

# DP IB Chemistry: SL

  
Your notes

## 4.1 Ionic & Covalent Bonding

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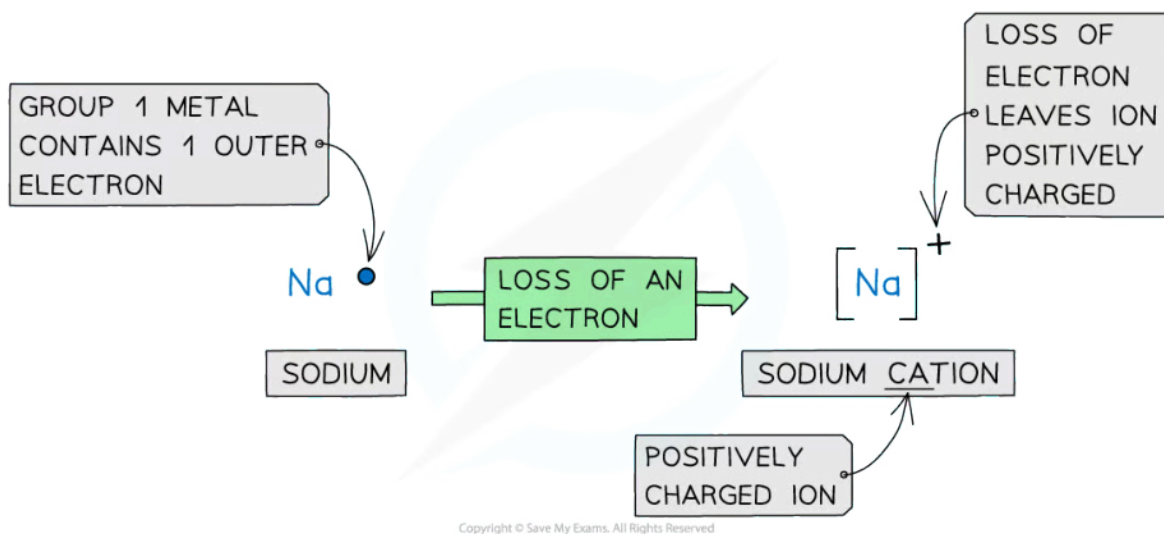


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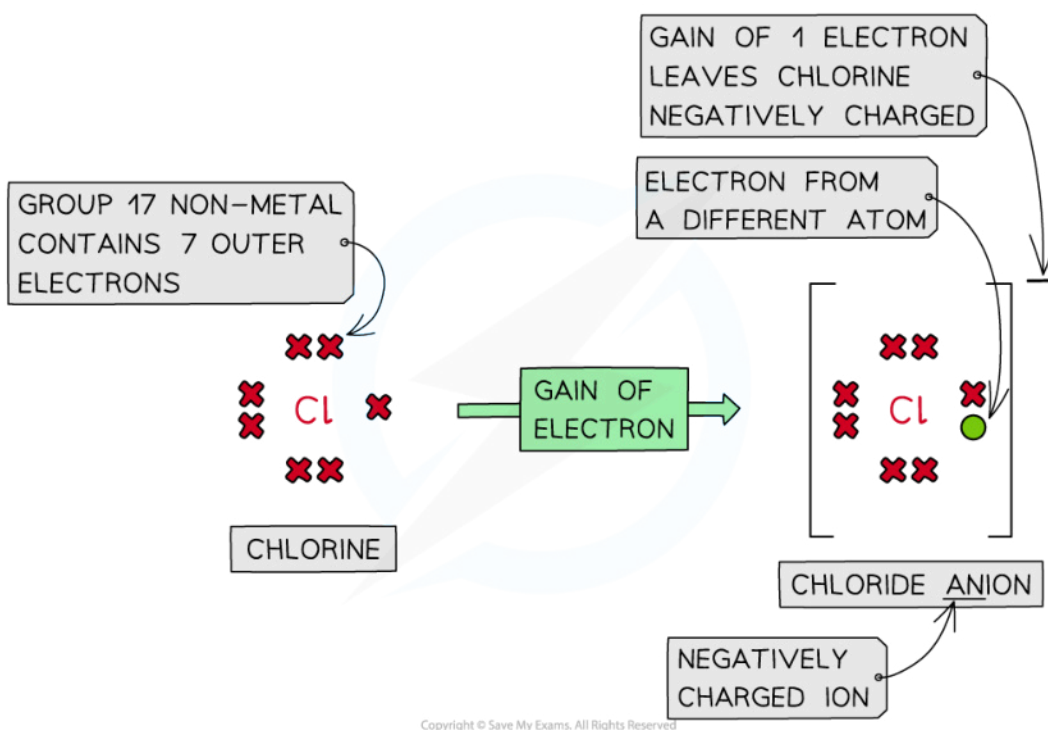
## 4.1.1 Forming Ions

