Name:

Chemical Bonding	Objectives
5. Chemical Bonding: Chemical Formulas	-understand that compounds can be represented by chemical formulas -relate the stability of noble gasses to their electron configurations -describe bonding and valency in terms of the attainment of a stable electronic structure -state the octet rule -explain its limitations -use the octet rule to predict the formulas of simple binary compounds of the first 36 elements (excluding d-block elements) and the hydroxides, carbonates, **nitrates, hydrogencarbonates, sulfites and sulfates** of these elements (where such exist). **-recognise that Cu, Fe, Cr and Mn have variable valencies** -relate the uses of helium and argon to their chemical unreactivity
Ionic Bonding	-define ion, positive ion, negative ion -appreciate the minute size of ions -explain ionic bonding in terms of electron transfer -represent ionic bonds using dot and cross diagrams -describe the structure of a sodium chloride crystal having reviewed models -associate ionic substances with their characteristics -outline two uses of ionic materials in everyday life
Covalent Bonding	-define molecule -appreciate the minute size of molecules -explain covalent bonding in terms of the sharing of pairs of electrons (Single, double and triple covalent bonds) -represent covalent bonds in molecules using dot and cross diagrams **-distinguish between sigma and pi bonding** -distinguish between polar and non-polar covalent bonding -test a liquid for polarity using a charged plastic rod -give examples of polar and non-polar materials in everyday life (two examples in each case) -associate covalent substances with their characteristics -test the solubility of ionic and covalent substances in different solvents
Electronegativity	-define electronegativity -recognise the trends in electronegativity values down a group and across a period -explain the general trends in electronegativity values • down a group • across a period. -relate differences in electronegativity to polarity of bonds -predict bond type using electronegativity differences
Shapes of Molecules and Intermolecular Forces	-describe the shapes of simple molecules -use appropriate modeling techniques to illustrate molecular shape **-explain the basis for electron pair repulsion theory use electron pair repulsion theory to explain the shapes of molecules of type AB_n for up to four pairs of electrons around**

	the central atom refer to bond angles (Shapes of molecules with pi bonds not to be considered) **-explain the relationship between symmetry and polarity in a molecule (dipole moments not required)** **-describe and distinguish between intramolecular bonding and intermolecular forces (van der Waals', dipole-dipole, hydrogen bonding)** **-describe the effects of intermolecular forces on the boiling point of covalent substances** **-relate the differences in boiling points of H_2 and O_2 , C_2H_2 and HCHO and of H_2O and H_2S to the effect of intermolecular forces**

Def^n: A **compound** is a substance that is made up of two or more different elements combined together chemically.

The Octet Rule: When bonding occurs, atoms tend to reach an electron arrangement with eight electron in the outermost energy level.

Exceptions to the octet rule:

1. Transition Metals usually do not obey the octet rule.
2. Hydrogen, Lithium and Beryllium tend to achieve two electrons in the outermost energy level instead of eight.

Ionic Bonding:

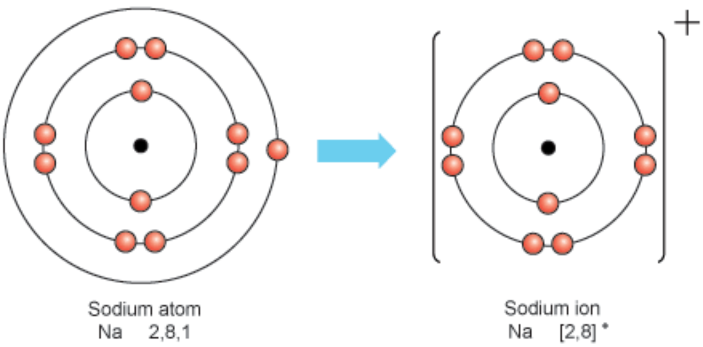
Def^n: An **ion** is a charged atom or group of atoms.

Elements in Group I of the Periodic Table form ions with a charge of +1 as they lose their 1 outermost electron.

Elements in Group II of the Periodic Table form ions with a charge of +2 as they lose their 2 outermost electrons.

Elements in Group III of the Periodic Table form ions with a charge of +3 as they lose their 3 outermost electrons.

