Chemical Bonding

Chemical Bonding
O' Level Knowledge and Skills

Ionic bonding

(a) Describe the formation of ions by electron loss/gain in order to obtain the electronic configuration of a noble gas.

(b) Describe the formation of ionic bonds between metals and non-metals, *e.g.* $NaCl$, $MgCl_2$.

(c) State that ionic materials contain a giant lattice in which the ions are held by electrostatic attraction, e.g. $NaCl$ (candidates will not be required to draw diagrams of ionic lattices).

(d) Deduce the formulae of other ionic compounds from diagrams of their lattice structures, limited to binary compounds.

(e) Relate the physical properties (including electrical property) of ionic compounds to their lattice structure.

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Chemical Bonding

O' Level Knowledge and Skills

Covalent bonding

(a) Describe the formation of a covalent bond by the sharing of a pair of electrons in order to gain the electronic configuration of a noble gas.

(b) Describe, using 'dot-and-cross' diagrams, the formation of covalent bonds between non-metallic elements, *e.g.* H_2, O_2, H_2O, CH_4, CO_2.

(c) Deduce the arrangement of electrons in other covalent molecules.

(d) Relate the physical properties (including electrical property) of covalent substances to their structure and bonding.

Metallic bonding

(a) Describe metals as a lattice of positive ions in a 'sea of electrons.'

(b) Relate the electrical conductivity of metals to the mobility of the electrons in the structure.

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Chemical Bonding

O' Level Knowledge and Skills

Structure and properties of materials

(a) describe the differences between elements, compounds and mixtures.

(b) Compare the structure of simple molecular substances, *e.g.* methane, iodine, with those of giant molecular substances, *e.g.* poly(ethene), sand (silicon dioxide), diamond, graphite in order to deduce their properties.

(c) Compare the bonding and structures of diamond and graphite in order to deduce their properties such as electrical conductivity, lubricating or cutting action (candidates will not be required to draw the structures).

(d) Deduce the physical and chemical properties of substances from their structures and bonding and vice versa.

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Chemical Bonding

Main Menu *(click to link)*

Chemical Bonding

Main Menu *(click to link)*

Chemical Bonding

Introduction

1. Why do the atoms of the different chemical elements *react* and *bond together* ?

Chemical Bonding

Introduction

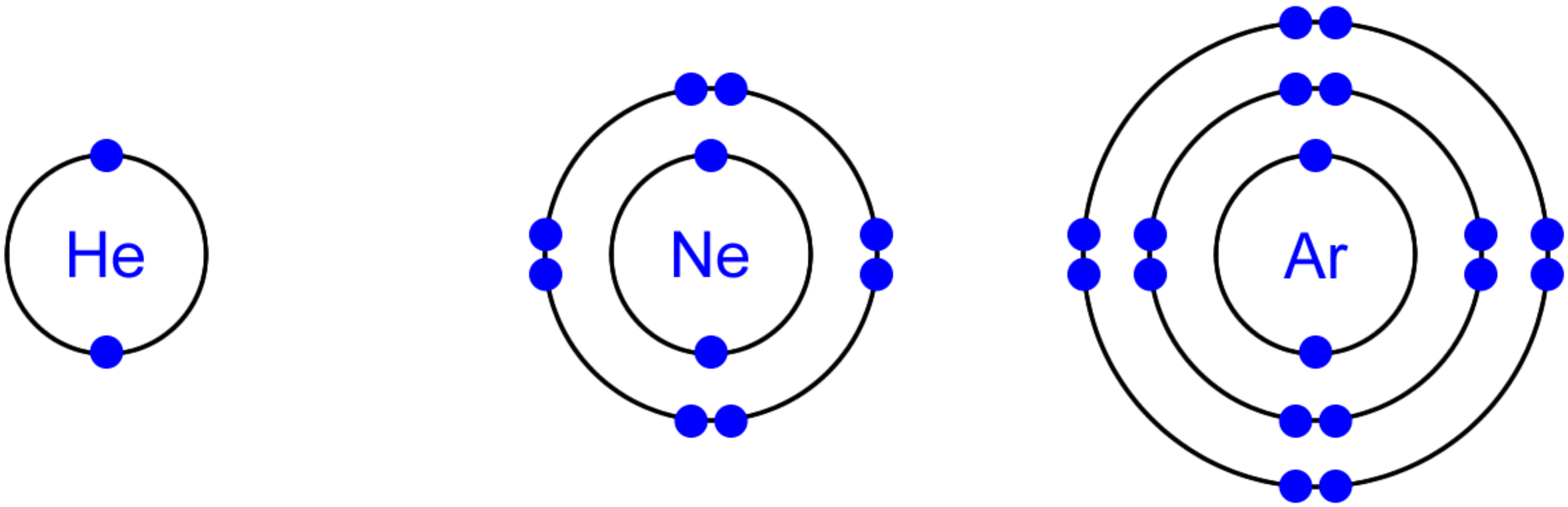
Essential Understanding

- During the formation of a chemical bond, atoms combine together by gaining, losing or sharing electrons in such a way that they acquire the electronic configuration of the nearest noble gas.

Chemical Bonding

Introduction

- Helium, neon and argon are noble gases which are all inert (unreactive) chemical elements.

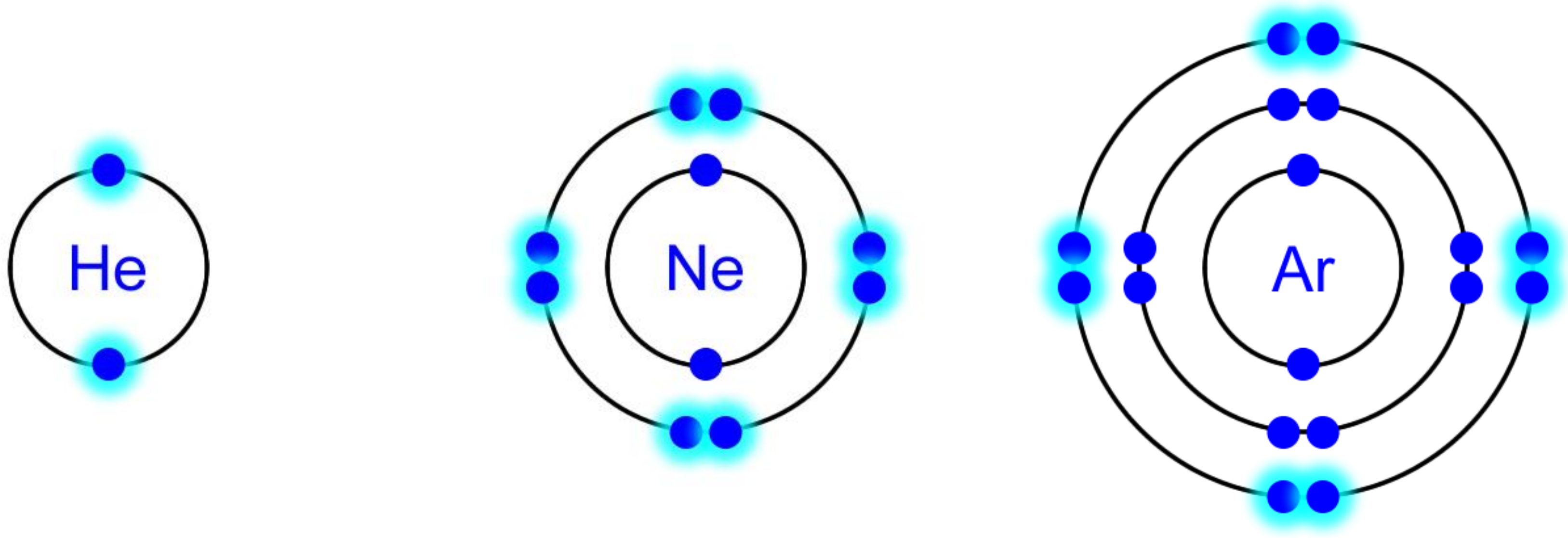


- What do their electronic configurations all have in common?
- What aspect of their electronic configurations could cause them to be so unreactive?

Chemical Bonding

Introduction

- Helium, neon and argon are noble gases which are all inert (unreactive) chemical elements.



- They all have *complete* (full) *valence shells*.
- Complete valence shells cause the atoms of helium, neon and argon to be *stable*, *low energy* and hence *unreactive*.

Chemical Bonding

Introduction

- Unlike helium, neon and argon, sodium is a very reactive chemical element.

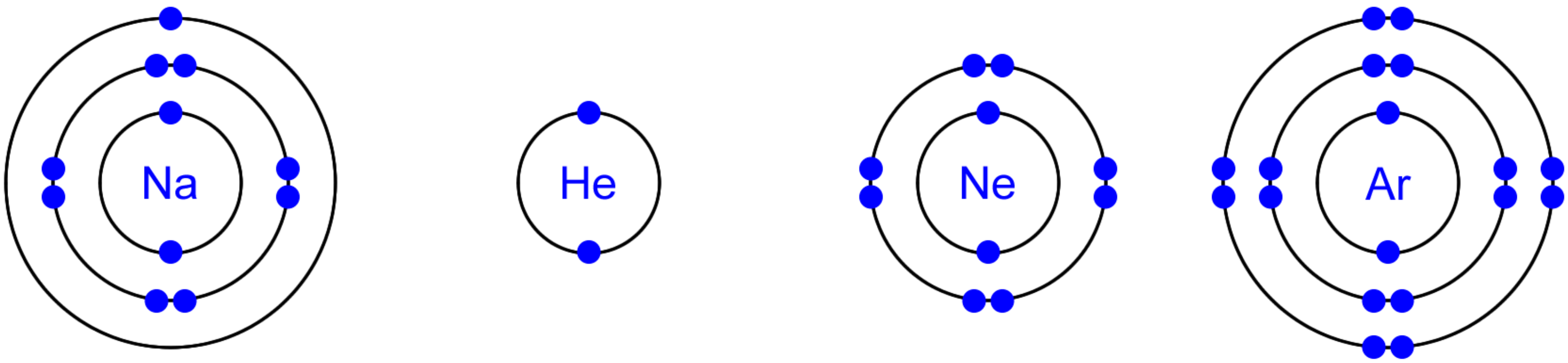
Sodium metal reacting with water at room temperature:

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

Chemical Bonding

Introduction

- Unlike helium, neon and argon, sodium is a very reactive chemical element.

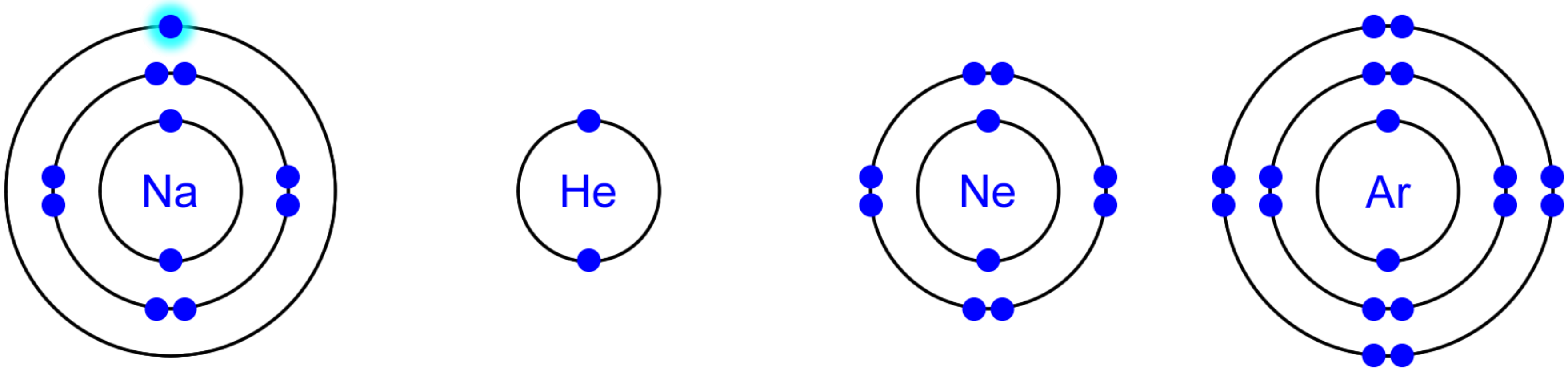


- In what way(s) is sodium's electronic configuration different to that of helium, neon and argon?
- Why is sodium a reactive chemical element?
- What do atoms of sodium achieve by reacting?

Chemical Bonding

Introduction

- Unlike helium, neon and argon, sodium is a very reactive chemical element.

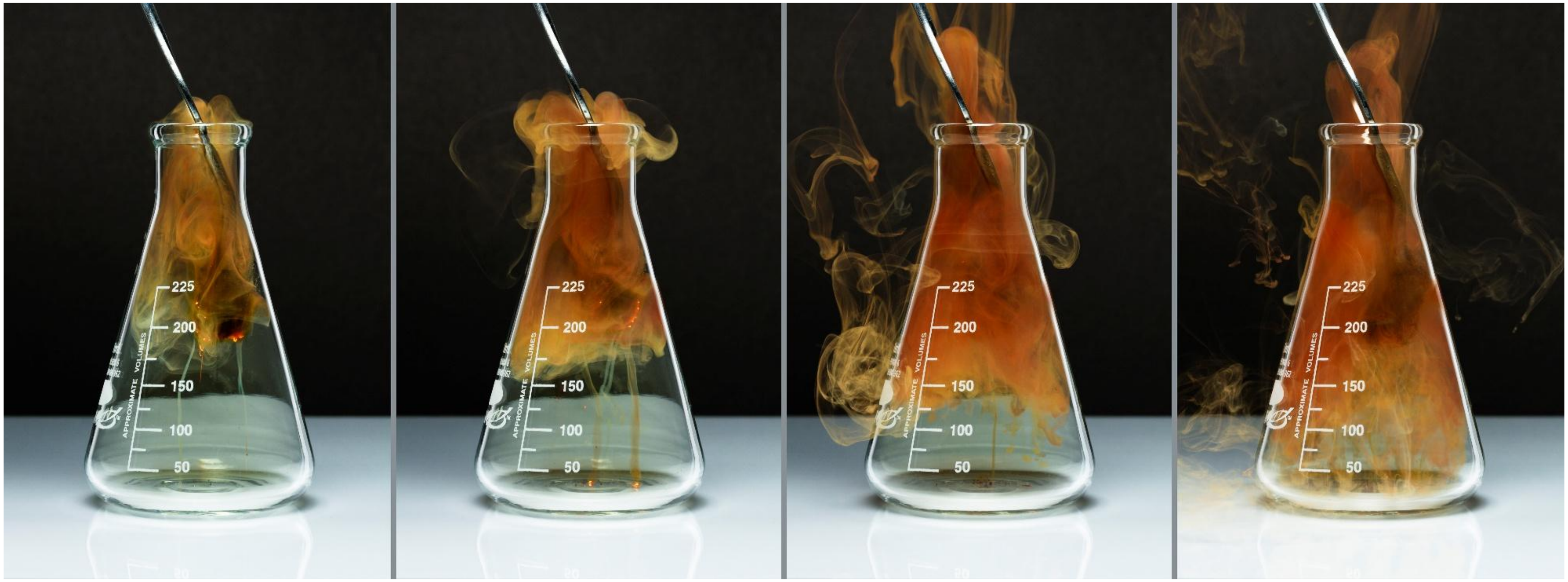


- A sodium atom has an *incomplete* valence shell.
- Due to its incomplete valence shell, a sodium atom is relatively *unstable*, *high energy* and hence *reactive*.
- A sodium atom will react in order to obtain a *noble gas electronic configuration*.

Chemical Bonding

Introduction

- Unlike helium, neon and argon, chlorine is a very reactive chemical element.



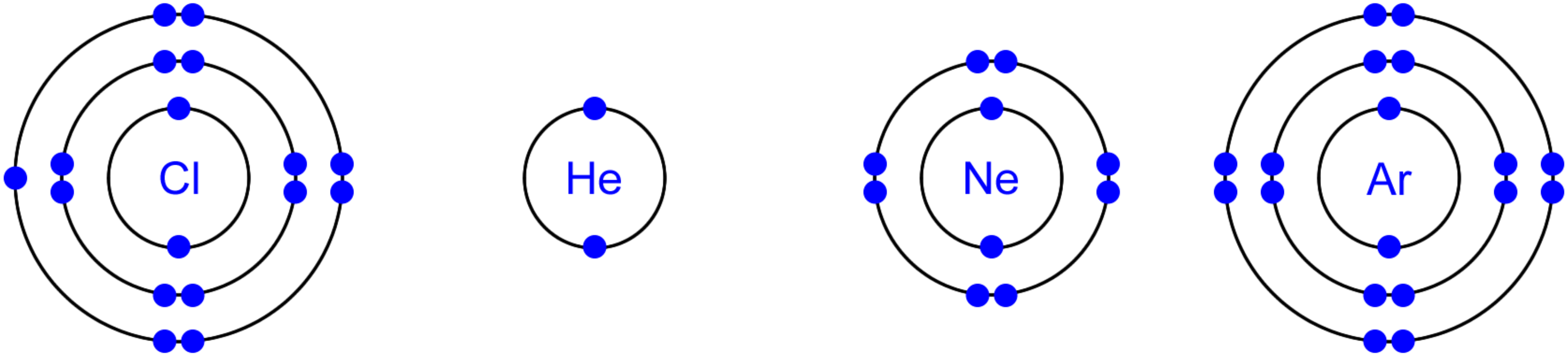
Iron wool reacting with chlorine gas at room temperature:

$$2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$$

Chemical Bonding

Introduction

- Unlike helium, neon and argon, chlorine is a very reactive chemical element.

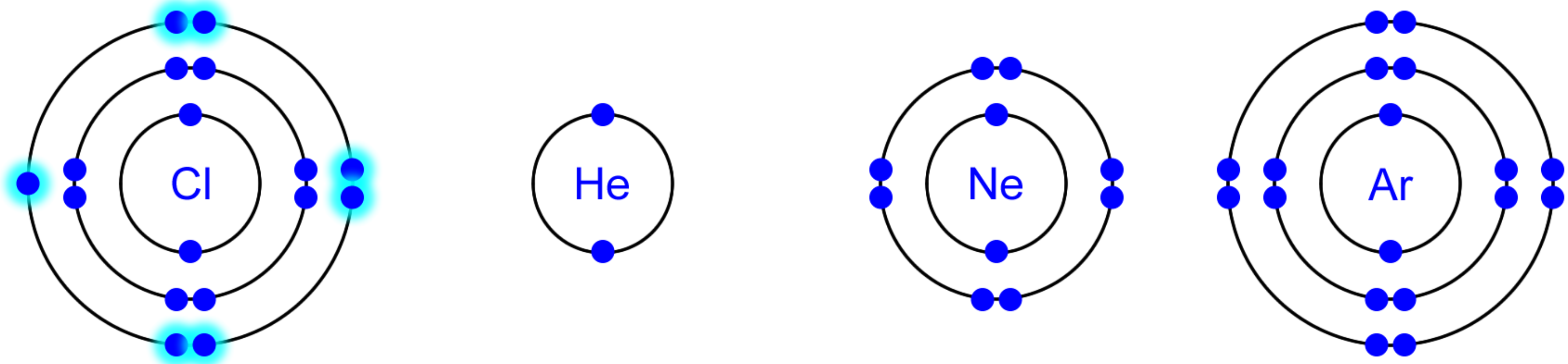


- In what way(s) is chlorine’s electronic configuration different to that of helium, neon and argon?
- Why is chlorine a reactive chemical element?
- What do atoms of chlorine achieve by reacting?

Chemical Bonding

Introduction

- Unlike helium, neon and argon, chlorine is a very reactive chemical element.



- A chlorine atom has an *incomplete* valence shell.
- Due to its incomplete valence shell, a chlorine atom is relatively *unstable*, *high energy* and hence *reactive*.
- A chlorine atom will react in order to obtain a *noble gas electronic configuration*.

Chemical Bonding

Introduction

- We will see how *sodium* and *chlorine* react to form a compound in which particles of each chemical element have stable electronic configurations shortly.

Chemical Bonding

Introduction

Objective of chemical bonding:

• At the end of a chemical reaction, all of the atoms or ions that are present *will have noble gas electronic configurations*, *i.e.* the outermost electron shell of every atom or ion must be filled with its maximum number of electrons.

• Inner electron shell = 2.
• Second Electron Shell = 8.
• Third Electron Shell = 8.

• By obtaining a noble gas electronic configuration, the atom of the chemical element becomes a *stable*, *low energy* particle.

Chemical Bonding

Introduction

There are three different types of chemical bonding:

- Ionic bonding (metal and non-metal).
- Covalent bonding (two or more non-metals).
- Metallic bonding.

Note: It is very important to remember that chemical bonding only involves rearranging an atom's *valence shell electrons*. The numbers of protons and neutrons in the nucleus of the atom are not affected by chemical bonding.

Chemical Bonding

Ionic Bonding – Metals and Non-metals

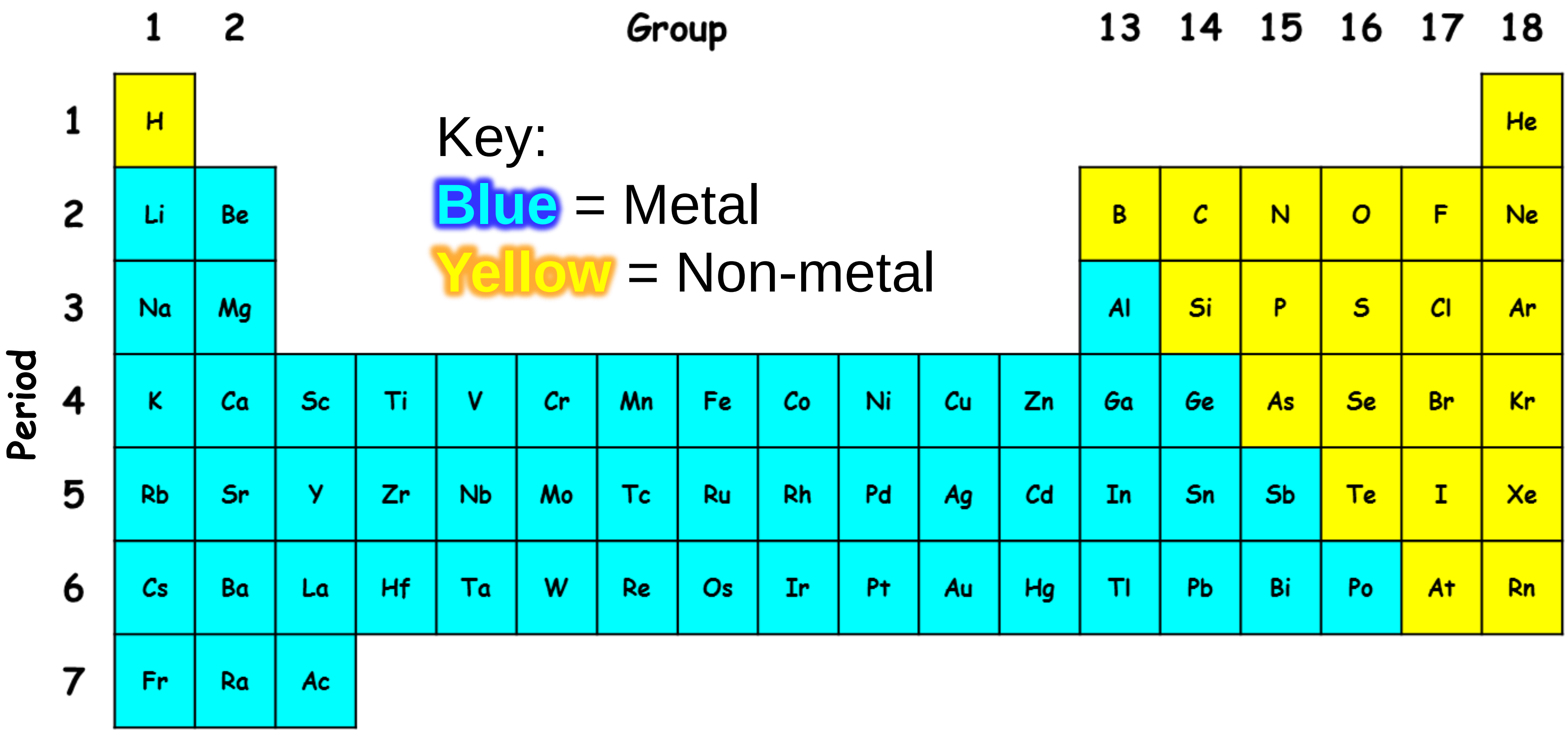
2. How do the atoms of a *metallic chemical element* and a *non-metallic chemical elements* react together to obtain noble gas electronic configurations?

Main Menu

Chemical Bonding

Ionic Bonding – Metals and Non-metals

- Ionic bonding occurs in compounds that are formed when a *metal* reacts with a *non-metal*.



Chemical Bonding

Ionic Bonding – Sodium Chloride

- The reaction between sodium and chlorine to form sodium chloride. Sodium is a very reactive metal and chlorine is a very reactive non-metal, yet the product of their reaction – sodium chloride – is safe enough to eat.

Sodium metal reacting chlorine gas at room temperature:
$2Na_{(s)} + Cl_{2(g)} \quad 2NaCl_{(s)}$

Chemical Bonding

Ionic Bonding – Sodium Chloride

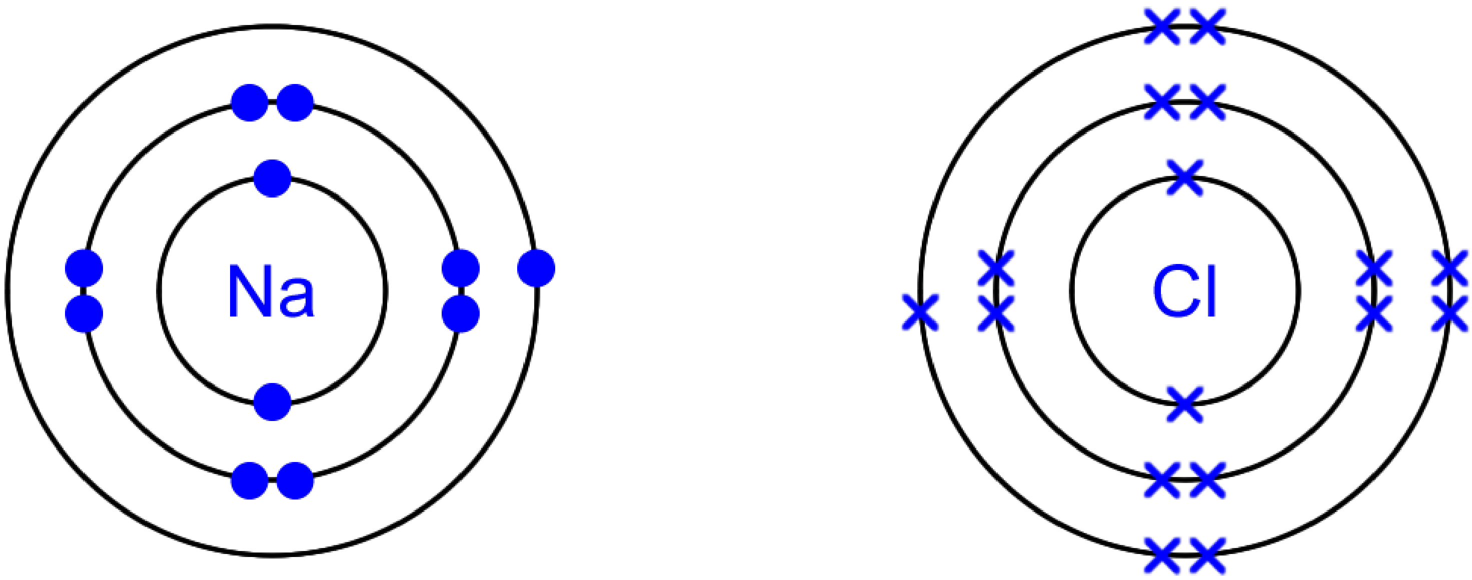
- Notice that the reaction between sodium and chlorine is highly *exothermic*, *i.e.* energy is lost to the surroundings as the sodium atoms and chlorine atoms react to obtain noble gas electronic configurations and become more stable.

Sodium metal reacting chlorine gas at room temperature:

$2Na_{(s)} + Cl_{2(g)} \quad 2NaCl_{(s)}$

Chemical Bonding

Ionic Bonding – Sodium Chloride

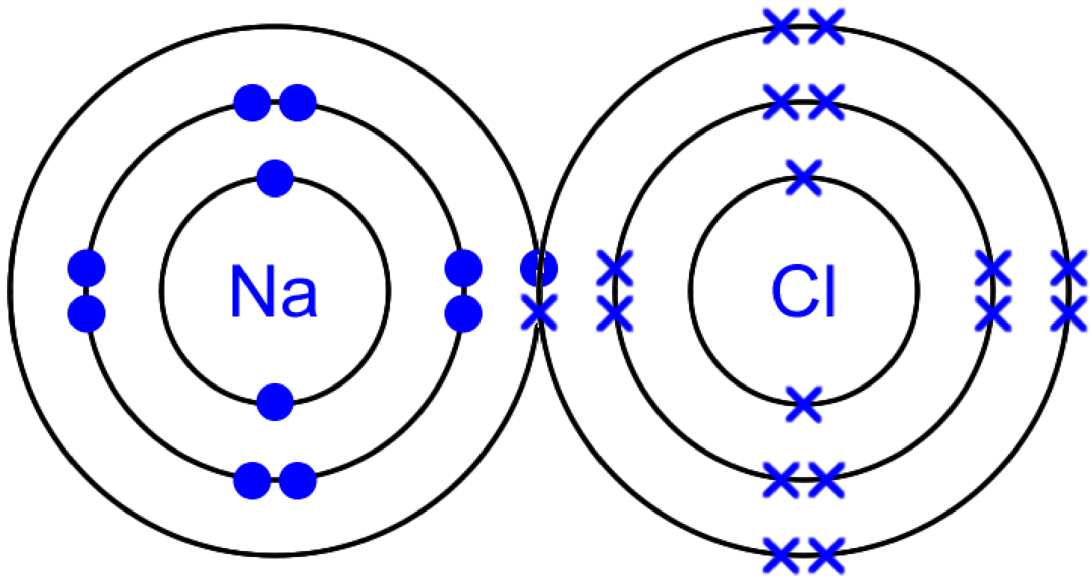


- Can an atom of sodium and an atom of chlorine *both* obtain noble gas electronic configurations by *sharing electrons*?

• = Electron of sodium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Sodium Chloride

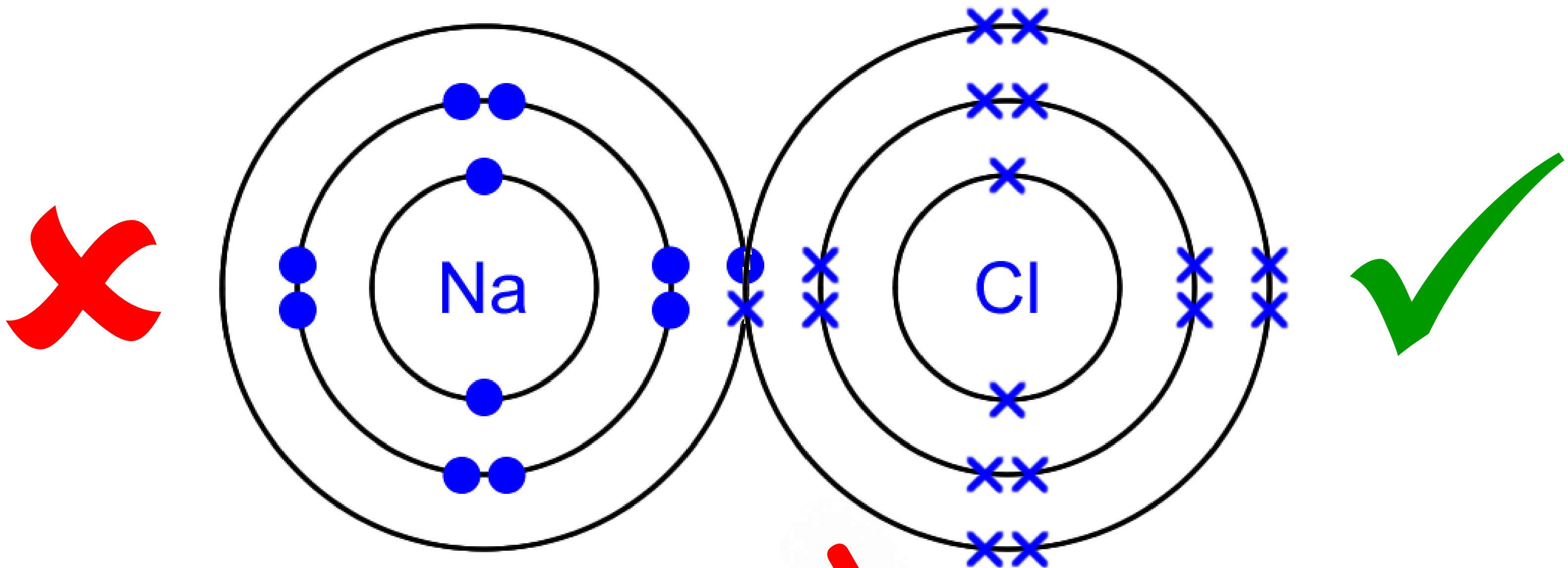


- Can an atom of sodium and an atom of chlorine *both* obtain noble gas electronic configurations by *sharing electrons*?

• = Electron of sodium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Sodium Chloride



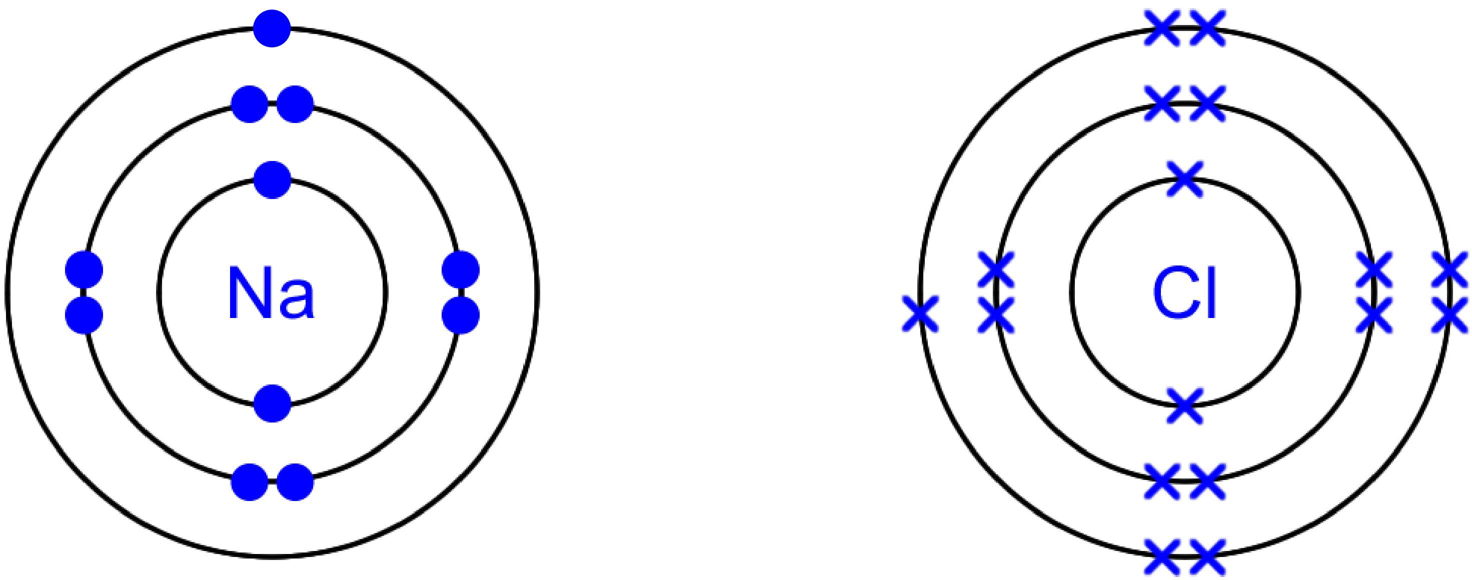
- Can an atom of sodium and an atom of chlorine *both* obtain noble gas electronic configurations by *sharing electrons*?

No!

• = Electron of sodium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Sodium Chloride

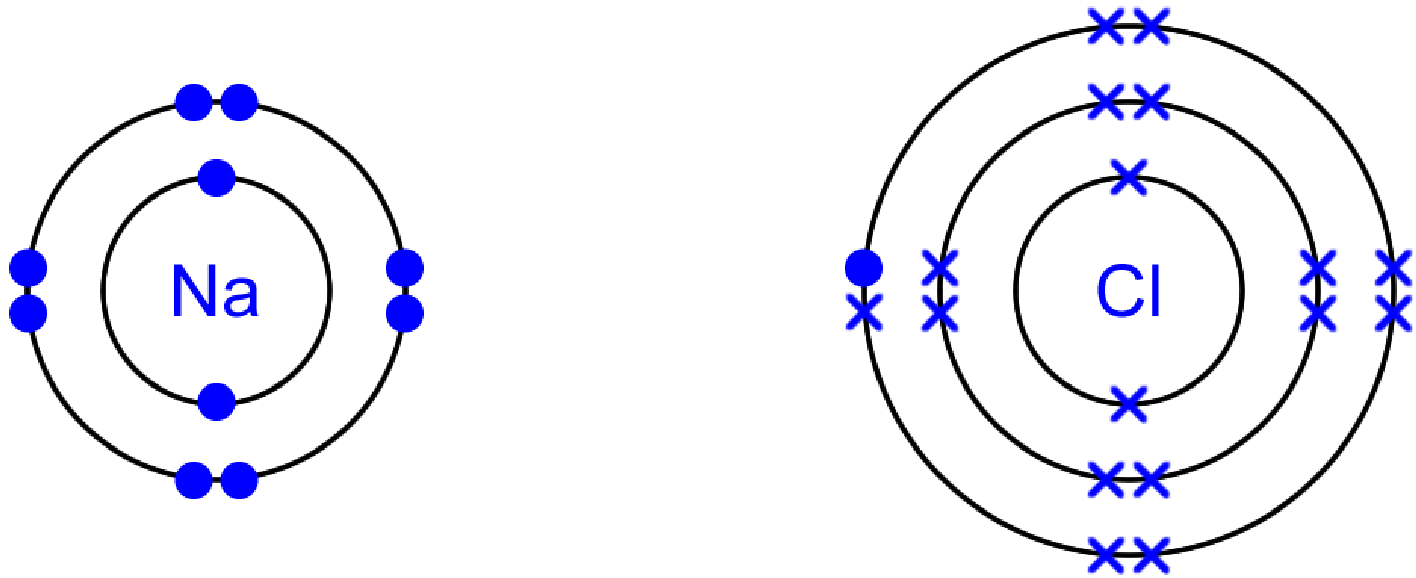


- In ionic bonding, the atom of the metallic element *transfers* it valence electron(s) into the valence shell of the non-metallic element.

• = Electron of sodium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Sodium Chloride



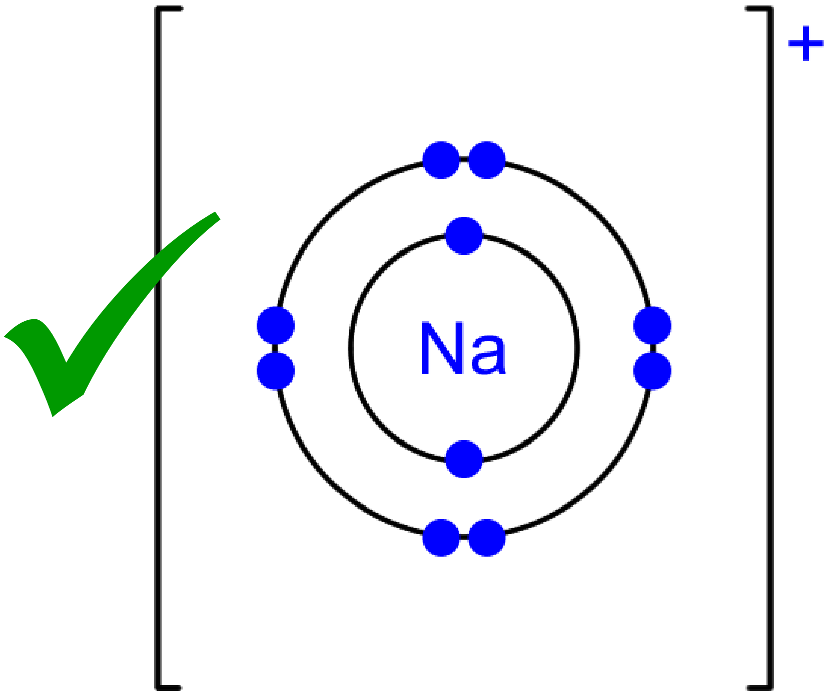
• In ionic bonding, the atom of the metallic element *transfers* it valence electron(s) into the valence shell of the non-metallic element.

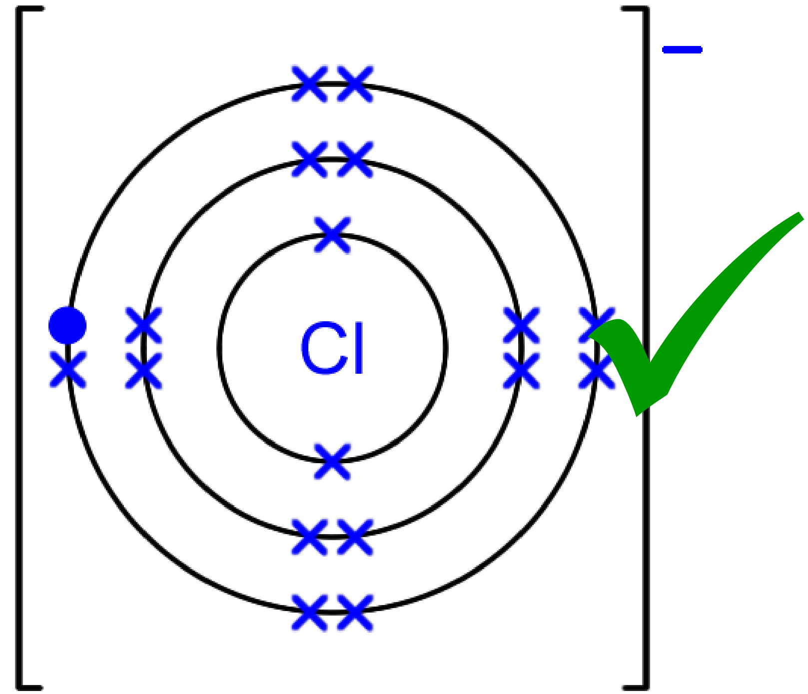
• = Electron of sodium.

× = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Sodium Chloride – NaCl





Positive Sodium Ion (Cation)

- Number of Protons = 11
- Number of Electrons = 10
- Overall Charge = (+11) + (-10) = +1

Negative Chloride Ion (Anion)

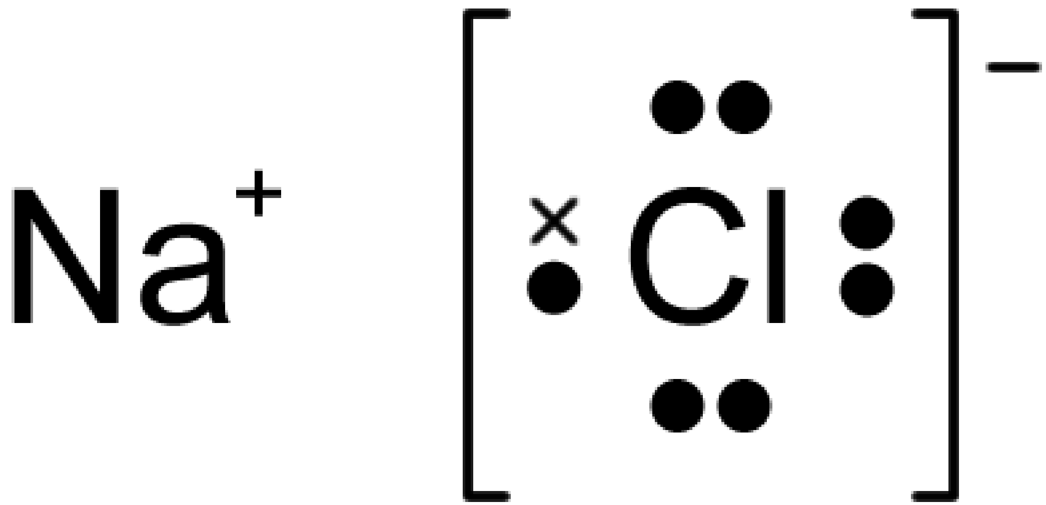
- Number of Protons = 17
- Number of Electrons = 18
- Overall Charge = (+17) + (-18) = -1

• = Electron of sodium.

× = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Sodium Chloride – NaC*l*

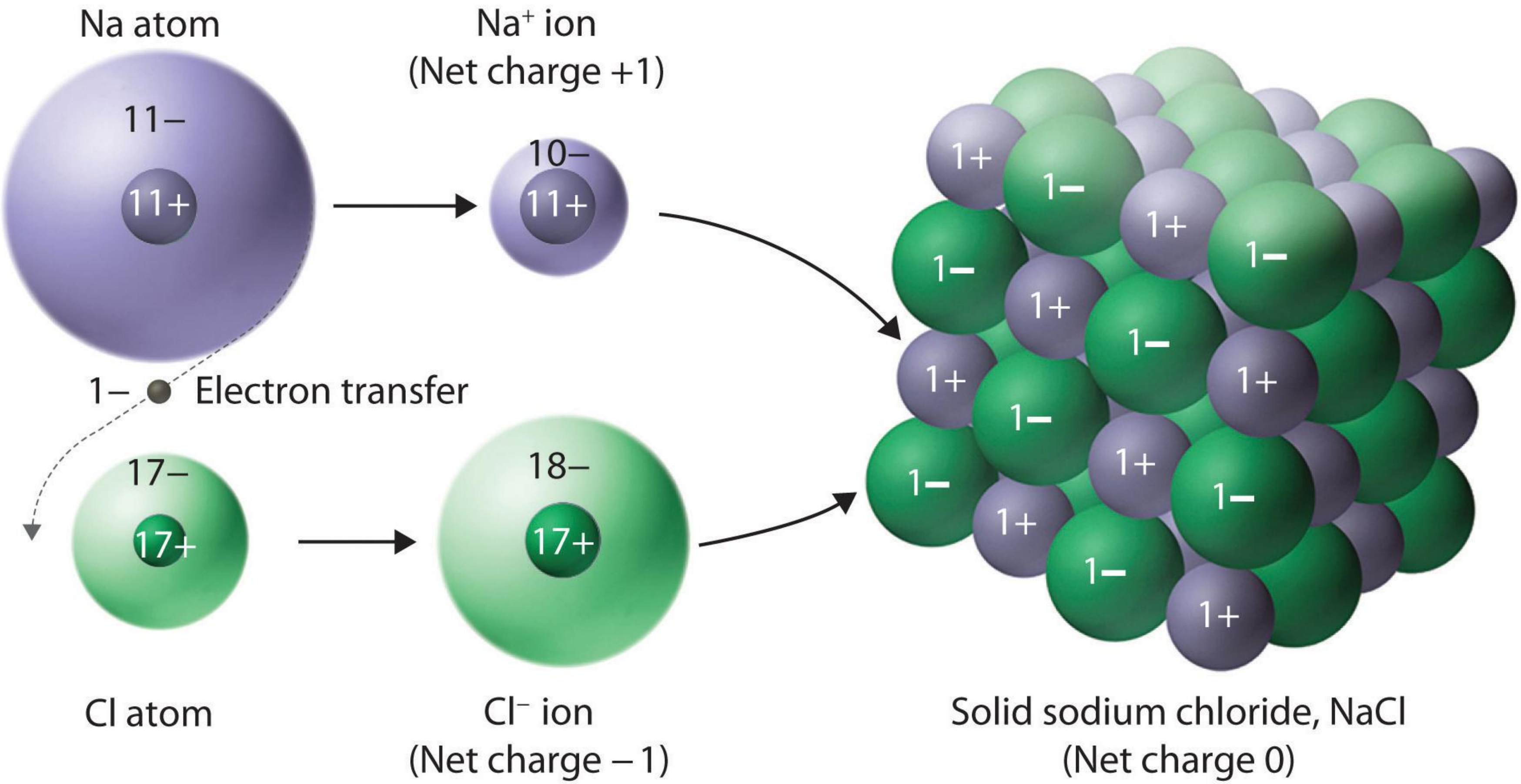


- The dot-and-cross diagram for sodium chloride that should be used as an answer in an examination.

• = Electron of chlorine. × = Electron of sodium.

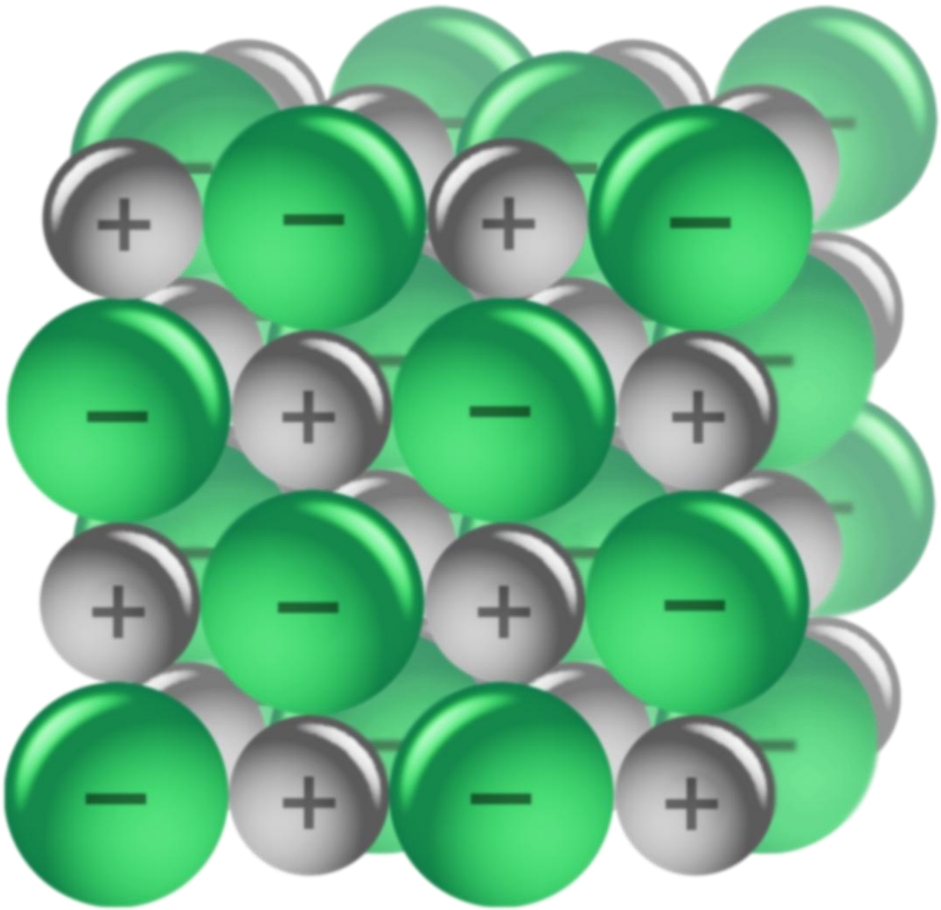
Chemical Bonding

Ionic Bonding – Sodium Chloride – NaC*l*



Chemical Bonding

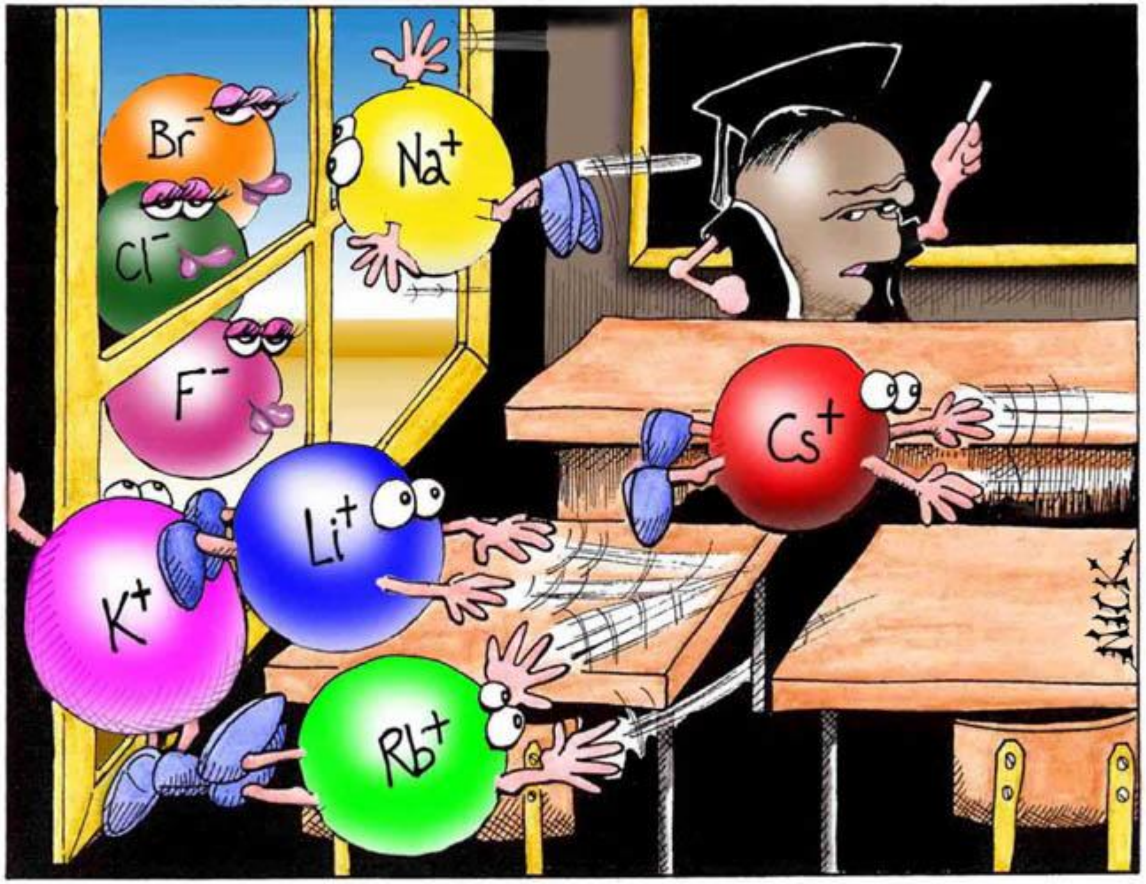
Ionic Bonding – Sodium Chloride – NaC*l*



Ions of opposite charge are attracted towards each other in a three dimensional arrangement. A strong *electrostatic force of attraction* holds the ions together in a close packed regular arrangement known as a *crystal lattice*. Each sodium ion is surrounded by six chloride ions and vice-versa. **Note**: *Never* describe any ionic compound as a "*molecule*!"

Chemical Bonding

Ionic Bonding – Sodium Chloride – NaC*l*



"Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive..?"

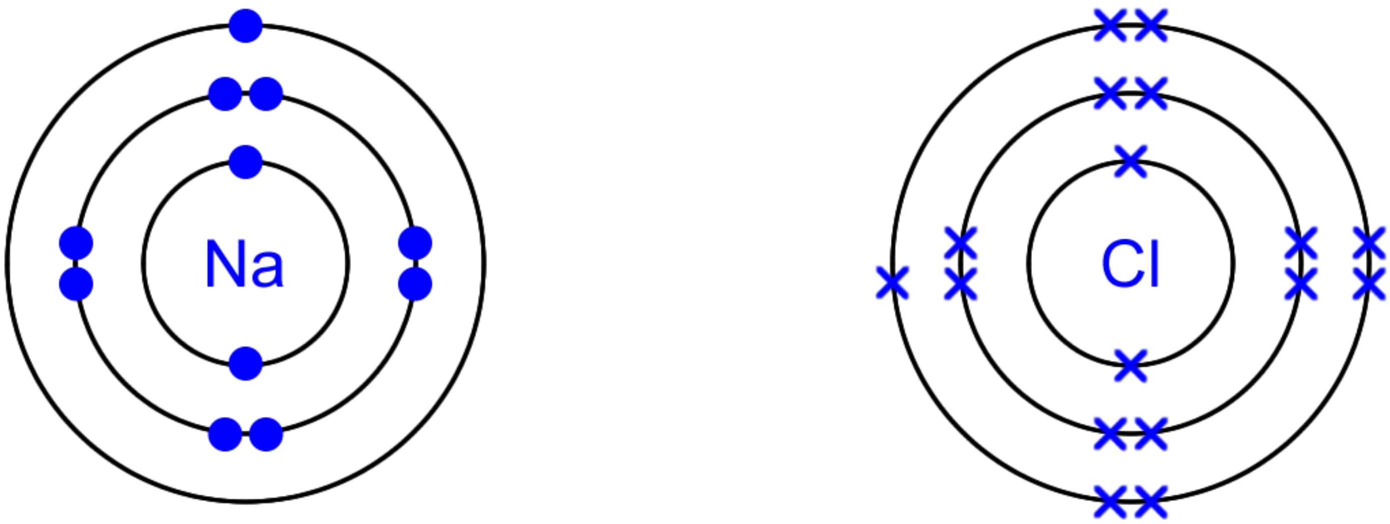
Chemical Bonding

Ionic Bonding – Sodium Chloride – NaCl

Why are electrons transferred from the *metal* to the *non-metal* and not from the *non-metal* to the *metal*?

Chemical Bonding

Ionic Bonding – Sodium Chloride – NaC*l*

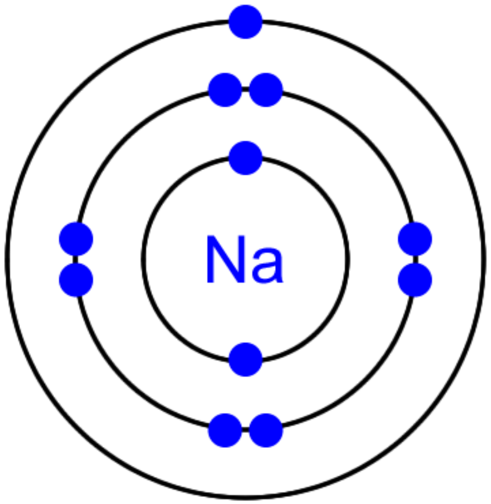


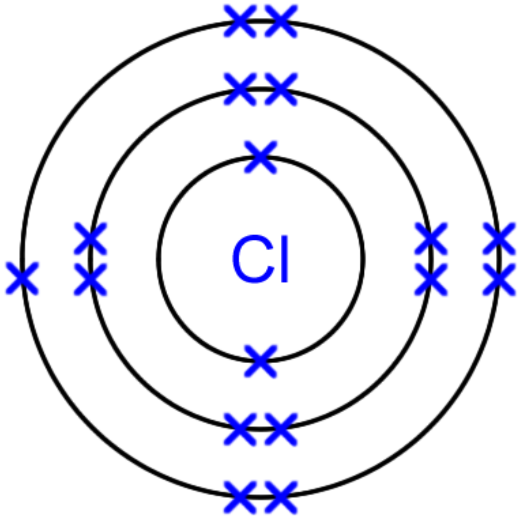
- Consider the reaction between sodium (metal) and chlorine (non-metal).

- Only *one* electron needs to be transferred from the sodium to the chlorine in order for both atoms to obtain noble gas electronic configurations.

Chemical Bonding

Ionic Bonding – Sodium Chloride – NaC*l*



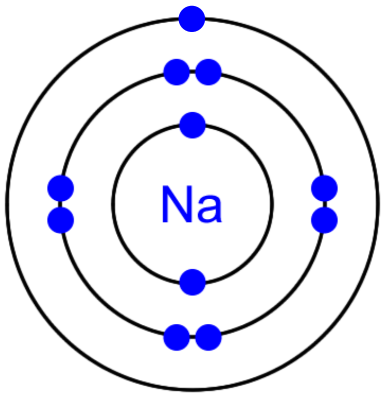


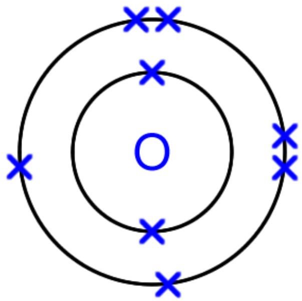
- However, *seven* electrons would need to be transferred from the chlorine to the sodium in order for both atoms to obtain a noble gas electronic configuration.

- It is easier to transfer *one* electron from sodium to chlorine compared to transferring *seven* electrons from chlorine to sodium.

Chemical Bonding

Ionic Bonding – Sodium Oxide



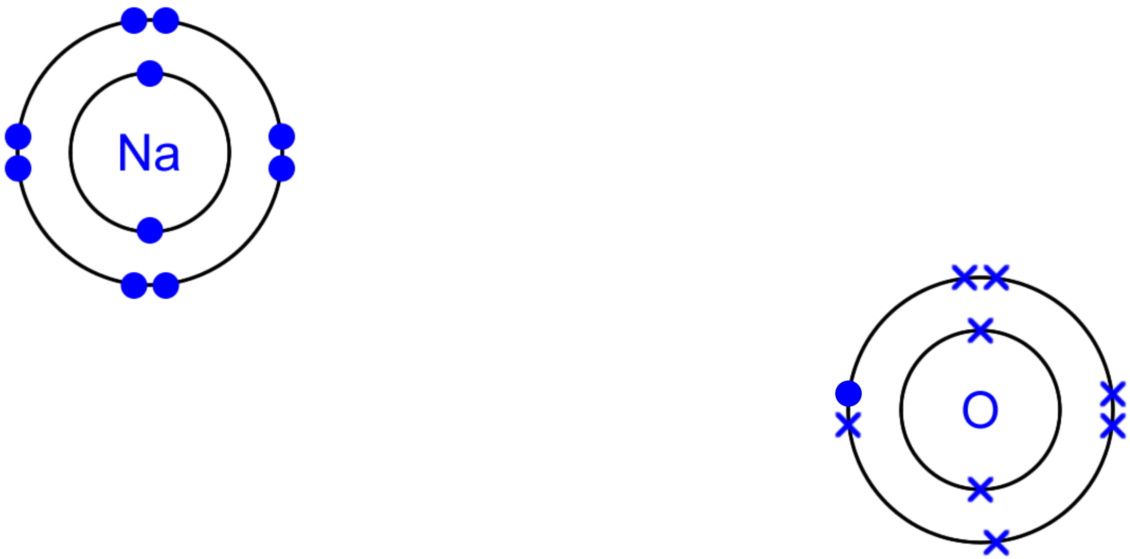


● = Electron of sodium.

× = Electron of oxygen.

Chemical Bonding

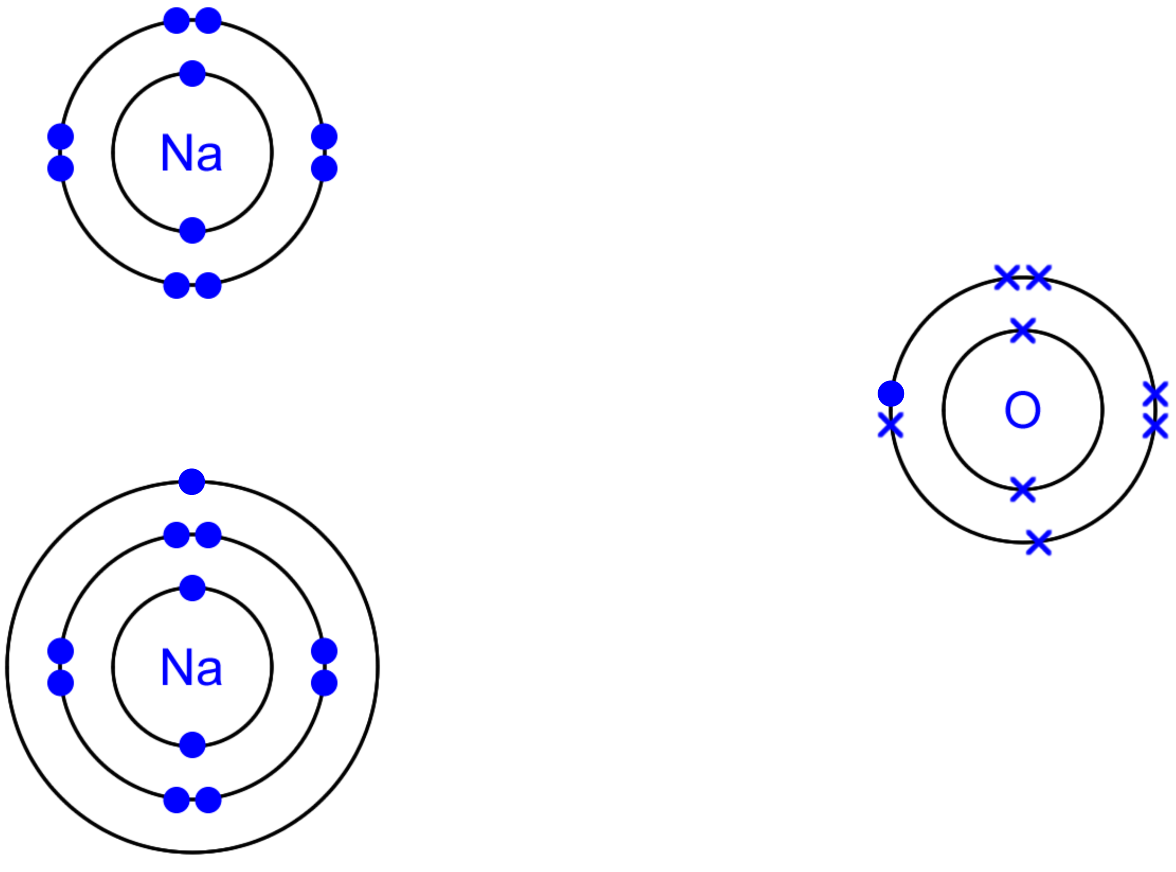
Ionic Bonding – Sodium Oxide



● = Electron of sodium. × = Electron of oxygen.

Chemical Bonding

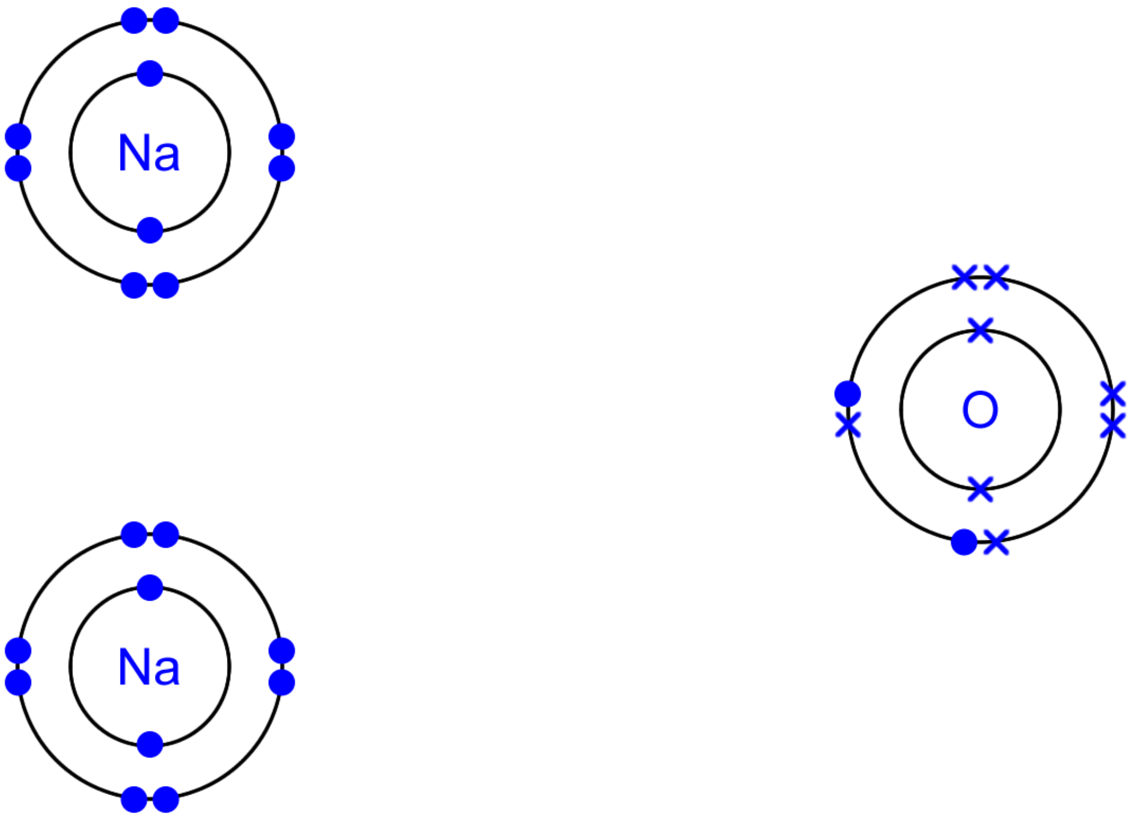
Ionic Bonding – Sodium Oxide



● = Electron of sodium. × = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Sodium Oxide

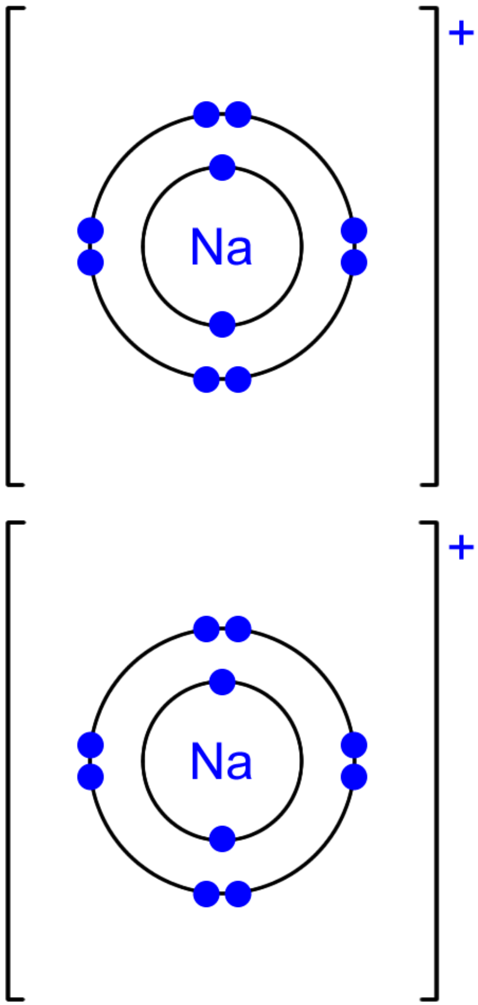


• = Electron of sodium.

× = Electron of oxygen.

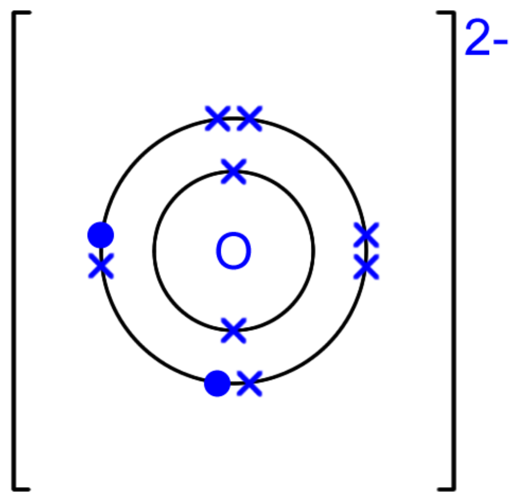
Chemical Bonding

Ionic Bonding – Sodium Oxide – Na_2O



Positive Sodium Ion (Cation)

- Number of Protons = 11
- Number of Electrons = 10
- Overall Charge = (+11) + (-10) = +1



Negative Oxide Ion (Anion)

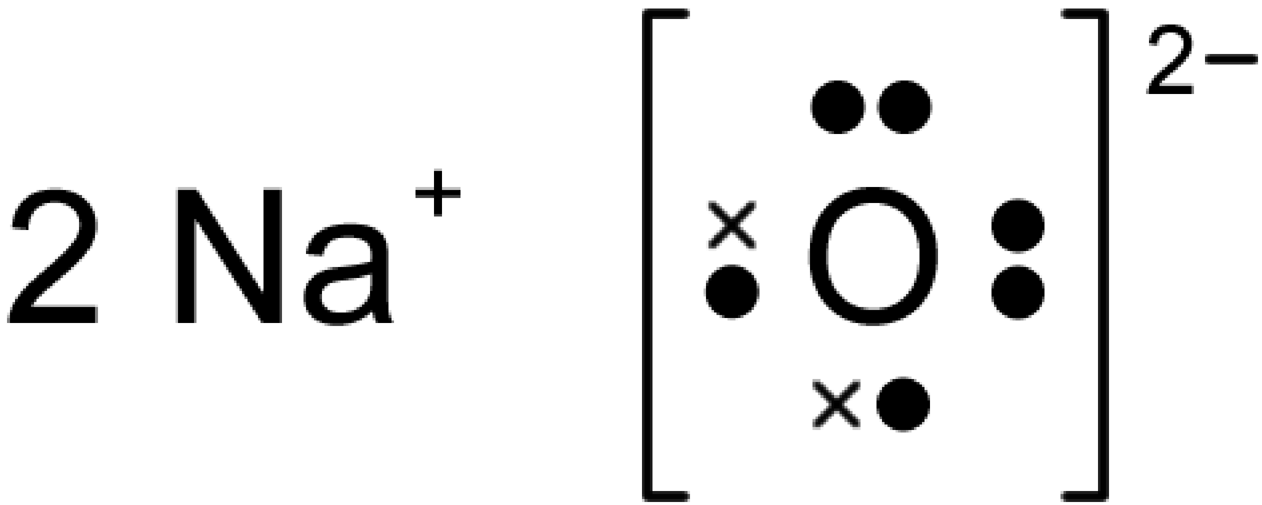
- Number of Protons = 8
- Number of Electrons = 10
- Overall Charge = (+8) + (-10) = -2

• = Electron of sodium.

× = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Sodium Oxide – Na_2O

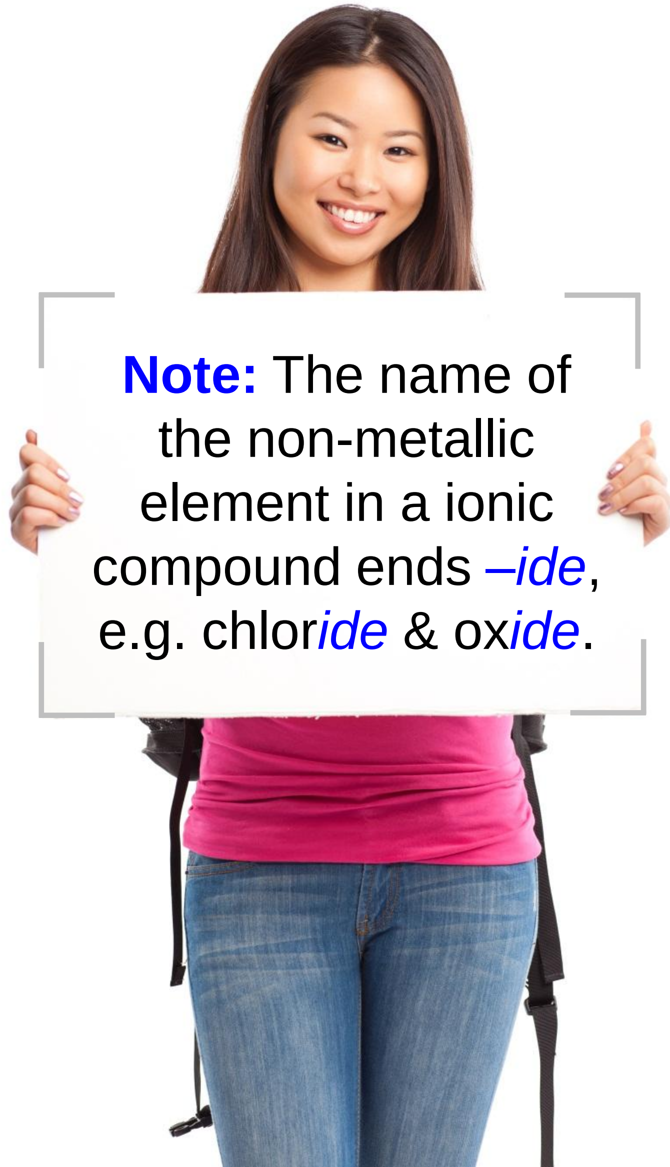


- The dot-and-cross diagram for sodium oxide that should be used as an answer in an examination.

