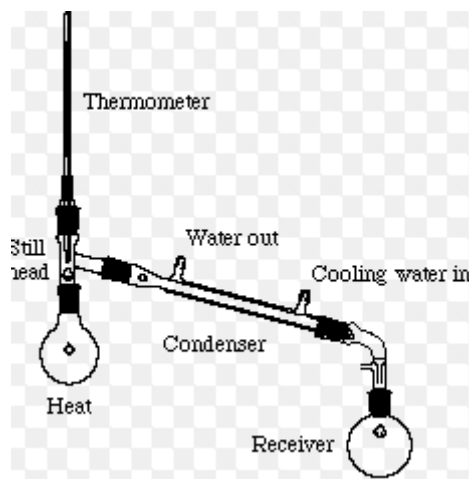
Equipment needed: gauze, tripod, Bunsen burner and gas, tongs, a 250 cm³ beaker, access to water

Place the gauze on the tripod; place the Bunsen burner under the tripod; half-fill the beaker with water; place the beaker on the gauze; light the Bunsen burner; heat the water on a medium flame

- as the water is heated, small bubbles start to form due to evaporation; the number of bubbles increases as the water gets hotter; then the bubbles start to get very large
- the water is boiling and the bubbles become large
- some of the water evaporates before it boils

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- The standard boiling point of a liquid can be determined as follows:
 - add a small amount of liquid to a flask
 - connect the flask to a three-way head with a thermometer inserted into it
 - attach the head to a condenser set up for distillation
 - ensure that the bulb of the thermometer is just above the level of the liquid
 - record the maximum temperature reached – this is the boiling point of the liquid



Demonstration 9.2: Determine the boiling point of water

- (a) Observe your teacher boil water using the procedure described above

What is the boiling point of water?

Equipment needed: thermometer with range up to 150 °C, quickfit 250 cm³ round-bottomed flask, quickfit three-way head, quickfit bung with hole for thermometer, quickfit condenser, 100 cm³ beaker, tripod, gauze, Bunsen burner, access to gas, 2 clamp stands, each with boss and clamp, 2 pipes to connect water to tap and sink

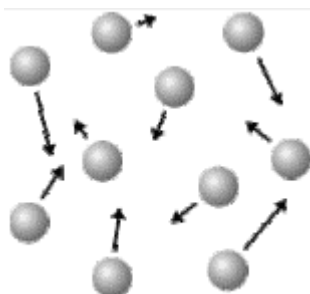
Set up close to a sink; place a gauze on a tripod; place a Bunsen burner connected to gas underneath the tripod; pour 100 cm³ of water into the flask; place the flask on top of the gauze and clamp into position using a stand, boss and clamp; insert the three way head into the top of the flask, place the thermometer through the quickfit bung with hole, insert the bung into the top of the head; adjust the height of thermometer so it is well above the level of the water; attach condenser to head; clamp condenser into position using another stand, boss and clamp; place beaker underneath lower end of condenser to catch water; use water tube to connect tap to lower water valve on condenser; use another water tube to connect upper water valve on condenser to sink; turn tap on and ensure that condenser is full of water and that there is a steady flow of water into the sink; then light Bunsen burner and heat on medium flame until temperature has reached a steady level

- The boiling point of water is 100 °C

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Lesson 10 - What are gases and what are their properties?

- In gases:
 - the particles are moving around rapidly and all bonds between them have been broken
 - there are significant spaces between the particles



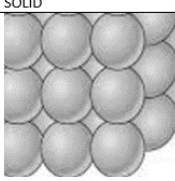
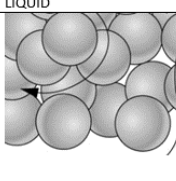
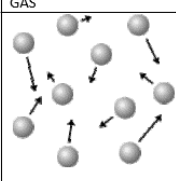
- gases have no fixed shape and will fill the entire volume available to them, because the particles are able to move rapidly and freely
- gases also have no fixed volume; they can be compressed or squashed because there are spaces between the particles
- a gas is a state with the lowest order (highest disorder) and most randomness
- in most gases, the forces in between the particles are negligible and can be ignored; as a result, unlike solids and liquids, the physical properties of a gas do not depend on the type of particle involved or the forces between them; all gases therefore have very similar physical properties



Test your knowledge 10.1: Understanding the structure and properties of solid, liquids and gases

Copy and complete the following table:

	SOLID	Change of state	LIQUID	Change of state	GAS
Arrangement of particles (Diagram)		→		→	
Arrangement of particles (description)		←		←	
Properties (volume and shape)					
Bonds between particles (intact or broken)					
Ordered or disordered?					

	SOLID	Change of state	LIQUID	Change of state	GAS
Arrangement of particles (diagram)		melting →		boiling →	
Arrangement of particles (description)	Particles are packed close together in a regular arrangement and cannot move; particles vibrate about fixed positions		Particles are packed close together but not in a regular arrangement; there are some spaces so particles can move		Particles are far apart and moving freely
Properties (volume and shape)	Fixed volume Fixed shape	← freezing	Fixed volume No fixed shape	← condensing	No fixed volume No fixed shape
Bonds between particles (intact or broken)	All bonds are intact		Most bonds are intact; some are broken		All bonds are broken
Degree of disorder	High order Low randomness		Less order/more randomness than solids but more order/less randomness than gases		Low order High randomness

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- It is often important to indicate the physical state of a material; this is shown by showing an abbreviation for the physical state of the substance in brackets, after the chemical formula; this is known as the **state symbol**:

State	Symbol
Solid	S
Liquid	L
Gas	G

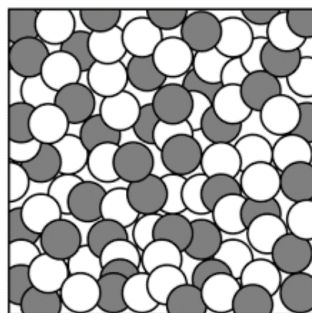
- Eg Ice is represented by the formula $\text{H}_2\text{O} (\text{s})$
- Eg Water is represented by the formula $\text{H}_2\text{O} (\text{l})$
- Eg Steam or water vapour is represented by the formula $\text{H}_2\text{O} (\text{g})$

(iv) Effect of impurities on solids and liquids

- Every pure substance has a fixed melting point and standard boiling point, unless it breaks down chemically before reaching that temperature; the melting points and boiling points of elements and compounds can be found in databooks and on the internet
- Mixtures are not pure substances and do not have fixed melting points; mixtures tend to melt and boil over a range of temperatures and the melting and boiling point will vary with the composition of the mixture
- It follows, therefore, that impurities in a substance will change its melting point; impurities in solid structures tend to weaken the bonds between the particles in the solid; this makes the bonds easier to break, which lowers the melting point; impurities also cause the solid to melt more gradually; the melting point can therefore be used as a test for the purity of a solid; impure solids melt at a lower temperature than the melting point of the pure solid, and melt over a range of temperatures; pure solids melt suddenly at their known melting point
- The boiling point can also be used as a test for the purity of liquids, although it is harder to do this accurately; impurities tend to reduce the vapour pressure of the pure liquid and hence increase the boiling point of the liquid

(v) Solutions

- Pure substances can exist either as solids, liquids or gases, depending on the temperature and the nature of the forces between the particles
- It is also common for two or more different substances to mix uniformly together; this usually happens when a solid, liquid or gas dissolves in another liquid; the particles in the dissolving substance separate and distribute themselves evenly throughout the liquid; the resulting mixture is known as a **solution**



UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- The minor component in a solution is called the **solute**; the major component in a solution is called the **solvent**; the most common solvent is water
- Solutions in which the solvent is water are known as **aqueous** solutions; solutes dissolved in water are shown by a special state symbol – aq
 - Eg a solution of sodium chloride in water is represented by the formula NaCl(aq); the sodium and chloride ions separate and distribute themselves throughout the water molecules



Test your knowledge 10.2: Understanding mixtures and solutions

- Suggest a simple procedure for determining whether a sample of aspirin is pure
 - State the particles present in an aqueous solution of magnesium chloride and explain how they are arranged; identify the solute and the solvent
- Take melting point by placing a small sample in a capillary tube and strapping to a thermometer; immerse in a suitable liquid and heat until it melts; note temperature at which it starts and finishes melting; compare melting point with that of aspirin; if sample melts sharply and at same temperature it is pure; if it melts gradually over a range of lower temperatures it is not pure
 - Mg²⁺ ions, Cl⁻ ions and water molecules; the Mg²⁺ and Cl⁻ are distributed evenly throughout the water molecules; the solute is magnesium chloride (or Mg²⁺ and Cl⁻) and the solvent is water

Lesson 11 – What evidence is there to support the kinetic model of matter?

(vi) Evidence for the Kinetic Model of Matter

- Although atoms, ions and most molecules are too small to see, there is some visible evidence for the kinetic model of matter, even without using a microscope; this evidence comes from **Brownian motion** and **diffusion**
- **Brownian motion** is the random movement of particles suspended in a liquid or gas as a result of collisions with the liquid or gas particles they are suspended in:
 - when particles large enough to be visible are suspended in a liquid or a gas, they are observed to move randomly, sometimes changing direction suddenly
 - these particles move because they are colliding with gas or liquid particles which are too small to see
 - Brownian motion therefore provides evidence for the particle theory of matter in liquids and gases



Demonstration 11.1: Observe Brownian motion in pollen grains or powdered sulphur

Your teacher prepared three slides for you to view under a microscope:

- one slide contains some pollen grains in water
- one slide contains some sulphur grains in water
- one slide contains some smoke in a smoke cell

Look at the pollen grains, powdered sulphur or smoke particles under the microscope. Describe what you see.