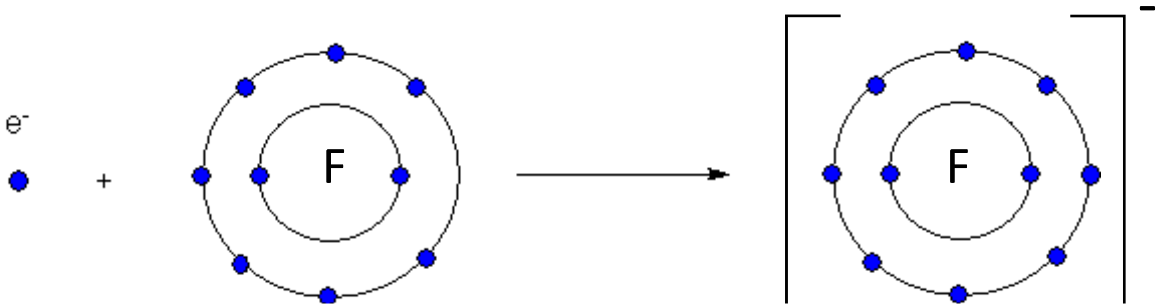
These positive ions are called **cations**.



Elements in Group VI of the Periodic Table form ions with a charge of -2 as they gain 2 electrons.

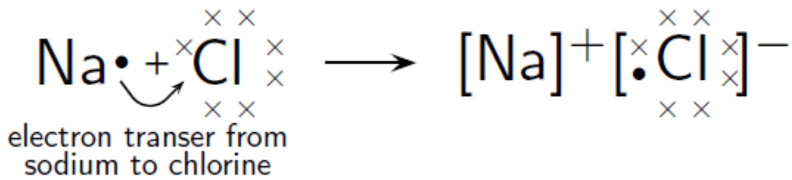
Elements in Group VII of the Periodic Table form ions with a charge of -1 as they gain 1 electron.

These negative ions are called **anions**.



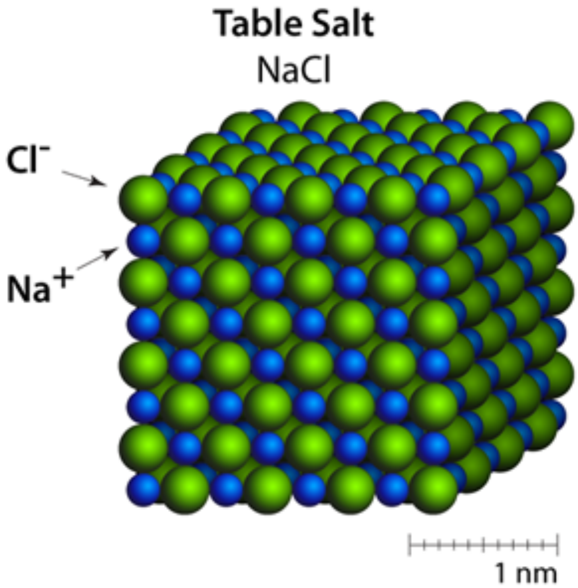
Def^n: An ionic **bond** is the force of attraction between oppositely charged ions in a compound. Ionic bonds are always formed by the complete transfer of electrons from one atom to another.

To show the ionic bonding in a compound, use a **dot-and-cross diagram**:



Note: The charges in an ionic compound always cancel each other out to be neutral overall.

Crystal Structure of Sodium Chloride:



Notes:

The three-dimensional arrangement of ions is called a **crystal lattice**.

Each sodium ion is surrounded by 6 chloride ions.

Each chloride ion is surrounded by 6 sodium ions.

Writing Formulas of Ionic Compounds:

Note: In the formula and the name of any ionic compound, the *positive* ion goes *first*.

Example: Write the formula for the compound Aluminium Oxide.

This compound, from its name, is made up of Aluminium, Al, and Oxygen, O.

Al is in Group III, so it forms a +3 ion. O is in Group VI so it forms a -2 ion.

The LCM of 3 and 2 is 6. We need to have a +6 charge and a -6 charge so that the compound is neutral overall.

So we need 2 Al^{3+} and we need 3 O^{2-}.

Our formula is $\mathbf{Al_2O_3}$.

Complex Ions (Group Ions)

A complex ion is an ion made up of two or more different atoms.

There are a number of complex ions whose *names, formulas and charges* you need to know by heart:

Name	Formula	Charge
Hydroxide ion Nitrate ion Hydrogencarbonate ion Permanganate ion	OH^- NO_3^- HCO_3^- MnO_4^-	One Negative Charge
Carbonate ion Chromate ion Dichromate ion Sulphate ion Sulphite ion Thiosulphate ion	CO_3^{2-} CrO_4^{2-} $Cr_2O_7^{2-}$ SO_4^{2-} SO_3^{2-} $S_2O_3^{2-}$	Two Negative Charges
Phosphate ion	PO_4^{3-}	3 Negative Charges
Ammonium ion	NH_4^+	1 Positive Charge

Example: Write the chemical formula for the ionic compound Calcium Hydrogencarbonate.