• = Electron of oxygen. × = Electron of sodium.

Chemical Bonding

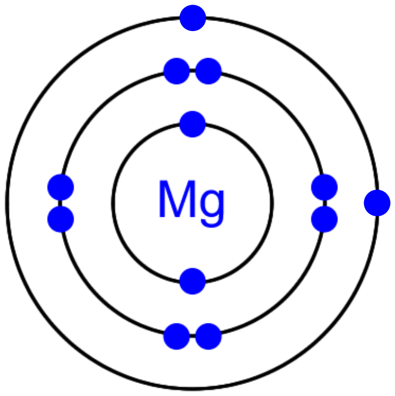
Ionic Bonding – A Note About Names

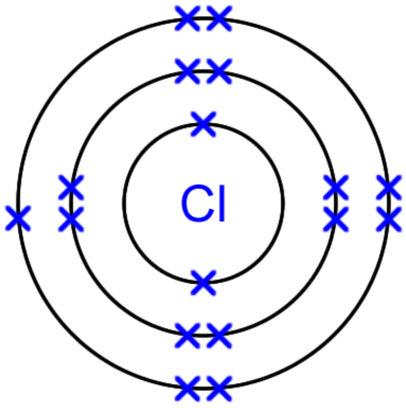


Note: The name of the non-metallic element in a ionic compound ends *–ide*, e.g. chlor*ide* & ox*ide*.

Chemical Bonding

Ionic Bonding – Magnesium Chloride

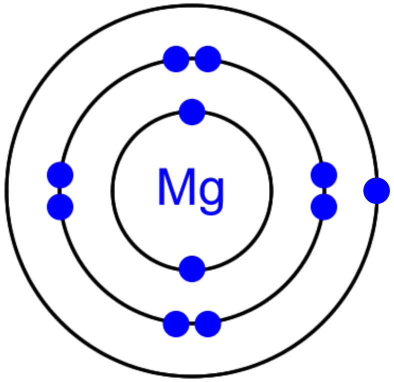


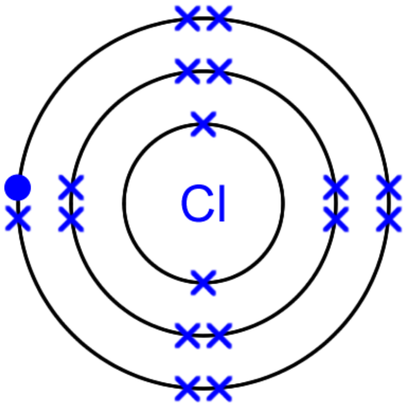


● = Electron of magnesium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Magnesium Chloride

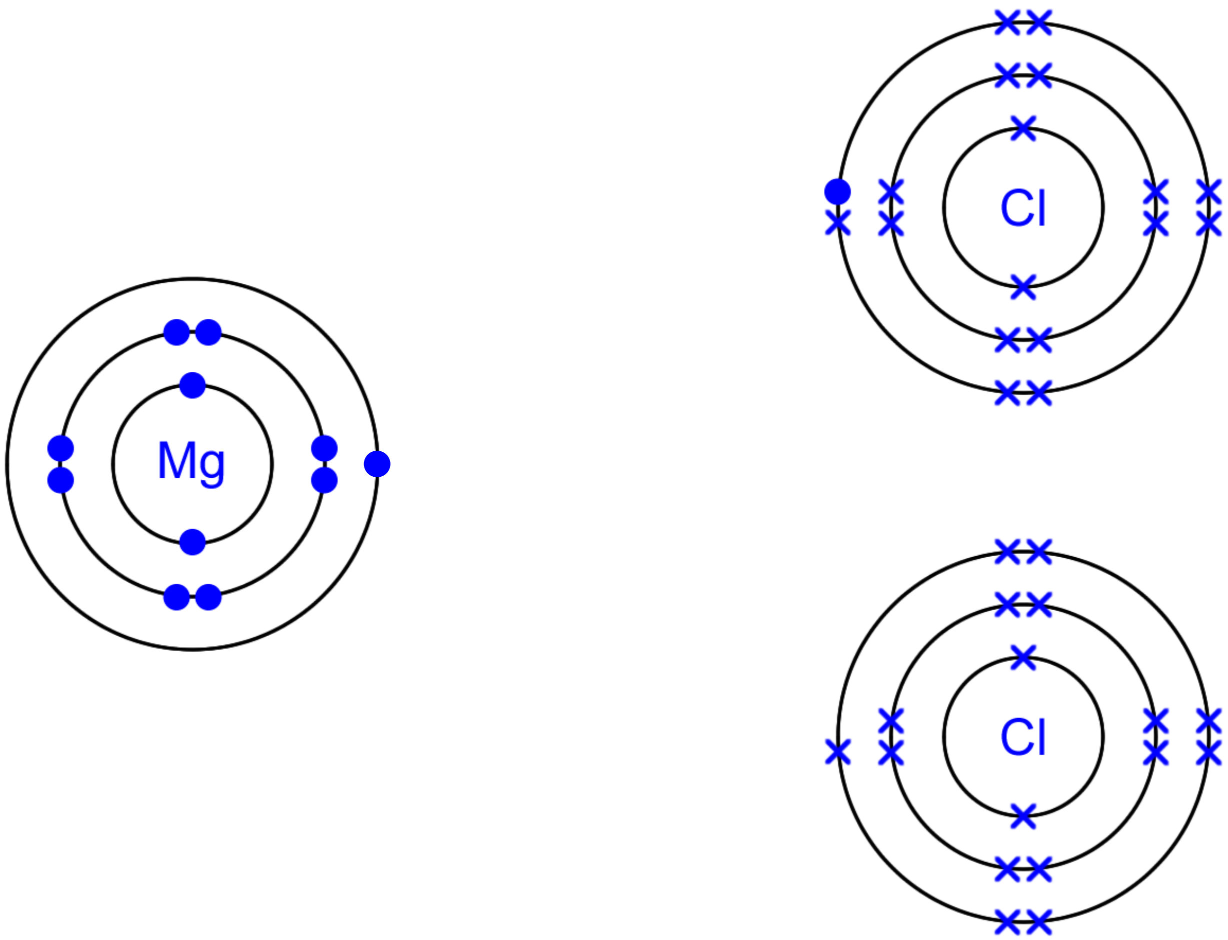




● = Electron of magnesium. × = Electron of chlorine.

Chemical Bonding

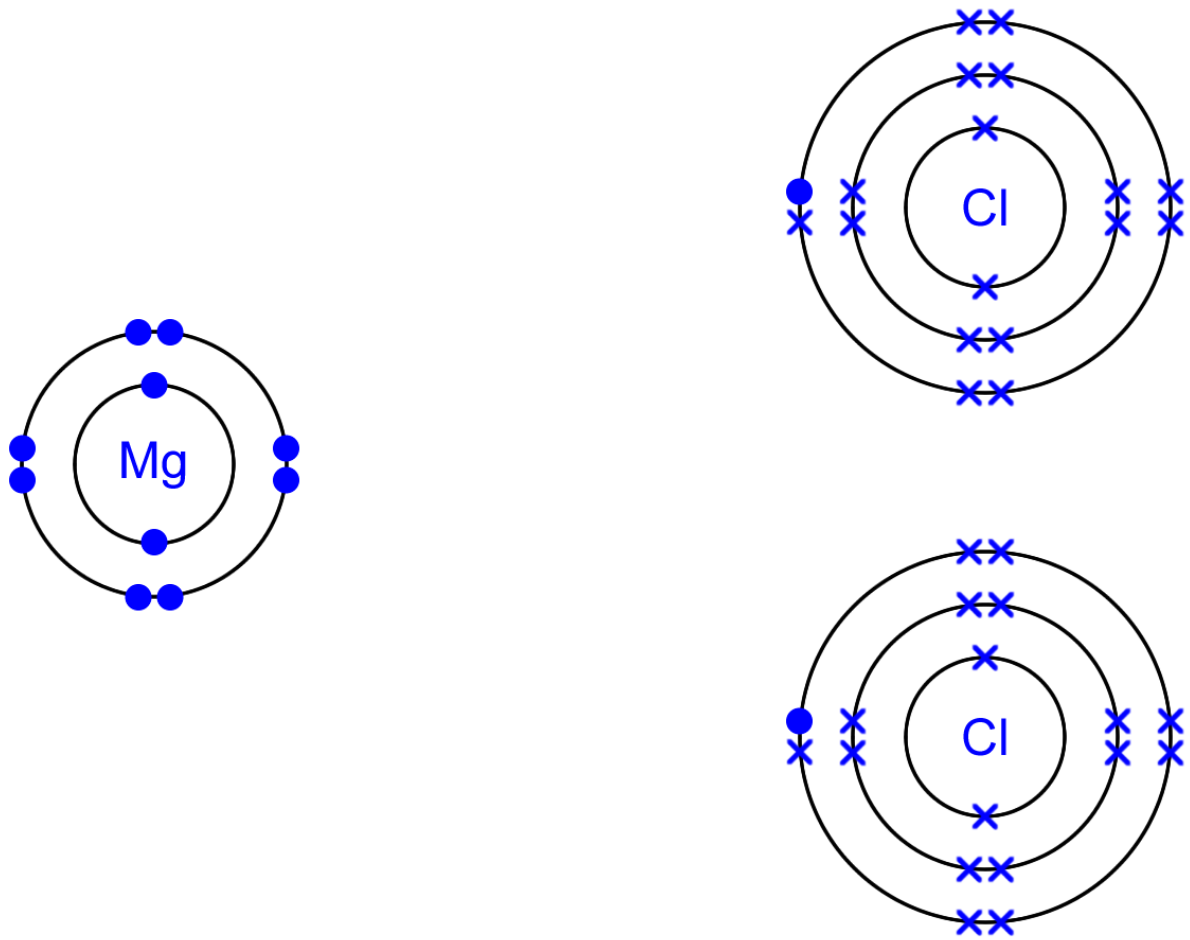
Ionic Bonding – Magnesium Chloride



• = Electron of magnesium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Magnesium Chloride



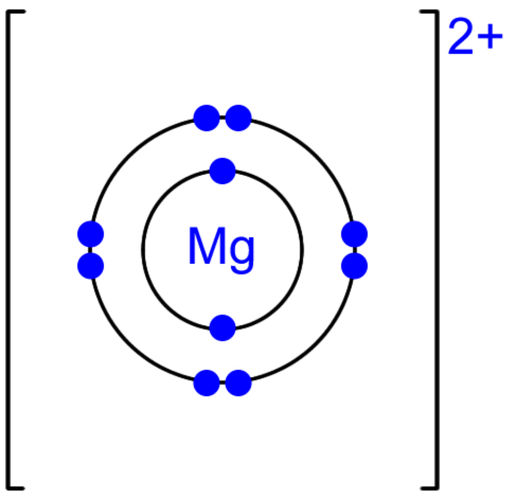
• = Electron of magnesium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Magnesium Chloride – $MgCl_2$

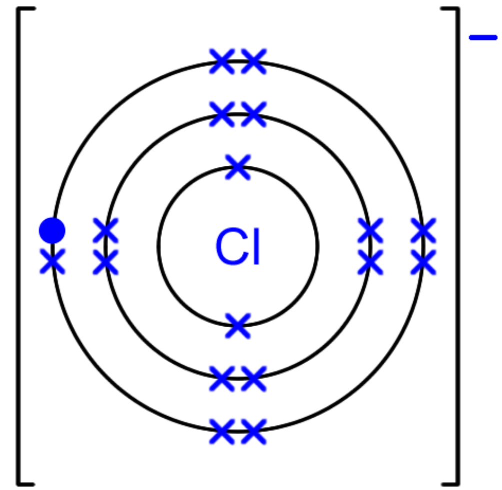
Positive Magnesium Ion (Cation)

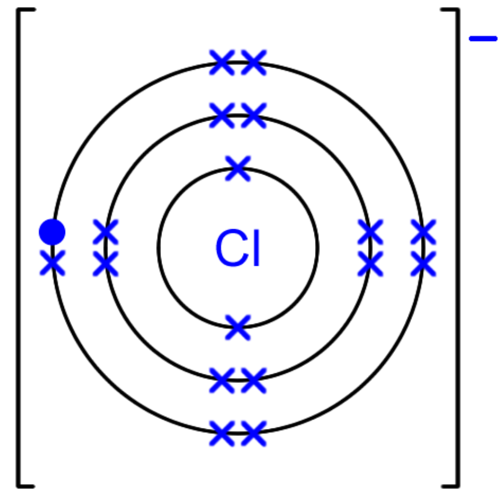
- Number of Protons = 12
- Number of Electrons = 10
- Overall Charge = (+12) + (-10) = +2



Negative Chloride Ion (Anion)

- Number of Protons = 17
- Number of Electrons = 18
- Overall Charge = (+17) + (-18) = -1

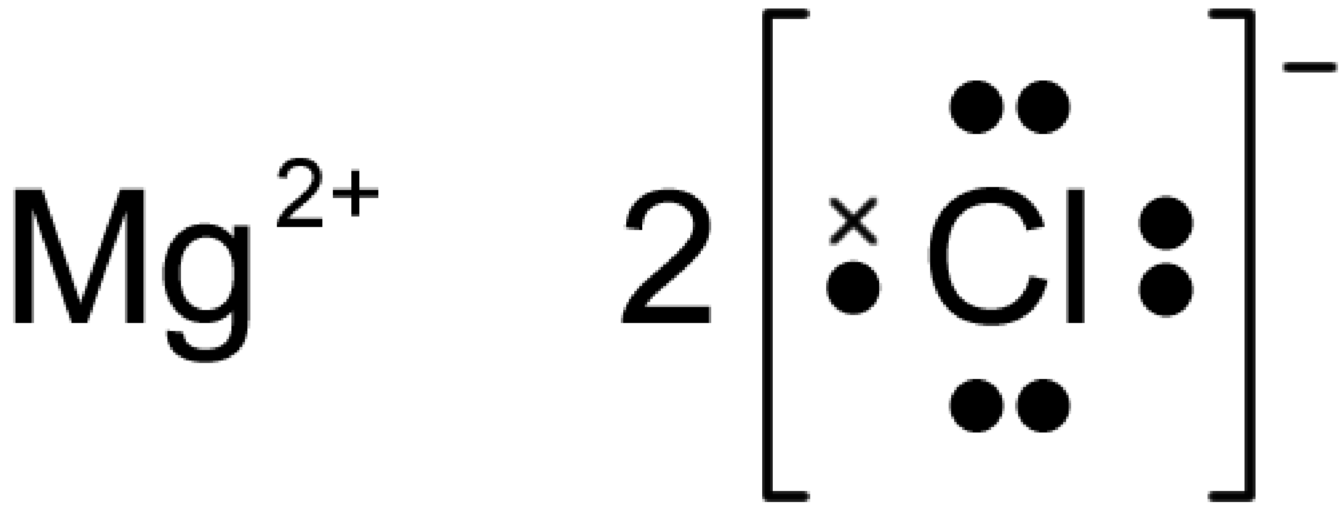




• = Electron of magnesium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Magnesium Chloride – $MgCl_2$

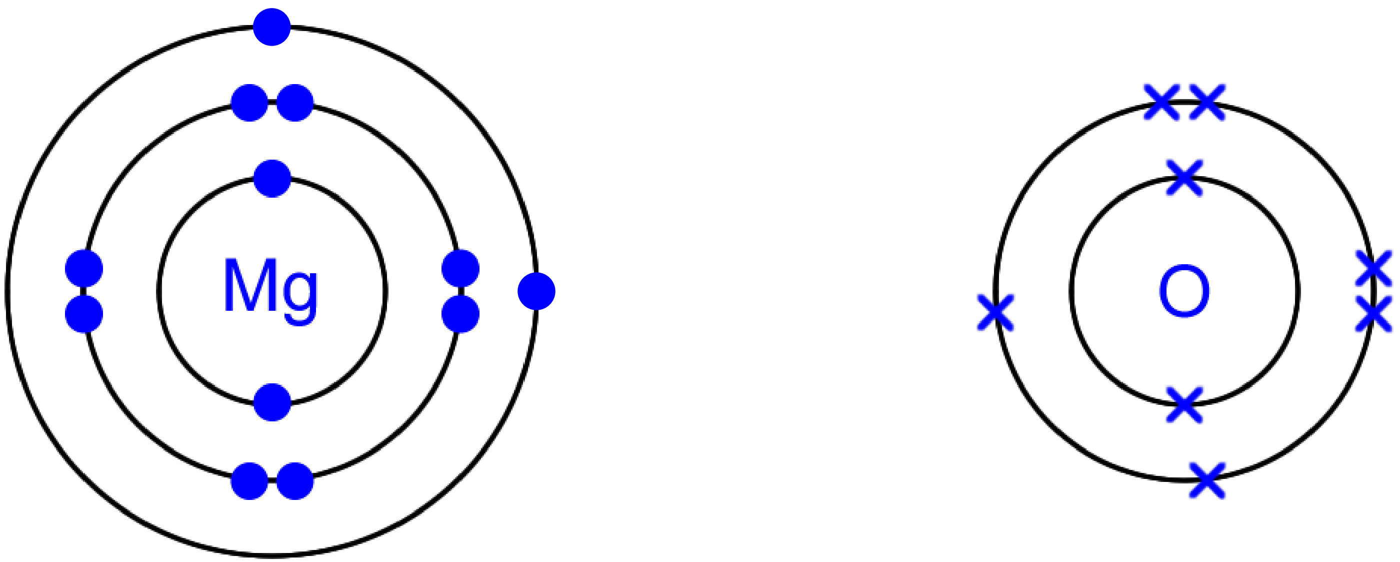


- The dot-and-cross diagram for magnesium chloride that should be used as an answer in an examination.

• = Electron of chlorine. × = Electron of magnesium.

Chemical Bonding

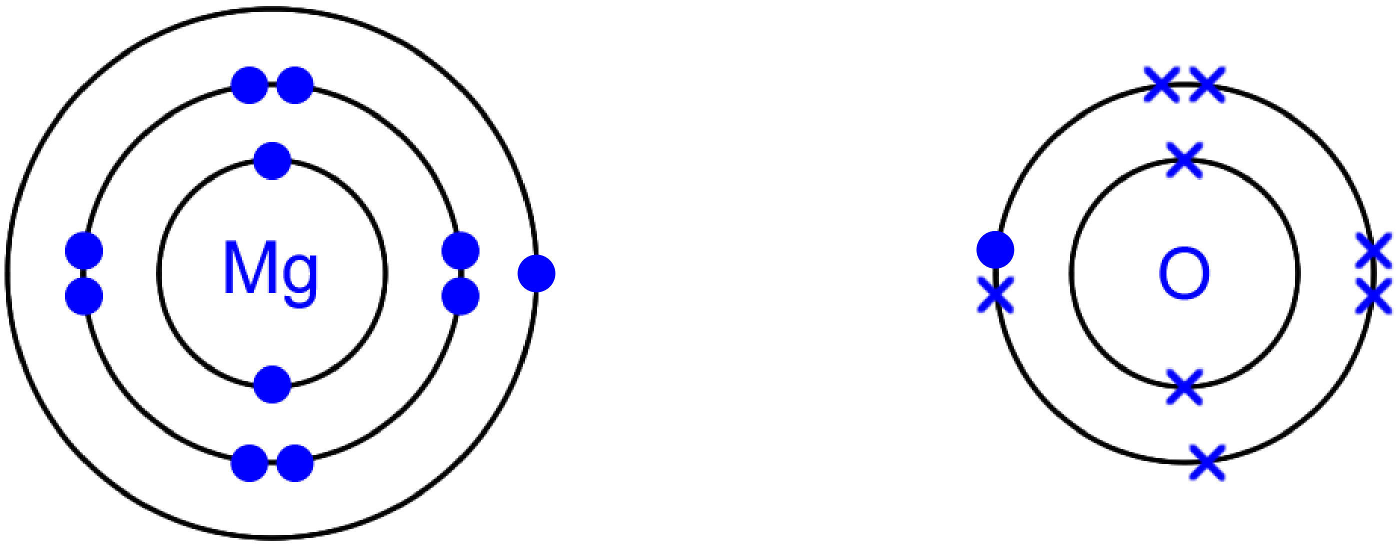
Ionic Bonding – Magnesium Oxide



• = Electron of magnesium. × = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Magnesium Oxide



● = Electron of magnesium. × = Electron of oxygen.

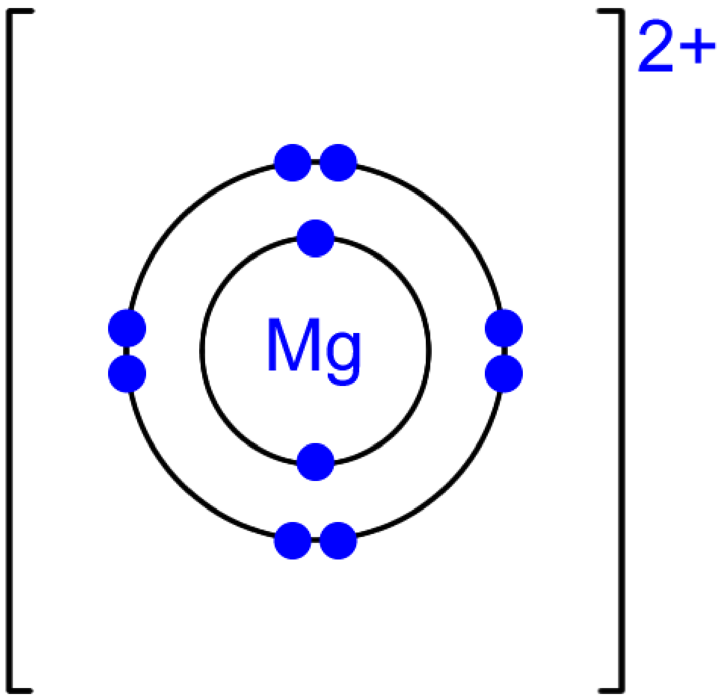
Chemical Bonding

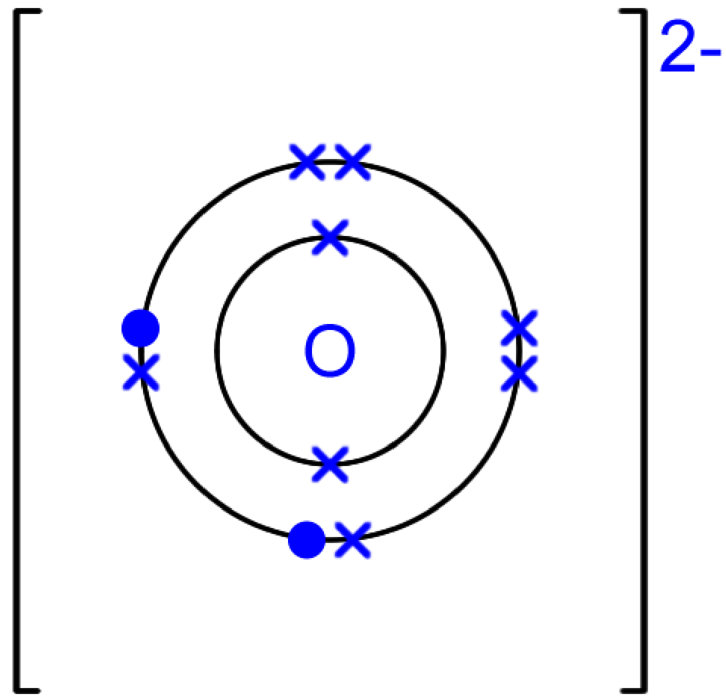
Ionic Bonding – Magnesium Oxide

● = Electron of magnesium. × = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Magnesium Oxide – MgO





Positive Magnesium Ion (Cation)

- Number of Protons = 12
- Number of Electrons = 10
- Overall Charge = (+12) + (-10) = +2

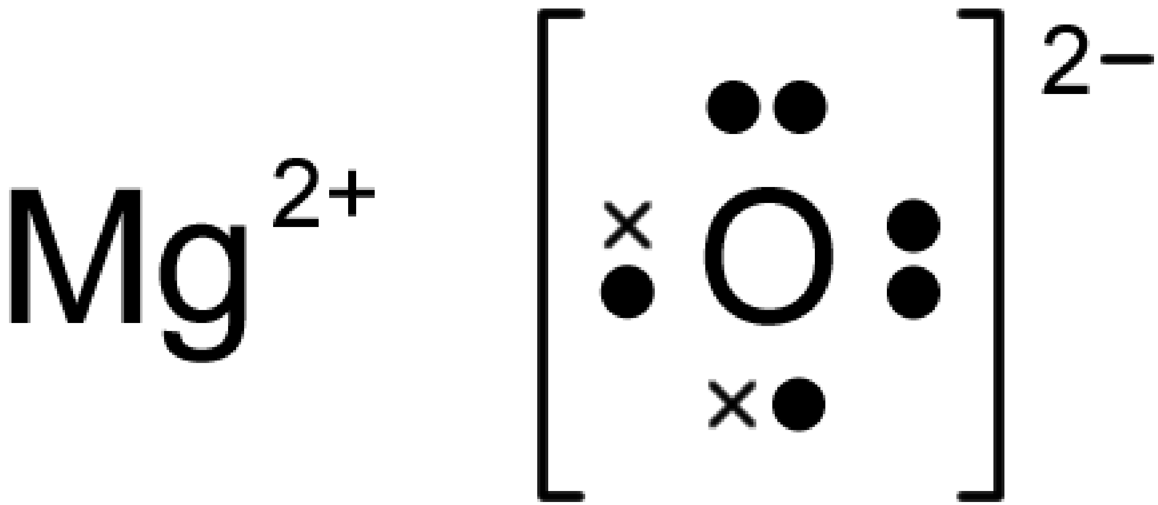
Negative Oxide Ion (Anion)

- Number of Protons = 8
- Number of Electrons = 10
- Overall Charge = (+8) + (-10) = -2

• = Electron of magnesium. × = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Magnesium Oxide – MgO

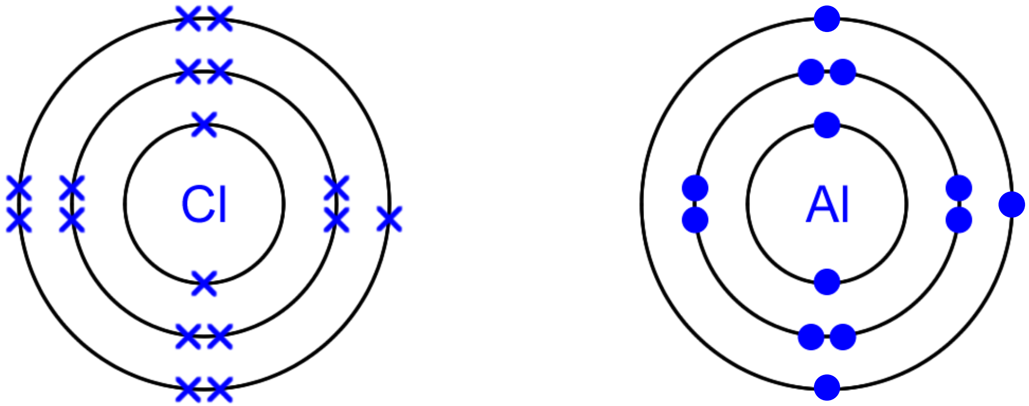


- The dot-and-cross diagram for magnesium oxide that should be used as an answer in an examination.

• = Electron of oxygen. × = Electron of magnesium.

Chemical Bonding

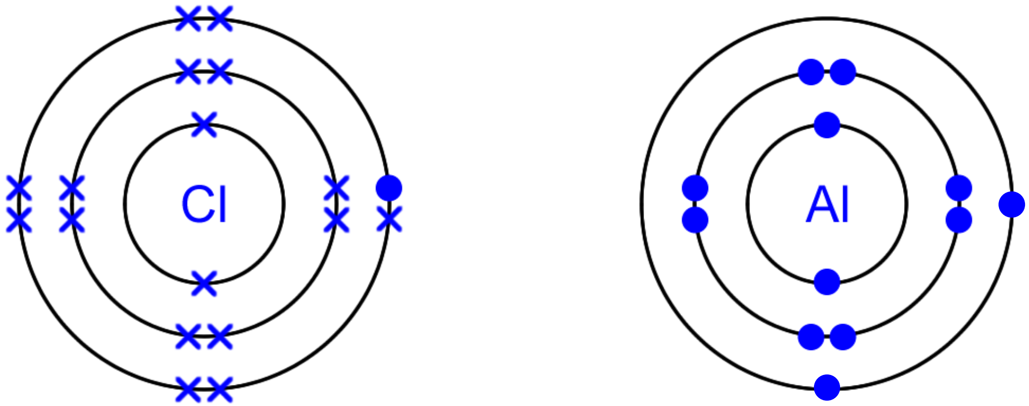
Ionic Bonding – Aluminium Chloride



● = Electron of aluminium. × = Electron of chlorine.

Chemical Bonding

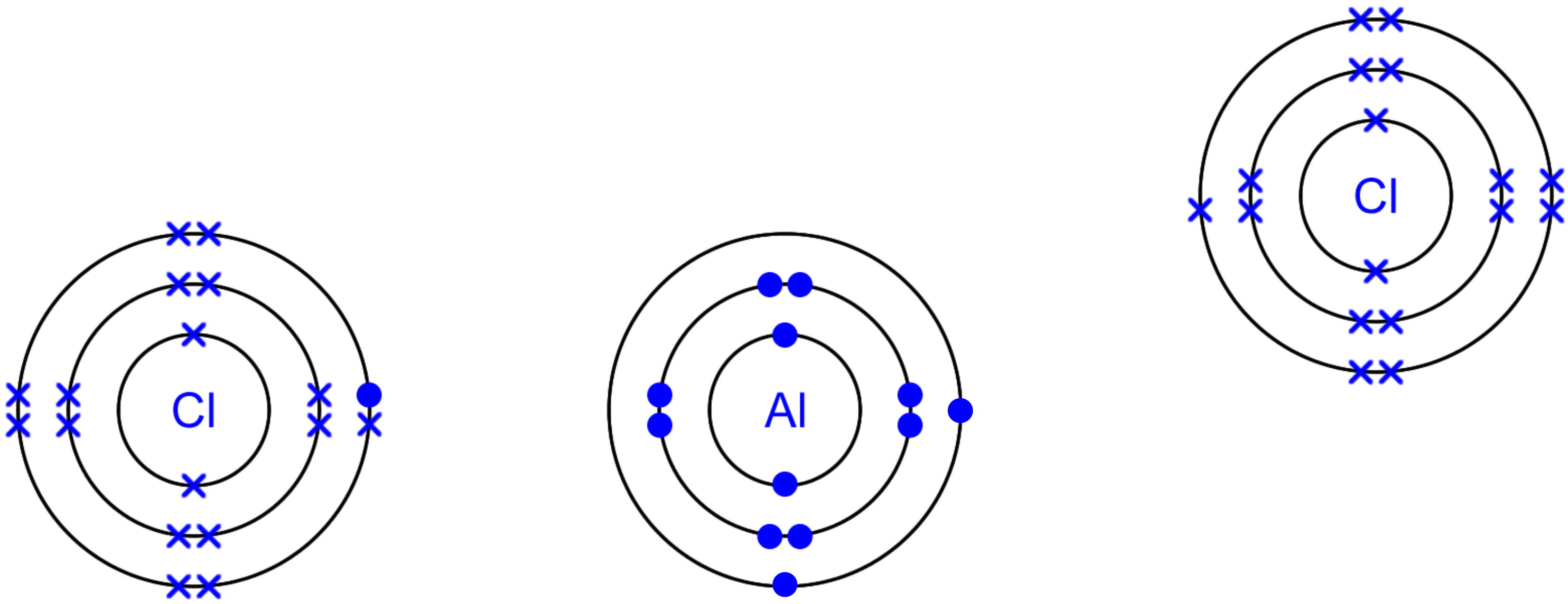
Ionic Bonding – Aluminium Chloride



• = Electron of aluminium. × = Electron of chlorine.

Chemical Bonding

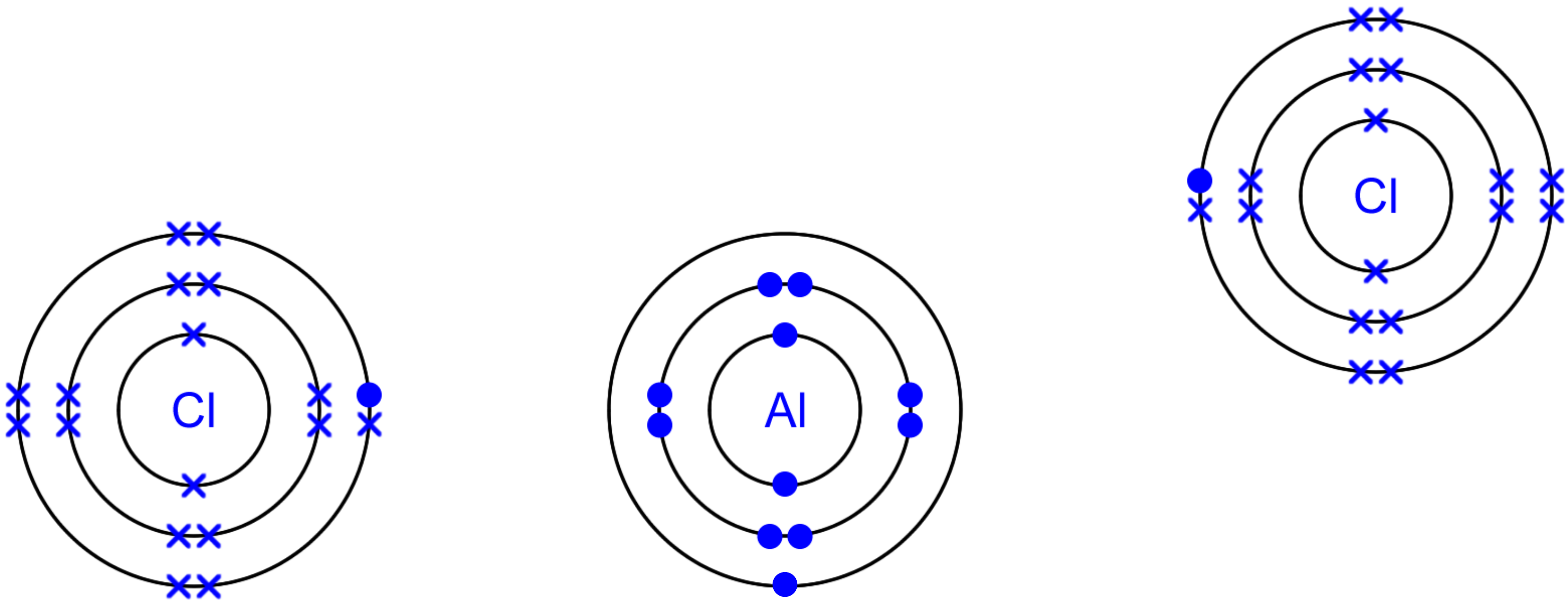
Ionic Bonding – Aluminium Chloride



• = Electron of aluminium. × = Electron of chlorine.

Chemical Bonding

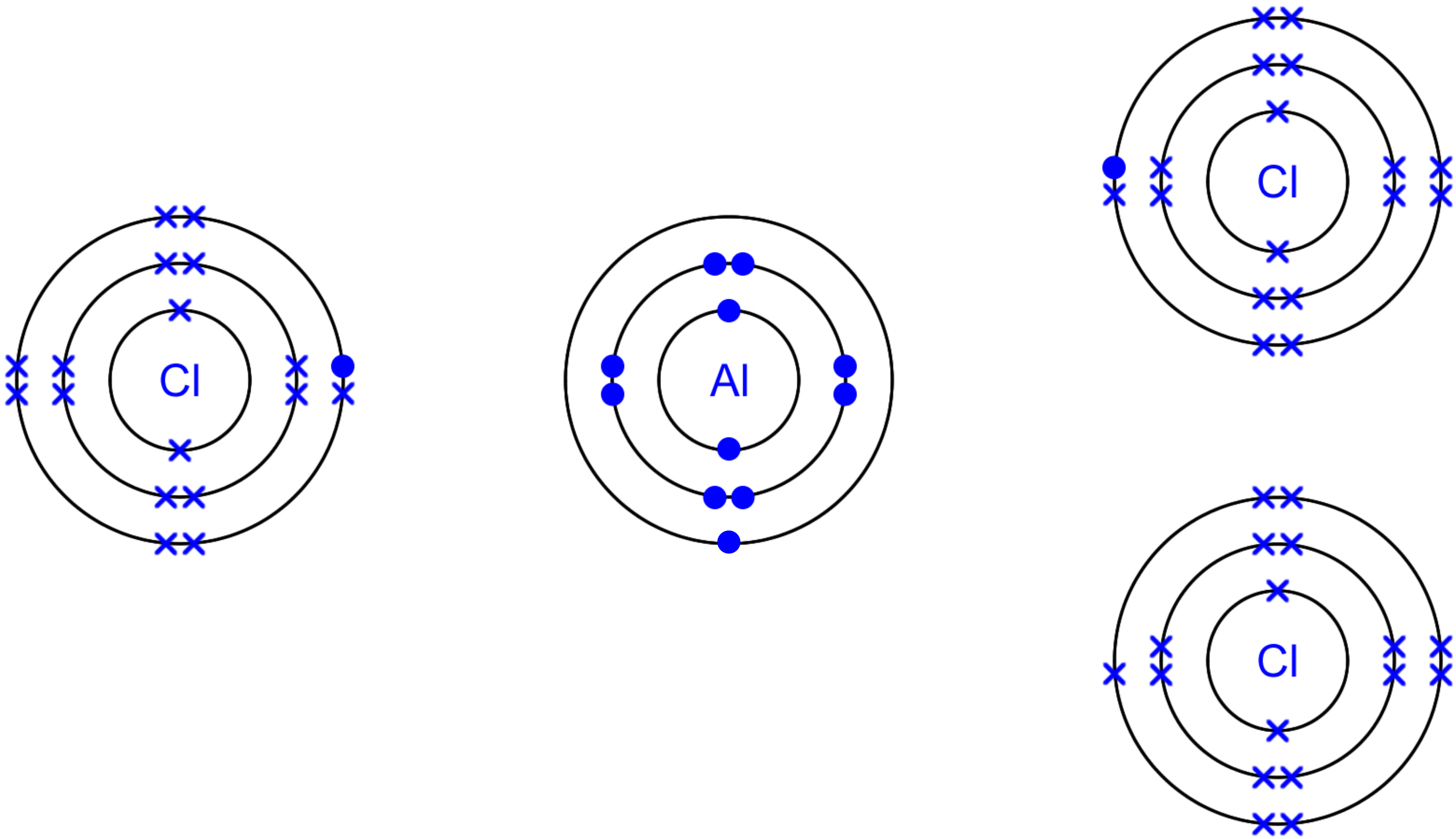
Ionic Bonding – Aluminium Chloride



• = Electron of aluminium. × = Electron of chlorine.

Chemical Bonding

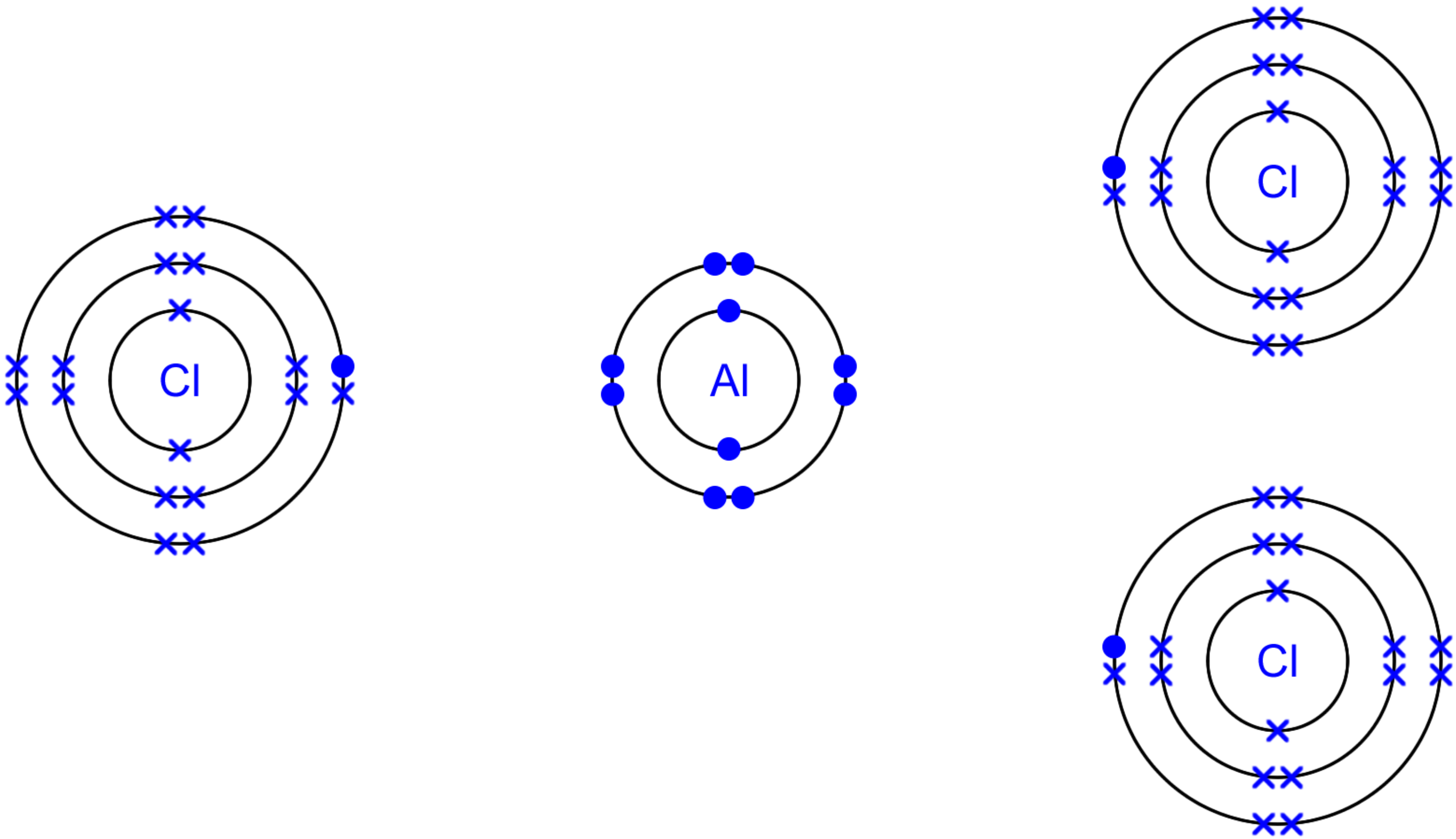
Ionic Bonding – Aluminium Chloride



• = Electron of aluminium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Aluminium Chloride



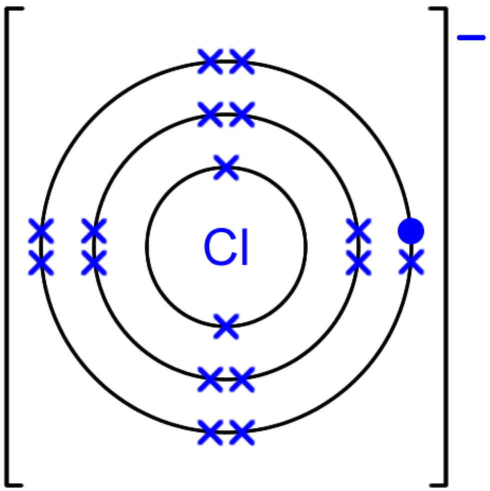
• = Electron of aluminium. × = Electron of chlorine.

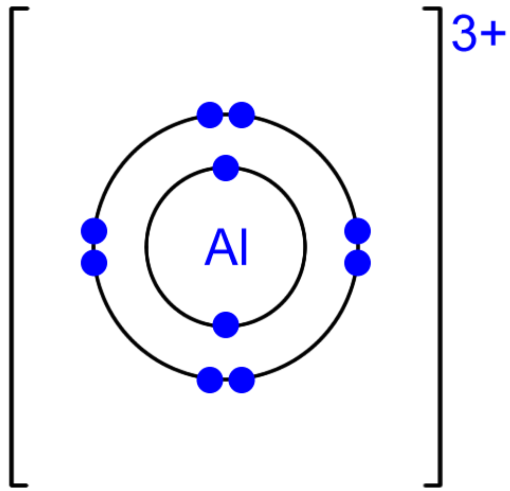
Chemical Bonding

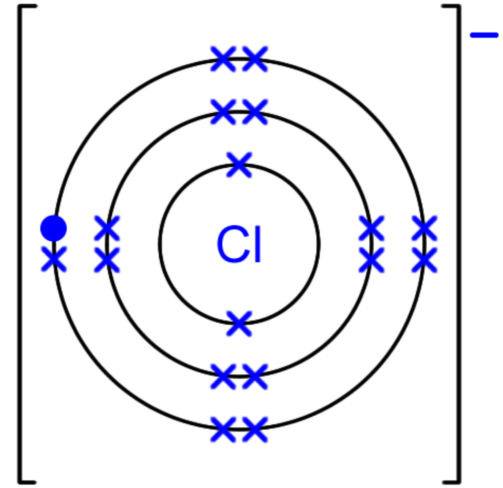
Ionic Bonding – Aluminium Chloride – $AlCl_3$

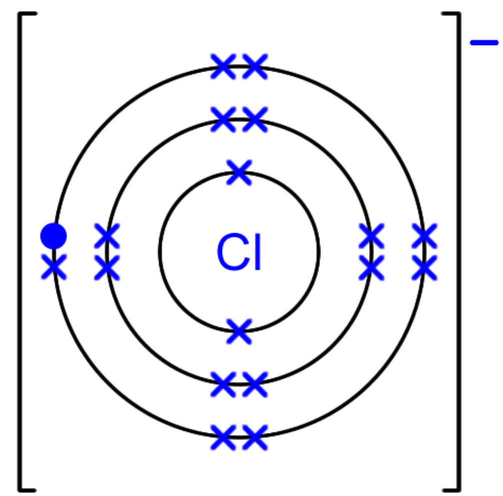
Positive Aluminium Ion (Cation)

- Number of Protons = 13
- Number of Electrons = 10
- Overall Charge = (+13) + (-10) = +3









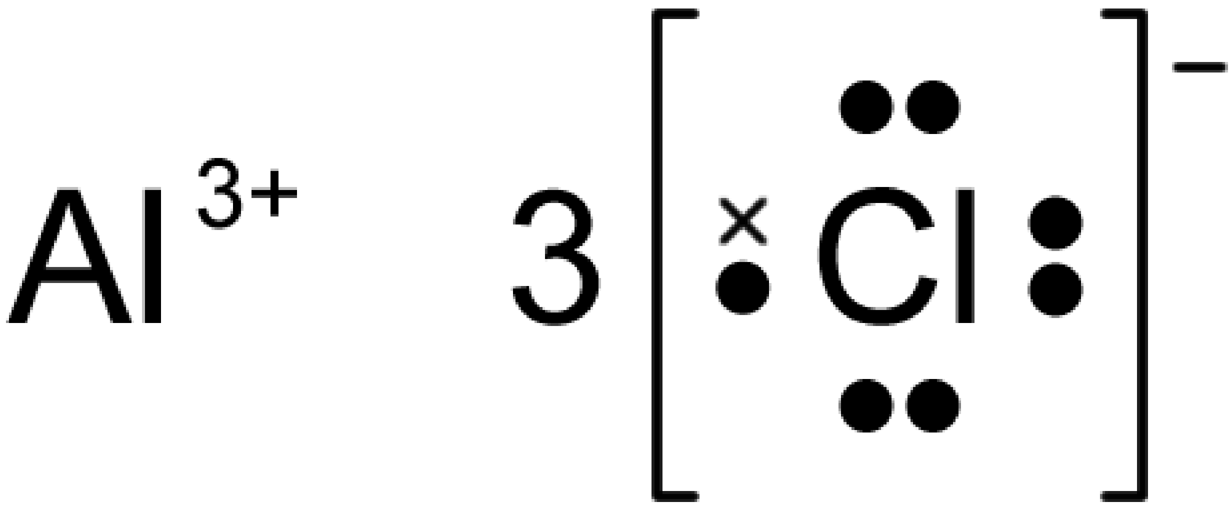
Negative Chloride Ion (Anion)

- Number of Protons = 17
- Number of Electrons = 18
- Overall Charge = (+17) + (-18) = -1

• = Electron of aluminium. × = Electron of chlorine.

Chemical Bonding

Ionic Bonding – Aluminium Chloride – $AlCl_3$

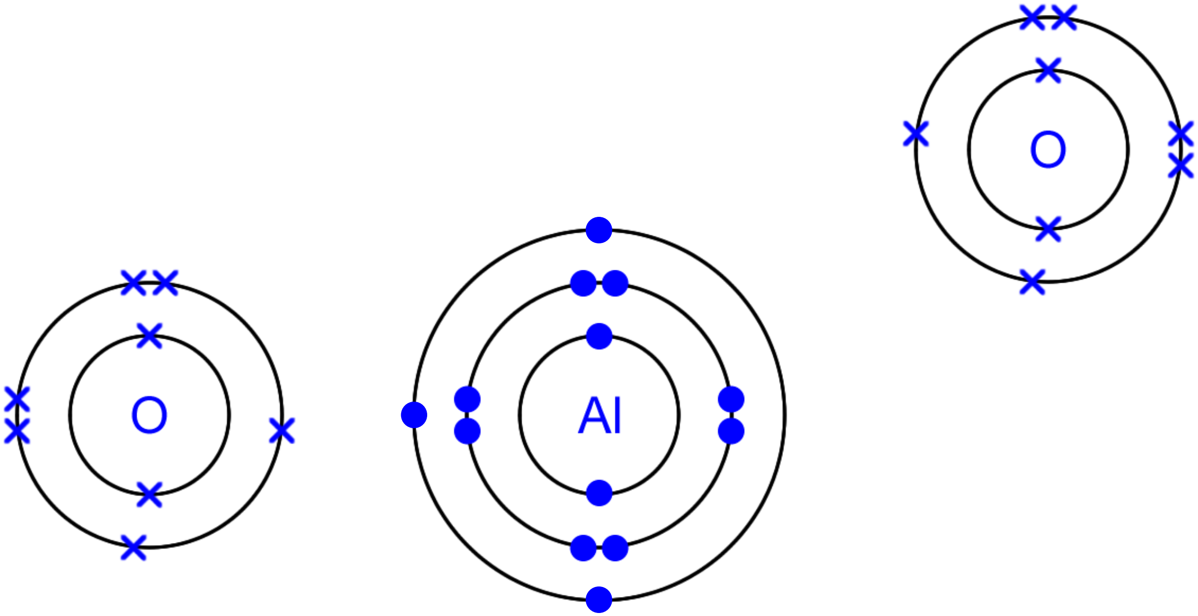


- The dot-and-cross diagram for aluminium chloride that should be used as an answer in an examination.

• = Electron of chlorine. × = Electron of aluminium.

Chemical Bonding

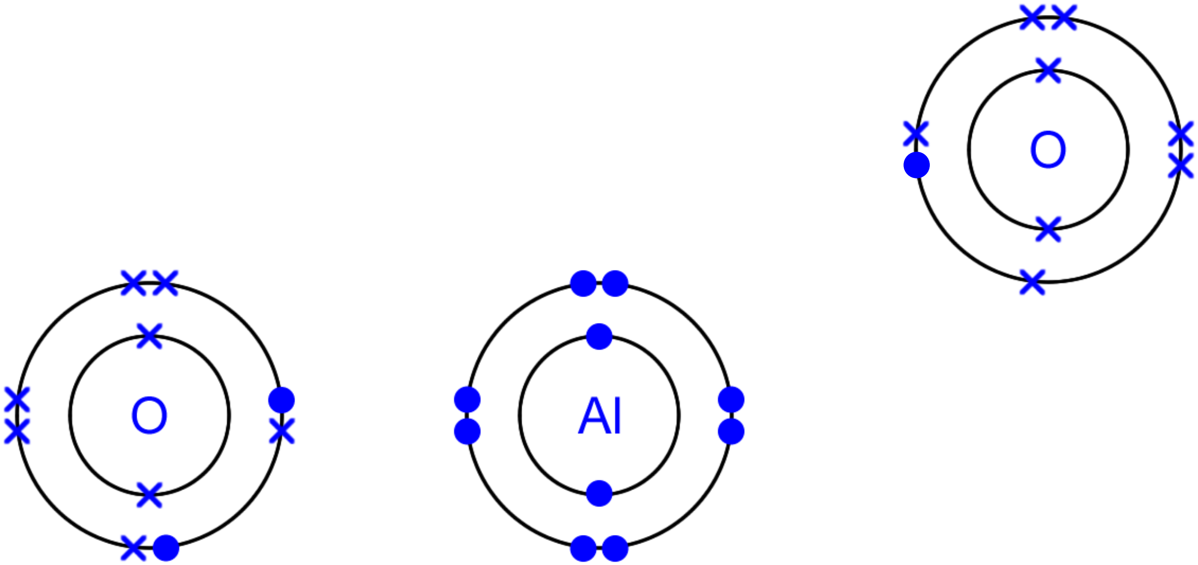
Ionic Bonding – Aluminium Oxide



• = Electron of aluminium. × = Electron of oxygen.

Chemical Bonding

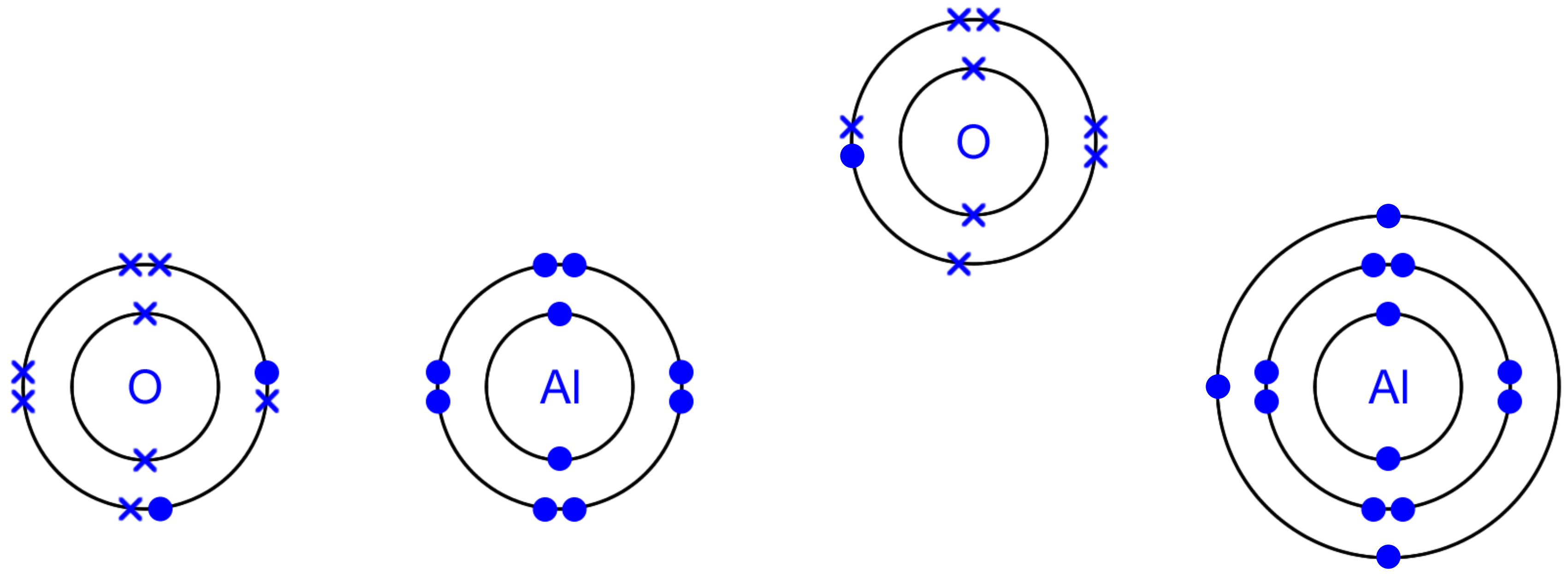
Ionic Bonding – Aluminium Oxide



• = Electron of aluminium.   × = Electron of oxygen.

Chemical Bonding

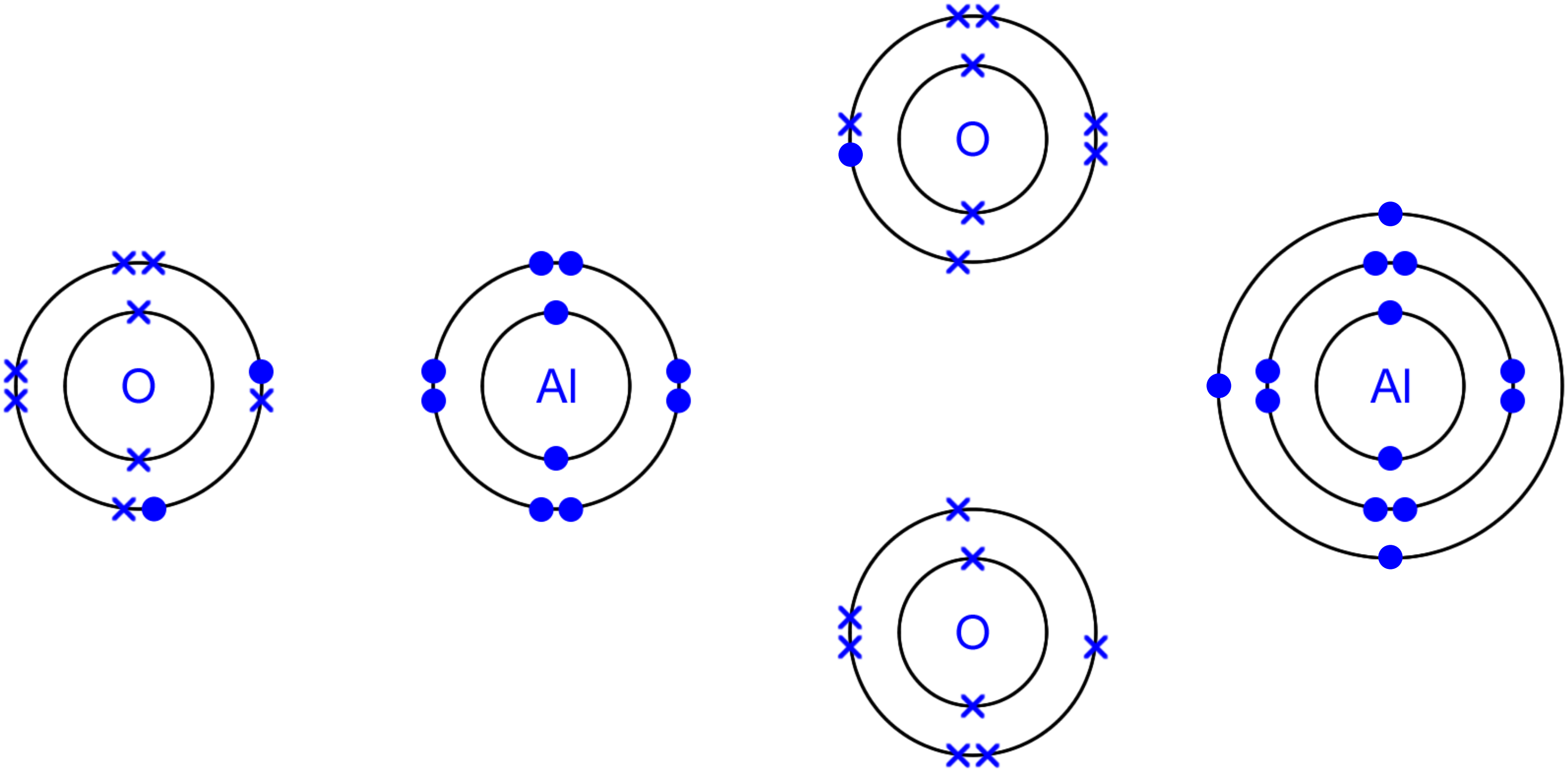
Ionic Bonding – Aluminium Oxide



● = Electron of aluminium. × = Electron of oxygen.

Chemical Bonding

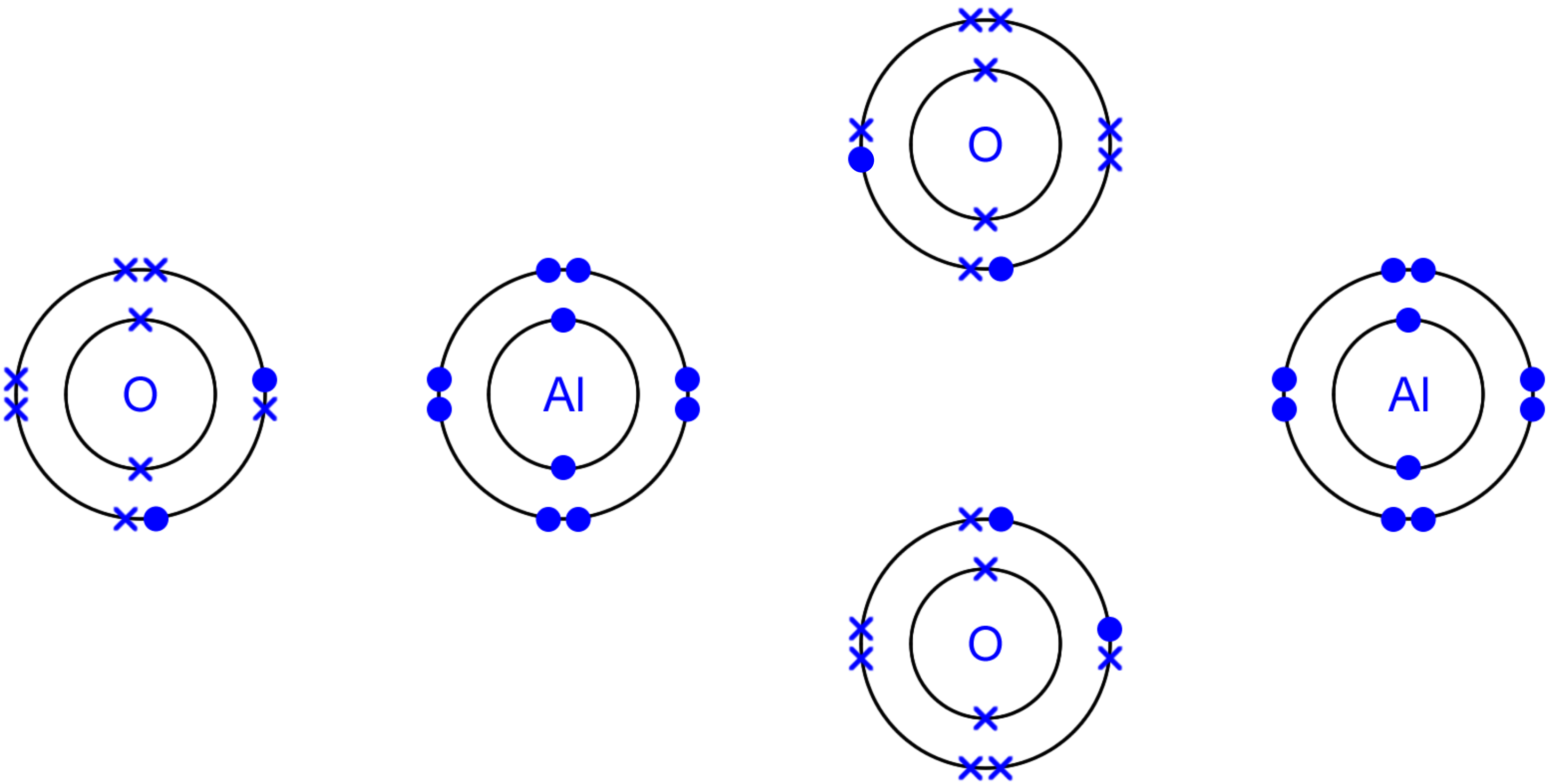
Ionic Bonding – Aluminium Oxide



• = Electron of aluminium. × = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Aluminium Oxide

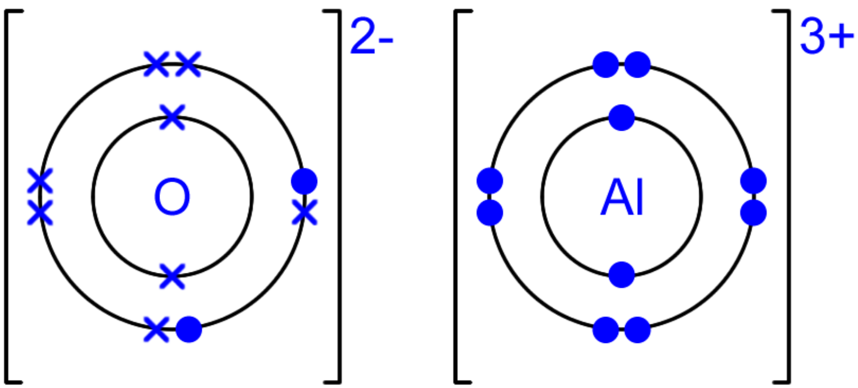


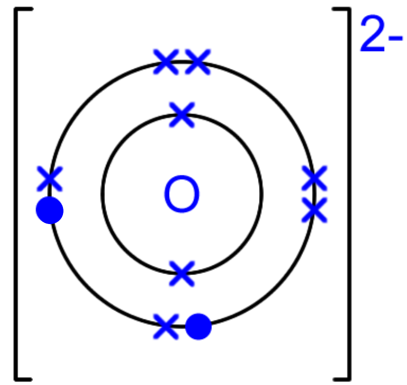
● = Electron of aluminium. × = Electron of oxygen.

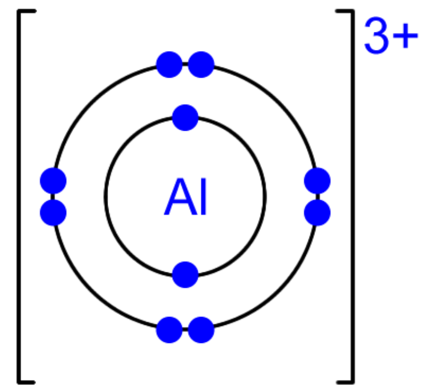
Chemical Bonding

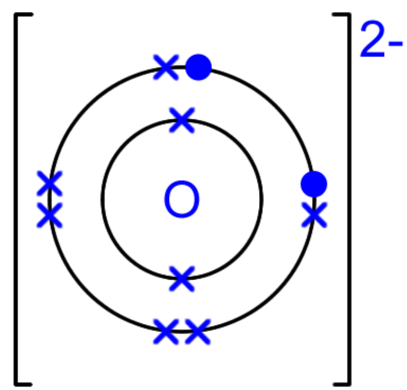
Ionic Bonding – Aluminium Oxide – Al_2O_3

Positive Aluminium Ion (Cation)
- Number of Protons = 13
- Number of Electrons = 10
- Overall Charge = (+13) + (-10) = +3







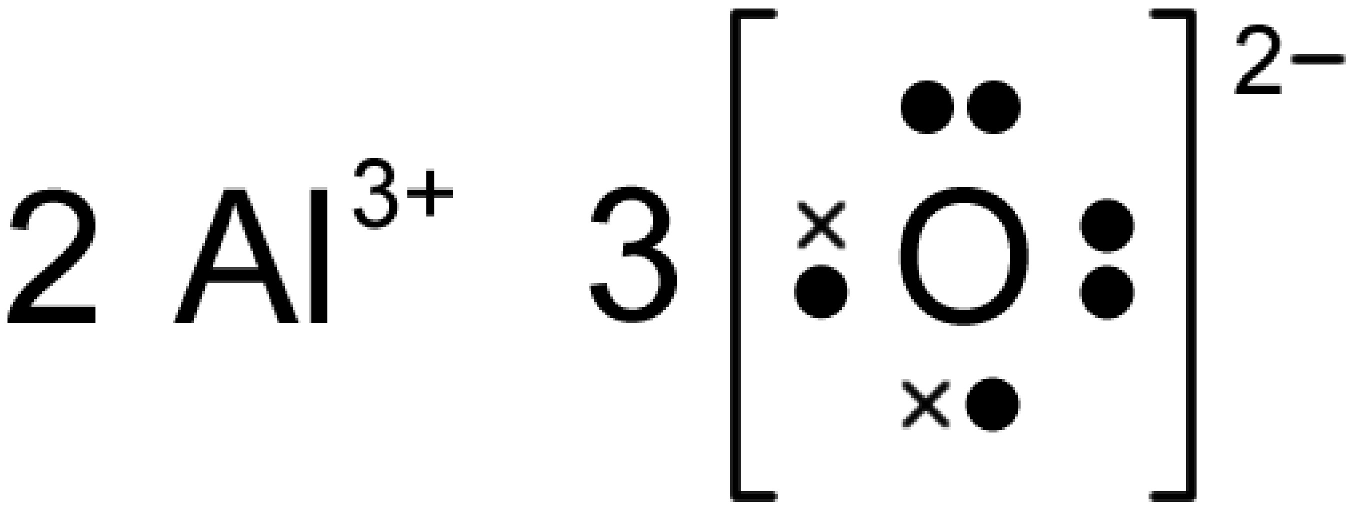


Negative Oxide Ion (Anion)
- Number of Protons = 8
- Number of Electrons = 10
- Overall Charge = (+8) + (-10) = -2

• = Electron of aluminium. × = Electron of oxygen.

Chemical Bonding

Ionic Bonding – Aluminium Oxide – Al_2O_3



- The dot-and-cross diagram for aluminium oxide that should be used as an answer in an examination.

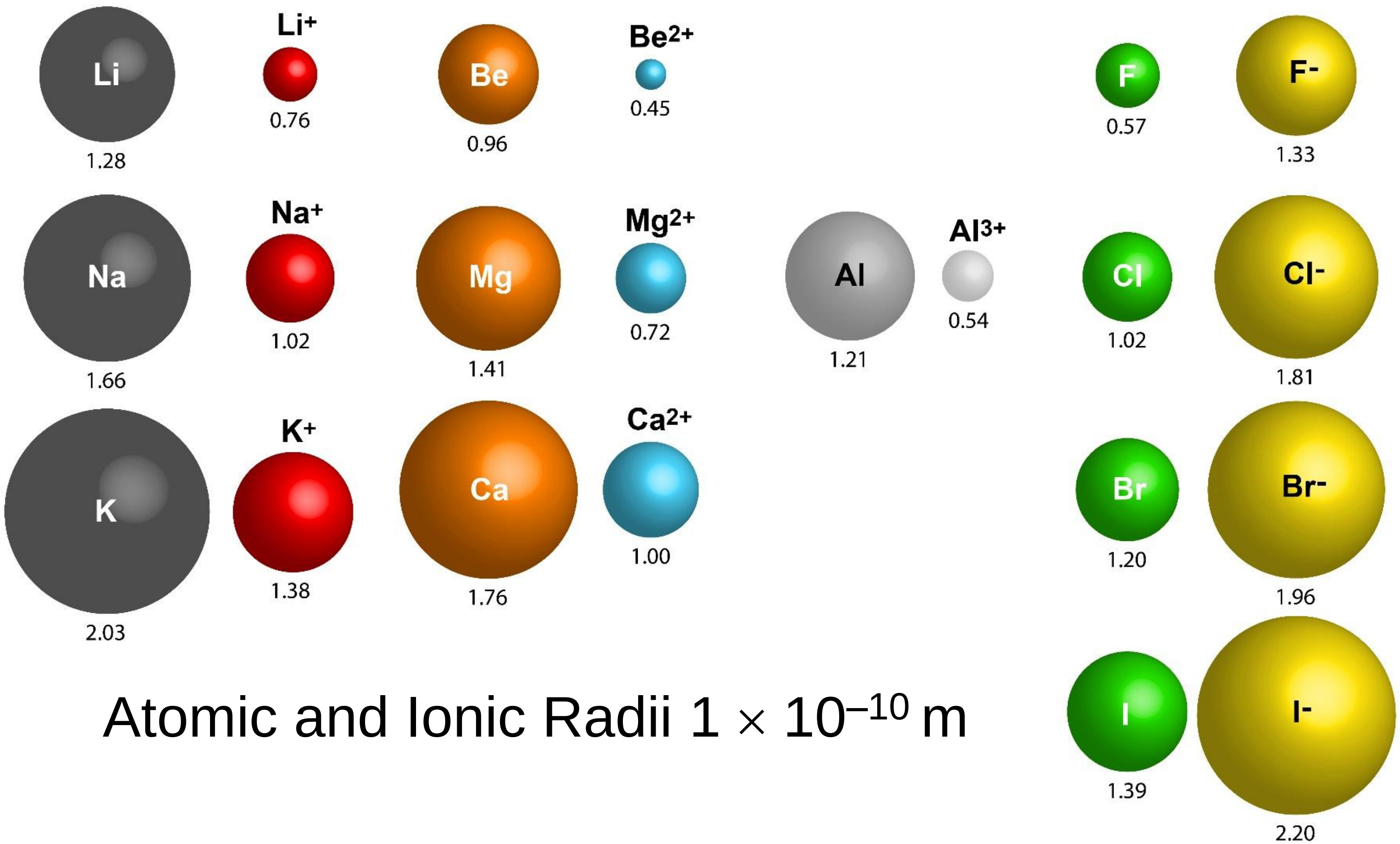
• = Electron of oxygen. × = Electron of aluminium.