### 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 ( $\text{Na}^+$ ) has the same electronic configuration as neon: [2,8]
  - A chloride ion ( $\text{Cl}^-$ ) also has the same electronic configuration as argon: [2,8,8]

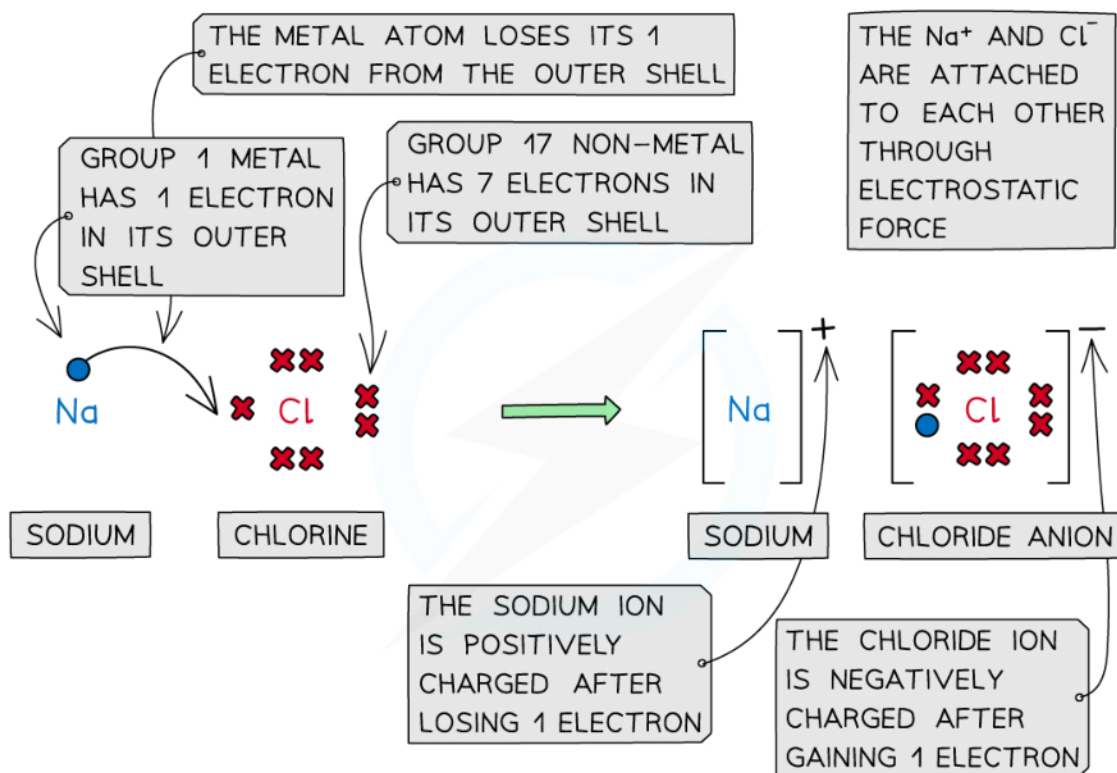


*Forming cations by the removal of electrons from metals*



**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



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

### Examiner 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.

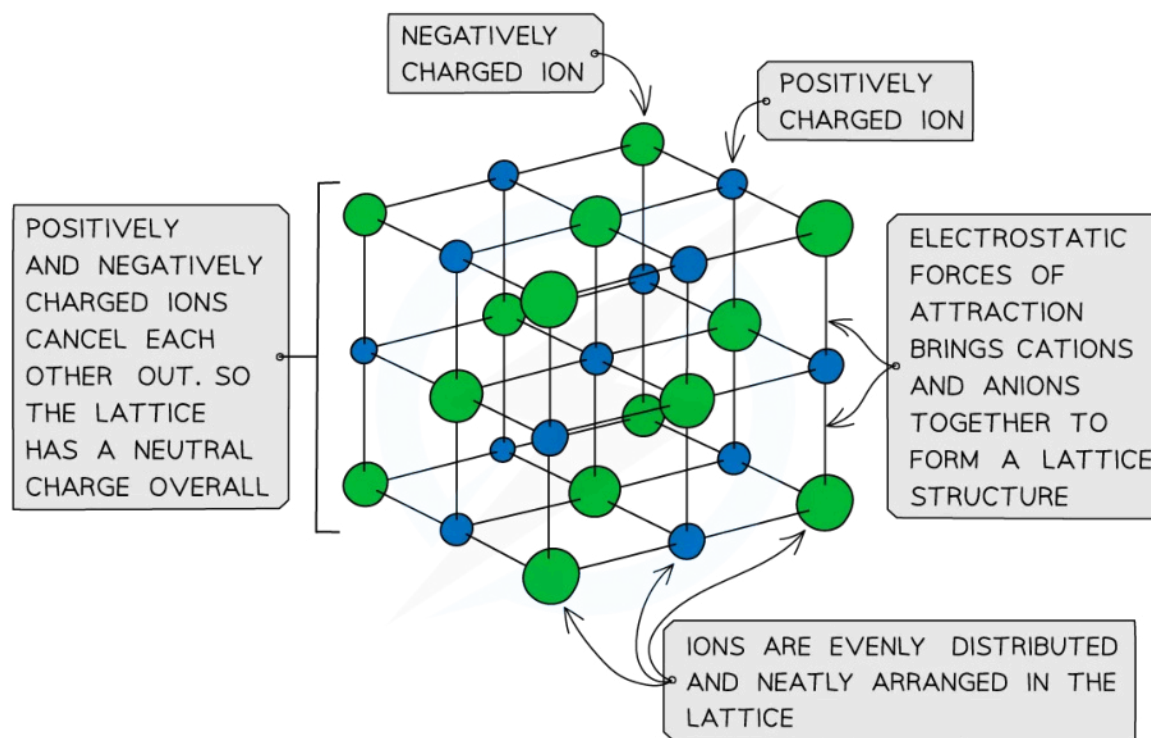


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## 4.1.2 Ionic Compounds

### 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**



*Ionic solids are arranged in lattice structures*



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

<b>Forces</b>	Electrostatic attraction between ions	Delocalised electrons attracting positive ions	Weak intermolecular forces and covalent bonds within a molecule	Electrons in covalent bonds between atoms
<b>Particles</b>	Ions	Positive ions in a sea of electrons	Small molecules	Atoms
<b>Examples</b>	NaCl	Copper	Br <sub>2</sub>	Graphite, silicon(IV) oxide



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### 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
<b>X</b>	839	Good	Soluble
<b>Y</b>	95	Very poor	Almost insoluble
<b>Z</b>	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**





