Beckfoot

Subject: Chemistry	Topic: Bonding 3.1.3	Year Group: 12

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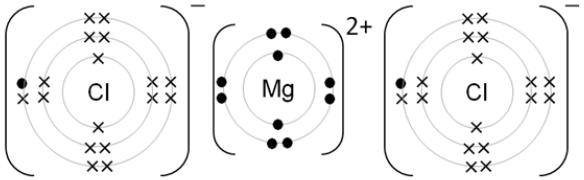
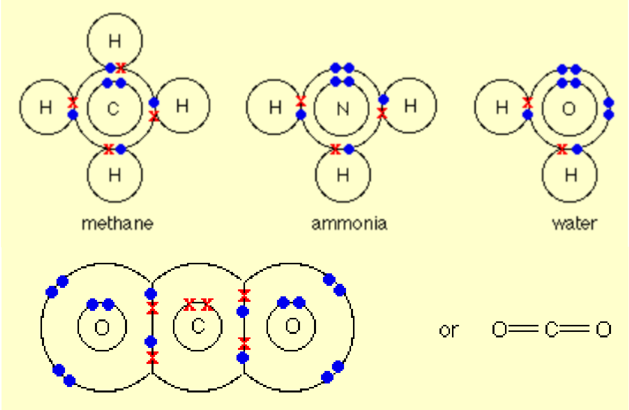
Ionic Bonding

1	Ionic bonding occurs because a metal atom loses electrons to become positively charges and a non metal gains electrons to become negatively charged. The electrostatic attraction between the oppositely charged ions is an ionic bond.
2	A metal loses electrons and a non metal gains electrons. **sodium atom** Na $1s^2\ 2s^2\ 2p^6\ 3s^1$ → **sodium ion** Na^+ $1s^2\ 2s^2\ 2p^6$ **chlorine atom** Cl $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ → **chloride ion** Cl^- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
3	In $MgCl_2$ the Mg loses 2 electrons and each chlorine gains 1 electron. 
4	An ionic solid has the ions arranged in a 3D structure called an ionic lattice. In the lattice for NaCl, each Na+ is surrounded by 6 Cl– ions and each Cl-is surrounded by 6 Na+.

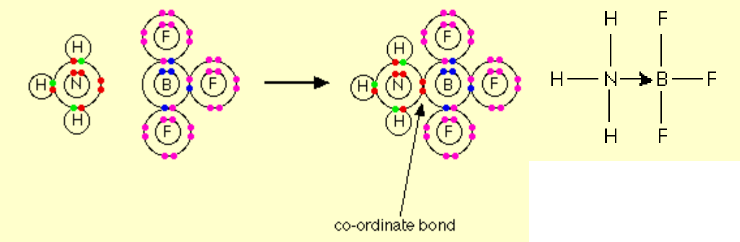
Metallic Bonding

1	Metal elements have a giant metallic lattice structure where the positive metal ions are attracted to the delocalised electrons.

Covalent bonding

1	Molecules are held together by strong covalent bonds which can be single, double or triple
2	In a covalent bond electrons are shared so they both have a full outer shell of electrons. 

Dative Covalent bonding

1	In a dative covalent bond or a coordinate bond, both electrons in the pair come from one atom. The bond is represented by an arrow pointing away from the donor atom. 

Key Vocabulary

1	Ion	A charged particle formed when one or more electrons are lost or gained by an atom or molecule.
2	Cation	Positively charged ion
3	Anion	Negatively charged ion
4	Lattice	Regular repeated 3D arrangement of atoms, ions or molecules in a metal or other crystalline solid.
5	Electronegativity	The ability to attract the bonding pair of electrons in a covalent bond.
6	Dipole	Difference in charge between two atoms caused by a shift in electron density.
7	Lone pair	An unshared pair of electrons in the outer shell of an atom.

Molecular Ions

1 Some ionic compounds contain molecular ions that contain covalent bonds but the compound is still ionic and will exhibit ionic properties.