Chemical Bonding

3. How does the *radius* of an *atom* differ from the *radius* of an *anion* and the *radius* of a *cation* ?

Chemical Bonding

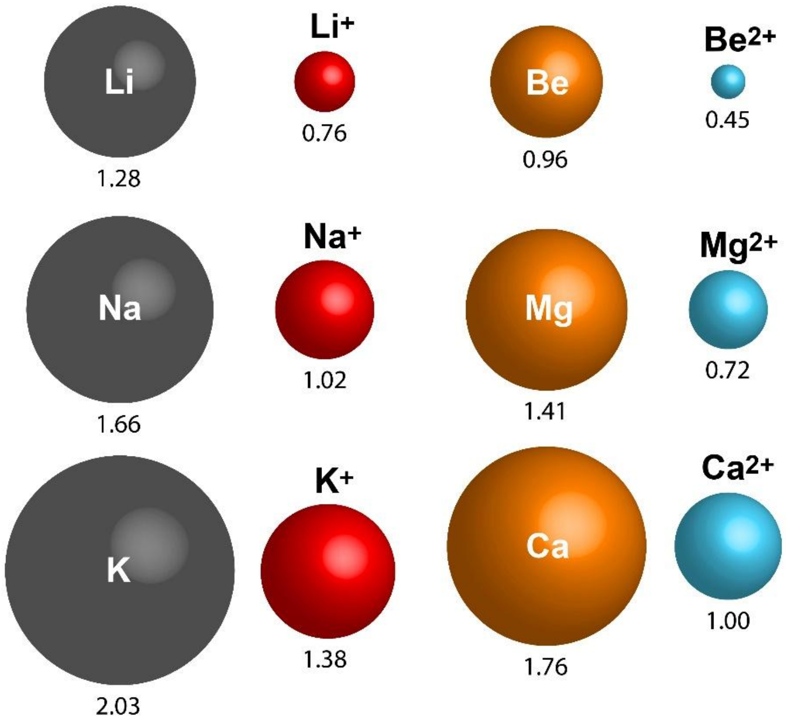
Ionic Bonding – Atomic and Ionic Radii

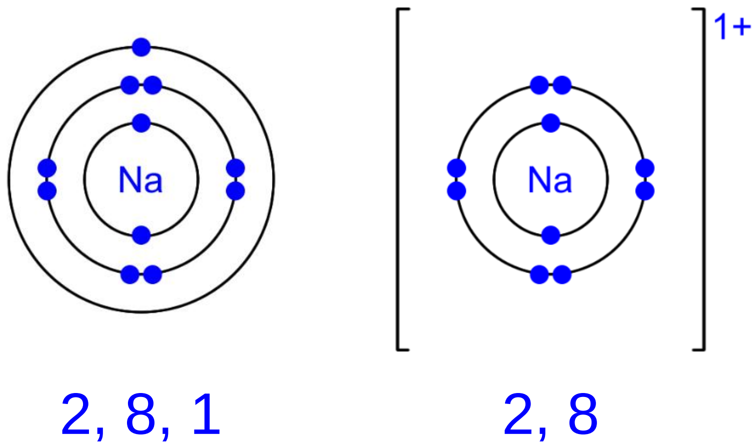


Atomic and Ionic Radii 1×10^{-10} m

Chemical Bonding

Ionic Bonding – Atomic and Ionic Radii

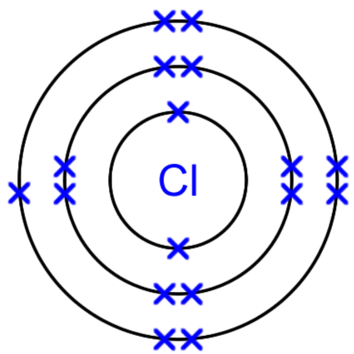


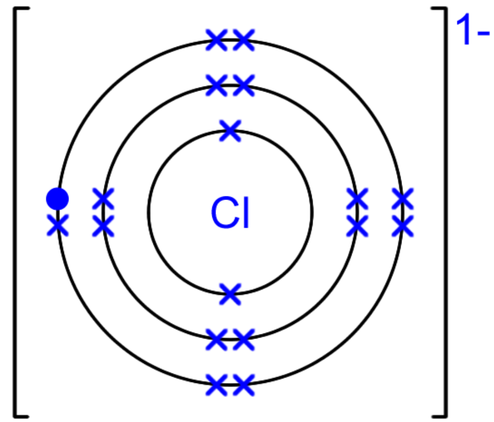


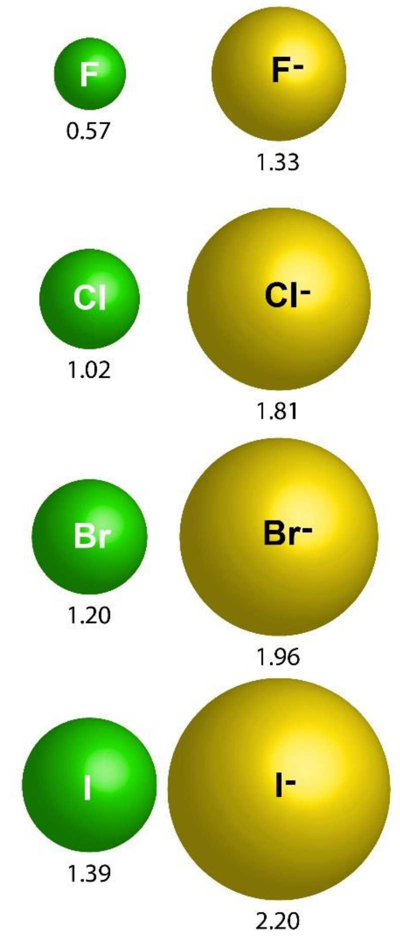
- For *metals*, the positively charged cation has a *smaller* radius than the original atom. This is because the cation has *one electron shell less* than the original atom, *e.g.* a sodium atom has three electron shells while a sodium ion has two electron shells.

Chemical Bonding

Ionic Bonding – Atomic and Ionic Radii



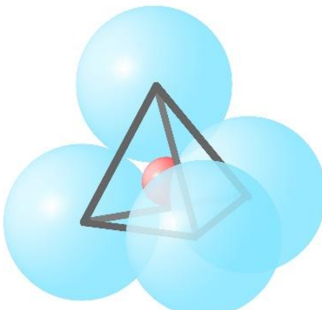




- For *non-metals*, the negatively charged anion has a *larger* radius than the original atom. This is because the anion has *gained* electrons in order to obtain the electronic configuration of a noble gas. The electrostatic force of *repulsion* between the negatively charged electrons causes the anion to be larger than the original atom.

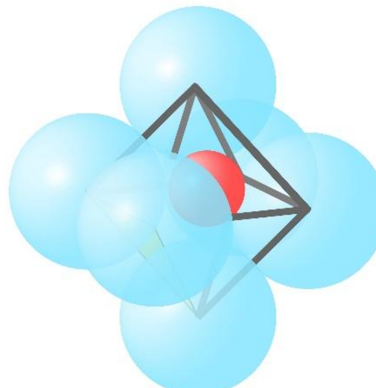
Chemical Bonding

Ionic Bonding – Atomic and Ionic Radii



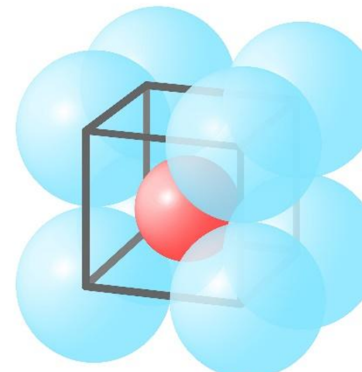
$$\frac{r_{cation}}{r_{anion}} = 0.225$$

coordination number 4



$$\frac{r_{cation}}{r_{anion}} = 0.414$$

coordination number 6



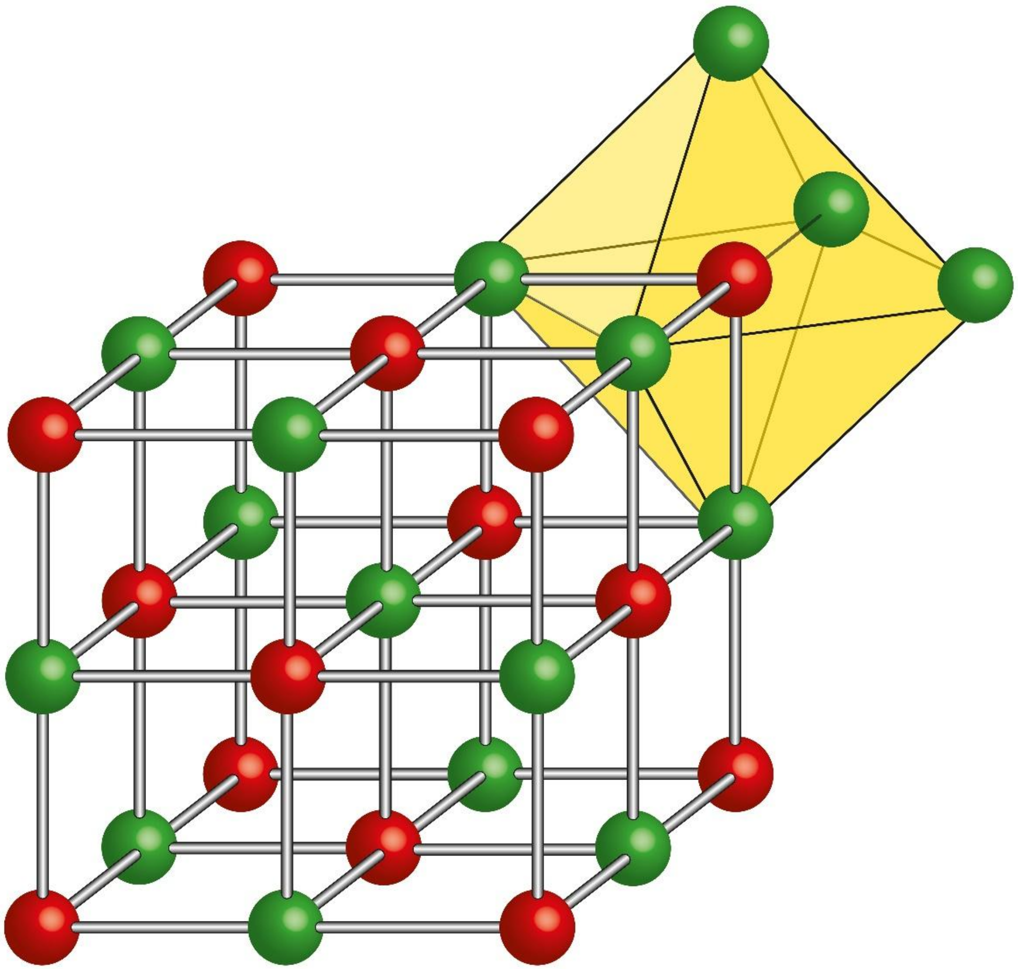
$$\frac{r_{cation}}{r_{anion}} = 0.732$$

coordination number 8

- Anions and cations can pack together in different arrangements, with different *coordination numbers*, depending upon the ionic radii of the anions and cations. The coordination number is a measure of how many anions surround a single neighbouring cation, and vice-versa.

Chemical Bonding

Ionic Bonding – Atomic and Ionic Radii



- Sodium chloride has a coordination number of *6*. *6* chloride ions surround a single sodium ion and *6* sodium ions surround a single chloride ion.

Chemical Bonding

4. Can I derive the *formula* of a compound based upon where the chemical elements are positioned in the Periodic Table?

Main Menu

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons							
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration							
Valency of Element							

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration							
Valency of Element							

Chemical Bonding

Group	**1**	**2**	**13**	**14**	**15**	**16**	**17**
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration	1	2	3	4	3	2	1
Valency of Element							

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration	1	2	3	4	3	2	1
Valency of Element	1	2	3	4	3	2	1

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration	1	2	3	4	3	2	1
Valency of Element	1	2	3	4	3	2	1

- The formula of a compound can be easily derived by *swapping* the *valencies* of the two elements.

e.g. potassium (Group 1) and oxygen (Group 16)

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration	1	2	3	4	3	2	1
Valency of Element	1	2	3	4	3	2	1

- The formula of a compound can be easily derived by *swapping* the *valencies* of the two elements.

e.g. potassium (Group 1) and oxygen (Group 16)

$$\overset{1}{K} \quad \overset{2}{O}$$

…becomes…

$$K_2O$$

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration	1	2	3	4	3	2	1
Valency of Element	1	2	3	4	3	2	1

- The formula of a compound can be easily derived by *swapping* the *valencies* of the two elements.

e.g. calcium (Group 2) and nitrogen (Group 15)

Chemical Bonding

Group	1	2	13	14	15	16	17
Number of Valence Electrons	1	2	3	4	5	6	7
Number of Electrons Lost or Gained to Obtain Noble Gas Electronic Configuration	1	2	3	4	3	2	1
Valency of Element	1	2	3	4	3	2	1

- The formula of a compound can be easily derived by *swapping* the *valencies* of the two elements.

e.g. calcium (Group 2) and nitrogen (Group 15)

$$\overset{2}{Ca}\ \overset{3}{N}$$

…becomes…

$$Ca_3N_2$$

Chemical Bonding

Ionic Bonding

5. Self-assessment

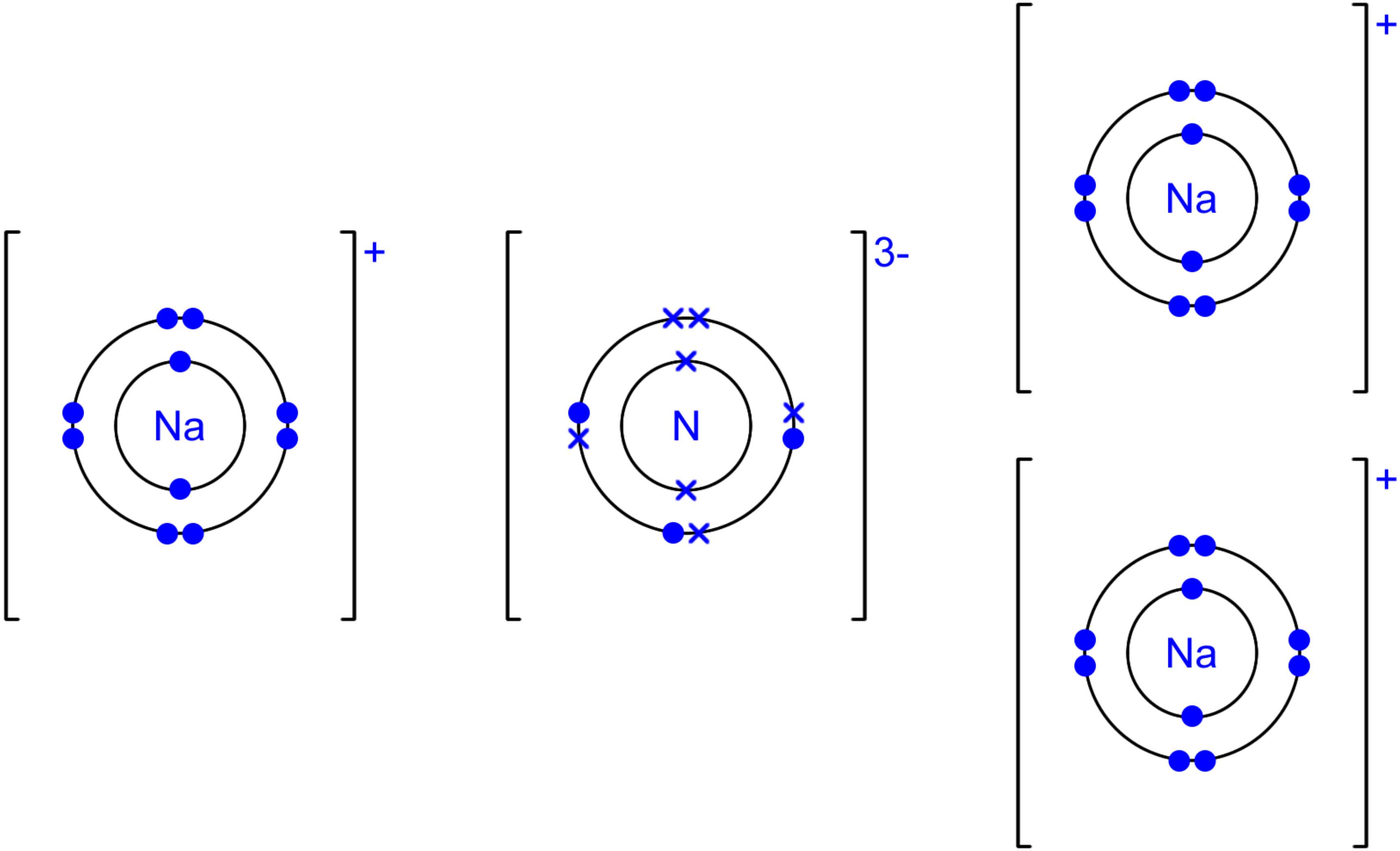
- Draw dot-and-cross diagrams to show the bonding in:

a) Sodium nitride.
b) Magnesium nitride.
c) Aluminium nitride.

Main Menu

Chemical Bonding

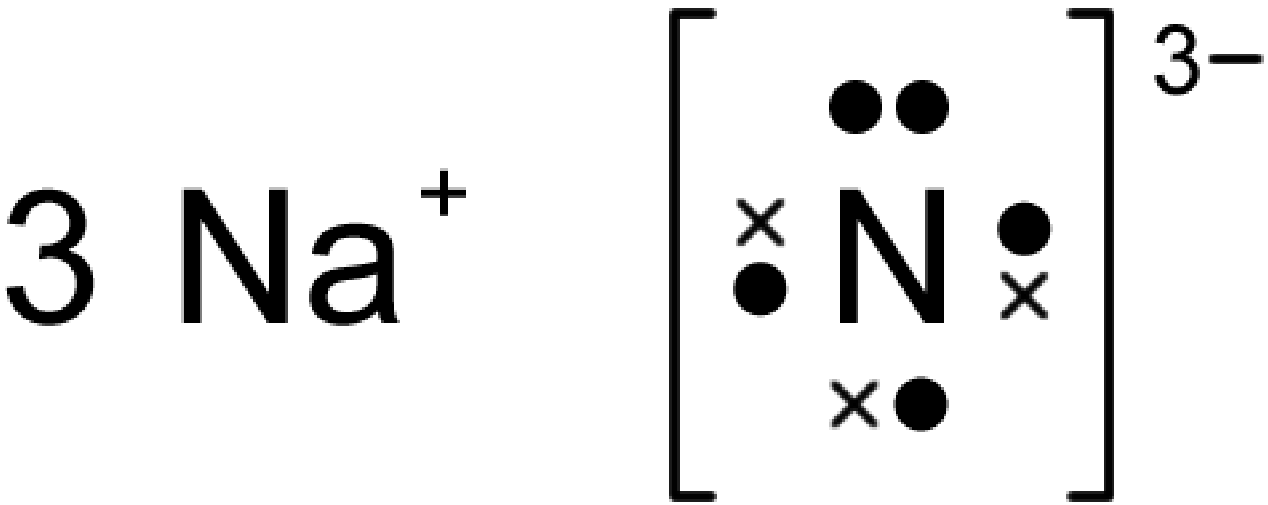
Ionic Bonding – Sodium Nitride – Na_3N



• = Electron of sodium. × = Electron of nitrogen.

Chemical Bonding

Ionic Bonding – Sodium Nitride – Na_3N

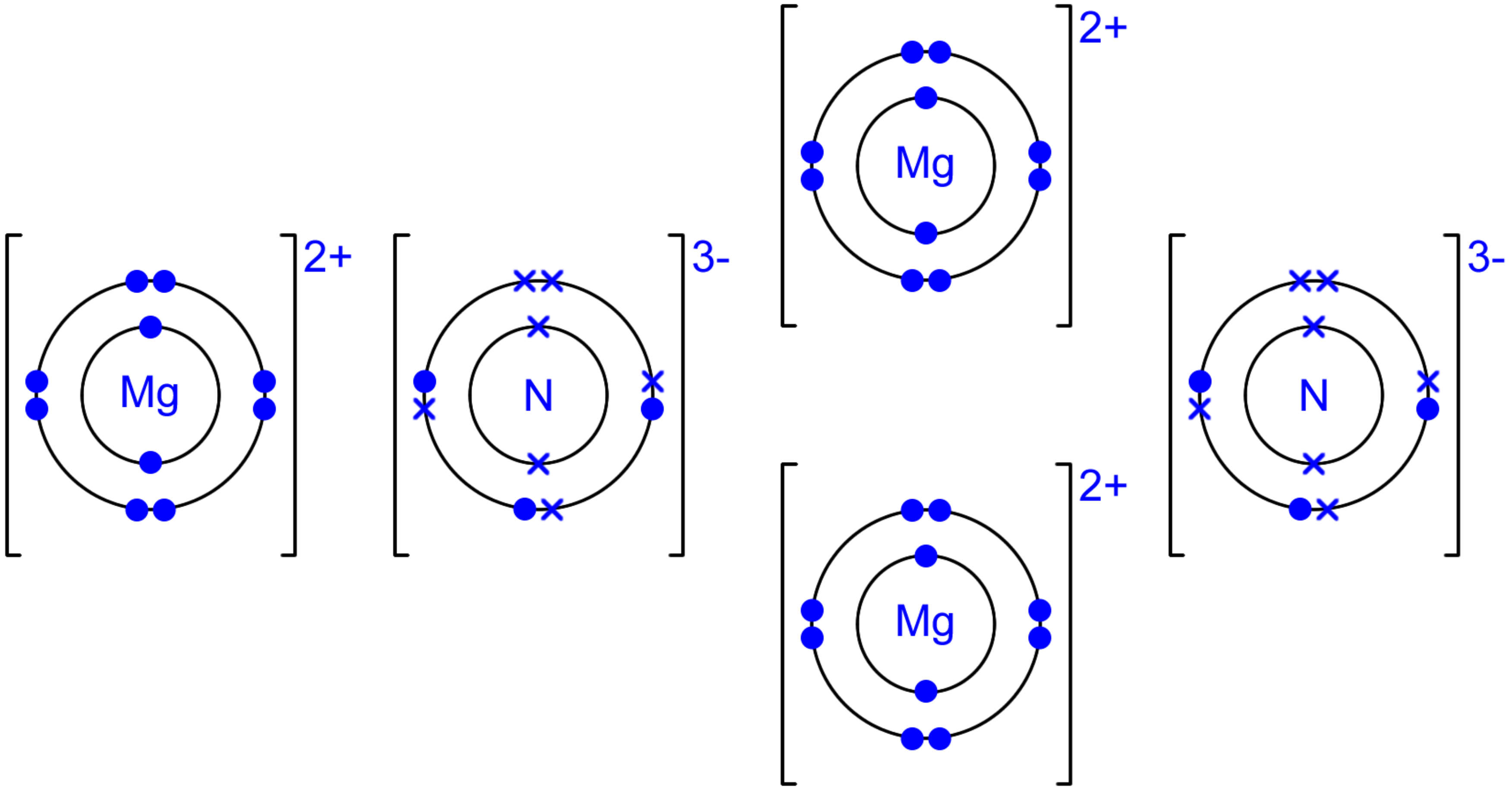


- The dot-and-cross diagram for sodium nitride that should be used as an answer in an examination.

• = Electron of nitrogen. × = Electron of sodium.

Chemical Bonding

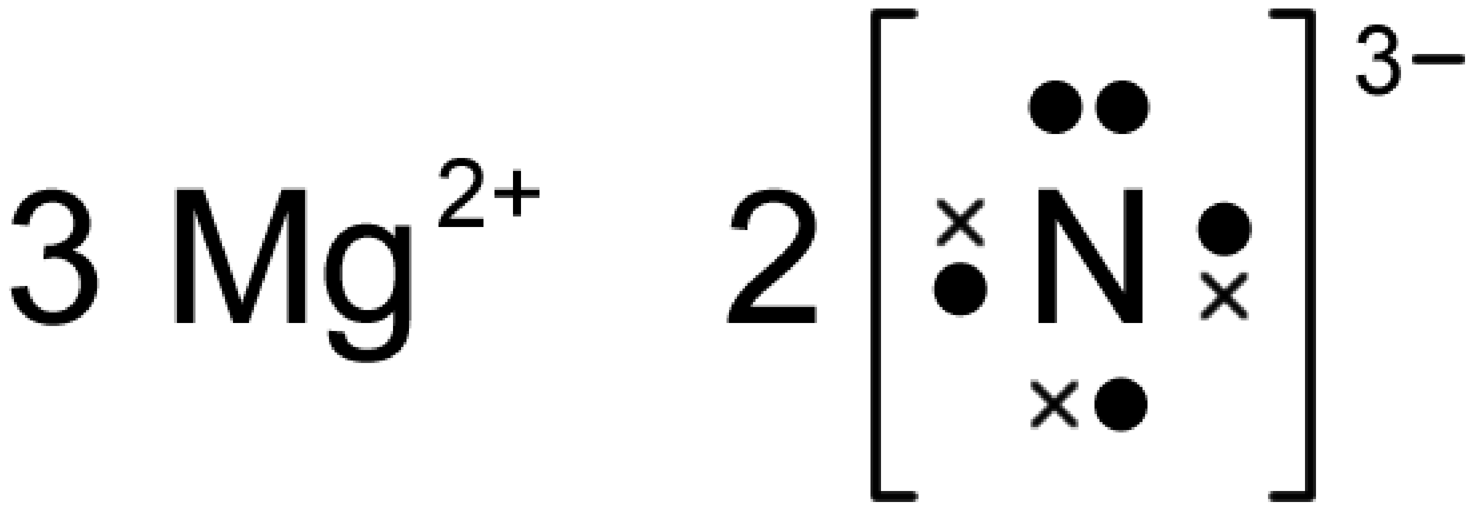
Ionic Bonding – Magnesium Nitride – Mg_3N_2



● = Electron of magnesium. × = Electron of nitrogen.

Chemical Bonding

Ionic Bonding – Magnesium Nitride – Mg_3N_2

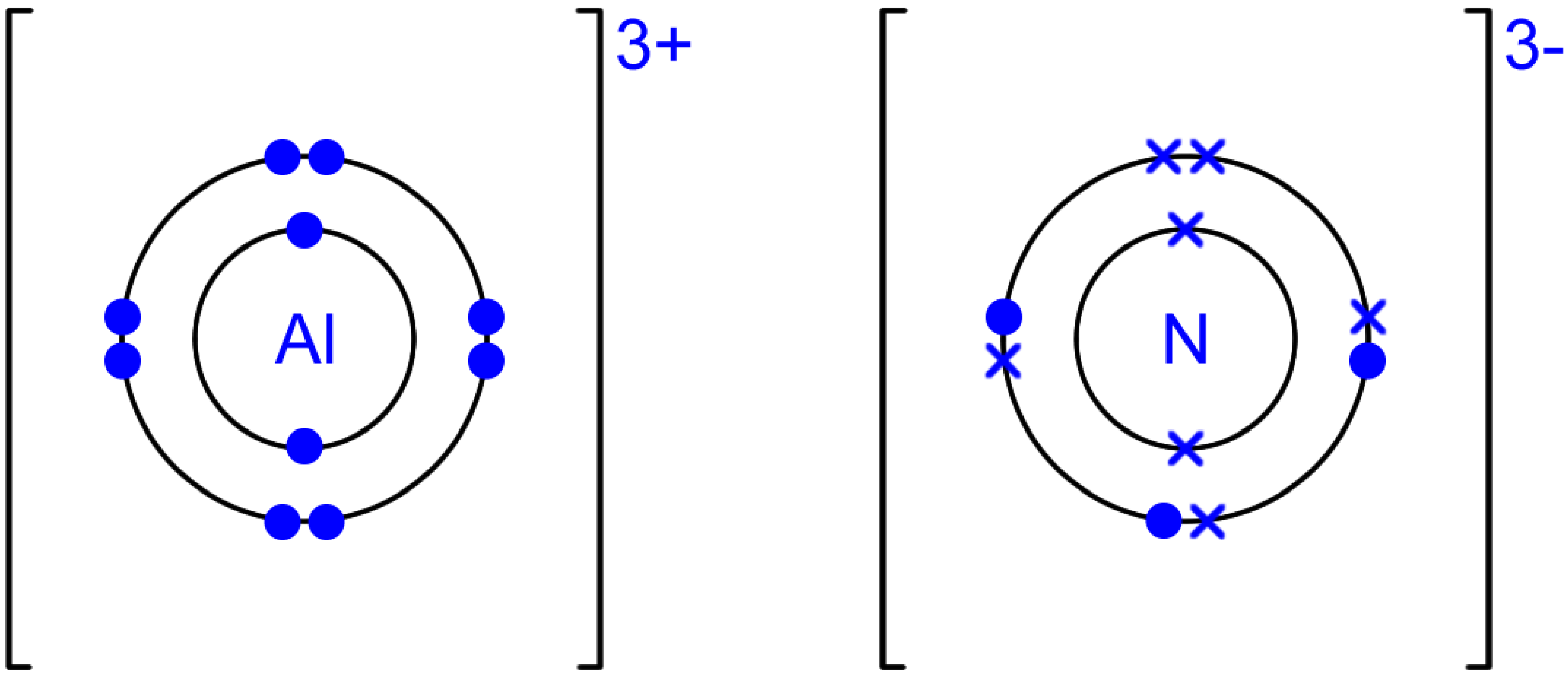


- The dot-and-cross diagram for magnesium nitride that should be used as an answer in an examination.

• = Electron of nitrogen. × = Electron of magnesium.

Chemical Bonding

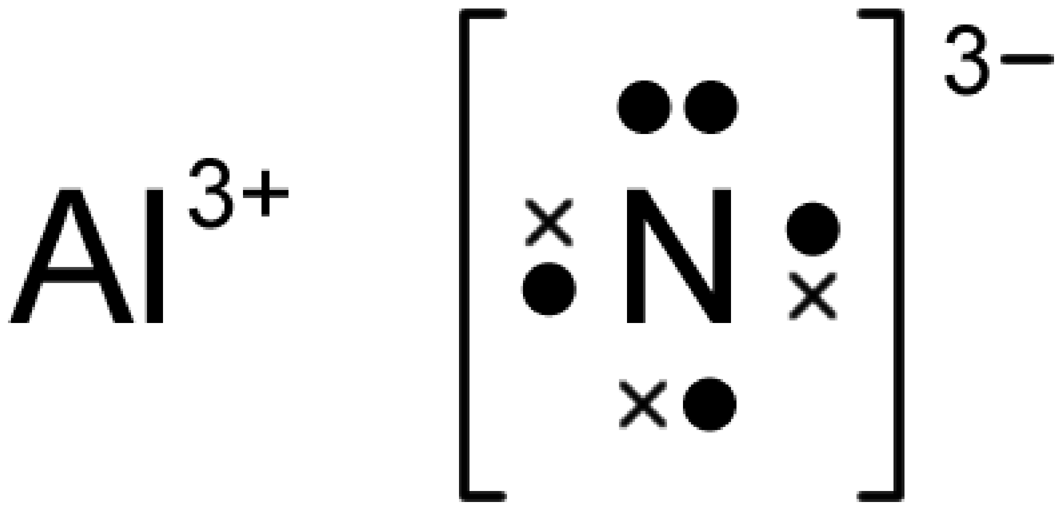
Ionic Bonding – Aluminium Nitride – AlN



● = Electron of aluminium. × = Electron of nitrogen.

Chemical Bonding

Ionic Bonding – Aluminium Nitride – A*l*N



- The dot-and-cross diagram for aluminium nitride that should be used as an answer in an examination.

• = Electron of nitrogen. × = Electron of aluminium.

Chemical Bonding

Ionic Bonding – Summary

- Ionic bonding occurs between a *metallic element* and a *non-metallic element*.
- The metal *transfers* its valence electron(s) into the valence shell of the non-metal.
- Both the metal and the non-metal obtain *noble gas electronic configurations*.
- Due to the loss of negatively charged electrons, the metal atom transforms into a *positive ion* (cation).
- Due to the gain of negatively charged electrons, the non-metal atom transforms into a *negative ion* (anion).
- Oppositely charged ions attract towards each other and arrange themselves in a 3D *crystal lattice*.

Chemical Bonding

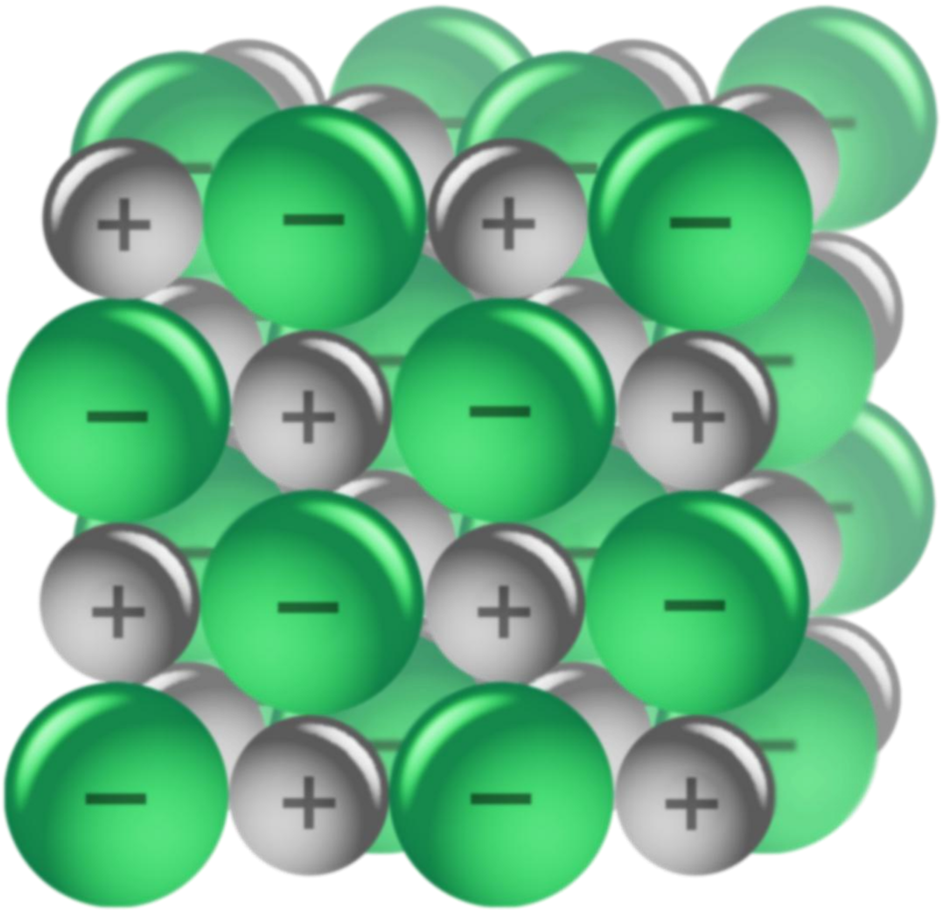
Ionic Bonding – Properties

6. What are the *general properties* of ionic compounds?

Main Menu

Chemical Bonding

Ionic Bonding – Properties



- Due to the very strong *electrostatic force of attraction* that hold the *anions* and *cations* together in a crystal lattice structure, ionic compounds are all *solids* at room temperature. All ionic compounds have *very high melting points* and *boiling points*.

Chemical Bonding

Ionic Bonding – Properties

- The melting point of an ionic compound varies with the strength of the electrostatic force of attraction between the anion (negative ion) and the cation (positive ion).

- The *stronger the electrostatic force of attraction* between the anion and the cation, the *greater the thermal energy* required to weaken the strong electrostatic force of attraction, and the *higher the melting point* of the ionic compound.

Chemical Bonding

Ionic Bonding – Properties

- Two variables affect the strength of the electrostatic force of attraction between an anion and a cation...

a) The size of the ion, referred to as the *ionic radius*.

b) The *amount of charge* carried by the ion.

- These two variables combine together give rise to a new variable called *charge density* – the amount of charge that an ion carries per unit volume.
- The greater the charge density on the anion and the cation, the greater the electrostatic force of attraction between the oppositely charge ions. More thermal energy is required to weaken the strong force of attraction between the oppositely charged ions, hence the ionic compound has a higher melting point.

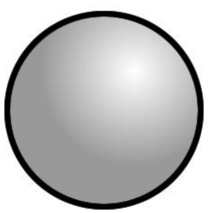
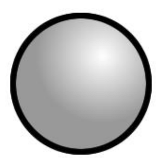
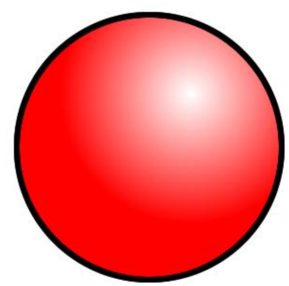
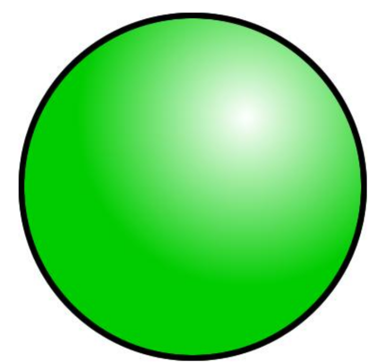
Chemical Bonding

Ionic Bonding – Properties

- Use the information presented in the table below to predict which compound – $\mathrm{NaC}l$, Na_2O or MgO will have…

a) The highest melting point?

b) The lowest melting point?

Na^{+}	Mg^{2+}	O^{2-}	$\mathrm{C}l^{-}$
			
ionic radius 1.02×10^{-10} m	ionic radius 0.72×10^{-10} m	ionic radius 1.40×10^{-10} m	ionic radius 1.81×10^{-10} m

Chemical Bonding

Ionic Bonding – Properties

- The magnesium ion, Mg^{2+}, has a greater charge and smaller ionic radius than the sodium ion, Na^+. *The magnesium ion therefore has a greater charge density than the sodium ion*.
- The oxide ion, O^{2-}, has a greater charge and smaller ionic radius than the chloride ion, Cl^-. *The oxide ion therefore has a greater charge density than the chloride ion*.
- The *strongest* electrostatic force of attraction will be between Mg^{2+} and O^{2-}, and so *MgO* will have the *highest* melting point (m.p. of MgO = +2852°C).
- The *weakest* electrostatic force of attraction will be between Na^+ and Cl^-, and so *NaCl* will have the *lowest* melting point (m.p. of NaCl = +801°C).

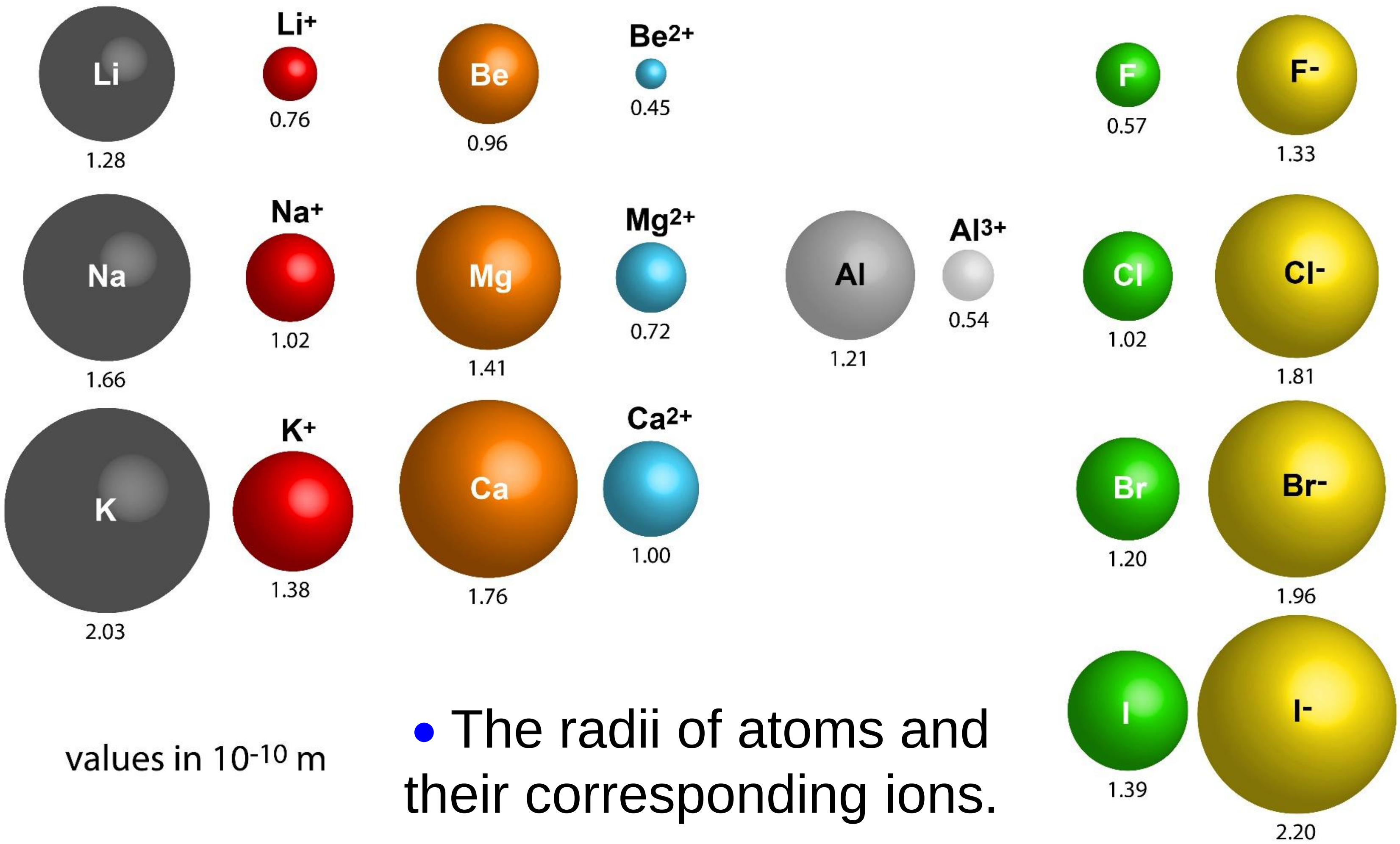
Chemical Bonding

Ionic Bonding – Properties

- Another approach which is maybe less scientific, but easier to understand, is to consider the *difference in charge* between the anion and the cation.
- As a general rule, the *greater* the difference in charge between the anion and cation, the *stronger* the electrostatic force of attraction between the oppositely charged ions and hence the *higher* the ionic compound's melting point and boiling point. **Note**: This approach *does not* take ionic radius or charge density into account.
- For example, consider MgO and NaC*l*. The difference in charge between Mg^{2+} and O^{2-} is *4*. The difference in charge between Na^{+} and Cl^{-} is *2*. It can therefore be predicted that MgO will have a higher melting point than NaC*l*.

Chemical Bonding

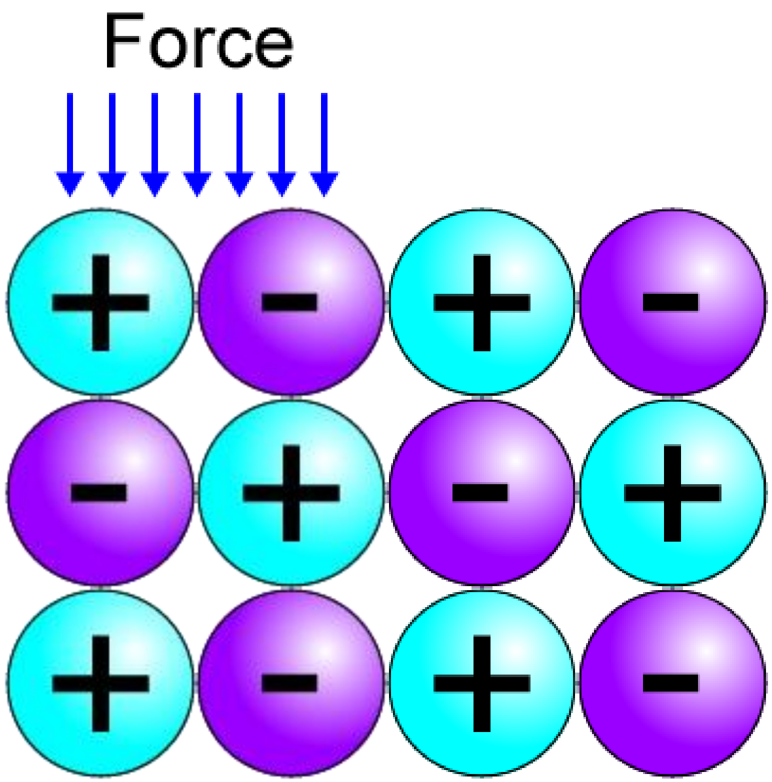
Ionic Bonding – Properties



- The radii of atoms and their corresponding ions.

Chemical Bonding

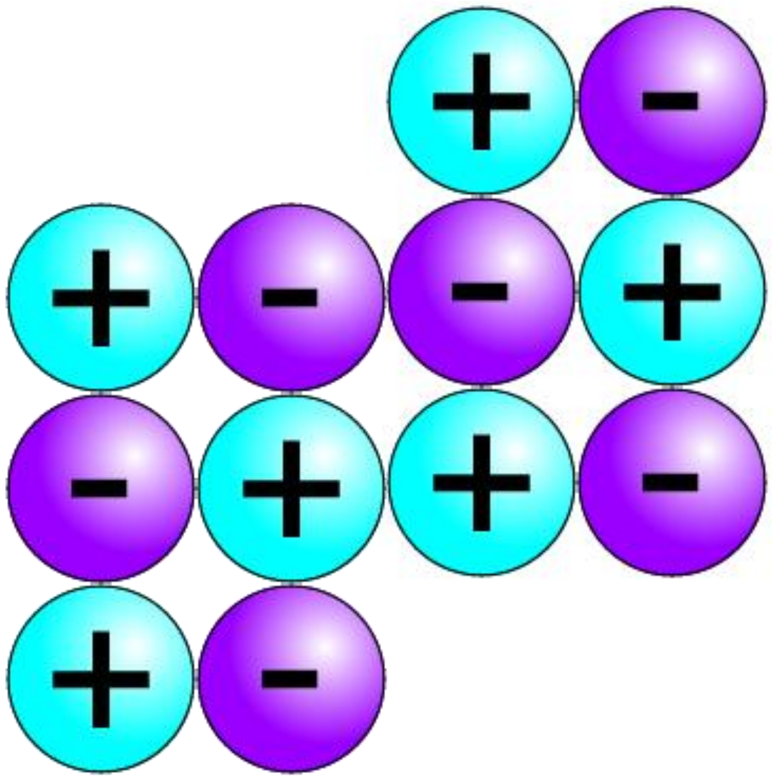
Ionic Bonding – Properties



- Ionic compounds are hard but *brittle*.

Chemical Bonding

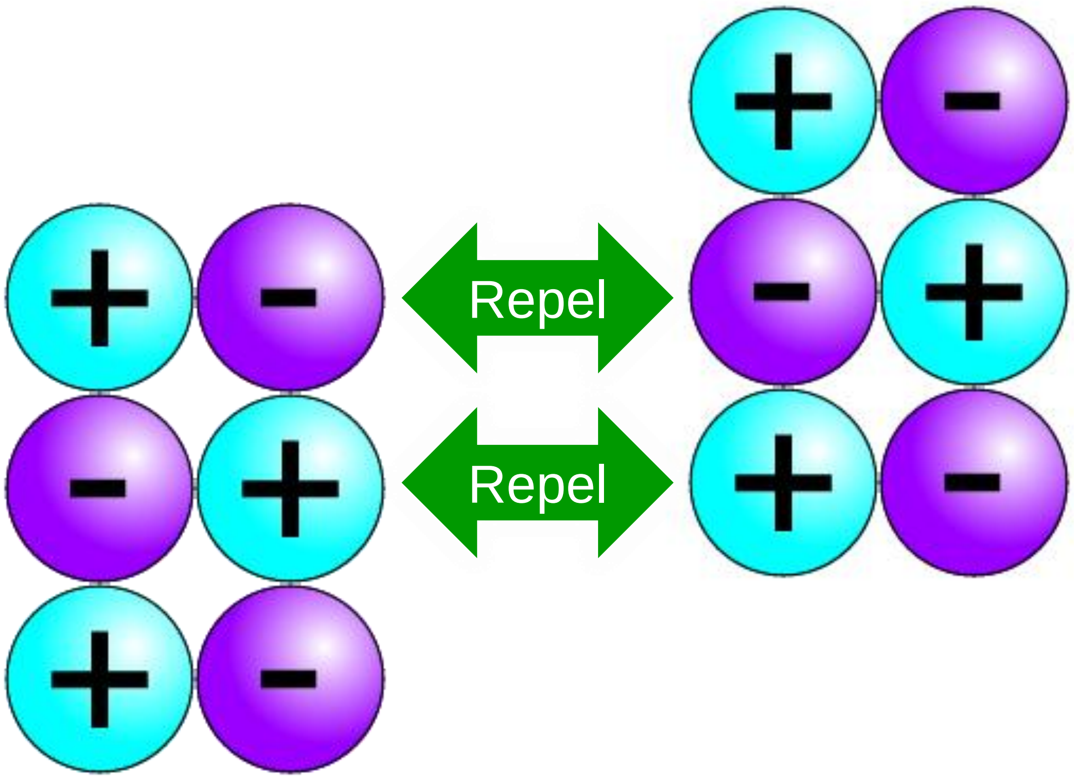
Ionic Bonding – Properties



- When ions of a *similar charge* are forced together, there will be an *electrostatic force of repulsion* between them, causing the ionic crystal to *shatter*.

Chemical Bonding

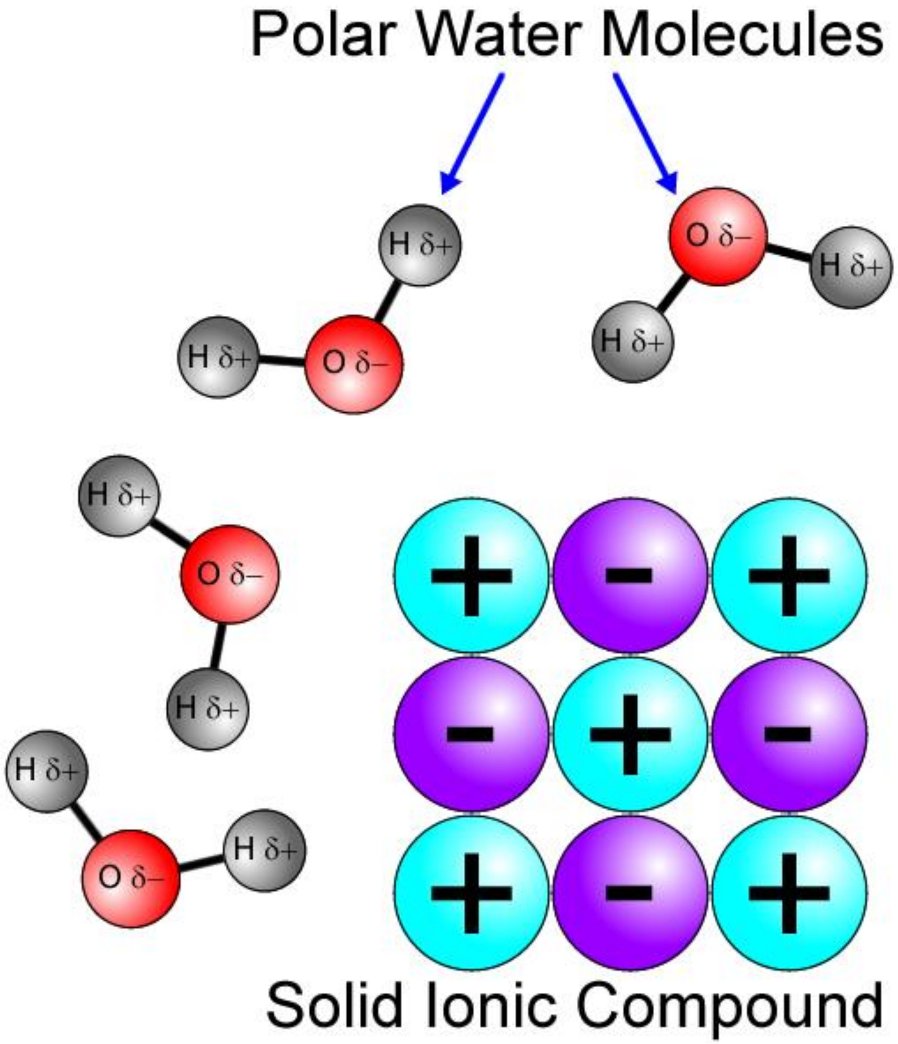
Ionic Bonding – Properties



- When ions of a *similar charge* are forced together, there will be an *electrostatic force of repulsion* between them, causing the ionic crystal to *shatter*.

Chemical Bonding

Ionic Bonding – Properties

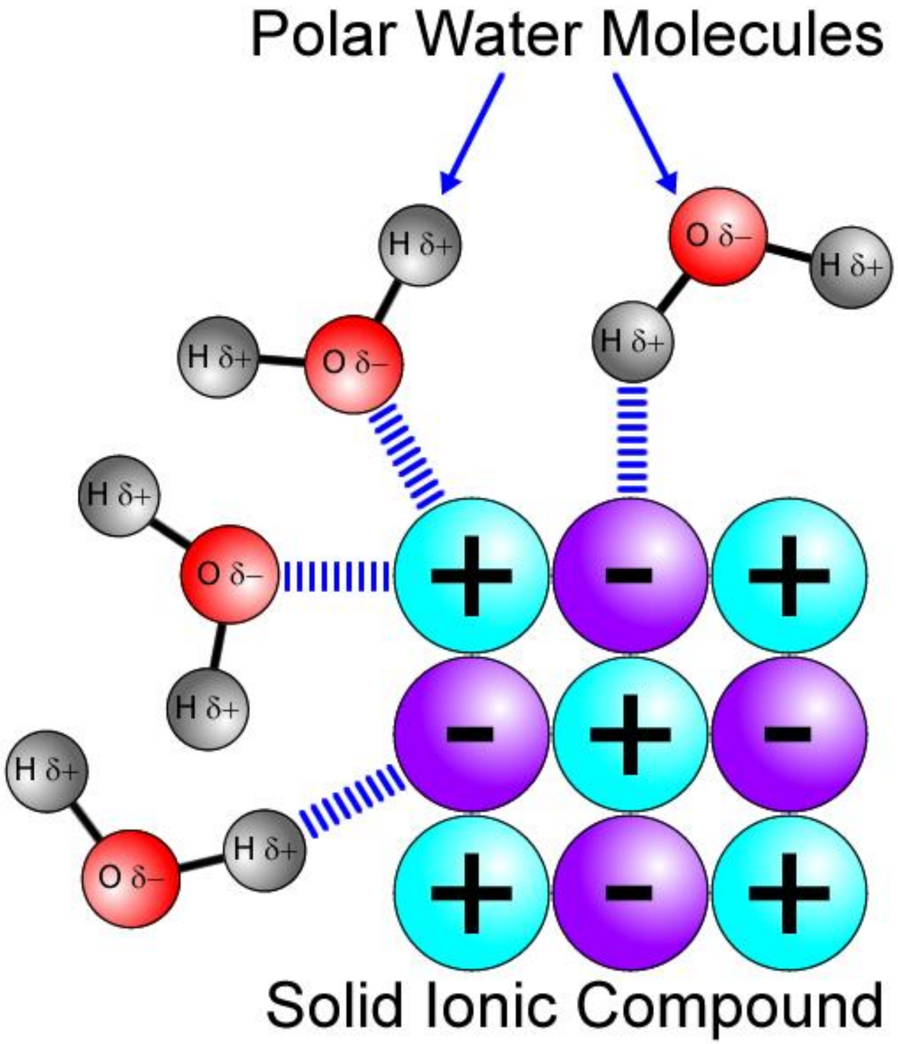


- Ionic compounds are *soluble* in *polar solvents* such as water.

- Ionic compounds are *insoluble* in *non-polar solvents* such as oil and hexane.

Chemical Bonding

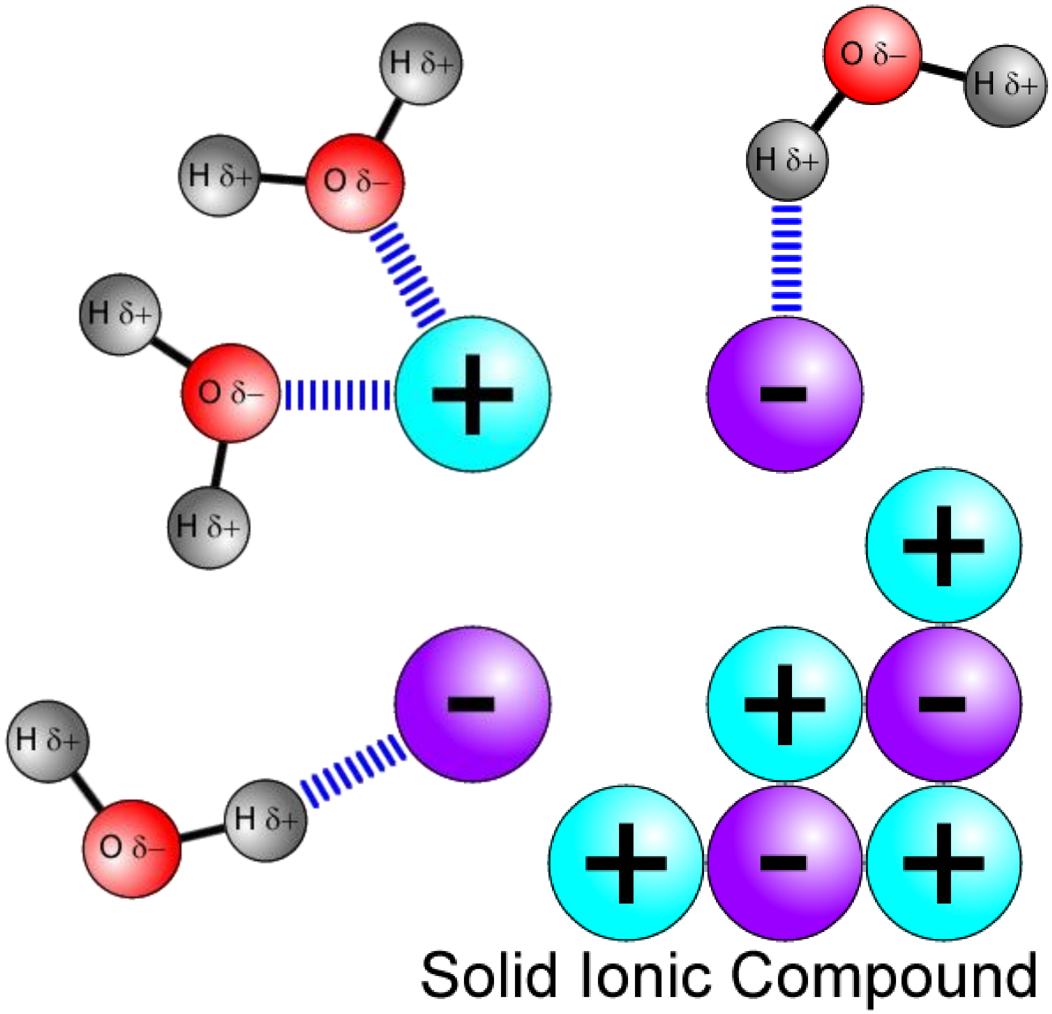
Ionic Bonding – Properties



- Water molecules are described as being *polar*. This means that there is a small distribution of charge over the water molecule (*δ+ H* and *δ– O*). As a result, water molecules are attracted towards positive and negative ions.

Chemical Bonding

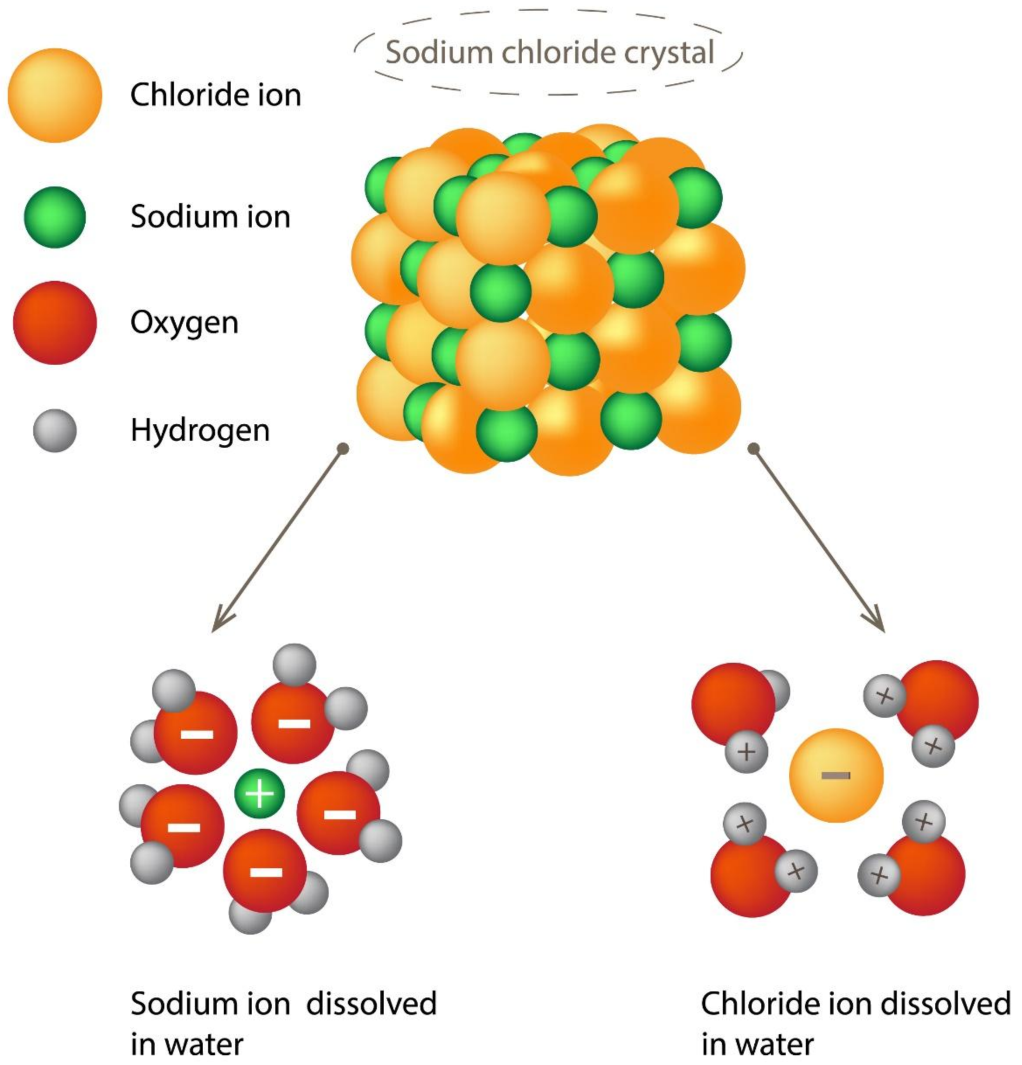
Ionic Bonding – Properties



- There is an *electrostatic force of attraction* between the *δ+ H* of water and *negative anions*, and an *electrostatic force of attraction* between the *δ– O* of water and *positive cations*. Hence ionic compounds *dissolve in water*.

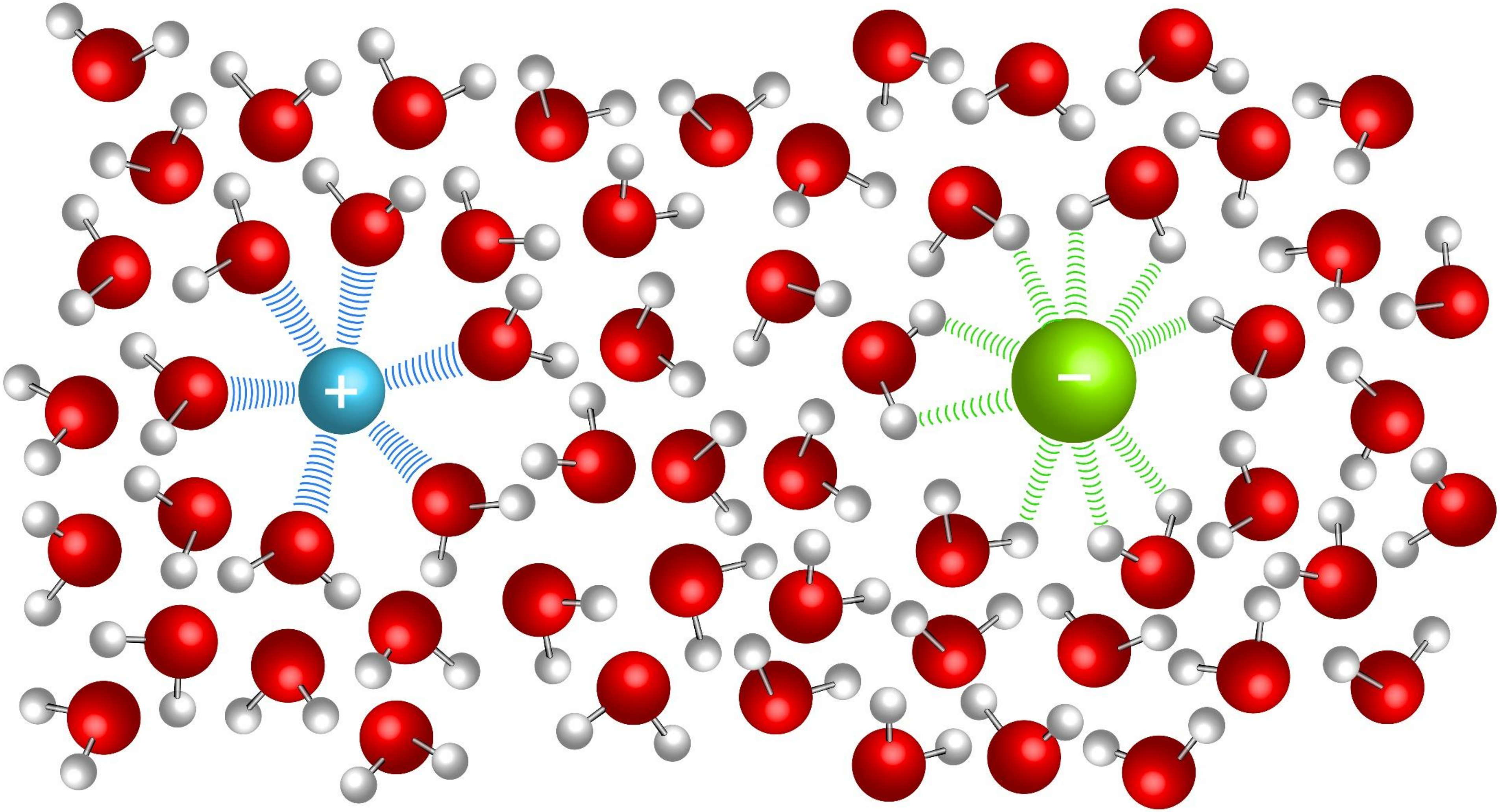
Chemical Bonding

Ionic Bonding – Properties



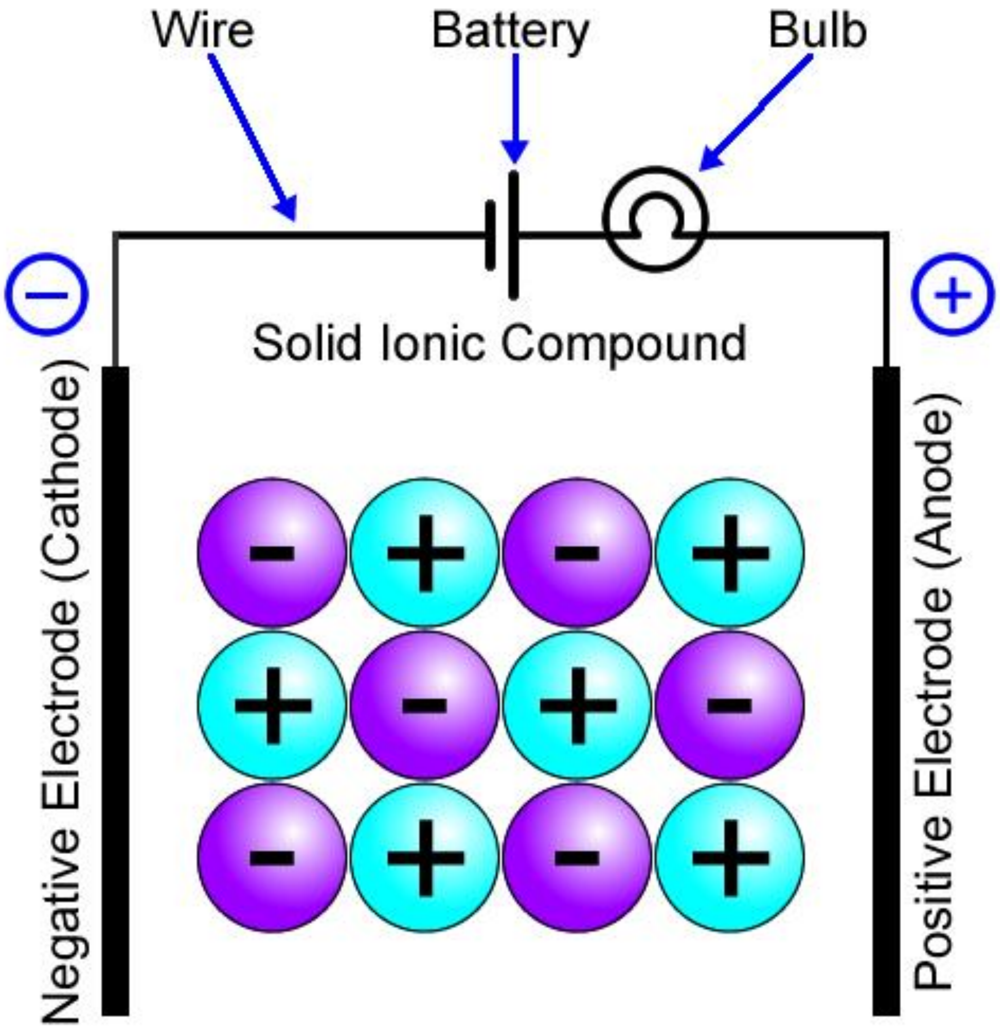
Chemical Bonding

Ionic Bonding – Properties



Chemical Bonding

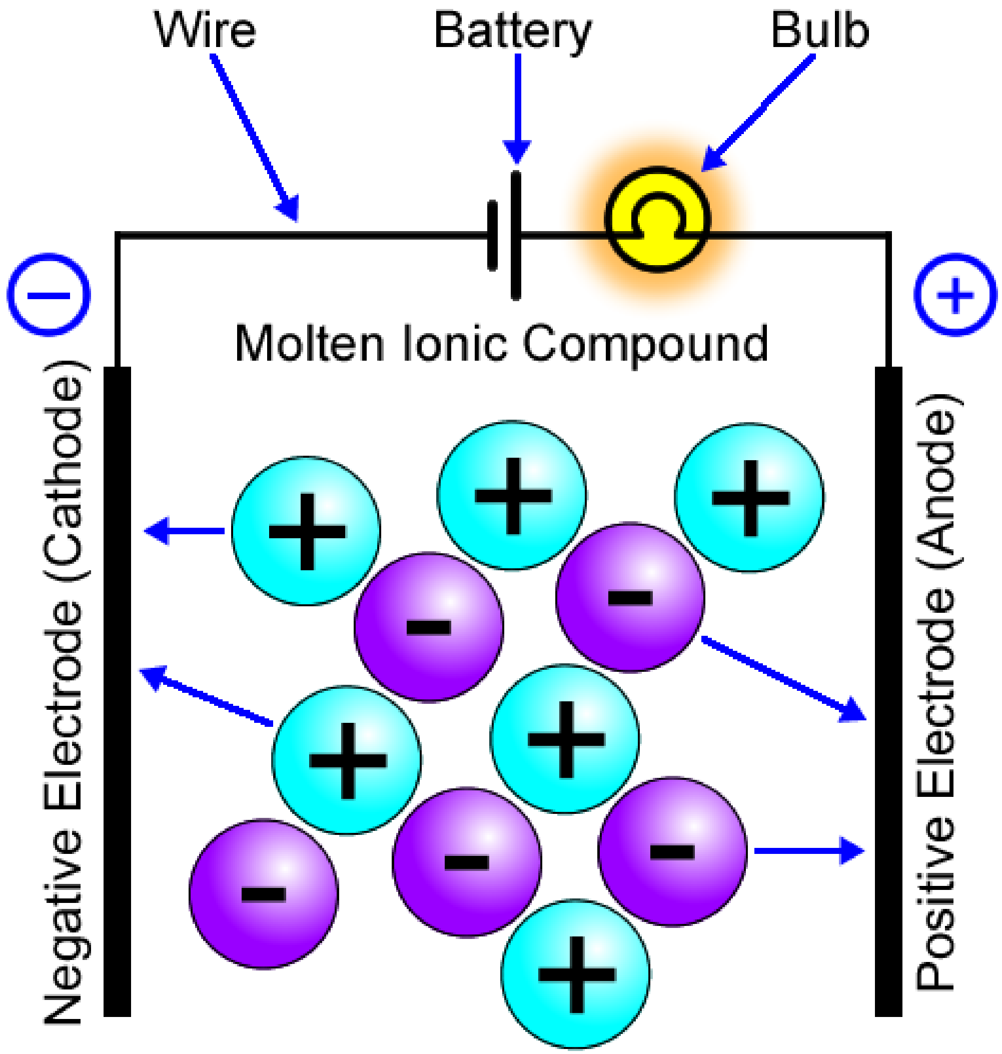
Ionic Bonding – Properties



- Ionic compounds are *electrolytes*. They do not conduct electricity in the solid form, but do conduct electricity when *molten* or when *dissolved in water*.

Chemical Bonding

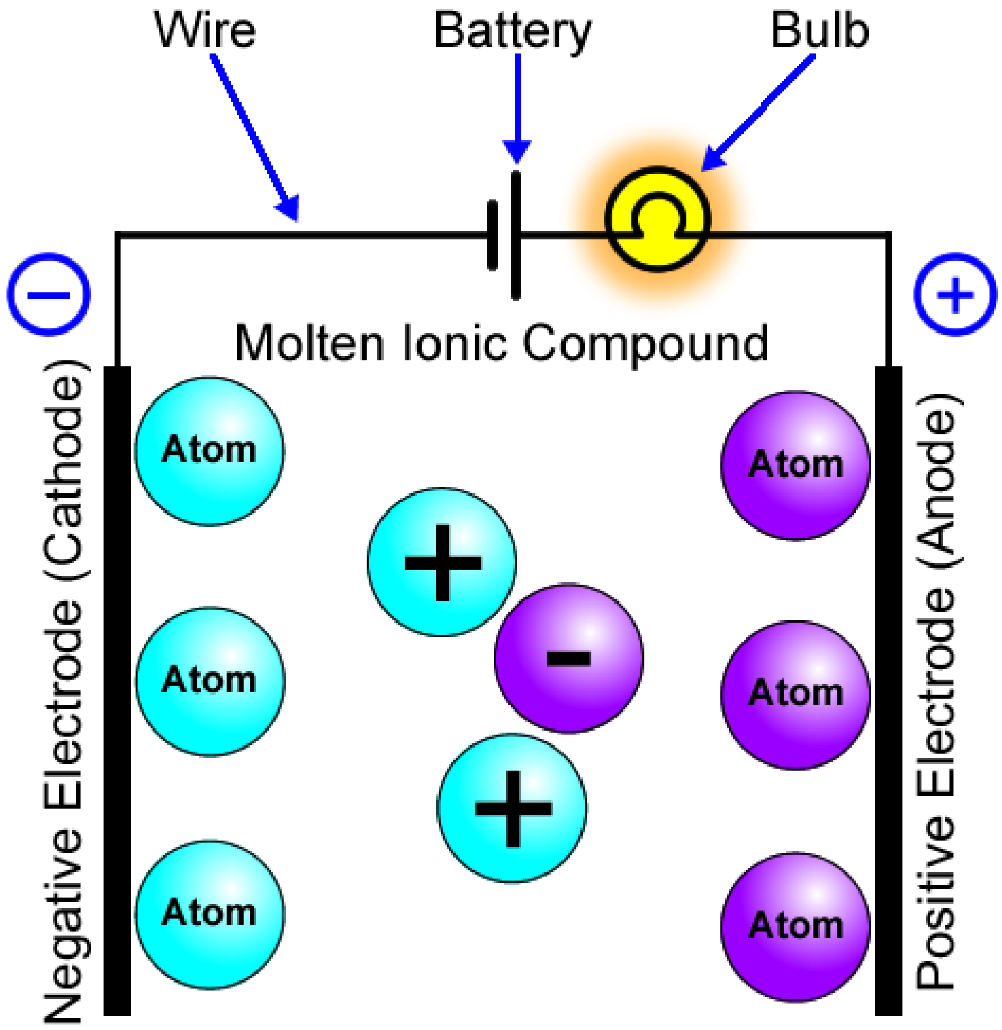
Ionic Bonding – Properties



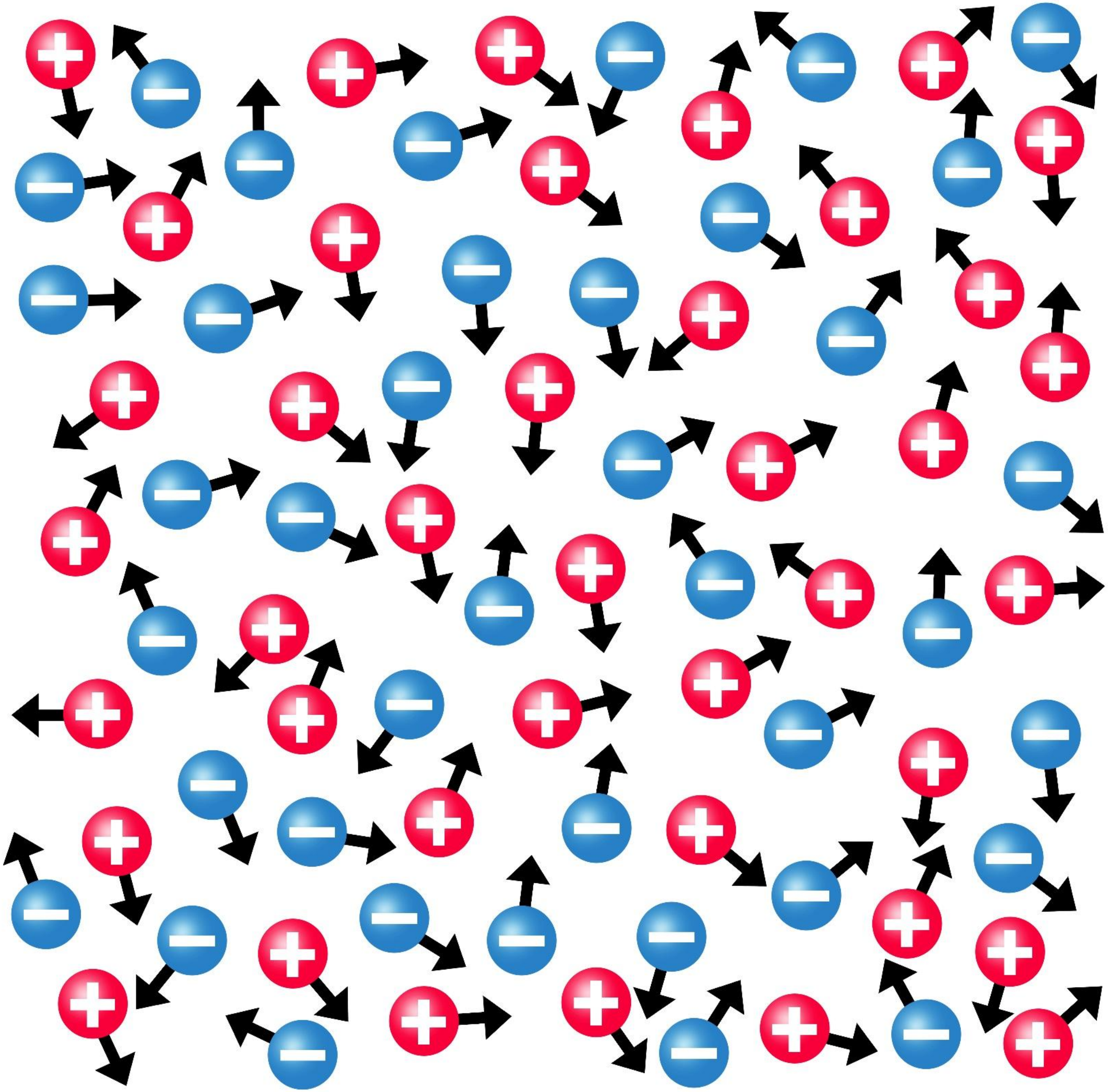
- Kinetic particle theory states that in a solid ionic compound, the positive and negative ions *vibrate about a fixed position*. They are unable to move towards the electrode of opposite charge.

