Chemical bonding

Most of the materials around us are made from combinations of different elements.

Two or more different elements combined together is called a **compound.**

A compound is a substance that is made up of two or more different elements combined together chemically.

Attractive forces hold these elements together, known as **chemical bonds.**

We will learn the different type of chemical bonds and how to write the chemical formulas.

The octet rule

Noble gases are very unreactive and form practically no compounds.

Because of this they are called inert.

The reasons for noble gases being unreactive is due to the fact all of them (except Helium) have 8 electrons in their outer energy level.

The rule was called the octet rule (octet = eight)

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

Elements react together to try and have an electron arrangement with 8 electrons in the outer energy level.

Exceptions to the octet rule

Transition metals do not usually obey the octet rule.

Elements near helium (hydrogen, lithium and beryllium) try and achieve the same electron arrangement as helium.

Ionic bonding - Transfer of electrons

When an atom gains or loses an electron it becomes a charged atom.

This is referred to as an **ion.**

An ion is a charged atom or group of atoms.

Elements in Group 1 tend to lose one electron.

Group 2 tend to lose 2 electrons

These elements form **cations** (positive ions)

Example:

$$\underset{2,8,1}{Na} \longrightarrow \underset{2,8}{Na^+} + e^- \qquad \underset{2,8,8,2}{Ca} \longrightarrow \underset{2,8,8}{Ca^{2+}} + 2e^-$$

Elements in groups 16 & 17 tend to gain electrons.

These elements form **anions** (negative ions).

Example:

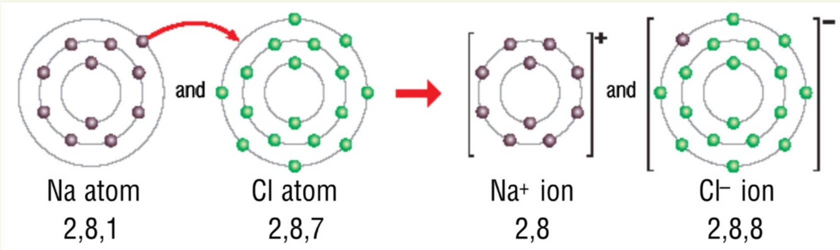
$$\underset{2,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-} \qquad \underset{2,6}{O} + 2e^- \longrightarrow \underset{2,8}{O^{2-}}$$

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.

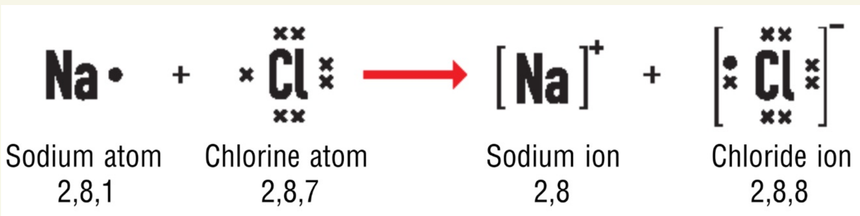
How to show the formation of ionic bonding

There are two ways of showing the formation of ionic bonding.

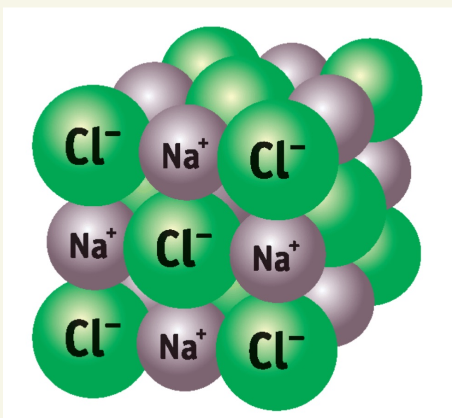
1. Bohr-type circle diagram



2. Dot-and-cross



Sodium chloride crystal structure



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

How to write the formulas of ionic compounds

For your examination you must know how to write the formulas of ionic compounds of the first 36 elements.

Ionic compounds are usually formed between **metals** (group 1 & 2) and **non-metals (**group 16 & 17).

Question

Write the formula of calcium chloride.

Answer

The calcium ion is Ca^{2+}. The chloride ion is Cl^-.
The formula $Ca^{2+}Cl^-$ is not correct since there are two positive charges but only one negative charge. To get two negative charges we must have two Cl^- ions.