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## 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,  $\text{NH}_4^+$ , and hydrogen,  $\text{H}^+$
- 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.  $\text{Br}^-$
- Elements in group 16 gain 2 electrons so have a 2- charge, eg.  $\text{O}^{2-}$
- Elements in group 15 gain 3 electrons so have a 3- charge, eg.  $\text{N}^{3-}$
- There are also more **polyatomic** or **compound negative ions**, which are negative ions made up of more than one type of atom

GROUP								18
1	2	H <sup>+</sup>	13	14	15	16	17	NONE
Li <sup>+</sup>	Be <sup>2+</sup>					O <sup>2-</sup>	F <sup>-</sup>	NONE
Na <sup>+</sup>	Mg <sup>2+</sup>		Al <sup>3+</sup>			S <sup>2-</sup>	Cl <sup>-</sup>	NONE
K <sup>+</sup>	Ca <sup>2+</sup>	TRANSITION ELEMENTS	Ga <sup>3+</sup>				Br <sup>-</sup>	NONE
Rb <sup>+</sup>	Sr <sup>2+</sup>						I <sup>-</sup>	NONE

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*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 <sub>4</sub> <sup>+</sup>
Hydroxide	OH <sup>-</sup>
Nitrate	NO <sub>3</sub> <sup>-</sup>
Sulfate	SO <sub>4</sub> <sup>2-</sup>
Carbonate	CO <sub>3</sub> <sup>2-</sup>
Hydrogen carbonate	HCO <sub>3</sub> <sup>-</sup>
Phosphate	PO <sub>4</sub> <sup>3-</sup>

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### 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<sub>2</sub>**

#### **Answer 2: Aluminium oxide**

- Aluminium 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 **Al<sub>2</sub>O<sub>3</sub>**