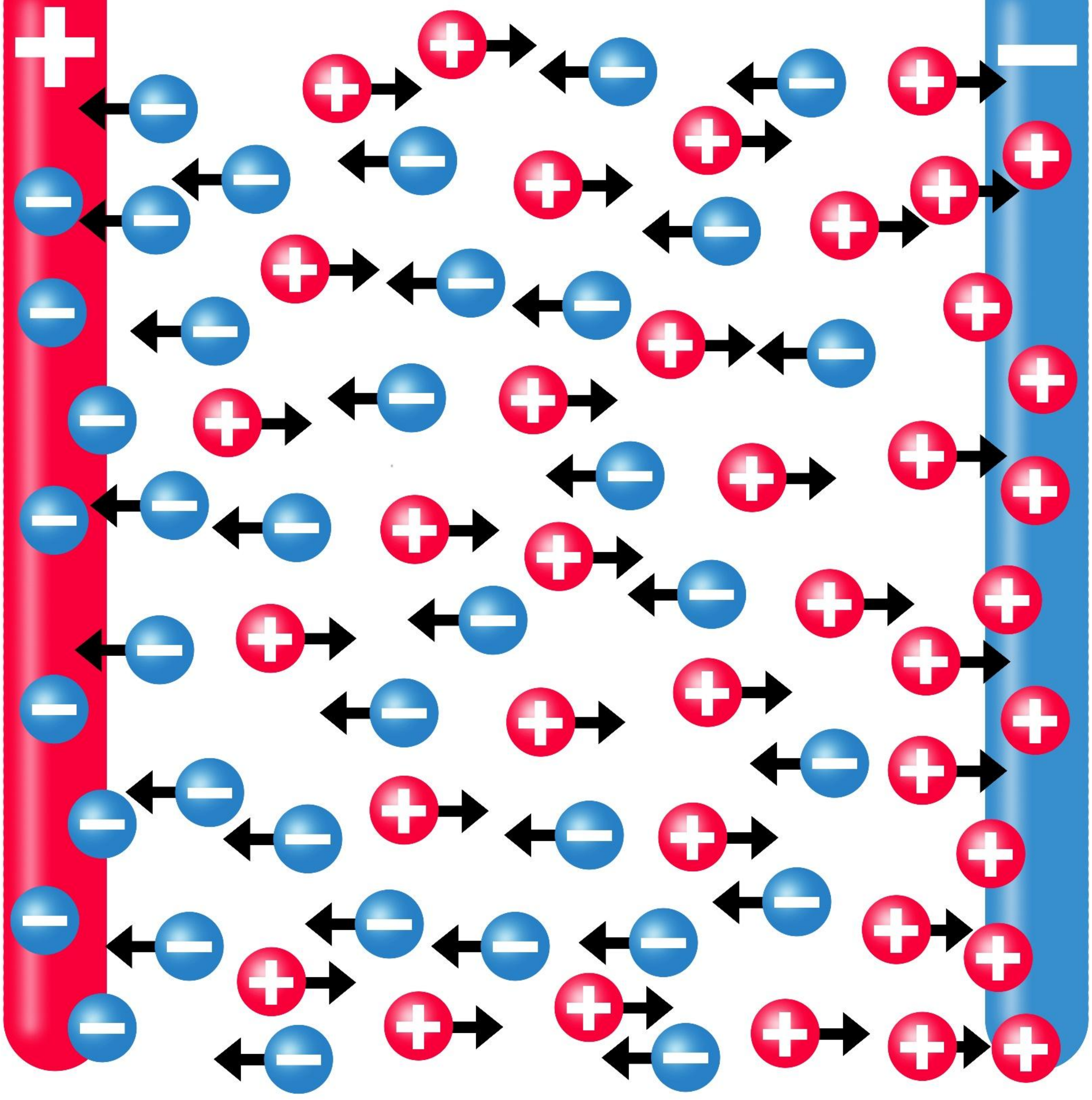
Chemical Bonding

Ionic Bonding – Properties



- When the ionic compound is *molten* or when it is *dissolved in water*, the positive and negative ions become *mobile* and are free to move towards the electrode of opposite charge, thus *conducting electricity*.





Chemical Bonding

Ionic Bonding – Properties – Summary

- Ionic compounds have *high melting points* and *boiling points* due to the strong electrostatic force of attraction that holds the positive and negative ions together.

- Ionic compounds are *hard but brittle*. When ions with a similar charge are forced together they repel, and the ionic crystal shatters.

- Ionic compounds are *soluble in polar solvents* such as water, but are *insoluble in non-polar solvents* such as oil and hexane.

- Ionic compounds are *electrolytes*. They do *not* conduct electricity when in the *solid state*, but *do* conduct electricity when either *molten* or *dissolved in water*.

Online Search

- Identify an ionic compound that is present in *toothpaste*. How is its function related to its structure and bonding?

- What ionic compound is *bone* composed of? How are its properties related to its structure and bonding?

- What ionic compound are the shells of *molluscs* composed of? How are its properties related to its structure and bonding?

Chemical Bonding

Covalent Bonding – Non-metals Only!

7. How do two or more atoms of *non-metallic chemical elements* react together to obtain noble gas electronic configurations?

Chemical Bonding

Covalent Bonding – Non-metals Only!

- Covalent bonding occurs when an atom of a *non-metallic* element bonds to another atom of a *non-metallic* element.

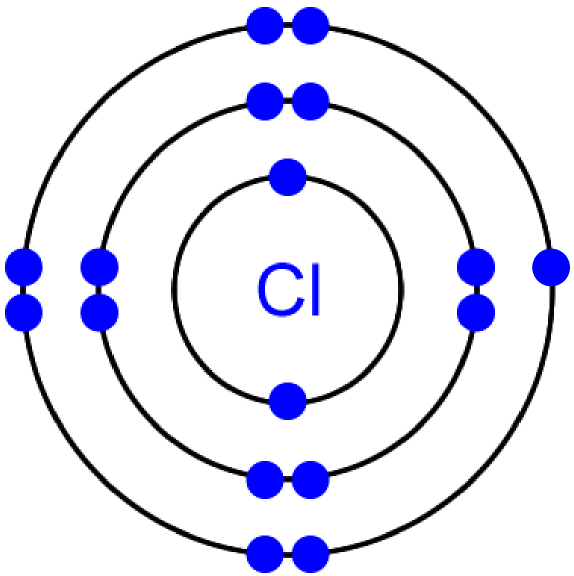
Group

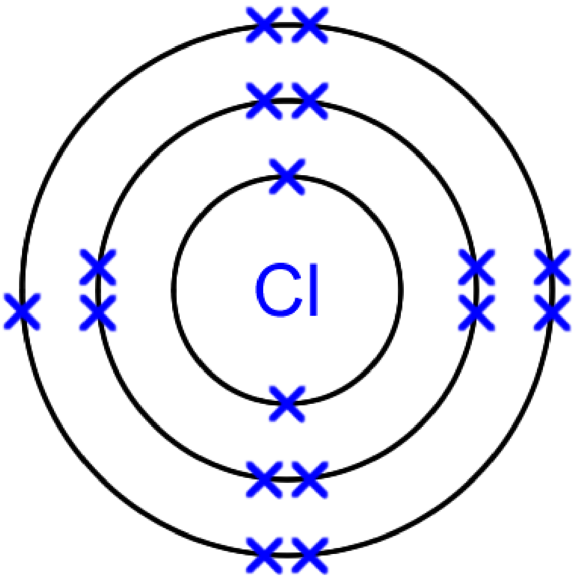
Period	1	2											13	14	15	16	17	18
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac															

Key:
Blue = Metal
Yellow = Non-metal

Chemical Bonding

Covalent Bonding



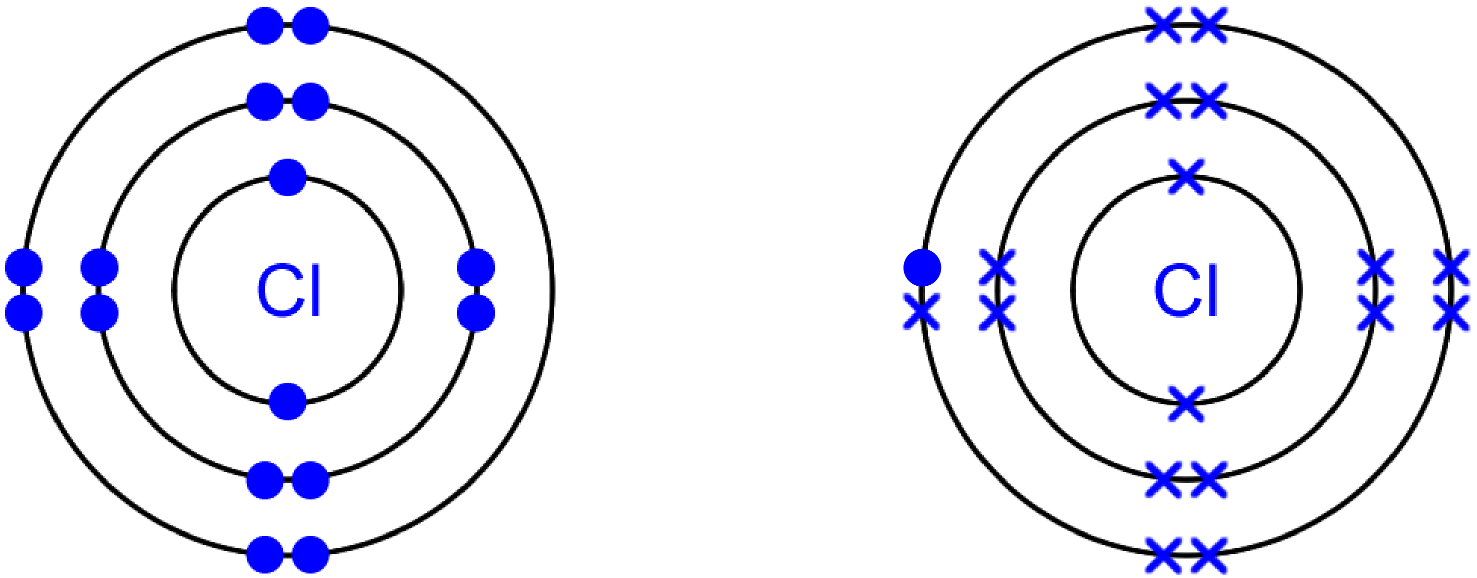


- Can two atoms of chlorine *both* obtain noble gas electronic configurations through *electron transfer*?

• = Electron of chlorine (left-hand-side). × = Electron of chlorine (right-hand-side).

Chemical Bonding

Covalent Bonding

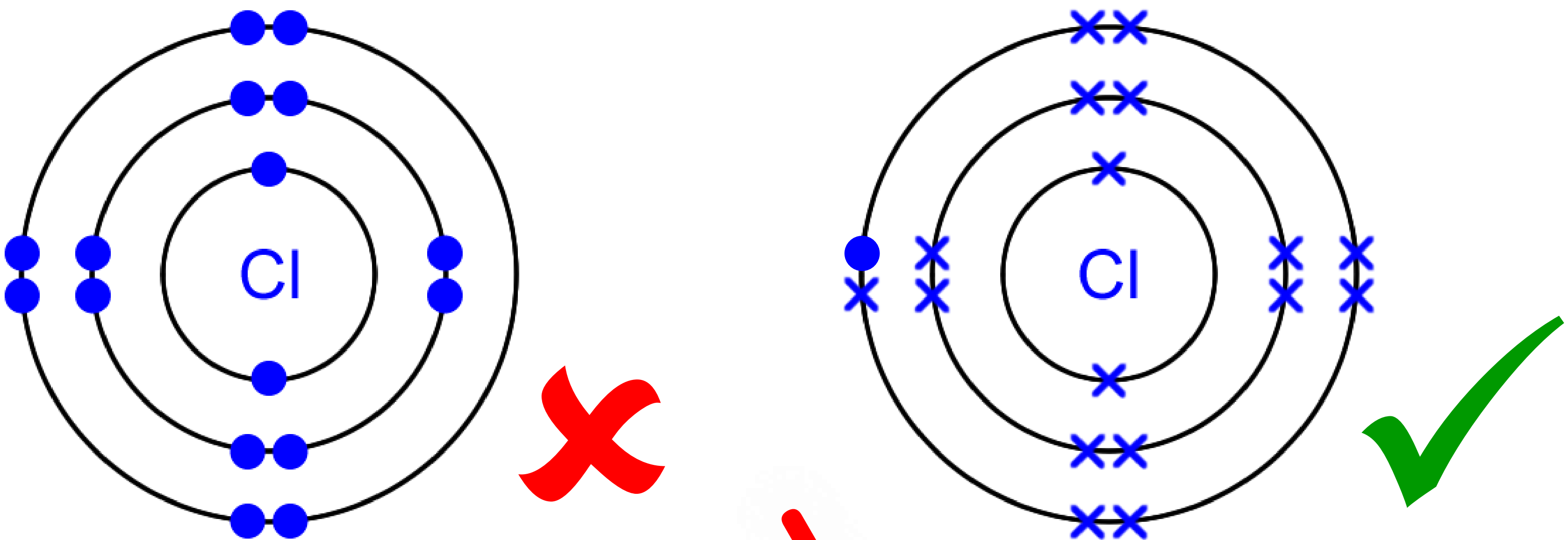


• Can two atoms of chlorine *both* obtain noble gas electronic configurations through *electron transfer*?

• = Electron of chlorine (left-hand-side). × = Electron of chlorine (right-hand-side).

Chemical Bonding

Covalent Bonding



- Can two atoms of chlorine *both* obtain noble gas electronic configurations through *electron transfer*?

No!

• = Electron of chlorine (left-hand-side). × = Electron of chlorine (right-hand-side).

Chemical Bonding

Covalent Bonding – Fluorine

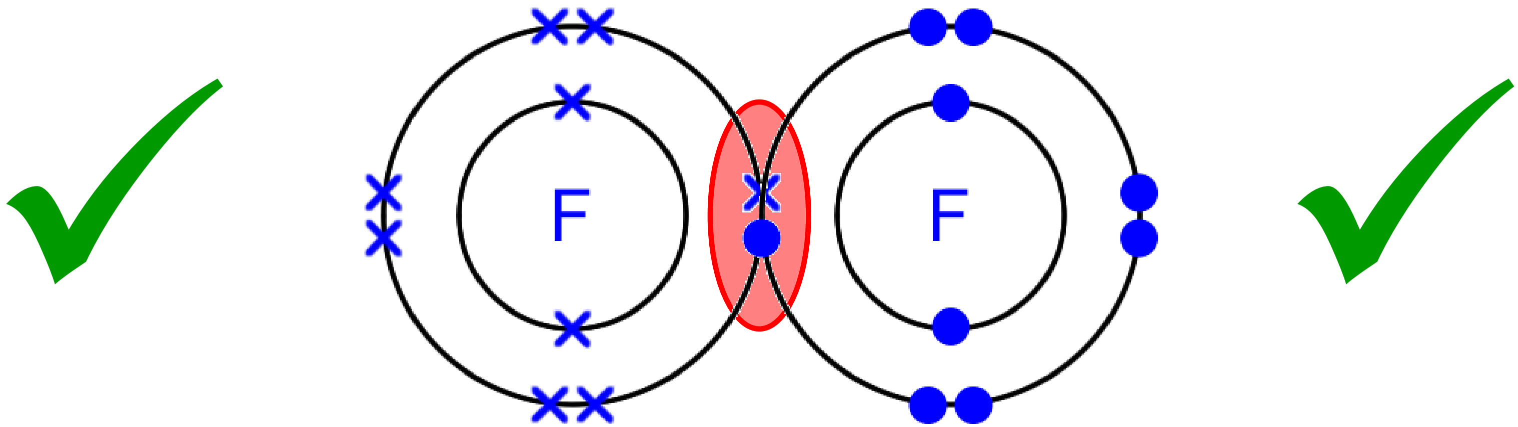
- In *covalent bonding*, atoms of the *non-metallic* elements must *join together* and *share electrons* in order for every atom to obtain a *noble gas electronic configuration*. **Note:** a covalent bond is a *shared pair of electrons*.

• = Electron of fluorine (right-hand-side). × = Electron of fluorine (left-hand-side).

Chemical Bonding

Covalent Bonding – Fluorine

Molecule of fluorine, formula: F_2



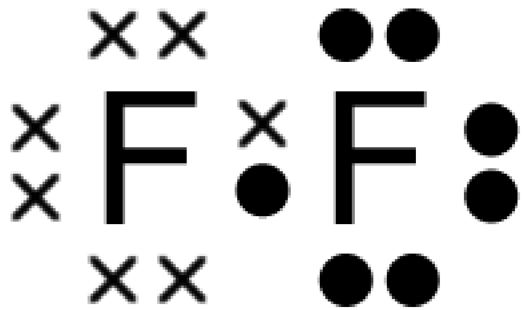
• In *covalent bonding*, atoms of the *non-metallic* elements must *join together* and *share electrons* in order for every atom to obtain a *noble gas electronic configuration*. **Note:** a covalent bond is a *shared pair of electrons*.

• = Electron of fluorine (right-hand-side). × = Electron of fluorine (left-hand-side).

Chemical Bonding

Covalent Bonding – Fluorine

Molecule of fluorine, formula: F_2



- The dot-and-cross diagram for molecular fluorine that should be used as an answer in an examination.

• = Electron of fluorine (right-hand-side). × = Electron of fluorine (left-hand-side).

Chemical Bonding

Covalent Bonding – Hydrogen

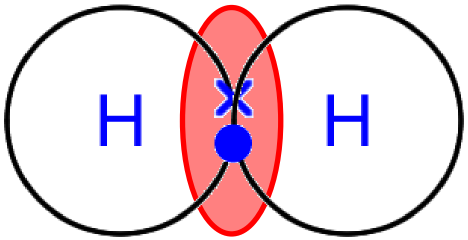
• In *covalent bonding*, atoms of the *non-metallic* elements must *join together* and *share electrons* in order for every atom to obtain a *noble gas electronic configuration*. **Note:** a covalent bond is a *shared pair of electrons*.

• = Electron of hydrogen (right-hand-side). × = Electron of hydrogen (left-hand-side).

Chemical Bonding

Covalent Bonding – Hydrogen

Molecule of hydrogen, formula: H_2



- In *covalent bonding*, atoms of the *non-metallic* elements must *join together* and *share electrons* in order for every atom to obtain a *noble gas electronic configuration*. **Note:** a covalent bond is a *shared pair of electrons*.

• = Electron of hydrogen (right-hand-side). × = Electron of hydrogen (left-hand-side).

Chemical Bonding

Covalent Bonding – Hydrogen

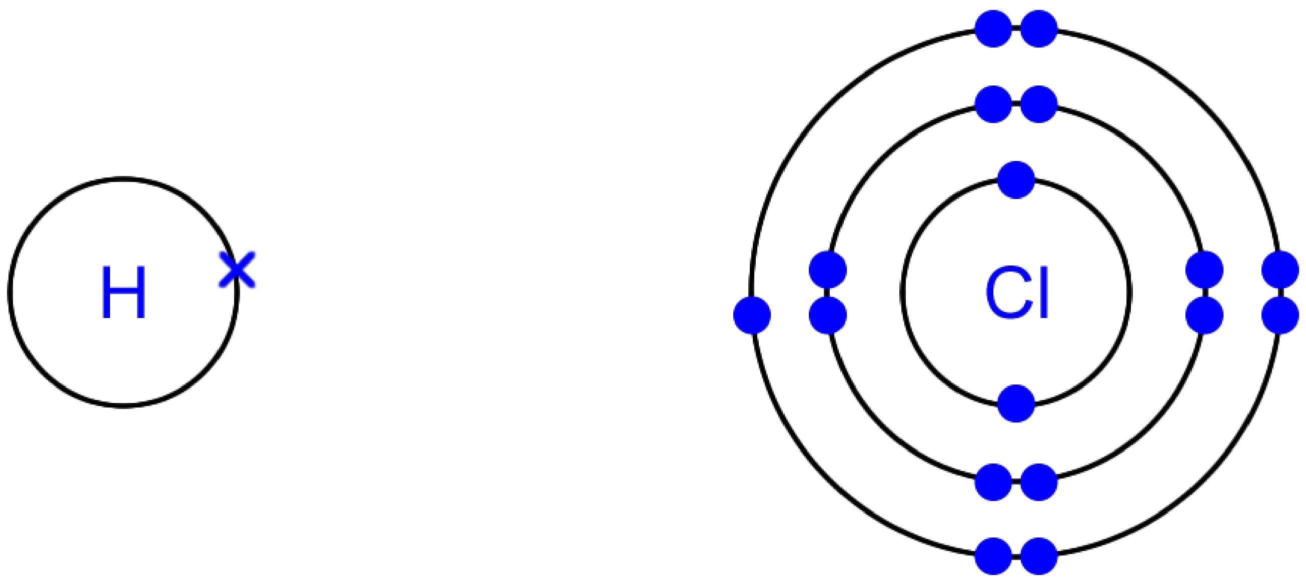
Molecule of hydrogen, formula: H_2

- The dot-and-cross diagram for molecular hydrogen that should be used as an answer in an examination.

• = Electron of hydrogen (right-hand-side). × = Electron of hydrogen (left-hand-side).

Chemical Bonding

Covalent Bonding – Hydrogen Chloride



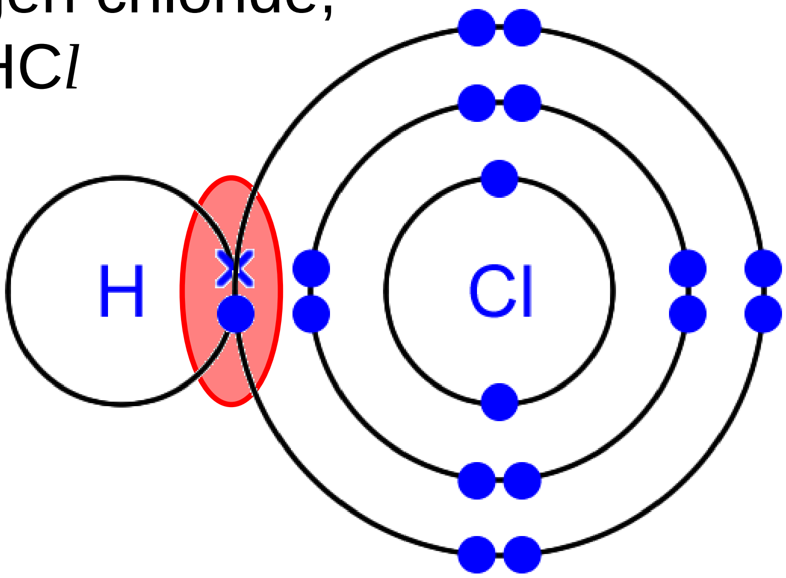
- In *covalent bonding*, atoms of the *non-metallic* elements must *join together* and *share electrons* in order for every atom to obtain a *noble gas electronic configuration*. **Note:** a covalent bond is a *shared pair of electrons*.

• = Electron of chlorine. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Hydrogen Chloride

Molecule of hydrogen chloride,
formula: HC*l*



• In *covalent bonding*, atoms of the *non-metallic* elements must *join together* and *share electrons* in order for every atom to obtain a *noble gas electronic configuration*. **Note:** a covalent bond is a *shared pair of electrons*.

• = Electron of chlorine. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Hydrogen

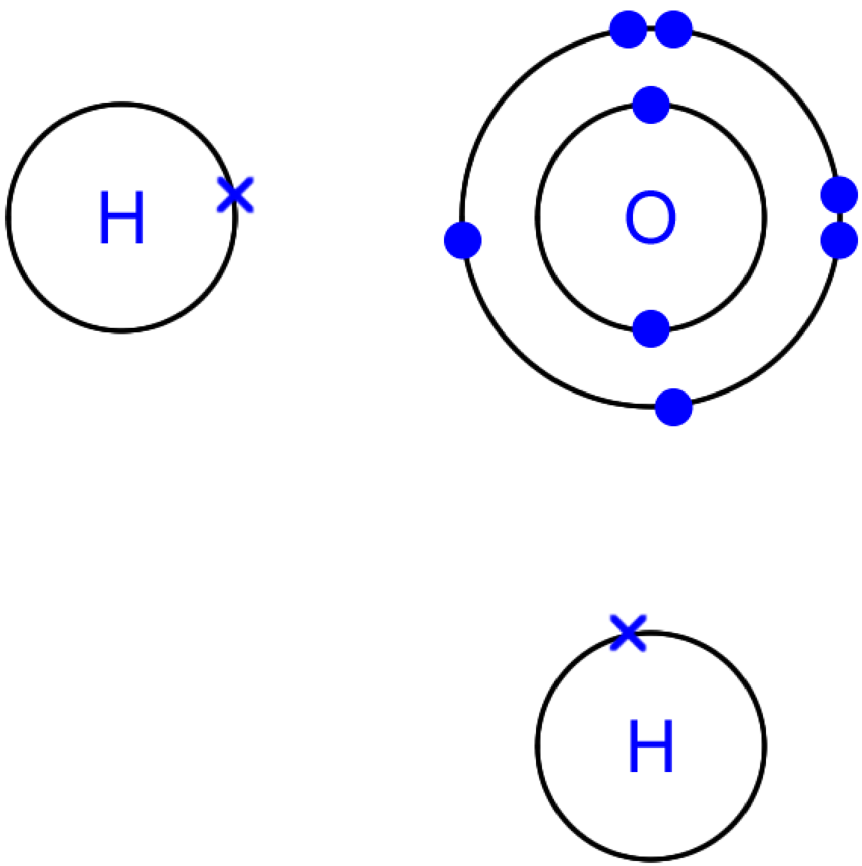
Molecule of hydrogen chloride, formula: HC*l*

- The dot-and-cross diagram for hydrogen chloride that should be used as an answer in an examination.

• = Electron of chlorine. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Water

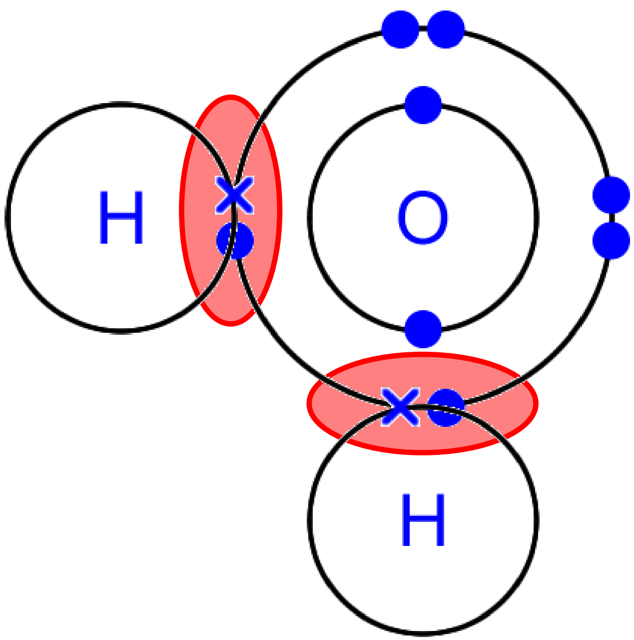


● = Electron of oxygen. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Water

Molecule of water, formula: H_2O

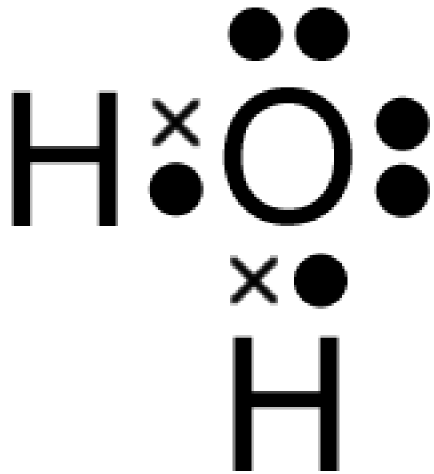


● = Electron of oxygen. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Water

Molecule of water, formula: H_2O

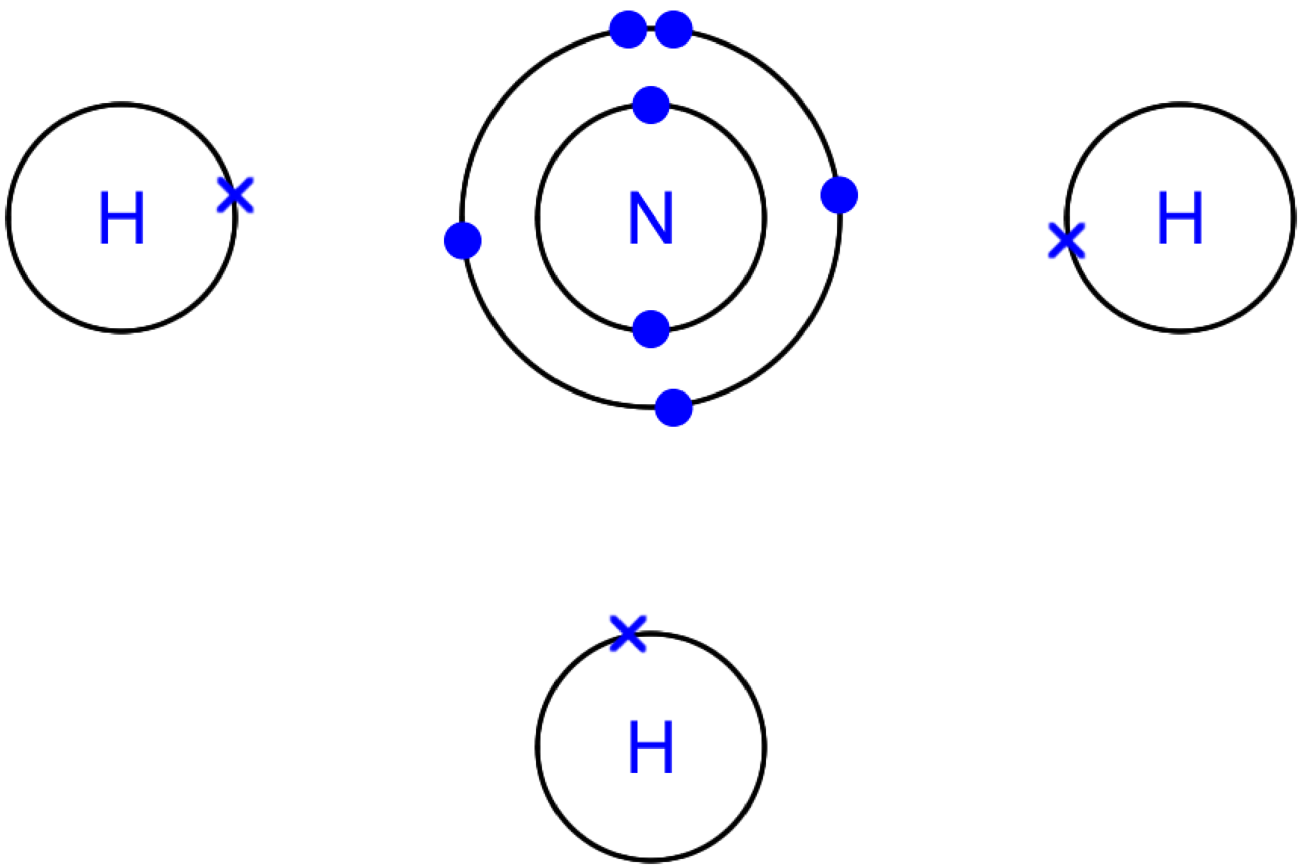


- The dot-and-cross diagram for water that should be used as an answer in an examination.

• = Electron of oxygen. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Ammonia

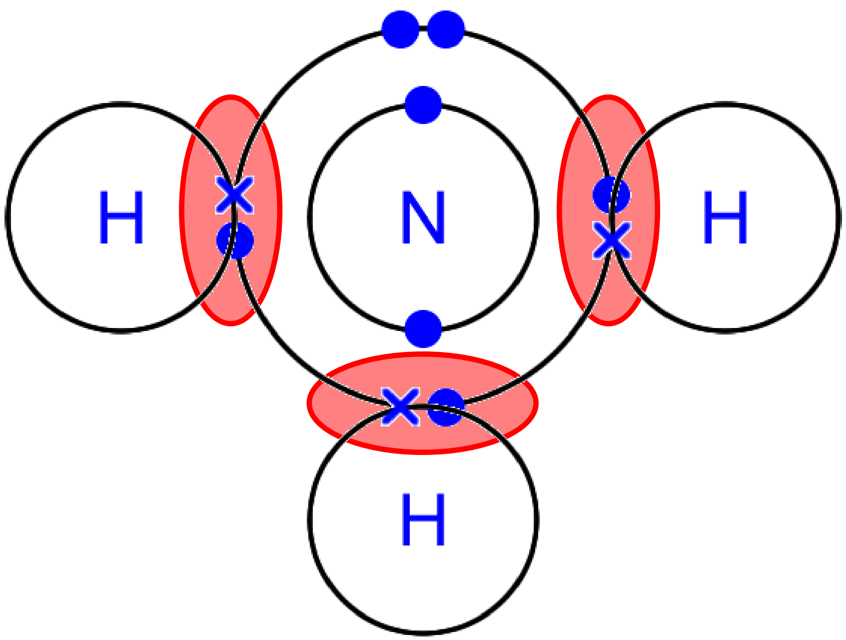


● = Electron of nitrogen. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Ammonia

Molecule of ammonia, formula: NH_3

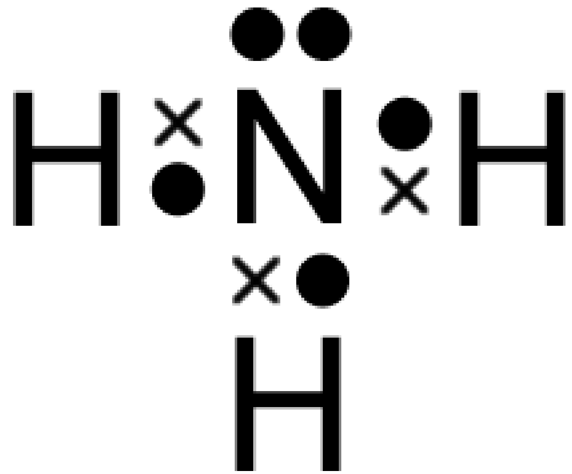


● = Electron of nitrogen. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Ammonia

Molecule of ammonia, formula: NH_3

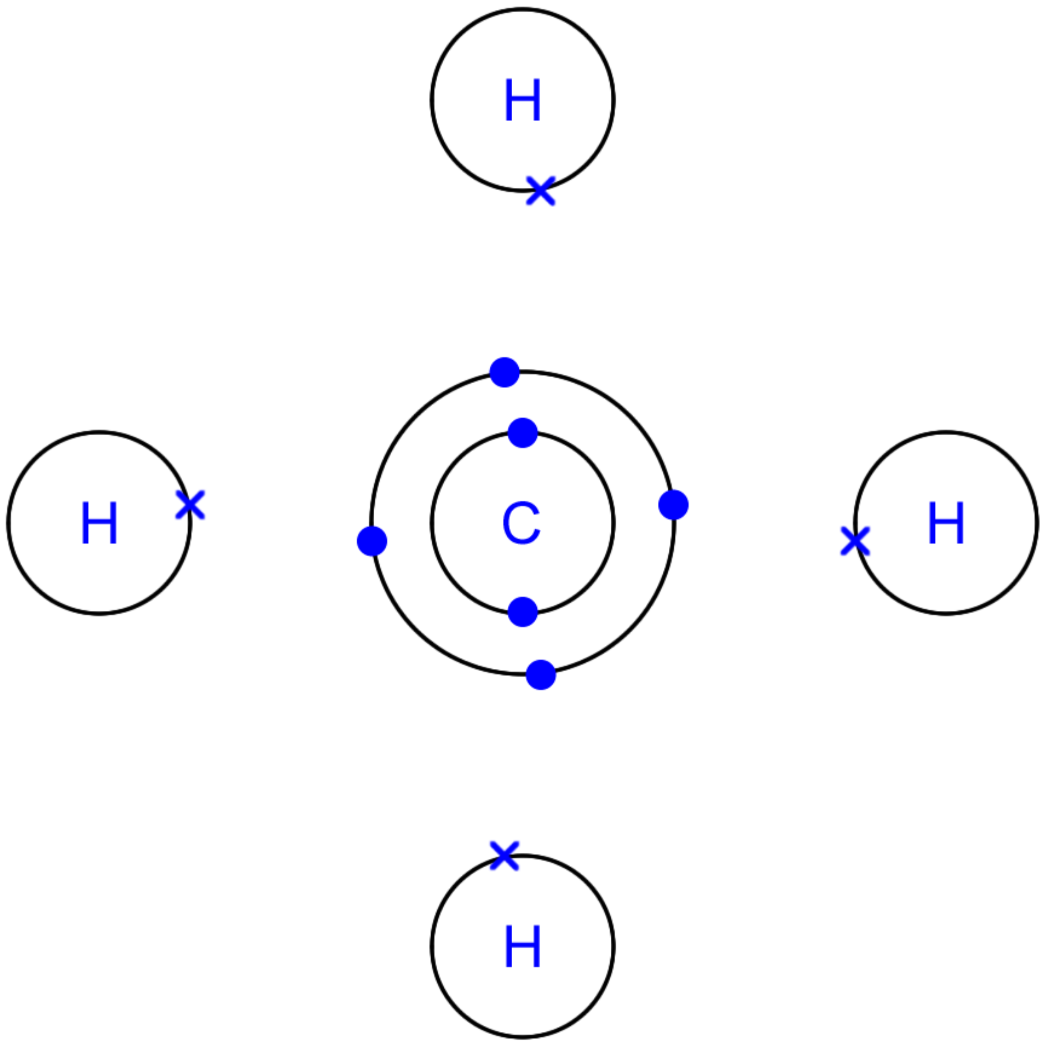


- The dot-and-cross diagram for ammonia that should be used as an answer in an examination.

• = Electron of nitrogen. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Methane

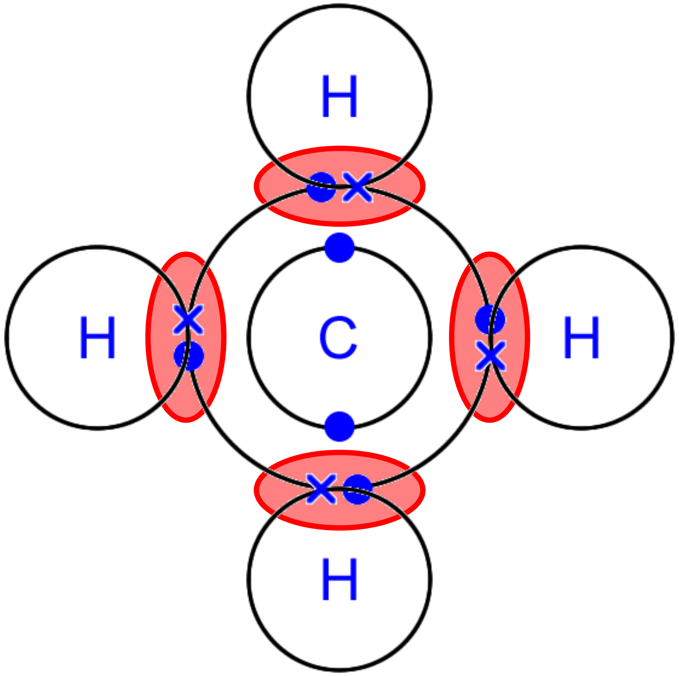


• = Electron of carbon. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Methane

Molecule of methane, formula: CH_4

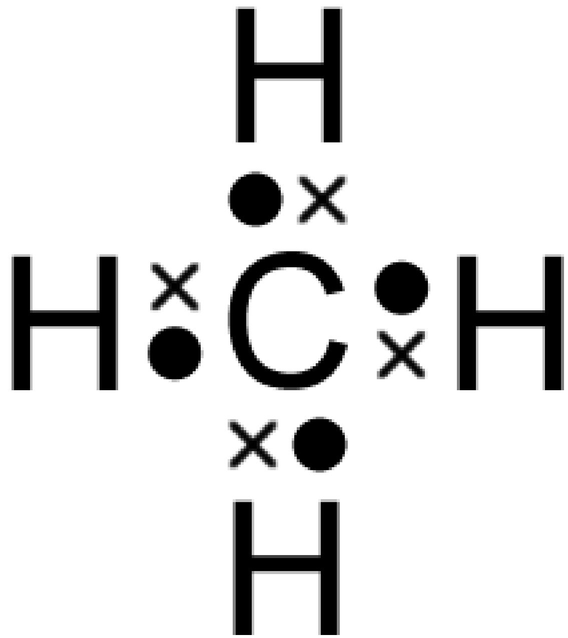


● = Electron of carbon. × = Electron of hydrogen.

Chemical Bonding

Covalent Bonding – Methane

Molecule of methane, formula: CH_4

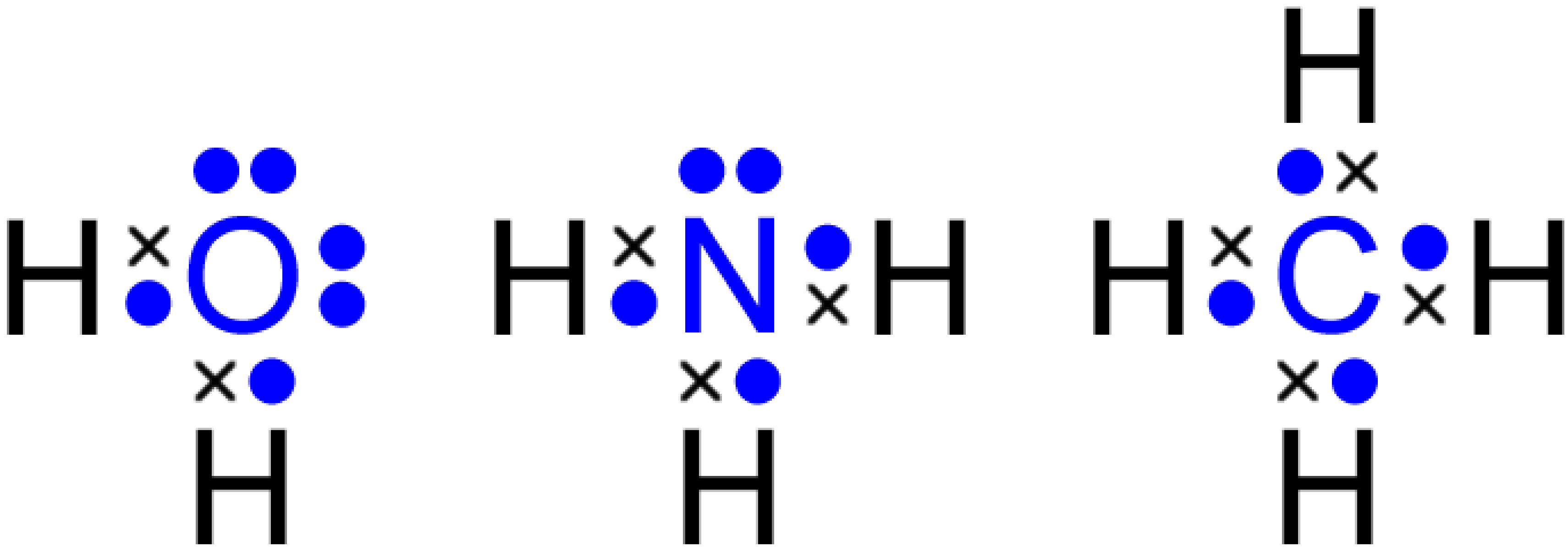


- The dot-and-cross diagram for methane that should be used as an answer in an examination.

• = Electron of carbon. × = Electron of hydrogen.

Chemical Bonding

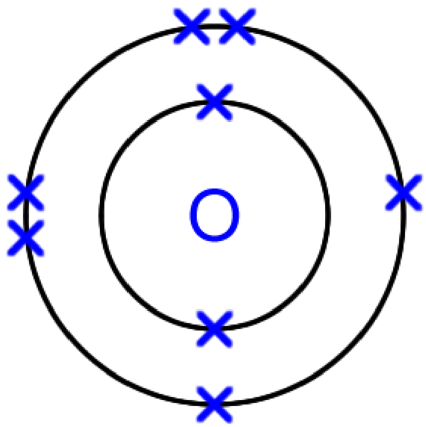
Covalent Bonding

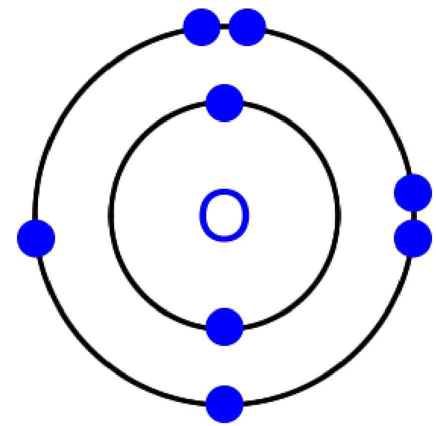


- **Note**: For *covalent compounds*, atom(s) of the chemical element present in the *smallest* number are usually drawn at the *centre* of the molecule. Atoms of the chemical element(s) present in *larger* numbers are then bonded around the *outside*.

Chemical Bonding

Covalent Bonding – Oxygen

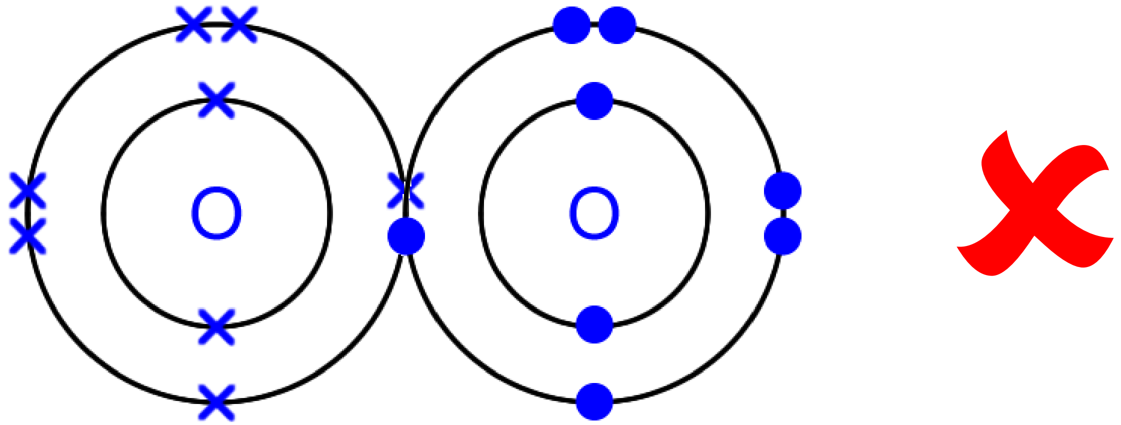




• = Electron of oxygen (right-hand-side). × = Electron of oxygen (left-hand-side).

Chemical Bonding

Covalent Bonding – Oxygen

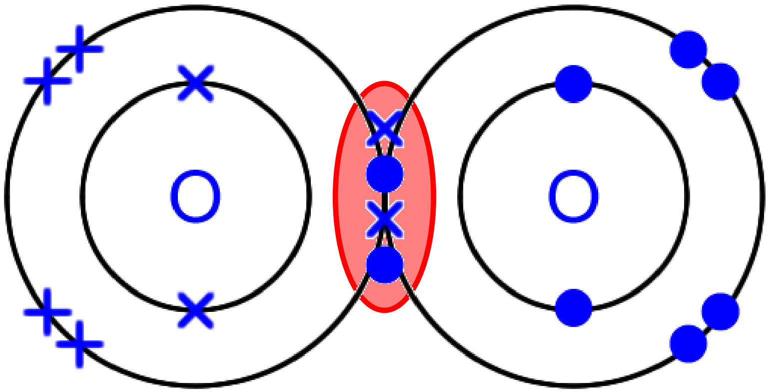


● = Electron of oxygen (right-hand-side). × = Electron of oxygen (left-hand-side).

Chemical Bonding

Covalent Bonding – Oxygen

Molecule of oxygen, formula: O_2



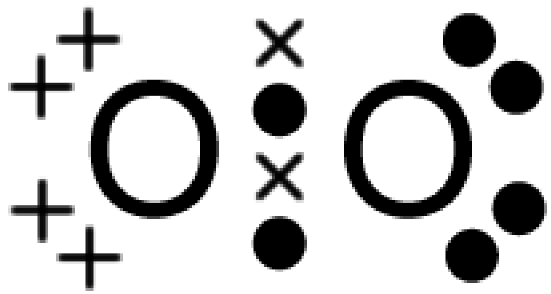
- **Note:** *Double covalent bond*.

• = Electron of oxygen (right-hand-side). × = Electron of oxygen (left-hand-side).

Chemical Bonding

Covalent Bonding – Oxygen

Molecule of oxygen, formula: O_2



- The dot-and-cross diagram for molecular oxygen that should be used as an answer in an examination.

• = Electron of oxygen (right-hand-side). × = Electron of oxygen (left-hand-side).

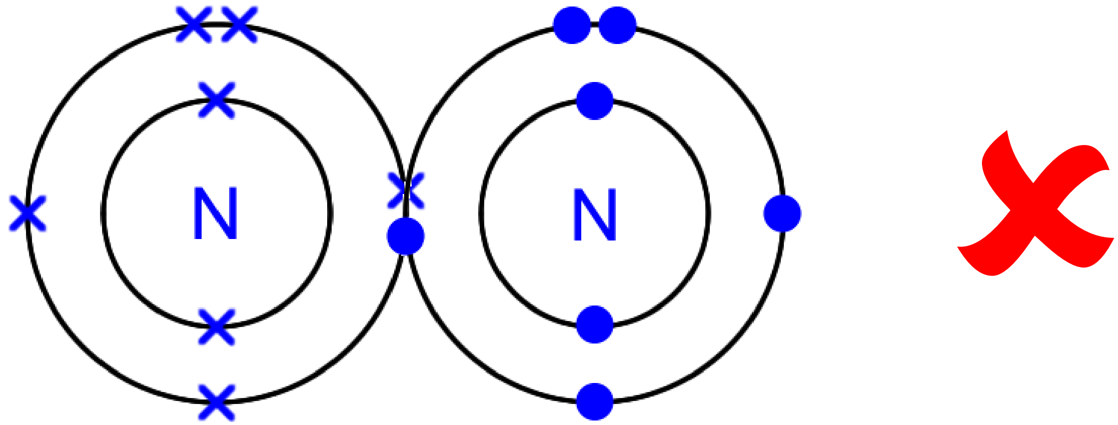
Chemical Bonding

Covalent Bonding – Nitrogen

● = Electron of nitrogen (right-hand-side). × = Electron of nitrogen (left-hand-side).

Chemical Bonding

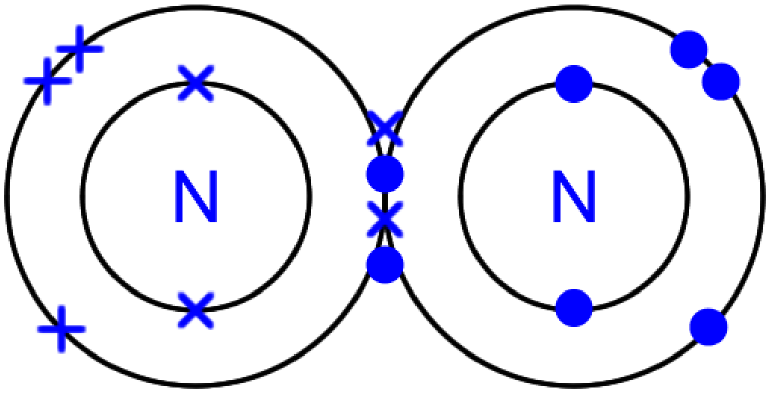
Covalent Bonding – Nitrogen



• = Electron of nitrogen (right-hand-side). × = Electron of nitrogen (left-hand-side).

Chemical Bonding

Covalent Bonding – Nitrogen

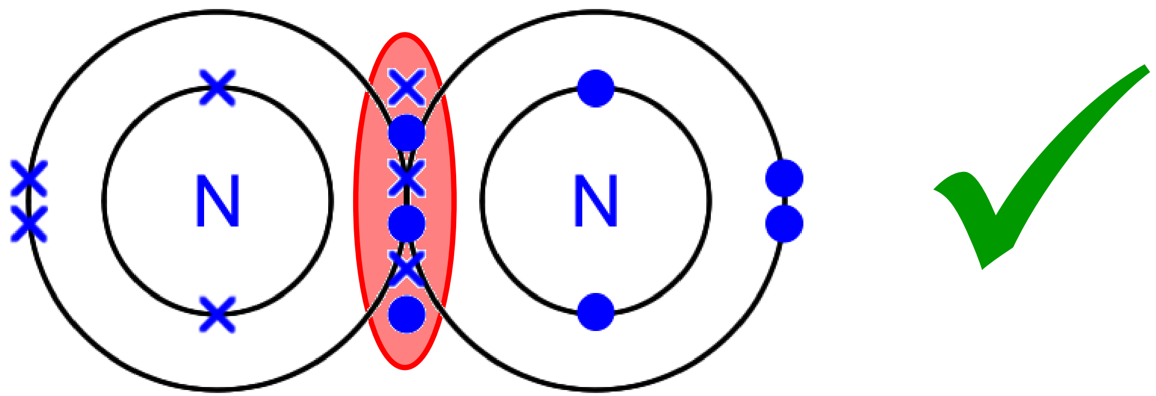


• = Electron of nitrogen (right-hand-side).   × = Electron of nitrogen (left-hand-side).

Chemical Bonding

Covalent Bonding – Nitrogen

Molecule of nitrogen, formula: N_2



- **Note:** *Triple covalent bond*.

• = Electron of nitrogen (right-hand-side). × = Electron of nitrogen (left-hand-side).

Chemical Bonding

Covalent Bonding – Nitrogen

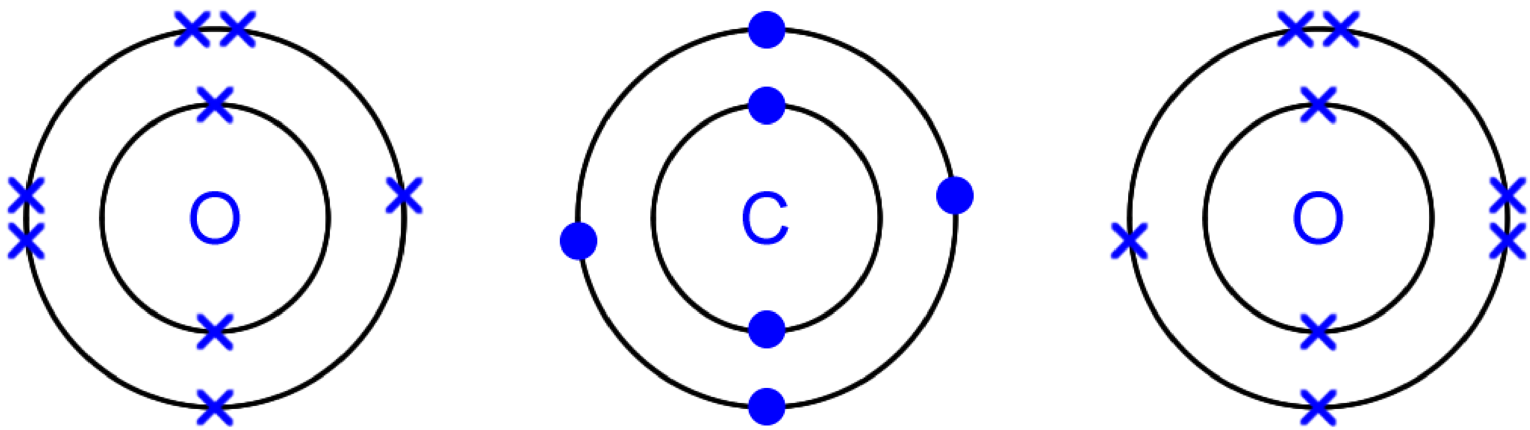
Molecule of nitrogen, formula: N_2

- The dot-and-cross diagram for molecular nitrogen that should be used as an answer in an examination.

• = Electron of nitrogen (right-hand-side). × = Electron of nitrogen (left-hand-side).

Chemical Bonding

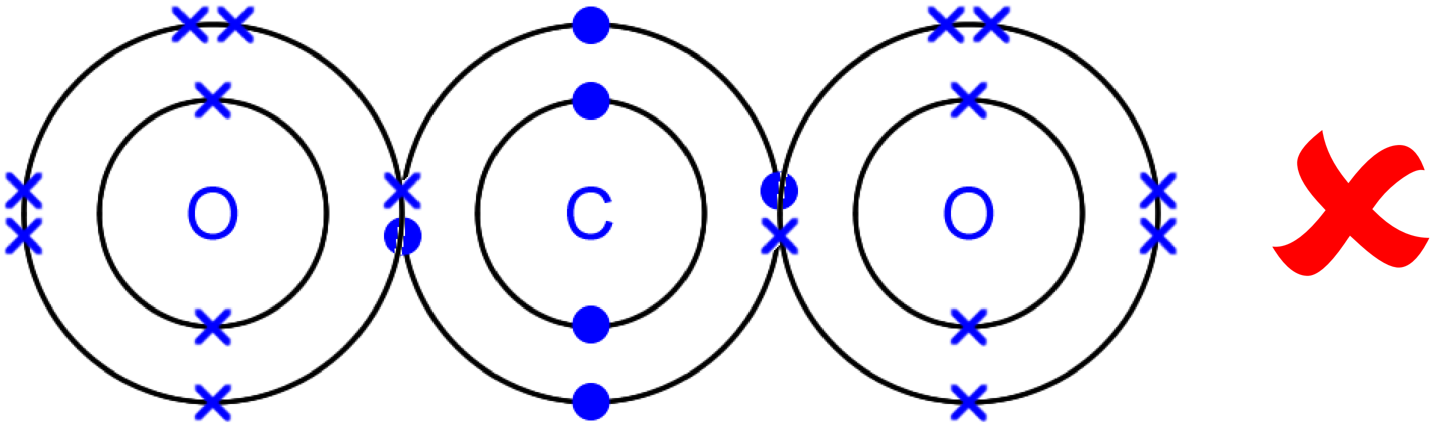
Covalent Bonding – Carbon Dioxide



● = Electron of carbon.　　× = Electron of oxygen.

Chemical Bonding

Covalent Bonding – Carbon Dioxide

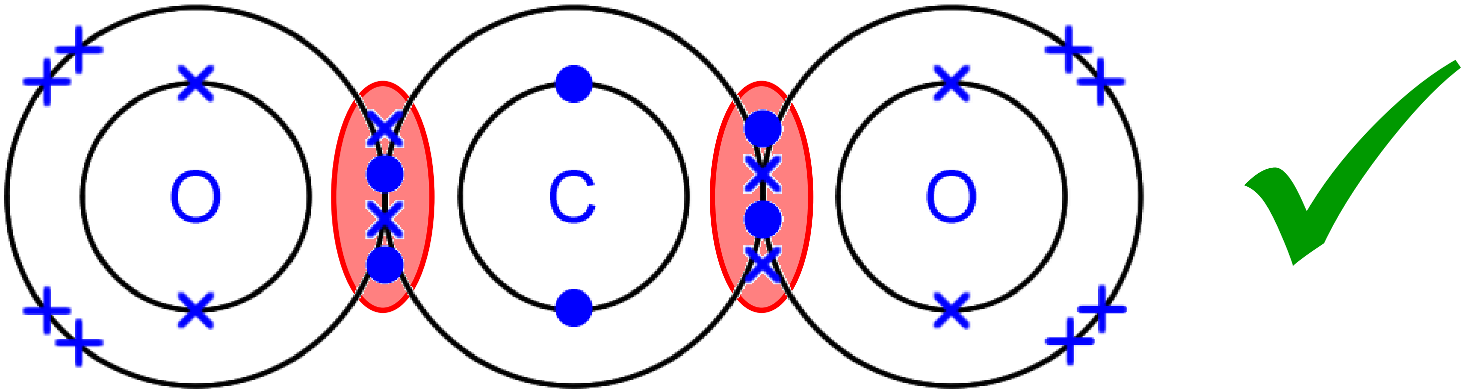


● = Electron of carbon. × = Electron of oxygen.

Chemical Bonding

Covalent Bonding – Carbon Dioxide

Molecule of carbon dioxide, formula: CO_2



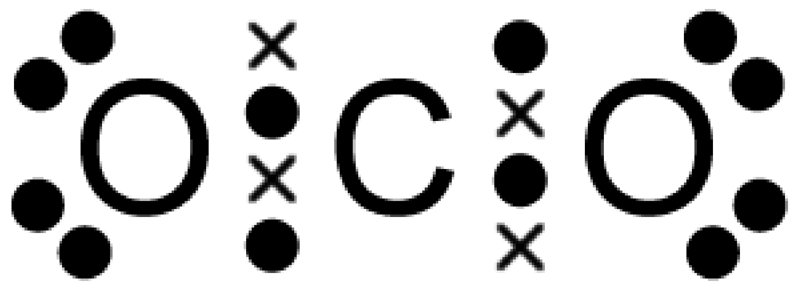
- **Note:** Two *double covalent bonds*.

• = Electron of carbon. × = Electron of oxygen.

Chemical Bonding

Covalent Bonding – Carbon Dioxide

Molecule of carbon dioxide, formula: CO_2



- The dot-and-cross diagram for carbon dioxide that should be used as an answer in an examination.

• = Electron of oxygen. × = Electron of carbon.

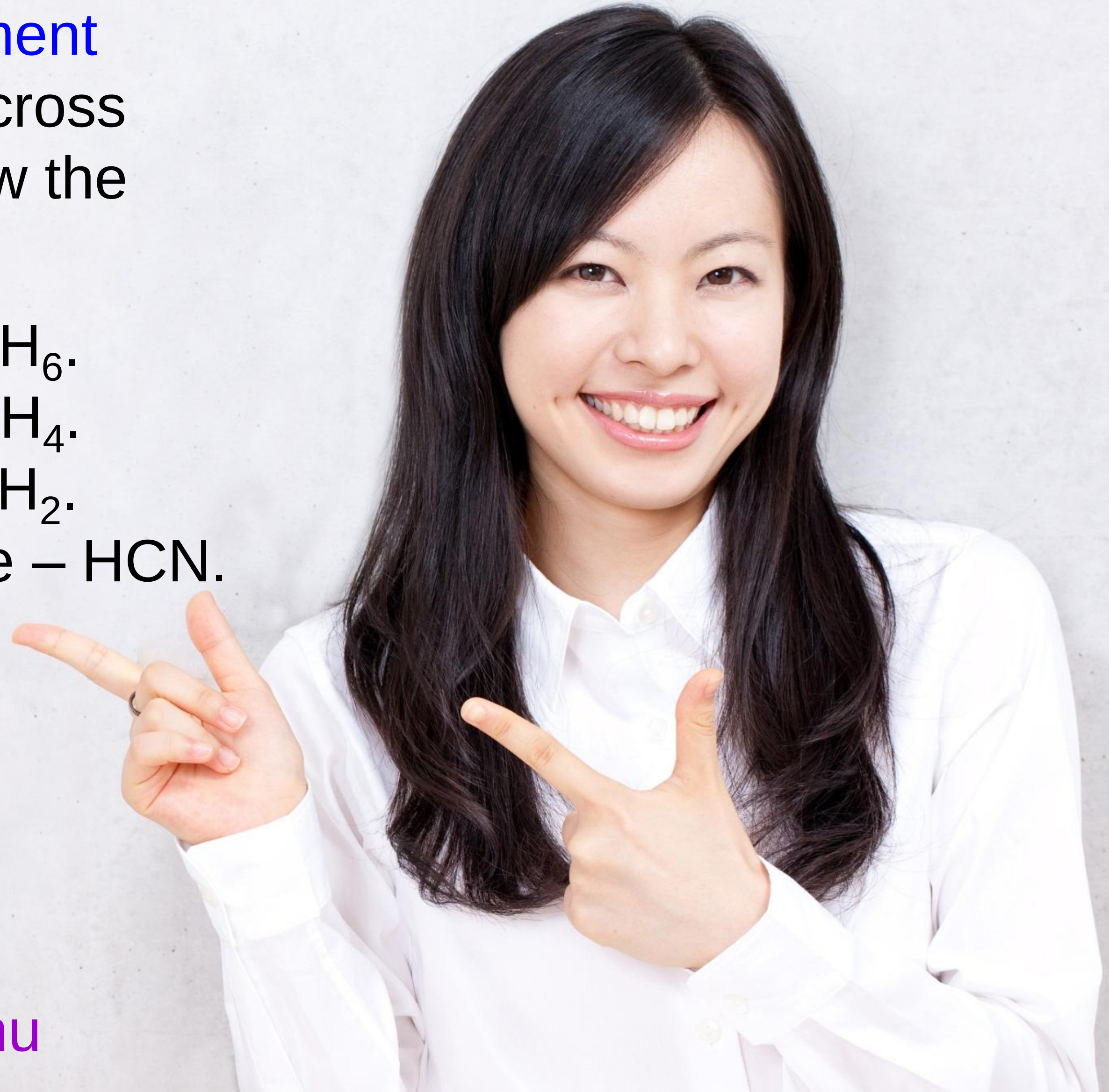
Chemical Bonding

Covalent Bonding

8. Self-assessment

- Draw dot-and-cross diagrams to show the bonding in:

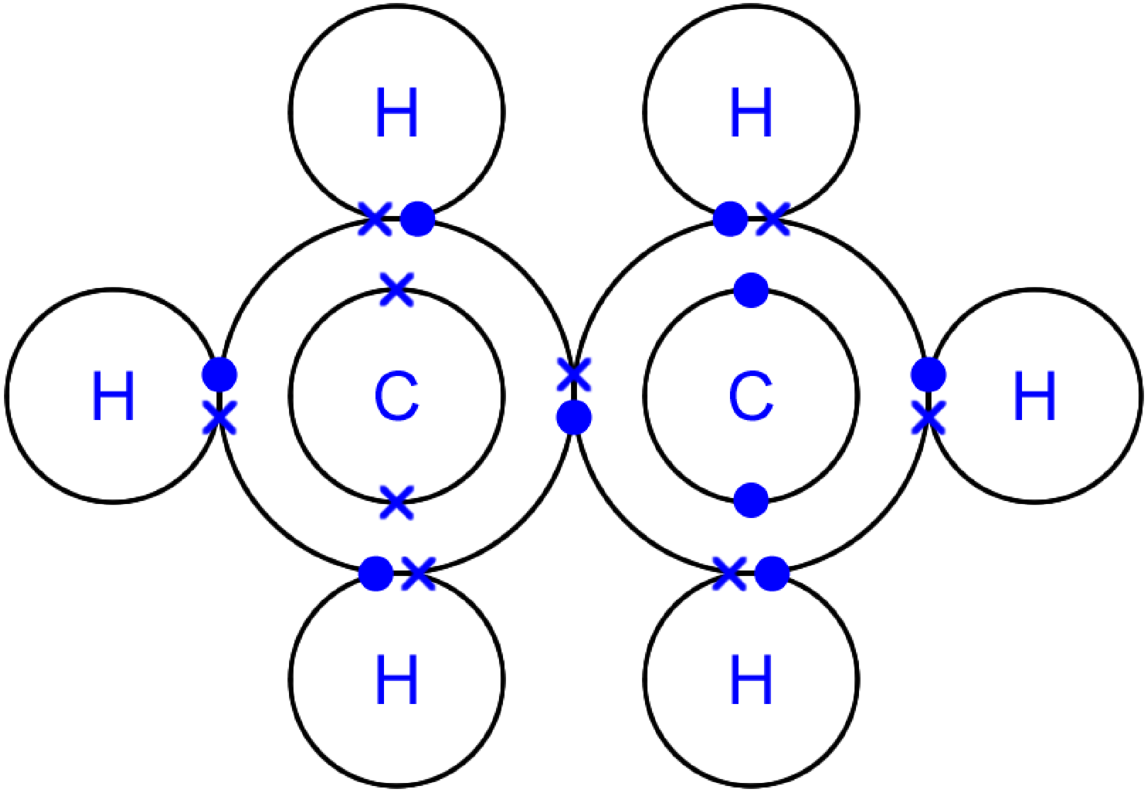
a) Ethane – C_2H_6.
b) Ethene – C_2H_4.
c) Ethyne – C_2H_2.
d) Hydrogen cyanide – HCN.



Main Menu

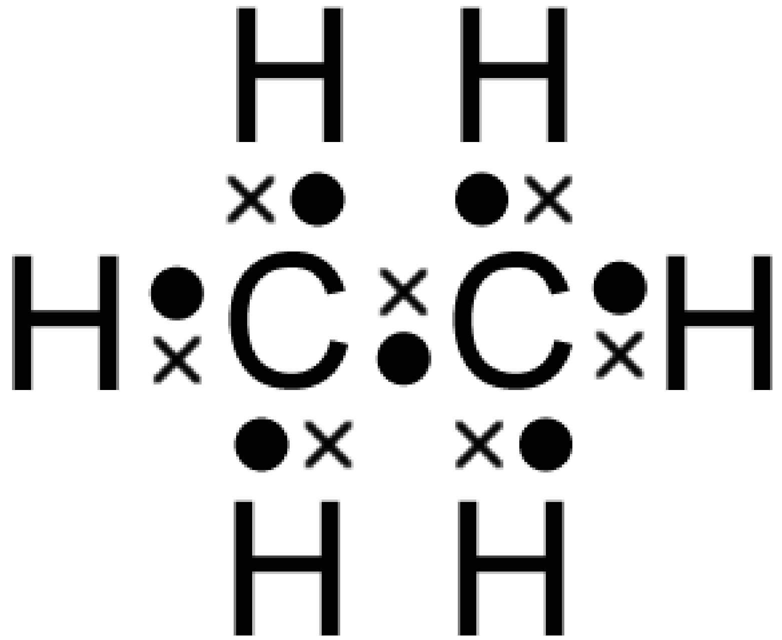
Chemical Bonding

Covalent Bonding – Ethane – C_2H_6



Chemical Bonding

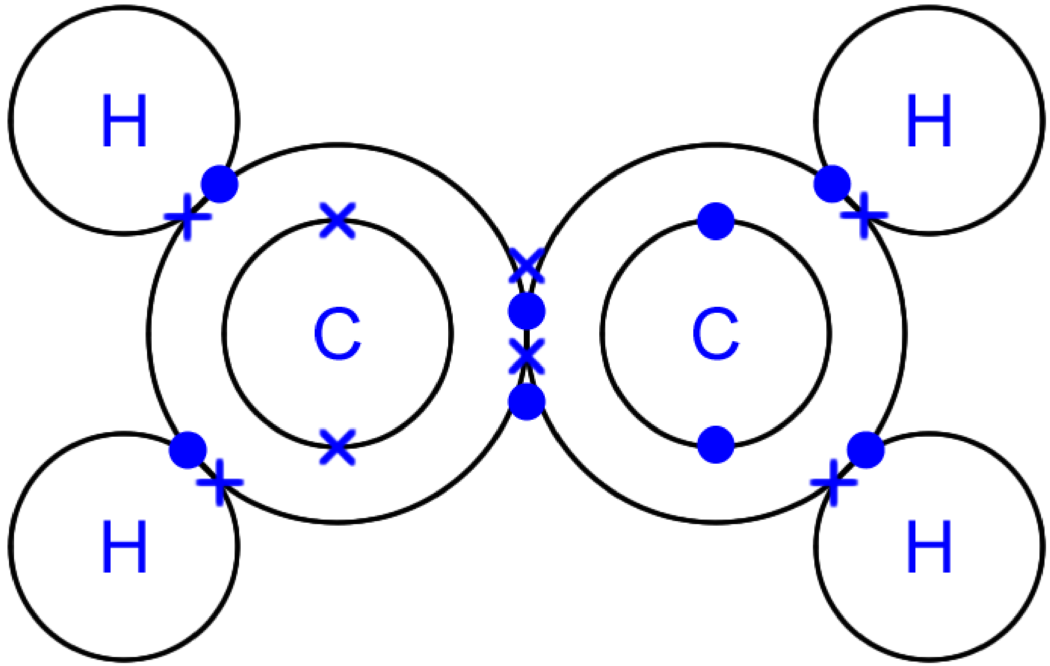
Covalent Bonding – Ethane – C_2H_6



- The dot-and-cross diagram for ethane that should be used as an answer in an examination.