Can't view this demonstration? Watch it here: www.youtube.com/watch?v=hy-clLi8gHg)

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Equipment needed: microscope, slide, slide cover, smoke cell, pollen grains, dropping pipettes, access to water
Take a micro-spatula measure of pollen grains and place them on a slide; add a drop of water and then add the slide cover; remove any pollen/water which gets squeezed out when the slide cover is added.

Repeat using sulphur powder instead of pollen grains.

Light a small piece of wood or paper and collect some of the smoke using a dropping pipette; inject the smoke from the dropping pipette into the smoke cell and place a slide cover over it to keep the smoke in.

Place all three under a microscope in turn. Students should be able to see Brownian motion

- pollen grains, smoke particles and sulphur grains can be seen moving randomly and changing direction suddenly (Brownian motion)



Activity 11.2: Observe Brownian motion in the classroom

Find a dusty room in the school which has a window you can see through from outside.

Sweep the room then look into the room from outside. What can you see?

Can't do this? Watch it here: www.youtube.com/watch?v=7orMPKrYPwE

If there is direct sunlight coming into the room, students should be able to observe Brownian motion in the dust particles

- dust particles moving randomly, changing direction constantly

- **Diffusion** is the net movement of particles from a region of high concentration to a region of low concentration; particles added to a specific region of a liquid or a gas will generally spread out until they are uniformly distributed throughout the liquid or gas
- Diffusion results from the random movement of particles; the higher the concentration of particles in a space, the more likely particles are to move out of that space; it can often be a slow process because the particles are constantly colliding with the other particles in the liquid or gas, which causes them to change direction; diffusion occurs more quickly in gases than in liquids, because the particles are less densely packed and so collide with each other less frequently



Demonstration 11.3: Observe bromine or nitrogen dioxide diffuse through air

You will see two gas jars: one contains bromine vapour (orange) and the other contains air (colourless)

The two gas jars will be placed horizontally facing each other; then the lids will be removed and the two jars connected so the two gases can mix.

What do you observe?

(As an alternative, nitrogen dioxide (brown) can be used instead of bromine)

Can't do this experiment? Watch it here: www.youtube.com/watch?v=H7QsDs8ZRMI

Note: This demonstration should be carried out in a fume cupboard.

Equipment needed: two gas jars with lids, access to liquid bromine or copper (II) nitrate

If using bromine: take one dropping pipette full of bromine and drop in into a gas jar inside a fume cupboard; when most of the gas jar is full of bromine, place a lid on the gas jar. Carefully place the gas jar horizontally and position another gas jar close by. Remove the lid and quickly connect the gas jars, ensuring that the bromine cannot escape.

If using NO_2 : Place 1 g of copper nitrate in a boiling tube. Heat the copper nitrate until it starts to decompose, giving off a brown gas. When the boiling tube is full of the brown gas, remove it from the heat. Then connect the boiling tube to another empty boiling tube upside down so that the air in the two tubes can mix.

- The orange/brown colour will gradually fill the other jar or tube, although it does take some time



Activity 11.4: Observe diffusion using smell

Concentrated ammonia has a very strong smell. Your teacher will open a bottle of concentrated ammonia in one corner of the classroom.

How long before you can smell it? Why can you smell it?

Note: this demonstration should be carried out either close to the end of a lesson or in another empty classroom so the room can be evacuated immediately afterwards. Take care that the smell does not overpower the students closest to the bottle. Replace the lid as quickly as possible.

- Students closest to the ammonia bottle will smell it quickly; students further away will take longer
- The ammonia molecules evaporate and diffuse through the room; they move slowly because they keep bumping into air particles

- Not all particles diffuse at the same rate; at the same temperature, all particles have the same average kinetic energy; but because $\text{KE} = \frac{1}{2}mv^2$, then particles with a larger mass must be moving more slowly, and particles with a larger mass must be moving more quickly, if both particles are at the same temperature
- Shown mathematically, if $\text{KE} = \frac{1}{2}mv^2$ and all particles at the same temperature have the same energy, then:

$$m_1v_1^2 = m_2v_2^2, \text{ so } \frac{v_2^2}{v_1^2} = \frac{m_1}{m_2} \text{ and } \frac{v_2}{v_1} = \sqrt{\frac{m_1}{m_2}}$$
- In other words, **the relative rate of diffusion of two gases is inversely proportional to the square root of their relative masses**; this is known as **Graham's Law of Diffusion**

Example: Argon atoms have a relative atomic mass of 40. Neon atoms have a relative atomic mass of 20. How fast will neon diffuse compared to Argon?

Answer: $(m_1/m_2) = 40/20 = 2$, so $\sqrt{(m_1/m_2)} = 1.4 = v_2/v_1$. **So Neon will diffuse 1.4 times faster than Argon.**



Test your knowledge 11.5: Using Graham's Law of Diffusion

Helium, neon, argon and krypton have relative atomic masses of 4, 20, 40 and 84 respectively.

- Which gas will diffuse the most quickly and which gas will diffuse the most slowly?
- How many times faster will helium diffuse compared to argon?
- How many times faster will neon diffuse compared to krypton?

- (a) He most quickly, Kr most slowly
- (b) $\sqrt{40/4} = 3.2$ times faster
- (c) $\sqrt{84/20} = 2.0$ times faster

Lesson 12 – What are the gas laws?

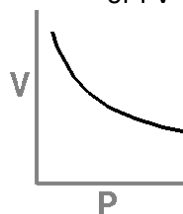
g) The Gas Laws

- Because the particles in gases do not interact with each other, most gases behave in very similar ways; the basic behaviour of gases can be predicted according to some simple laws called the **gas laws**; these laws deal with the relationship between the temperature, pressure and volume of gases
- All gases exert a **pressure** on the walls around them:
 - gas particles move around rapidly
 - inside a container, the particles will be constantly colliding with the sides of the container
 - when they do this, they exert a force on the sides of the container
 - the force exerted per unit area by a gas on the sides of a container is called the **pressure** of the gas
 - pressure has units of N/m^2 or Pascals (Pa)
 - the pressure of air in the atmosphere on the earth's surface is around 100,000 Pa or 100 kPa; this is often referred to as 1 atmosphere pressure or "standard pressure"
- Gases always occupy the entire **volume** available to them; if a container has a fixed volume, the gas particles will spread out until they are evenly distributed throughout the container; if a container has flexible volume, then its volume will increase or decrease depending on the pressure of the gas and its temperature
- The **temperature** of the gas is a measure of the kinetic energy of the particles; the higher the temperature, the faster the gas particles are moving and the more energy they have

(i) Boyle's Law

- If the volume occupied by a fixed amount of gas increases at a fixed temperature, the particles will collide with the sides of the container less often and therefore the pressure of the gas will increase; similarly, if a fixed amount of gas is compressed into a smaller volume, its pressure will increase; **Boyle's Law** states that "**the pressure exerted by a given mass of an ideal gas is inversely proportional to the volume it occupies, if the temperature remains unchanged in a closed system**" (a closed system is a system from which particles cannot escape, and into which particles cannot enter)

- Mathematically: $P \propto \frac{1}{V}$ or $PV = k$ or $P_1V_1 = P_2V_2$



- Graphically:

Example: If a sample of gas exerts a pressure of 100,000 Pa in a container of fixed volume 0.05 m^3 , what pressure will it exert if the volume is compressed to 0.02 m^3 ?

Answer: $P_1V_1 = 100,000 \times 0.05 = 5,000 = P_2V_2$, so $P_2 = 5000/V = 5000/0.02 = \mathbf{250,000 \text{ Pa}}$

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

(ii) Charles' Law

- If a fixed amount of gas is heated at constant pressure, the particles have more energy and move faster; this will cause the gas to spread out more; similarly, if a fixed amount of gas is cooled at constant pressure, the particles will have less energy and move more slowly, causing the gas to contract; **Charles' Law** states that **“the volume occupied by a given mass of an ideal gas is directly proportional to its absolute temperature, if the pressure remains unchanged in a closed system”**

- Mathematically: $V \propto T$ or $V/T = k$ or $V_1/T_1 = V_2/T_2$



- Graphically:

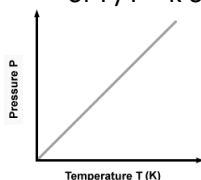
Example: If a sample of gas occupies a volume of 100 dm³ at a temperature of 300 K, what volume will the gas occupy if it is heated to a temperature of 360 K? at constant pressure?

Answer: $V_1/T_1 = 100/300 = 1/3 = V_2/T_2$, so $V_2 = 1/3 \times 360 = \mathbf{120 \text{ dm}^3}$

(iii) Gay-Lussac's Law

- If a fixed amount of gas is heated at constant volume, the particles have more energy and move faster; this will cause them to collide with the sides of the container with more force, resulting in an increase in pressure; similarly, a fixed amount of gas is cooled at constant volume, the particles will collide with the sides of the container with less force, resulting in a decrease in pressure; **Gay-Lussac's Law** states that **“the pressure exerted by a given mass of an ideal gas is directly proportional to its absolute temperature, if the volume remains unchanged in a closed system”**

- Mathematically: $P \propto T$ or $P/T = k$ or $P_1/T_1 = P_2/T_2$



- Graphically:

Example: If a sample of gas occupies exerts a pressure of 100,000 Pa at a temperature of 300 K, what pressure will the gas exert if it is heated to a temperature of 450 K whilst maintaining its volume constant?