This compound is made up of Calcium, Ca, and the complex ion, Hydrogencarbonate, HCO_3^-.

Calcium is in Group II, so it forms a +2 ion. Hydrogencarbonate is a -1 complex ion.

The LCM of 2 and 1 is 2. We need to have a +2 charge and a -2 charge to be neutral overall.

So, we need only one Ca^{2+} ion. We need two HCO_3^- ions, which we write like $(HCO_3)_2$.

Our formula is $\mathbf{Ca(HCO_3)_2}$.

Ionic bonding with the Transition Metals:

Transition metals are the elements found in the d-block (except for Scandium and Zinc). They are special for a number of reasons:

1. Transition metals are used as catalysts.
2. Each transition metal can make different ions with different charges. For instance, Mn (Manganese) can form ions with a charge of either +2, +4 or +7. We show which charge the ion has by using Roman Numerals in brackets after the name of the element. So an Mn^{7+} ion is called Manganese (VII).
3. Transition metals form coloured compounds. (One transition metal can form compounds of different colours depending on its charge – e.g. Mn^{7+} is purple, but Mn^{2+} is completely colourless.)

Example: Write the chemical formula for the ionic compound Iron (II) Carbonate.

We have Iron with a +2 charge, as told by the "(II)". So we have Fe^{2+}. We also have Carbonate, CO_3^{2-}, a -2 charged ion.

The LCM of 2 and 2 is 2, so we need a +2 and a -2 charge for the compound to be neutral.

So we need only one Fe^{2+} and only one CO_3^{2-}.

Our formula is $\mathbf{FeCO_3}$

Example: Name the compound $Cr_2(SO_4)_3$.

For this example we need to work out the charge on the $(SO_4)_3$ to find the charge on the Cr (which is a transition metal, so it needs a roman numeral in brackets in its name so that its charge is clearly known).

SO_4 is the sulphate ion, which has a charge of -2.

We have three sulphate ions so the total negative charge is -6.

Therefore our Cr_2 must have a +6 charge (for **two** Cr atoms). This means that each Cr ion has a charge of +3.

So the name of our compound is **Chromium (III) Sulphate.**

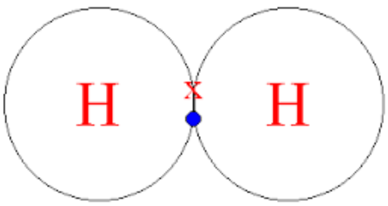
Covalent Bonding:

Covalent bonding involves electrons being *shared* between atoms, rather than electrons being fully transferred.

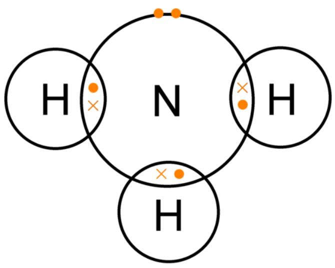
Def^n^: A **molecule** is a group of atoms joined together. It is the smallest particle of an element or compound that can exist independently.

We can show covalent bonding using dot-and-cross diagrams:

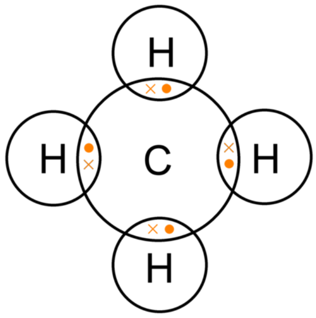
Hydrogen, H_2



Ammonia, NH_3



Methane, CH_4



Def^n^: The **valency** of an element is defined as the number of atoms of hydrogen which the element will combine with.

