**Unit 1.5/1.6 Chemical bonding**

Different substances form when atoms join to each other. These new substances have new physical and chemical properties. The atoms need to hold on to one another for these new substances to form. For example, for a sodium atom (Na) and a chlorine atom (Cl) to form sodium chloride (NaCl) the two atoms have to stick together in some way. This process is called **chemical bonding.**

In order to understand the nature of chemical bonds you need to understand electron configuration, so now you are ready to look at chemical bonding.

**Electrons and chemical bonds**

Remember the following points:

* Core electrons occupy the filled inner principal energy levels.
* Valence electrons occupy the outermost principal energy level.
* The valence electrons are responsible for forming the chemical bonds that holds two or more atoms together.
* The inert Noble Gases have filled outermost principal energy levels.
* A filled outermost principal energy level can carry a maximum of 8 electrons (in the *s* and *p* sublevels), except in Period 1 where only 2 electrons can be accommodated (in the *s* sublevel).



Elements which do not have filled outermost principal energy levels are *unstable* and *reactive*. They therefore need to react with other atoms and form chemical bonds in such a way that they become more stable. They do this by obeying the octet rule.

**The octet rule**

Atoms bond together in such a way that each of the atoms involved in the bond ends up with 8 electrons in itsoutermost principal energy level.

Atoms achieve this in several ways:

* Ionic bonding
* Covalent bonding
* Metallic bonding

**Lewis dot diagrams**

Lewis dot diagrams show the valence electrons of an element. This is very useful for understanding chemical bonding since it is only the valence electrons, and not the core electrons that are involved in bond formation.

Lewis dot diagrams are drawn using the following rules:



* The symbol of the element represents the nucleus and the core electrons of the atom.
* A valence electron is represented by a dot ●.
* Use the Periodic Table to determine the number of valence electrons in any particular element.
* The maximum number of valence electrons is 8, so a maximum of 4 pairs of dots can surround the central symbol, one pair on each side and top and bottom of the symbol.
* Place the first 4 valence electrons around the four sides of the symbol before starting to double up the dots.

**Activity 1: Practising Lewis dot diagrams**

Write Lewis dot diagrams for the following elements:

1. Boron
2. Carbon
3. Nitrogen
4. Oxygen
5. Flourine
6. Neon
7. Sulfur
8. Argon
9. Sodium
10. Calcium

**Ionic bonds**

Ionic bonds form between a **metal atom** and a **non-metal atom** (or a metal atom and a group of non-metal atoms called polyatomic anions). Particles held together by ionic bonds are called **formula units**.

Ionic bonds involve the formation of positively charged particles (called **cations**) and negatively charged particles (called **anions**). The electron configuration of cations and anions look the same as the electron configuration of a noble gas – they are obeying the octet rule.

The particles of opposite charge are then attracted to each other in an ionic bond. This is called **electrostatic attraction**.

If you look at the valence electrons of the metals you will notice that there are relatively few. For example, the alkali metals all have only one valence electron and the alkaline earth metals only have 2 valence electrons.

On the other hand, the non-metals have a larger number of valence electrons. For example, all the halogens have 7 valence electrons.

In order to end up with filled outermost principal energy levels the metal atoms donate the valence electrons to the non-metals which are looking for ways to fill the gaps. This movement of electrons, which are negatively charged particles, leads to the formation of ions carrying charges.

Each ion has the same number of electrons as a particular noble gas, so it has the stability of a noble gas. It is said to be **isoelectronic** with the noble gas.

Ionic bonds form in the following way.





The alkaline earth metals have 2 valence electrons. They would donate 2 electrons and therefore get a 2+ charge. The size of the charge is determined by the number of electrons donated.

 Mg (insert dots) → Mg2+ + 2e-

If Mg2+ combines with chlorine it would need to combine with 2 Cl- ions to balance the 2+ charge, giving MgCl2 for magnesium chloride.

The subscript after the element indicates the number of ions present for that element.

You now know how to write balanced chemical formulae!

Example

Name the noble gas with an electron configuration the same as the following ions:

1. Li+
2. F-

Solution

1. Li has 3 electrons. It loses 1 to for the ion Li+, leaving 2 electrons. He is the noble gas with 2 electrons.
2. F has 9 electrons. It gains 1 to for the ion F- giving a total of 10 electrons. Ne is the noble gas with 10 electrons.

Key concepts

* Metal elements lose electrons and become positively charged cations. The number of electrons lost determines the size of the positive charge.
* Non-metal elements gain electrons and become negatively charged anions. The number of electrons gained determines the size of the negative charge.
* Ionic bonds form when ions of opposite charge are attracted to one another to give a compound with zero net charge

**Activity 2: Electron configuration and ions**

1. Write out an equation for the formation of the ions of the following elements:
	1. Potassium
	2. Calcium
	3. Bromine
	4. Oxygen
	5. Nitrogen (5)
2. Name the noble gas with an electron configuration the same as the following ions:
	1. S2-
	2. Na+
	3. Al3+
	4. Br-
	5. P3- (5)

**Example**

Write the formula for the compound formed between sodium and oxygen.