Answer: $P_1/T_1 = 100000/300 = 1000/3 = P_2/T_2$, so $P_2 = 1000/3 \times 450 = \mathbf{150,000 \text{ Pa}}$



Test your knowledge 12.1: Using the gas laws

- If a sample of gas exerts a pressure of 120,000 Pa in a container of fixed volume 0.03 m³, what pressure will it exert if the volume is compressed to 0.02 m³?
- If a sample of gas occupies a volume of 40 dm³ at a temperature of 300 K, what volume will the gas occupy if it is heated to a temperature of 375 K?
- If a sample of gas occupies exerts a pressure of 100,000 Pa at a temperature of 300 K, what pressure will the gas exert if it is heated to a temperature of 900 K?
- If a sample of gas exerts a pressure of 250,000 Pa in a container of fixed volume 0.05 m³, what pressure will it exert if the volume is expanded to 0.10 m³?
- If a sample of gas occupies a volume of 200 dm³ at a temperature of 300 K, what volume will the gas occupy if it is cooled to a temperature of 270 K?
- If a sample of gas occupies exerts a pressure of 100,000 Pa at a temperature of 270 K, what pressure will the gas exert if it is heated to a temperature of 1350 K?

(a) 180,000 Pa; (b) 50 dm³; (c) 300,000 Pa; (d) 125,000 Pa; (e) 180 dm³; (f) 500,000 Pa

(iv) Combined Gas Law

- Boyle's Law, Charles' Law and Gay-Lussac's Law can be combined as follows:

$$PV/T = k \text{ or } P_1V_1/T_1 = P_2V_2/T_2$$

This law is called the **combined gas law**

Example: If a sample of gas occupies exerts a pressure of 100,000 Pa in a volume of 5 dm³ at a temperature of 300 K, what pressure will the gas exert if it is heated to a temperature of 450 K whilst also being compressed into a volume of 4 dm³?

Answer: $P_1V_1/T_1 = 100000 \times 5/300 = 5000/3 = P_2V_2/T_2$, so $P_2 = 5000/3 \times 450/4 = \mathbf{187,500 \text{ Pa}}$



Test your knowledge 12.2: Using the combined gas law

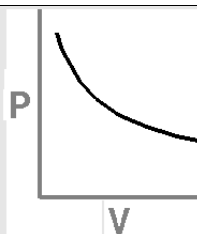
- If a sample of gas occupies exerts a pressure of 200,000 Pa at a temperature of 350 K, what pressure will the gas exert if it is heated to a temperature of 400 K whilst maintaining its volume constant?
- If a sample of gas exerts a pressure of 150,000 Pa in a container of fixed volume 0.04 m³, what pressure will it exert if the volume is expanded to 0.06 m³?
- If a sample of gas occupies a volume of 50 dm³ at a temperature of 300 K, what volume will the gas occupy if it is cooled to a temperature of 275 K?
- If a sample of gas occupies exerts a pressure of 100,000 Pa in a volume of 2 dm³ at a temperature of 320 K, what pressure will the gas exert if it is cooled to a temperature of 300 K whilst also being compressed into a volume of 1.5 dm³?
- A hot air balloon has a volume of 100 dm³ at 300 K and 100,000 Pa. When the air in the balloon is heated, the volume increases to 120 dm³ and the pressure increases to 110,000 kPa. What is the new temperature of the air in the balloon?
- A tyre has a volume of 10 dm³ when the temperature of the air inside is 300 K and the pressure is 100,000 Pa. Later that day, the temperature of the air in the tyre has increased to 400 K and the pressure of the tyre has increased to 120,000 Pa. What is the new volume of the tyre?
- In each of questions (a) – (d), state which Law you used to calculate your answer.

(a) 229,000 Pa; (b) 100,000 Pa; (c) 45.8 dm³; (d) 125,000 Pa; (e) 396 K; (f) 11.1 dm³, (g) Gay-Lussac, Boyle, Charles, Combined



Extension 12.3: Applying the gas laws

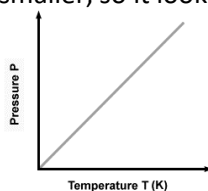
- How does volume change with pressure? Can you sketch a graph to show this relationship (put P on the y-axis and V on the x-axis). Why do balloons burst when they are squeezed?
- How does volume change with temperature? Can you sketch a graph to show this relationship (put V on the y-axis and T on the x-axis). Why do tyres sometimes look deflated on cold mornings?
- How does pressure change with temperature? Can you sketch a graph to show this relationship (put P on the y-axis and T on the x-axis). Why do tyres sometimes burst on hot road surfaces?



- (a) volume decreases as pressure increases;
when you reduce the volume of a balloon the pressure increases until it is more than the balloon can bear



- (b) volume increases as temperature increases,
at low temperatures the volume of the tyre is smaller, so it looks deflated



- (c) pressure increases as temperature increases,
as the air temperature in the tyre increases, so does the pressure in the tyre, on very hot days the pressure will be very high and maybe too much for the tyre to bear

(v) Dalton's Law of Partial Pressures

- The gas laws apply even if the gas is a mixture of different gases; the **partial pressure** of a gas in a mixture of gases is the pressure which would be exerted if that gas alone occupied the entire gaseous volume
- For example, dry air consists of a mixture of different gases, mainly nitrogen, oxygen and argon; at normal atmospheric pressure, the partial pressure of each gas is as follows:

Nitrogen:	79,000 Pa
Oxygen:	20,000 Pa
Argon:	1000 Pa
- According to **Dalton's Law of Partial Pressures**, the **total pressure in a system is the sum of the partial pressures of the different gases in the system**
 - so the overall pressure of air at normal atmospheric pressure is $79000 + 20000 + 1000 = 100,000$ Pa
 - like the other gas laws, this law works because the gas particles do not interact with each other


Test your knowledge 12.4: Using Dalton's Law of Partial Pressures

- (a) On a day when the atmospheric pressure was found to be 102 kPa, a sample of air was found to contain 2.3% H₂O, 1.0% Ar, 77.4% N₂ and 19.3% O₂.

Calculate the partial pressure of each gas in the air sample.

- (a) H₂O: $2.3/100 \times 102 = 2.35$ kPa; Ar: $1.0/100 \times 102 = 1.02$ kPa; N₂ = $77.4/100 \times 102 = 78.95$ kPa ; O₂ = $19.3/100 \times 102 = 19.7$ kPa

Lesson 13 – What are giant ionic structures?

h) Chemical Structures

- The way in which individual atoms, ions and molecules are attracted to each other and hence join together is called **bonding**
- The way in which atoms, molecules or ions arrange themselves on a larger scale is called **structure**
- Most structures in the solid state consist of a regular, repeating pattern, usually in three dimensions; this regular, repeating pattern is called a **lattice**
- The physical and chemical properties of a solid or liquid material depend on:
 - the type of particles they contain (atoms, molecules or ions)
 - the type of bonding between the particles (ionic, covalent, metallic or intermolecular)
 - the type of **structure** which the material adopts
- There are four main types of structure:
 - **Giant ionic lattice structure** – this is the structure adopted by oppositely charged ions held together by ionic bonds
 - **Giant metallic lattice structure** – this is the structure adopted by atoms when they are held together by metallic bonds
 - **Simple molecular structure** – this is the structure adopted when atoms combine with covalent bonds to form molecules, which are held together by intermolecular bonds
 - **Giant covalent (or macromolecular) lattice structure** – this is the structure adopted when atoms combine with covalent bonds but cannot form stable molecules

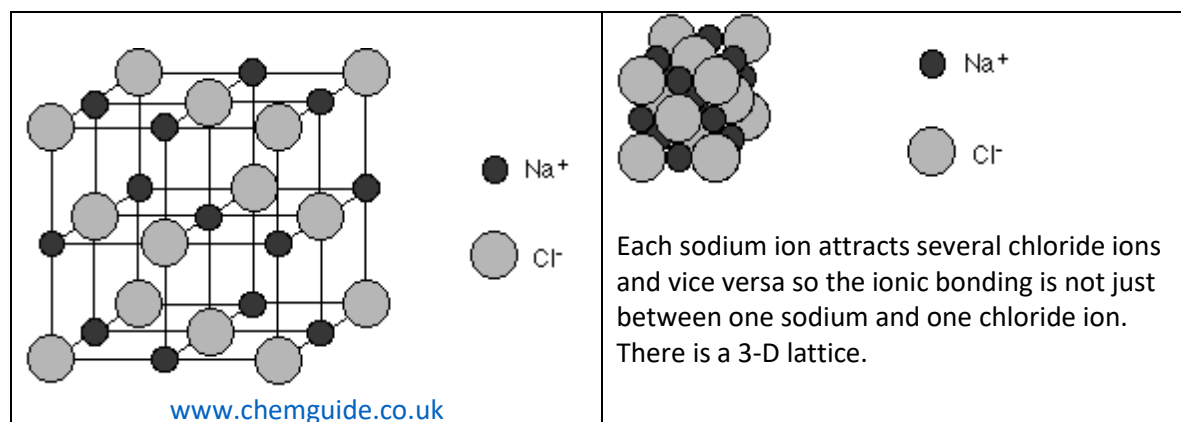
(i) Giant Ionic Structures

- An ionic bond is an attraction between oppositely charged ions; after the ions are formed, they all come together to form a **lattice**; all of the anions are surrounded by cations and all of the cations are surrounded by anions; a **giant ionic structure** consists of a large number of cations and anions arranged in a regular lattice

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Example – sodium chloride

- In sodium chloride, each sodium ion is surrounded by six chloride ions and vice versa
- The diagram below shows the structure of sodium chloride; the pattern repeats in this way and the structure extends (repeats itself) in all directions throughout a single crystal
- Remember - this diagram represents only a tiny part of a single sodium chloride crystal



- Giant ionic structures tend to have a high melting point:
 - The attractions between the oppositely charged ions (ie the ionic bonds) are very strong; a lot of energy is thus required to overcome them