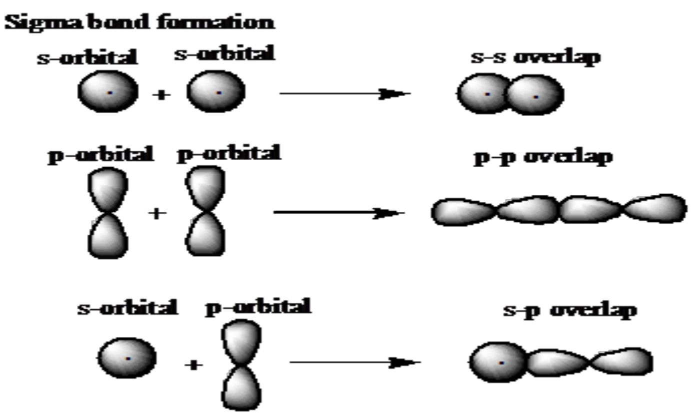
Example: Carbon belongs to Group 4 of the Periodic Table so it has 4 outer electrons. To achieve an outer shell of 8 electrons (Octet Rule) it would need to bond with 4 hydrogen atoms. Therefore Carbon's valency is 4.

Sigma (σ) and Pi (π) Bonding:

Covalent Bonding can be expressed in terms of overlapping atomic orbitals. There are two ways this can happen:

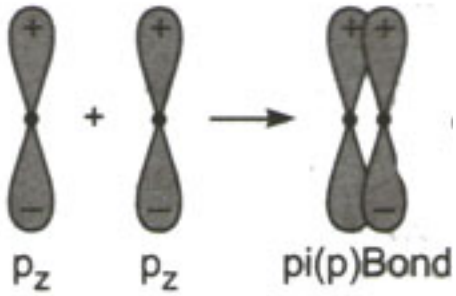
1. **Sigma (σ) Bonding:**

   Def^n^: A **sigma bond** is formed by the *head-on* overlap of atomic orbitals.



2. Pi (π) Bonding:

Def^n: A **pi bond** is formed by the *sideways* overlap of p orbitals.

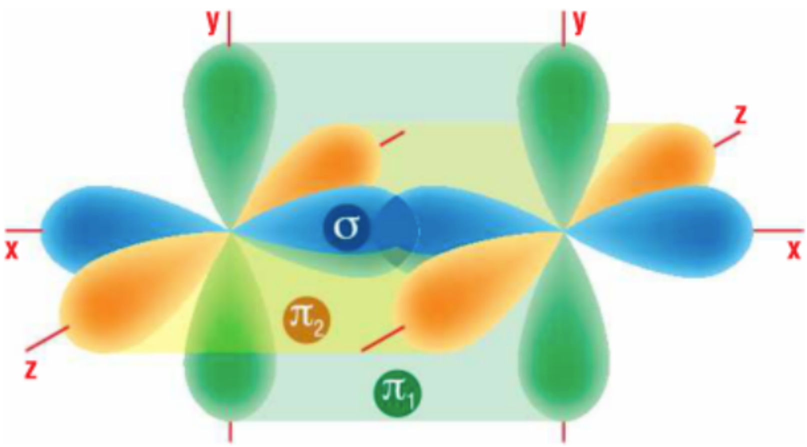


Rule of Thumb: A single bond always consists of a sigma bond.

A double bond always consists of one sigma and one pi bond.

A triple bond always consists of one sigma bond and two pi bonds.

Look at the triple bond in Nitrogen gas, N_2:



Different properties between ionic and covalent compounds:

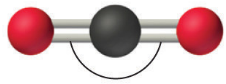
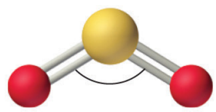
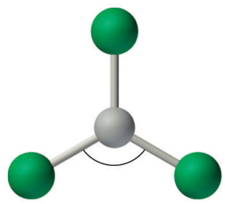
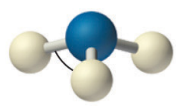
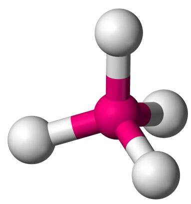
	Ionic	Covalent
1.	Contain a network of ions in the crystal	Contain individual molecules
2.	Ususally hard and brittle	Usually soft
3.	High melting and boiling points	Low melting and boiling points
4.	Ususally solid at room temperature	Usually liquids, gases, or soft solids at room temperature
5.	Conduct electricity when molten or dissolved in water	Do not conduct electricity

Shapes of Covalent Molecules:

The theory used to predict the shapes of covalent molecules is called VSEPR (Valence Shell Electron Pair Repulsion Theory).