#### **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 sulfate is **(NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub>**

### Examiner Tip

Remember: **polyatomic ions** are ions that contain more than one type of element, such as OH<sup>-</sup>

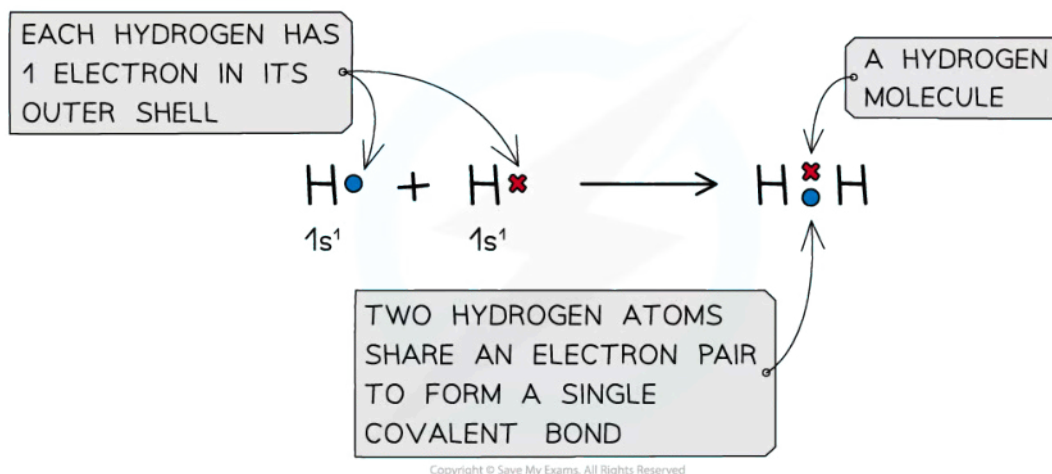


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## 4.1.4 Covalent Bonds

### 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

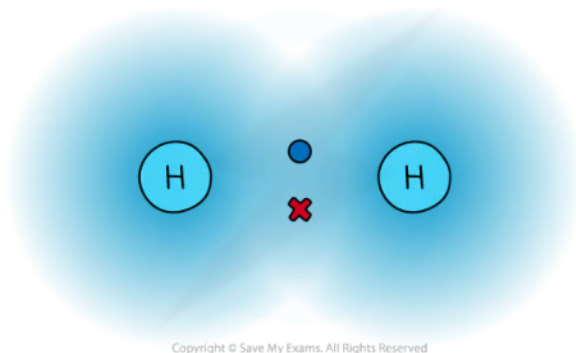


***The positive nucleus of each atom has an attraction for the bonding electrons shared in the covalent bond***

- In a normal covalent bond, each atom provides 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**



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**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 that the central atom is '**electron deficient**'
  - Some examples of this can be found in the section on Lewis structures

### Examiner 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.  $\text{H}_2$ ,  $\text{O}_2$  and  $\text{Cl}_2$
- 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