Compound	NaCl	MgO
Melting point/°C	801	2852

- The higher the charge on the ions, and the smaller they are, the stronger the attraction between them and the higher the melting and boiling points
- Eg In MgO, the ions have a 2+ and 2- charge and thus the attraction between them is stronger than in NaCl, in which the ions have a +1 and a -1 charge, so the melting point is higher
- Giant ionic structures conduct electricity in the molten and aqueous states but not in the solid state:
 - since giant ionic structures contain ions, they are attracted by electric fields and ions will, if possible, move towards the electrodes and thus conduct electricity
 - in the solid state, however, the ions are not free to move since they are tightly held in place by each other, so giant ionic structures do not conduct electricity in the solid state – they are insulators
 - in the liquid state, the ions are free to move and so can move towards their respective electrodes; thus giant ionic structures can conduct electricity in the liquid state
- Giant ionic structures tend to be hard and brittle:
 - as soon as the ions are moved out of place, the structure breaks down because the oppositely charged ions are no longer adjacent to each other

UNIT 2 – PARTICLES, BONDING AND STRUCTURES



Test Your Knowledge 13.1: Understanding the physical properties of giant ionic structures

Copy and complete the following table about the physical properties of giant ionic lattice structures:

Property	Explanation
high melting point	
non-conductor of electricity in solid state	
conductor of electricity in molten or aqueous state	
brittle	

Property	Explanation
high melting point	Strong ionic bonds Need a large amount of energy to break
non-conductor of electricity in solid state	Ions Are not free to move
conductor of electricity in molten or aqueous state	Ions Are free to move
brittle	Layers cannot slide over each other As it would cause repulsion between the ions

- Giant ionic structures are always compounds as they always contain different types of atom; they are represented by their **unit formula**; the unit formula of a giant ionic structure is the simplest whole number ratio of the ions in the compound, shown by writing the cation followed by the anion:

Name of compound	Formula	Meaning of formula
Sodium chloride	NaCl	Ratio of Na ⁺ to Cl ⁻ ions is 1:1
Sodium oxide	Na ₂ O	Ratio of Na ⁺ to O ²⁻ ions is 2:1
Magnesium chloride	MgCl ₂	Ratio of Mg ²⁺ to Cl ⁻ ions is 1:2
Aluminium oxide	Al ₂ O ₃	Ratio of Al ³⁺ to O ²⁻ ions is 2:3

- the charges on the ions always cancel out, so the formula has no overall charge

- Some ions contain two or more atoms – these are known as polyatomic ions; common examples are:

ammonium	NH ₄ ⁺
nitrate	NO ₃ ⁻
hydroxide	OH ⁻
sulphate	SO ₄ ²⁻
carbonate	CO ₃ ²⁻

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- these ions combine with other ions to form ionic compounds with giant ionic structures; if a unit formula contains more than one of these ions, the formula of the ion is written inside brackets and the number of ions is written outside the brackets:

Name of compound	Formula	Meaning of formula
sodium carbonate	Na_2CO_3	Ratio of Na^+ to CO_3^{2-} ions is 2:1
ammonium sulphate	$(\text{NH}_4)_2\text{SO}_4$	Ratio of NH_4^+ to SO_4^{2-} ions is 2:1
ammonium nitrate	NH_4NO_3	Ratio of NH_4^+ to NO_3^- ions is 1:1
magnesium nitrate	$\text{Mg}(\text{NO}_3)_2$	Ratio of Mg^{2+} to NO_3^- ions is 1:2
aluminium sulphate	$\text{Al}_2(\text{SO}_4)_3$	Ratio of Al^{3+} to SO_4^{2-} ions is 2:3



Test Your Knowledge 13.2: Deducing the unit formula of giant ionic structures

Deduce the unit formula of the following giant ionic structures:

- lithium oxide
- barium chloride
- lithium nitride
- ammonium chloride
- potassium nitrate
- aluminium nitrate
- magnesium hydroxide
- magnesium sulphate
- ammonium nitrate
- ammonium sulphate

(a) Li_2O ; (b) BaCl_2 ; (c) Li_3N ; (d) NH_4Cl ; (e) KNO_3 ; (f) $\text{Al}(\text{NO}_3)_3$; (g) $\text{Mg}(\text{OH})_2$; (h) MgSO_4 ; (i) NH_4NO_3 ; (j) $(\text{NH}_4)_2\text{SO}_4$



Extension 13.3: Comparing the melting points of giant ionic structures

- State, with reasons, whether NaCl or KCl will have a higher melting point
- State, with reasons, whether NaCl or MgCl_2 will have a higher melting point
- State, with reasons, whether MgCl_2 or MgO will have a higher melting point
- State, with reasons, whether NaCl or NaF will have a higher melting point

- NaCl because Na^+ is smaller than K^+ so stronger attraction to Cl^-
- MgCl_2 because Mg^{2+} is smaller and more highly charged than Na^+ so stronger attraction to Cl^-
- MgO because O^{2-} is more highly charged than Cl^- so stronger attraction to Mg^{2+}
- NaF because F^- is smaller than Cl^- so stronger attraction to Na^+

- At room temperature, giant ionic structures are found either in the solid state or dissolved in water; the physical state of the structure can be shown by the corresponding state symbol after the formula:
 - $\text{NaCl}(\text{s})$ means a solid sample of sodium chloride
 - $\text{NaCl}(\text{aq})$ means a sample of sodium chloride dissolved in water

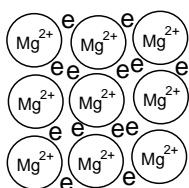
Lesson 14 – What are giant metallic structures?

(ii) Giant Metallic Structures

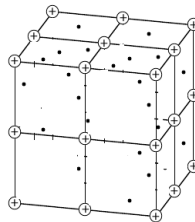
- Metallic bonding is the attraction between cations and a sea of delocalised electrons; the cations are arranged to form a lattice, with the electrons free to move between them; this is a giant metallic structure

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- A material which adopts a giant metallic structure is called a **metal**
- A simplified form of a magnesium lattice is shown below:



This is a simplified 2D form of the metal lattice



This is a simplified 3D lattice for sodium; the + are sodium ions and the dots are delocalised electrons

- Giant metallic structures conduct electricity:
 - the electrons in a metal are delocalised, they are free to move throughout the crystal in a certain direction when a potential difference is applied and metals can thus conduct electricity in the solid state
 - the delocalised electron system is still present in the liquid state, so metals can also conduct electricity well in the liquid state
- Giant metallic structures tend to have high melting and boiling points:
 - metallic bonds (the attractions between cations and delocalised electrons) are relatively strong, although usually not as strong as ionic bonding

Metal	Na	K	Be	Mg
Melting point/ °C	98	64	1278	649

- Smaller ions, and those with a high charge, attract the electrons more strongly and so have higher melting points than larger ions with a low charge
- Eg Na^+ is smaller than K^+ so Na has a higher melting point than K
- Eg Mg^{2+} has a higher charge than Na^+ and is also smaller, so Mg a higher melting point than Na
- Giant metallic structures tend to be malleable and ductile:
 - it possible to move the positive ions around without destroying the metallic bonding, because wherever the ions are, they will still be surrounded by delocalised electrons
 - this means that metals can be bent into different shapes (ie they are malleable) and turned into wires (they are ductile) without destroying their structure
- Almost all giant metallic structures are elements; they are therefore represented simply by the atom, and they are named after the atom from which they are made

Name of giant metallic structure	Formula	Meaning of formula
Magnesium	Mg	Only Mg atoms present
Aluminium	Al	Only Al atoms present

- Giant metallic structures are almost always solids at room temperature; to distinguish between an individual atom of magnesium and a sample of solid magnesium metal, the solid state symbol is used:
 - Mg(s) means a sample of solid magnesium metal
 - Mg(g) means an individual magnesium atom

UNIT 2 – PARTICLES, BONDING AND STRUCTURES



Test Your Knowledge 14.1: Understanding the physical properties of giant metallic structures

Copy and complete the following table about the physical properties of giant metallic lattice structures:

Property	Explanation
High melting point	
Conductor of electricity	
Malleable	

Property	Explanation
High melting point	Strong metallic bonds Require lots of energy to break
Conductor of electricity	Delocalised electrons Are free to move
Malleable	Ions can be moved around Without breaking the metallic bonds



Extension 14.2: Comparing the physical properties of giant metallic structures

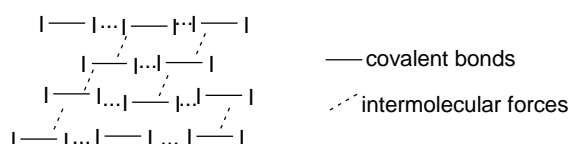
- State, with reasons, whether Li or Be will have a higher melting point
- State, with reasons, whether Mg or Be will have a higher melting point
- Suggest one other way in which the physical properties of Mg and Be will be different

- Be, because the Be^{2+} ions are smaller and more highly charged than Li^+ , so attract the delocalised electrons more strongly
- Be because the Be^{2+} ions are smaller than Mg^{2+} , so attract the delocalised electrons more strongly
- Be will be harder/less malleable and ductile

(iii) Simple Molecular Structures

- A covalent bond is a pair of electrons shared between two atoms; the electron pair in the covalent bond is attracted to the nuclei of both atoms, and this holds the atoms together; these bonds can be very strong
- In most cases, atoms bond with one or a few other atoms to form a molecule, after which the bonding capacity of each atom has been reached; in such cases, the structure of the material consists of a large number of individual **molecules** with no strong bonds holding the molecules together
- The molecules are held together only by **intermolecular forces** (Van der Waal's forces or hydrogen bonding), which are much weaker than covalent bonds, but their strength is very variable and intermolecular forces are sometimes strong enough to keep the substance in the solid or liquid state at room temperature
- In the gaseous state, the intermolecular forces are broken but the bonds within the molecule remain intact; the gas phase therefore consists of molecules, not atoms
- There are many examples of simple molecular substances: CH_4 , I_2 , He, S_8 , P_4 , O_2 , H_2O , NH_3