How to use this theory:

1. Draw a dot-and-cross diagram of the molecule in question.
2. Count how many Bond Pairs of electrons are around the central atom of the molecule.
3. Count how many Lone Pairs of electrons are around the central atom of the molecule.
4. Use the number of Bond Pairs and Lone Pairs from parts 2. and 3. to find the shape using the table below:

Shape	No. of Bond Pairs	No. of Lone Pairs	Diagram	Bond Angles
Linear	2	0	CO_2 	180°
V-Shaped	2	2	SO_2 	104.5°
Triangular Planar	3	0	BCl_3 	120°
Pyramidal	3	1	NH_3 	107°
Tetrahedral	4	0		109.5°

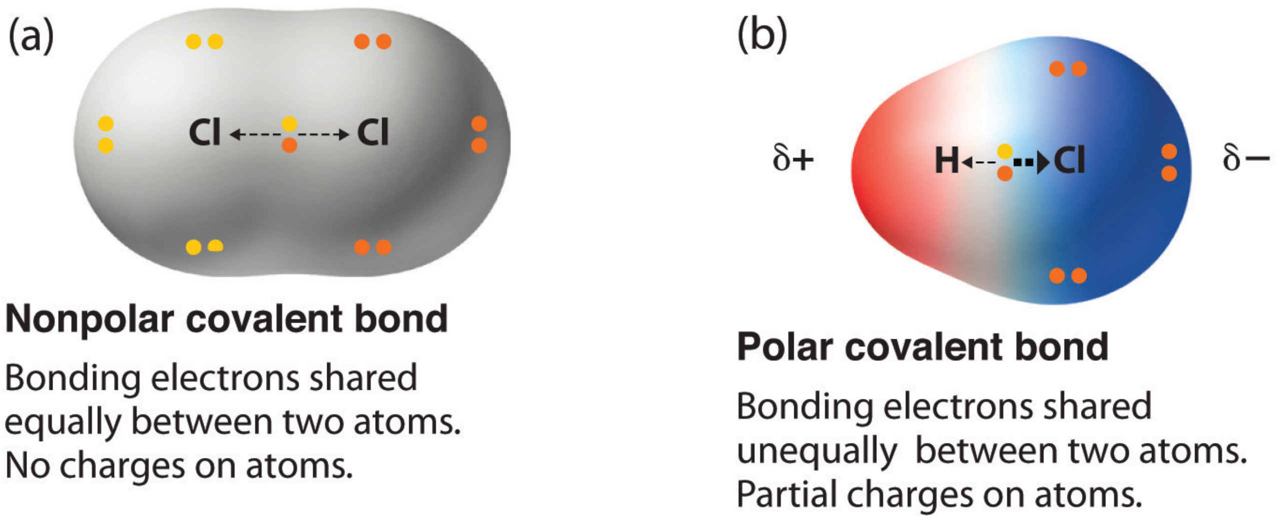
Note: This table needs to be known by heart – you will not have this supplied during the exam.

Electronegativity:

Def^n^: **Electronegativity** is the relative attraction that an atom in a molecule has for the shared pair of electrons in a covalent bond.

In covalent bonds, the electrons are not always equally shred. One atom in the molecule has a greater "pull" on the electrons in the bond – this atom with the greater "pull" is the more electronegative atom. This leads to having two main types of bond in a covalent molecule:

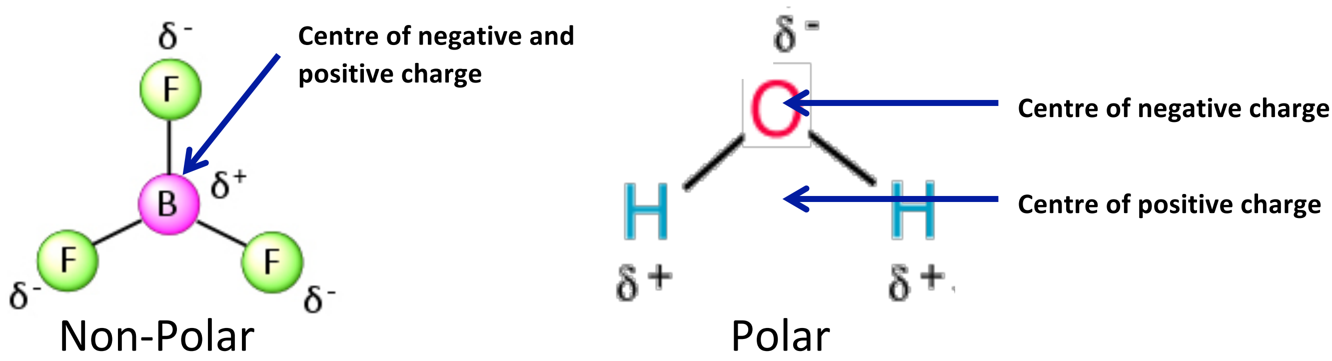
(a) Non-polar covalent: The atoms in the molecule all share electrons equally.
(b) Polar covalent: The atoms in the molecule do not share the electrons equally. This causes one end of the bond to be slightly positive (δ+) and the other ened to be slightly negative (δ-).



