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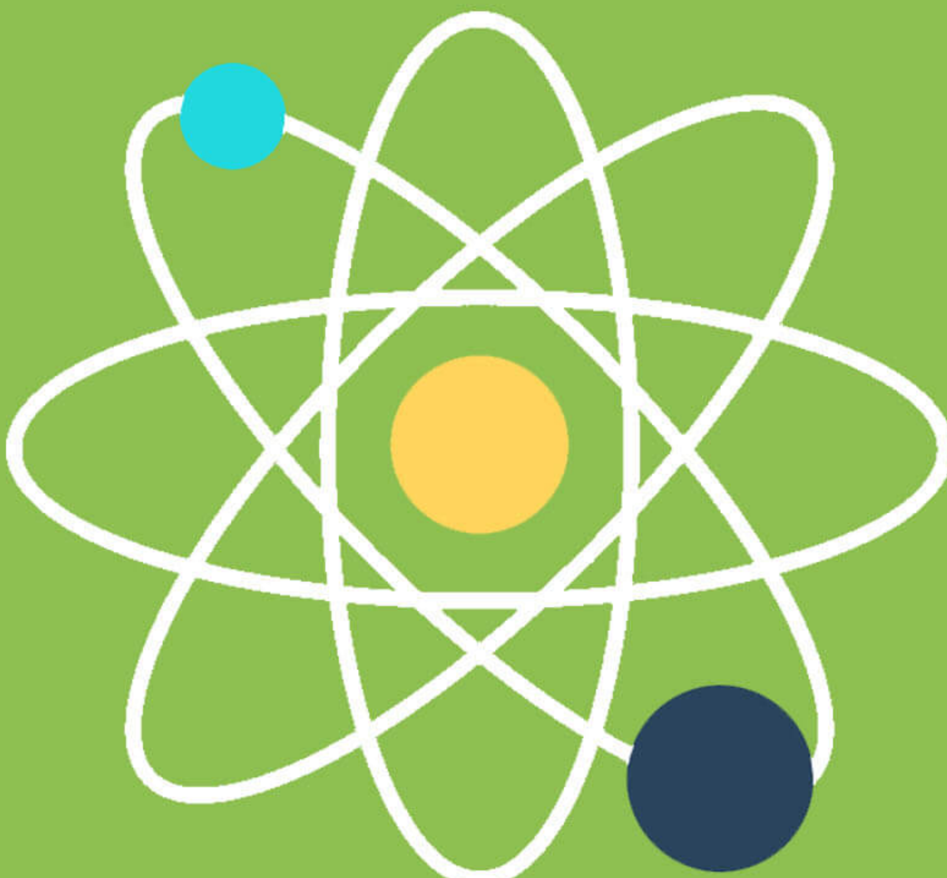
4.1 Ionic & Covalent Bonding