Answer: The formula of calcium chloride is $CaCl_2$

Question

Write the fomula of sodium sulfide.

Answer

The sodium ion is Na^+. The sulfide ion is S^{2-}.
The formula Na^+S^{2-} is not correct since there is one positive charge but two negative charges. To get two positive charges we must have two Na^+ ions.

Answer: The formula of sodium sulfide is Na_2S.

Writing formulas of compounds with group ions

It is not possible to predict the formulas of group Iona from the periodic table, so the following table must be learned off.

Name	Formula	
Hydroxide ion	OH^-	One negative charge
Nitrate ion	NO_3^-	
Hydrogencarbonate ion	HCO_3^-	
Permanganate ion	MnO_4^-	
Carbonate ion	CO_3^{2-}	Two negative charges
Chromate ion	CrO_4^{2-}	
Dichromate ion	$Cr_2O_7^{2-}$	
Sulfate ion	SO_4^{2-}	
Sulfite ion	SO_3^{2-}	
Thiosulfate ion	$S_2O_3^{2-}$	
Phosphate ion	PO_4^{3-}	Three negative charges
Ammonium ion	NH_4^+	One positive charge

Writing formulas of compounds containing transition metals

The d-block elements, some of which are transition metals. Have a **variable valency.**

Iron combines with chlorine to form either FeCl2 or Fe Cl3.

These compounds are **ionic** compounds.

The charge on iron in FeCl2 is +2.

This is why we call FeCl2 **iron (II) chloride.**

The charge on iron in FeCl3 is +3

So it is called **iron (III) chloride.**

Copper combines with oxygen to form Cu2O or CuO.

Cu2O is called **copper (I) oxide**

CuO is called **copper (II) oxide**

Transition metals exhibit variable valency because of the small energy difference between the 4s and 3d sublevels.

Compounds that end in **-ide** contains just two elements

Compounds that end in **-ate** contain oxygen as well as two other elements.

d-Block Elements and Transition Elements

Transition metals have variable valency

Transition metals usually form coloured compounds

Transition metals are widely used as catalysts

A transition metal is one that forms at least one ion with a partially filled d sublevel

Covalent bonding - sharing of electrons

Atoms that achieve noble gas configurations by sharing electrons

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

Since hydrogen atom cannot combine with more than one atom of any other element, we say **it has a valency of one.**

The valency of an element is defined as the number of atoms of hydrogen or any other monovalent element with which each atom of the element combines.

Since chlorine combines with one atom of hydrogen (HCl), we say it has a valency of one.

Since oxygen combines with two atoms of hydrogen (H2O) we say it has a valency of two.

Double and triple bonds

A single bond is formed when **one** pair of electrons is shared between two atoms.

A double bond is formed when **two** pairs of electrons are shared between two atoms.

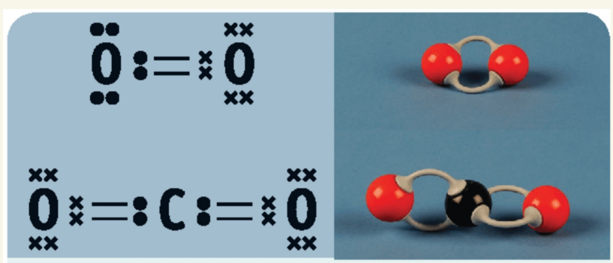


Fig. 5.14 Both the oxygen molecule and the carbon dioxide molecule contain double bonds.

A triple bond is formed when three pairs of electrons are shared.

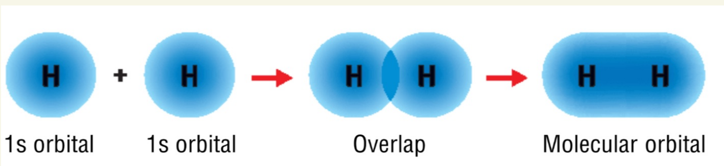
The nitrogen molecule is an example of a molecule that contains a triple bond.

Fig. 5.15 The nitrogen molecule contains a triple bond.

Sigma and Pi Bonding

It is also possible to describe covalent bonding in terms of atomic orbitals.