**Solution**

Sodium forms the Na+ cation

Oxygen forms the O2- anion

To balance the charges 2 sodium ions are required for each oxygen ion.

The compound is Na2O

**Activity 3: Writing ionic chemical formulae**

Write the formula for the compound formed between the following pairs of elements:

1. Potassium and iodine
2. Aluminium and chlorine
3. Lithium and sulfur
4. Magnesium and oxygen
5. Sodium and nitrogen (10)

**Naming binary ionic compounds**

Binary ionic compounds are made up of only two different kinds of elements.

There are simple rules for naming binary ionic compounds.

Using NaCl as an example:

* Identify the metal and write down its name: sodium
* Identify the non-metal and write down the first part of its name (called the stem): chlorine → chlor
* Replace the second part of the name (ine in this case) with the suffix ide: chloride
* Put the two parts together: sodium chloride

**Activity 4: Naming ionic compounds**

Name these binary compounds and list each element and how many are involved in each formula.

1. KBr
2. CaCl2
3. LiF
4. Al2O3
5. Na2S (10)

**Ions and the Periodic Table**

The Periodic Table also tells us a lot about the formation of ions. You already know that all the alkali metals are in Group 1 and have the same general electron configuration with 1 electron in the outermost principal energy level. This means that all Group 1 elements would lose 1 electron to become ions with a -1 charge. So it would not be necessary to draw out Bohr structures in order to know the charge of the ion.

**Activity 5: Using the Periodic Table to determine ionic charges**

Using what you already know from your work in Activity 1, state the ionic charge for ions of elements in the following Groups on the Periodic Table:

1. Group 18
2. Group 2
3. Group 15
4. Group 17
5. Group 16 (5)

**Polyatomic ions**

The ions you have looked at so far have all been made of one atom only, for example the chloride ion, Cl-. They are therefore monatomic ions, although we just call them ions.

A polyatomic ion, also known as a molecular ion, is an ion composed of **two** **or more** atoms covalently bonded together, and together carrying a charge. An example would be the hydroxide ion, -OH- which is made up of an O atom and an H atom covalently bonded together. Together they have a charge of -1.

Here is a table of some of the more common polyatomic ions:

|  |  |
| --- | --- |
| **Name of ion** | **Formula of ion** |
| ammonium | NH4+ |
| hydroxide | OH- |
| nitrate | NO3- |
| nitrite | NO2- |
| sulfite | SO32- |
| sulfate | SO42- |
| hydrogen sulfate | HSO4- |
| carbonate | CO32- |
| hydrogen carbonate | HCO3- |
| phosphate | PO43- |
| hydrogen phosphate | HPO42- |
| permanganate | MnO4- |
| acetate | C2H3O2- |

Note: There is only one **cation** (positively charged polyatomic ion) on the list!

The best thing to do now is to learn them off by heart! All chemists have had to do this.

Note: Remember to treat the polyatomic ion as a whole unit. So if you need to double the number of negative charges to balance the positive charges when writing a formula put a bracket around the polyatomic ion. Look at these two examples:

* For Calcium chloride the calcium ion has a +2 charge and the chloride ion has a -1 charge. Therefore two chloride ions are needed to combine with one calcium ion to form the compound. The formula is therefore CaCl2
* For calcium calcium hydroxide the calcium ion has a +2 charge and the hydroxide ion has a -1 charge. Therefore two hydroxide ions are needed to combine with one calcium ion to form the compound. The formula is therefore Ca(OH)2

Activity : Complete the following table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Compound name** | **Compound formula** | **Cation** | **Anion** |
| Ammonium hydroxide |  |  |  |
|  |  | Na+ | CO32- |
| Potassium permanganate |  | K+ |  |
| Magnesium nitrate |  |  | NO3- |
|  | Ca(OH)2 |  |  |
| Sodium phosphate |  |  |  |

**Answers:**

|  |  |  |  |
| --- | --- | --- | --- |
| **Compound name** | **Compound formula** | **Cation** | **Anion** |
| Ammonium hydroxide | NH4OH | NH4+ | OH- |
| Sodium carbonate | Na2 CO3 | Na+ | CO32- |
| Potassium permanganate | K MnO4 | K+ | MnO4- |
| Magnesium nitrate | Mg(NO3-)2 | Mg2+ | NO3- |
|  | Ca(OH)2 |  |  |
| Sodium phosphate | Na3 PO4 | Na+ | PO43- |

**Ionic crystals**

Ionic compounds do not usually exist as single formula units (like one Na and one Cl only in NaCl). They rather group together in large structures called **crystal lattices** or **ionic crystals**. The salt that is big enough to see is salt in crystalline form in a crystal lattice.

Insert pic

P180 Fig 8

Many millions of formula units pack together in a three dimensional structure as shown in the diagram. This results in a very strong network of ionic bonds.