IB Chemistry - Revision Notes

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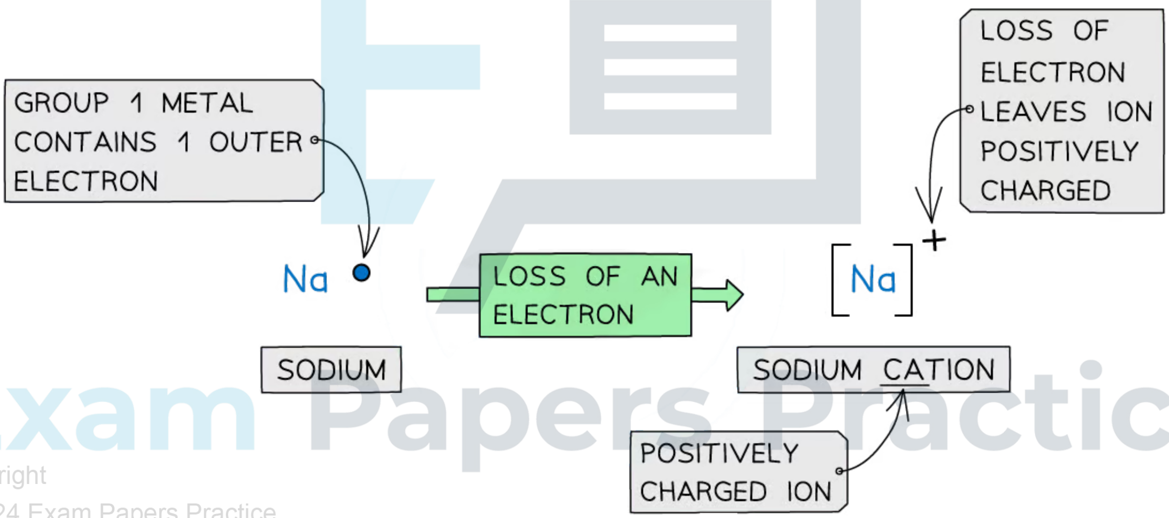
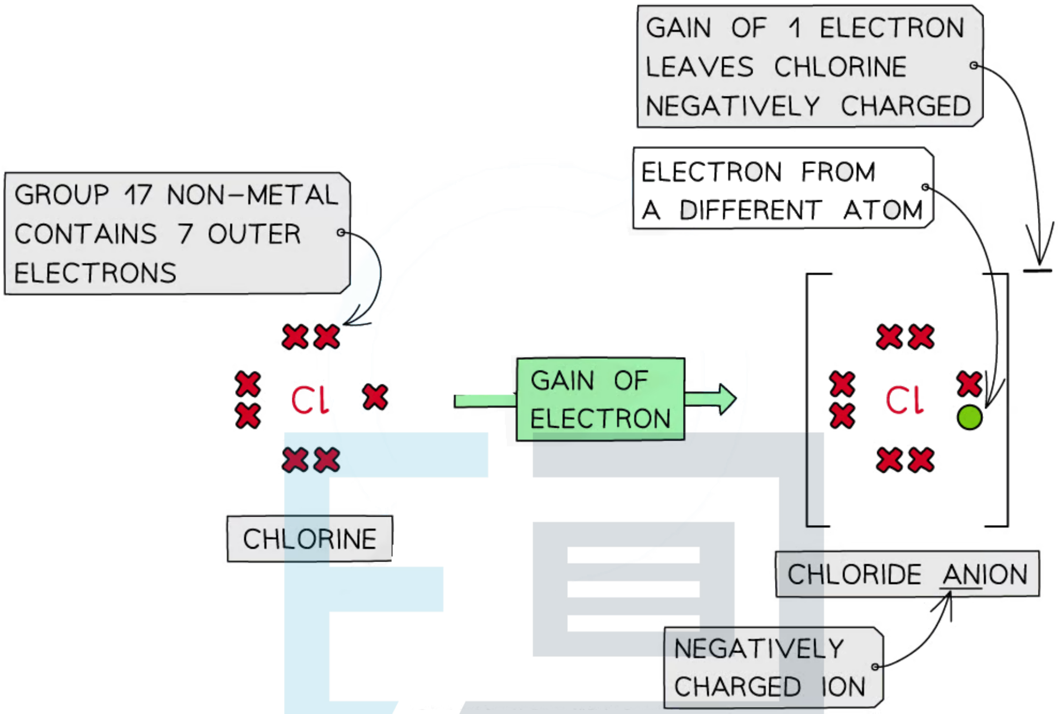
4.1.1 Forming Ions

Forming Ions

- As a general rule, **metals** are on the **left** of the Periodic Table and **non-metals** are on the **right-hand** side
- **Ionic bonds** involve the **transfer** of electrons from a **metallic** element to a **non-metallic** element
- Transferring electrons usually leaves the metal and the non-metal with a **full outer shell**
- Metals **lose** electrons from their valence shell forming positively charged **cations**
- Non-metal atoms **gain** electrons forming negatively charged **anions**
- Once the atoms become ions, their electronic configurations are the same as a noble gas.
  - A sodium ion (Na^+) has the same electronic configuration as neon: [2,8]
  - A chloride ion (Cl^-) also has the same electronic configuration as argon: [2,8,8]