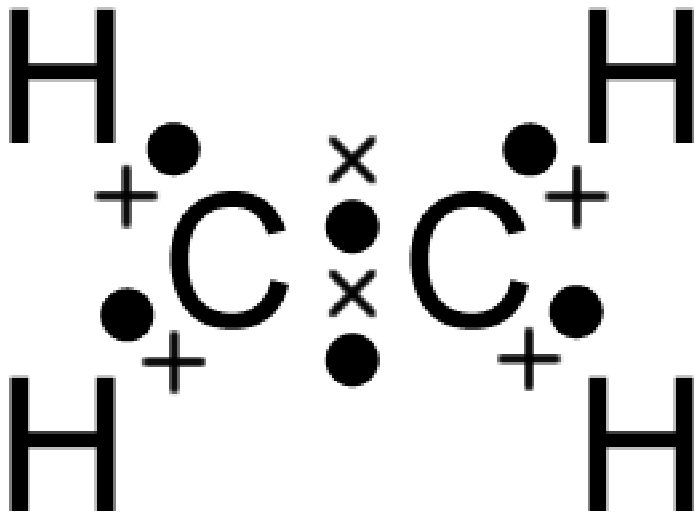
Chemical Bonding

Covalent Bonding – Ethene – C_2H_4



Chemical Bonding

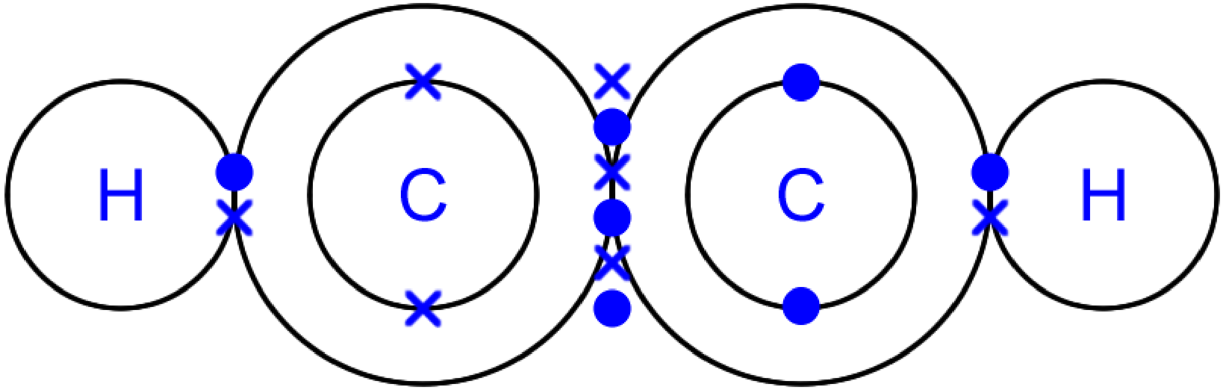
Covalent Bonding – Ethene – C_2H_4



- The dot-and-cross diagram for ethene that should be used as an answer in an examination.

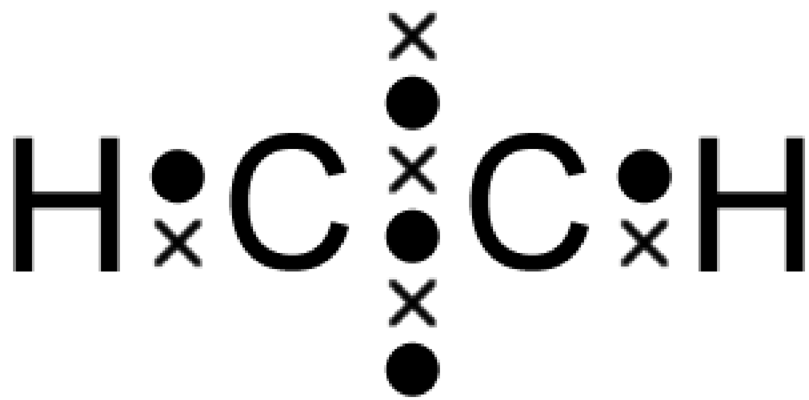
Chemical Bonding

Covalent Bonding – Ethyne – C_2H_2



Chemical Bonding

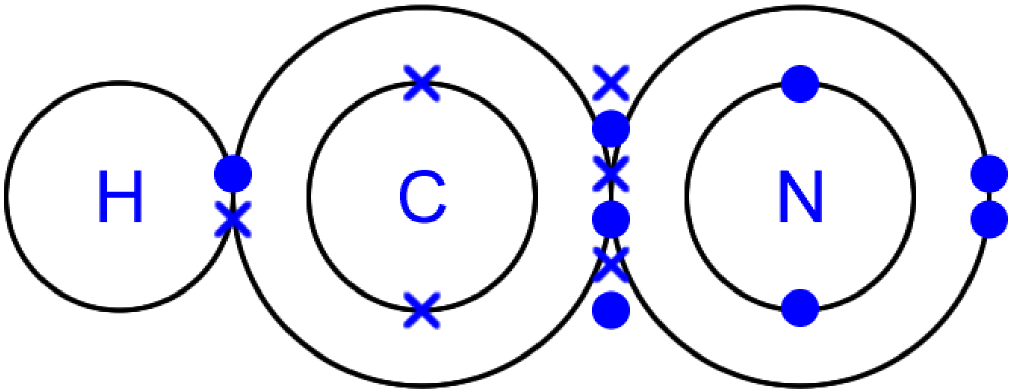
Covalent Bonding – Ethyne – C_2H_2



- The dot-and-cross diagram for ethyne that should be used as an answer in an examination.

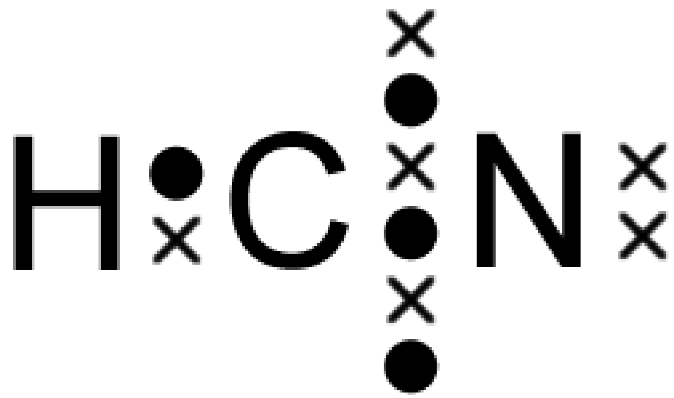
Chemical Bonding

Covalent Bonding – Hydrogen Cyanide – HCN



Chemical Bonding

Covalent Bonding – Hydrogen Cyanide – HCN



- The dot-and-cross diagram for hydrogen cyanide that should be used as an answer in an examination.

Chemical Bonding

Covalent Bonding – Summary

- Covalent bonding occurs between atoms of *non-metallic elements*.
- Atoms of the non-metallic elements *share* their valence electrons.
- Atoms of all of the non-metallic elements obtain *noble gas electronic configurations*.
- A pair of electrons that is shared equally between two atoms is referred to as a *covalent bond*.
- A covalent bond exists because the positive nuclei of two atoms are attracted towards the pair of negatively charged electrons that are equally shared between them.
- Chemicals that are composed of a relatively simple covalent structure are called *molecules*.

Chemical Bonding

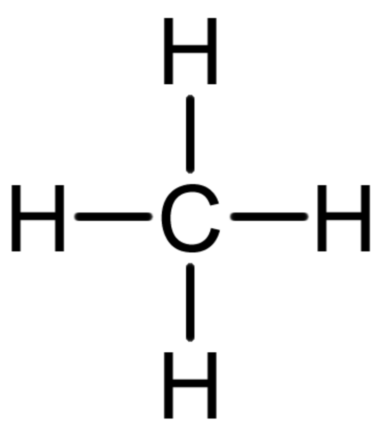
9. In addition to dot-and-cross diagrams, what other ways of *drawing molecules* are there?

Main Menu

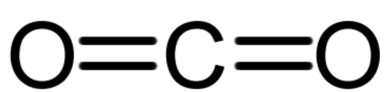
Chemical Bonding

Covalent Bonding – Structural Formulae

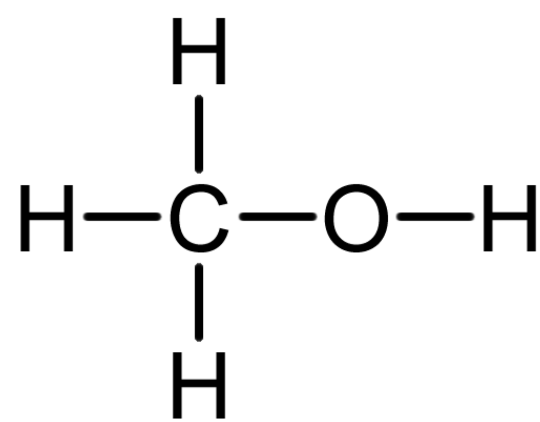
- A structural formula is a graphical representation of the arrangement of atoms in a molecule.
- Lines are used to represent covalent bonds, *i.e.* shared pairs of electrons (dots and crosses).
- The shape of the molecule is represented by the angles between the lines / bonds.



Methane



Carbon Dioxide

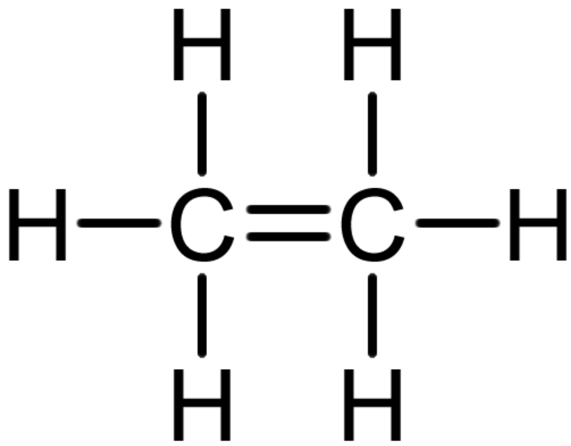


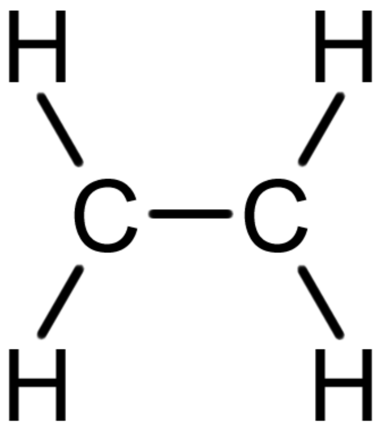
Methanol

Chemical Bonding

Covalent Bonding – Structural Formulae

- **Question:** What is wrong with the two structural formulae given below?

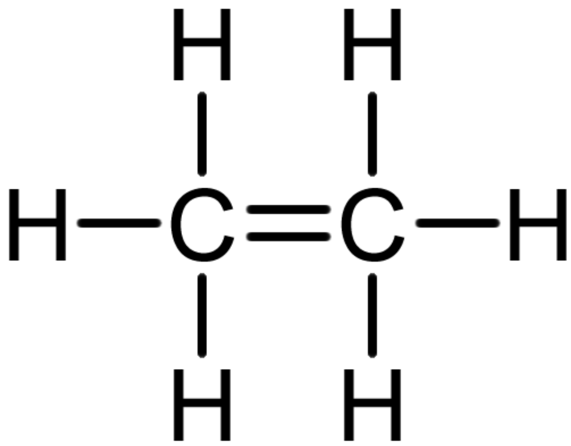


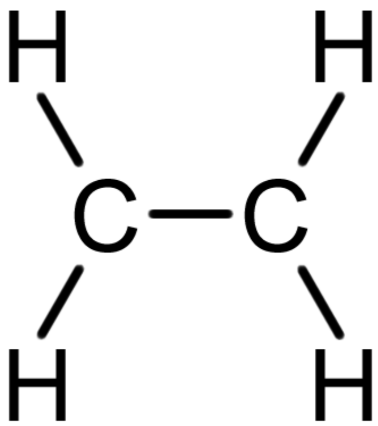


Chemical Bonding

Covalent Bonding – Structural Formulae

- **Question:** What is wrong with the two structural formulae given below?





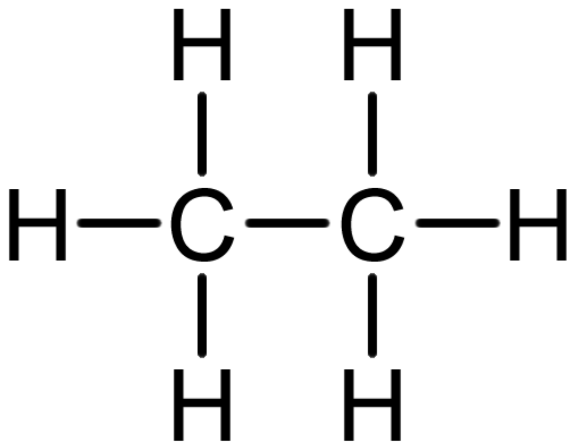
- Each carbon atom is making *five* covalent bonds. Carbon has a valency of *four* and should make *four* covalent bonds.

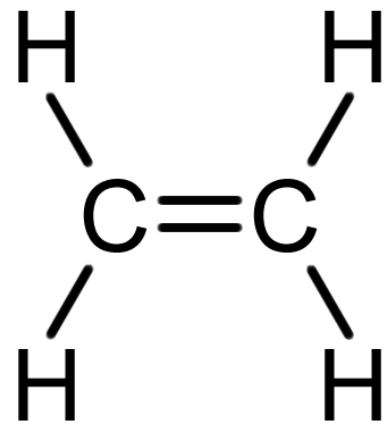
- Each carbon atom is making *three* covalent bonds. Carbon has a valency of *four* and should make *four* covalent bonds.

Chemical Bonding

Covalent Bonding – Structural Formulae

- The *correct* structural formulae are given below:





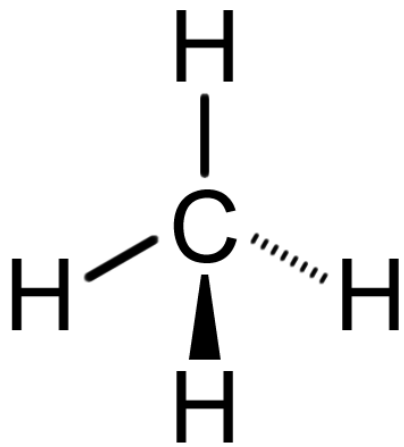
- The *correct* structural formula for ethane, C_2H_6. All of the carbon atoms are *tetravalent*, *i.e.* making four covalent bonds.

- The *correct* structural formula for ethene, C_2H_4. All of the carbon atoms are *tetravalent*, *i.e.* making four covalent bonds.

Chemical Bonding

Covalent Bonding – Structural Formulae

- More information can be given about the shape or geometry of a molecule by combining *solid straight lines* with *triangular lines* and *dotted lines*.



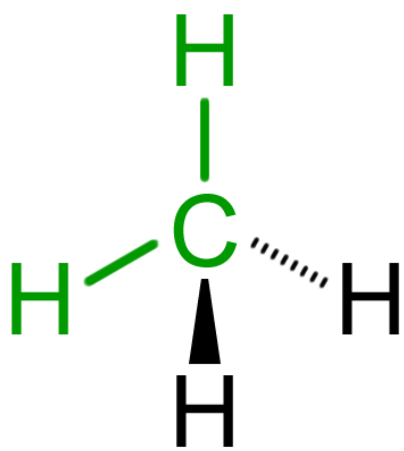
Methane

Chemical Bonding

Covalent Bonding – Structural Formulae

- More information can be given about the shape or geometry of a molecule by combining *solid straight lines* with *triangular lines* and *dotted lines*.

→ *Solid straight lines* represent a bond in the plane of the paper.



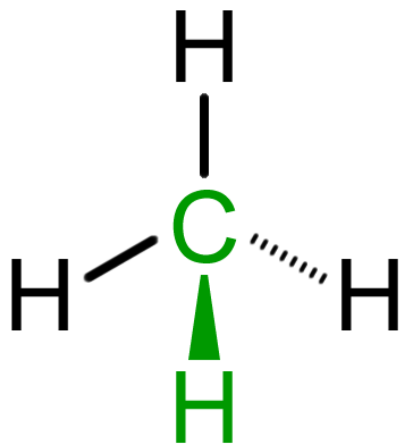
Methane

Chemical Bonding

Covalent Bonding – Structural Formulae

- More information can be given about the shape or geometry of a molecule by combining *solid straight lines* with *triangular lines* and *dotted lines*.

→ *Triangular lines* represent a bond emerging from the paper towards the observer.



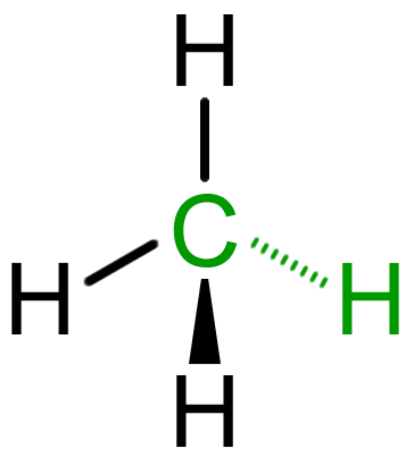
Methane

Chemical Bonding

Covalent Bonding – Structural Formulae

- More information can be given about the shape or geometry of a molecule by combining *solid straight lines* with *triangular lines* and *dotted lines*.

→ *Dotted lines* represent a bond receding into the paper away from the observer.

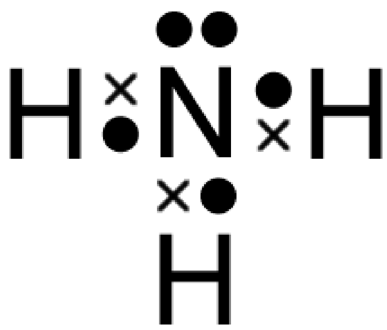


Methane

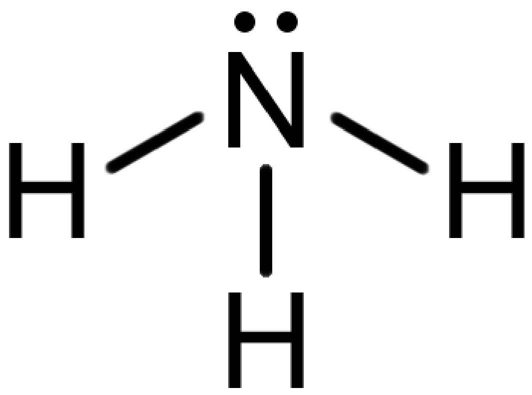
Chemical Bonding

Covalent Bonding – Lewis Dot Diagrams

- Lewis dot diagrams are structural formulae that use a pair of dots (• •) to represent a pair of electrons in the valence shell of an atom that are *not* involved in bonding. Such electrons are called *lone-pair* electrons or *non-bonding* electrons.



The Dot-and-cross Diagram
for Ammonia



The Lewis Dot Diagram
for Ammonia

Chemical Bonding

Covalent Bonding – Lewis Dot Diagrams

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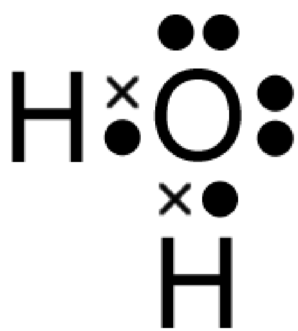
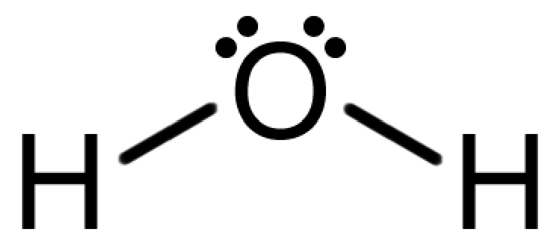


The Dot-and-cross Diagram for Water



The Lewis Dot Diagram for Water

Chemical Bonding

Covalent Bonding – Lewis Dot Diagrams

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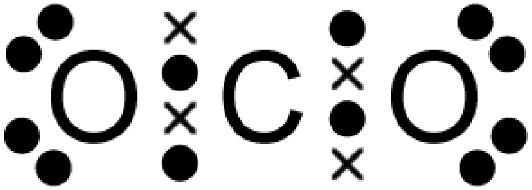
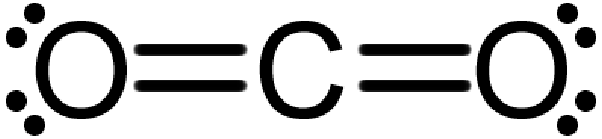


The Dot-and-cross Diagram
for Carbon Dioxide



The Lewis Dot Diagram
for Carbon Dioxide

Chemical Bonding

Covalent Bonding – Lewis Dot Diagrams

- **Question:** For the compound methylamine, formula CH_3NH_2, draw:

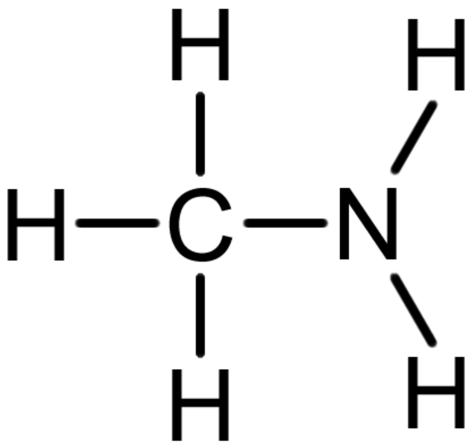
a) the structural formula.

b) the Lewis dot diagram.

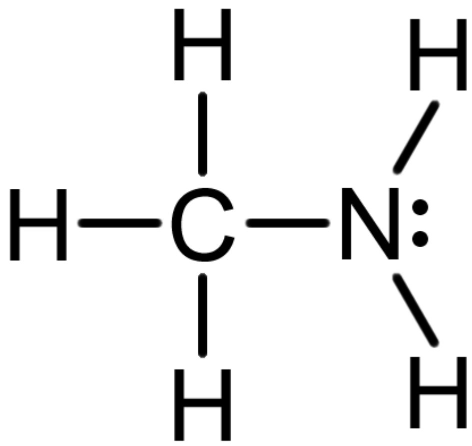
Chemical Bonding

Covalent Bonding – Lewis Dot Diagrams

- **Question:** For the compound methylamine, formula CH_3NH_2, draw:
  a) the structural formula.
  b) the Lewis dot diagram.



a) Structural formula of methylamine.



b) Lewis dot diagram of methylamine.

Chemical Bonding

10. Which chemical elements are *diatomic*?

Main Menu

Chemical Bonding

Diatomic Elements:

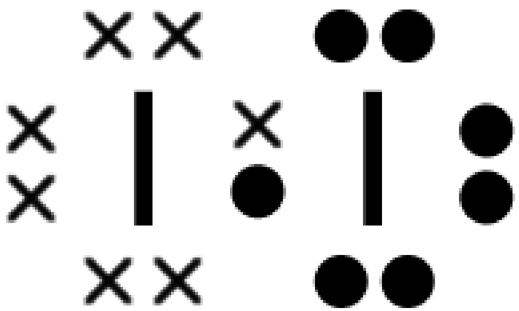
I Have No Bright Or Clever Friends

Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever Friends

Iodine – I_2

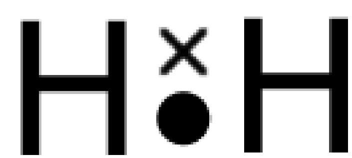


Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever Friends

Hydrogen – H_2



Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever Friends

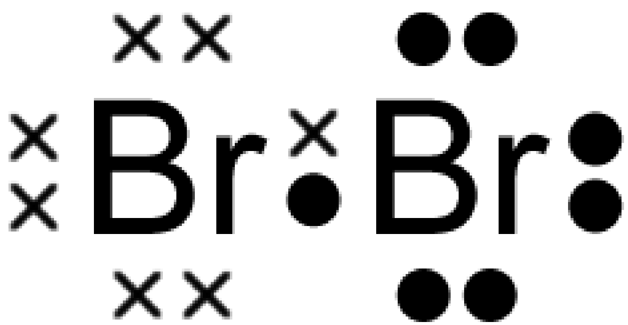
Nitrogen – N_2

Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever Friends

Bromine – Br_2

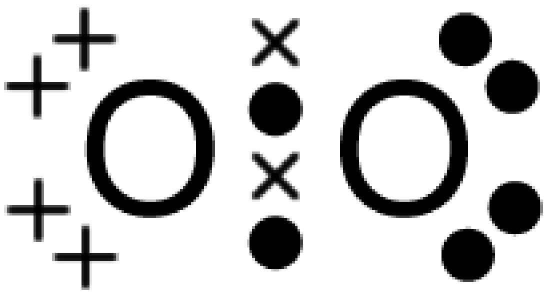


Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever Friends

Oxygen – O_2

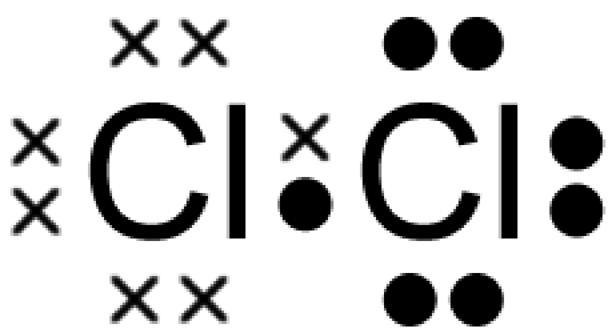


Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever Friends

Chlorine – Cl_2

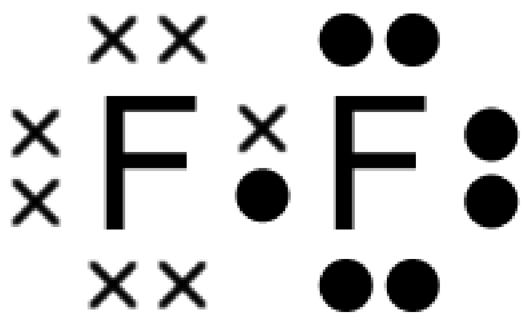


Chemical Bonding

Diatomic Elements:

I Have No Bright Or Clever **F**riends

Fluorine – F_2



Chemical Bonding

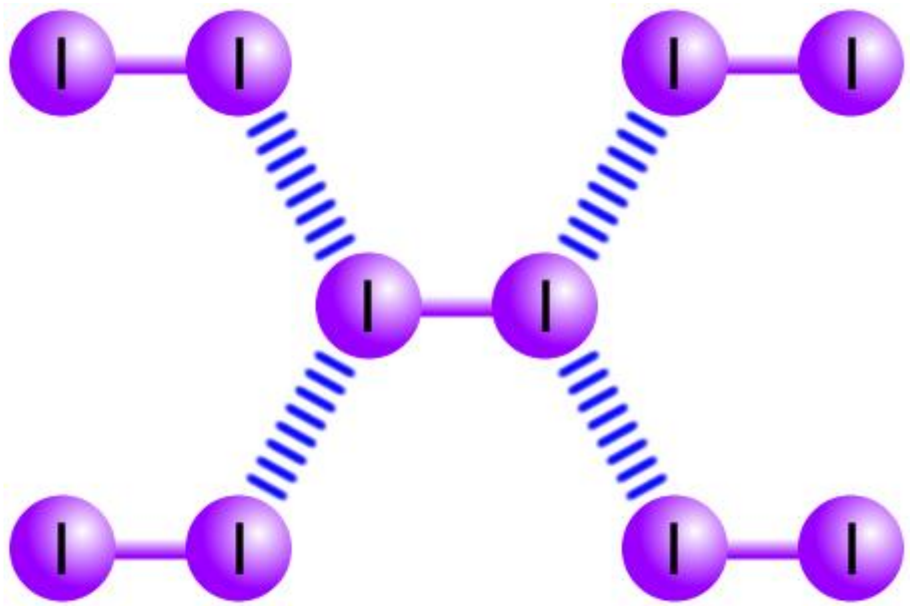
Covalent Bonding – Simple Molecular – Properties

11. What are the *general properties* of chemicals with simple molecular structures?

Main Menu

Chemical Bonding

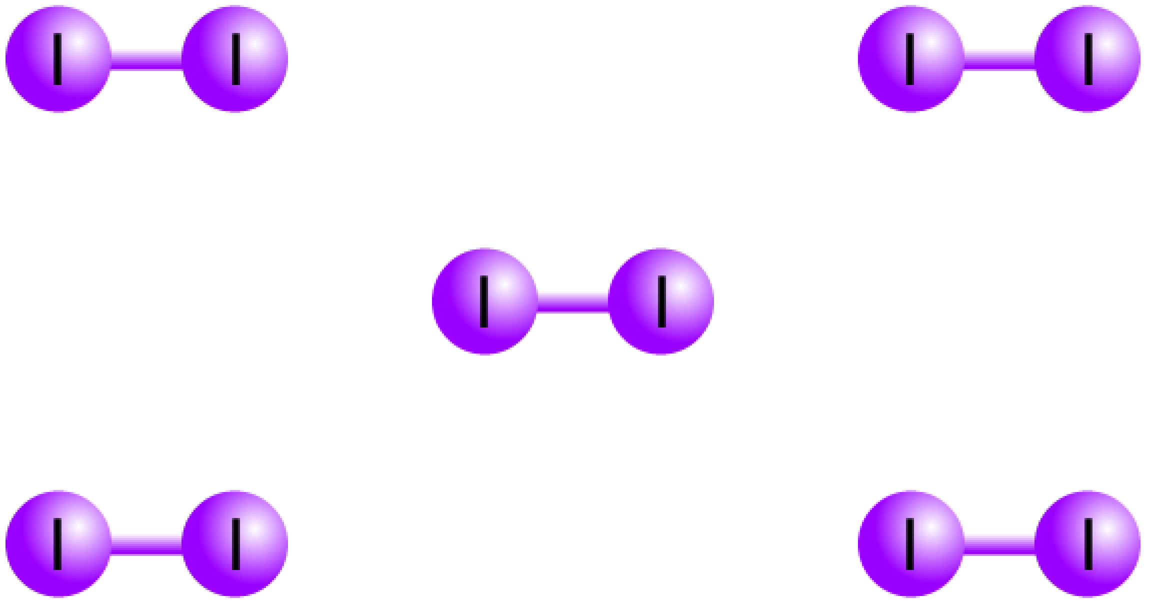
Covalent Bonding – Simple Molecular – Properties



- Chemicals with a simple molecular structure, for example iodine, have relatively *low* melting points and boiling points.

Chemical Bonding

Covalent Bonding – Simple Molecular – Properties



- The low melting points and boiling points are *not* because the covalent bonds within the molecules are weak but because the *forces of attraction between the molecules* (*intermolecular forces / van der Waals forces**) *are weak*.

*__Note:__ There are different types of intermolecular forces of attraction.

Chemical Bonding

Covalent Bonding – Simple Molecular – Properties

- Chemicals with a simple molecular structure have relatively *low melting points and boiling points* because the intermolecular force of attraction (or van der Waals force of attraction) between the molecules is weak and therefore only requires a small amount of thermal energy to weaken (for melting) or overcome (for boiling).

- Chemicals with a simple molecular structure (except those that are crystalline, *e.g.* iodine) tend to be *soft*.

Chemical Bonding

Covalent Bonding – Simple Molecular – Properties

- Chemicals with a simple molecular structure are generally *insoluble in polar solvents such as water*. They are, however, generally *soluble in non-polar solvents such as oil or hexane*.

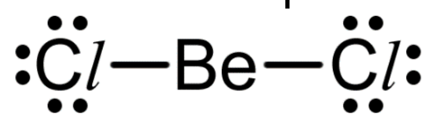
- Chemicals with a simple molecular structure are *electrical insulators*, *i.e.* they do not conduct electricity in any form (solid, liquid or solution) because *they do not contain any mobile electrons or mobile ions* (*no mobile charge carrying particles*).

Chemical Bonding

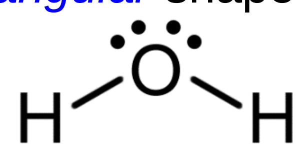
Valence Shell Electron Pair Repulsion Theory

12. What are the different *shapes* of simple covalent molecules?

- Why does the $BeCl_2$ molecule have a *linear* shape?



- But the H_2O molecule has an *angular* shape?



Main Menu

Chemical Bonding

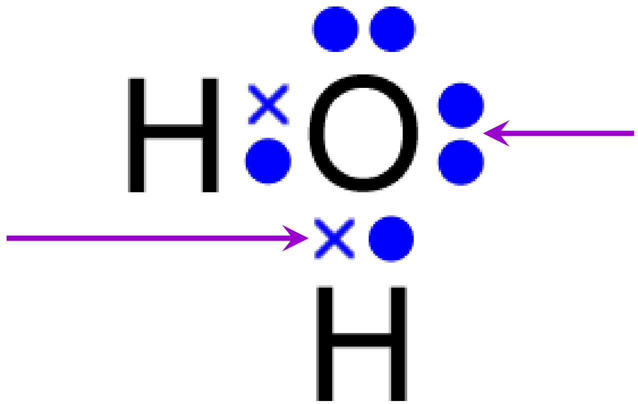
Valence Shell Electron Pair Repulsion Theory

- *Valence shell electron pair repulsion theory* (*VSEPRT*) states that pairs of electrons (both *bonding pairs* of electrons and *non-bonding* or *lone pairs* of electrons) in the valence shell of an atom *repel* each other and *arrange* themselves to be as *far away* from each other as possible in *three-dimensional space*. This theory can be used to both *explain* and *predict* the *shapes of molecules*.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- Bonding pair of electrons in H_2O.



- Non-bonding or lone pair of electrons in H_2O.

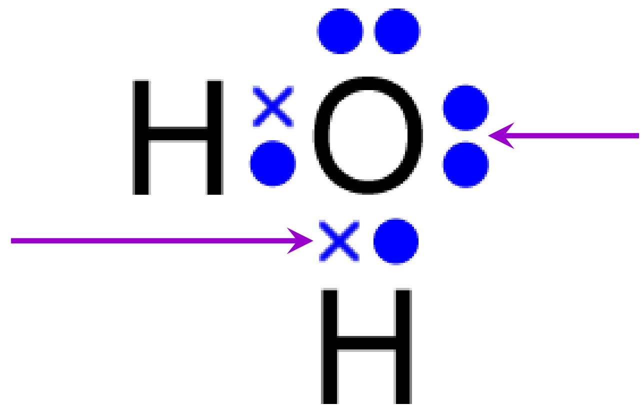
- Non-bonding or lone pairs of electrons *influence* the *shape* of a molecule, but are *ignored* when *describing* the *shape* of the molecule.

- Only the arrangement of the *bonding pairs* of electrons are referred to when *describing* the *shape* of a molecule.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- Bonding pair of electrons in H_2O.



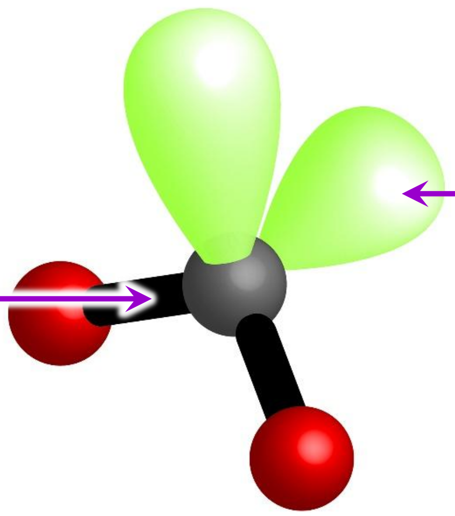
- Non-bonding or lone pair of electrons in H_2O.

- For example, consider a single molecule of *water*. There are *four pairs of electrons* in the valence shell of the central oxygen atom, *two bonding pairs of electrons* and *two non-bonding* or *lone pairs of electrons*.

Chemical Bonding

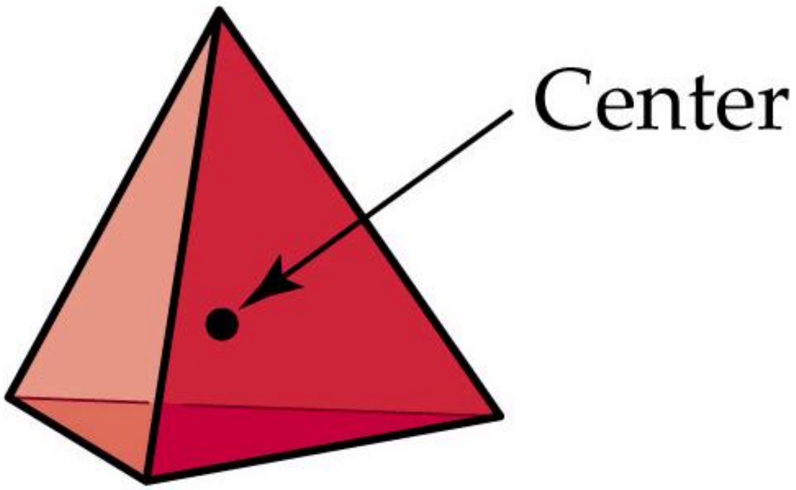
Valence Shell Electron Pair Repulsion Theory

- Bonding pair of electrons in H_2O.



- Non-bonding or lone pair of electrons in H_2O.

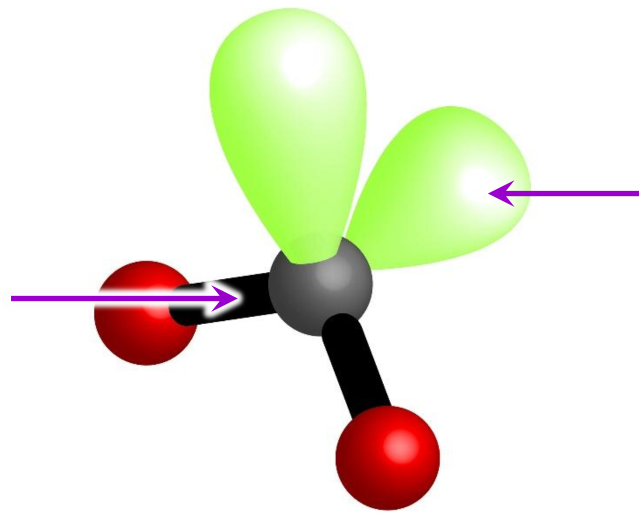
- The four pairs of negatively charged electrons repel each other, taking-up a three dimensional *tetrahedral* arrangement. A *tetrahedron* can be visualised as a triangular pyramid.



Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- Bonding pair of electrons in H_2O.

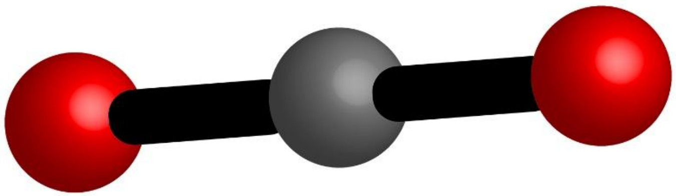


- Non-bonding or lone pair of electrons in H_2O.

- When considering the actual shape of the water molecule, only the *oxygen atom* and *two hydrogen atoms*, and the *bonding pairs of electrons* in-between them, are taken into consideration. The shape of the water molecule is described as *angular*.

Chemical Bonding

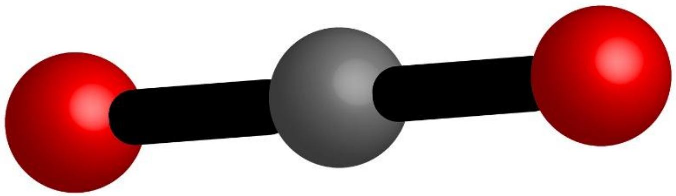
Valence Shell Electron Pair Repulsion Theory

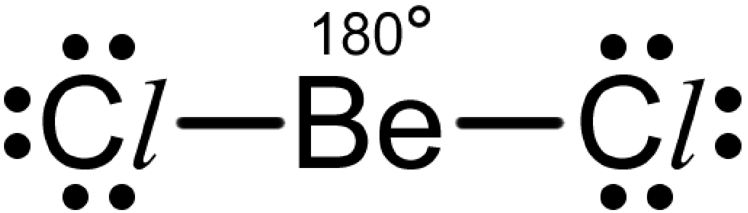


• In beryllium chloride, $BeCl_2$, there are *two bonding pairs of electrons* and *zero lone pairs of electrons* in the valence shell of the central beryllium atom. *Electrostatic forces of repulsion* between the negatively charged pairs of electrons results in beryllium chloride having a *linear* shape.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

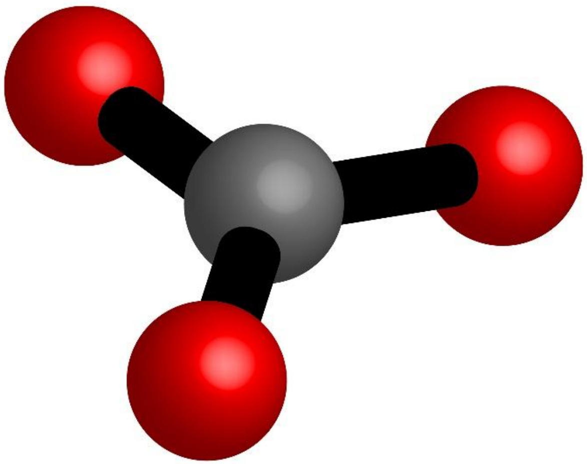


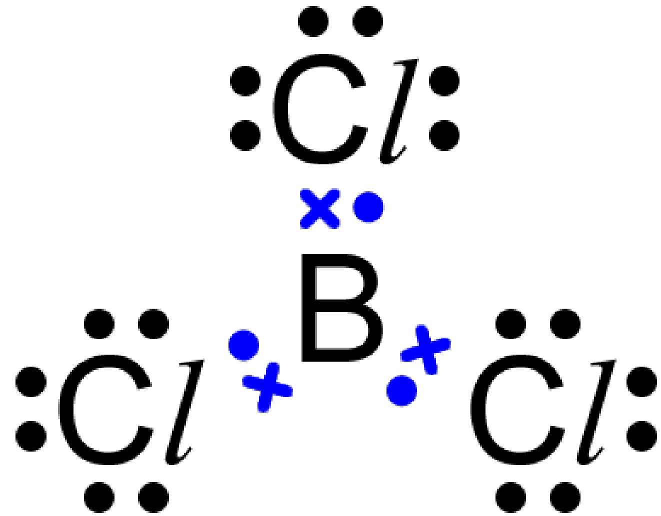


- In beryllium chloride, $BeCl_2$, there are *two bonding pairs of electrons* and *zero lone pairs of electrons* in the valence shell of the central beryllium atom. *Electrostatic forces of repulsion* between the negatively charged pairs of electrons results in beryllium chloride having a *linear* shape.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

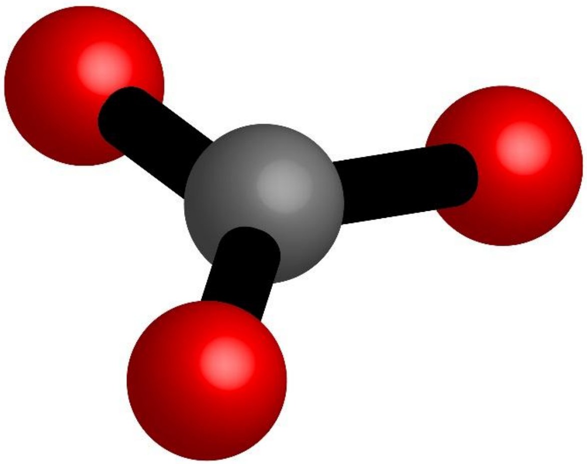


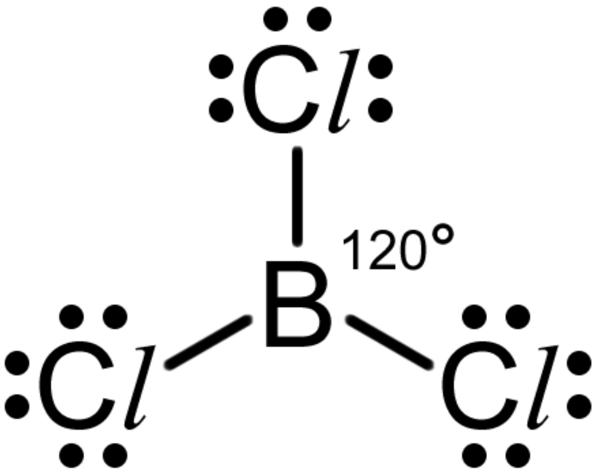


- In boron trichloride, BCl_3, there are *three bonding pairs of electrons* and *zero lone pairs of electrons* in the valence shell of the central boron atom. *Electrostatic forces of repulsion* between the negatively charged pairs of electrons results in boron trichloride having a *trigonal planar* shape.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory





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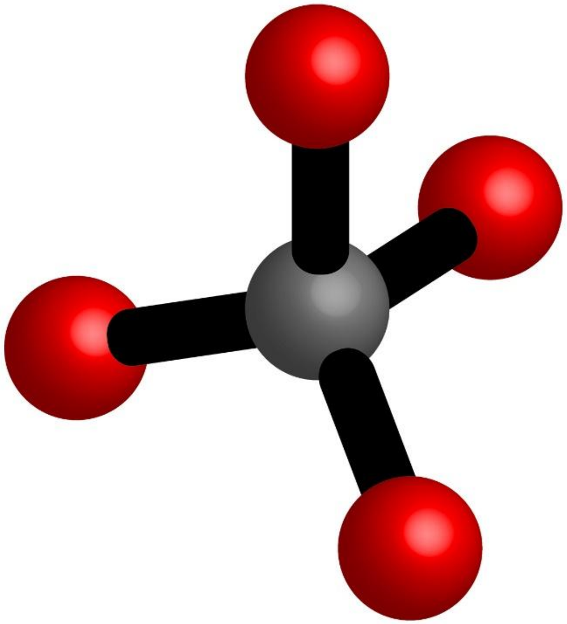
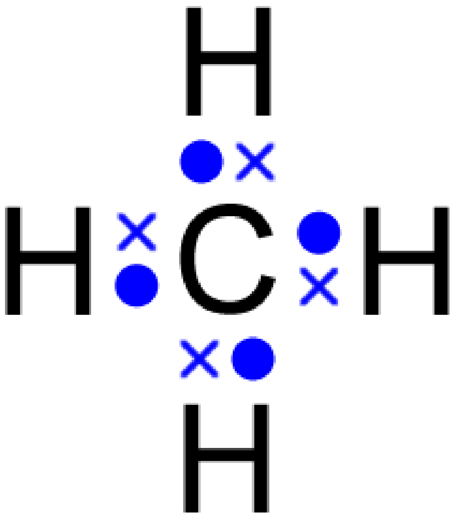
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

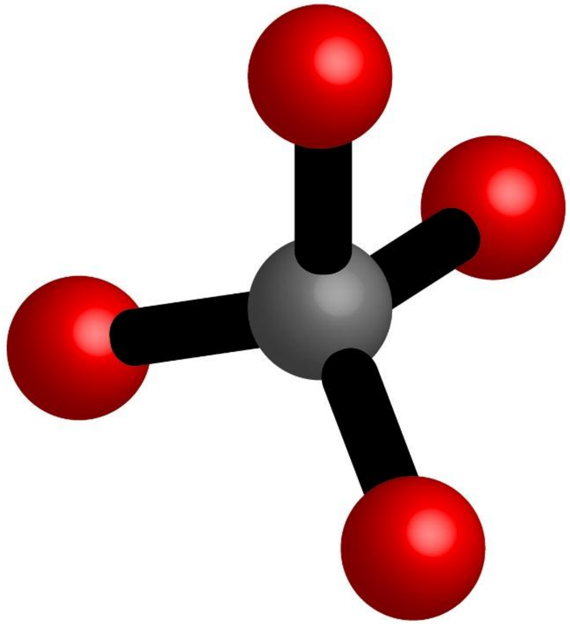


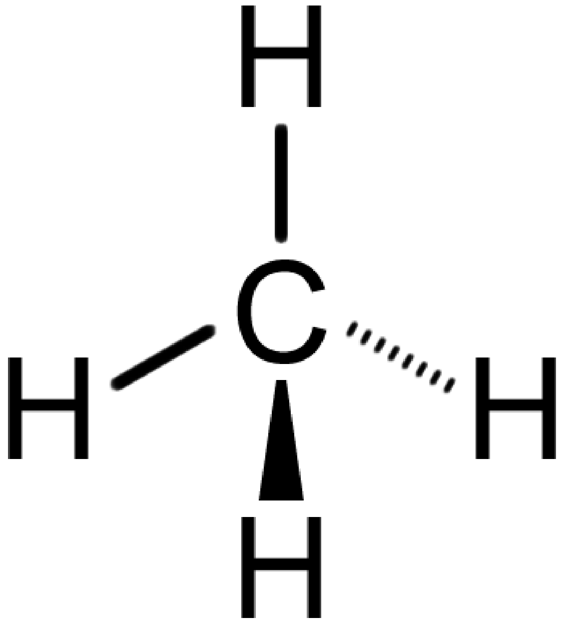


• In methane, CH_4, there are *four bonding pairs of electrons* and *zero lone pairs of electrons* in the valence shell of the central carbon atom. *Electrostatic forces of repulsion* between the negatively charged pairs of electrons results in methane having a *tetrahedral* shape.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory





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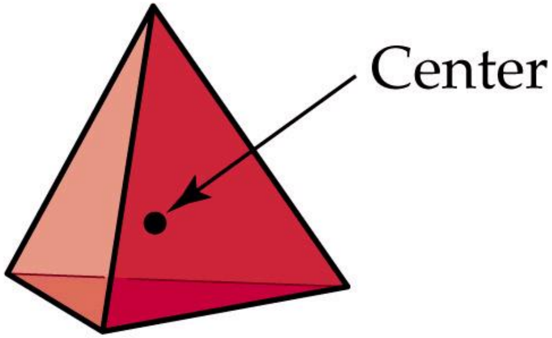
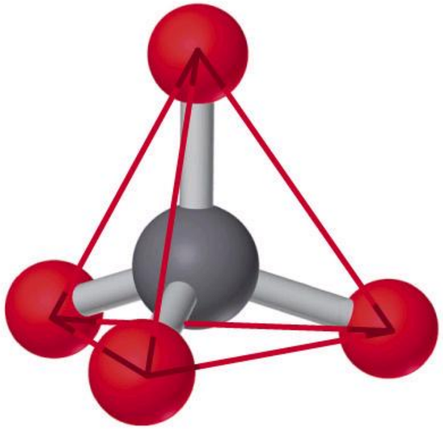
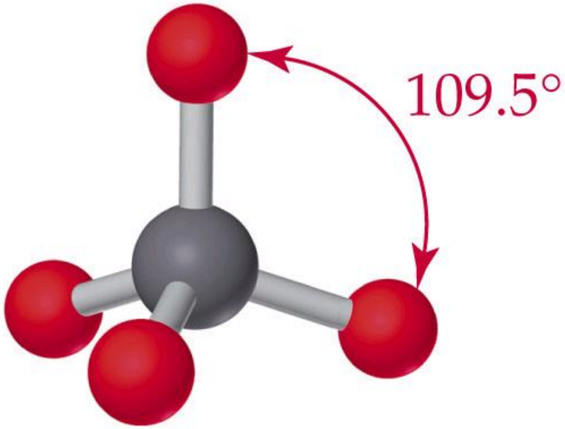
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory



A regular tetrahedron



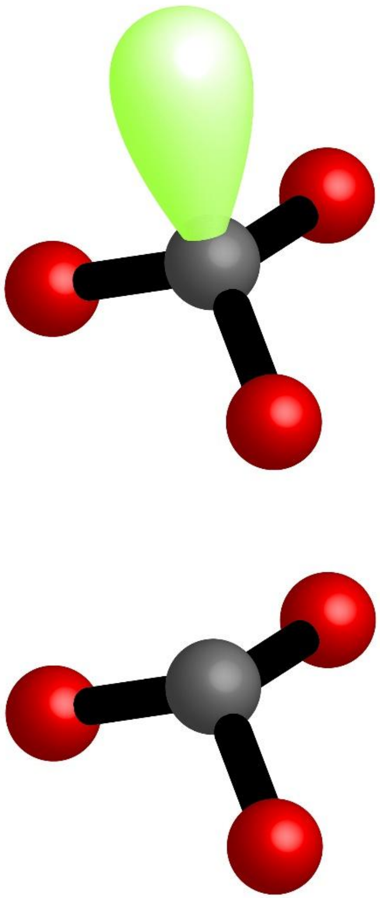


A tetrahedral molecule

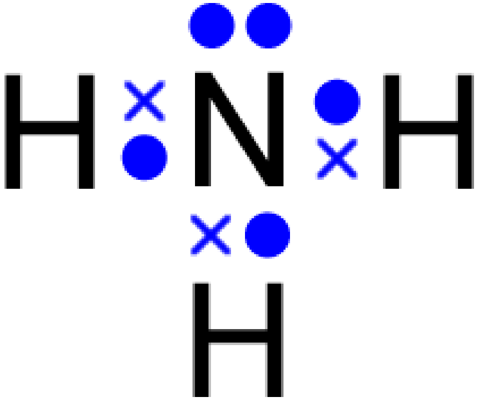
- Tetrahedral: A simple covalent molecule in which the central atom has *four bonding pairs* of electrons *only* in its valence shell. Example – methane, CH_4.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory



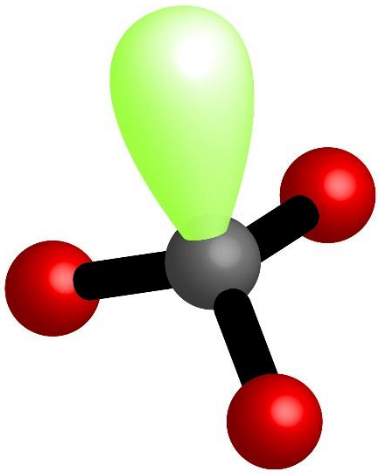
trigonal pyramidal

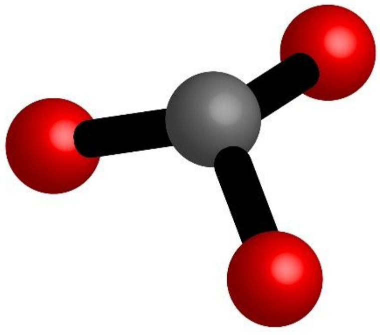


- In ammonia, NH_3, there are *four pairs* of electrons in the valence shell of the central nitrogen atom, *three bonding pairs* and *one non-bonding / lone pair*. *Electrostatic forces of repulsion* between the four negatively charged pairs of electrons results in a *tetrahedral* arrangement, *with the three bonding pair of electrons taking-up a trigonal pyramidal arrangement*.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory





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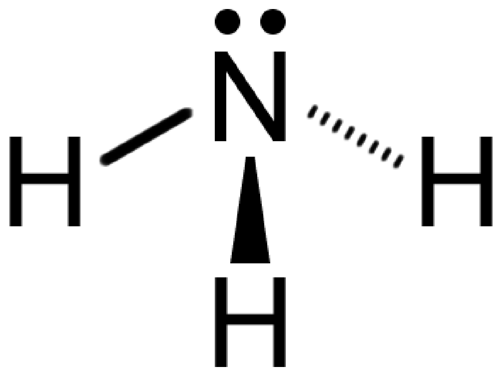

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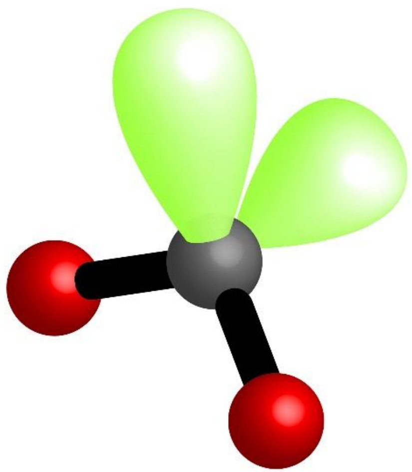
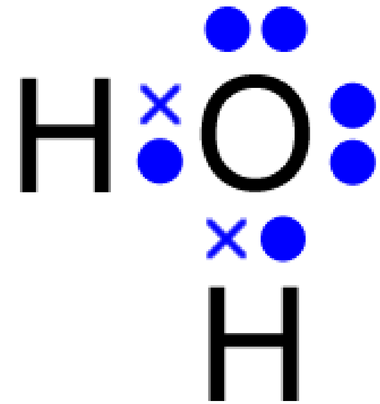
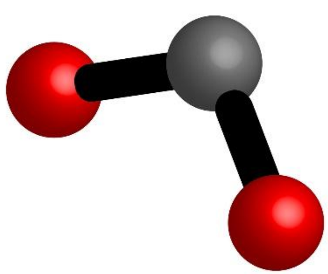
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory





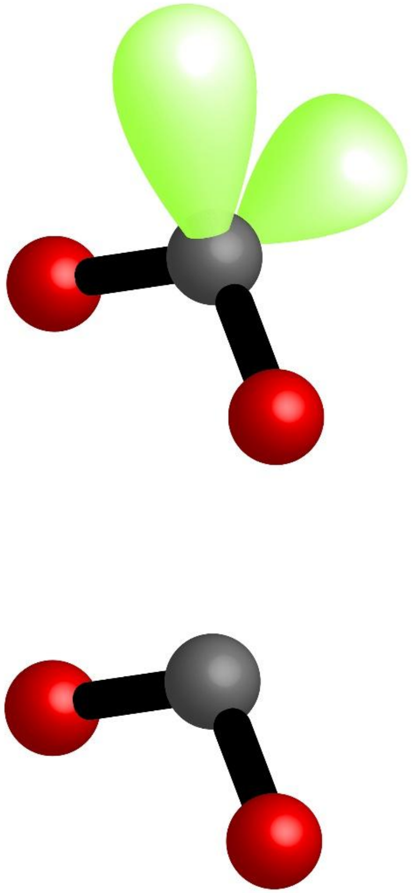
- In water, H_2O, there are *four pairs* of electrons in the valence shell of the central oxygen atom, *two bonding pairs* and *two non-bonding / lone pairs*. *Electrostatic forces of repulsion* between the four negatively charged pairs of electrons results in a *tetrahedral* arrangement, *with the two bonding pairs of electrons taking-up an angled / angular arrangement*.



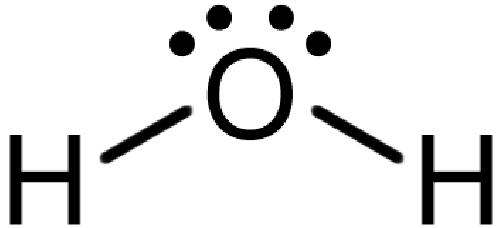
angled

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory



angled



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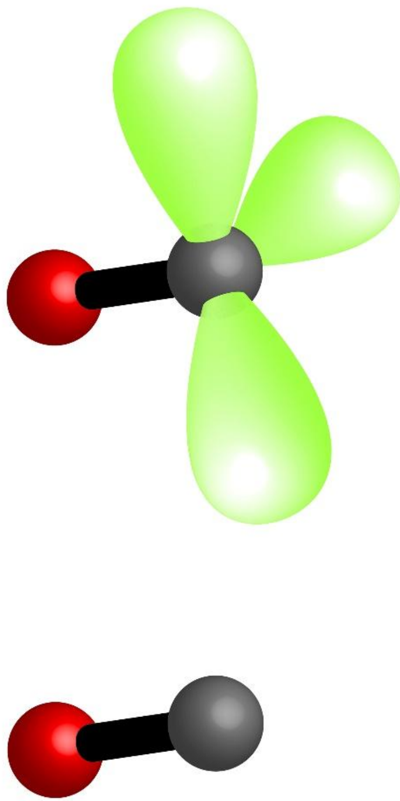
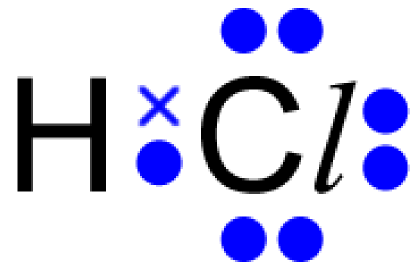
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory



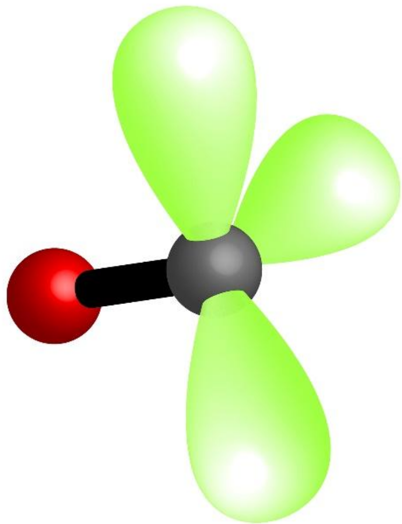
linear

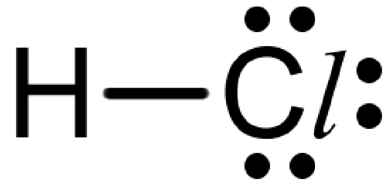


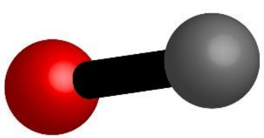
- In hydrogen chloride, HC*l*, there are *four pairs* of electrons in the valence shell of the central chlorine atom, *one bonding pair* and *three non-bonding / lone pairs*. *Electrostatic forces of repulsion* between the four negatively charged pairs of electrons results in a *tetrahedral* arrangement, *with the single pair of bonding electrons taking-up a linear arrangement*.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory







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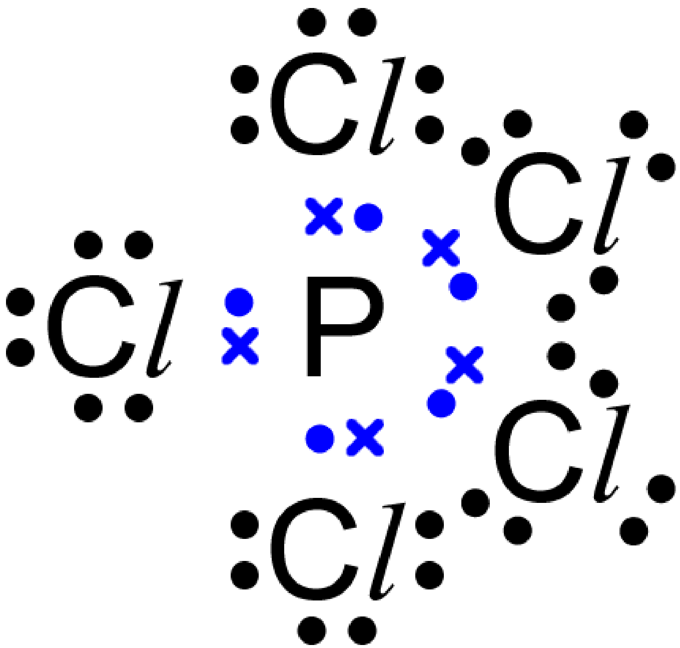
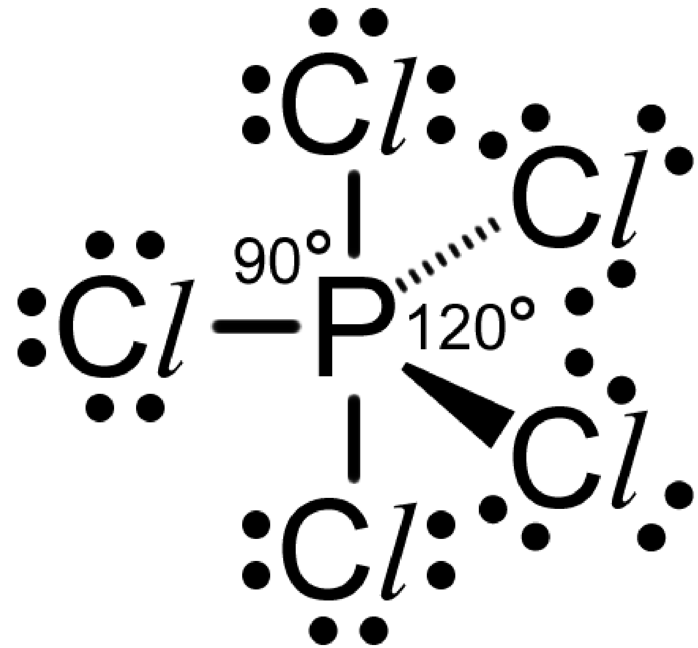
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

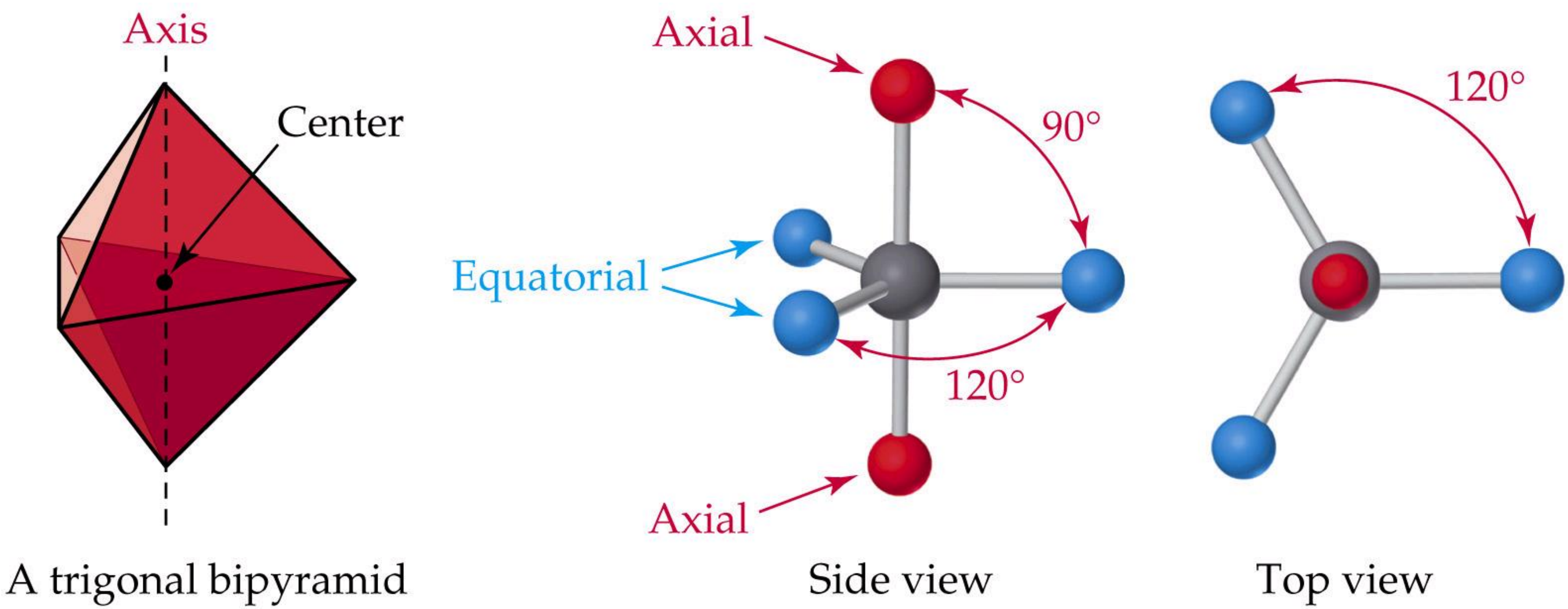




- Trigonal Bipyramidal: A simple covalent molecule in which the central atom has *five bonding pairs* of electrons *only* in its valence shell. Example – phosphorus(V) chloride, PCl_5.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory



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Chemical Bonding

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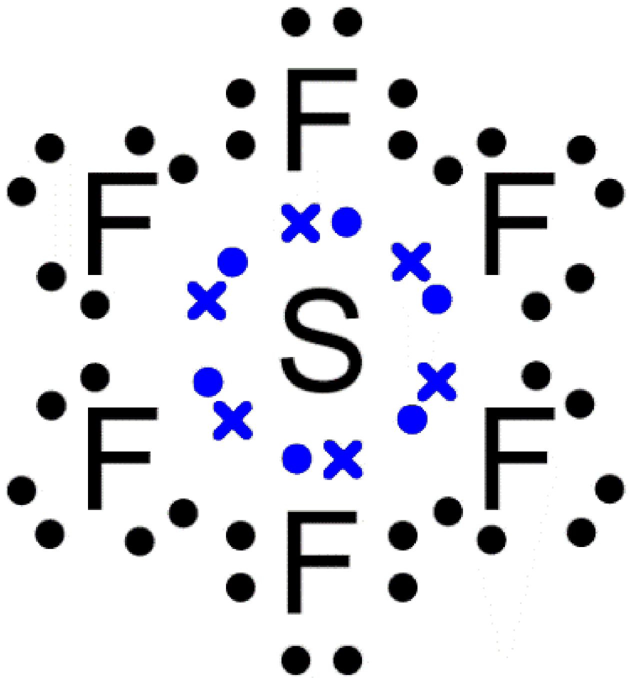
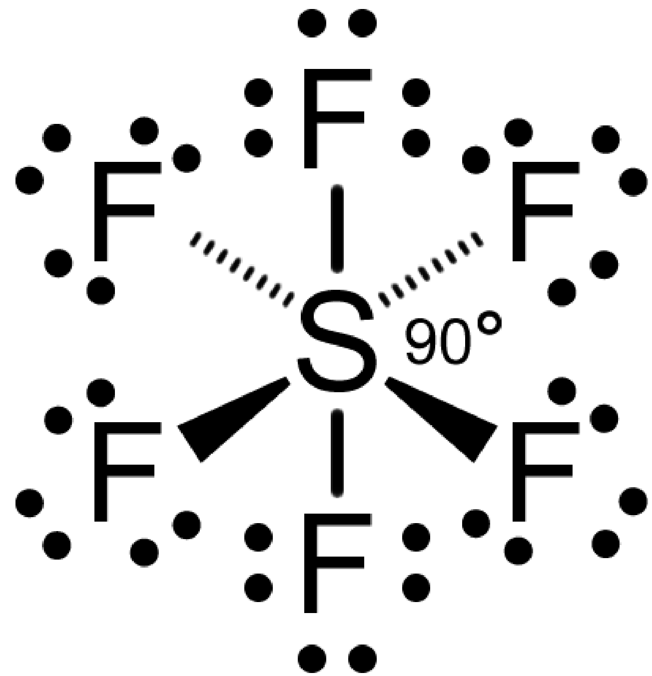
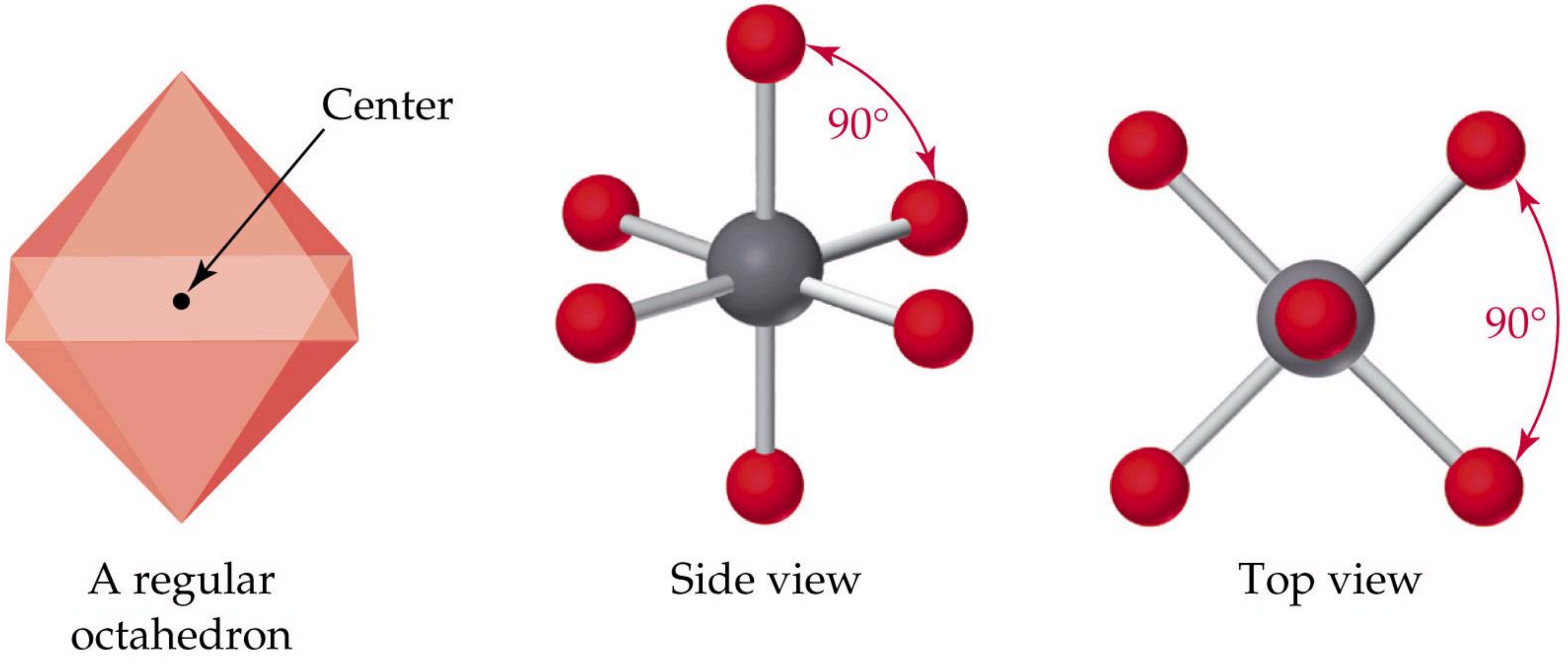




• Octahedral: A simple covalent molecule in which the central atom has *six bonding pairs* of electrons *only* in its valence shell. Example – sulfur(VI) fluoride, SF_6.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

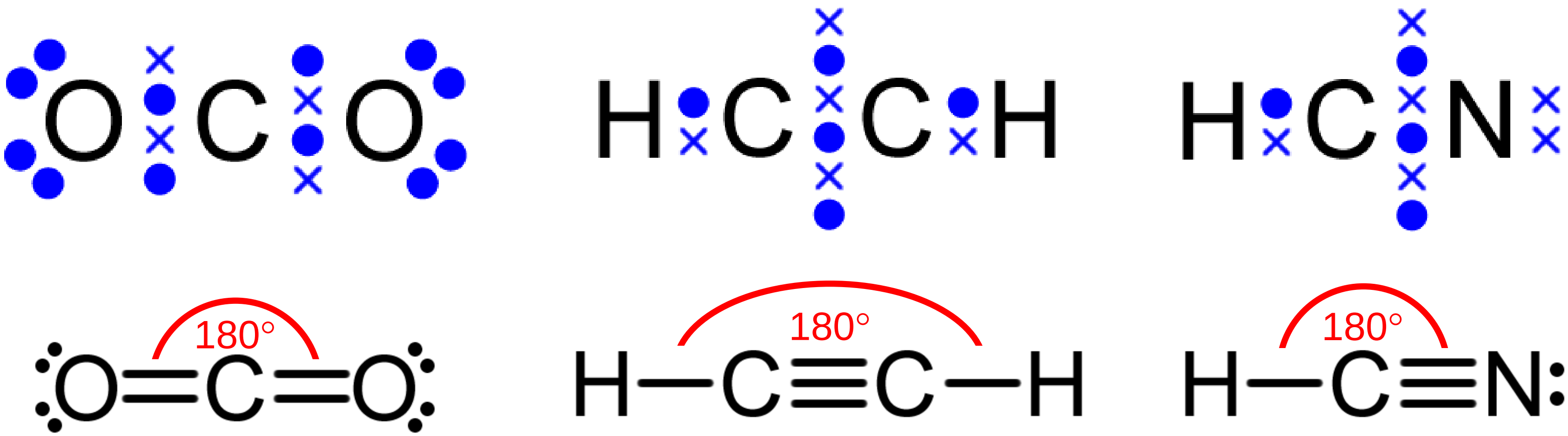


• Octahedral: A simple covalent molecule in which the central atom has *six bonding pairs* of electrons *only* in its valence shell. Example – sulfur(VI) fluoride, SF_6.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- Examples of molecules with *linear* geometry.