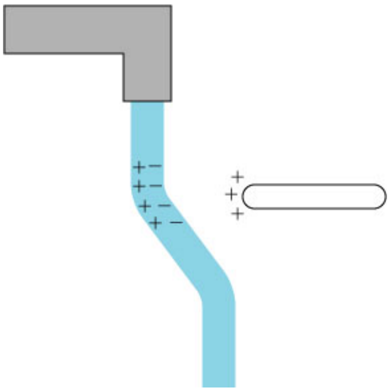
Even though many molecules have polar bonds, it *does not mean* that the molecule itself is polar.

To find out if a molecule is polar we need to:

1. Find the centre of positive charge.
2. Find the centre of negative charge.
3. If the centres from 1. and 2. are in exactly the same place, the molecule is non-polar (even though the bonds are polar). If the centres from 1. and 2. are not in exactly the same place, the molecule is polar overall.



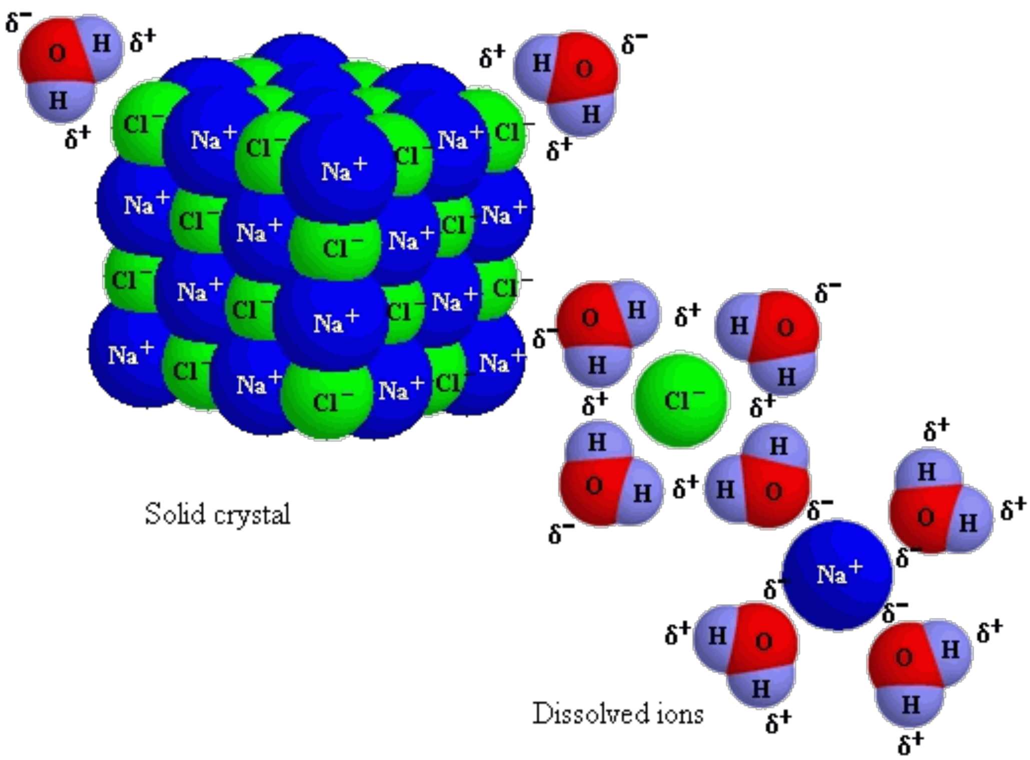
We can prove that water is polar, by placing a charged plastic rod near a thin stream of water (from a burette). We see that whether the plastic rod has a positive or negative charge, the water is always attracted to (and never repelled from) the rod.



Explanation: If the rod is positively charged, the water molecules will spin so that the negative end of the water molecule is facing the rod, causing an attraction. Likewise, if the rod is negatively charged, the water molecules will spin so that the positive end of the molecule will face the rod, causing attraction

Dissolving Ionic Compounds in Water:

We will look at the example of dissolving NaCl, Sodium Chloride (Table Salt) in water:



Ionic compounds are made up of positively and negatively charged ions.