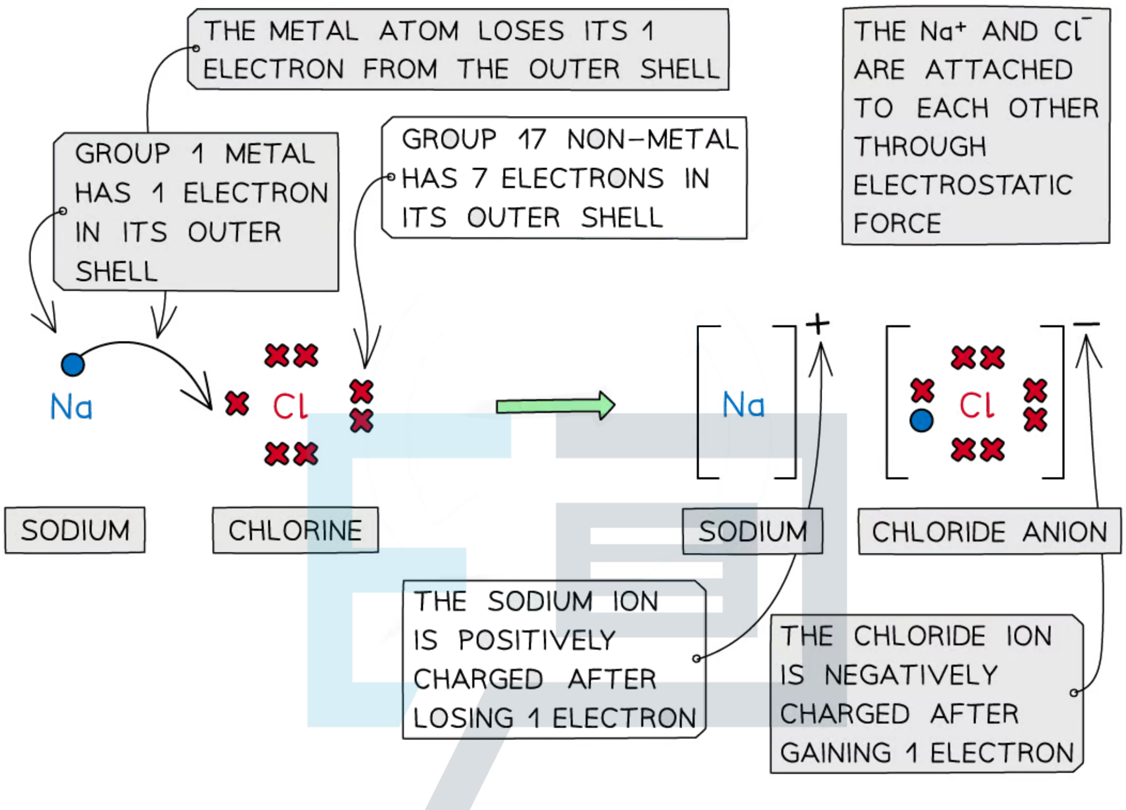


Forming cations by the removal of electrons from metals



Forming anions by the addition of electrons to nonmetals

- **Cations** and **anions** are oppositely charged and therefore attracted to each other
- **Electrostatic attractions** are formed between the oppositely charged ions to form **ionic compounds**
- This form of attraction is very **strong** and requires a lot of energy to overcome
  - This causes high melting points in ionic compounds



Cations and anions bond together using strong electrostatic forces, which require a lot of energy to overcome

Exam Tip

Metals usually **lose** all electrons from their outer valence shell to become **cations.** You can make use of the groups on the periodic table to work out how many electrons an atom is likely to lose or gain by looking at the **group** an atom belongs to.

4.1.2 Ionic Compounds

Ionic Lattices

- The ions form a **lattice structure** which is an evenly distributed **crystalline** structure
- Ions in a lattice are arranged in a **regular repeating pattern** so that positive charges cancel out negative charges
- Therefore the final lattice is overall electrically **neutral**

Properties of Ionic Compounds

- Different types of **structure** and **bonding** have different effects on the **physical properties** of substances such as their **melting** and **boiling points**, **electrical conductivity** and **solubility**

Ionic bonding & giant ionic lattice structures

- Ionic compounds are **strong**
  - The **strong electrostatic forces** in ionic compounds keep the ions held strongly together
- They are **brittle** as ionic crystals can split apart
- Ionic compounds have **high melting** and **boiling points**
  - The strong electrostatic forces between the ions in the lattice act in all directions and keep them strongly together
  - Melting and boiling points increase with the charge density of the ions due to the **greater electrostatic attraction** of charges
  - $Mg^{2+}O^{2-}$ has a higher melting point than $Na^{+}Cl^{-}$
- Ionic compounds are **soluble** in water as they can form **ion-dipole bonds**
- Ionic compounds only **conduct electricity** when **molten** or **in solution**
  - When molten or in solution, the ions can freely move around and conduct electricity
  - As a solid, the ions are in a fixed position and unable to move around

Table comparing the characteristics of giant ionic lattices with other structure types

	Giant ionic	Giant metallic	Simple covalent	Giant covalent
Melting / boiling point	High	Moderately high to high	Low	Very high
Electrical conductivity	Only when molten or in solution	When solid or liquid	Do not conduct electricity	Do not conduct electricity (except **graphite**)
Solubility	Soluble	Insoluble but some may react	Usually insoluble unless they are polar	Insoluble
Hardness	Hard, brittle	Hard, malleable	Soft	Very hard (**diamond and silica**) or soft (**graphite**)
Physical state at room	Solid	Solid	Solid, liquid or gas	Solid

temperature				
Forces	Electrostatic attraction between ions	Delocalised electrons attracting positive ions	Weak intermolecular forces and covalent bonds within a molecule	Electrons in covalent bonds between atoms
Particles	Ions	Positive ions in a sea of electrons	Small molecules	Atoms
Examples	NaCl	Copper	Br_2	Graphite, silicon(IV) oxide

Worked example

The table below shows the physical properties of substances **X**, **Y** and **Z**.

Substance	Melting point (°C)	Electrical conductivity when molten	Solubility in water
X	839	Good	Soluble
Y	95	Very poor	Almost insoluble
Z	1389	Good	Insoluble

Which one of the following statements about **X**, **Y** and **Z** is completely true?