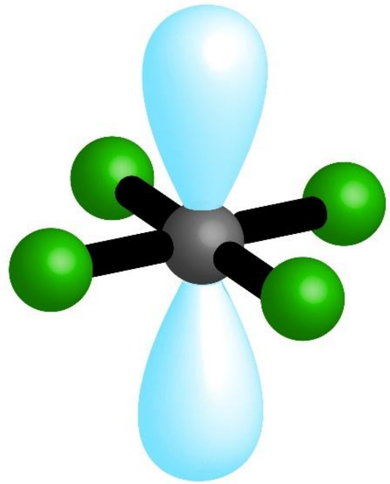
Carbon Dioxide
CO_2

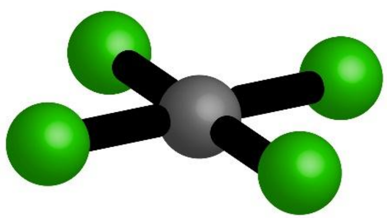
Ethyne
C_2H_2

Hydrogen Cyanide
HCN

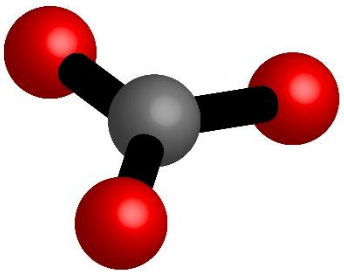
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory





square planar

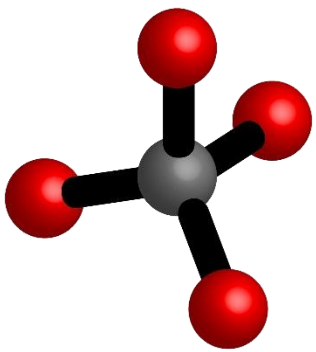


trigonal planar

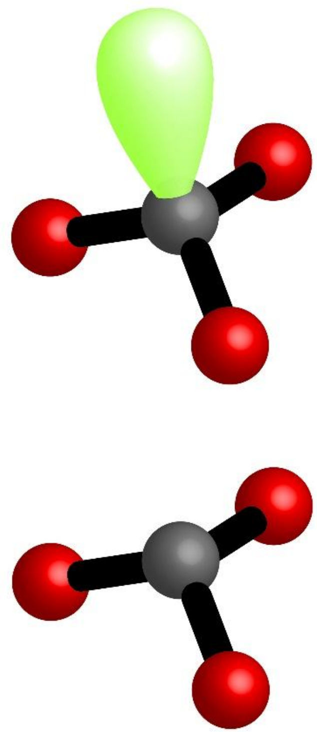
- An example of molecule with *square planar* geometry is xenon(IV) fluoride, XeF_4.
- Examples of molecules and ions with *trigonal planar* geometry include BF_3, CO_3^{2-} and NO_3^{-}.

Chemical Bonding

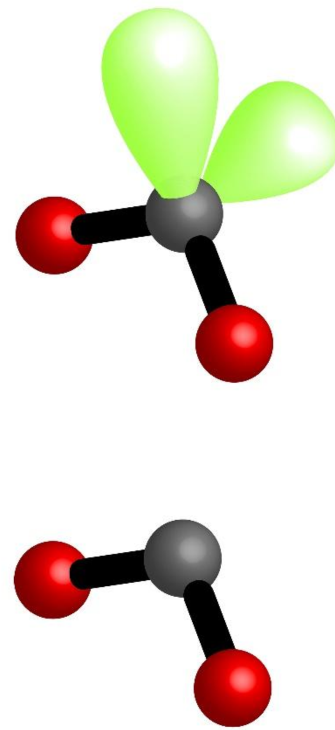
Valence Shell Electron Pair Repulsion Theory
Summary of Common Shapes



tetrahedral



trigonal pyramidal



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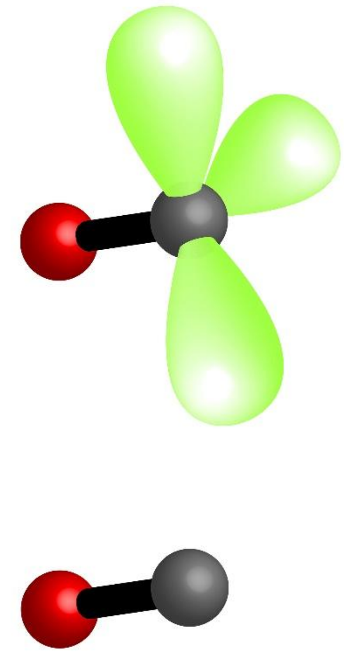
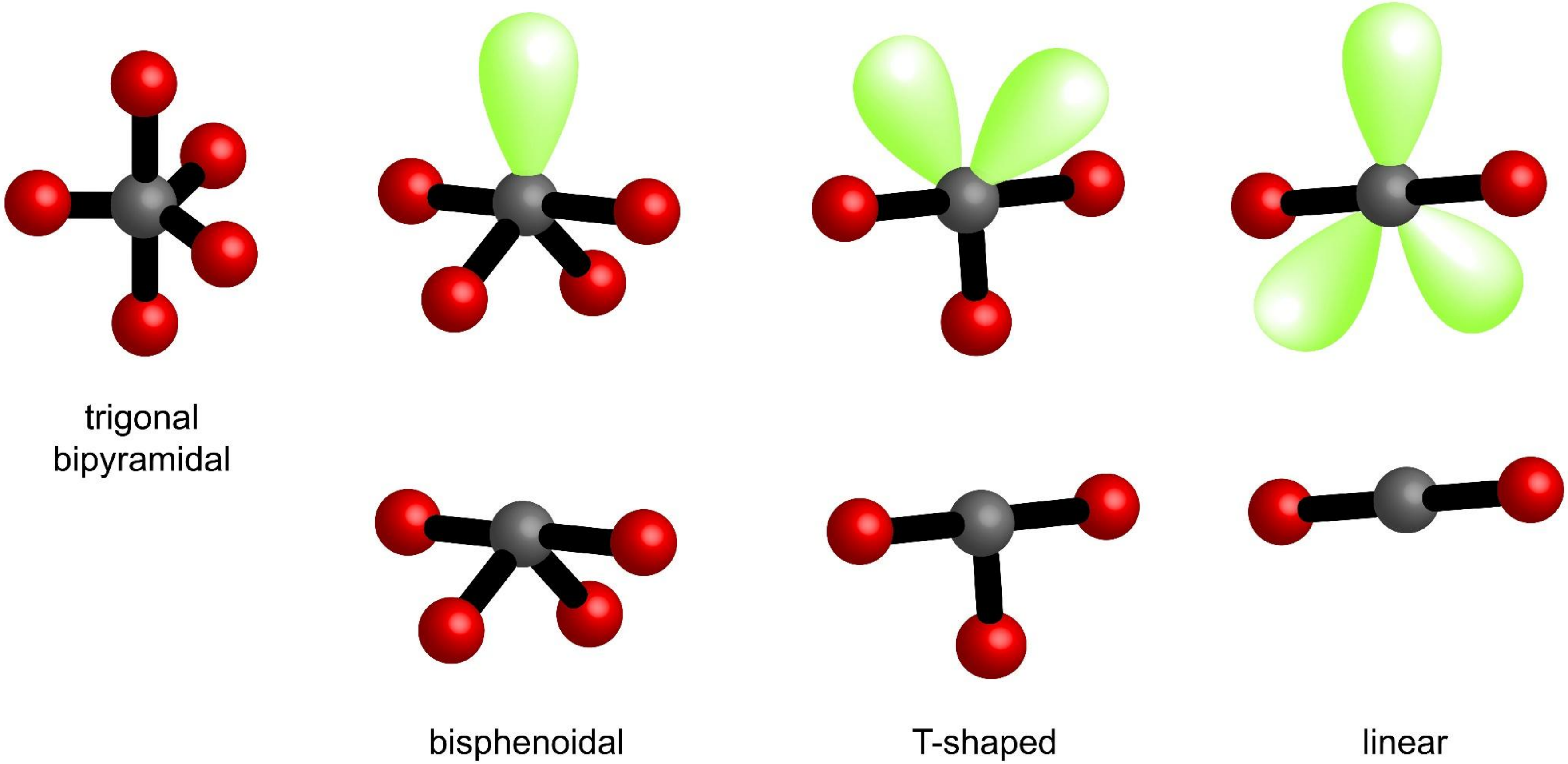


linear

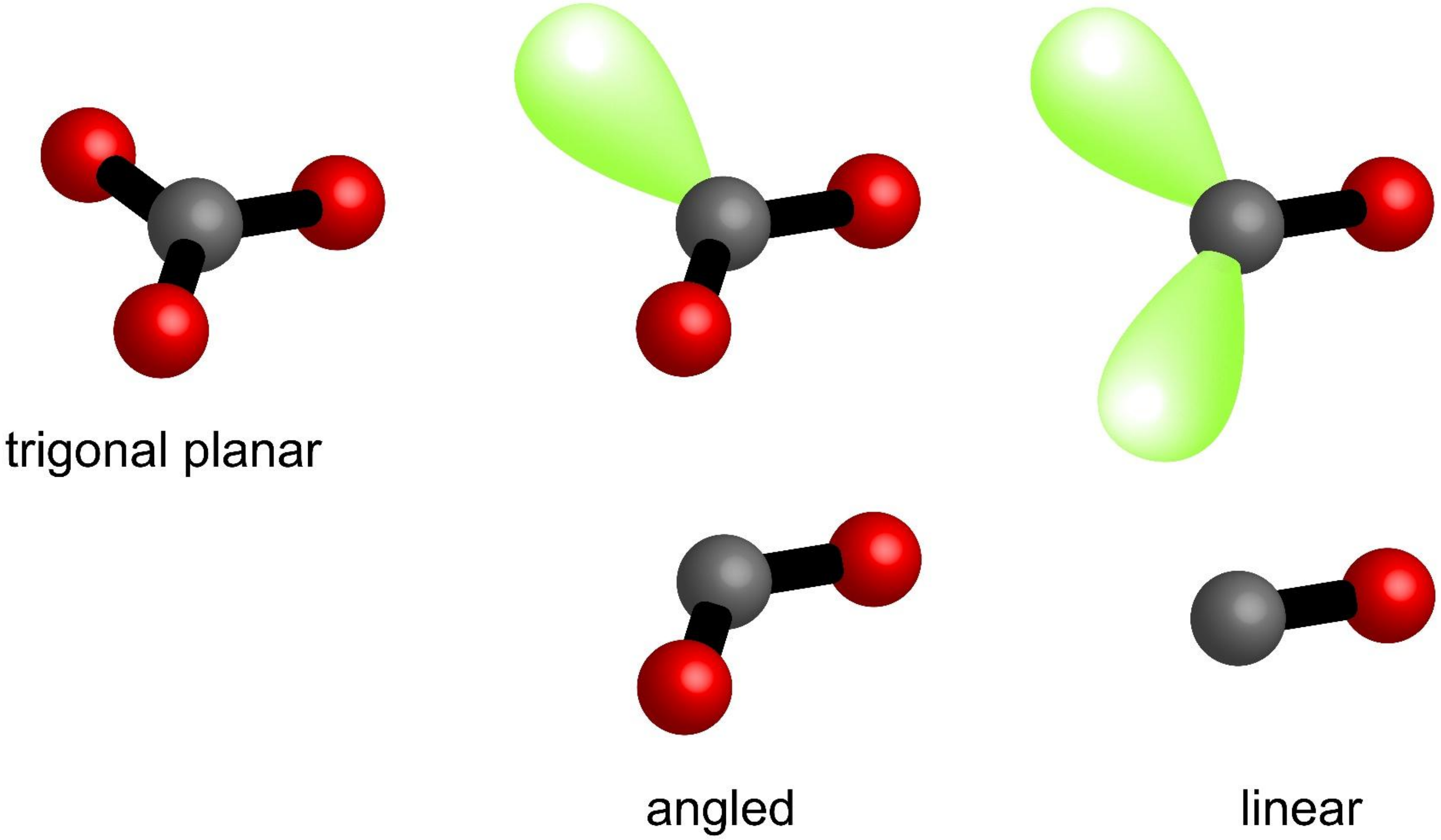
Chemical Bonding

Valence Shell Electron Pair Repulsion Theory
Summary of Common Shapes



Chemical Bonding

Valence Shell Electron Pair Repulsion Theory
Summary of Common Shapes



Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- **Question 1.**

What is the shape / geometry of a sulfur dichloride molecule, SCl_2?

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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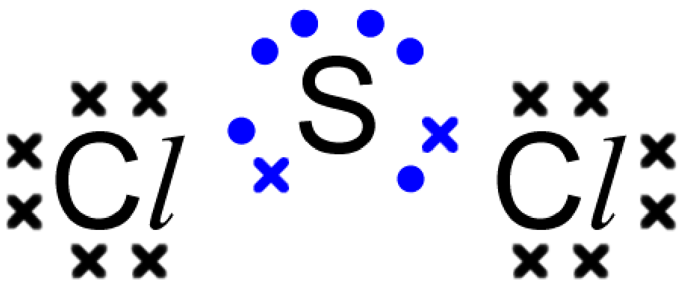

- Angular:
- 2 bonding pairs of electrons
- 2 non-bonding / lone pairs of electrons

Chemical Bonding

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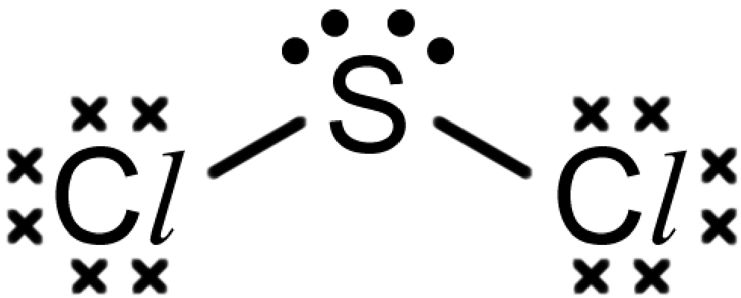

- Angular:
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Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- **Question 2.**

What is the shape / geometry of a phosphine molecule, PH_3?

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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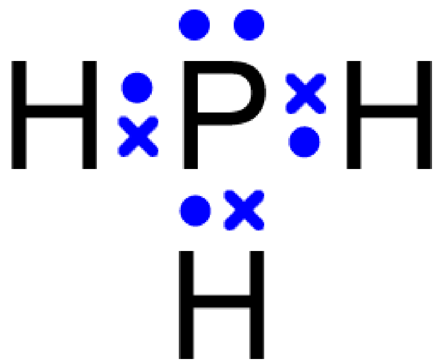
What is the shape / geometry of a phosphine molecule, PH_3?



- Trigonal Pyramidal:
- 3 bonding pairs of electrons
- 1 non-bonding / lone pair of electrons

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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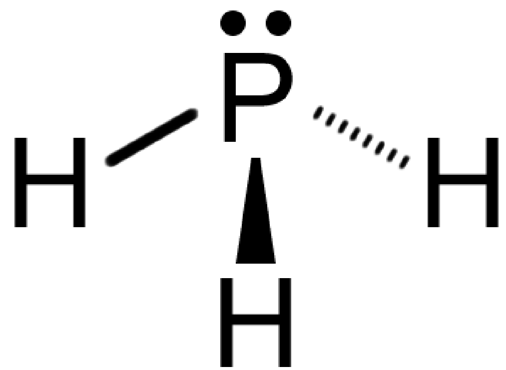

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Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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What is the shape / geometry of a silane molecule, SiH_4?

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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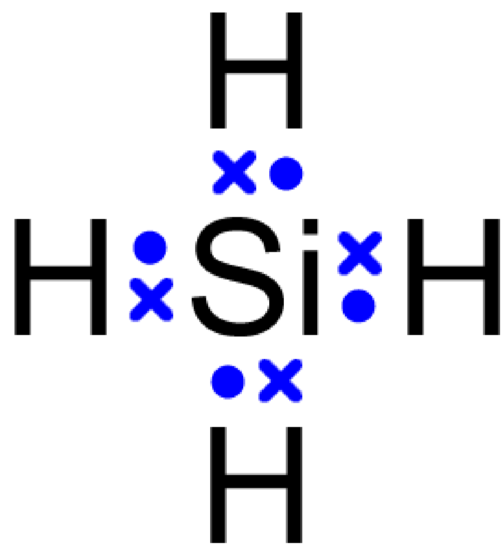

- Tetrahedral:
- 4 bonding pairs of electrons
- 0 non-bonding / lone pair electrons

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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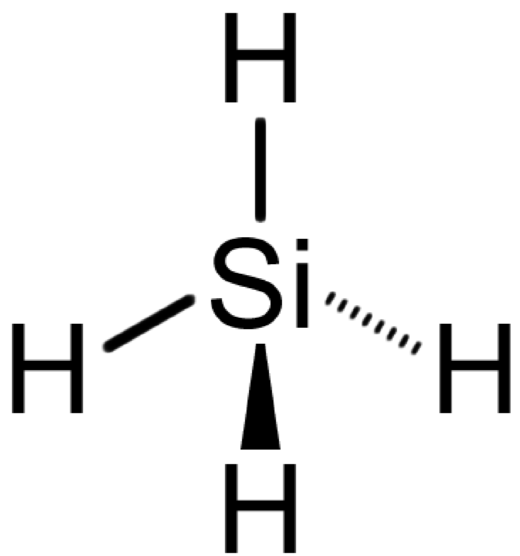

- Tetrahedral:
- 4 bonding pairs of electrons
- 0 non-bonding / lone pair electrons

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- **Question 4.**

What is the shape / geometry of an ethene molecule, CH_2CH_2?

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

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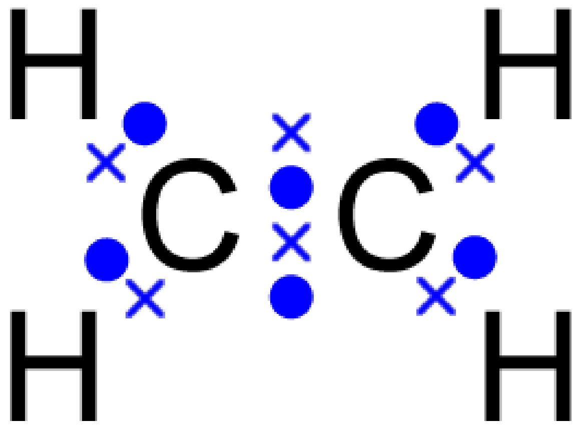


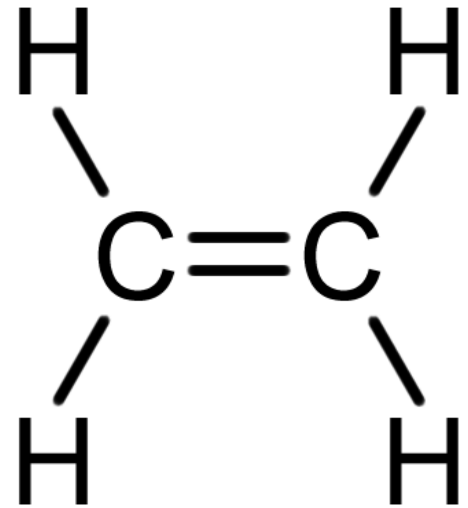
- The geometry around each carbon atom is trigonal planar.

Chemical Bonding

Valence Shell Electron Pair Repulsion Theory

- **Question 4.**

What is the shape / geometry of an ethene molecule, CH_2CH_2?



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Chemical Bonding

Intermolecular Forces of Attraction

13. Covalent bonds exist *within* molecules (*intramolecular*). What forces of attraction exist *between* molecules (*intermolecular*)?

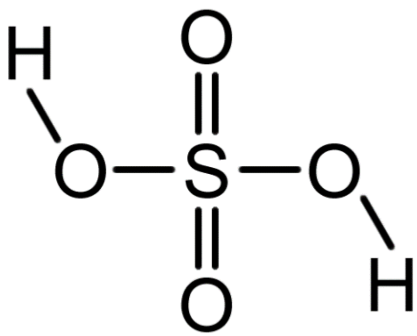
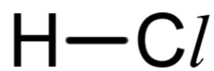
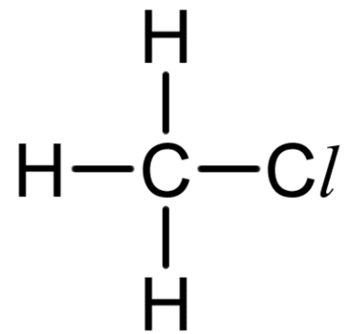
- Why is *methane*, CH_4, a *gas* at room temperature and pressure, while *water*, H_2O, is a *liquid*?

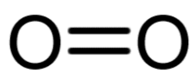
Main Menu

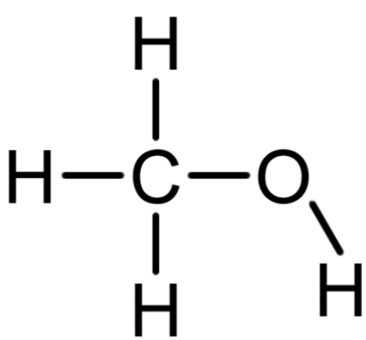
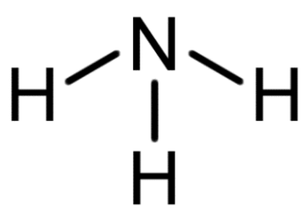
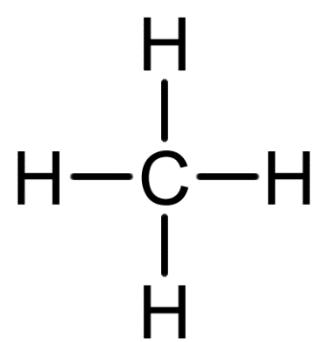
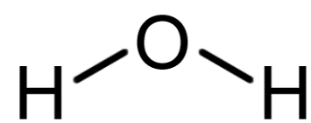
Chemical Bonding

Intermolecular Forces of Attraction

- Which intermolecular force of attraction is most significant for each one of the following chemicals?



- Click here to view answers on slides 289, 290 and 291.

Chemical Bonding

Intermolecular Forces of Attraction

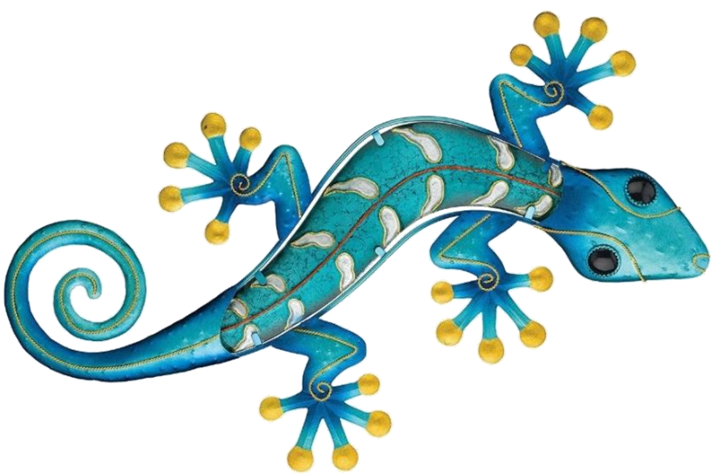
- *Intermolecular forces* are the forces that are responsible for the interactions between molecules. Intermolecular forces may be forces of *attraction* or *repulsion* between a molecule and a neighbouring particle, such as another molecule, an atom or an ion.

- *Intermolecular forces* are weak relative to *intramolecular forces*, *i.e.* the forces that hold a single molecule together. For example, a *covalent bond* – the shared pair of electrons between two atoms – is much stronger than the intermolecular force of attraction between two neighbouring molecules.

Chemical Bonding

Intermolecular Forces of Attraction

- When molecules pack together in the liquid or solid state, there must be forces of attraction holding the molecules close to each other. J. D. van der Waals postulated the existence of forces of attraction that are neither ionic nor covalent. Such forces arise in a number of ways and are collectively called *van der Waals forces*. One phenomenon due to van der Waals forces of attraction is a gecko's ability to walk up a vertical glass surface.



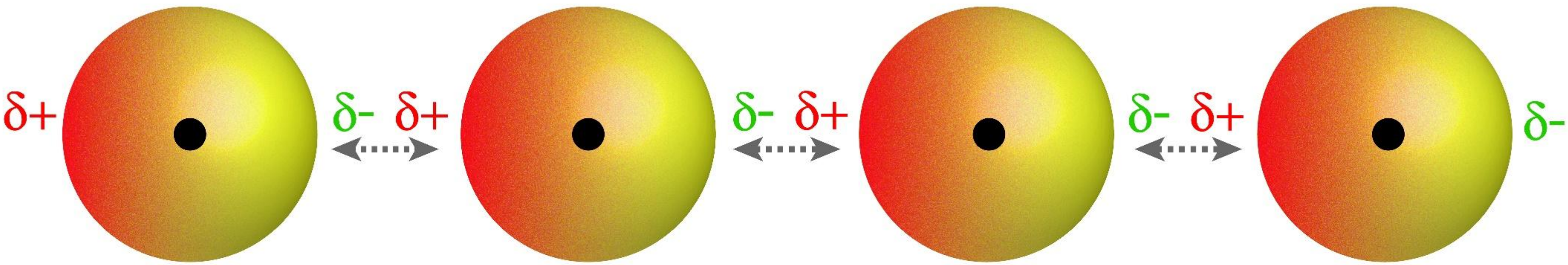
Chemical Bonding

Intermediate Forces of Attraction

- The gecko's feet are covered with many microscopic hairs or *setae* which, when combined together, have a *very large surface area*. When weak van der Waals forces are exerted over a large surface area, their strength becomes very significant – strong enough to support the gecko's body weight as it climbs walls and ceilings.

Chemical Bonding

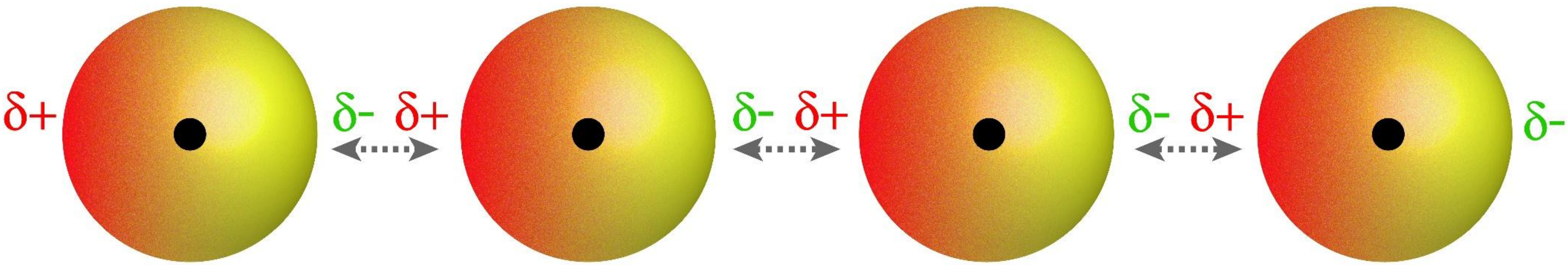
Intermolecular Forces – London Dispersion Forces



- There are different types of van der Waals forces of attraction. Attraction between *polar molecules* is one type of van der Waals force. Attractive forces also exist between *non-polar molecules*. Even atoms of the noble gases are attracted towards each other to a small extent, which explains why the noble gases can be liquified.

Chemical Bonding

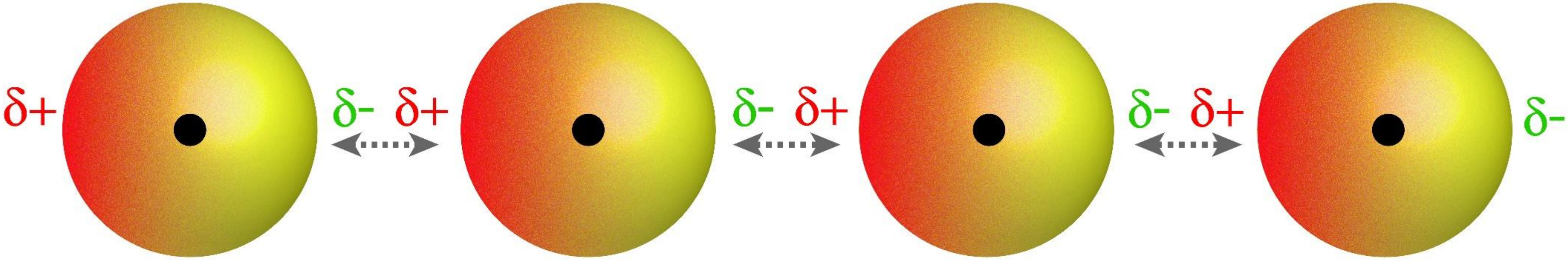
Intermolecular Forces – London Dispersion Forces



- Consider two non-polar molecules that are very close together. They are non-polar because the electrons within each molecule are *evenly distributed* throughout the molecule (the distribution of electrons is said to be *symmetrical*). However, at any given moment, because the electrons are *moving randomly* within each molecule, the electron distribution within one molecule may suddenly become *unsymmetrical*.

Chemical Bonding

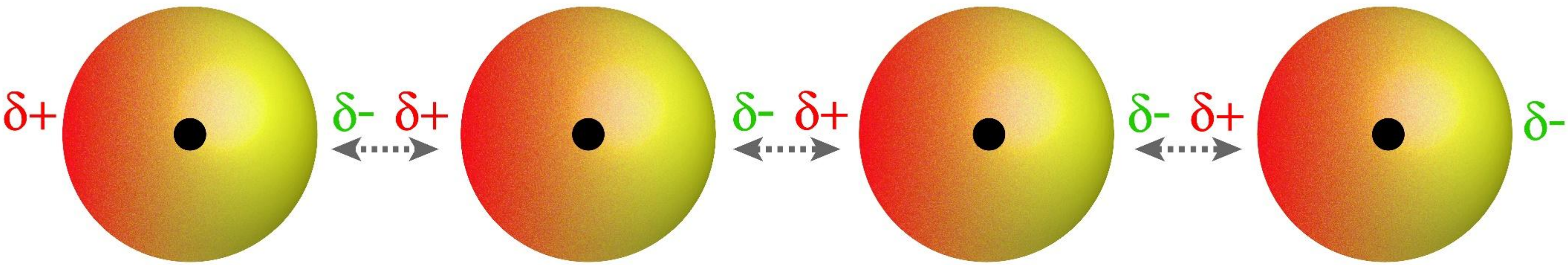
Intermolecular Forces – London Dispersion Forces



- Random movement of electrons, leading to their unsymmetrical distribution in a molecule, leads to one area of the molecule having a *surplus of negatively charged electrons* while another area of the molecule has a *deficit of negatively charged electrons*. For an instant, the area of the molecule with a surplus of electrons gains a *slight negative charge* (δ–) and the region with a deficit of electrons gains a *slight positive charge* (δ+).

Chemical Bonding

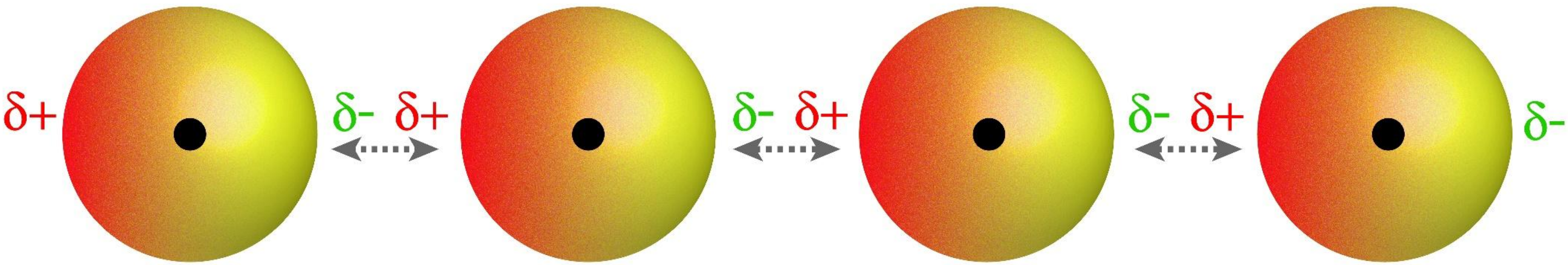
Intermolecular Forces – London Dispersion Forces



- When one area of the molecule has a δ– charge and another area of the molecule has a δ+ charge, the molecule is described as having a *instantaneous dipole*. The diagram given above shows how the instantaneous dipole of one molecule can *influence the electron cloud of a neighbouring molecule*, the δ+ area *attracting* electrons of a neighbouring molecule towards it, and the δ– area *repelling* electrons of a neighbouring molecule away from it.

Chemical Bonding

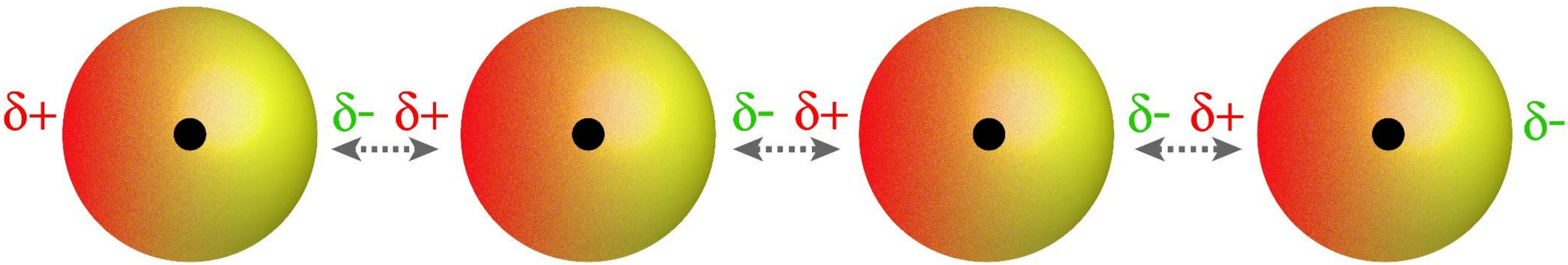
Intermolecular Forces – London Dispersion Forces



- Once a molecule has an instantaneous dipole (*i.e.* temporary regions of δ+ and δ–), it can *induce dipoles* in neighbouring molecules. This means than many molecules can have dipoles, and the directions of the dipoles will be such that they attract one another. Since electrons move about at high speed, the attraction only has a momentary existence. New dipoles will form, and they will induce diploes in the neighbouring molecules that result in attraction.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

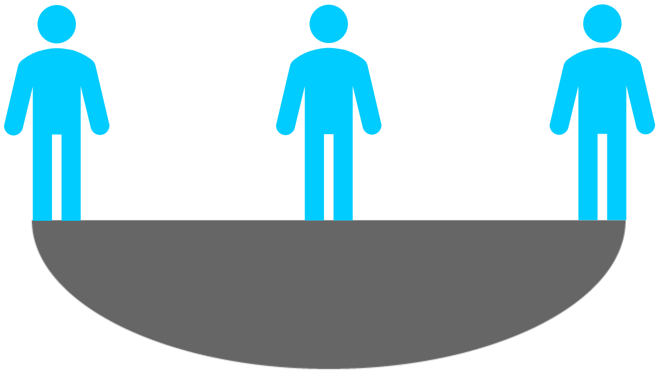


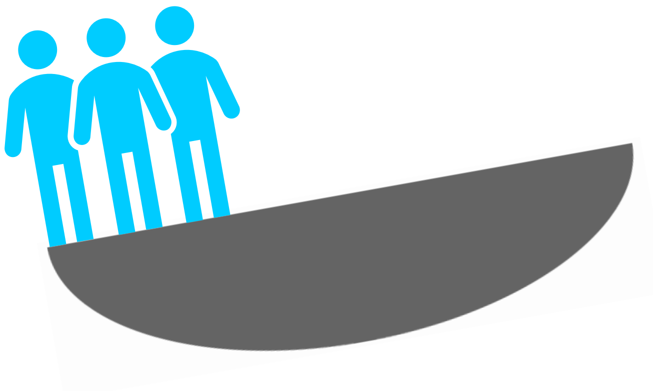
- The dipoles are only *temporary*, but the intermolecular forces of attraction are always present *somewhere* in the chemical.

- The ease with which an electron cloud is distorted, and therefore the ease with which a dipole is induced in a molecule, is called *polarisability*. Polarisability *increases with the number of electrons* in a molecule, and so the strength of the intermolecular force of attraction *increases with increasing relative molecular mass*.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

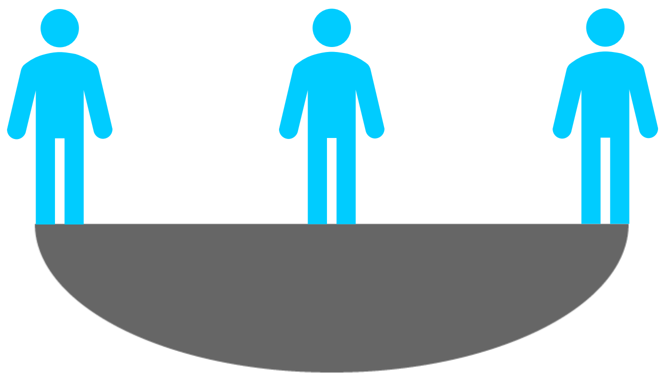


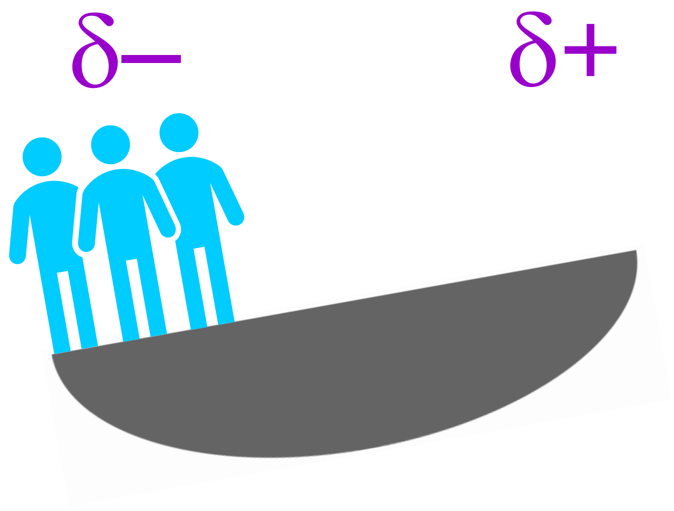


- Imagine a *small number* of people standing on a boat in the water. If the people are evenly distributed across the boat, then the boat will sit level in the water, but if the people all move to one side of the boat, then the boat will start to *tilt slightly* in the water.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces



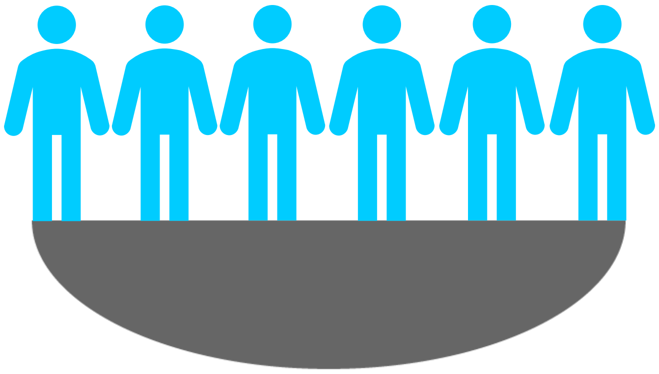


- Now imagine that the boat is a *molecule*, and that the people are *electrons*.

- A molecule with a *small number* of electrons (small electron cloud) experiences a small amount of polarizability, leading to *relatively weak London dispersion forces* between molecules.

Chemical Bonding

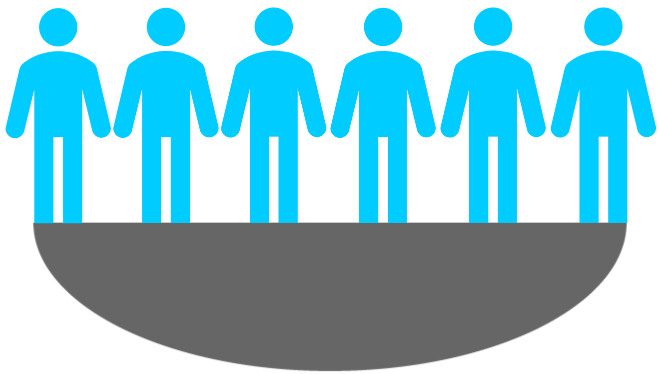
Intermolecular Forces – London Dispersion Forces

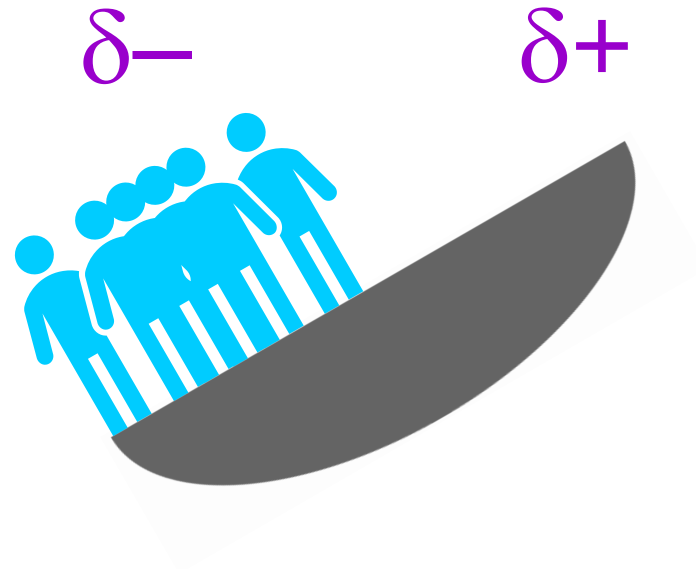


- Now imagine a *large number* of people standing on a boat in the water. If the people are evenly distributed across the boat, then the boat will sit level in the water, but if the people all move to one side of the boat, then the boat will *tilt significantly* in the water.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

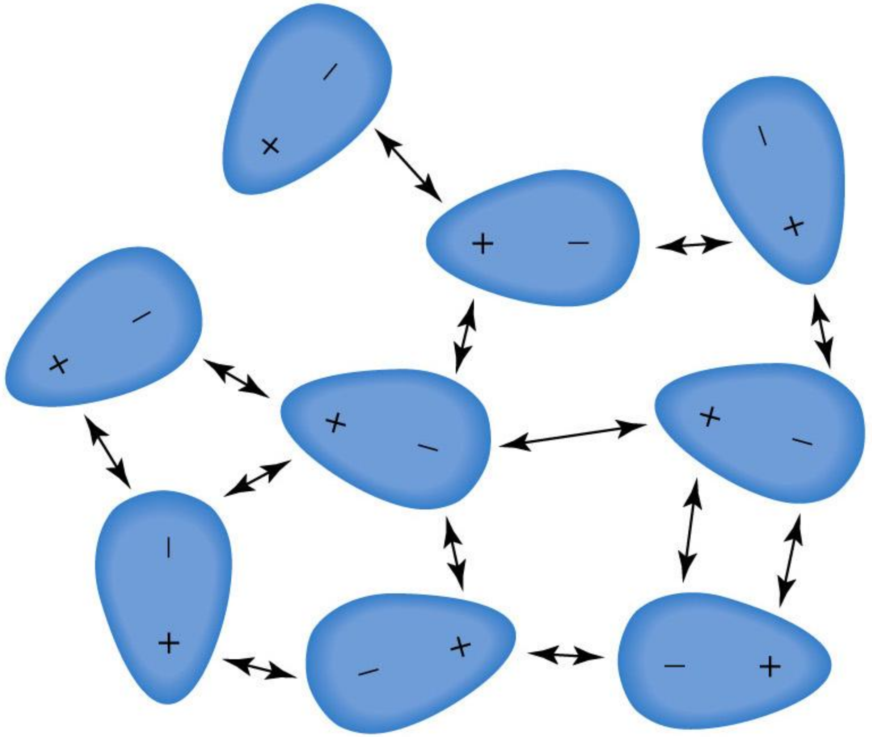




- Now imagine that the boat is a *molecule*, and that the people are *electrons*.
- A molecule with a *large number* of electrons (large electron cloud) experiences a large amount of polarizability, leading to *stronger London dispersion forces* between molecules.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces



• *London dispersion forces*: An *instantaneous dipole* in one molecule, caused by the random movement of electrons within the molecule, *induces dipoles* in the surrounding molecules, which results in an *intermolecular force of attraction*.

Chemical Bonding

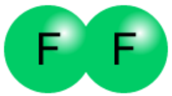
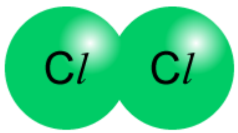
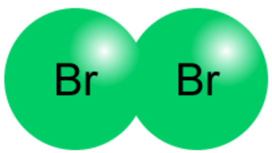
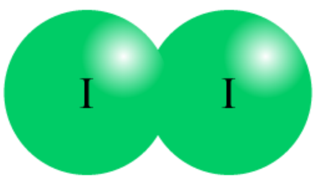
Intermolecular Forces – London Dispersion Forces

Halogen	melting point /°C	boiling point /°C	covalent radius / pm	total number of electrons	relative molecular mass
F F	-220	-188	71	9 + 9 = 18	38.0
C*l* C*l*	-101	-35.0	99	17 + 17 = 34	71.0
Br Br	-7.20	58.8	114	35 + 35 = 70	159.8
I I	114	184	133	53 + 53 = 106	253.8

- For Group 17, as the *number of electrons* (*size of the electron cloud*) within a single molecule of the element *increases*, so the *polarisability* of the molecule increases. This increases the strength of the *London dispersion forces*, hence increasing melting and boiling points.

Chemical Bonding

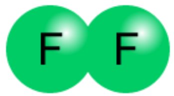
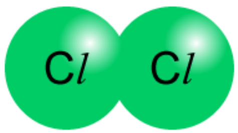
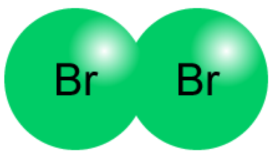
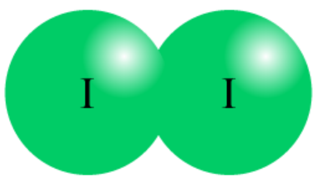
Intermolecular Forces – London Dispersion Forces

Halogen	melting point /°C	boiling point /°C	covalent radius / pm	total number of electrons	relative molecular mass
	-220	-188	71	9 + 9 = 18	38.0
	-101	-35.0	99	17 + 17 = 34	71.0
	-7.20	58.8	114	35 + 35 = 70	159.8
	114	184	133	53 + 53 = 106	253.8

- As the covalent radii of the atoms increases, so the *surface area* of the molecules also *increases*. The larger the surface area available for London dispersion forces to operate over, the stronger they become, resulting in *higher* melting and boiling points.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

Halogen	melting point /°C	boiling point /°C	covalent radius / pm	total number of electrons	relative molecular mass
 F F	-220	-188	71	9 + 9 = 18	38.0
 C*l* C*l*	-101	-35.0	99	17 + 17 = 34	71.0
 Br Br	-7.20	58.8	114	35 + 35 = 70	159.8
 I I	114	184	133	53 + 53 = 106	253.8

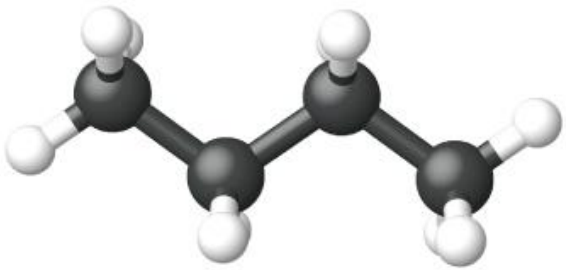
- As the *relative molecular mass* of the diatomic molecules *increases*, so the melting points and boiling points also *increase*. More energy is required to overcome the *inertia** of the molecules with the greater relative molecular mass.

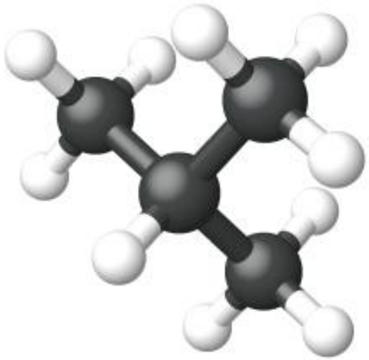
***Note:** *Inertia* is the natural tendency of an object to stay at rest, or remain in motion.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

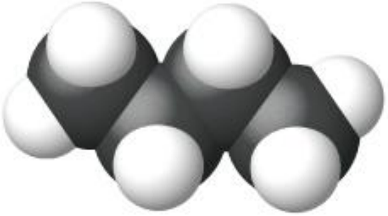
- Butane, C_4H_{10}
- Ball-and-stick model.





- 2-Methylpropane, C_4H_{10}
- Ball-and-stick model.

- Butane, C_4H_{10}
- Space-filled model.



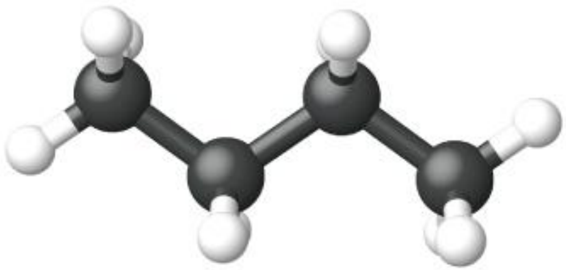
- 2-Methylpropane, C_4H_{10}
- Space-filled model.

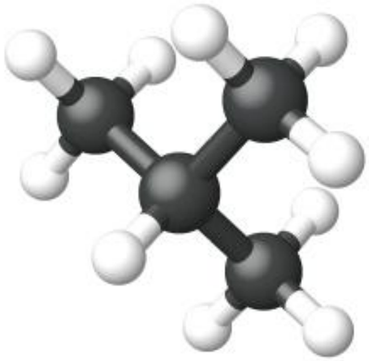
- Butane and 2-methylpropane have the same molecular formulae, but different structural formulae (they are *isomers*).
- Compared to butane, 2-methylpropane has a more *compact structure*, with a *smaller surface area*.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

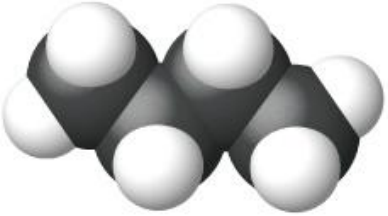
- Butane, C_4H_{10}
- Ball-and-stick model.





- 2-Methylpropane, C_4H_{10}
- Ball-and-stick model.

- Butane, C_4H_{10}
- Space-filled model.



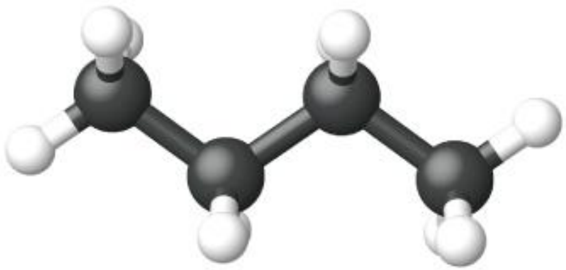
- 2-Methylpropane, C_4H_{10}
- Space-filled model.

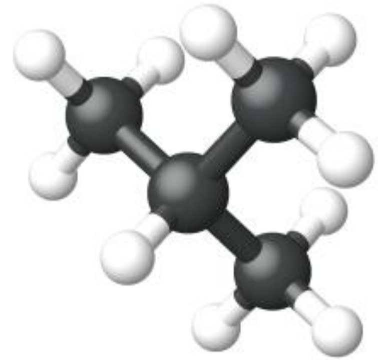
- As a consequence, the area of contact between two molecules of 2-methylpropane is *smaller* than the area of contact between two molecules of butane, resulting in a *weaker London dispersion forces* between the molecules of 2-methylpropane.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

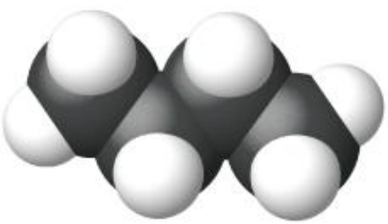
- Butane, C_4H_{10}
- Ball-and-stick model.

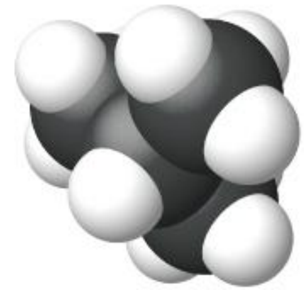




- 2-Methylpropane, C_4H_{10}
- Ball-and-stick model.

- Butane, C_4H_{10}
- Space-filled model.





- 2-Methylpropane, C_4H_{10}
- Space-filled model.

- This explains why the boiling point of 2-methylpropane is *lower* (–11.7°C) than the boiling point of butane (–1.00°C), because the *weaker London dispersion forces* in 2-methylpropane require *less thermal energy* to overcome.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

- Spaghetti

- macaroni

- Imagine 500 g of *spaghetti* and 500 g of *macaroni*.

- The spaghetti is analogous to the straight-chain molecules of butane, while the macaroni is analogous to more compact branched-chain molecules of 2-methylpropane.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

• Spaghetti

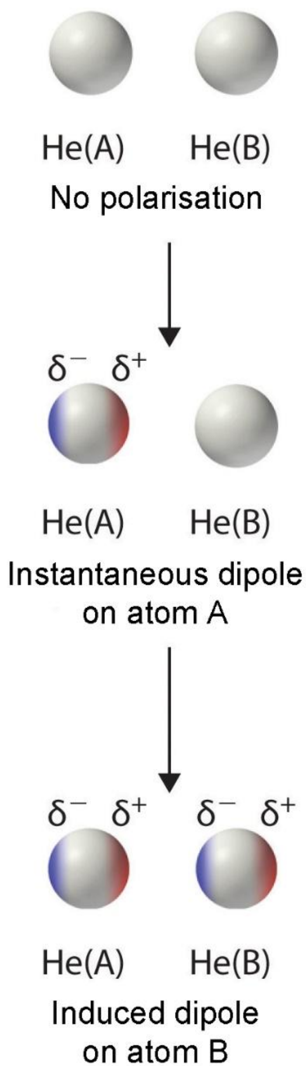
• macaroni

• The long and straight strands of spaghetti can pack closely together, and therefore have a relatively *large surface area in contact* with each other (analogous to stronger London dispersion forces).

• For the same mass of macaroni, the pieces of pasta are more compact and have a relatively *small surface area in contact* with each other (analogous to weaker London dispersion forces).

Chemical Bonding

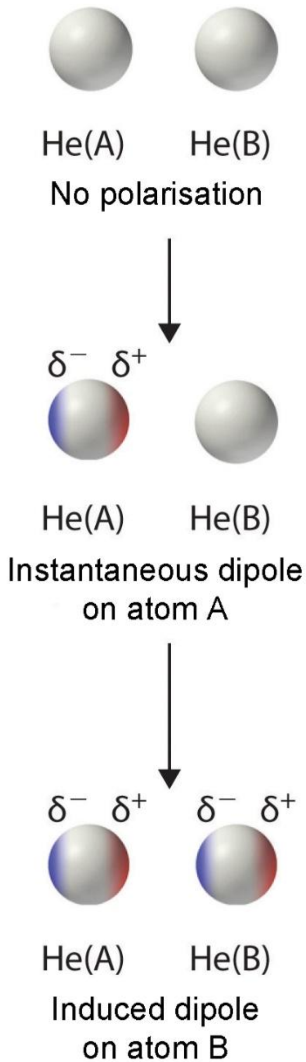
Intermolecular Forces – London Dispersion Forces



- Temporary fluctuations in the electron distributions within atoms and non-polar molecules can result in the formation of short lived instantaneous dipole moments, which produce forces of attraction known as *London dispersion forces* (named after the German Physicist, Fritz London) between otherwise non-polar substances.

Chemical Bonding

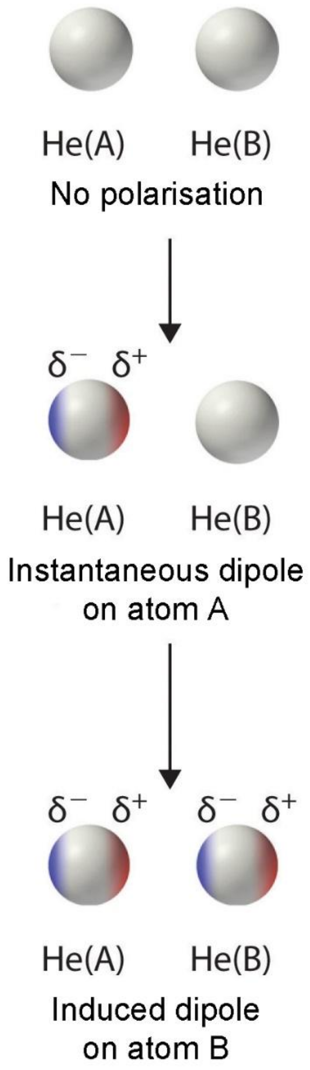
Intermolecular Forces – London Dispersion Forces



- Consider a pair of adjacent He atoms, for example. On average, the two electrons in each He atom are uniformly distributed around the nucleus. Because the electrons are in constant motion, however, their distribution in one atom is likely to be *asymmetrical* at any given instant, resulting in an *instantaneous dipole moment*.

Chemical Bonding

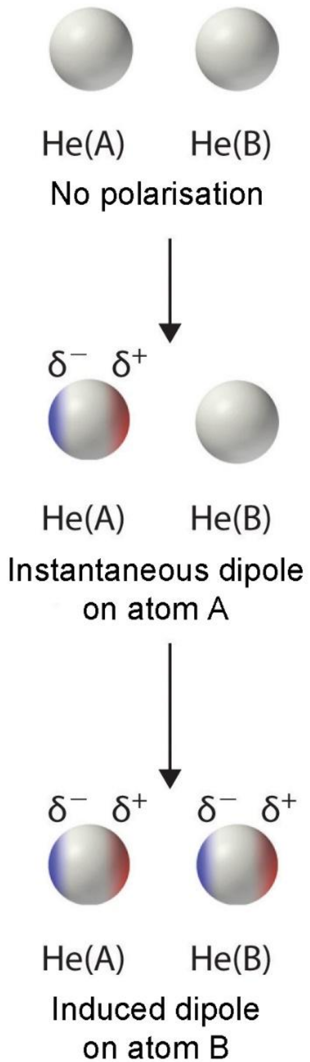
Intermolecular Forces – London Dispersion Forces



- The *instantaneous dipole moment* on one atom can interact with the electrons in an adjacent atom, pulling them toward the positive end of the instantaneous dipole or repelling them from the negative end.

Chemical Bonding

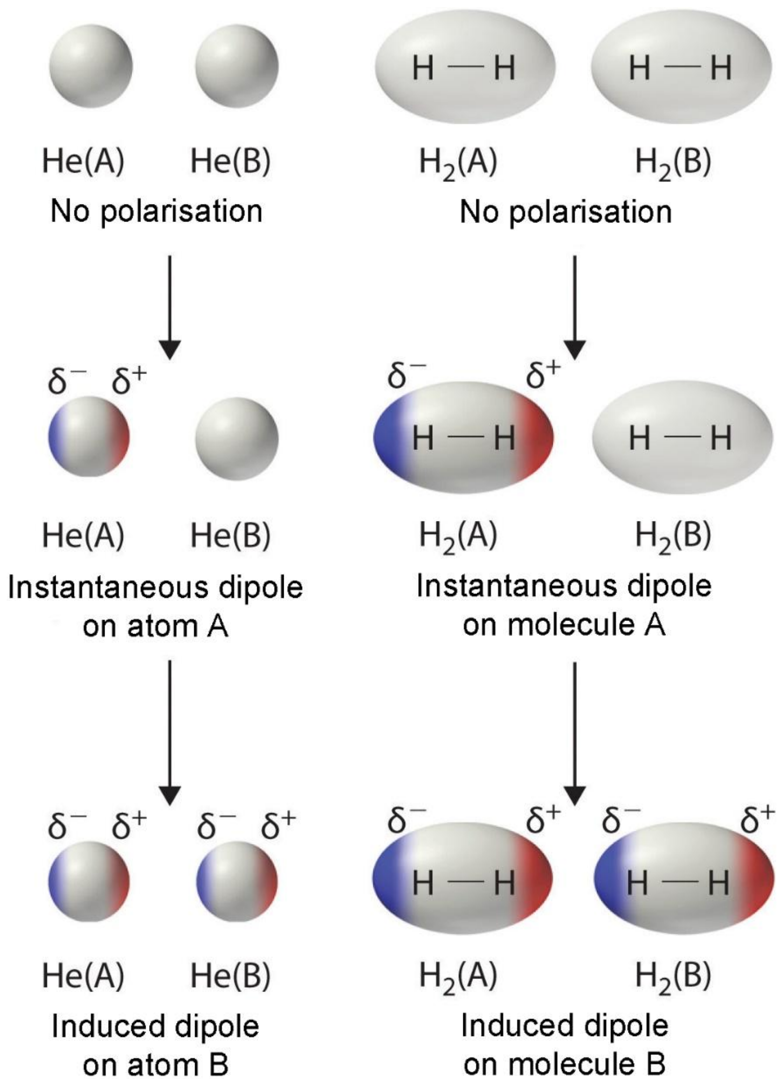
Intermolecular Forces – London Dispersion Forces



- The net effect is that the first atom causes the temporary formation of a dipole, called an *induced dipole*, in the second. Interactions between these temporary dipoles cause atoms to be attracted to one another. These attractive interactions are weak and fall off rapidly with increasing distance.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces

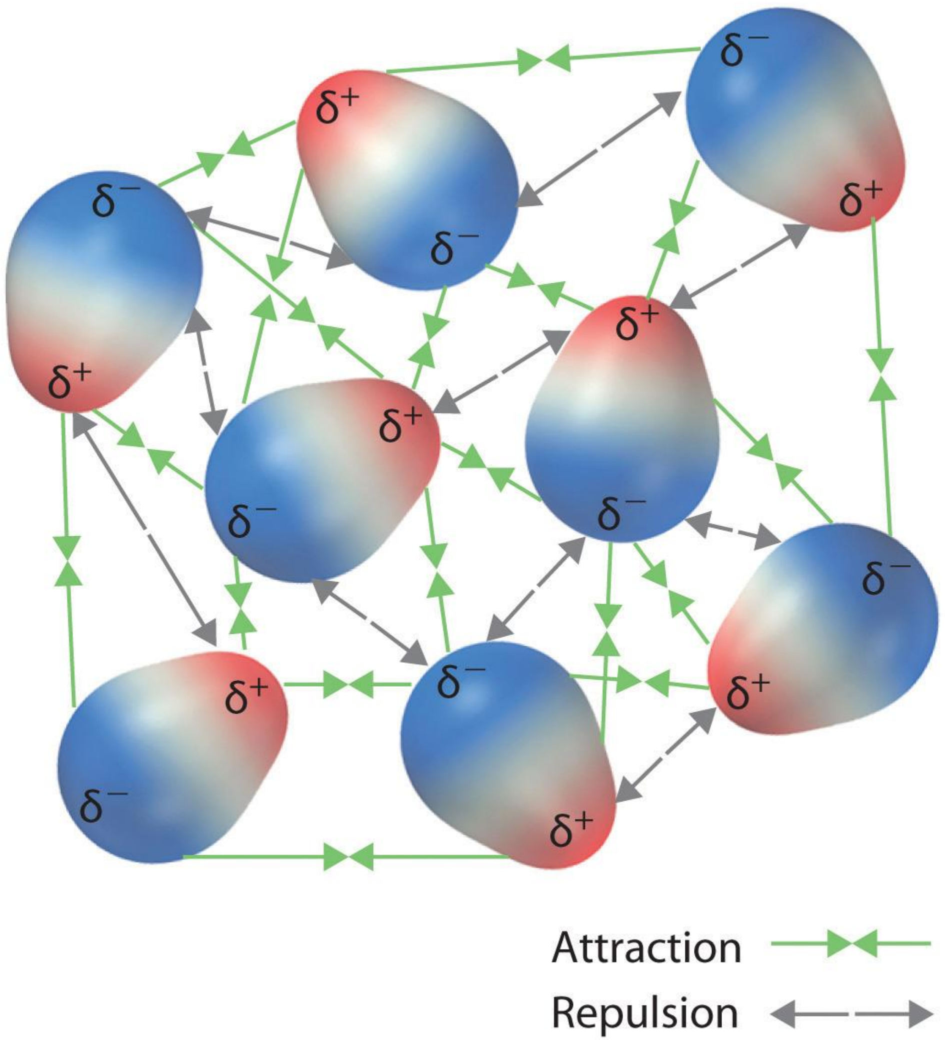


- London dispersion forces are known as *instantaneous dipole – induced dipole* interactions.

- London dispersion forces can occur between different combinations of atoms and non-polar molecules.

Chemical Bonding

Intermolecular Forces – London Dispersion Forces



- ***London dispersion forces***: An *instantaneous dipole* in one molecule, caused by the random movement of electrons within the molecule, *induces dipoles* in the surrounding molecules, which result in an *intermolecular force of attraction*.

Chemical Bonding

Intermolecular Forces – Electronegativity

- Definition of *Electronegativity*

→ Electronegativity is a measure of the relative tendency of an atom to attract a *bonding pair of electrons*.

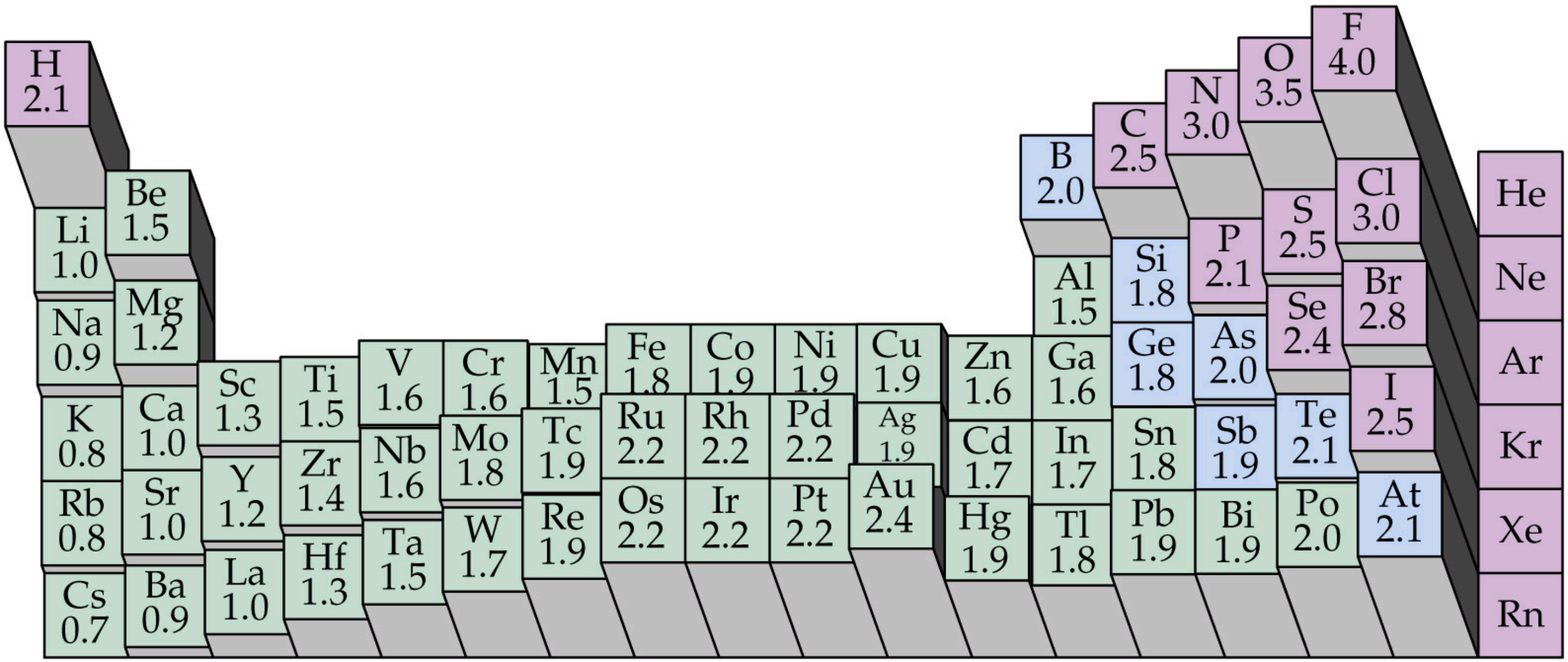
→ Electronegativity values of the chemical elements are given on a scale of 0 to 4, with 4 being the most electronegative.

→ Electronegativity values are *relative* and do not have any units.

Chemical Bonding

Intermolecular Forces – Electronegativity

- Electronegativity Values of the Chemical Elements (Pauling Scale)

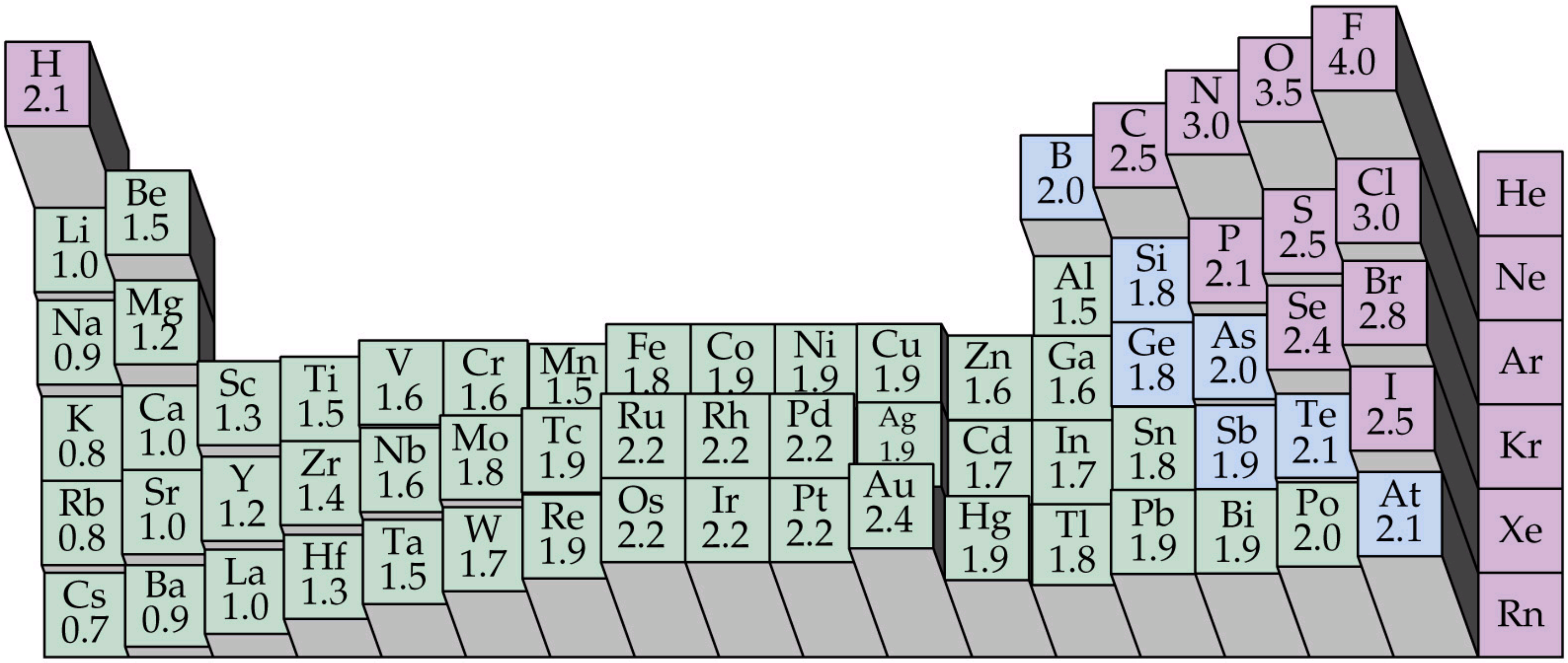


- Why are no electronegativity values assigned to the Noble Gases (Group 18)?

Chemical Bonding

Intermolecular Forces – Electronegativity

- Electronegativity Values of the Chemical Elements (Pauling Scale)



- Noble Gases are unreactive. They do not form covalent bonds, *i.e.* they do not share pairs of electrons with other atoms.

Chemical Bonding

Intermolecular Forces – Electronegativity

→ When two atoms of different elements are held together by a covalent bond, if the difference in electronegativity values of the two elements is greater than ~0.5, then the covalent bond will be *polar*.

→ The atom of the element with the *greater* electronegativity value, because it is pulling the negatively charged bonding pair of electrons *towards* it, will gain a *slight negative charge*, written as $\delta-$ (delta negative).

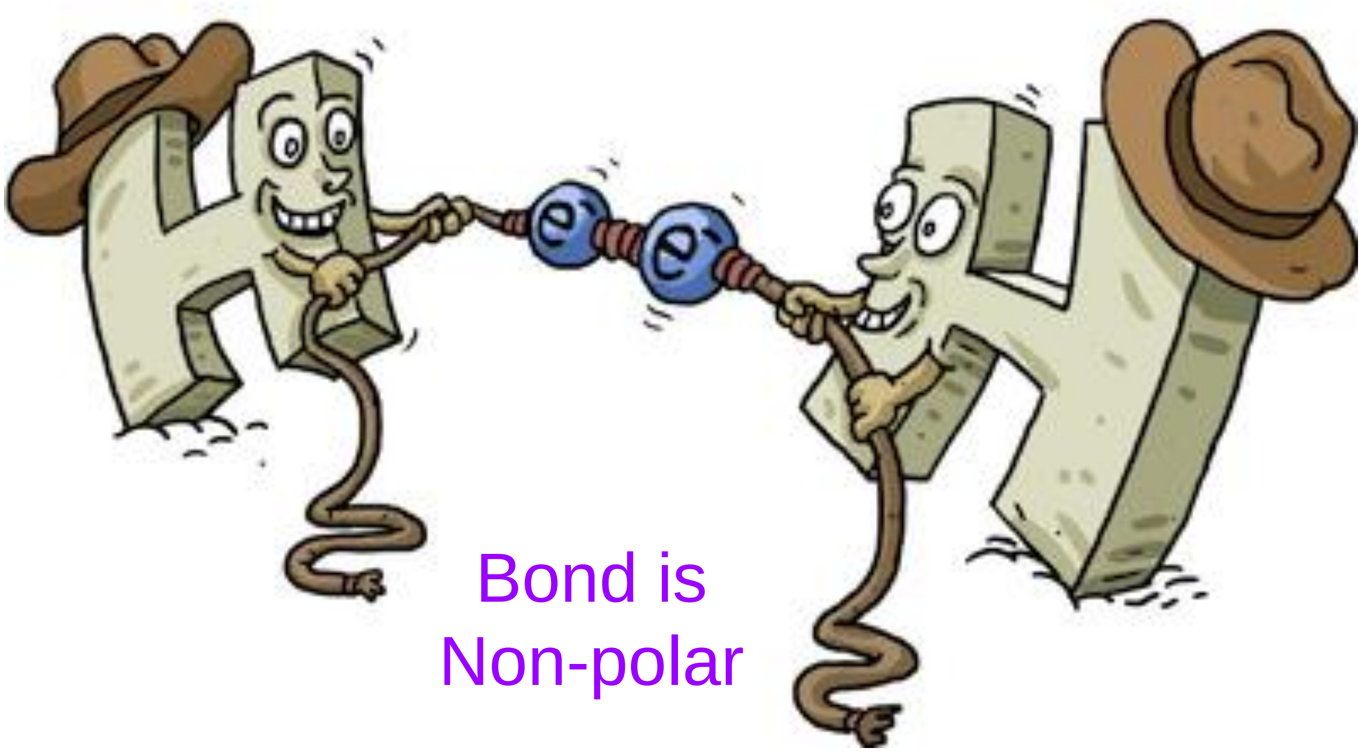
→ The atom of the element with the *smaller* electronegativity value, because it has the negatively charged bonding pair of electrons pulled *away* from it, will gain a *slight positive charge*, written as $\delta+$ (delta positive).