The formation of the H2 molecule can be described by saying the 1s orbitals of two hydrogen atoms overlap as shown.



When two atomic orbitals overlap they form a **molecular orbital.**

The covalent bond formed by head-on overlap of the two atomic orbitals is called a **sigma bond.**

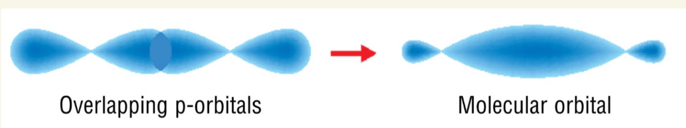


Fig. 5.18 The covalent bond in Cl_2 being formed by the head-on overlap of two p orbitals.

A **pi bond** is formed by the sideways over of p orbitals.

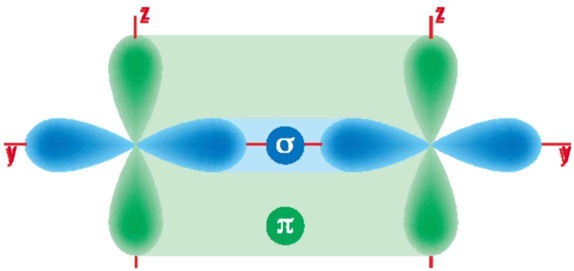
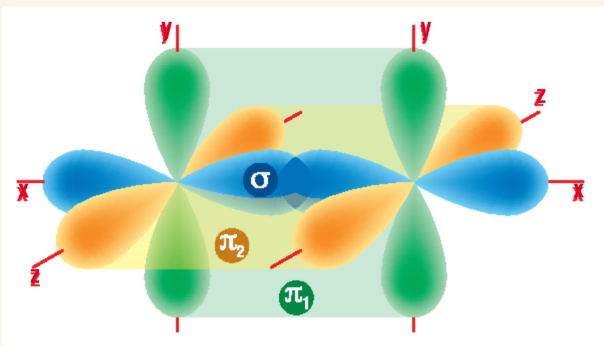


Fig. 5.19 The double bond in oxygen consists of one sigma and one pi bond.

If we look at the electron configuration for Nitrogen

$$N = 1s^2,\ 2s^2,\ 2p_x^1,\ 2p_y^1,\ 2p_z^1$$

We have 3 half-filled p-orbitals.



The Px orbitals overlap head on to form a **sigma bond**

The Py and Pz orbitals overlap sideways to form **two pi bonds.**

Sigma bonds are stronger than pi bonds as there is more overlapping.

Characteristics of Ionic and Covalent bonds

Hardness

Ionic

Ionic compounds are usually difficult to cut.
This is because each ion is held in a crystal lattice.

Covalent

Usually soft because they consists of molecules.

Melting and boiling points

Ionic

High melting and boiling points.

Covalent

Low melting and boiling points

Conduction of electricity

Ionic

Only conduct electricity if melted or dissolved in water.

Covalent

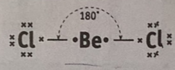
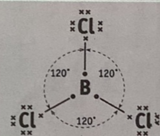
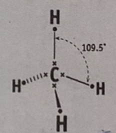
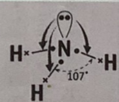
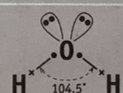
Do not conduct electricity.

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

Shapes of covalent molecules

To theory to account for the shapes of covalent molecules is known as the **Valence Shell Electron Pair Repulsion Theory** or **VSEPR Theory**

This theory states that the shape of a molecule depends on the number of pairs of electrons around the central atom.

Compound	Central atom	No. of electrons in outer shell of central atom	Bonding present	No. of electron pairs around central atom	Shape
$BeCl_2$	Beryllium (group 2)	2		2 bond pairs	Linear
BCl_3	Boron (group 3)	3		3 bond pairs	Triangular Planar
CH_4	Carbon (group 4)	4		4 bond pairs	Tetrahedral
NH_3	Nitrogen (group 5)	5		3 bond pairs + 1 lone pair	Pyramidal
H_2O	Oxygen (group 6)	6		2 bond pairs + 2 lone pairs	V-shaped

Lone pair/lone pair > lone pair/bond pair > bond pair/bond pair

Since lone pairs are not involved in bonding they are closer to the nucleus.