Statement 1: **X** has a giant ionic structure, **Y** has a giant molecular structure, **Z** is a metal

Statement 2: **X** is a metal, **Y** has a simple molecular structure, **Z** has a giant molecular structure

Statement 3: **X** is a metal, **Y** has a simple molecular structure, **Z** has a giant ionic structure

Statement 4: **X** has a giant ionic structure, **Y** has a simple molecular structure, **Z** is a metal

Answer:

- Compound **X** has a relatively high melting point, is soluble in water and conducts electricity when molten
  - This suggests that **X** has a giant ionic structure
- Compound **Y** has a low melting point which suggests that little energy is needed to break the lattice
  - This suggests that **Y** is a simple molecular structure
  - This is further supported by its low electrical conductivity and it being almost insoluble in water
- Compound **Z** has a very high melting point, which is characteristic of either metallic, giant ionic lattices or giant covalent / molecular lattices
  - However since it is insoluble in water, compound **Z** must be a metal
- Therefore, the correct answer is **Statement 4**

4.1.3 Formulae & Names of Ionic Compounds

Formulae & Names of Ionic Compounds

- **Ionic compounds** are formed from a **metal** and a **nonmetal** bonded together
- Ionic compounds are electrically neutral; the positive charges equal the negative charges

Charges on positive ions

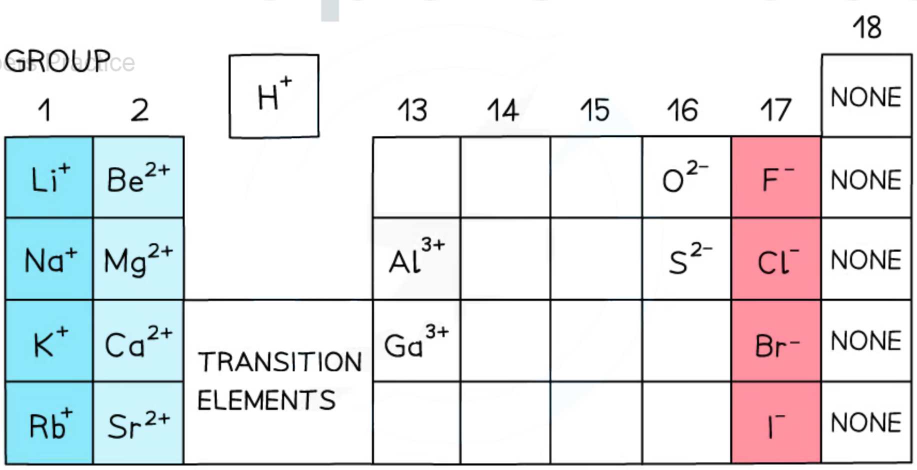
- All metals form **positive** ions
  - There are some non-metal positive ions such as ammonium, NH_4^+, and hydrogen, H^+

- The **metals** in Group 1, Group 2 and Group 13 have a charge of 1+ and 2+ and 3+ respectively
- The charge on the ions of the **transition elements can vary** which is why **Roman numerals** are often used to indicate their charge
- This is known as **Stock notation** after the German chemist Alfred Stock
- **Roman numerals** are used in some compounds formed from transition elements to show the **charge** (or **oxidation state**) of metal ions
  - Eg. in copper (II) oxide, the copper ion has a charge of 2+ whereas in copper (I) nitrate, the copper has a charge of 1+

Non-metal ions

- The **non-metals** in group 15 to 17 have a negative charge and have the suffix '**ide**'
  - Eg. nitride, chloride, bromide, iodide

- Elements in group 17 gain 1 electron so have a 1– charge, eg. Br^-
- Elements in group 16 gain 2 electrons so have a 2– charge, eg. O^{2-}
- Elements in group 15 gain 3 electrons so have a 3– charge, eg. N^{3-}
- There are also more **polyatomic** or **compound negative ions**, which are negative ions made up of more than one type of atom



The charges of simple ions depend on their position in the Periodic Table

- There are seven polyatomic ions you need to know for IB Chemistry:

Formulae of Polyatomic Ions Table

Ion	Formula and Charge
Ammonium	NH_4^+
Hydroxide	OH^-
Nitrate	NO_3^-
Sulfate	SO_4^{2-}
Carbonate	CO_3^{2-}
Hydrogen carbonate	HCO_3^-
Phosphate	PO_4^{3-}

Worked example

Determine the formulae of the following ionic compounds

1. magnesium chloride
2. aluminium oxide
3. ammonium sulfate

Answer:

Answer 1: Magnesium chloride

- Magnesium is in group 2 so has a charge of 2+
- Chlorine is in group 17 so has a charge of 1–
- Magnesium needs two chlorine atoms for each magnesium atom to be balanced so the formula is **$MgCl_2$**

Answer 2: Aluminium oxide

- Aluminum is in group 13 so the ion has a charge of 3+
- Oxygen is in group 16 so has a charge of 2–
- The charges need to be equal so 2 aluminium to 3 oxygen atoms will balance electrically, so the formula is $\mathbf{Al_2O_3}$

Answer 3: Ammonium sulfate

- Ammonium is a polyatomic ion with a charge of 1+
- Sulfate is a **polyatomic ion** and has a charge of 2–
- The polyatomic ion needs to be placed in a bracket if more than 1 is needed
- The formula of ammonium nitrate is $\mathbf{(NH_4)_2SO_4}$