Example - Iodine



UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- Noble gases do not form any bonds at all and so do not form molecules; the individual atoms are held together by intermolecular forces, so the structure is called **simple atomic** but the materials behave in a manner similar to simple molecular structures
- Simple molecular structures can be elements or compounds; both are represented by their **molecular formula**; this shows the actual number of atoms of each element in one molecule of the substance; the more electropositive atom is often, but not always, shown first:
- Molecules which are elements are named after the atom from which they are made; there are systematic methods for naming molecular compounds which will be studied later in the course (see Unit 6 – Redox Reactions and Unit 7 – Introduction to Organic Chemistry); most simple molecules also have common names

Name of simple molecular structure	Formula	Meaning of formula
Chlorine	Cl ₂	Each molecule contains 2 Cl atoms
Oxygen	H ₂	Each molecule contains 2 O atoms
Water	H ₂ O	Each molecule contains 2 H atoms and 1 O atom
Ethane	C ₂ H ₆	Each molecule contains 2 C atoms and 6 H atoms
Carbon Dioxide	CO ₂	Each molecule contains 1 C atom and 2 O atoms
Ammonia	NH ₃	Each molecule contains 1 N atom and 3 H atoms
Glucose	C ₆ H ₁₂ O ₆	Each molecule contains 6 C, 12 H and 6 O atoms



Test Your Knowledge 14.3: Exploring different simple molecular structures

The following is a list of the molecular formulae of some substances which have simple molecular structures: Cl₂, H₂O, CO₂, O₂, N₂, NH₃, CH₄, Br₂, I₂, C₁₂H₂₂O₁₁

Some are elements and some are compounds

Some are gases, some are liquids and some are solids at room temperature.

Copy the table below and put each formula into the correct box:

State at room temperature	Type of substance	
	Element	Compound
SOLID		
LIQUID		
GAS		

State at room temperature	Type of substance	
	Element	Compound
SOLID	I ₂	C ₁₂ H ₂₂ O ₁₁
LIQUID	Br ₂	H ₂ O
GAS	Cl ₂ O ₂ N ₂	CO ₂ NH ₃ CH ₄

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

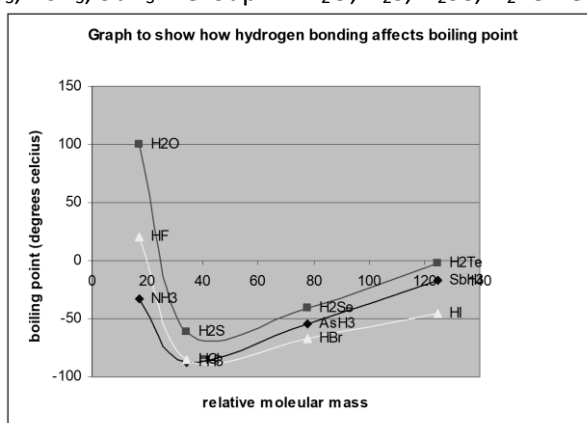
Lesson 15 – What are the physical properties of simple molecular substances?



Summary Activity 15.1: Intermolecular forces

- What are Van der Waal's forces and what is the origin of Van der Waal's forces?
 - What is hydrogen bonding and what is the origin of hydrogen bonding?
 - Why are the Van der Waal's forces between bromine molecules stronger than those between chlorine molecules?
 - Why are the intermolecular forces between water molecules stronger than those between carbon dioxide molecules?
- Temporary dipoles, caused by random movement of electrons in molecules, induce temporary dipoles in adjacent molecules; the resulting attraction between the molecules is a Van der Waal's force
 - Hydrogen bonding is the attraction between an electropositive hydrogen atom on one molecule and an electronegative atom (O, N or F) on another molecule; it arises when H is attached to a very electronegative atom, which makes it very positive and very small
 - Br₂ has more electrons in molecule so stronger Van der Waal's forces between molecules
 - Water has hydrogen bonding between molecules but carbon dioxide does not

- Simple molecular substances tend to have low melting points, although there is a wide variation:
 - Van der Waal's forces and hydrogen bonds are weaker than ionic and metallic bonds, and so much less energy is required to break them
 - the strength of the intermolecular forces varies greatly from molecule to molecule, depending on the number of electrons in the molecule, the surface area of the molecule and whether there is a permanent dipole (Van der Waal's forces) and on whether there is hydrogen bonding between the molecules
 - The greater the number of electrons in a molecule, and the larger its surface area, the stronger the Van der Waal's forces between molecules will be, so more energy will be needed to separate the molecules and the boiling point will be higher
 - Eg Br₂ has a higher boiling point than Cl₂
 - Eg Ar has a higher boiling point than Ne
- Molecules which form hydrogen bonds, such as HF, H₂O and NH₃, have much higher boiling points than similarly sized molecules which do not have hydrogen bonding; this is most clearly seen in the boiling points of the hydrides of elements in groups V, VI and VII of the periodic table:
 - Group V: NH₃, PH₃, AsH₃, SbH₃ Group VI: H₂O, H₂S, H₂Se, H₂Te Group VII: HF, HCl, HBr, HI



- The general increase in boiling points of the hydrides on descending the group results from the greater number of electrons in the molecules, which causes stronger Van der Waal's forces
- The unusually high boiling point of the Period 2 hydrides (HF, H₂O and NH₃) results from the ability of these molecules to form hydrogen bonds with each other

UNIT 2 – PARTICLES, BONDING AND STRUCTURES



Test Your Knowledge 15.2: Explaining the boiling points of simple molecular structures:

- (a) Explain why I_2 is a solid at room temperature but Br_2 is a liquid
- (b) Explain why O_2 has a higher boiling point than N_2
- (c) Explain why HBr has a higher boiling point than HCl
- (d) Explain why HF has a higher boiling point than HCl
- (a) More electrons in one molecule of I_2 than in one molecule of Br_2 , so stronger Van der Waal's forces between molecules of I_2 , which are strong enough to hold I_2 molecules together as a solid at room temperature
- (b) More electrons in one molecule of O_2 than in one molecule of N_2 , so stronger Van der Waal's forces between molecules of O_2
- (c) More electrons in one molecule of HBr than in one molecule of HCl , so stronger Van der Waal's forces between molecules of O_2
- (d) Hydrogen bonding between HF molecules but not between HCl molecules, so overall intermolecular forces stronger between HF molecules, despite stronger Van der Waal's forces between HCl molecules

- Simple molecular substances do not conduct electricity:
 - there are no ions and no delocalised electrons in simple molecular substances, so these substances do not show electrical conductivity in either solid or liquid state
- Simple molecular substances tend to be soft and not very strong
 - the weak intermolecular bonds are easily broken, meaning it usually easy to break the structure apart; simple molecular structures tend to be soft and not very strong



Test Your Knowledge 15.3: Explaining the physical properties of simple molecular structures:

Copy and complete the following table about the physical properties of simple molecular structures:

Property	Explanation
Low melting point	
Non-conductor of electricity	

Property	Explanation
Low melting point	Only the weak intermolecular forces need to be broken Not a lot of energy is needed to break them
Non-conductor of electricity	There are no ions and no free electrons So no mobile charged particles



Thinkabout 15.4: The special properties of water

In addition to its high boiling point, water has a number of special properties which result from the strong hydrogen bonds between the molecules. Can you think of some of these properties? Can you explain why hydrogen bonding causes these special properties?

Strong hydrogen bonding causes strong surface tension in water, allowing some insects to walk on water
Strong hydrogen bonding causes ice to form an open lattice structure, causing it to be less dense than water

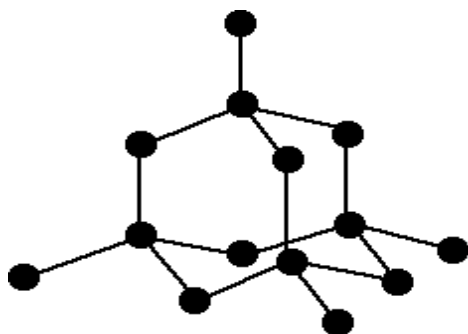
UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- The physical states of simple molecular structures can vary widely and the appropriate state symbol is often used to clarify the physical state in which the substance exists
 - $\text{H}_2\text{O} (\text{s})$ means ice
 - $\text{H}_2\text{O} (\text{g})$ means steam or water vapour
 - $\text{I}_2 (\text{s})$ means solid iodine
 - $\text{I}_2 (\text{g})$ means iodine vapour
 - $\text{I}_2 (\text{aq})$ means iodine dissolved in water

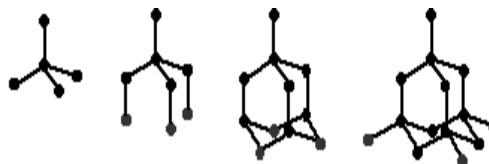
Lesson 16 – What are giant covalent structures?

(iv) Giant Covalent Structures

- In some structures involving covalent bonds, it is not possible to satisfy the bonding capacity of the atoms by forming a simple molecule; the covalent bonds between atoms therefore continue indefinitely, and a large lattice is formed, with covalent bonds existing between all adjacent atoms
- Such substances are called **giant covalent** (or **giant atomic** or **macromolecular**) substances, and the most important examples are C, B, Si and SiO_2 ; the giant structure can be two dimensional or three dimensional
- The most common examples of giant covalent structures are **diamond** and **graphite**; diamond and graphite are both **allotropes** of carbon:
 - In diamond, each C atom forms four covalent bonds with adjacent atoms
the bonds arrange themselves in a tetrahedral shape
the network of covalent bonds stretches throughout the crystal in three dimensions.