When these ions are put in water, the water molecules arrange themselves around the ions so that the crystal is "pulled apart", or dissolved.

The negative (oxygen) end of the water will surround the positive ions (Na^+) in the crystal.

The postive (hydrogen) end of the water will surround the negative ions (Cl^-) in the crystal.

Using Electronegativity Values to Predict the Type of Bonding in a Molecule:

For this you need the Electronegativity Tables (Log Tables, Page 81). These table give us a number for each element, letting us know exactly how electronegative that element is.

To see what type of bonding is occuring in a molecule:

1. Find the electonegativity values (from the Log Tables) of the elements in the molecule.
2. Subtract the two values (if you get a negative answer, ignore the minus).
3. Compare your answer with the table on the right to get the type of bonding in the molecule.

Value	Bonding
0 - 0.4	Non-Polar Covalent
0.4 - 1.7	Polar Covalent
Over 1.7	Ionic

Example: What type of bonding occurs in a molecule of HCl (Hydrogen Chloride)?

From the tables, the Electronegativity Value for H is 2.20.

From the tables, the Electronegativity Value for Cl is 3.16.

3.16 – 2.2 = 0.96

Since 0.96 is between 0.4 and 1.7, the type of bonding in HCl is **Polar Covalent.**

<u>Intramolecular and Intermolecular Bonding:</u>

Intramolecular Bonding:

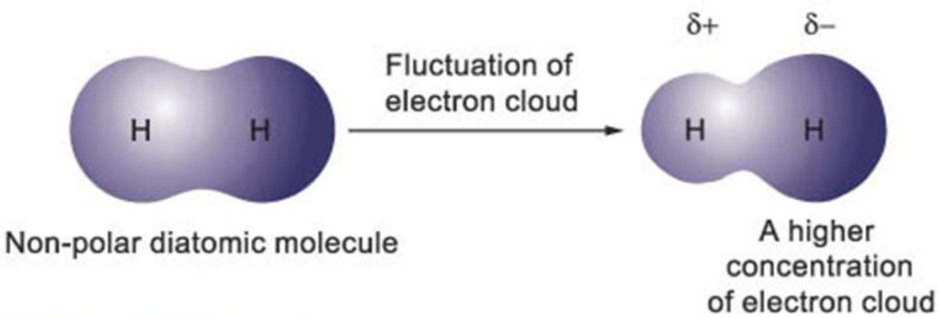
Def^n: **Intramolecular Bonding** is bonding that takes place *within* a molecule i.e. it holds the atoms in a molecule together. Covalent bonding is an example of intramolecular bonding.

Intermolecular Forces:

Def^n: **Intermolecular Forces** are the forces of attraction that exist *between* molecules. Van der Waals forces, dipole-dipole forces and hydrogen bonding are examples of intermolecuar forces.

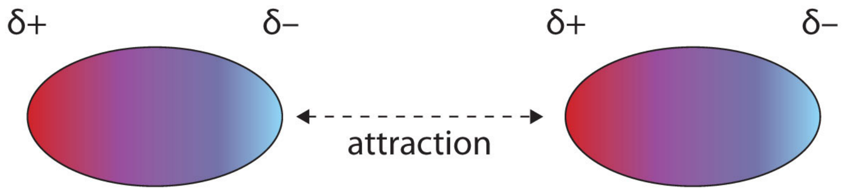
In detail, the three types of intermolecular forces are (from weakest to strongest):

1. **<u>Van der Waals Forces:</u>**
   Van der Waals forces occur between *non-polar* molecules.



We will look at Hydrogen, H_2, a non-polar molecule. The electrons around the hydrogen atoms are constantly moving around. This means that at one moment, the electrons may both be on one side of the molecule. This gives the molecule a slight negative charge (δ-) on the side with the electrons and leaves a slight positive charge (δ+) on the side with no electrons. This dipole (difference in charge) is only temporary as the electrons are constantly moving.

This formation of a temporary dipole can induce a similar dipole in a nearby molecule. There is now an attraction between the two molecules. This attraction is called Van der Waals force.



Def^n: **Van der Waals forces** are weak attractive forces between molecules resulting from the formation of temporary dipoles. They are the only forces of attraction between non-polar molecules.