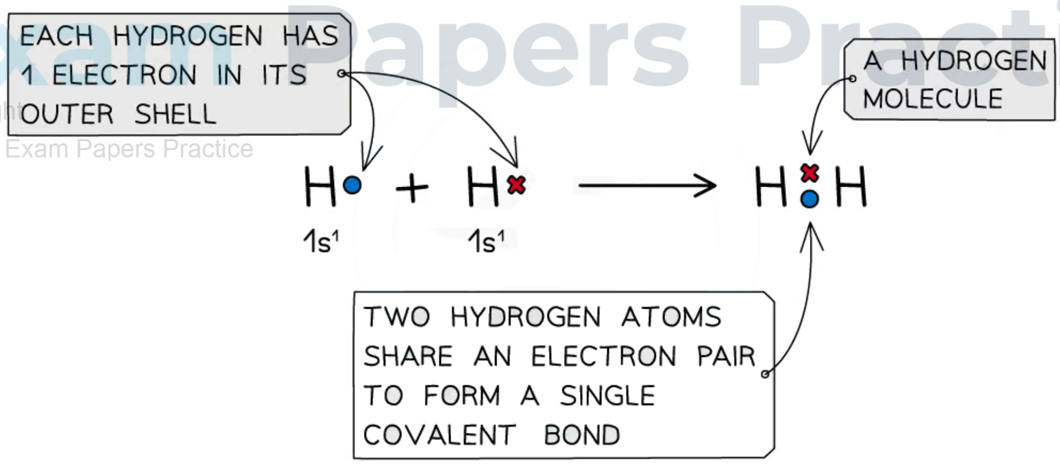
Exam Tip

Remember: **polyatomic ions** are ions that contain more than one type of element, such as OH^-

4.1.4 Covalent Bonds

Covalent Bonds

- **Covalent** bonding occurs between two **non-metals**
- A covalent bond involves the **electrostatic attraction** between nuclei of two atoms and the electrons of their outer shells
- **No electrons** are **transferred** but only **shared** in this type of bonding
- When a covalent bond is formed, two **atomic orbitals** overlap and a **molecular orbital** is formed
- Covalent bonding happens because the electrons are more stable when attracted to two nuclei than when attracted to only one



The positive nucleus of each atom has an attraction for the bonding electrons shared in the covalent bond

- In a normal covalent bond, each atom provide one of the electrons in the bond. A covalent bond is represented by a short straight line between the two atoms, H-H
- Covalent bonds should not be regarded as shared electron pairs in a fixed position; the electrons are in a state of constant motion and are best regarded as **charge clouds**

A representation of electron charge clouds. The electrons can be found anywhere in the charge clouds

- **Non-metals** are able to **share** pairs of electrons to form different types of covalent bonds
- Sharing electrons in the covalent bond allows each of the 2 atoms to achieve an electron configuration similar to a noble gas
  - This makes each atom more stable

- In some instances, the central atom of a covalently bonded molecule can accommodate **more** or **less** than 8 electrons in its outer shell
  - Being able to accommodate **more** than 8 electrons in the outer shell is known as **'expanding the octet rule'**
  - Accommodating **less** than 8 electrons in the outer shell means than the central atom is **'electron deficient'**
  - Some examples of this can be found in the section on Lewis structures

Exam Tip

Covalent bonding takes place between two nonmetal atoms.Remember to use the periodic table to decide how many electrons are in the outer shell of a nonmetal atom.

Predicting Covalent Bonding

- The differences in Pauling electronegativity values can be used to predict whether a bond is **covalent** or **ionic** in character

Electronegativity & covalent bonds

- In **diatomic molecules** the electron density is shared equally between the two atoms
  - Eg. H_2, O_2 and Cl_2

- Both atoms will have the same electronegativity value and have an **equal attraction** for the bonding pair of electrons leading to formation of a **covalent** bond