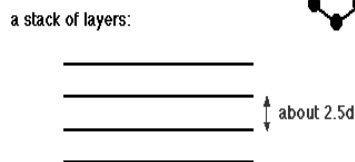
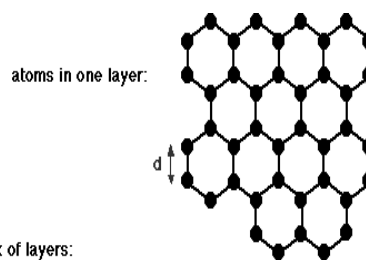
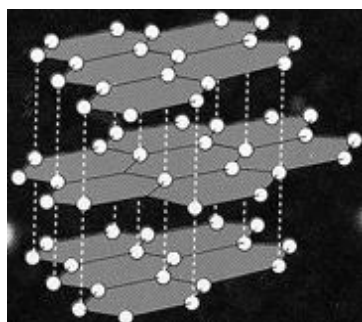


This is just a tiny part of a giant structure extending on all 3 dimensions.



- In graphite, each C atom forms three covalent bonds with adjacent atoms
- The bonds arrange themselves in a planar shape to form layers, which stretch throughout the crystal in two dimensions
- the spare electron is delocalized and occupies the space in between the layers
- all atoms in the same layer are held together by strong covalent bonds
- the different layers are held together by intermolecular forces

UNIT 2 – PARTICLES, BONDING AND STRUCTURES



or

- Giant covalent structures have a very high melting point:
 - covalent bonds must be broken before the lattice can break down and the atoms separated
 - this is true of both three-dimensional lattices (like diamond) and two-dimensional lattices (like graphite)
 - covalent bonds are usually very strong and breaking them requires a lot of energy

- Giant covalent structures are usually non-conductors of electricity:
 - giant covalent substances contain no ions
 - most also contain no delocalised electrons, so there is no electrical conductivity in either solid or liquid state
 - diamond and silicon dioxide (sand) do not conduct electricity
 - graphite is an important exception – the delocalised electrons in graphite make it a good conductor of electricity (so it is used in electrodes)

- The mechanical properties depend on whether the lattice is two-dimensional or three-dimensional:
 - in 3D lattices such as diamond, the atoms are held strongly in place in all three dimensions with no movement possible; these substances are hard and brittle (diamond is used in glass cutting)
 - in 2D lattices such as diamond, the layers are only held together by intermolecular forces and so the layers can slip over each other; 2D giant covalent substances are therefore usually soft (graphite is used in pencils and as a lubricant)



Test Your Knowledge 16.1: Explaining the physical properties of giant covalent structures:

Copy and complete the following table about the physical properties of giant covalent lattice structures:

Property	Reason
High melting point	
Non-conductor of electricity (except graphite)	
Hard if the lattice is 3D, soft if the lattice is 2D	

Property	Reason
High melting point	Strong covalent bonds between the atoms Require lots of energy to break
Non-conductor of electricity (except graphite)	No free electrons and no ions Graphite has delocalised electrons
Hard if the lattice is 3D, soft if the lattice is 2D	3D lattices do not have layers which can slid over each other but 2D lattice do

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- Giant covalent structures can be elements or compounds
 - giant covalent compounds, like ionic compounds, are represented by their **unit formula**; the unit formula of a giant covalent compound is the simplest whole number ratio of the atoms in the compound; the more electropositive atom is often, but not always, shown first

Name of compound	Formula	Meaning of formula
silicon dioxide	SiO ₂	Ratio of Si to O atoms is 1:2

- Giant covalent elements are represented simply by the atom:

Name of element	Formula	Meaning of formula
Boron	B	Only B atoms present

- Giant covalent substances are almost always found in the solid state; this is often used to distinguish between the element and the atom:
 - Si (s) means a solid sample of silicon
 - Si (g) means an individual atom of silicon

Test Your Knowledge 16.2: Understanding different types of chemical formula

- What is meant by the term “unit formula”? Which substances are generally represented by a unit formula? Give two examples of a unit formula.
- What is meant by the term “molecular formula”? Which substances are usually represented by a molecular formula? Give two examples of a molecular formula.
- Which types of element are represented just using the symbol of the atom?

- Simplest whole number ratio of ions (in an ionic compound) or atoms (in a giant covalent compound); it is used for ionic compounds and giant covalent compounds; eg KCl, Na₂CO₃, SiO₂
- Number of atoms of each type in one molecule of the substance; it is used for elements and compounds with simple molecular structures; eg Cl₂, CO₂, O₂, H₂O
- Elements with giant covalent, giant metallic or simple atomic structures (ie all elements except those with simple molecular structures)

Lesson 17 – Why do some substances dissolve in water but not others?

(v) Effect of structure on solubility

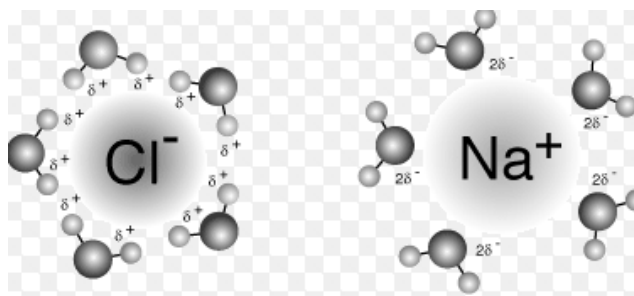


Thinkabout 17.1: Which substances dissolve in water and which do not?

- Can you think of any substances which dissolve in water?
 - Can you think of any substances which do not dissolve in water?
 - Can you think of any other solvents we use in everyday life?
-
- sodium chloride (salt), sucrose (sugar), vinegar all dissolve in water
 - most oils, calcium carbonate (chalk) and most rocks do not dissolve in water
 - many paints use ethanol or methanol or a mixture of the two (methylated spirits) as a solvent

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- Solvents are very useful in everyday life; it is often convenient to dissolve a substance in a solvent to make a solution of that substance:
 - Water is a very widely used solvent; many substances dissolve in it but many other things do not; water molecules are very polar, and water is therefore often described as a polar solvent
 - ethanol is another common solvent; it is also a polar solvent, although its molecules are less polar than those of water
 - Hexane and tetrachloromethane are also common solvents; the molecules in these substances are not polar and hexane and tetrachloromethane are therefore often described as non-polar solvents
 - Different substances dissolve in different types of solvent; some substances dissolve only in water; some dissolve only in non-polar solvents; some dissolve in both and some dissolve in neither
- When giant ionic structures are added to water, the cations are attracted to the negative dipole on a water molecule and anions are attracted to the positive dipole on a water molecule; sometimes this attraction is enough to break up the ionic lattice, and many (but not all) giant ionic structures are therefore soluble in water:



Source: www.quora.com

- sodium chloride dissolves in water; this reaction is represented $\text{NaCl(s)} \rightarrow \text{Na}^{\text{+}}(\text{aq}) + \text{Cl}^{\text{-}}(\text{aq})$; this means that the attraction between the ions and the water molecules is strong enough to break up the giant ionic lattice in NaCl
- some ionic compounds, such as AgCl, do not dissolve in water; in the case of AgCl the attraction between the ions and the water molecules is not strong enough to break up the giant ionic lattice in AgCl
- ionic compounds very rarely dissolve in solvents other than water; non-polar solvents are not able to attract anions and cations in the same way
- Non-polar simple molecular substances, such as iodine and methane, do not generally dissolve in water, but do dissolve in non-polar solvents such as hexane and ether:
 - the molecules in the solute can form Van der Waal's forces with the molecules in the solvent
 - if the solvent is also non-polar, the van der Waal's forces between the solute and the solvent will usually be of similar strength to those in the pure solute and solvent and the two substances will mix
 - if the solvent is water, the molecules in the solute cannot form hydrogen bonds with water, so they would disrupt the hydrogen bonding in the water if they dissolved in it, and water will generally reject the solute and the two substances will not mix
- Polar simple molecular substances, such as ammonia and sulphur dioxide, generally dissolve in water better than they do in non-polar solvents
 - these molecules can form permanent dipole interactions and sometimes hydrogen bonds with water
 - the hydrogen bonds between the solute and water will usually be of similar strength to those in the pure solute and water and therefore the solute is likely to be soluble in water
- Giant covalent structures do not contain ions and so are generally insoluble in water; the strong covalent bonds throughout the structure are hard to break and so these substances are generally insoluble in most solvents

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- Giant metallic structures contain delocalised electrons; this delocalised system would be disrupted by water and so most metals do not dissolve in water (although some react with water)



Test your knowledge 17.2: Predicting Solubility in water and in non-polar solvents

State, with a reason, whether the substances below are likely to be more soluble in water or in hexane:

(a) BH_3 ; (b) CH_4 ; (c) H_2 ; (d) HCl ; (e) NH_3 ; (f) HF ; (g) CO_2 ; (h) NaCl

BH_3 , CH_4 , H_2 , CO_2 non-polar so more soluble in hexane

HCl highly polar so more soluble in water

NH_3 , HF can form hydrogen bonds with water so more soluble in water

NaCl ionic so more soluble in water

Lesson 18 – How can we relate the physical properties of a substance to its structure?