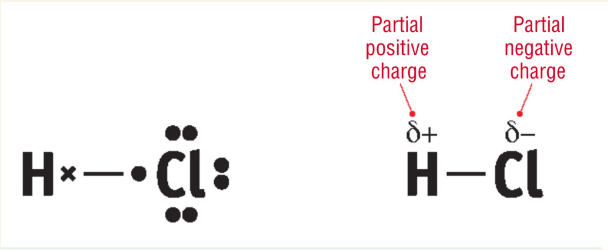
Because the lone pairs are closer to the nucleus, they exert a greater force of repulsion on the bond pairs.

Electronegativity

Electronegativity is like the ‘tug of war’ for the electrons.

In a covalent bond between identical atoms, the electrons are shared equally.

In a bond between different atoms, the electron is attracted closer to one atom than the other.



The ability for an atom to attract an electron is called the **electronegativity** of an atom.

Since the ‘electron pulling power’ of chlorine is stronger than hydrogen, the electron is closer to the chlorine.

The correct way to refer to the HCl bond is a **polar covalent bond.**

The word polar means ‘not shared equally’.

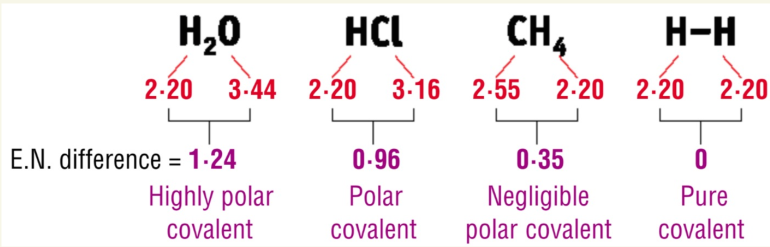
The term **pure covalent** is referred to a covalent bond where the electrons are shared equally.

Use of electronegativity values

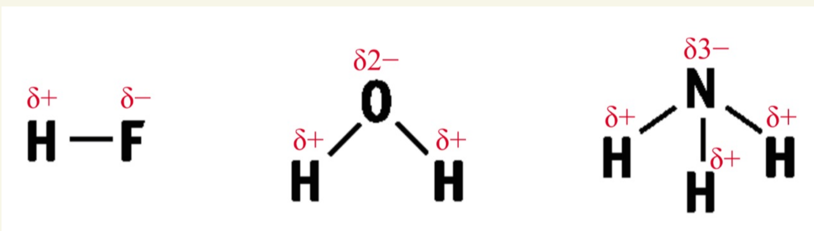
There are two main uses of electronegativity values:

1. **The predict the polarity of covalent bonds**

2. **To predict which compounds are ionic and which are covalent.**



In most molecules that have polar covalent bonds, the molecules themselves are also polar, i.e. the overall molecules have a partial positive and partial negative pole.



There are some molecules which, even though they are polar covalent bonds, are not polar molecules. These are usually symmetrical molecules

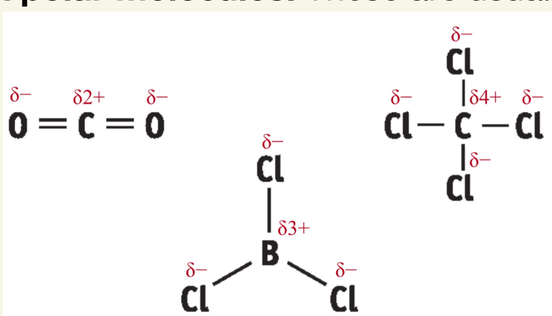


Fig. 5.33 Since the 'centre' of the partial negative charges coincides with that of the partial positive charges, none of these molecules are polar molecules.