Notes on Van der Waals Forces:

- Van der Waals forces get stronger for bigger molecules. This is because bigger molecules have more electrons and therefore have a bigger chance of making a temporary dipole (having more electrons on one side of the molecule than the other).
- The stronger the Van der Waals forces on a molecule, the higher its boiling point and melting point. E.g. H_2 has a boiling point of -253°C, while O_2 has a boiling point of -183°C. Even though both molecules are non-polar and both have Van der Waals forces between molecules, the fact that oxygen atoms are bigger than hydrogen atoms means that the Van der Waals forces for Oxygen are much stronger. This causes the molecules to be pulled together more strongly and have a higher boiling point.

2. Dipole-Dipole Forces:

Dipole-Dipole forces are very similar to Van der Waals forces, except that they occur between *polar* covalent molecules i.e. molecules with a permanent dipole.

Defⁿ: **Dipole-Dipole Forces** are forces of attraction between the negative pole of one polar molecule and the positive pole of another polar molecule.

In the diagram above we have two molecules of HCl, a polar compound. Cl is more electronegative than H so there is a slight negative charge (δ-) over the Cl, leaving a slight positive charge (δ+) over the H. The δ- of the Cl attracts the δ+ on the H of a nearby HCl molecule, causing an attraction. This attraction is the dipole-dipole force and is represented by the dotted line.

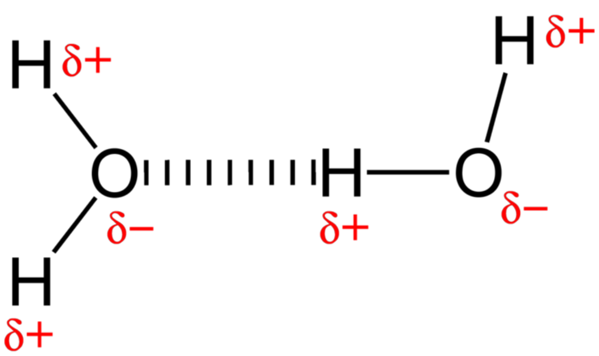
Note on Dipole-Dipole Forces:

- These forces are much stonger than Van der Waals forces. This causes polar molecules to have much higher boiling and melting points than similar non-polar molecules. E.g. the boiling point of HCl (-85°C) is much higher than that of H_2 (-253°C) as the stronger dipole-dipole forces of HCl need more heat energy to be broken (boil) compared to the extremely weak Van der Waals forces of H_2.

3. Hydrogen Bonding:

Def^n: **Hydrogen Bonds** are particular types of dipole-dipole attractions between molecules in which Hydrogen atoms are bonded to Nitrogen, Oxygen or Fluorine. The Hydrogen atom carries a partial positive charge and is attracted to the electronegative atom (N, O or F) in another molecule.

When H is bonded to either N, O or F, the partial charges are particularly strong as there is a big electronegativity difference between the atoms. This means that Hydrogen Bonding is the *strongest* form of intermolecular force.



In the diagram, we can see that there are intermolecular forces between the two water molecules. Because each water molecule has H bonded to O, we have hydrogen bonding, which is the strongest type of intermolecular force. That is why water has a particularly high boiling point (100°C), as these strong forces need to be broken by a lot of heat in order for the liquid water to boil. The dotted line represents the hydrogen bonding between the water molecules.

Important Note on Intermolecular Bonding:

Even though we talk of stonger and weaker forces (hydrogen bonding vs dipole-dipole, for example), in reality these forces are very weak – MUCH weaker than the covalent bonds holding the atoms in each molecule together.

Approach to Exam Questions:

You will often be given a list of compounds along with a list of boiling points. You are asked to match up each compound with the correct boiling point. The way to approach this is by identifying the types of intermolecular forces between the molecules. So identify which molecule has Van der Waals forces, which has Dipole-Dipole forces and which has Hydrogen Bonding. As you now have them ordered from weakest to strongest intermolecular forces, you can assign the boiling points (lowest to highest).

Dissolving Compounds in Water:

A good rule to remember when trying to make a solution up is "LIKE DISSOLVES LIKE". This means that a polar solute will only dissolve in a polar solvent, and a non-polar solute will only dissolve in a non-polar solvent.

As water is a highly polar molecule, only polar and ionic substances can dissolve in it.

The more polar the substance, the more easily it will dissolve in water.

If a compound has Hydrogen Bonding, it will dissolve even more easily in water.

Non-polar compounds will NOT dissolve in water. Non-polar solutes are usually dissolved in *cyclohexane*, a non-polar solvent.