(vi) Summary of different types of chemical structure and their properties



Demonstration 18.1: Compare the physical properties of different materials

Observe the melting point and electrical conductivity of materials A – F and use your observations to complete the following table:

Property	A copper	B water	C salt	D sugar	E sand	F graphite
Melting Point						
Electrical Conductivity						
Structure						

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Equipment needed: tongs, Bunsen burner, access to gas, three test tubes, three spatulas, three small beakers (100 cm³), two crocodile clips, two electrical wires, one small light bulb in socket connectible to wires, 12V power supply, strip of copper, 2 graphite electrodes, 1 graphite rod, a gram of sugar, a gram of sand, a gram of salt, access to water

- Hold the copper and graphite strips separately with tongs and place it in a medium Bunsen flame for one minute; nothing will happen
- Pour a small amount (1 cm depth) of sugar, sand and salt separately into test tubes, hold the test tubes with tongs and place each in a medium Bunsen flame; the sugar will quickly melt but the sand and salt will not
- Connect crocodile clips to each end of the copper strip and graphite rods, use a wire to connect one end of the strip to one side of the bulb, and use two more wires to connect the other end of the strip and the other side of the bulb to a 12V power supply (set the power lower than this); in both cases the bulb will light and shine brightly
- Place the graphite electrodes 5 cm apart in a beaker half filled with water with the circuit otherwise complete and the power supply set at 12V; the bulb will not light;
- Add a small spatula measure of sand to the water and stir gently; the sand will not dissolve and the bulb will not light
- Repeat, but adding sugar; the sugar will dissolve in the water but the bulb will not light
- Repeat, but adding salt; the salt will dissolve in the water and the bulb will light

Property	A copper	B water	C salt	D sugar	E sand	F graphite
Melting Point	high	low	high	low	high	high
Electrical Conductivity	yes	no	In aqueous solution only	no	no	Yes
Structure	Giant metallic	Simple molecular	Giant ionic	Simple molecular	Giant covalent	Giant covalent

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- The basic physical properties of all substances (melting point, electrical conductivity, malleability, solubility in water etc) can be explained by their structures; a summary of the relationship between physical properties and structure can be shown below:

Bond Type	Type of Structure and Description	Physical properties
IONIC	<p>GIANT IONIC STRUCTURE (eg NaCl) Attraction between oppositely charged ions. Infinite lattice of oppositely charged ions in three dimensions</p>	<p>High mpt, bpt Usually dissolve in water Rarely dissolve in non-polar solvents Good conductors in liquid and aqueous states Poor conductors in solid state Hard and brittle</p>
METALLIC	<p>GIANT METALLIC STRUCTURE (eg NaCl) Attraction between cations and delocalised electrons. Infinite lattice of cations in three dimensions, with delocalized electrons in the spaces</p>	<p>High mpt, bpt Good conductors in solid state Good conductors in liquid state Malleable and ductile Generally insoluble in solvents</p>
COVALENT	<p>Either SIMPLE MOLECULAR STRUCTURE (Eg I₂) Discrete molecules. Atoms in molecule held together by covalent bonds. Weak intermolecular forces between molecules.</p>	<p>Low mpt, bpt Poor conductors in solid state Poor conductors in liquid state Soft and weak If polar, soluble in water and non-polar solvents If non-polar, soluble in non-polar solvents only</p>
	<p>Or GIANT COVALENT STRUCTURE (eg diamond and graphite) Infinite lattice of atoms linked by covalent bonds in two or three dimensions.</p>	<p>Very high mpt, bpt Poor conductors in solid state (unless delocalised electrons present) Poor conductors in liquid state Hard and brittle Generally insoluble in solvents</p>



Test your knowledge 18.2: Summarising the relationship between physical properties and structure

Copy and complete the following table about the physical properties of different structures:

Structure	Property	Reason
Giant ionic lattice	Melting point:	
	Electrical Conductivity:	
	Mechanical properties:	
Giant metallic lattice	Melting point:	
	Electrical Conductivity:	
	Mechanical properties:	
Simple molecular	Melting point:	
	Electrical Conductivity:	
	Mechanical properties:	
Giant covalent lattice (3D)	Melting point:	
	Electrical Conductivity:	
	Mechanical properties:	

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Structure	Property	Reason
Giant ionic lattice	Melting point: high	strong ionic bonds (attraction between oppositely charged ions) require lots of energy to break
	Electrical Conductivity: low in solid state, but high in molten or aqueous states	Ions cannot move in the solid state, but can move in molten or aqueous states
	Mechanical properties: hard and brittle	Layers of ions cannot slip over each other due to repulsion between ions with the same charge
Giant metallic lattice	Melting point: high	strong metallic bonds (attraction between cation and delocalised electrons) require lots of energy to break
	Electrical Conductivity: high	Delocalised electrons can move
	Mechanical properties: malleable	Layers of cations can slip over each other without affecting metallic bonds
Simple molecular	Melting point: low	Weak intermolecular forces (between molecules) do not require much energy to break
	Electrical Conductivity: low	No ions or free electrons
	Mechanical properties: soft and weak	Weak intermolecular forces can be easily broken
Giant covalent lattice (3D)	Melting point: high	Strong covalent bonds (attraction between shared electrons and nuclei) require lots of energy to break
	Electrical Conductivity: low	No ions or free electrons
	Mechanical properties: hard and brittle	Layers of atoms cannot slip over each other without breaking strong covalent bonds



Test your knowledge 18.3: Using structures to explain physical properties

- Name the four main chemical structures and give one example of each.
- Why does sodium chloride have a high melting point?
- Why does sodium chloride conduct electricity when molten but not when solid?
- Why does sodium chloride dissolve in water?
- Why do metals conduct electricity?
- Why are metals malleable?
- Why does magnesium have a high melting point?
- Why does chlorine have such a low boiling point?
- Why is ammonia more soluble in water than carbon dioxide?
- Why do diamond and graphite have such high melting points?
- Why does graphite conduct electricity but not diamond?
- Why is diamond hard and why is graphite soft?
- Why is diamond used in drills?
- Why is graphite used in pencils?

(a) giant ionic - NaCl; giant metallic - Mg; simple molecular - Cl₂; giant covalent - SiO₂; (b) strong ionic bonds require lots of energy to break; (c) ions can move when molten but not when solid; (d) Na⁺ and Cl⁻ ions attracted to polar water molecules; (e) delocalised electrons can move; (f) layers of cations can slip over each other without breaking the metallic bonds; (g) strong metallic bonds require lots of energy to break; (h) intermolecular forces between Cl₂ molecules don't require much energy to break; (i) strong covalent bonds require lots of energy to break; (j) graphite has delocalised electrons, diamond does not; (l) layers of atoms in diamond cannot slip over each other without breaking covalent bonds; layers of atoms in graphite can slip over each other - only intermolecular forces need to be overcome; (m) it is very hard; (n) it is very soft and the layers can slip off

Lesson 19 – How does structure and bonding in the elements change across the Periodic Table?

i) Structure and Bonding in the Periodic Table



Summary Activity 19.1: Periodic properties of atoms

- State and explain how atomic size changes across a Period
- State and explain how atomic size changes down a Group
- State and explain how electronegativity changes across a Period
- State and explain how electronegativity changes down a Group

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- Atomic size decreases across a Period; nuclear charge increases, shielding stays the same, no nuclear attraction increases and outer shell electrons are pulled in more strongly
- Atomic size increases down a Group; number of shells increases so shielding increases; nuclear attraction decreases (despite increase in nuclear charge) and outer shell electrons are pulled in less strongly
- Electronegativity increases across a Period; nuclear attraction increases so attraction to bonding electrons increases
- Electronegativity decreases down a Group; nuclear attraction decreases so attraction to bonding electrons decreases

- Atoms in the periodic table show **periodic** properties (trends in properties which repeat themselves across each period); these properties include:
 - a decrease in atomic and ionic size across a period, and an increase down a group
 - an increase in electronegativity across a period, and a decrease down a group

- Elements are substances made up of only one type of atom; for each type of atom, therefore, there is one corresponding element; elements in the Periodic Table show periodic properties just like the atoms do

(i) Structure and Bonding in Elements

- All the atoms in a single element have the same electronegativity, so the bond type cannot be ionic – it has to be either metallic or covalent
- Atoms with low electronegativities form metallic bonds with themselves; the resulting elements therefore have giant metallic structures (ie they are metals)
- Atoms with higher electronegativities form covalent bonds with themselves; the resulting elements therefore have either simple molecular structures or giant covalent structures (ie they are non-metals)
- Noble gases form no bonds with themselves; the resulting elements therefore have simple atomic structures (ie they are non-metals)
- A non-metal is an element which does not form a giant metallic structure
- As electronegativity increases across a period, **the elements therefore become less metallic across a period** (sodium and magnesium are metals but chlorine and argon are non-metals)
- As electronegativity decreases down a group, **the elements therefore become more metallic down a group** (carbon and silicon are non-metals but tin and lead are metals)

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- The periodic variation in structure, and hence metallic character, can be shown in the table below:

H (H ₂) simple molecular (non-metal)							He simple atomic (non-metal)
Li giant metallic (metal)	Be giant metallic (metal)	B giant covalent (non-metal)	C giant covalent (non-metal)	N (N ₂) simple molecular (non-metal)	O (O ₂) simple molecular (non-metal)	F (F ₂) simple molecular (non-metal)	Ne simple atomic (non-metal)
Na giant metallic (metal)	Mg giant metallic (metal)	Al giant metallic (metal)	Si giant covalent (non-metal)	P (P ₄) simple molecular (non-metal)	S (S ₈) simple molecular (non-metal)	Cl (Cl ₂) simple molecular (non-metal)	Ar simple atomic (non-metal)

- Lithium, beryllium, sodium, magnesium and aluminium have **giant metallic structures** (ie they are metals); they consist of an infinite lattice of cations held together by a sea of delocalised electrons; metallic bonding is typically strong and as a result metals tend to have fairly high melting points and boiling points; the melting points increase with increasing charge and decreasing size and thus increase across a period:

Element	Sodium	Magnesium	Aluminium
Mpt/°C	98	669	680

- Boron, carbon (both graphite and diamond) and silicon have **giant covalent structures** and are hence non-metals; silicon has a 3D tetrahedral structure identical to diamond and boron has a 2D layered structure similar to graphite; in giant covalent structures, strong covalent bonds link all atoms in the lattice and these elements have high melting points; with the exception of graphite, these elements do not conduct electricity as they have no free ions or electrons
- Hydrogen, nitrogen, oxygen, fluorine, phosphorus, sulphur and chlorine form simple molecular structures; there are strong, covalent bonds within each molecule but the different molecules are only held together by weak Van der Waal's forces; separating the molecules thus requires little energy and the melting and boiling points of simple molecular elements is relatively low
- Helium, neon and argon form simple atomic structures with weak Van der Waal's forces holding the atoms together; they have very low boiling points



Test your knowledge 19.2: Explaining the bonding and structure of the Period 2 elements

Complete the following table by drawing diagrams to show the structures of the elements in Period 2. Draw six atoms in each structure.