Ion	Formula
Sulfate	SO_4^{2-}
Nitrate	NO_3^-
Hydroxide	OH^-

Ion	Formula
carbonate	CO_3^{2-}
hydrogen carbonate	HCO_3^-
ammonium	NH_4^+

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Bond Polarity

1	Electronegativity is measured on the Pauling scale. A higher number means the atom can better attract the bonding pair of electrons. F, O, N and Cl are very electronegative.
2	If two atoms in a covalent bond have different electronegativities, the bonding electrons are pulled towards the most electronegative atom making this atom delta (slightly) negative and the other atom delta positive. This is a polar bond, which causes a dipole.
3	A polar molecule will have a permanent dipole which is caused by an uneven distribution of charge over the whole molecule. If the polar bonds in a molecule are arranged symmetrically the dipoles cancel out so the molecule wont be polar overall.

Forces Between Molecules

1	Intermolecular forces are much weaker than bonds and form between molecules.
2	Van der Waals forces (or induced dipole-dipole) are the weakest intermolecular force and cause all atoms and molecules to be attracted to each other. This is because electron clouds are always moving so in any particular moment the electrons are likely to be more on one side than the other making it slightly negative. The atom will have a temporary dipole and induces a dipole in the opposite direction on a neighboring atom. The two dipoles are attracted to each other.
3	Permanent dipole-dipole forces occur in a substance made up of molecules that have permanent dipoles. This is because there are weak electrostatic forces of attraction between the delta positive and delta negative charges on neighboring molecules.
4	Hydrogen bonding is the strongest intermolecular force and only happens when hydrogen is covalently bonded to F, N or O because they are very electronegative so draw the electrons away from the hydrogen. The hydrogen then forms weak bonds with lone pairs on the F, N or O of other molecules.

Bonding and Physical Properties

1 In exam questions you will be asked to compare melting and boiling points of a range of substances and you need to be able to use the structure and bonding to explain this.

2

Type of crystalline substance	**Metals**	**Ionic Compounds**	**Molecular (simple covalent)**	**Macromolecular (giant covalent)**
Common examples	Magnesium	Sodium chloride	Ice, iodine	Diamond, graphite
Bonding	Metallic	Ionic	Covalent within the molecules and IMF between the molecules	Covalent (graphite has weak bonds between the layers in its structure)
Electrical conductivity	Conducts because delocalised electrons carry the charge.	Conducts when molten or dissolved in water because the ions are free to move and can carry the charge.	Does not conduct	Diamond does not conduct but graphite does because delocalised electrons carry the charge.
Melting point	High due to strong metallic bond.	High due to strong electrostatic attraction between the ions.	Low due to weak forces between molecules.	High due to strong covalent bonds.
Solubility in water	Insoluble	Generally soluble	Mostly insoluble (some polar substances will dissolve in water)	Insoluble

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Finding the Number of Electron Pairs

1	To work out the shape you need to know how many bonding pairs and lone pairs are around the central atom.
2	Follow these steps: 1. Find the central atom and work out how many electrons are in the outer shell. 2. Add 1 electron for every atom that the central atom is bonded to. 3. If an ion, add 1 electon for each negative charge or subtract 1 for each positive charge. 4. Divide the total by 2 to give the number of electron pairs. 5. Compare the number of electron pairs to the number of bonds to find the number of bonding pairs and lone pairs on the central atom.
3	When drawing shapes, use wedges to show the bond pointing towards you and a dotted line to show a bond pointing away from you.

Shapes of Molecules

1

Total number of electron pairs around central atom	Number of bonding pairs of electrons	Number of lone pairs of electrons	Shape	Bond angle	Examples
2	2	0	linear	180°	$BeCl_2$, CO_2
3	3	0	trigonal planar	120°	BF_3
4	4	0	tetrahedral	109.5°	CH_4, NH_4^+
4	3	1	pyramidal	107°	NH_3, H_3O^+
4	2	2	bent	104.5°	H_2O, BrF_2^+
5	5	0	trigonal bipyramid	90° and 120°	PF_5
5	3	2	T-shaped	86°	BrF_3
5	2	3	linear	180°	XeF_2
6	6	0	octahedral	90°	SF_6
6	4	2	square planar	90°	BrF_4^-

Electron Pair Repulsion

1	Electrons are negatively charged so charge clouds repel each other until they are as far apart as possible.
2	Lone pair charge clouds repel more than bonding pair charge clouds which means bonds angles are reduced because bonding pairs are pushed together by lone pair repulsion.