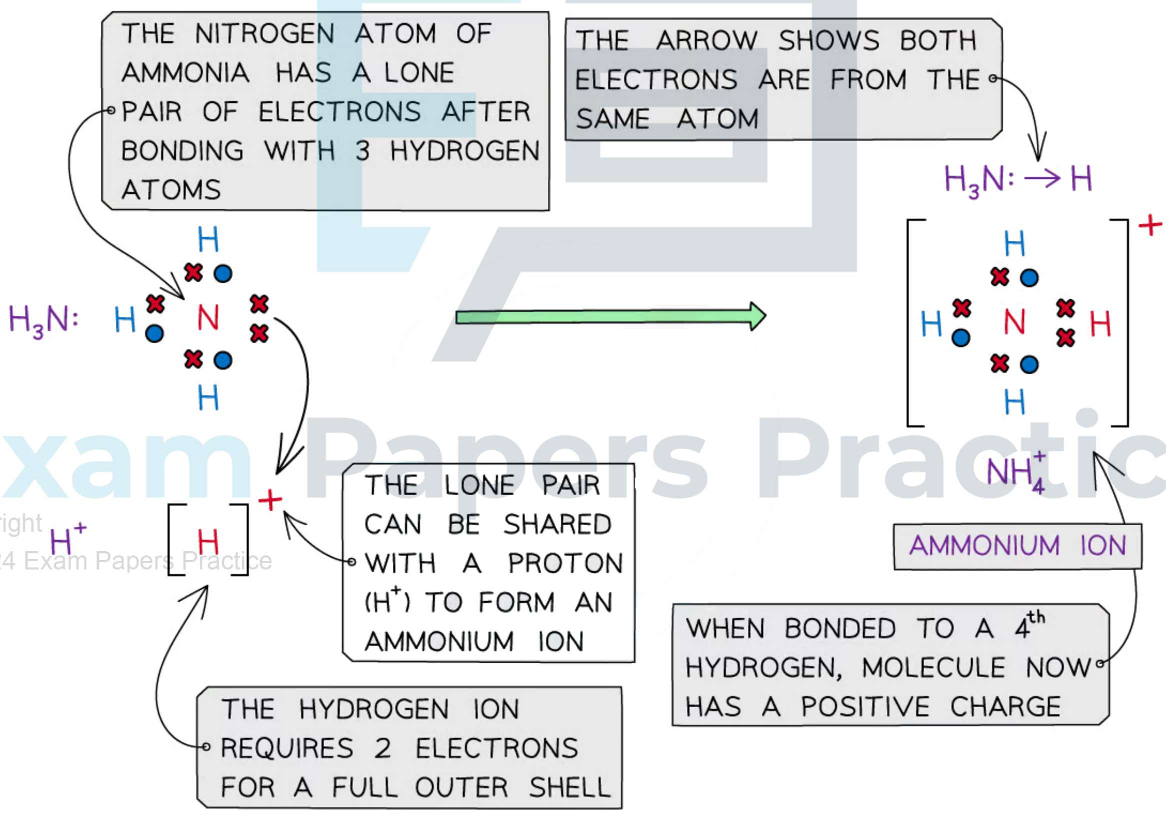
- A difference of less than around **1.0** in electronegativity values will be associated with covalent bonds, although between 1.0 and 2.0 can be considered polar covalent:

You can use the Pauling scale to decide whether a bond is polar or nonpolar:

Difference in Electronegativity	Bond Type
< 1.0	Covalent
1.0 – 2.0	Polar Covalent
> 2.0	Ionic

Coordinate Bonds

- In **simple covalent bonds** the two atoms involved share electrons
- Some molecules have a **lone** pair of electrons that can be donated to form a bond with an **electron-deficient** atom
    - An electron-deficient atom is an atom that has an **unfilled outer orbital**
- So **both electrons** are from the **same atom**
- This type of bonding is called **dative covalent bonding** or **coordinate bond**
- An example of a dative bond is in an **ammonium ion**
    - The hydrogen ion, H^+ is **electron-deficient** and has space for two electrons in its shell
    - The nitrogen atom in ammonia has a lone pair of electrons which it can donate to the hydrogen ion to form a dative covalent bond



Ammonia (NH_3) can donate a lone pair to an electron-deficient proton (H^+) to form a charged ammonium ion (NH_4^+)

- More examples of coordinate bonding can be found in the section on **Lewis Structures**

Multiple Bonds

- **Non-metals** are able to **share** more than one pair of electrons to form different types of covalent bonds
- Sharing electrons in the covalent bond allows each of the 2 atoms to achieve an electron configuration similar to a noble gas
  - This makes each atom more stable
- It is not possible to form a quadruple bond as the repulsion from having 8 electrons in the same region between the two nuclei is too great

Covalent Bonds & Shared Electrons Table

Type of covalent bond	Number of electrons shared
Single (C – C)	2
Double (C = C)	4
Triple (C ≡ C)	6

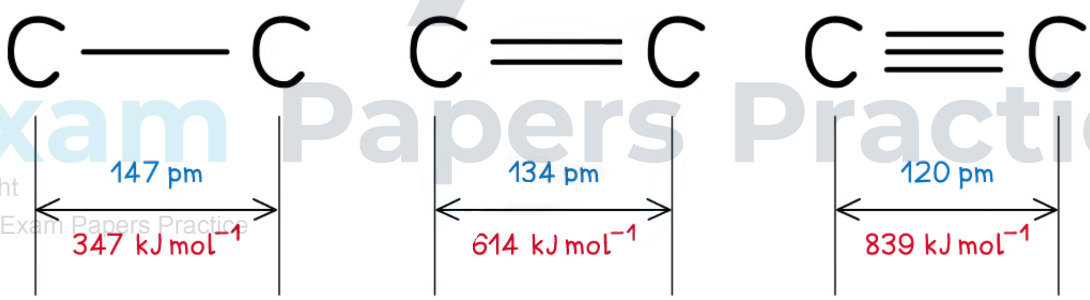
Bond Length & Strength

Bond energy

- The **bond energy** is the energy required to **break** one mole of a particular covalent bond in the gaseous states
  - Bond energy has units of kJ mol^{-1}
- The **larger** the bond energy, the **stronger** the covalent bond is

Bond length

- The **bond length** is **internuclear distance of two covalently bonded atoms**
  - It is the distance from the nucleus of one atom to another atom which forms the covalent bond
- The **greater** the forces of attraction between electrons and nuclei, the more the atoms are pulled closer to each other
- This **decreases** the **bond length** of a molecule and **increases** the **strength** of the covalent bond
- **Triple bonds** are the **shortest** and **strongest** covalent bonds due to the large electron density between the nuclei of the two atoms
- This increase the forces of attraction between the electrons and nuclei of the atoms
- As a result of this, the atoms are pulled closer together causing a shorter bond length
- The increased forces of attraction also means that the covalent bond is **stronger**



Triple bonds are the shortest covalent bonds and therefore the strongest ones

- Test your knowledge of covalent bonding:

Worked example

Which molecules react together to form a dative covalent bond?