Chemical Bonding

Intermolecular Forces – Electronegativity

Hydrogen = 2.1

Hydrogen = 2.1



Bond is
Non-polar

Chemical Bonding

Intermolecular Forces – Electronegativity

Chlorine = 3.0 Hydrogen = 2.1

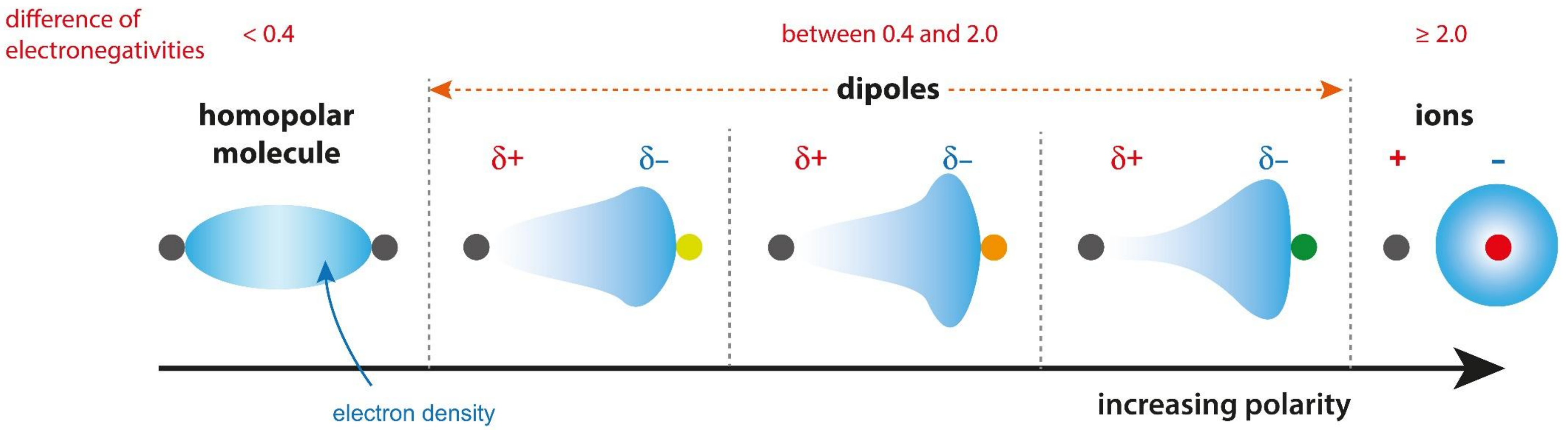
$\delta-$

$\delta+$

Bond has a Permanent Dipole

Chemical Bonding

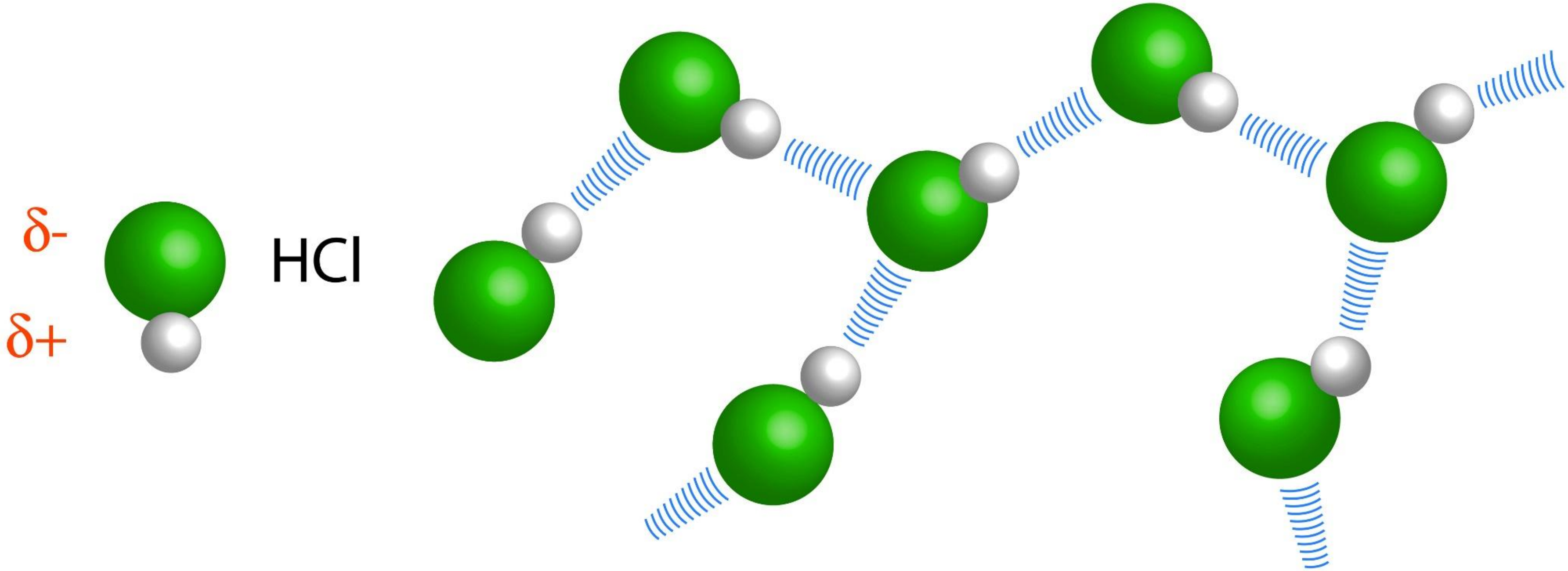
Permanent Dipole Interactions



• Ionic and covalent bonding can be considered to exist at opposites ends of a *spectrum*. In between covalent bonds and ionic bonds are *dipoles* or *polar covalent bonds* (bonds with δ+ and δ– charges). The *greater* the difference in *electronegativity* values between the two bonded atoms, the *greater* the *polarity* of the bond between them.

Chemical Bonding

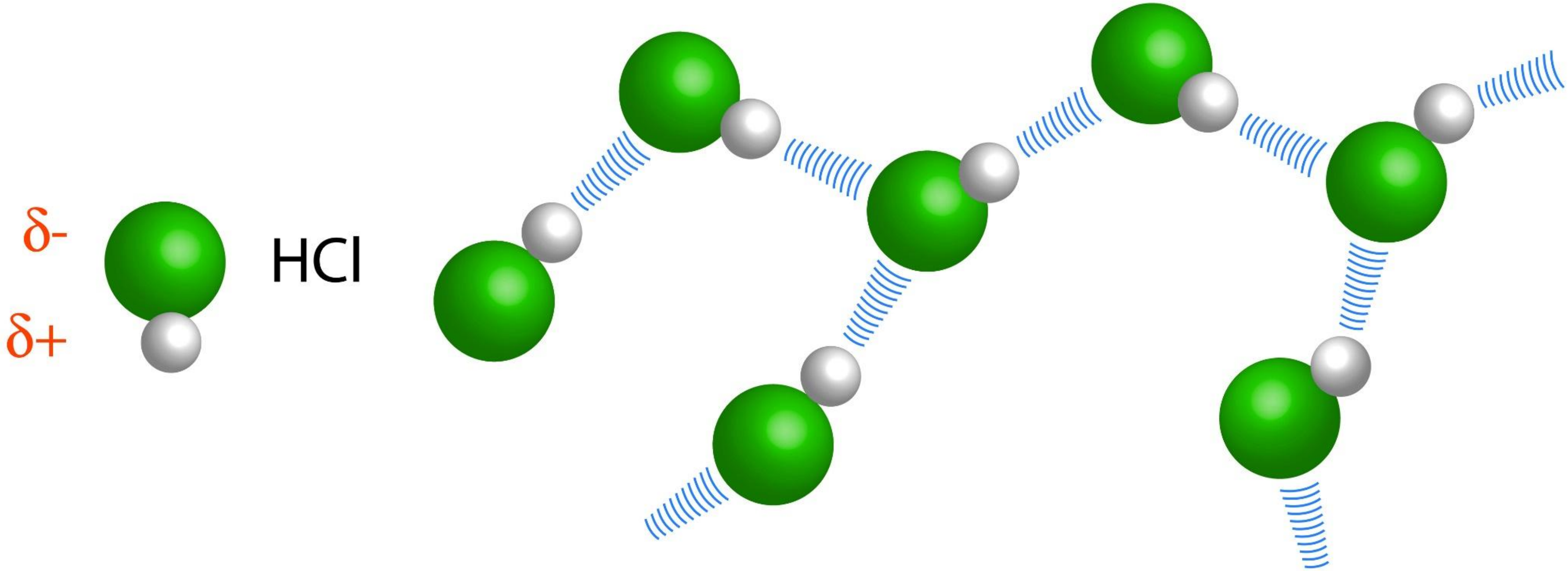
Permanent Dipole Interactions



- Permanent dipole-permanent dipole interactions are electrostatic forces of attraction between molecules that have *permanent dipoles* (permanent polar covalent bonds). Molecules align themselves with regions of opposite charge attracting each other (regions with the same charge will repel).

Chemical Bonding

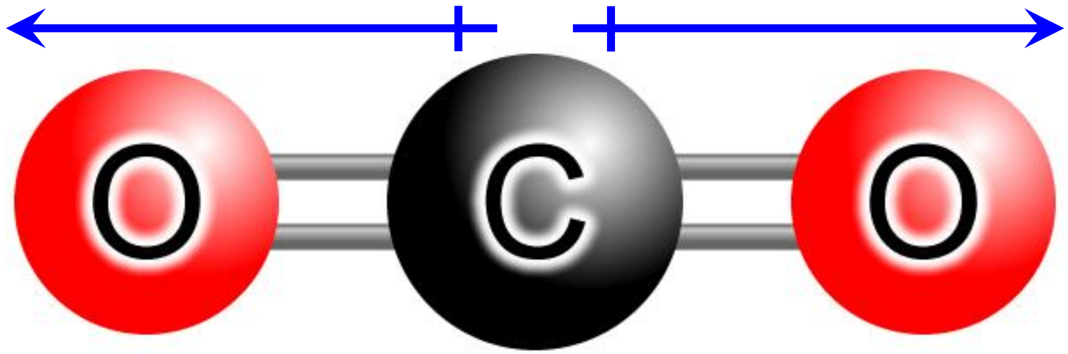
Permanent Dipole Interactions



- An example of a permanent dipole-permanent dipole interaction can be seen in hydrogen chloride. The δ+ *hydrogen atom* of one molecule will attract the δ– *chlorine atom* of a neighbouring molecule, resulting in a net force of attraction between the two neighbouring molecules.

Chemical Bonding

Permanent Dipole Interactions

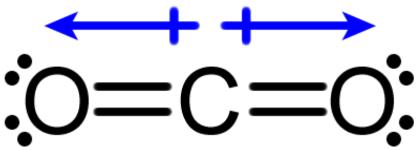


- Although the electronegativity value of *carbon* is *2.5*, and the electronegativity value of *oxygen* is *3.5*, the carbon-to-oxygen covalent bond is not a permanent dipole / is not polar. This is because the molecule *linear* and *symmetrical*.

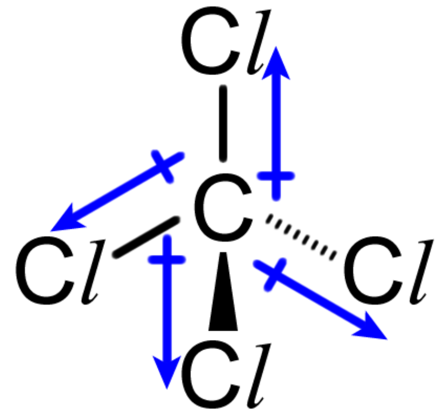
- We can imagine that as the oxygen atom on the left-hand-side pulls the bonding pairs of electrons *away from carbon*, the oxygen atom on the right-hand-side pull the pairs of electrons *back towards the carbon* again. With no uneven distribution of electrons, there are *no permanent dipoles in* CO_2.

Chemical Bonding

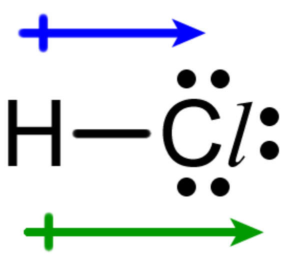
Permanent Dipole Interactions



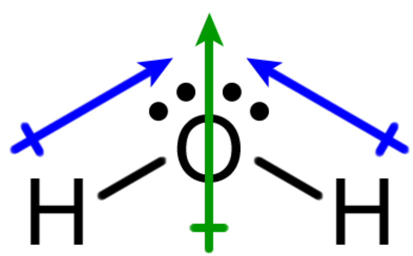
- No permanent *molecular dipole* moment in CO_2 (bond dipoles cancel).



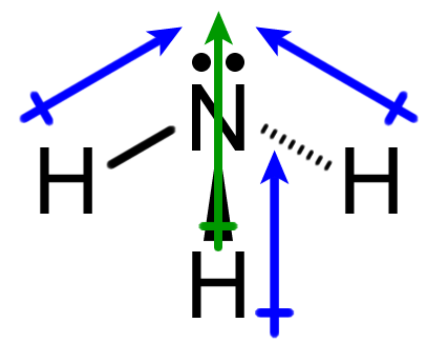
- No permanent *molecular dipole* moment in CCl_4 (bond diploes cancel).



- Permanent *molecular dipole* moment in HC*l*.



- Permanent *molecular dipole* moment in H_2O.



- Permanent *molecular dipole* moment in NH_3.

Chemical Bonding

Intermediate Forces – van der Waals Forces

- All types of van der Waals forces of attraction (instantaneous dipoles and permanent dipoles) between small molecules are weak. Between molecules with *longer chains of atoms*, giving a larger surface area over which many points of contact exist, *van der Waals forces are stronger*. For example, ethane (C_2H_6) is a *gas* at room temperature and pressure, while hexane (C_6H_{14}) is a *liquid* and octadecane ($C_{18}H_{38}$) is a *solid*.
- Another factor to consider when studying the strength of the van der Waals force of attraction is the *shape* of the molecule. *Elongated* molecules are more easily polarised, and hence have *stronger* van der Waals forces of attraction than molecules that are *compact* and *symmetrical*.

Chemical Bonding

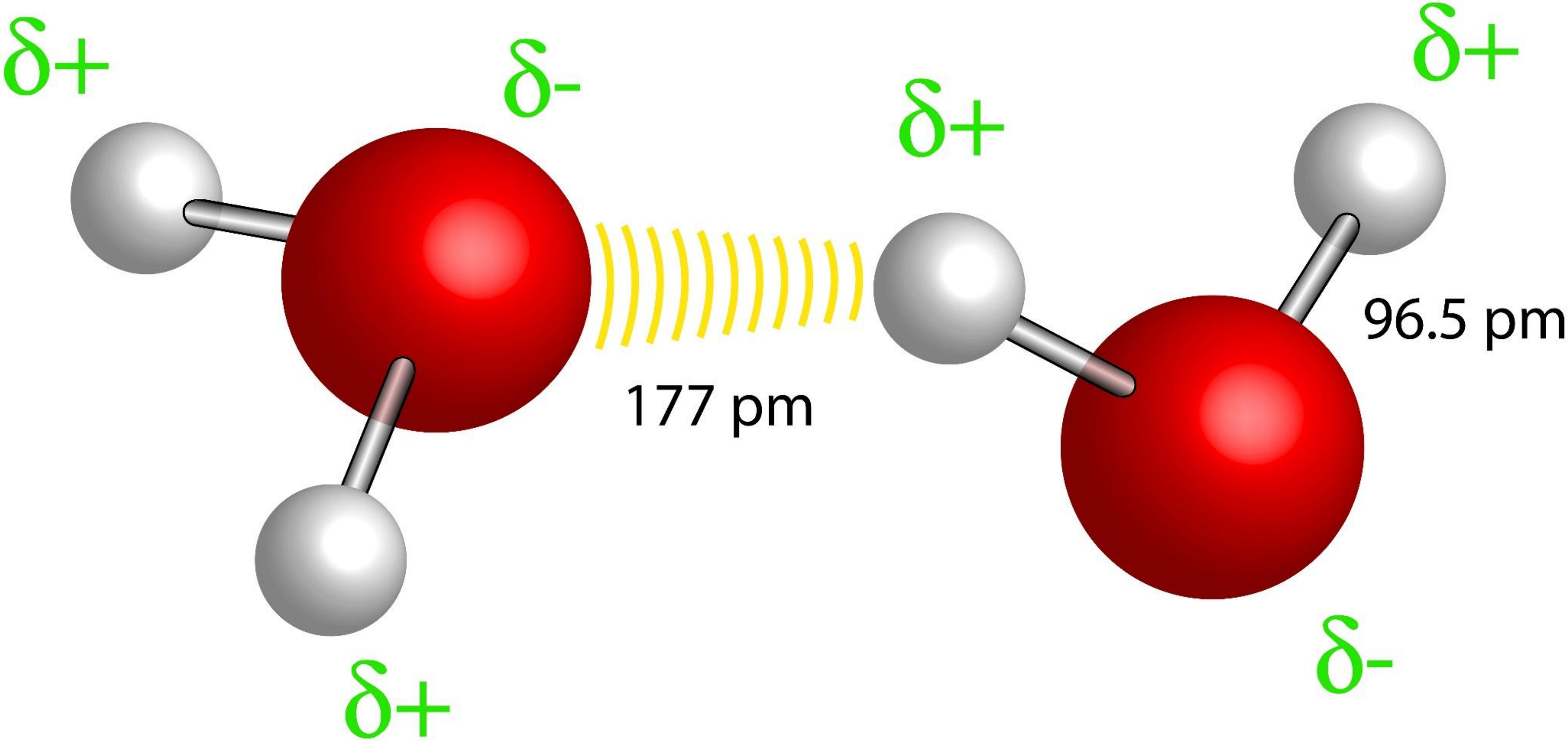
Intermediate Forces – Hydrogen Bonding

- A hydrogen bond is the attraction between the lone pair electrons of an electronegative atom and a *hydrogen* atom that is bonded to either *nitrogen*, *oxygen* or *fluorine*.

- The hydrogen bond is often described as a strong electrostatic dipole-dipole interaction. However, it also has some features in common with covalent bonding; *(i)* it is directional *(ii)* it is stronger than a van der Waals interaction *(iii)* molecules interact with a limited number of neighbouring molecules, which can be interpreted as a type of valency.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding



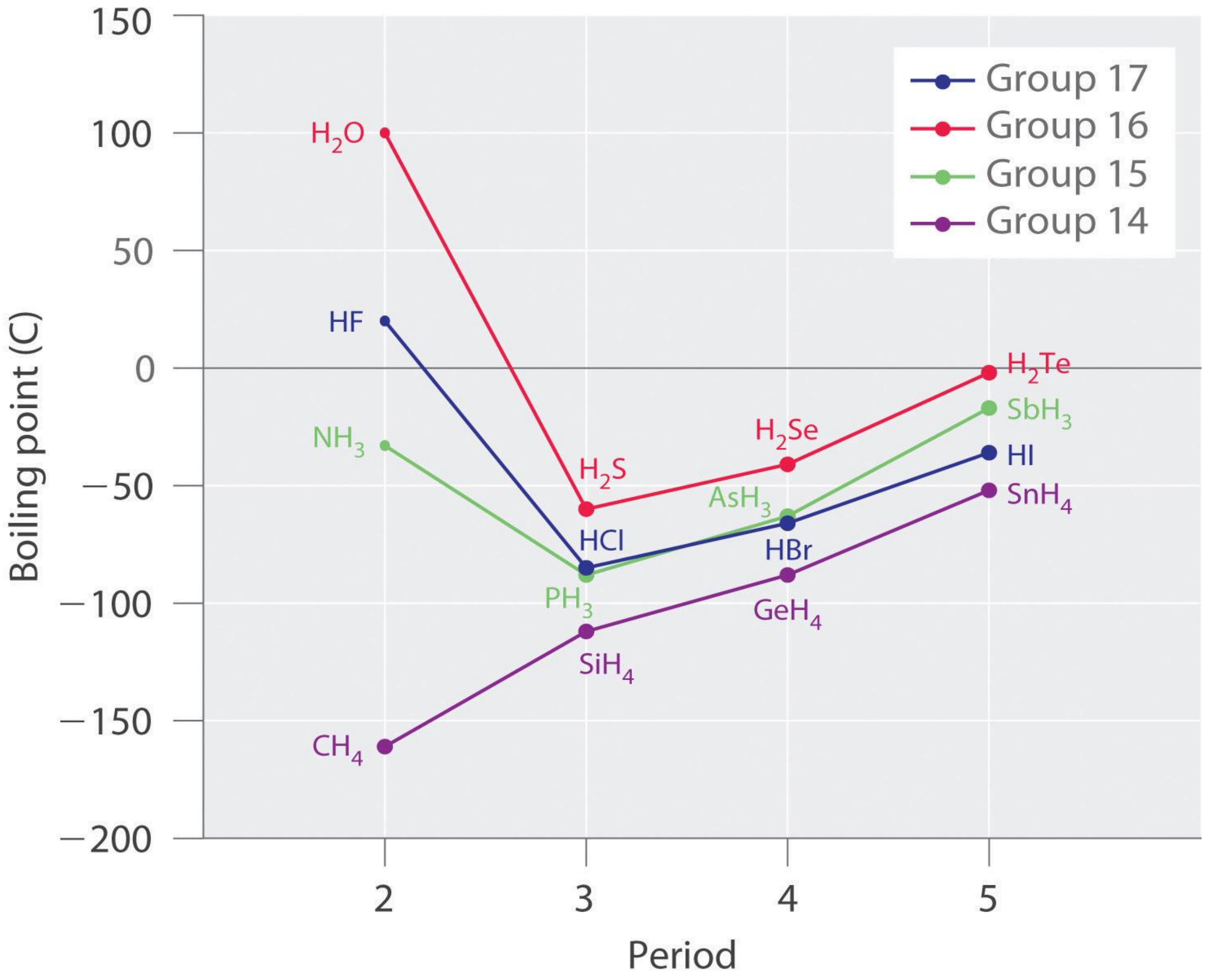
Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- Some evidence for the existence of the hydrogen bond can be seen when comparing the boiling points of *hydrogen fluoride* (HF), *water* (H_2O) and *ammonia* (NH_3) with the boiling points of the other hydrides in the same Groups of the Periodic Table.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding



Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- The boiling points of HF, H_2O and NH_3 are *much higher* than those of the other hydrides in the same groups.

- The molecules of HF, H_2O and NH_3 must be held together by intermolecular forces which are stronger than those between molecules of the other hydrides.

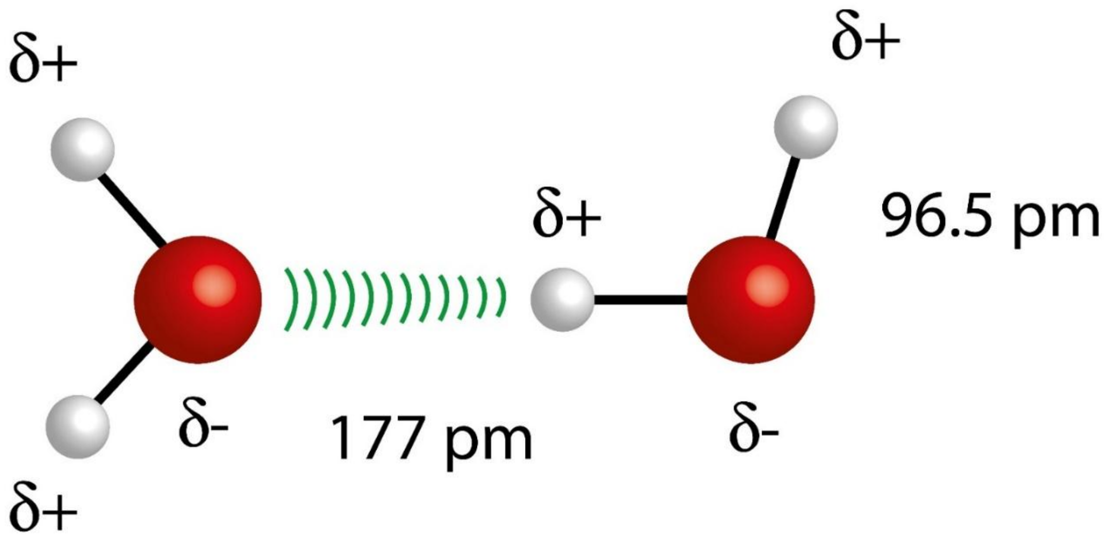
- Since fluorine, oxygen and nitrogen are the most electronegative elements in their Groups, the intermolecular forces of attraction are thought to be *hydrogen bonds*.

- Hydrogen bonds are estimated to be $^{1}/_{10}$ to $^{1}/_{20}$ the strength of a covalent bond.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- Hydrogen Bonding Between Water Molecules

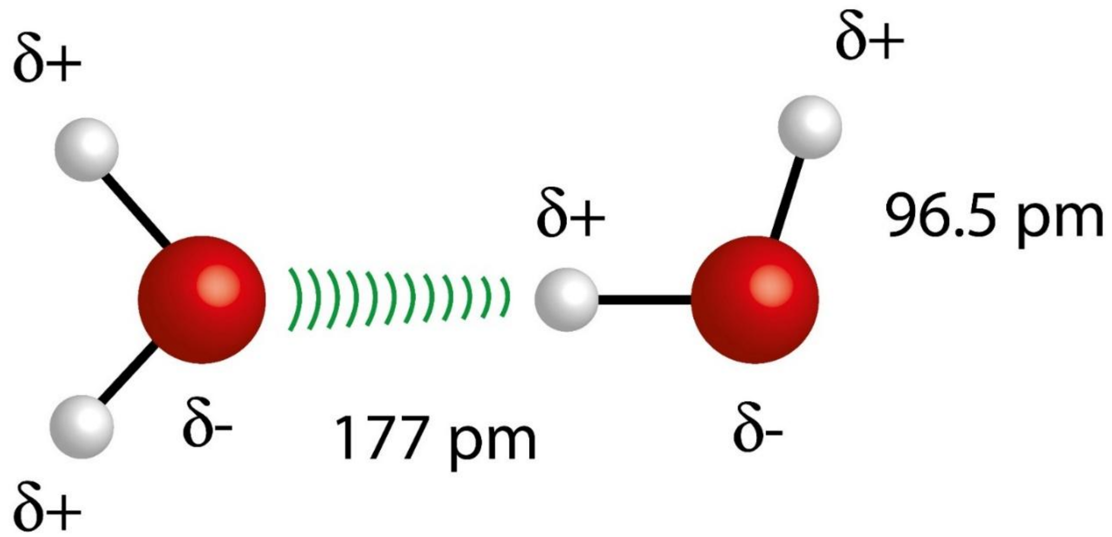


→ The bond between oxygen (electronegativity value = *3.5*) and hydrogen (electronegativity value = *2.1*) in water is *polar*.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- Hydrogen Bonding Between Water Molecules

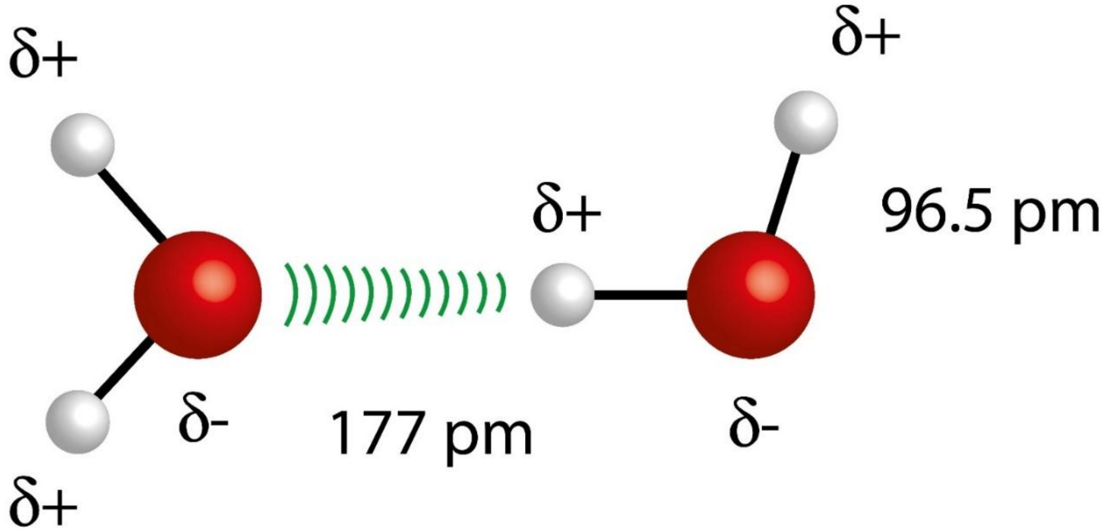


→ The oxygen atom attracts the bonding pair of electrons towards itself, away from the hydrogen atom.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- Hydrogen Bonding Between Water Molecules

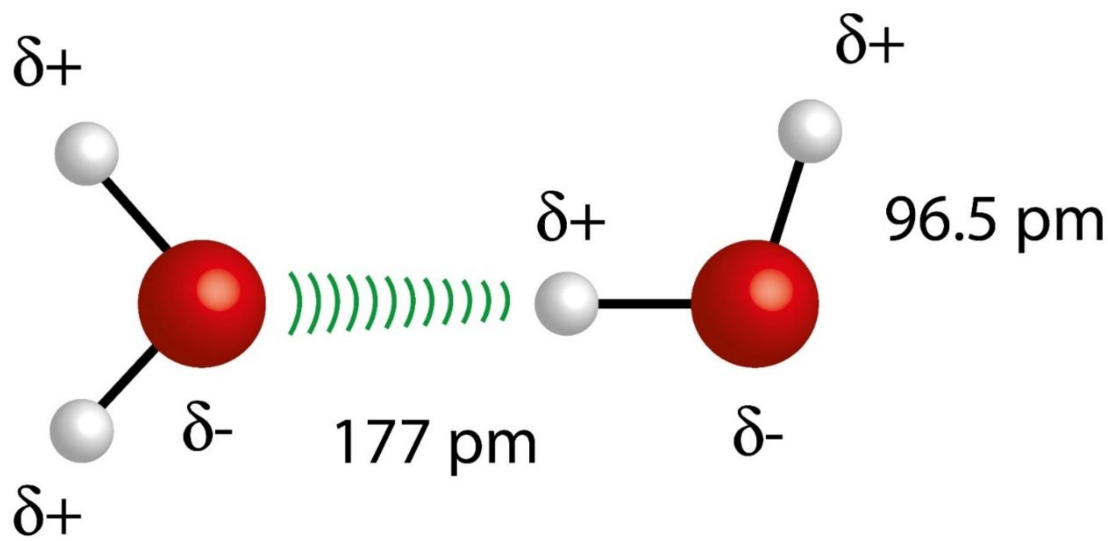


→ The oxygen atom gains a slight negative charge written δ– (delta negative) while the hydrogen atom gains a slight positive charge, written δ+ (delta positive).

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- Hydrogen Bonding Between Water Molecules

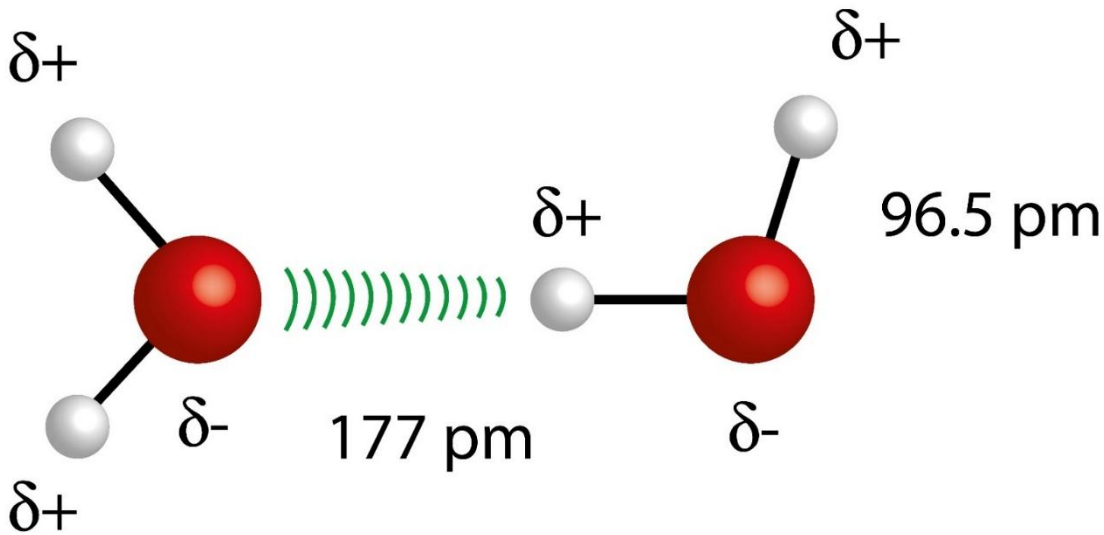


→ A weak electrostatic force of attraction between the lone pair electrons of the δ– oxygen atom of one water molecule and the δ+ hydrogen atom of another water molecule creates an intermolecular force of attraction between the two water molecules.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

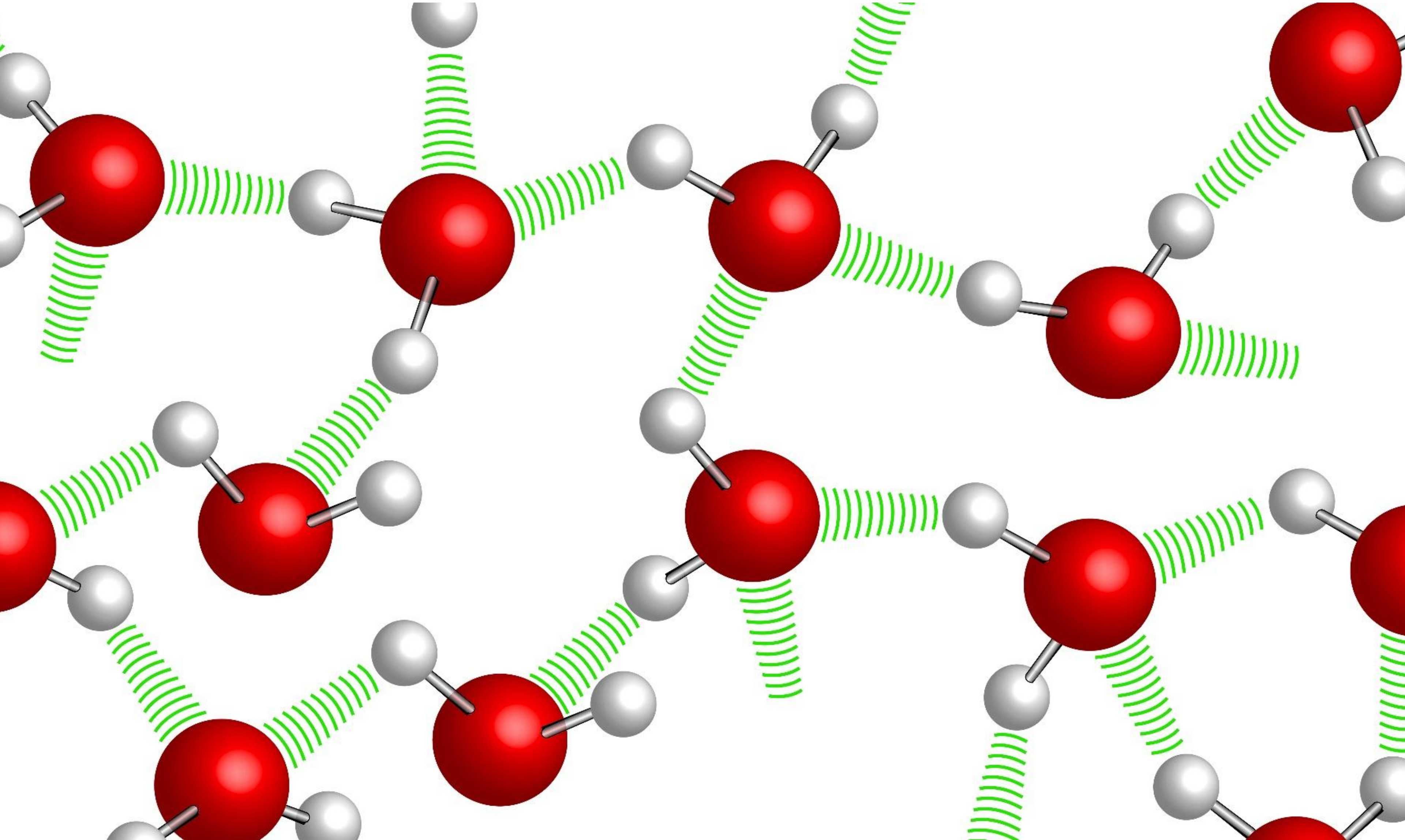
- Hydrogen Bonding Between Water Molecules



→ This special intermolecular force of attraction is called a *hydrogen bond*. It is stronger that the London dispersion force and permanent dipole-permanent dipole intermolecular force of attraction, and requires more energy to overcome, causing water to be a liquid at room temperature and pressure.

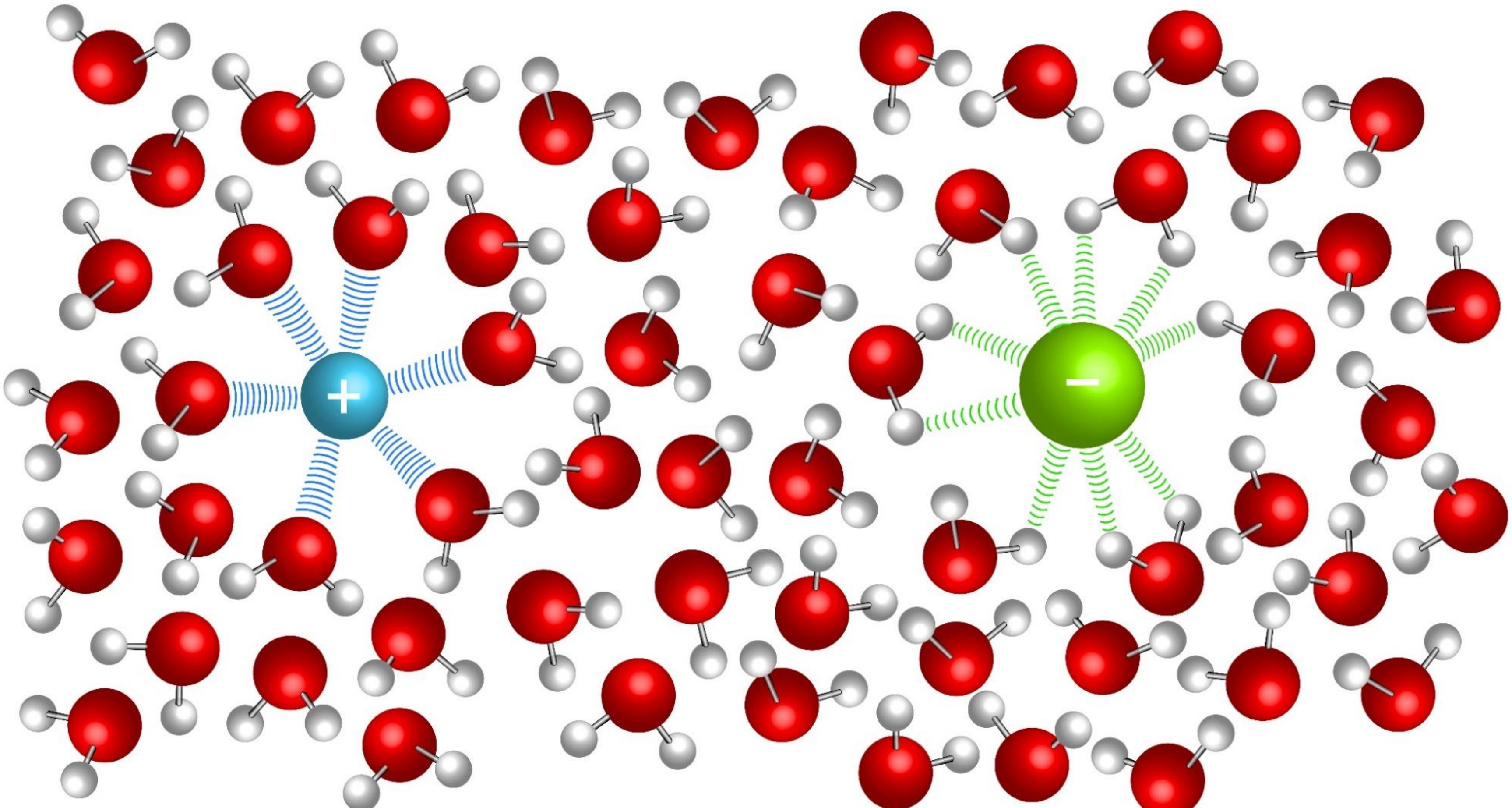
Chemical Bonding

Intermolecular Forces – Hydrogen Bonding



Chemical Bonding

Intermolecular Forces – Hydrogen Bonding



Chemical Bonding

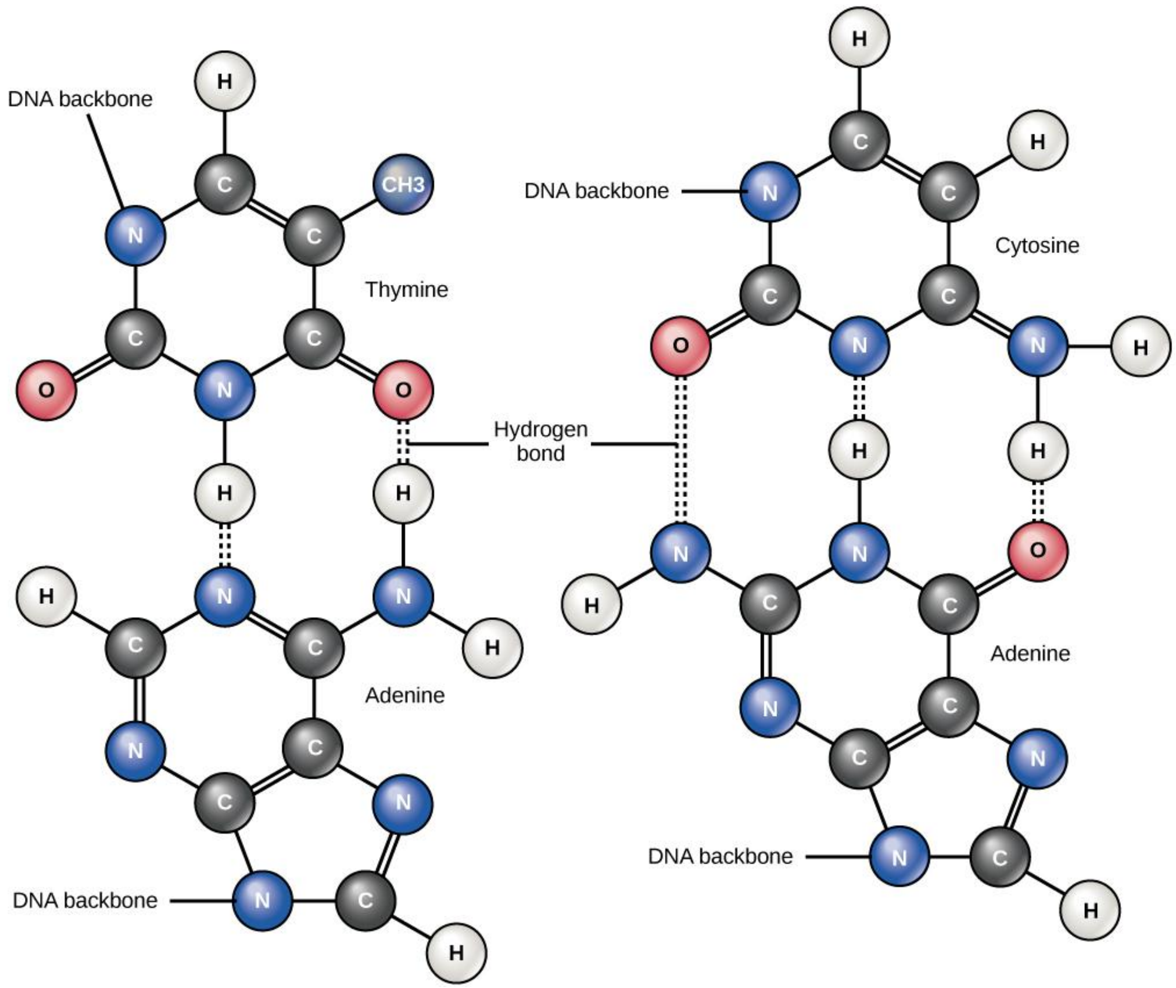
Intermolecular Forces – Hydrogen Bonding

- Ionic compounds dissolve in polar solvents such as water.
- The $\delta+$ *hydrogen atoms* of the water molecules are attracted towards the *negative anions* in the ionic lattice.
- The $\delta-$ *oxygen atoms* of the water molecules are attracted towards the *positive cations* in the ionic lattice.

Chemical Bonding

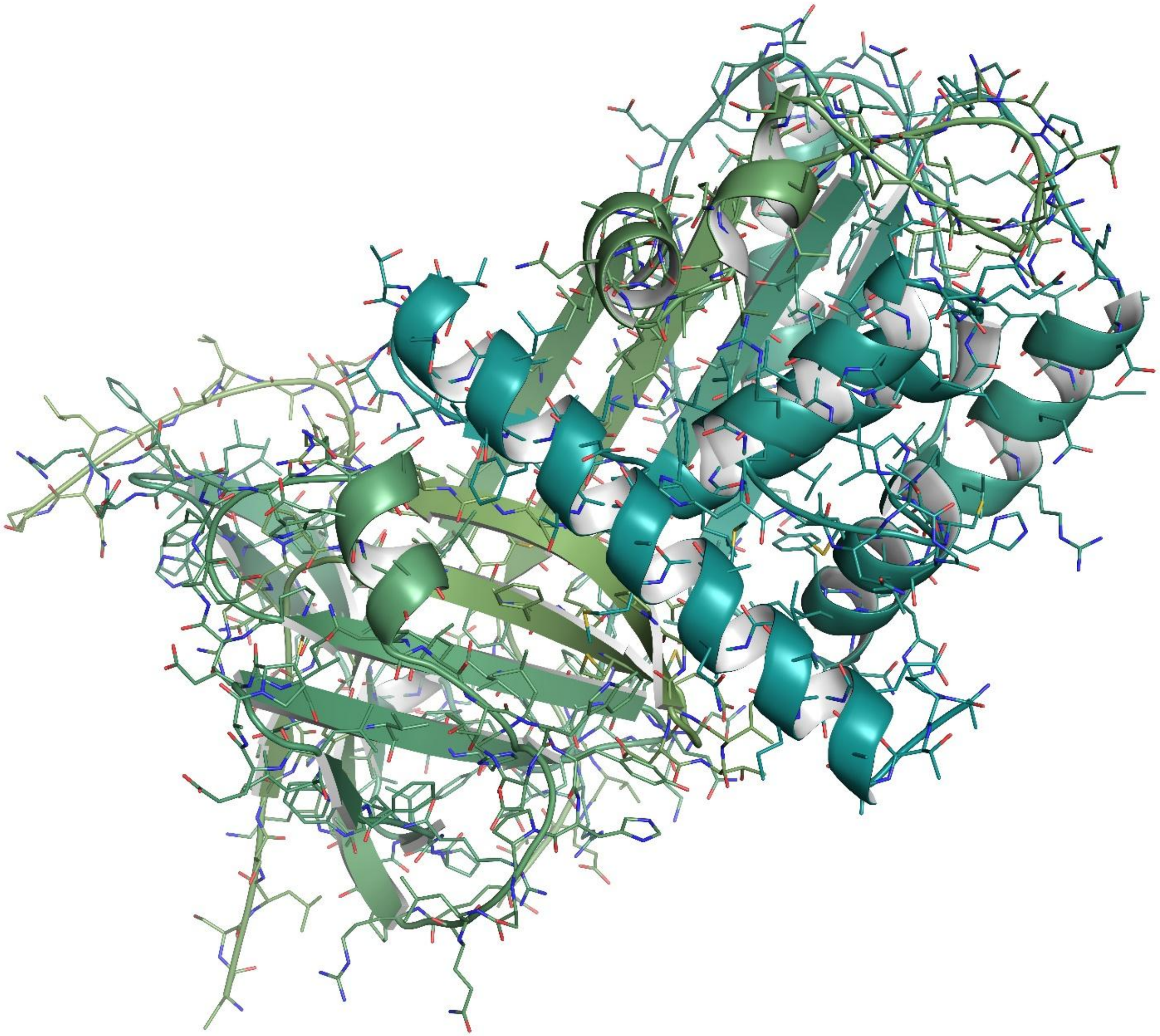
Intermolecular Forces – Hydrogen Bonding

- Hydrogen Bonding Between DNA Base Pairs



Chemical Bonding

Intermolecular Forces – Hydrogen Bonding



Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

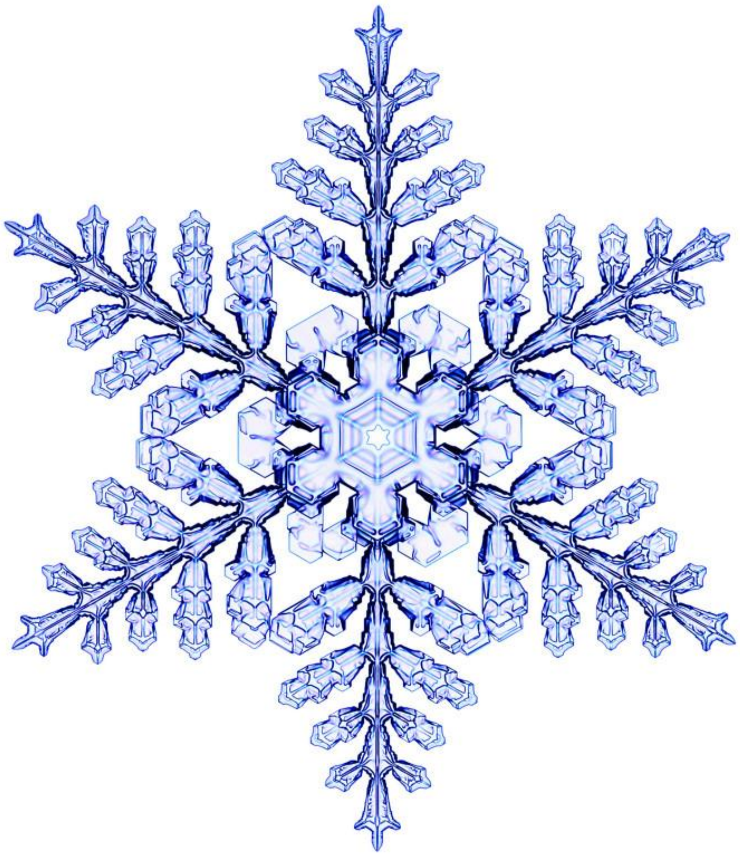
- Hydrogen bonds occur between $\delta-$ N–H $\delta+$ and $\delta+$ C=O $\delta-$ regions within proteins. Hydrogen Bonding contributes to the secondary structures (α-helix and β-pleated sheets) found in proteins.

- At elevated temperatures, hydrogen bonds within proteins are weakened and / or broken. This causes the protein to lose its unique three-dimensional shape, a process that biochemists refer to as *denaturation*.

- Due to a change in the shape of their active site, enzymes lose their catalytic activity when they are denatured.

Chemical Bonding

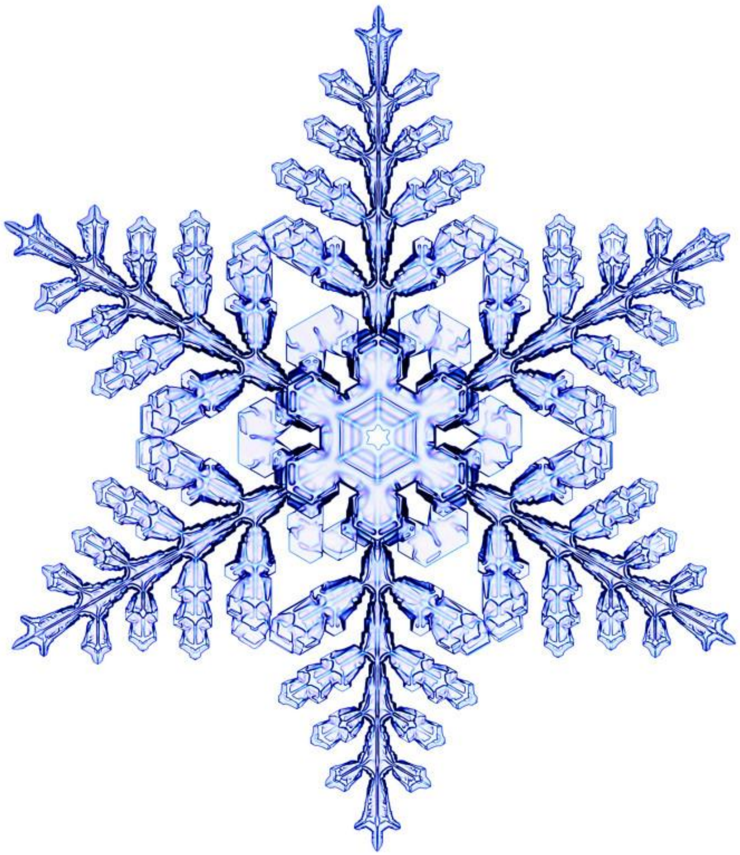
Intermolecular Forces – Hydrogen Bonding



- All snowflakes have the same basic hexagonal shape, although they vary in their fine structure.
- The hexagonal shape is due to the arrangement in which the water molecules crystallise, with the $\delta+$ hydrogen atoms being attracted towards the lone pairs of electrons in the valence shell of the $\delta-$ oxygen atoms.

Chemical Bonding

Intermediate Forces – Hydrogen Bonding

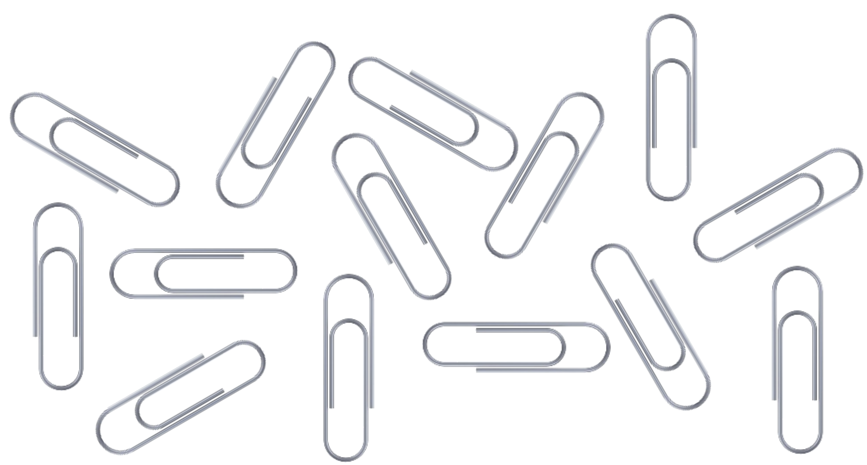


- Hence, the shape of the water molecule (based upon valence shell electron pair repulsion theory) influences the overall shape of the snowflake.

Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- An analogy of a polar covalent solid, *e.g.* glucose, dissolving in a polar solvent, *e.g.* water. Molecules of the polar covalent solid are represented by bar magnets, and the polar water molecules are represented by paper clips.



Chemical Bonding

Intermolecular Forces – Hydrogen Bonding

- In the solid state, polar molecules interact to form an ordered, crystalline arrangement through *hydrogen bonding*. The polar molecules pack in such a way that the *slightly positive* (δ+) region of one molecule is *adjacent* to the *slightly negative* (δ–) region of another molecule, just as magnets will arrange themselves with the *north pole* of one magnet *adjacent* to the *south pole* of another magnet.

Chemical Bonding

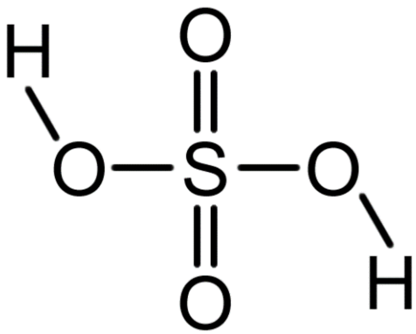
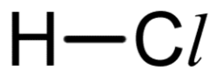
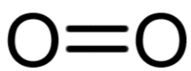
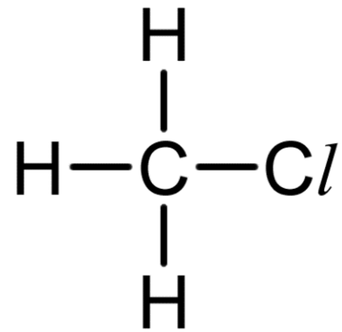
Intermolecular Forces – Hydrogen Bonding

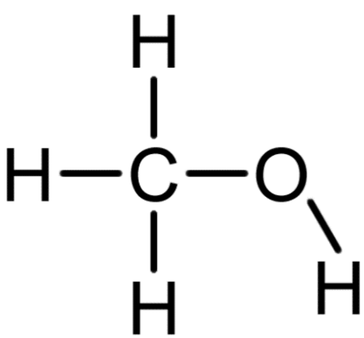
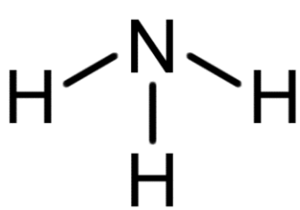
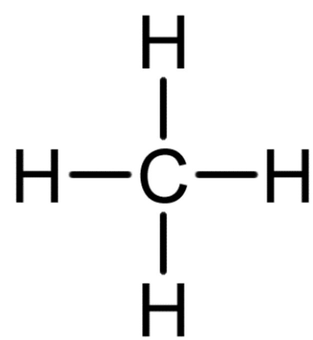
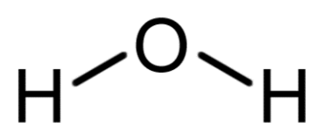
- When dissolving, hydrogen bonds between polar molecules of the solid are broken, and new hydrogen bonds form between polar molecules of the solid and polar water molecules.

Chemical Bonding

Intermolecular Forces of Attraction

- Which intermolecular force of attraction is most significant for each one of the following chemicals?

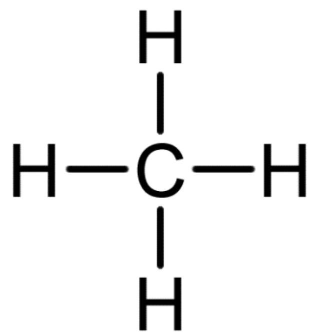





Chemical Bonding

Intermolecular Forces of Attraction

- Which intermolecular force of attraction is most significant for each one of the following chemicals?

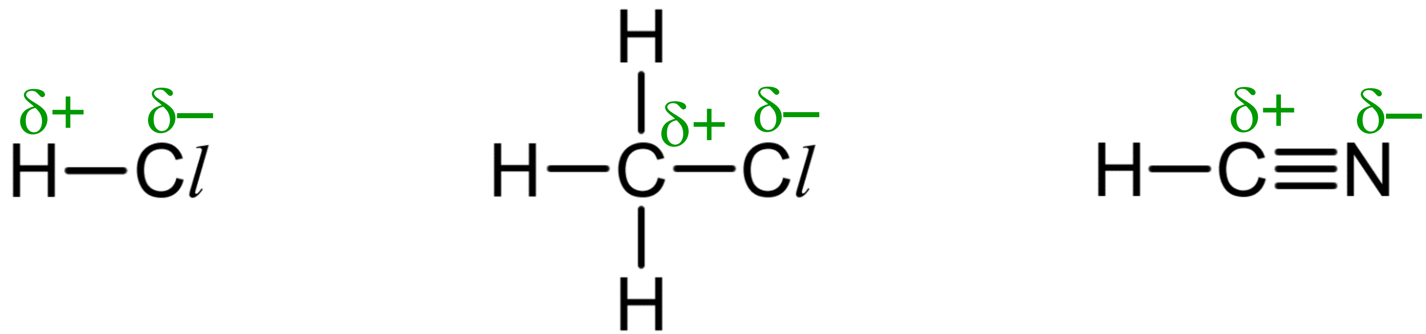


- In each molecule, there is *no significant difference* between the *electronegativity values* of the different elements (the difference is < 0.5) so there are *no permanent dipoles* / *no polar covalent bonds*.
- Molecules of O_2, CH_4, and C_2H_2 interact through **instantaneous dipole-induced dipole** intermolecular forces of attraction known as ***London dispersion forces***.

Chemical Bonding

Intermolecular Forces of Attraction

- Which intermolecular force of attraction is most significant for each one of the following chemicals?

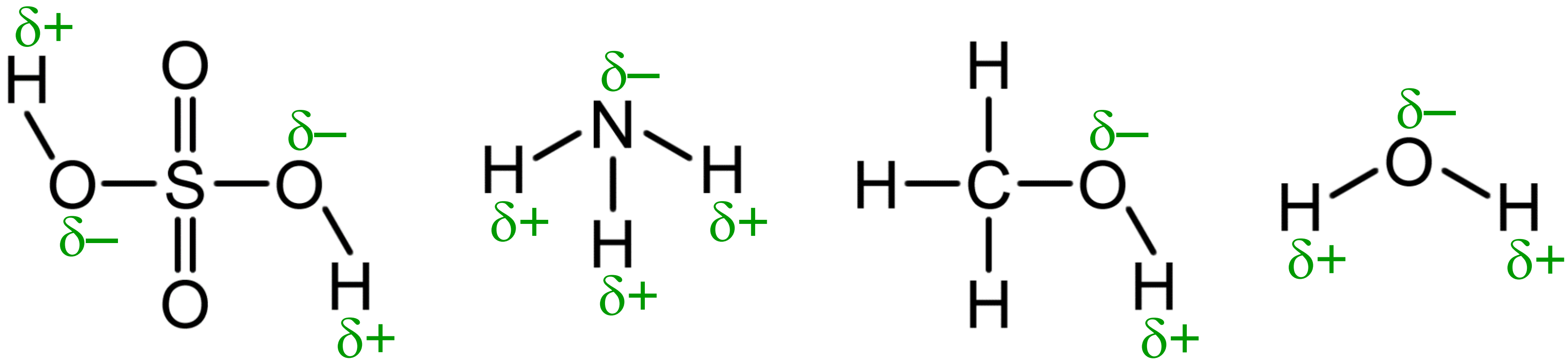


- In each molecule, there *is* a *significant difference* between the *electronegativity values* of the different elements (the difference is ≥ 0.5) which causes the formation of *permanent dipoles* / *polar covalent bonds*.
- Molecules of HC*l*, CH_3Cl and HCN, interact through ***permanent dipole-permanent dipole*** intermolecular forces of attraction.

Chemical Bonding

Intermediate Forces of Attraction

- Which intermolecular force of attraction is most significant for each one of the following chemicals?



- In each molecule, there *is* a *significant difference* between the *electronegativity values* of the different elements, which causes the formation of *permanent dipoles* / *polar covalent bonds*.
- In each molecule, a *highly electronegative element* is bonded to *hydrogen*. H_2SO_4, NH_3, CH_3OH and H_2O interact through an intermolecular force of attraction known as ***hydrogen bonds***.

Chemical Bonding

Intermolecular Forces of Attraction

Relative Strengths of Intermolecular Forces of Attraction						
Stronger	Hydrogen Bonds	**>**	Permanent Dipole – Permanent Dipole	**>**	London Dispersion Forces	**Weaker**

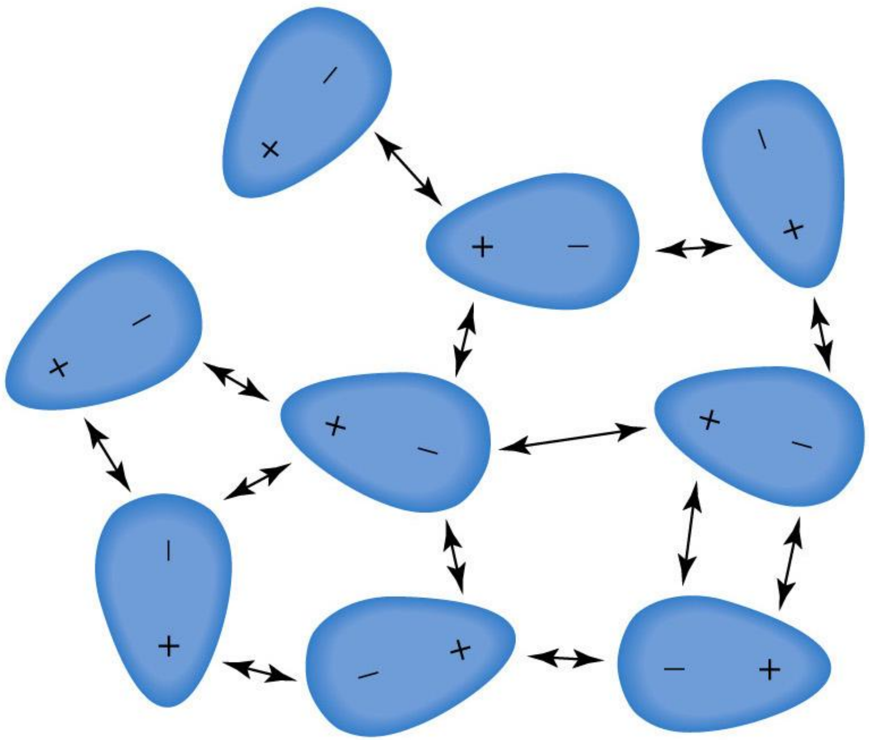
- Hydrogen Bonds: Exist between molecules that contain δ+ H–F δ– or δ+ H–O δ– or δ+ H–N δ– bonds.

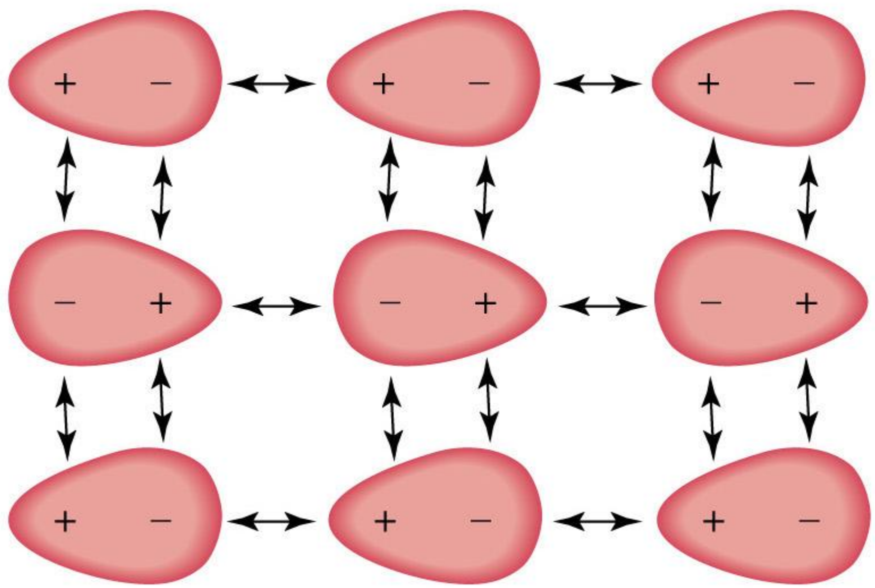
- Permanent Dipole – Permanent Dipole Interactions: Exist between molecules in which the atoms directly bonded together differ in their electronegativity values by *0.5 or more*, *e.g.* δ+ H–C*l* δ–.

- Instantaneous Dipole-Induced Dipole (London Dispersion Forces): Exist between molecules in which the atoms directly bonded together differ in their electronegativity values by *0.4 or less*, *e.g.* Cl_2, O_2 and most *hydrocarbons* (*e.g.* CH_4 and C_2H_6).

Chemical Bonding

Intermolecular Forces of Attraction



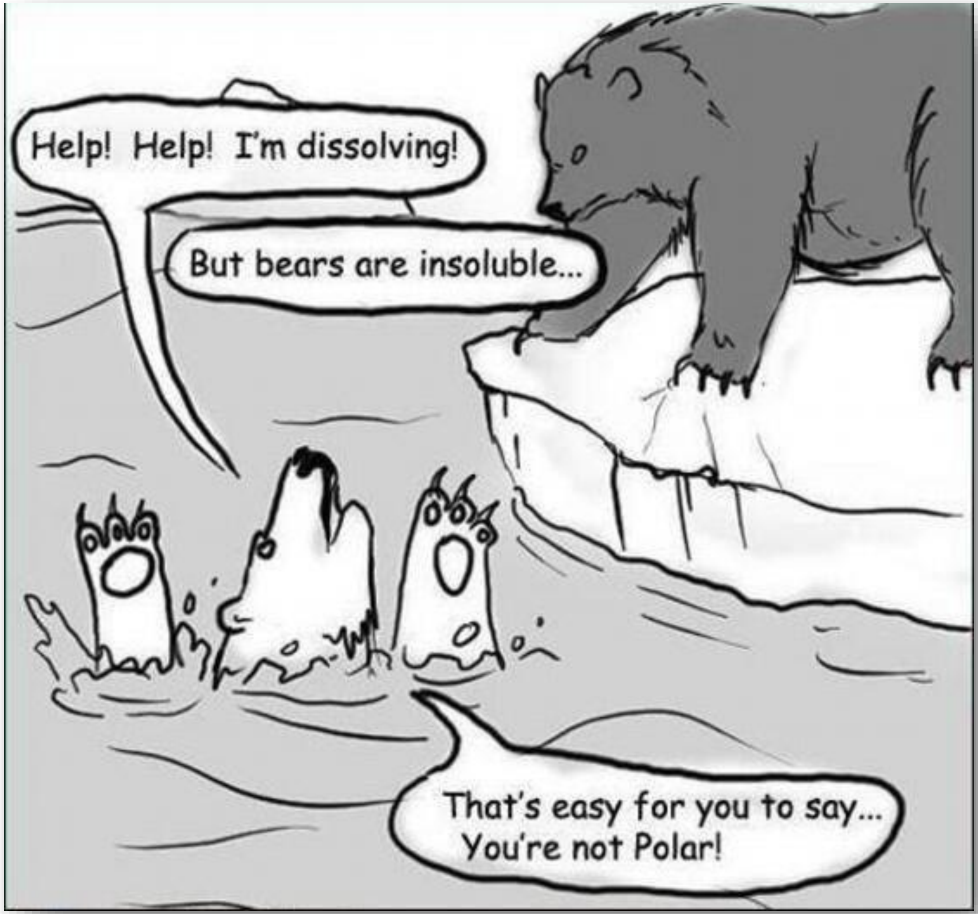


- Molecules with instantaneous dipoles – induced dipoles (London dispersion forces).

- Molecules with permanent dipoles.

Chemical Bonding

Intermolecular Forces of Attraction



- Click here to return to the start of the chapter on intermolecular forces of attraction.

Online Search

- What is the structure of *aspirin*? What are the chemical and physical properties of aspirin? How are these properties related to its structure and bonding?

- What is the structure of *glucose*? What are the chemical and physical properties of glucose? How are these properties related to its structure and bonding?

- What is the structure of caffeine? What are the chemical and physical properties of caffeine? How are these properties related to its structure and bonding?

Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

14. What are the *general properties* of chemicals with *giant molecular structures*?

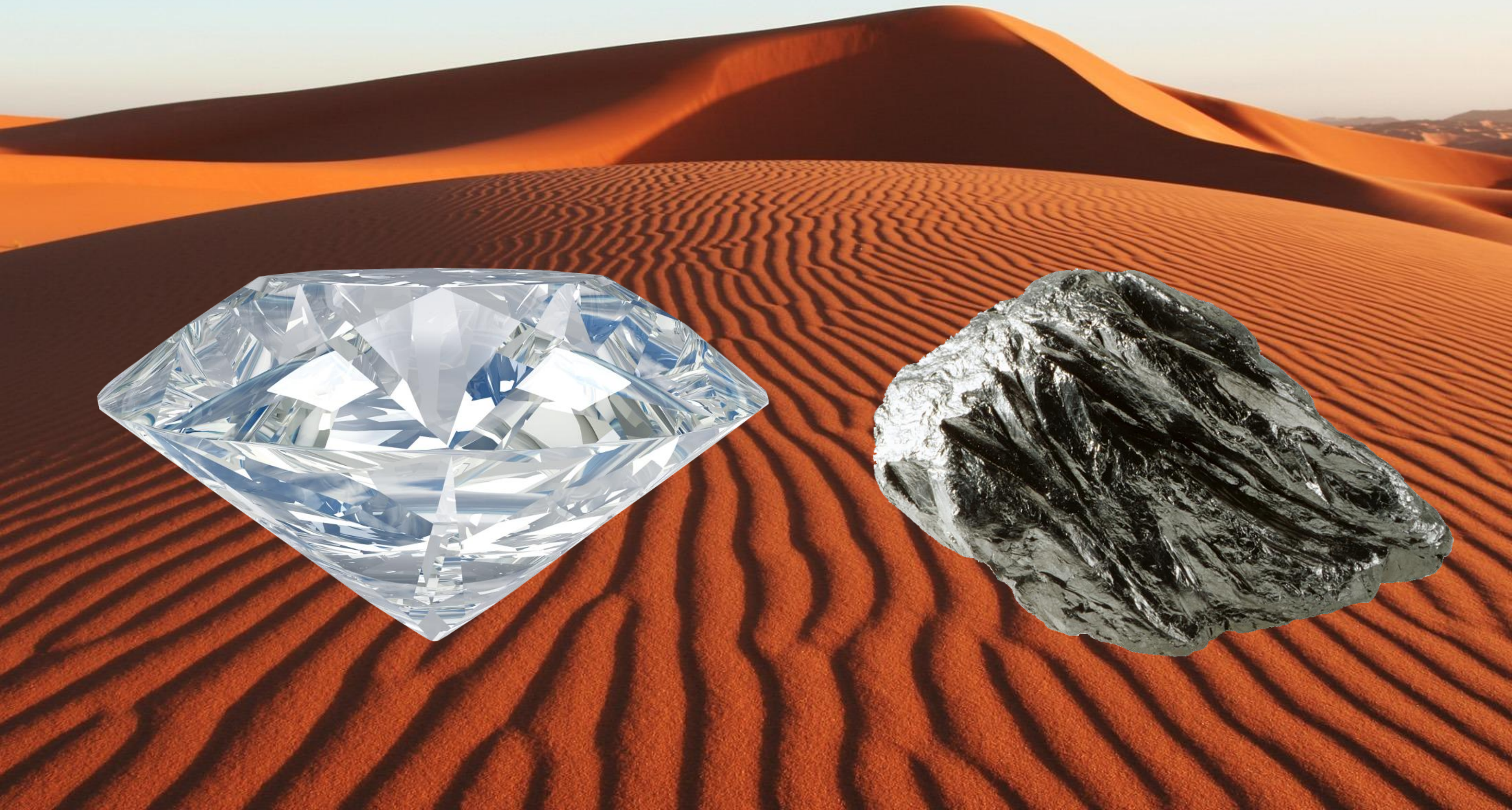
Note: The terms *giant molecular structure*, *giant covalent structure* and *macromolecular structure* are synonymous.