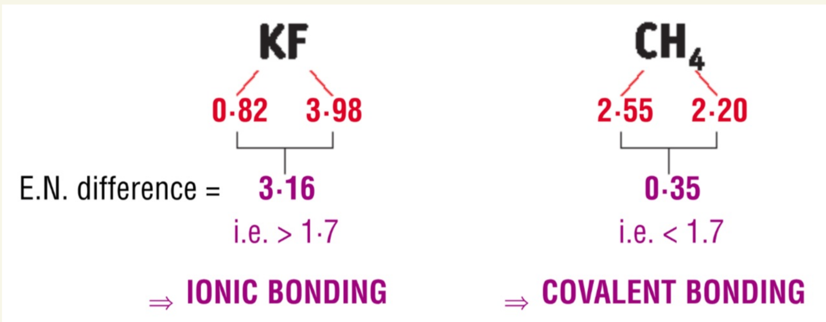
To predict which compounds are Ionic and which are Covalent

Electronegativity difference greater than 1.7 = Ionic bonding

Electronegativity difference less than or equal to 1.7 = covalent

Greater than 0.4 but less than 1.7 = polar covalent

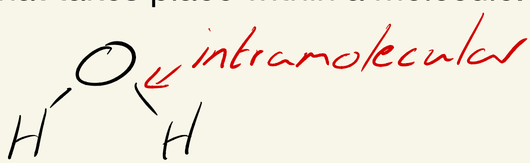
Electronegativity less than or equal to 0.4 = non-polar



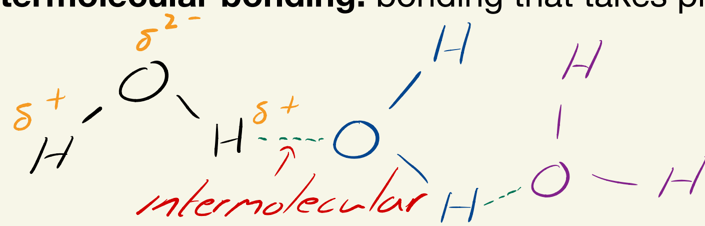
This is a general rule of thumb with some exceptions.

Intermolecular bonding and Intermolecular bonding

Intramolecular bonding: bonding that takes place within a molecule.



Intermolecular bonding: bonding that takes place between molecules



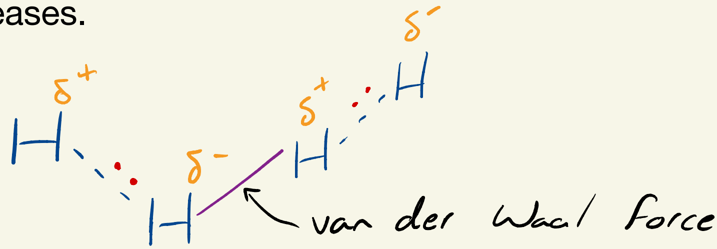
Van Der Waals Forces

Small attractive forces exist between molecules as a result of temporary shifts in the distribution of electrons in a molecule

The attraction between two molecules only exists for an instant. Therefore the attraction is very weak.

They are the only forces of attraction that exist between non-polar molecules.

The strength of van der Waals forces increase as the size of the molecule increases.



Cl2 is a gas at room temperature, Br2 is a liquid and I2 is a solid.

The reason for this is due to increasing van der Waals forces as the molecular masses increase, the number of electrons increase and the size of the electron cloud increases.

More electrons = stronger van der Waals forces.

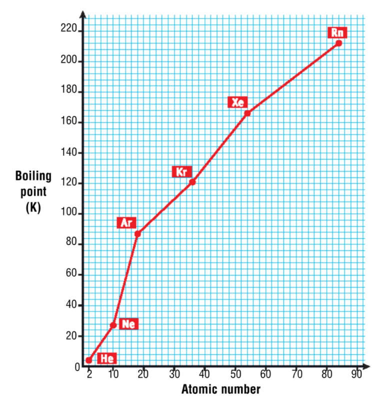
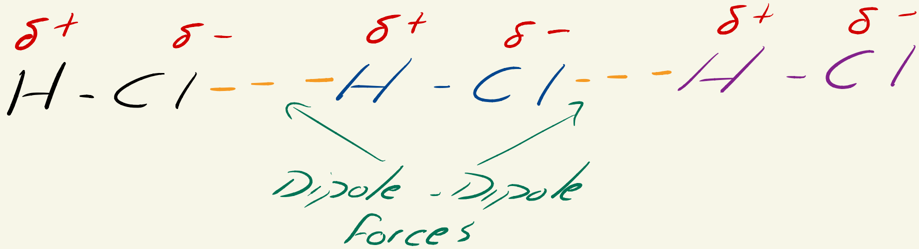


Fig. 5.40 The boiling points of the noble gases increase with increasing atomic number because the strength of the van der Waals forces increases as the atoms get bigger.