A. Cl_2 and HF

B. C_2H_2 and Cl_2

C. NH_3 and HF

D. CH_4 and NH_3

Answer:

The correct option is **C**.

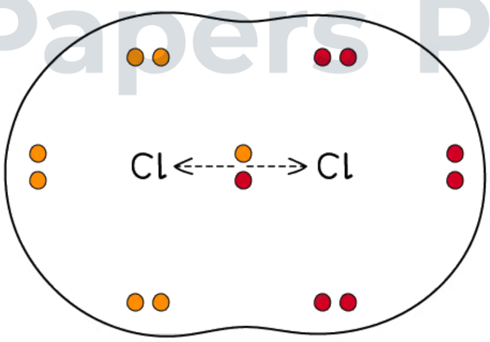
- To form a dative covalent bond one species must have a lone pair of electrons and the other must be electron deficient.
- NH_3 has a lone pair and HF splits into H^+ (electron deficient) and F^-

$$NH_3 + HF \rightarrow NH_4^+F^-$$

4.1.5 Bond Polarity

Bond Polarity

- When two atoms in a covalent bond have the **same electronegativity** the covalent bond is **nonpolar**

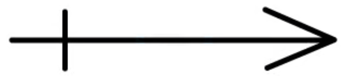


The two chlorine atoms have identical electronegativities so the bonding electrons are shared equally between the two atoms

- When two atoms in a covalent bond have **different electronegativities** the covalent bond is **polar** and the electrons will be drawn towards the **more electronegative** atom
- As a result of this:
  - The negative charge centre and positive charge centre do not **coincide** with each other
  - This means that the **electron distribution** is **asymmetric**
  - The **less electronegative** atom gets a partial charge of **δ+** (**delta positive**)
  - The **more electronegative** atom gets a partial charge of **δ–** (**delta negative**)
- The greater the difference in **electronegativity** the more polar the bond becomes

Dipole moment

- The **dipole moment** is a measure of how **polar** a bond is
- The **direction** of the dipole moment is shown by the following sign in which the **arrow** points to the **partially negatively charged end** of the dipole:



The sign shows the direction of the dipole moment and the arrow points to the delta negative end of the dipole

Worked example

The electronegativity values of four elements are given.

C = 2.6 N = 3.0 O = 3.4 F = 4.0

What is the order of **increasing** polarity of the **bonds** in the following compounds?

A. $CO < OF_2 < NO < CF_4$

B. $NO < OF_2 < CO < CF_4$

C. $CF_4 < CO < OF_2 < NO$

D. $CF_4 < NO < OF_2 < CO$

Answer:

The correct option is **B**.

- You have to calculate the difference in electronegativity for the bonds and then rank them from smallest to largest:

NO (3.4 - 3.0 = **0.4**)

OF_2 (4.0 - 3.4 = **0.6**)

CO (3.4 - 2.6 = **0.8**)

CF_4 (4.0 - 2.6 = **1.4**)

4.1.6 Lewis Structures

Lewis Structures

- **Lewis structures** are simplified electron shell diagrams and show pairs of electrons around atoms.
- A pair of electrons can be represented by dots, crosses, a combination of dots and crosses or by a line. For example, chlorine can be shown as:

Different Lewis Structures for chlorine molecules

- Note: Cl-Cl is not a **Lewis structure**, since it does not show all the electron pairs.
- The "**octet rule**" refers to the tendency of atoms to gain a valence shell with a total of 8 electrons