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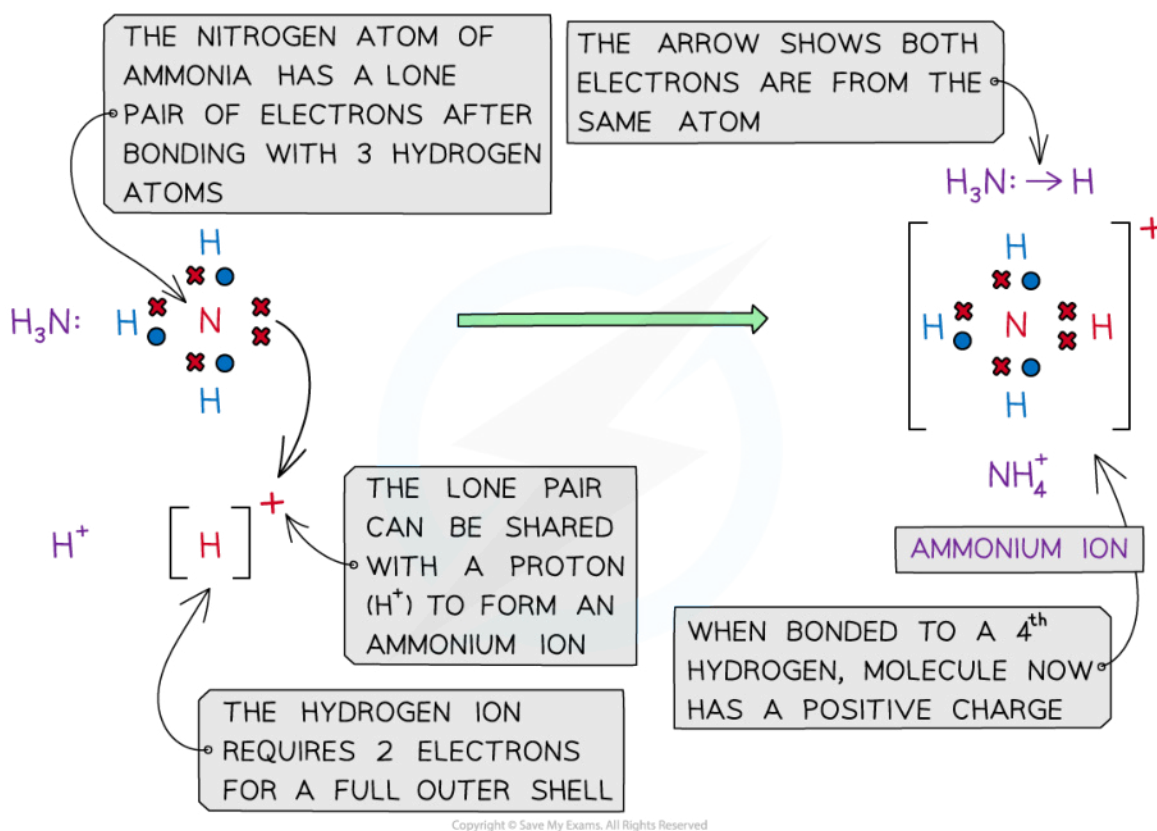
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## Coordinate Bonds

- 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



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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

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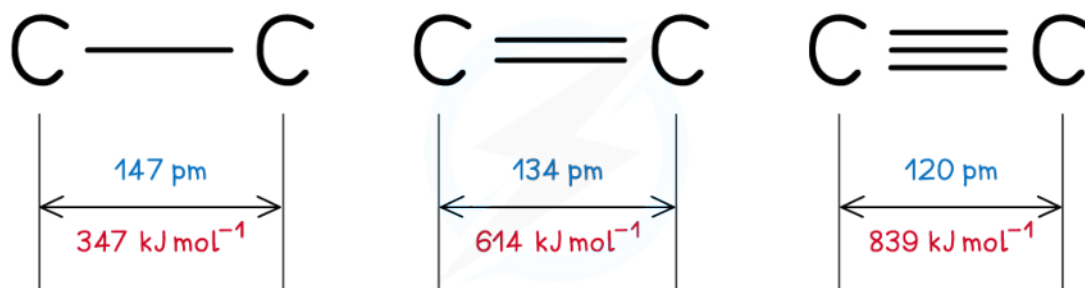
## Bond Length & Strength

### 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  $\text{kJ mol}^{-1}$
- 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**



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

- Test your knowledge of covalent bonding:



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 **Worked example**

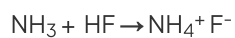
Which molecules react together to form a dative covalent bond?

- A.  $\text{Cl}_2$  and HF
- B.  $\text{C}_2\text{H}_2$  and  $\text{Cl}_2$
- C.  $\text{NH}_3$  and HF
- D.  $\text{CH}_4$  and  $\text{NH}_3$

**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.
- $\text{NH}_3$  has a lone pair and HF splits into  $\text{H}^+$  (electron deficient) and  $\text{F}^-$



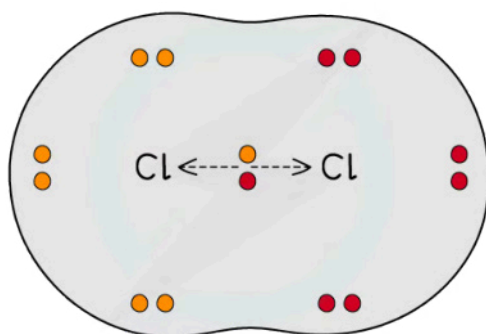
## 4.1.5 Bond Polarity



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### Bond Polarity

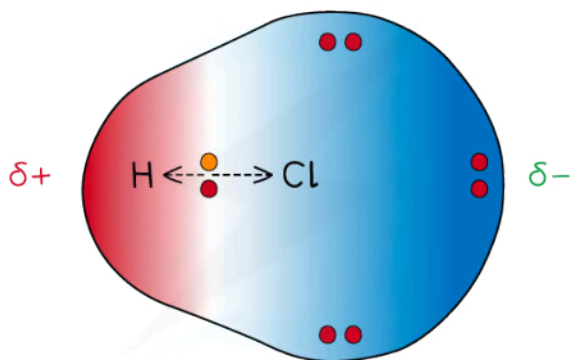
- When two atoms in a covalent bond have the **same electronegativity** the covalent bond is **nonpolar**