Main Menu

Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

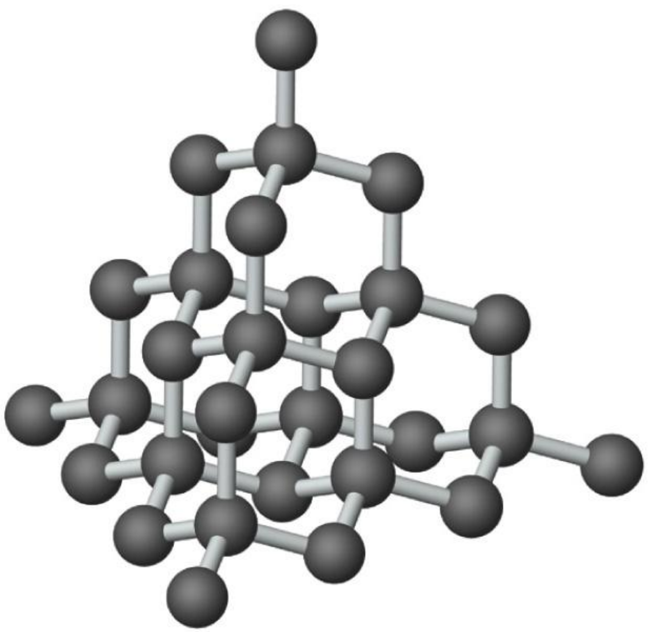
Diamond – C   Sand – SiO_2 Graphite – C



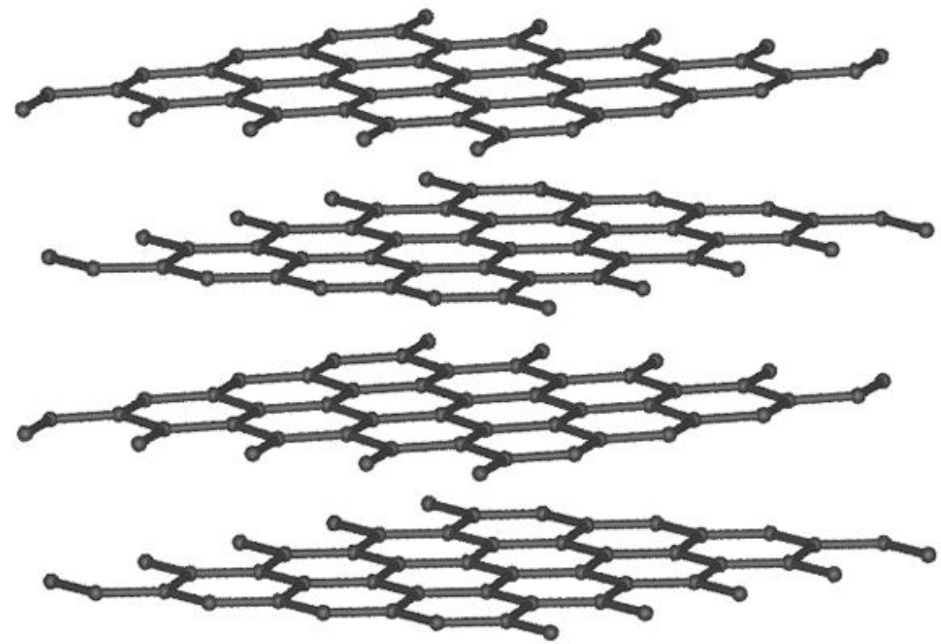
Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

Diamond – C



Graphite – C

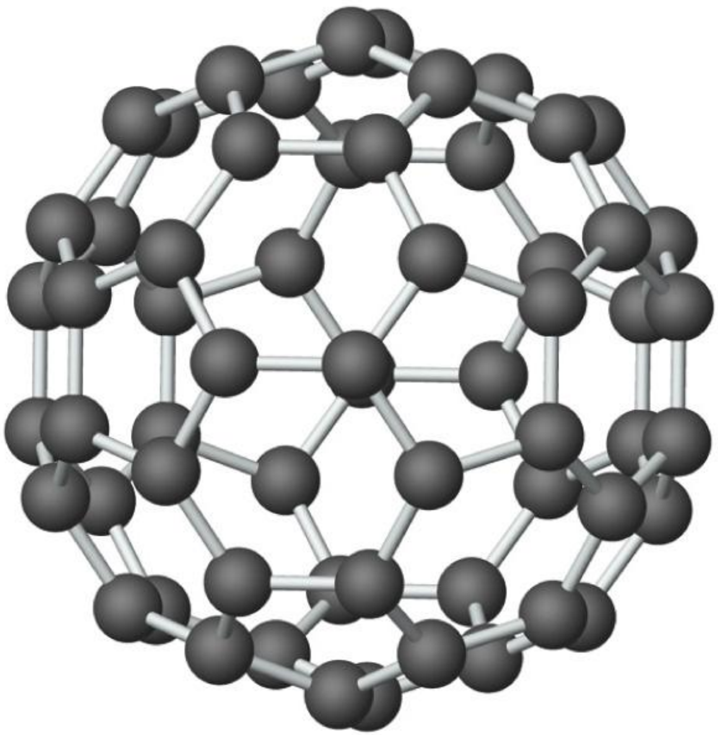


- Diamond and graphite are both *allotropes* of carbon, *i.e.* forms of the *same chemical element* in which the atoms are bonded together in *different arrangements*.

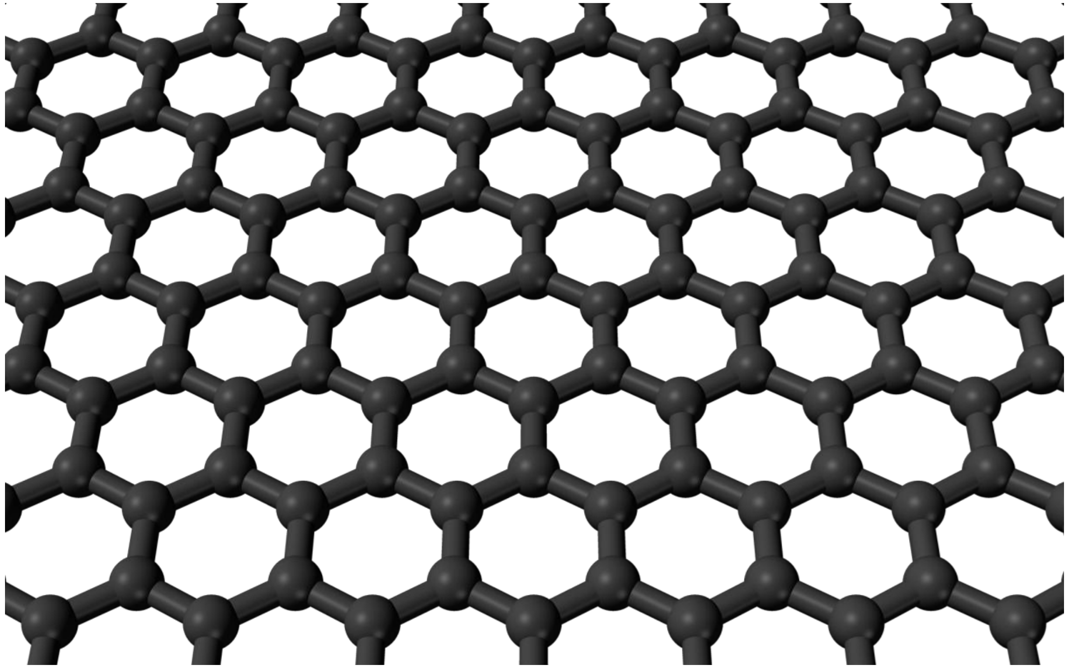
Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

Buckminsterfullerene
(Nobel Prize in Chemistry, 1996).



Graphene
(Nobel Prize in Physics, 2010).

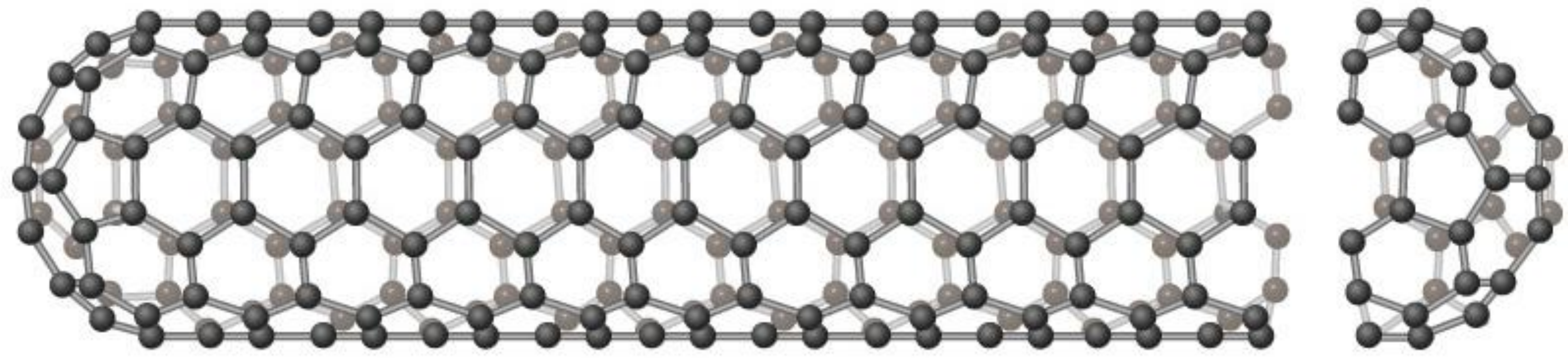


- *Buckminsterfullerene* (formula: C_{60}) and *Graphene* (a single layer of carbon atoms) are other interesting allotropes of carbon.

Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

- *Buckytube* or *carbon nanotube*.



Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

- Giant covalent / molecular structures are composed of many millions of non-metallic atoms all covalently bonded together in a single arrangement, *e.g*. diamond.

- Chemicals with a giant molecular structure have *very high melting points and boiling points* (sometimes even greater than those of ionic compounds and metals). This is because the entire structure is composed of atoms of a non-metallic element all joined together by *strong covalent bonds which require a large amount of thermal energy to weaken / break*. Because the structure is not composed of small covalent molecules, there are no weak inter-molecular forces of attraction present.

Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

- Giant molecular structures tend to be *extremely hard* (the hardest naturally occurring substance is diamond) because the entire structure is held together by *strong covalent bonds which require a large amount of energy to break*.

Note: Graphite is relatively soft because it is composed of layers of carbon atoms, with weak intermolecular forces of attraction between them, that can slide over each other.

Chemical Bonding

Covalent Bonding – Giant Molecular – Properties

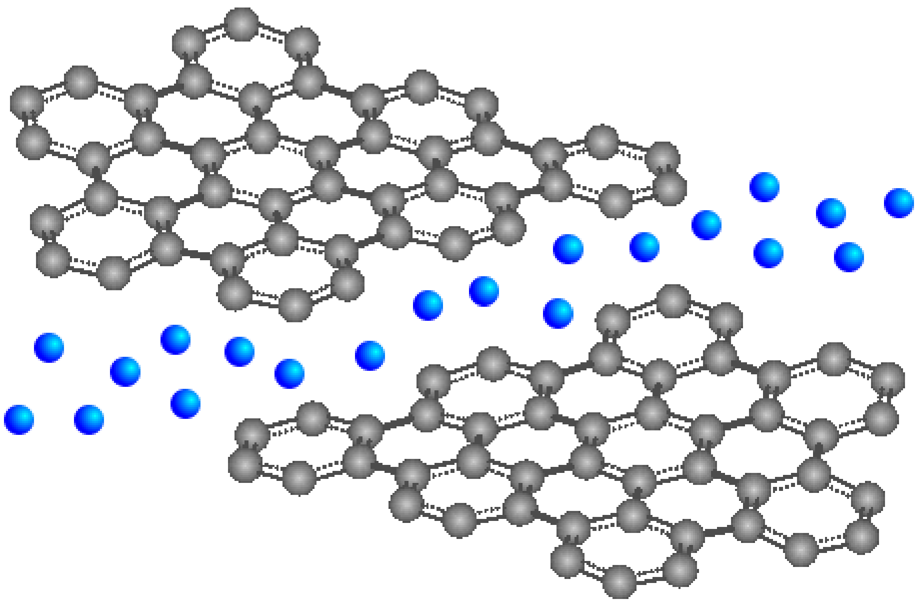
• Giant molecular structures are *insoluble* in both polar solvents and non-polar solvents. This is because any intermolecular forces of attraction that are formed between atoms of the giant molecular structure and molecules of the solvent are not strong enough to break / overcome the strong covalent bonds within the giant molecular structure.

• Giant molecular structures tend to be *electrical insulators*, *i.e*. they do not conduct electricity in any form because they do not contain any mobile charge carrying particles (electrons or ions). **Note:** Graphite does conduct electricity because it contains *delocalised* (mobile) *electrons* between the layers of carbon atoms.

Chemical Bonding

Covalent Bonding – Giant Covalent – Graphite

Special Properties of Graphite

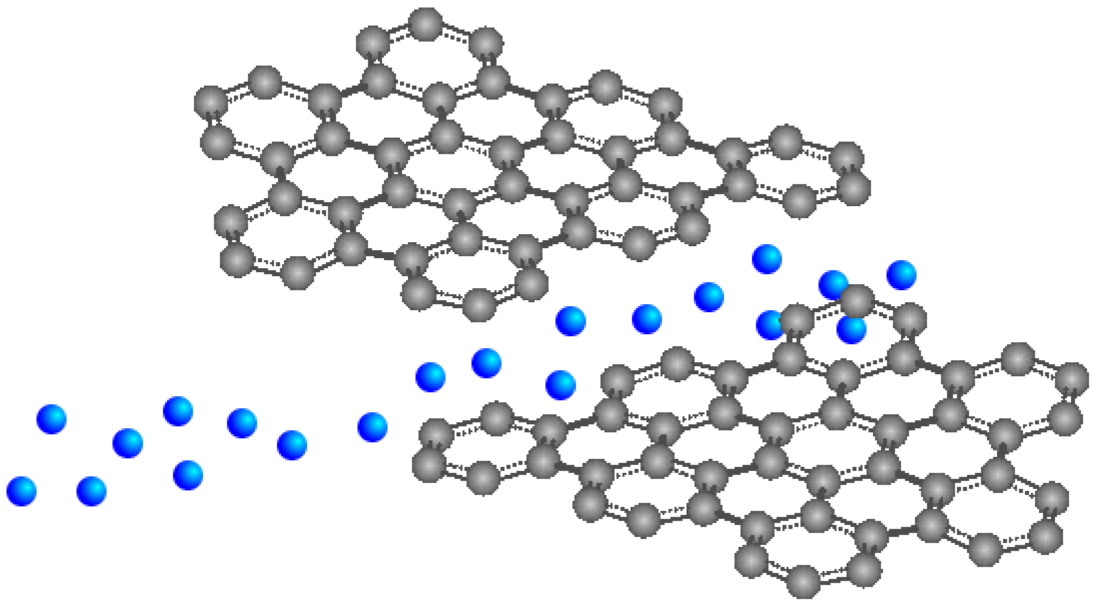


- Graphite is composed of many carbon atoms covalently bonded together in *hexagonal layers*.
- The shape around each carbon atom is *trigonal planar*.
- In graphite, many hexagonal layers of carbon atoms are stacked one-on-top-of-another.

Chemical Bonding

Covalent Bonding – Giant Covalent – Graphite

Special Properties of Graphite

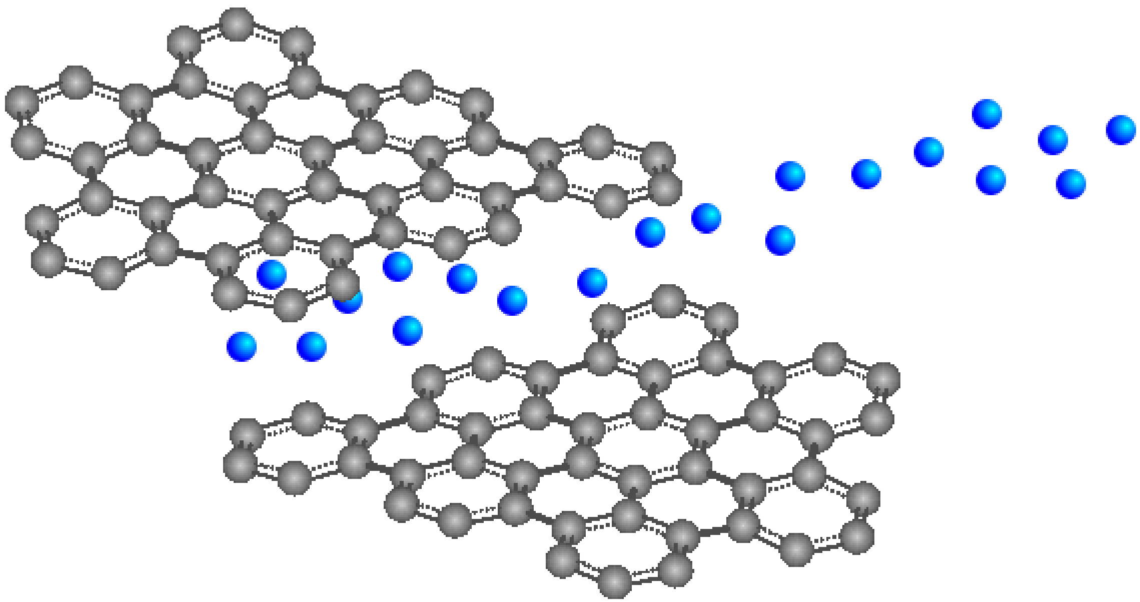


- Each carbon atom has four valence electrons.
- Three valence electrons are used to make covalent bonds to neighbouring carbon atoms within the same layer.

Chemical Bonding

Covalent Bonding – Giant Covalent – Graphite

Special Properties of Graphite

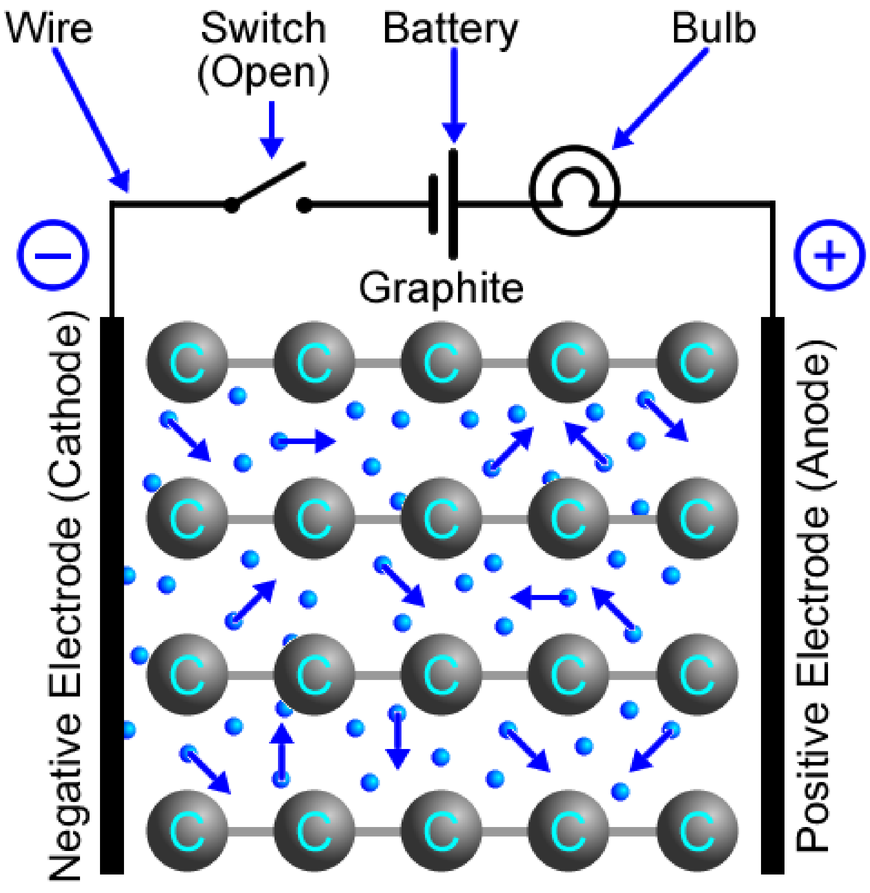


- The forth valence electron of each carbon atom is not involved in forming covalent bonds. Hence, each carbon atom in graphite has a *single unbonded valence electron* which is *delocalised* between the layers of carbon atoms.

Chemical Bonding

Covalent Bonding – Giant Covalent – Graphite

- Because graphite contains *mobile* or *delocalised* electrons, it is able to *conduct electricity* when solid. This is an unusual property for a substance that is classified as a non-metal!

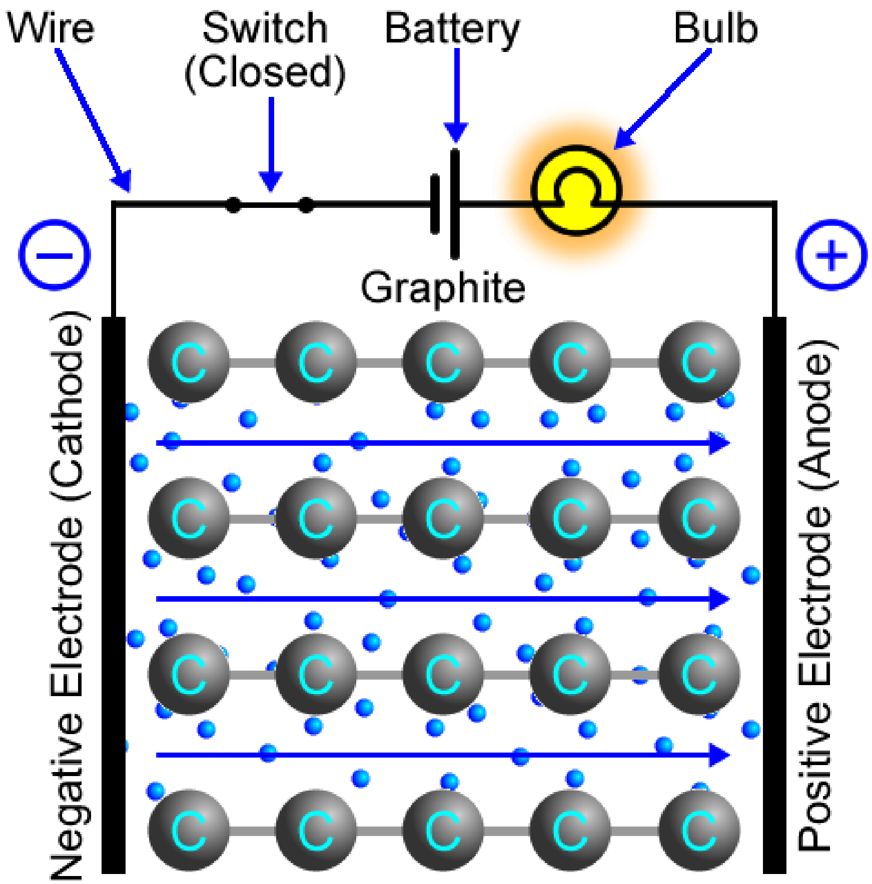


- When the circuit is closed, the *mobile* or *delocalised* electrons within the graphite are all repelled away from the negative electrode (cathode) and attracted towards the positive electrode (anode).

Chemical Bonding

Covalent Bonding – Giant Covalent – Graphite

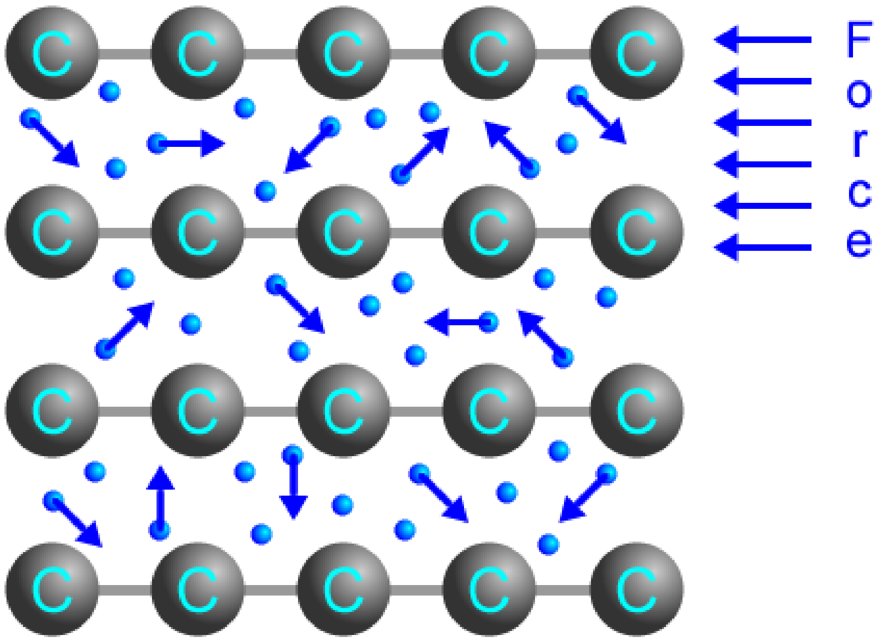
- Because graphite contains *mobile* or *delocalised* electrons, it is able to *conduct electricity* when solid. This is an unusual property for a substance that is classified as a non-metal!



- When the circuit is closed, the *mobile* or *delocalised* electrons within the graphite are all repelled away from the negative electrode (cathode) and attracted towards the positive electrode (anode).

Chemical Bonding

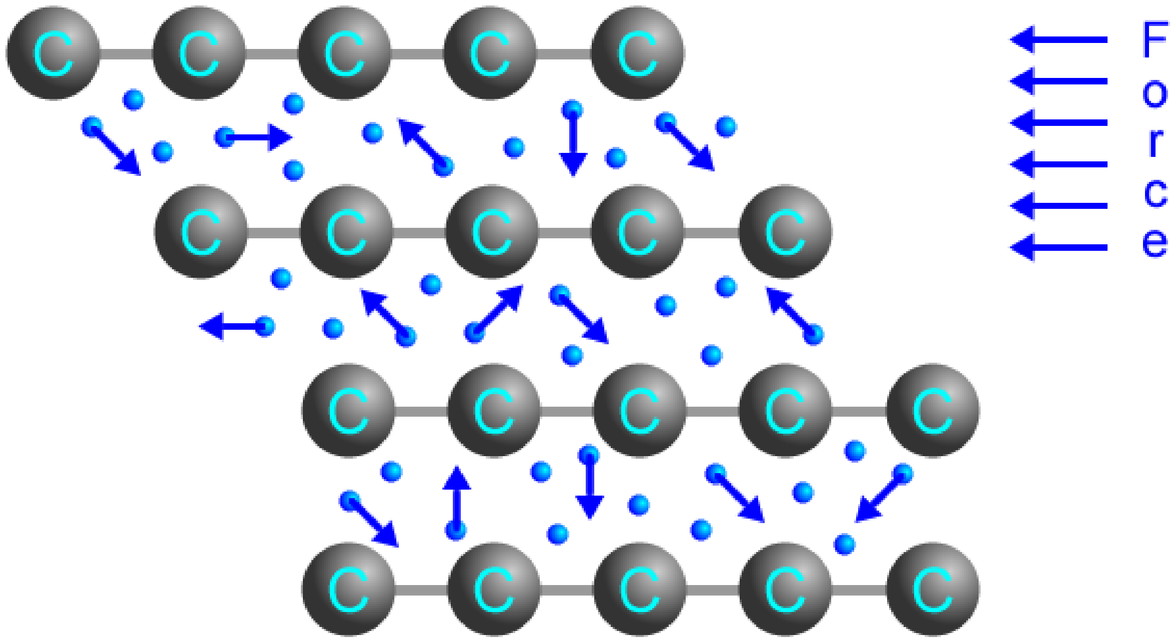
Covalent Bonding – Giant Covalent – Graphite



- Because the intermolecular forces between the layers of carbon atoms are relatively weak, the layers of carbon atoms in graphite are able to *slide over each other* when a force is applied. This gives graphite a "slippery" texture which allows it to be used as a *solid lubricant* in some areas of engineering.

Chemical Bonding

Covalent Bonding – Giant Covalent – Graphite



- Because the intermolecular forces between the layers of carbon atoms are relatively weak, the layers of carbon atoms in graphite are able to *slide over each other* when a force is applied. This gives graphite a “slippery” texture which allows it to be used as a *solid lubricant* in some areas of engineering.

Chemical Bonding

Covalent Bonding – Properties – Summary

- Substances with a *simple molecular structure* tend to have low melting points and boiling points. They tend to be soft. They do not dissolve in polar solvents but are soluble in non-polar solvents. They are electrical insulators and do not conduct electricity in any form (solid, liquid or solution).

- Substances with a *giant molecular structure* tend to have very high melting points and boiling points. They tend to be very hard. They do not dissolve in either polar solvents or non-polar solvents. They are electrical insulators and do not conduct electricity in any form. **Note:** Graphite is an exception to some of these rules.

Online Search

- What are the industrial applications of *diamond*? How are these applications related to its structure and bonding?

- What are the industrial applications of *graphite*? How are these applications related to its structure and bonding?

- What are the industrial applications of *silicon carbide*? How are these applications related to its structure and bonding?

Chemical Bonding

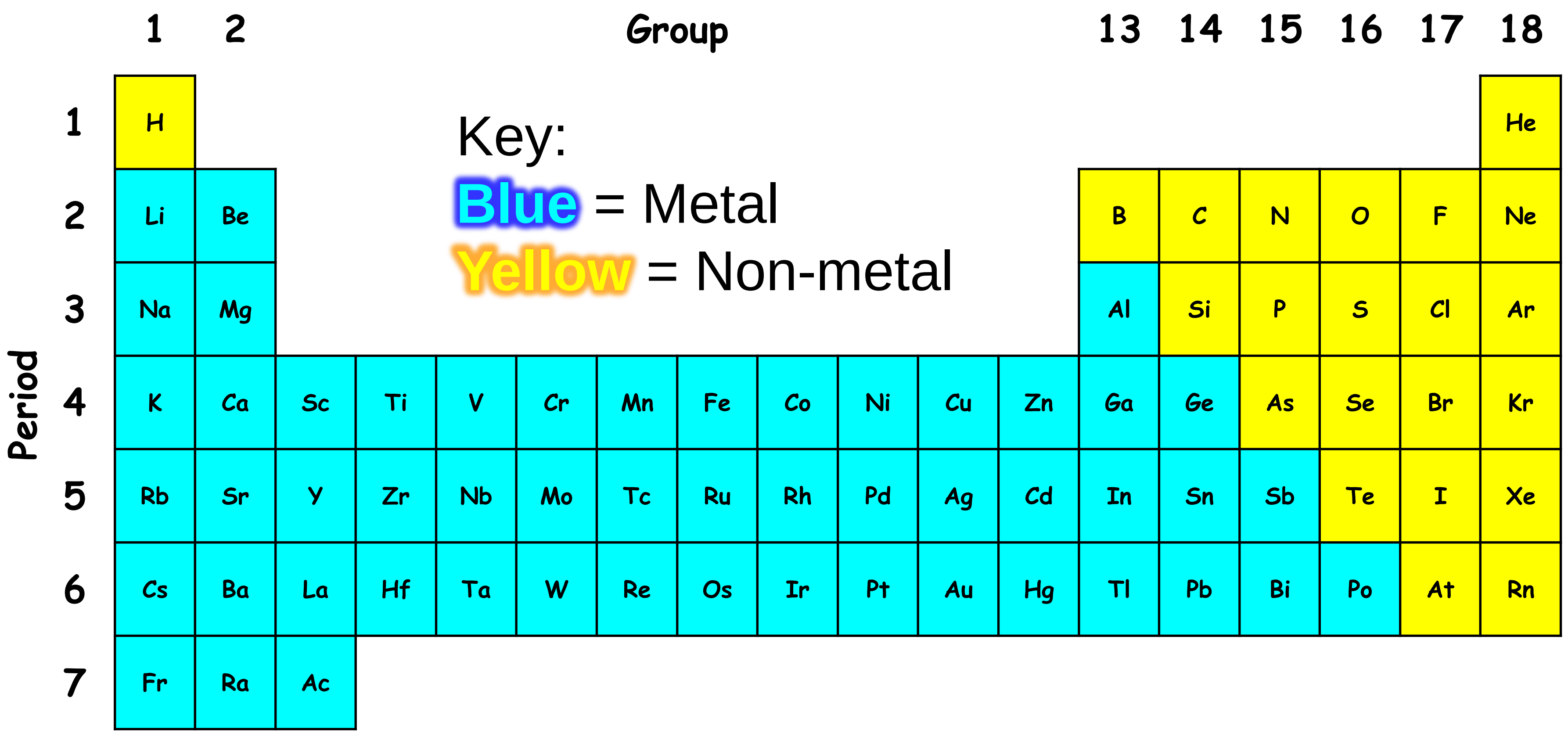
Metallic Bonding

15. What type of bonding exists within a *metallic* chemical element?

Main Menu

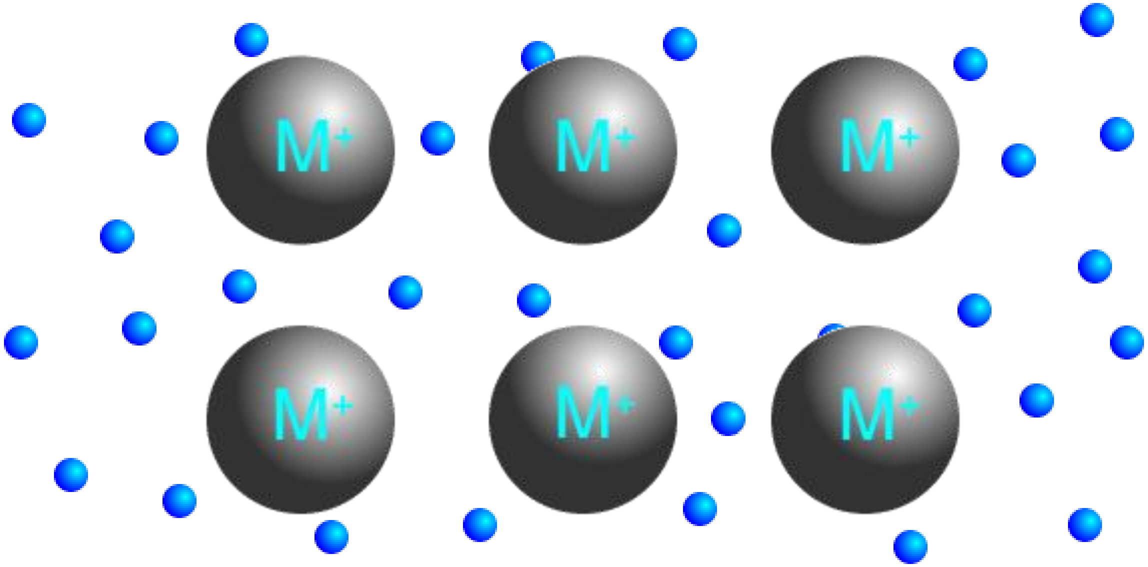
Chemical Bonding

Metallic Bonding



Chemical Bonding

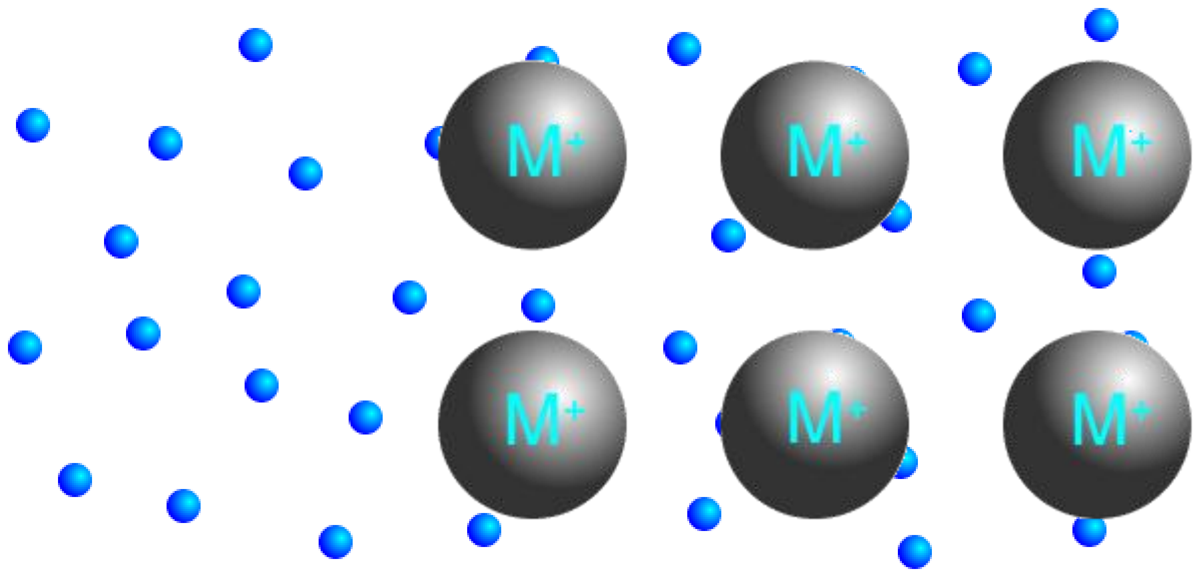
Metallic Bonding



- A metal is composed of positively charged metal ions (cations) surrounded by a “sea” of *mobile* or *delocalised* electrons. The “sea” of delocalised electrons is formed when the metal atoms lose their valence electrons in order to obtain a noble gas electronic configuration.

Chemical Bonding

Metallic Bonding



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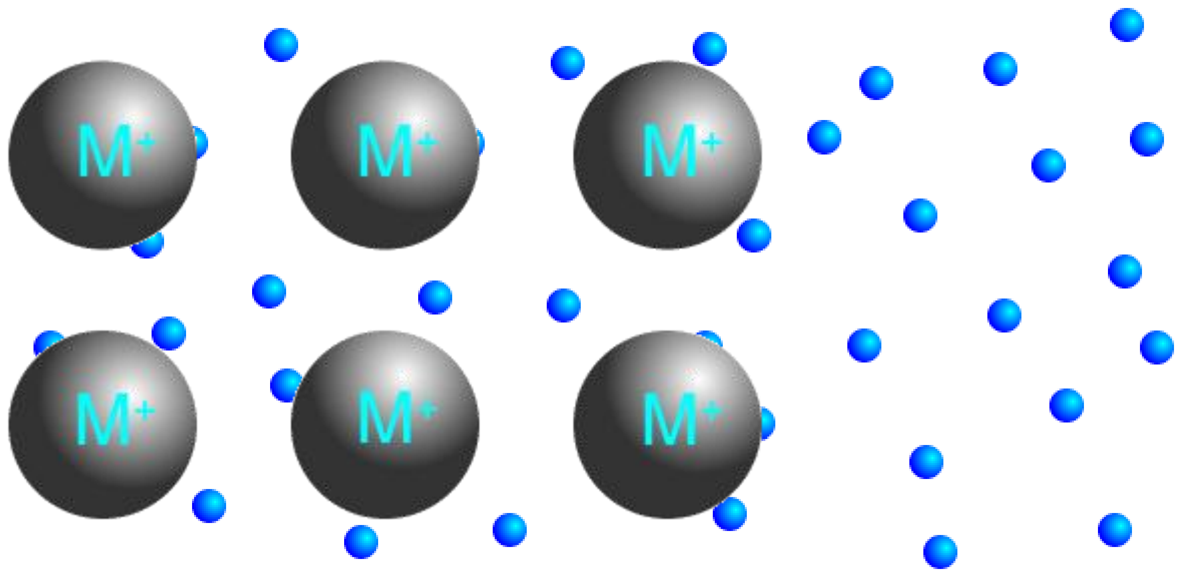
Chemical Bonding

Metallic Bonding



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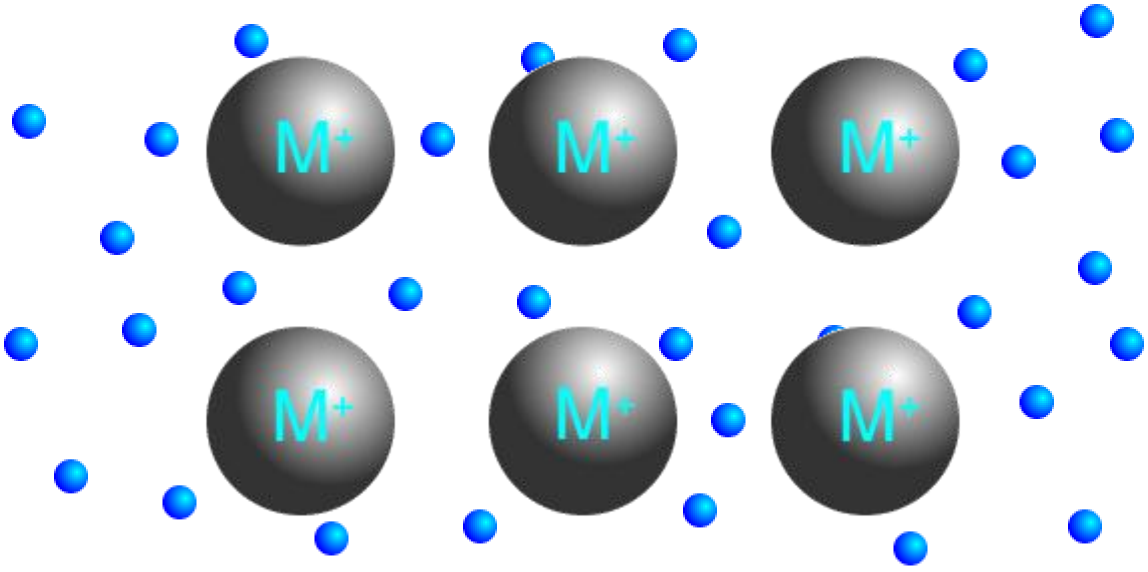
Chemical Bonding

Metallic Bonding



- The actual metallic bond is the *electrostatic force of attraction* that exists between the positively charged metal cations and the negatively charged electrons that are delocalised around them.

Chemical Bonding

Metallic Bonding

16. What are the *general properties* of metallic chemical elements?

Main Menu

Chemical Bonding

Metallic Bonding – Properties

- Metals tend to have *high melting points and boiling points* because the electrostatic force of attraction between the metal cations and delocalised electrons is strong and therefore requires a large amount of thermal energy to weaken / overcome.

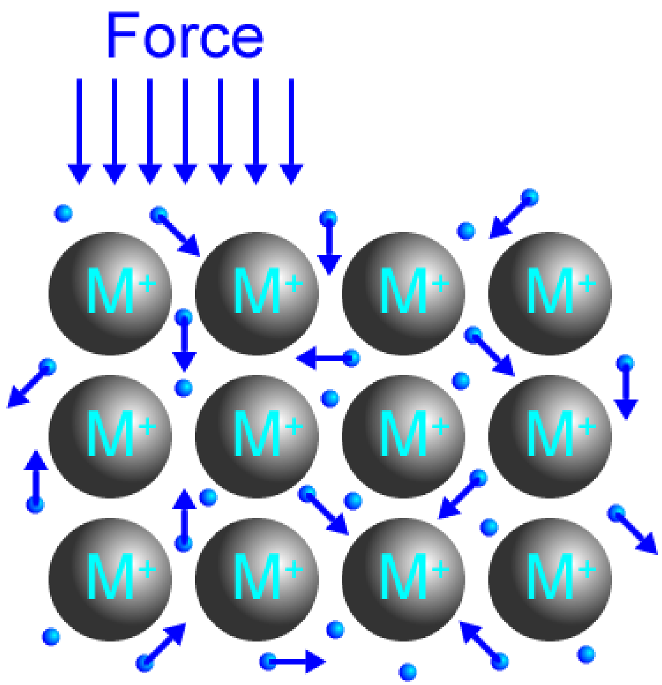
- Metals are *hard and strong*. They are also *malleable* (shape can be changed without breaking) and *ductile* (can be drawn out into a wire). Metals are also *sonorous* (ring when struck).

- Metals *do not dissolve in polar solvents* such as water. Metals are also *insoluble in non-polar solvents* such as oil and hexane. **Note:** Some metals, *e.g.* sodium, react with water.

- Metals are *very good conductors of electricity* in both the solid and molten states due to the “sea” of delocalised electrons.

Chemical Bonding

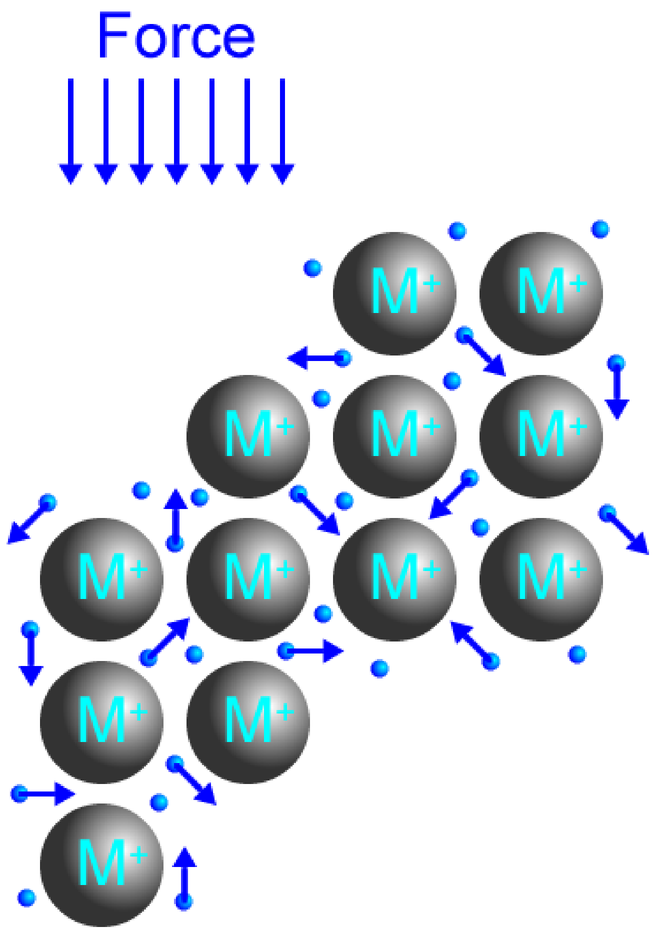
Metallic Bonding – Properties



- When a force is applied to a metal it will bend without breaking. This is because the “sea” of electrons within the metal is mobile, effectively making the metallic bonds themselves mobile and therefore allowing the layers of metal ions to slide over each other.

Chemical Bonding

Metallic Bonding – Properties

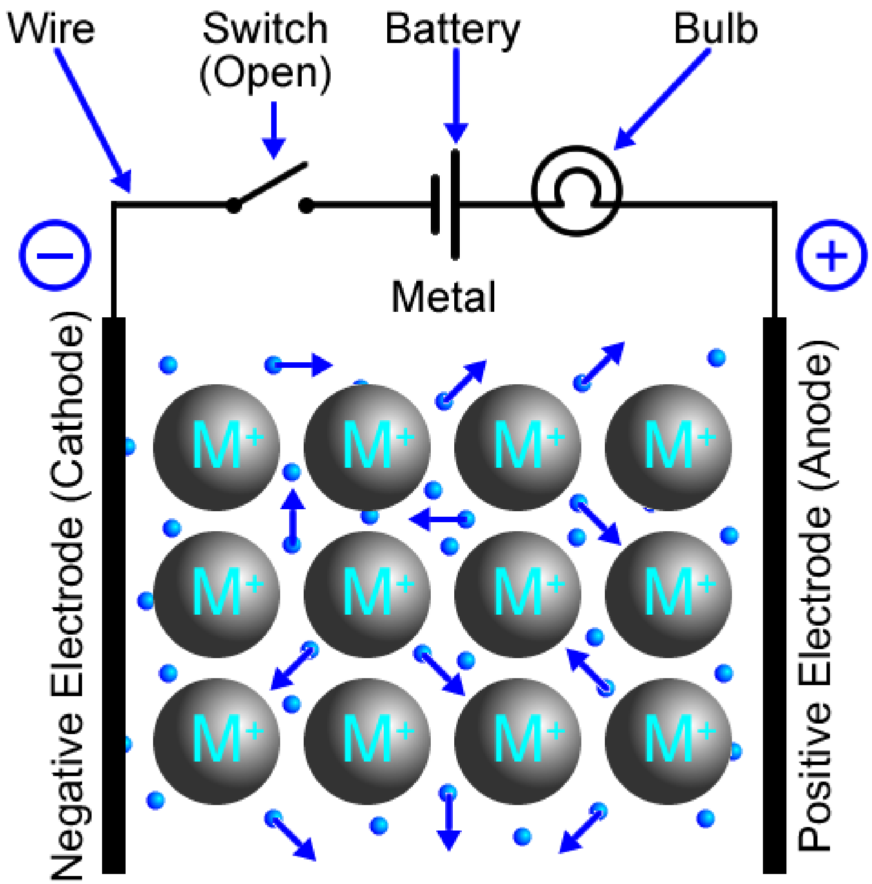


- When a force is applied to a metal it will bend without breaking. This is because the "sea" of electrons within the metal is mobile, effectively making the metallic bonds themselves mobile and therefore allowing the layers of metal ions to slide over each other.

Chemical Bonding

Metallic Bonding – Properties

- Because metals contain *mobile* or *delocalised* electrons, they are able to *conduct electricity* in both their solid and molten states.

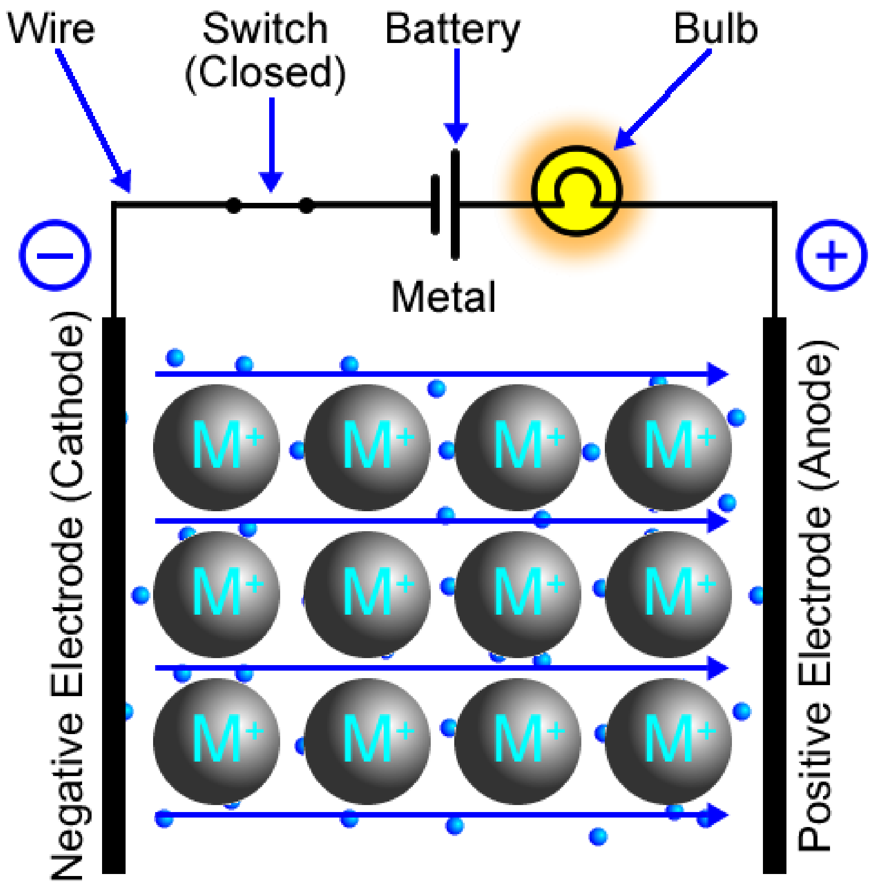


- When the circuit is closed, the *mobile* or *delocalised* electrons within the metal are all repelled away from the negative electrode (cathode) and attracted towards the positive electrode (anode).

Chemical Bonding

Metallic Bonding – Properties

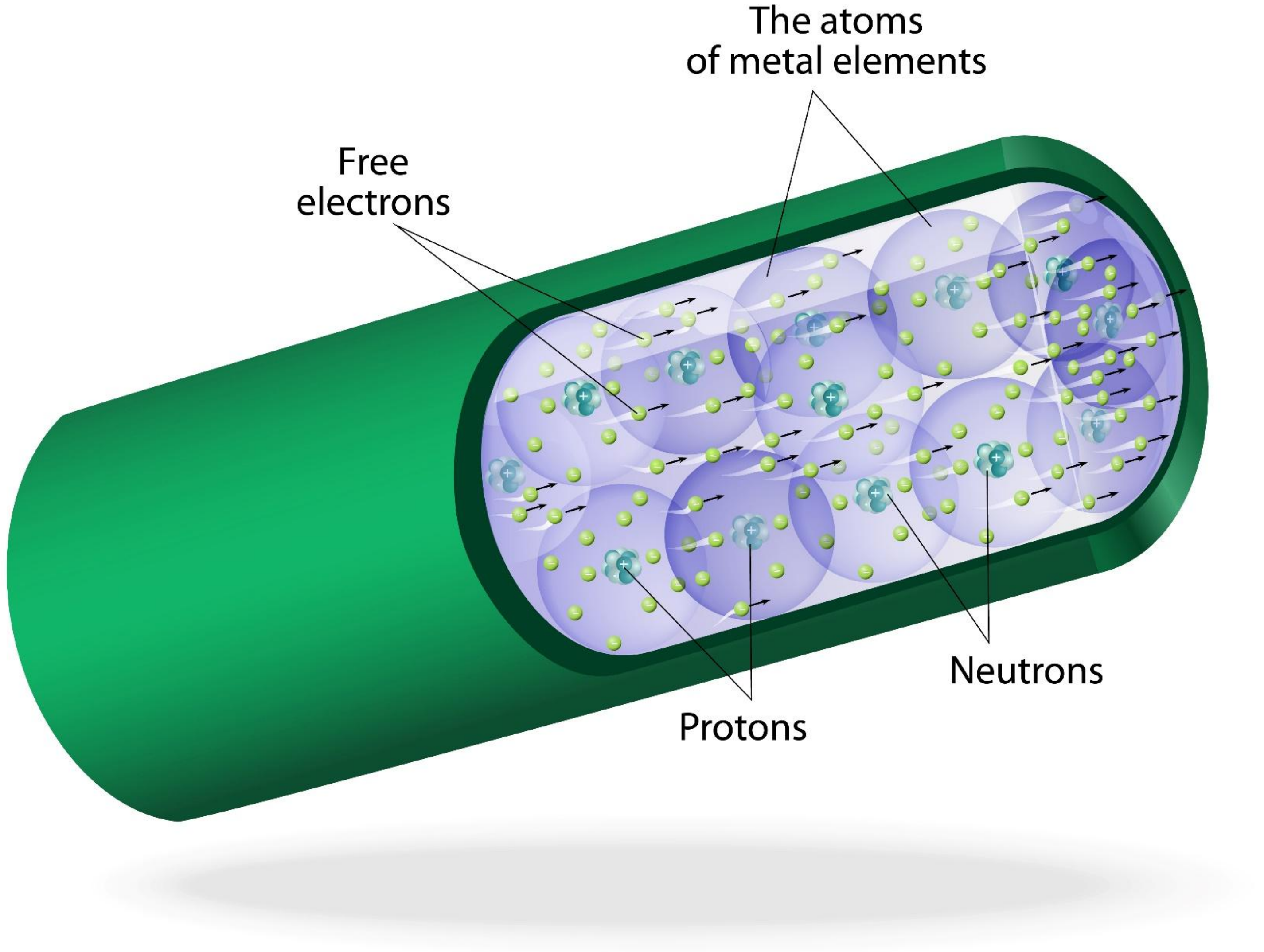
- Because metals contain *mobile* or *delocalised* electrons, they are able to *conduct electricity* in both their solid and molten states.



- When the circuit is closed, the *mobile* or *delocalised* electrons within the metal are all repelled away from the negative electrode (cathode) and attracted towards the positive electrode (anode).

Chemical Bonding

Metallic Bonding – Properties



Chemical Bonding

Metallic Bonding – Summary

- A metal is composed of positively charged metal ions (cations) surrounded by a "sea" of delocalised electrons.
- Metals have high melting points and boiling points.
- Metals are hard and strong. They are malleable and ductile.
- Metals do not dissolve in polar solvents and are also insoluble in non-polar solvents.
- Metals are very good conductors of electricity in both their solid states and molten states.

Online Search

- It what ways are the chemical and physical properties of *caesium* different from those of typical metals?

- Which materials are **a)** *cooking utensils* **b)** *electrical wires* commonly made of? Why are these materials used? What are the advantages and disadvantages of using these particular materials?

- What is an *alloy* and what particular uses do they have? Why is an alloy used for a certain application instead of a pure metallic element?

Chemical Bonding

17. Could I please have an *overview* of the different types of chemical bonding?

Chemical Bonding

Chemical Bonding $\xrightarrow{\text{Why?}}$ Atoms of the chemical elements react to obtain the stable electronic configurations of noble gases.

Chemical Bonding

Chemical Bonding

Why?

Atoms of the chemical elements react to obtain the stable electronic configurations of noble gases.

How?

Ionic Bonding

Ionic bonding takes place between a metallic element and a non-metallic element.
Atoms of the metallic element transfer electrons to atoms of the non-metallic element.
The metal forms a positively charged ion (cation) with noble gas electronic configuration.
The non-metal forms a negatively charged ion (anion) with noble gas electronic configuration.
Ions of opposite charge are held together in a highly ordered crystal lattice structure by strong electrostatic forces of attraction.

Chemical Bonding

Ionic Bonding

Ionic bonding takes place between a metallic element and a non-metallic element.
Atoms of the metallic element transfer electrons to atoms of the non-metallic element.
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Ions of opposite charge are held together in a highly ordered crystal lattice structure by strong electrostatic forces of attraction.

How?

Chemical Bonding

Why?

Atoms of the chemical elements react to obtain the stable electronic configurations of noble gases.

How?

Covalent Bonding

Covalent bonding occurs between the atoms of two or more non-metallic elements.
The atoms may be of the same non-metallic element or different non-metallic elements.
The valence shells of the atoms overlap and they share electrons to obtain the electronic configurations of noble gases. A shared pair of electrons is referred to as a covalent bond.
A mutual electrostatic force of attraction between the positively charged nuclei and negatively charged shared pair of electrons bonds the atoms together.

Chemical Bonding

Ionic Bonding

Ionic bonding takes place between a metallic element and a non-metallic element.
Atoms of the metallic element transfer electrons to atoms of the non-metallic element.
The metal forms a positively charged ion (cation) with noble gas electronic configuration.
The non-metal forms a negatively charged ion (anion) with noble gas electronic configuration.
Ions of opposite charge are held together in a highly ordered crystal lattice structure by strong electrostatic forces of attraction.

How?

Chemical Bonding

Why?

Atoms of the chemical elements react to obtain the stable electronic configurations of noble gases.

How?

Metallic Bonding

Positively charged metal cations are arranged in an ordered crystal lattice surrounded by a sea of delocalised electrons.
A mutual electrostatic force of attraction between the positively charged cations and negatively charged delocalised electrons bonds the cations together.

How?

Covalent Bonding

Covalent bonding occurs between the atoms of two or more non-metallic elements.
The atoms may be of the same non-metallic element or different non-metallic elements.
The valence shells of the atoms overlap and they share electrons to obtain the electronic configurations of noble gases. A shared pair of electrons is referred to as a covalent bond.
A mutual electrostatic force of attraction between the positively charged nuclei and negatively charged shared pair of electrons bonds the atoms together.

Chemical Bonding

Ionic Compounds:

- Ionic compounds are formed when a metallic element reacts with a non-metallic element. Atoms of the metallic element transfer their valence electron(s) into the valence shell of the non-metallic element forming positively charged metallic ions (cations) and negatively charged non-metallic ions (anions) with the electronic configurations of noble gases. A strong electrostatic force of attraction between the oppositely charged ions holds them together in a crystal lattice arrangement. Because the electrostatic force of attraction between the oppositely charged ions is strong, a large amount of thermal energy is required to weaken / break it, hence ionic compounds have relatively high melting and boiling points. Ionic compounds are electrolytes – they are electrical insulators in the solid state (because strong electrostatic forces of attraction hold the ions in fixed positions) but can conduct electricity when molten or in aqueous solution (because mobile ions are free to move to the electrode of opposite charge).

Chemical Bonding

Simple Covalent Molecules:

- A small number of atoms of non-metallic elements held together by shared pairs of electrons (covalent bonds). Atoms of non-metallic elements share pairs of electrons in order to obtain the electronic configuration of a noble gas. Covalent bonds within the molecules are strong, but intermolecular forces of attraction (*e.g.* London dispersion forces, permanent dipole-permanent-dipole interactions and hydrogen bonds) between the molecules are weak. Consequently, they have relatively low melting and boiling points because only a small amount of thermal energy is required to weaken / break the weak intermolecular force of attraction. They are electrical insulators in all states because they do not contain any mobile charge carrying particles, *i.e.* they do not contain any delocalised electrons or mobile ions. The exceptions are acids, which are simple covalent molecules that ionise when dissolved in water to produce a hydrogen ion and an anion.

Chemical Bonding

Giant Covalent Structures:

- A very large number of atoms of non-metallic elements held together by shared pairs of electrons (covalent bonds) throughout the entire structure. Atoms of non-metallic elements share pairs of electrons in order to obtain the electronic configuration of a noble gas. Covalent bonds throughout the entire structure are strong, and require a large amount of thermal energy to weaken / break, hence giant covalent structures have very high melting and boiling points. Note: there are no weak intermolecular forces of attraction in giant covalent structures, only strong covalent bonds throughout. Giant covalent structures are electrical insulators in all states because they are not composed of ions, and there are no delocalised electrons. One exception is graphite – each carbon atom in graphite has four valence electrons, of which three are used to form covalent bonds, the fourth electron being delocalised and thus able to conduct electricity.

Chemical Bonding

Metals:

- In metals, the atoms of metallic elements are closely packed in a regular crystal lattice. When packed closely together, each metal atom loses its valence electron(s) in order to obtain the electronic configuration of a noble gas. As a consequence, metals are composed of positively charged metal ions (cations) arranged in a regular crystal lattice, surrounded by a delocalised sea of electrons. The electrostatic force of attraction between the metal cations and delocalised electrons constitutes the metallic bond. Because the electrostatic force of attraction between the metal cations and delocalised electrons is strong, a large amount of thermal energy is required to weaken / break the metallic bond, and metals have high melting and boiling points. Metals are good conductors of electricity in both the solid and liquid (molten) states because the delocalised electrons are free to move and carry charge throughout the metallic structure.

Chemical Bonding

	Chemical **A**	Chemical **B**	Chemical **C**	Chemical **D**	Chemical **E**
Melting and Boiling Points	Low or Very Low	Low or Very Low	High or Very High	High or Very High	High or Very High
Solubility in Water	Insoluble	Soluble	Soluble	Insoluble	Insoluble
Electrical Conductivity of **Solid**	Electrical Insulator	Electrical Insulator	Electrical Insulator	Conducts Electricity	Electrical Insulator
Electrical Conductivity of **Aqueous Solution**	*Not Applicable*	Conducts Electricity	Conducts Electricity	*Not Applicable*	*Not Applicable*

- Based upon the properties described in the table above, predict the type of structure and bonding present in chemicals **A**, **B**, **C**, **D** and **E**.

Chemical Bonding

	Chemical **A**	Chemical **B**	Chemical **C**	Chemical **D**	Chemical **E**
Melting and Boiling Points	Low or Very Low	Low or Very Low	High or Very High	High or Very High	High or Very High
Solubility in Water	Insoluble	Soluble	Soluble	Insoluble	Insoluble
Electrical Conductivity of **Solid**	Electrical Insulator	Electrical Insulator	Electrical Insulator	Conducts Electricity	Electrical Insulator
Electrical Conductivity of **Aqueous Solution**	*Not Applicable*	Conducts Electricity	Conducts Electricity	*Not Applicable*	*Not Applicable*

A = Simple covalent molecule.

B = Simple covalent molecule which is probably an *acid*.

C = Ionic compound. **D** = Metal (could also be *graphite*).

E = Giant covalent structure / macromolecule (but *not graphite*).

Chemical Bonding

18. Could I please have a review of how ionic bonding and covalent bonding are *similar* and how they are *different*?

Chemical Bonding

How are Ionic Bonding and Covalent Bonding Similar to Each Other?

- Both ionic and covalent bonding result in the formation of particles that have the *electronic configurations of noble gases*, *i.e.* complete valence electron shells.

- Both ionic and covalent bonding result in the formation of *forces of attraction* that bind the resulting particles together.
  → In the case of ionic bonding, positively charged metal ions (cations) are attracted towards negatively charged non-metal ions (anions) by a *strong electrostatic force of attraction* – opposite charges attract.
  → In the case of covalent bonding, the nuclei of the atoms are *attracted towards the pair(s) of electrons* that are shared between them.

Chemical Bonding

How are Ionic Bonding and Covalent Bonding Different from Each Other?

- Ionic bonding occurs between the atom(s) of a *metallic element* and the atom(s) of a *non-metallic element*.

…but…

- Covalent bonding occurs between the atoms of *two or more non-metallic elements*. Covalent bonding may occur between atoms of the same non-metallic element, or between the atoms of different non-metallic elements.

Chemical Bonding

How are Ionic Bonding and Covalent Bonding Different from Each Other?

- Ionic bonding requires electrons to be *transferred* from atoms of the metallic element to atoms of the non-metallic element.

…but…

- Covalent bonding does not involve the transfer of electrons between atoms. Instead, a pair (or pairs) of electrons are *shared* between atoms. A shared pair of electrons is a covalent bond.

Chemical Bonding

How are Ionic Bonding and Covalent Bonding Different from Each Other?

- Ionic bonding results in the formation of *positively charged metal ions* (cations) and *negatively charged non-metal ions* (anions).

…but…

- Covalent bonding results in the formation *molecules* that are *neutral*, *i.e.* do not carry any overall charge.

Chemical Bonding

19. Could I please have some *questions* to practice my understanding?

Main Menu

Revision Questions

Question 1:

- Study the information given about the three different elements given below:

Particle	**Number of Protons**	**Number of Neutrons**	**Number of Electrons**
Q	9	11	9
R	12	14	12
S	8	8	10

a) Identify element **Q** and write its nuclide notation.

Revision Questions

Question 1:

- Study the information given about the three different elements given below:

Particle	Number of Protons	Number of Neutrons	Number of Electrons
Q	9	11	9
R	12	14	12
S	8	8	10

a) Identify element **Q** and write its nuclide notation.

Answer: Number of protons = 9 $\therefore$ atomic number = 9 $\therefore$ the element is fluorine. Nuclide notation = $^{20}_{9}F$.

Revision Questions

Question 1:

- Study the information given about the three different elements given below:

Particle	Number of Protons	Number of Neutrons	Number of Electrons
Q	9	11	9
R	12	14	12
S	8	8	10

b) Identify element **R** and write its nuclide notation.

Revision Questions

Question 1:

- Study the information given about the three different elements given below:

Particle	Number of Protons	Number of Neutrons	Number of Electrons
Q	9	11	9
R	12	14	12
S	8	8	10

b) Identify element **R** and write its nuclide notation.

Answer: Number of protons = 12 $\therefore$ atomic number = 12 $\therefore$ the element is magnesium. Nuclide notation = ${}^{26}_{12}Mg$.

Revision Questions

Question 1:

- Study the information given about the three different elements given below:

Particle	**Number of Protons**	**Number of Neutrons**	**Number of Electrons**
Q	9	11	9
R	12	14	12
S	8	8	10

c) Identify element **S** and write its nuclide notation.

Revision Questions

Question 1:

- Study the information given about the three different elements given below:

Particle	Number of Protons	Number of Neutrons	Number of Electrons
Q	9	11	9
R	12	14	12
S	8	8	10

c) Identify element **S** and write its nuclide notation.

Answer: Number of protons = 8 ∴ atomic number = 8 ∴ the element is oxygen. Nuclide notation = $^{16}_{8}O^{2-}$.

Revision Questions

Question 2:

- Study the information given about the four different elements given below:

Particle	**Electronic Configuration**
U	2, 8, 1
W	2, 8, 2
Y	2, 5
Z	2, 8, 7

a) Give the formula of the compound that is formed between element **U** and element **Y**. Is the bonding in this compound ionic or covalent?

Revision Questions

Question 2:

- Study the information given about the four different elements given below:

Particle	Electronic Configuration
U	2, 8, 1
W	2, 8, 2
Y	2, 5
Z	2, 8, 7

a) Give the formula of the compound that is formed between element **U** and element **Y**. Is the bonding in this compound ionic or covalent?

Answer: $\mathbf{U_3Y}$ – the bonding is ionic.

Revision Questions

Question 2:

- Study the information given about the four different elements given below:

Particle	Electronic Configuration
U	2, 8, 1
W	2, 8, 2
Y	2, 5
Z	2, 8, 7

b) Give the formula of the compound that is formed between element **Y** and element **Z**. Is the bonding in this compound ionic or covalent?

Revision Questions

Question 2:

- Study the information given about the four different elements given below:

Particle	Electronic Configuration
U	2, 8, 1
W	2, 8, 2
Y	2, 5
Z	2, 8, 7

b) Give the formula of the compound that is formed between element **Y** and element **Z**. Is the bonding in this compound ionic or covalent?

Answer: **YZ_3** – the bonding is covalent.

Revision Questions

Question 2:

- Study the information given about the four different elements given below:

Particle	Electronic Configuration
U	2, 8, 1
W	2, 8, 2
Y	2, 5
Z	2, 8, 7

c) Give the formula of the compound that is formed between element **W** and element **Y**. Is the bonding in this compound ionic or covalent?

Revision Questions

Question 2:

- Study the information given about the four different elements given below:

Particle	**Electronic Configuration**
U	2, 8, 1
W	2, 8, 2
Y	2, 5
Z	2, 8, 7

c) Give the formula of the compound that is formed between element **W** and element **Y**. Is the bonding in this compound ionic or covalent?

Answer: $\mathbf{W_3Y_2}$ – the bonding is ionic.

Revision Questions

Question 3:

a) Draw the dot-and-cross diagram to show the arrangement of the valence electrons, and hence the bonding, between the elements shown below:

P: electronic configuration = 2, 8, 1

Q: electronic configuration = 2, 8, 6

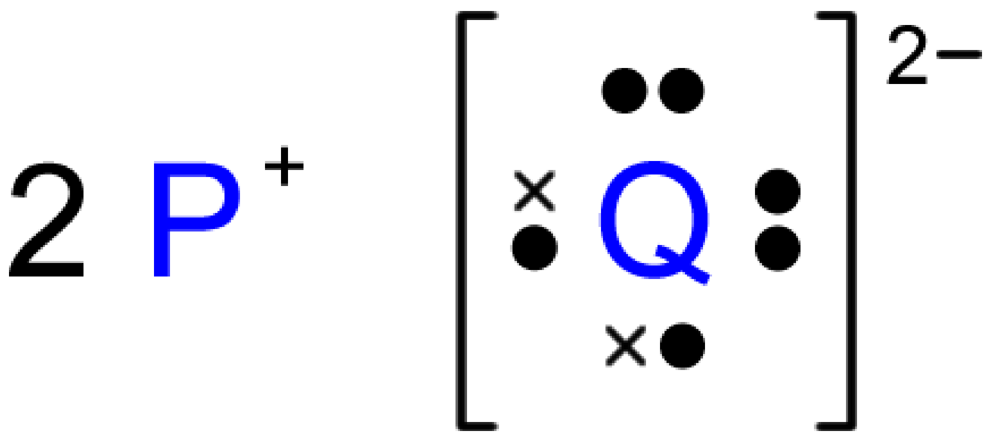
Revision Questions

Question 3:

a) Draw the dot-and-cross diagram to show the arrangement of the valence electrons, and hence the bonding, between the elements shown below:

P: electronic configuration = 2, 8, 1

Q: electronic configuration = 2, 8, 6



• = Electron of **Q**.   × = Electron of **P**.

Revision Questions

Question 3:

b) Draw the dot-and-cross diagram to show the arrangement of the valence electrons, and hence the bonding, between the elements shown below:

X: electronic configuration = 2, 4

Y: electronic configuration = 2, 8, 7

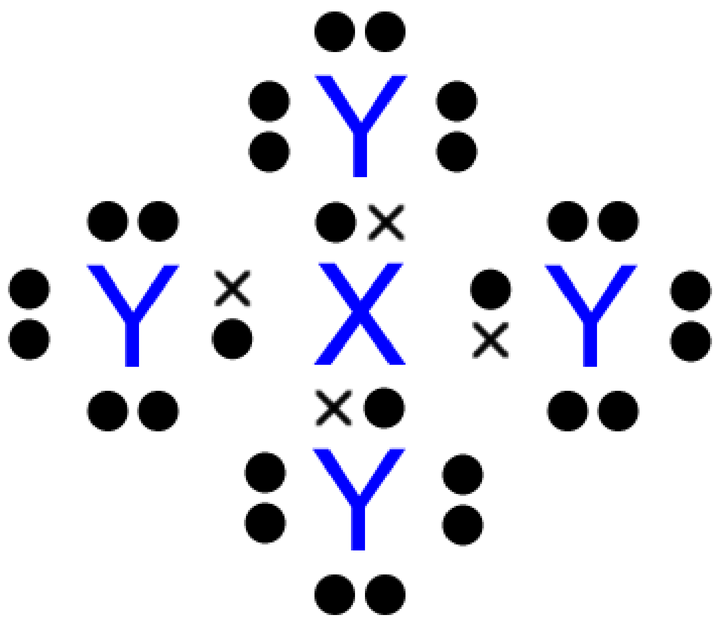
Revision Questions

Question 3:

b) Draw the dot-and-cross diagram to show the arrangement of the valence electrons, and hence the bonding, between the elements shown below:

X: electronic configuration = 2, 4

Y: electronic configuration = 2, 8, 7



• = Electron of **Y**.   × = Electron of **X**.

Chemical Bonding

20. Could I please have a *summary of the different terms* that have been used?

Glossary of Terms

Atom:

- An atom is the smallest part of a chemical element that demonstrates all of the typical chemical properties of that element.

Glossary of Terms

Anion:

- An anion is a negatively charged ion, *e.g.* chloride ions (Cl^-) and oxide ions (O^{2-}) are both anions.

Glossary of Terms

Atomic Number:

- The atomic number of a chemical element equals the number of protons present in the nucleus of a single atom of the element.

Glossary of Terms

Cation:

- A cation is a positively charged ion *e.g.* sodium ions (Na^+) and magnesium ions (Mg^{2+}) are both cations.

Glossary of Terms

Compound:

- A compound is a pure substance that is formed when two or more different chemical elements react and chemically bond together. The ratio of elements in a compound is fixed, and is given by the compound's formula, *e.g.* CO_2. A compound has different properties compared to the elements from which it was formed. The elements in a compound cannot be separated by a physical process, *e.g.* chromatography or distillation.

Glossary of Terms

Covalent Bond:

- A covalent bond is the type of chemical bond that is formed when a pair of electrons is shared between the atoms of two non-metallic elements.

Glossary of Terms

Diatomic Molecule:

- A diatomic molecule is a group of two non-metallic atoms that are joined together by a covalent bond(s). Diatomic molecules can be either elements, *e.g.* chlorine (Cl_2) or compounds, *e.g.* carbon monoxide (CO).

Glossary of Terms

Electron:

- An electron is a subatomic particle with a charge of –1 and a mass of $^{1}/_{1840}$ a.m.u. (atomic mass unit) that orbits the nucleus of an atom.

Glossary of Terms

Element:

- An element is a pure substance that cannot be converted into anything more simple by a chemical process / reaction.

Glossary of Terms

Ion:

- An ion is a positively or negatively charged particle. It is formed when an atom or group of atoms loses or gains electrons.

Glossary of Terms

Ionic Compound:

- An ionic compound is formed when a metal reacts with a non-metal. The metal transfers electrons to the non-metal so that both obtain a noble gas electronic configuration. The metal forms a positively charged ion (cation) and the non-metal forms a negatively charged ion (anion). The oppositely charged ions are held together in an ordered lattice / crystal structure by strong electrostatic forces of attraction.

Glossary of Terms

Ionic Bond:

- An ionic bond is the electrostatic force of attraction between positively charged ions (cations) and negatively charged ions (anions).

Glossary of Terms

Isotope:

- Isotopes are atoms of the same chemical element that contain the same number of protons (*i.e.* the same atomic number) but a different number of neutrons (*i.e.* a different mass number). Isotopes have the same chemical properties, but slightly different physical properties.

Glossary of Terms

Mass Number:

- The mass number of a chemical element equals the total number of protons and neutrons present in the nucleus of a single atom of the element.

Glossary of Terms

Metallic Bond:

- The metallic bond is the electrostatic force of attraction between positively charged metal ions (cations) and the negatively charged delocalised electrons that surround them.

Glossary of Terms

Mixture:

- A mixture is formed when two or more chemical elements / compounds are added together but do not react and do not chemically combine together. The components of a mixture can be easily separated by a physical process such as distillation, filtration or chromatography.

Glossary of Terms

Molecule:

- A group of two or more non-metallic atoms that are joined together by a covalent bond(s). Molecules can be elements, *e.g.* chlorine (Cl_2) or compounds, *e.g.* water (H_2O).

Glossary of Terms

Neutron:

- A neutron is a subatomic particle with a charge of 0 and a relative mass of 1 a.m.u. (atomic mass unit) that is located in the nucleus of an atom.

Glossary of Terms

Polyatomic Molecule:

- A polyatomic molecule is a group of three or more non-metallic atoms that are joined together by covalent bonds. Polyatomic molecules can be either elements, *e.g.* phosphorus (P_4) or compounds, *e.g.* methane (CH_4).

Glossary of Terms

Proton:

- A proton is a subatomic particle with a charge of +1 and a relative mass of 1 a.m.u. (atomic mass unit) that is located in the nucleus of an atom.

Glossary of Terms

Valence Electrons:

- Valence electrons are electrons in the outer electron shell of an atom that used by the atom for forming chemical bonds.

Chemical Bonding

Presentation on
Chemical Bonding
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23rd June 2015

Updated 24th May 2018