Dipole-Dipole Forces

In compounds whose molecules contain polar covalent bonds, forces called dipole-dipole forces exist permanently.

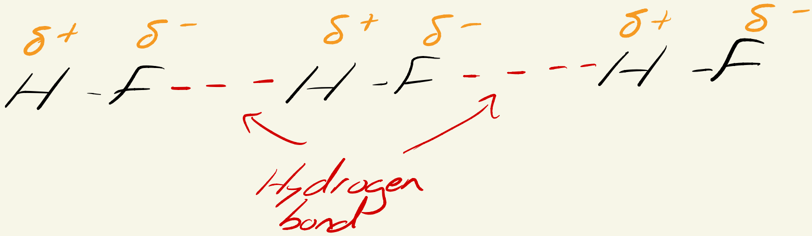
Dipole-dipole forces are forces of attraction between the negative pole of one polar molecule and the positive pole of another polar molecule.

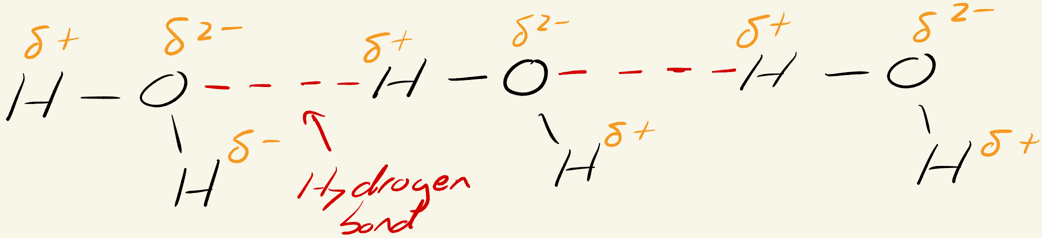


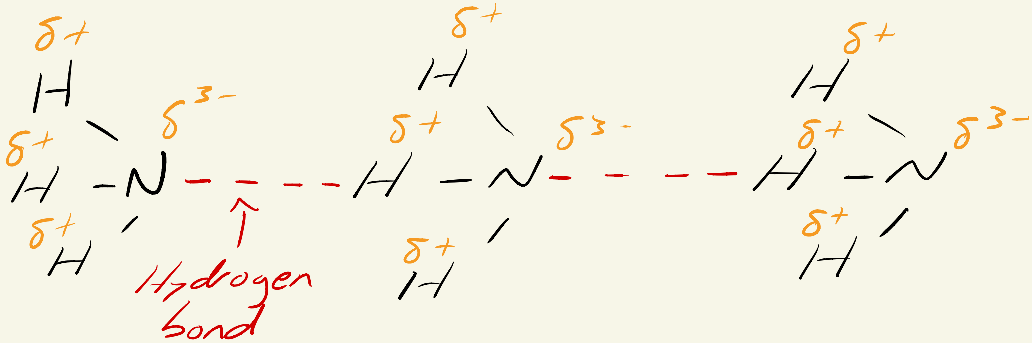
The dipole-dipole forces between the molecules give rise to higher boiling points.

Hydrogen bonding

Hydrogen bonding arises when an atom of hydrogen is bonded to an atom of **fluorine or oxygen or nitrogen.**







The hydrogen atom in each of the three compounds has a fairly strong positive charge because in each case it is bonded to a very electronegative element.

This causes a strong polarity.

The high boiling points of the three molecules is due to the considerable energy required to break the hydrogen bonds.

Hydrogen bonding continued

Hydrogen bonds are much stronger than either van dr Waals forces or dipole-dipole forces.

However, they are much weaker than a regular covalent bond.

The boiling point of H2O is higher than NH3 because oxygen is more electronegative than nitrogen, making the hydrogen bonding stronger.

Dissolving of covalent compounds in water

Since water is a polar covalent solvent, many compounds which are also polar covalent tend to dissolve in it.

Example: ammonia dissolving water.