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*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  $\delta+$  (delta positive)
  - The **more electronegative** atom gets a partial charge of  $\delta-$  (delta negative)
- The greater the difference in **electronegativity** the more polar the bond becomes



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*Cl has a greater electronegativity than H causing the electrons to be more attracted towards the Cl atom which becomes delta negative and the H delta positive*



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## 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.

$$\mathbf{C = 2.6 \quad N = 3.0 \quad O = 3.4 \quad F = 4.0}$$

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

- A.  $\text{CO} < \text{OF}_2 < \text{NO} < \text{CF}_4$
- B.  $\text{NO} < \text{OF}_2 < \text{CO} < \text{CF}_4$
- C.  $\text{CF}_4 < \text{CO} < \text{OF}_2 < \text{NO}$
- D.  $\text{CF}_4 < \text{NO} < \text{OF}_2 < \text{CO}$

#### 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:

$$\text{NO} (3.4 - 3.0 = \mathbf{0.4})$$

$$\text{OF}_2 (4.0 - 3.4 = \mathbf{0.6})$$

$$\text{CO} (3.4 - 2.6 = \mathbf{0.8})$$

$\text{CF}_4$  ( $4.0 - 2.6 = 1.4$ )



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## 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  $\text{CCl}_4$

**Answer:**



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1 TOTAL NUMBER OF VALENCE ELECTRONS = C + 4Cl = 4 + (4 × 7) = 32

2 DRAW THE SKELETAL POSITIONS

3 ADD THE BONDING PAIRS

ADD 24 LONE PAIR ELECTRONS

4 COMPLETED LEWIS STRUCTURE

$\begin{array}{c} \text{Cl} \\ \text{Cl} \text{ C } \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \end{array}$

$\begin{array}{c} \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \end{array}$

$\begin{array}{c} \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \end{array}$

4 BONDING PAIRS MEANS 32 - 8 = 24 ELECTRONS LEFT

$\begin{array}{c} \text{:Cl:} \\ \text{:Cl:} \\ \text{:Cl:} \\ \text{:Cl:} \\ \text{:Cl:} \\ \text{:Cl:} \end{array}$

LONE PAIRS

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### Steps in drawing the Lewis Structure for $\text{CCl}_4$

### Further examples of Lewis structures

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



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Molecule	Total number of valence electrons	Lewis structure
CH <sub>4</sub>	$C + 4H$ $4 + (4 \times 1) = 8$	$\begin{array}{c} H \\ \vdots \\ H : C : H \\ \vdots \\ H \end{array}$
NH <sub>3</sub>	$N + 3H$ $5 + (3 \times 1) = 8$	$\begin{array}{c} \vdots \\ H : N : H \\ \vdots \\ H \end{array}$
H <sub>2</sub> O	$2H + O$ $(2 \times 1) + 6 = 8$	$H : \ddot{O} : H$
CO <sub>2</sub>	$C + 2O$ $4 + (2 \times 6) = 16$	$: \ddot{O} : C : \ddot{O} :$
HCN	$H + C + N$ $1 + 4 + 5 = 10$	$H : C \equiv N :$

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## 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 as 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  $1s^2 2s^1$  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
$BeCl_2$	$Be + 2Cl =$ $2 + (2 \times 7) = 16$	$:\ddot{Cl}:Be:\ddot{Cl}:$
$BF_3$	$B + 3F =$ $3 + (3 \times 7) = 24$	$:\ddot{F}:B:\ddot{F}:$ $:\ddot{F}:$

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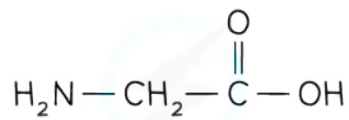
- Test your understanding of Lewis diagrams in the following example:



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 **Worked example**

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



2-AMINOETHANOIC ACID

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- 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<sub>2</sub> and NH<sub>2</sub> groups. Hydrogen does not follow the octet rule.