Element	Li	Be	B	C (diamond)	N	O	F	Ne
Bonding	metallic	metallic	covalent	covalent	covalent	covalent	covalent	-
Structure	giant metallic	giant metallic	giant covalent	giant covalent	simple molecular	simple molecular	simple molecular	simple atomic
Diagram								

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Element	Li	Be	B	C (diamond)	N	O	F	Ne
Bonding	metallic	metallic	covalent	covalent	covalent	covalent	covalent	-
Structure	giant metallic	giant metallic	giant covalent	giant covalent	simple molecular	simple molecular	simple molecular	simple atomic
Diagram								



Test your knowledge 19.3: Explaining the bonding and structure of the Period 1 and 3 elements

Complete the following table by drawing diagrams to show the structures of the elements in Periods 1 and 3. Draw at least six atoms in each structure. Represent phosphorus and sulphur as P₄ and S₈ respectively.

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Bonding	metallic	metallic	metallic	covalent	covalent	covalent	covalent	-
Structure	giant metallic	giant metallic	giant metallic	giant covalent	simple molecular	simple molecular	simple molecular	simple atomic
Diagram								

Element	H	He
Bonding	covalent	-
Structure	simple molecular	simple atomic
Diagram		

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Bonding	metallic	metallic	metallic	covalent	Covalent	covalent	covalent	-
Structure	giant metallic	giant metallic	giant metallic	giant covalent	simple molecular	simple molecular	simple molecular	simple atomic
Diagram								

Element	H	He
Bonding	covalent	-
Structure	simple molecular	simple atomic
Diagram		

Lesson 20 – How does structure and bonding in compounds change across the Periodic Table?

(ii) Bonding in Compounds – oxides and chlorides

- Atoms on the left of the Periodic Table have low electronegativities, and hence a large difference in electronegativity with oxygen and chlorine, both of which have high electronegativities; atoms on the right of the Periodic Table therefore tend to form ionic bonds with oxygen and chlorine, and the resulting compounds have **giant ionic** structures

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

- As you move from left to right in the Periodic Table, the electronegativities increase, so the difference in electronegativity between the atoms and oxygen or chlorine becomes smaller; so the atoms start to form covalent bonds with oxygen and chlorine, and the resulting compounds have **simple molecular** or **giant covalent** structures
- On moving across a Period from left to right, the type of bond formed with oxygen and chlorine becomes less ionic and more covalent in character

Type of structure	Examples of chlorides and oxides
Giant ionic	LiCl, NaCl, MgCl ₂ , Li ₂ O, Na ₂ O, BeO, Al ₂ O ₃
Simple molecular	HCl, H ₂ O, BeCl ₂ , AlCl ₃ , SiCl ₄ , PCl ₃ , SCl ₂ , NO, P ₄ O ₆
Giant covalent	SiO ₂



Test your knowledge 20.1: Describe the structure and bonding in oxides and chlorides

Copy and complete the following table to show the formulae and structure of the oxides and chlorides of the first 18 elements.

Giant ionic lattice	Giant covalent lattice	Simple molecular
	SiO ₂	NO, P ₄ O ₆ , SO ₂

Giant ionic lattice	Giant covalent lattice	Simple molecular
LiCl, Li ₂ O, BeO, Na ₂ O, NaCl, MgCl ₂ , MgO, Al ₂ O ₃	SiO ₂	NO, P ₄ O ₆ , SO ₂ , HCl, H ₂ O, BeCl ₂ , B ₂ O ₃ , BCl ₃ , CO ₂ , CCl ₄ , NCl ₃ , Cl ₂ O, OF ₂ , ClF, AlCl ₃ , SiCl ₄ , PCl ₃ , SCl ₂



Test your knowledge 20.2: Summarising structure and bonding in the Periodic Table

- Why do the elements become less metallic across a Period?
- From the first 18 elements in the Periodic Table, name five metals and five non-metals.
- Give the formulae of three chlorides of the first 18 elements which have giant ionic structures.
- Give the formulae of three oxides of the first 18 elements which have simple molecular structures.
- Explain why sodium chloride has a high melting point and conducts electricity when molten, but aluminium chloride has a low melting point and does not conduct electricity when molten.
- Explain why aluminium oxide and silicon dioxide both have high melting points, but only aluminium oxide conducts electricity when molten.

- Electronegativity increases so less tendency to form metallic bonds
- Metals - lithium, beryllium, sodium, magnesium, aluminium; Non-metals: any of the other 13!
- LiCl, NaCl and MgCl₂
- Any oxide in the simple molecular column of LA 2.16
- Sodium chloride is giant ionic but aluminium chloride is simple molecular
- Aluminium oxide is giant ionic but silicon dioxide is giant covalent

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

Lesson 21 – What have I learned about Particles, Bonding and Structure?

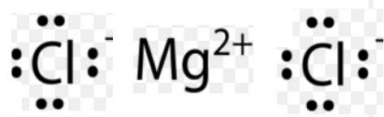
21.1 END-OF-TOPIC QUIZ

UNIT 2 – PARTICLES, BONDING AND STRUCTURES

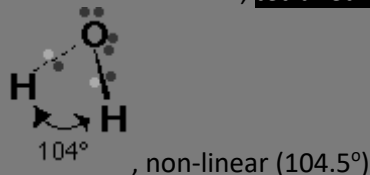
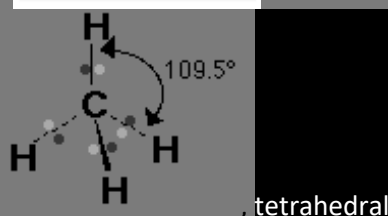
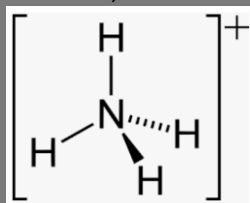


1. Draw a Lewis-dot structure to show the bonding in magnesium chloride (MgCl_2).
2. Draw a Lewis-dot structure to show the bonding in ammonia (NH_3).
3. Explain why the bonding in magnesium chloride is ionic but the bonding in ammonia is covalent.
4. Explain how an ammonia molecule can bond with a hydrogen ion and draw the structure of the species formed.
5. Draw a molecule of CH_4 ; state its shape and give the angle between the bonds.
6. Draw a molecule of H_2O ; state its shape and give the angle between the bonds.
7. Explain how a van der Waal's force is created between two molecules.
8. Explain why SO_2 is a polar molecule but CO_2 is not.
9. Describe the processes which take place when a solid is heated until it melts.
10. Explain what is meant by Brownian motion and explain how it provides evidence for the kinetic model of matter.
11. A tyre has a volume of 25 dm^3 when the temperature of the air inside is 285 K and the pressure is $2,000,000 \text{ Pa}$. Later that day, the temperature of the air in the tyre has increased to 400 K and the pressure of the tyre has increased to $2,400,000 \text{ Pa}$. What is the new volume of the tyre?
12. Describe the structure and bonding in magnesium. Explain why magnesium is a good conductor of electricity.
13. Describe the structure and bonding in diamond. Explain why diamond is hard and brittle.
14. Describe the structure and bonding in sodium chloride. Explain why sodium chloride is soluble in water.
15. Explain why the boiling point of H_2O is higher than that of H_2S .

UNIT 2 – PARTICLES, BONDING AND STRUCTURES



3. Difference in electronegativity between Mg and Cl is large; difference in electronegativity between N and H is much smaller
4. N on NH₃ has a lone pair; it can form a dative covalent bond with H⁺ by providing both of the shared electrons; the resulting species is NH₄⁺



8. Random movement of electrons causes temporary dipole, which induces a dipole in the adjacent molecule; the resulting dipoles attract
9. SO₂ is non-linear so the dipoles on the two bonds do not cancel out; CO₂ is linear so the dipoles on the two bonds do cancel out
10. Particles vibrate about fixed positions in a solid; on heating, the particles get more energy so vibrate more and faster; eventually the particles have enough energy to break some of the bonds between the particles and so the particles are able to move around
11. Brownian motion is the random movement in different directions of visible particles such as smoke, pollen or sulphur; it shows that the particles are constantly being hit by smaller particles and pushed in different directions
12. $P_1V_1/T_1 = P_2V_2/T_2$ so $V_2 = P_1V_1T_2/P_2T_1 = 29.2 \text{ dm}^3$
13. 3D lattice of Mg²⁺ cations, held together by a sea of delocalised electrons; the delocalised electrons are free to move
14. 3D lattice of C atoms, each attached to four others in a tetrahedral arrangement; all atoms are held in place by strong bonds and the atoms cannot move without breaking the bonds, so the structure is hard to break and cannot be bent
15. 3D lattice of Na⁺ and Cl⁻ ions; each Na⁺ surrounded by 6 Cl⁻ and vice versa; water is a polar molecule; the Na⁺ ions are attracted to the negative pole and the Cl⁻ ions are attracted to the positive pole
16. Water can form hydrogen bonds, H₂S can only form Van der Waal's forces, hydrogen bonds are generally stronger