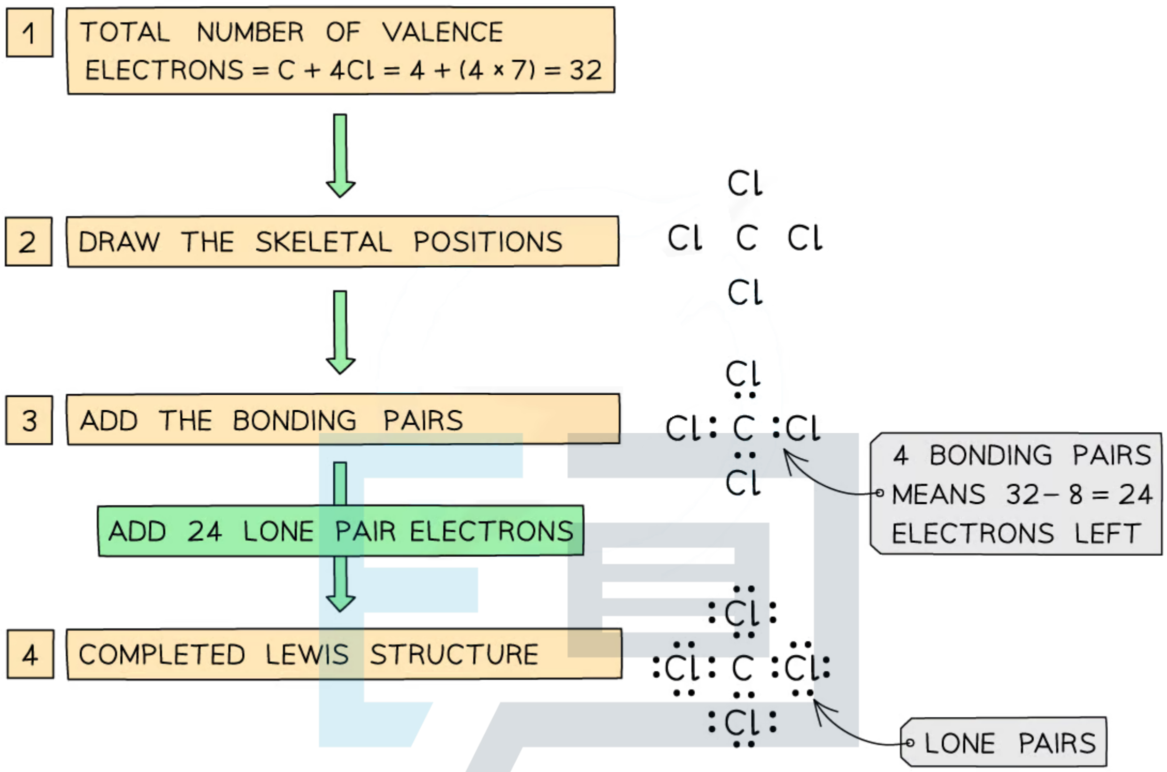
Steps for drawing Lewis Structures

1. Count the total number of **valence**
2. Draw the **skeletal structure** to show how many atoms are linked to each other.
3. Use a pair of crosses or dot/cross to put an electron pair in each bond between the atoms.
4. Add more electron pairs to complete the octets around the atoms (except H which has 2 electrons)
5. If there are not enough electrons to complete the octets, form double/triple bonds.
6. Check the total number of electrons in the finished structure is equal to the total number of **valence** electrons

Worked example

Draw a Lewis structure for CCl_4

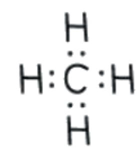
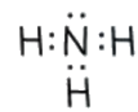
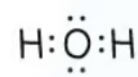
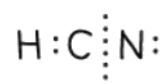
Answer:



Steps in drawing the Lewis Structure for CCl_4

Further examples of Lewis structures

- Follow the steps for drawing Lewis structures for these common molecules

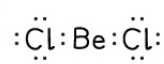
Molecule	Total number of valence electrons	Lewis structure
CH_4	C + 4H $4 + (4 \times 1) = 8$	
NH_3	N + 3H $5 + (3 \times 1) = 8$	
H_2O	2H + O $(2 \times 1) + 6 = 8$	
CO_2	C + 2O $4 + (2 \times 6) = 16$	
HCN	H + C + N $1 + 4 + 5 = 10$	

Incomplete Octets

- For elements below atomic number 20 the **octet rule** states that the atoms try to achieve 8 electrons in their valence shells, so they have the same electron configuration as a noble gas
- However, there are some elements that are exceptions to the **octet rule**, such a H, Li, Be, B and Al
    - H can achieve a stable arrangement by gaining an electron to become $1s^2$, the same structure as the noble gas helium
    - Li does the same, but losing an electron and going from $1s^22s^1$ to $1s^2$ to become a Li^+ ion
    - Be from group 2, has two valence electrons and forms stable compounds with just four electrons in the valence shell
    - B and Al in group 13 have 3 valence electrons and can form stable compounds with only 6 valence electrons

- There are two examples of **Lewis structures** with incomplete octets you should know, $BeCl_2$ and BF_3:

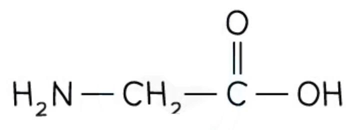
Incomplete Octets Examples

Molecule	Total number of valence electrons	Lewis structure
$BeCl_2$	Be + 2Cl = 2 + (2 × 7) = 16	
BF_3	B + 3F = 3 + (3 × 7) = 24	

- Test your understanding of Lewis diagrams in the following example:

Worked example

How many electrons are in the 2-aminoethanoic acid molecule?



2-AMINOETHANOIC ACID

A. 18

B. 20

C. 28

D. 30

Answer:

The correct option is **D**.

- You must count the lone pairs on N and O as well as the bonding pairs. There are 5 'hidden' pairs of bonding electrons in the OH, CH_2 and NH_2 groups. Hydrogen does not follow the octet rule.