The written formula, NaCl for example, represents a **formula unit.** It gives the smallest possible combining ratio for the elements making up the compound, namely 1Na:1Cl. Another name for this is the empirical formula.

**Covalent bonds**

When two non-metals join to form a molecule they cannot do so with ionic bonds as neither is able to form a positively charged ion. Instead, the two non-metal atoms **share** valence electrons.

Consider the formation of H2 (hydrogen gas)

The Lewis dot diagram or each H atom is H●

Each H atom is short of one electron to fill the first principal energy level with 2 electrons.

If the 2 atoms come together and share the Lewis dot diagram is H●●H.

One bond is made from 1 pair of shared electrons.

Another way of writing the formula for hydrogen gas is H-H where the – represents 2 shared electrons.

**Electronegativity (EN)**

Each element has the ability to attract electrons towards itself. What varies, however, is the strength of the force of attraction. An American scientist named Linus Pauling was able to measure these forces (compared with a force of 2.5bfor the C atom). These values are called electronegativity values. The higher the number, the greater the force of attraction.

Insert Periodic Table of Electronegativities.

Activity : Electron tug-of-war

1. Which element has the highest EN value?
2. How does this affect a bond involving this element?
3. Which other elements have high EN values?
4. Are these elements metals or non-metals?
5. Which element has the lowest EN value?
6. How does this affect a bond involving this element?
7. Which other elements have low EN values?
8. Are these elements metals or non-metals?
9. What do you notice about the EN values as you move down a Group?
10. What do you notice about the EN values as you move along a Period?
11. Study the following diagram which depicts a covalent bond between two hydrogen atoms:

Insert diagram

M&M U2.8

* 1. What is the electro-negativity of each of the atoms?
	2. Two electrons are involved in the covalent bond between the two hydrogen atoms, one from each atom. Would you expect the resulting electron pair to be closer to one atom than the other? Explain you answer.
	3. Draw a similar diagram depicting a hydrogen atom and a fluorine atom. Based on your understanding of electro-negativity, would you expect the electron pair to be closer to the hydrogen or the fluorine, or exactly halfway between the two as found for hydrogen?
	4. If the electron pair is closer to one of the nuclei than the other, would you expect the bond to be completely neutral i.e. showing no sign of charge at all? Explain your answer.

A covalent bond is something like a tug-of-war between the two atoms involved in the bond. If both nuclei are identical, or if their electro-negativities are very similar, the electron pair will be located halfway between the two and there will be no net charge associated with the bond. If, on the other hand, one nucleus is significantly more electron-negative than the other it will exert a stronger pull on the electrons. Both electrons will thus be situated closer to that nucleus than the other and the nucleus will become slightly negative in comparison to the other nucleus.

Insert tug-of-war picture

M&M U2.9

Cartoon or photo

Got it: A covalent chemical bond can be described as the net electrostatic force between two atoms which are sharing electrons.

# Non-polar and polar covalent bonds

You have now understood that covalent bonds form when two non-metal elements come together and share pairs of electrons. You have also met the concept of electro-negativity which is a measure of the force of attraction that a particular nucleus has for electrons. When the electro-negativities of two bonding atoms are equal (or close to equal) the pull of the electrons must be equal, and neither atom has more influence over the electrons than the other. This kind of bond is called a *non-polar bond*. Said another way, non-polar covalent bonds are bonds where the electrons are shared equally between the two atoms involved in the bond. An obvious example would be in a diatomic element where the electro-negativity of each of the two elements involved in the bond is the same.

Cartoon of 2 tug-of-war efforts, one having equally matched teams and the other mismatched teams

Can you think of other instances where the bonds might be non-polar? Can you give any examples?

On the other hand, when the electro-negativities of two bonding atoms are quite different from one another, the atom with the higher value will pull the electrons more strongly and an unequal pull will result. Such a bond would still be a covalent bond (sharing of electrons) but it would no longer be non-polar. It would now be called a *polar covalent bond*.

Ask yourself.

1. Can you name the elements which occur as diatomic molecules and are therefore non-polar?
2. Can you think of other instances where the bonds might be non-polar? Can you give any examples?
3. Can you find some examples of polar covalent bonds?

Answer:

1. O2, N2, H2, F2, Cl2, Br2, I2
2. When the electro-negativity value of one non-metal is the same as the electro-negativity of another different non-metal element e.g. C and S, B and As
3. When the electro-negativity value of one non-metal element is different from the electro-negativity of the other element in the bond the bond is polar, e.g. H and O in H2O and C and O in CO2

Got it: When the electro-negativity values of two non-metals involved in a covalent bond are equivalent the bond is termed **non-polar** because the electrons are equally shared between the two atoms.

Insert diagram

Cloud type pics for H2, Cl2

M&M U2.15

The difference in electro-negativity between two atoms is called ΔEN

What do you think happens in the case where the electro-negativity values of the two elements involved in the covalent bond are quite different from one another, such as in a bond between H and Cl or N and O?

What happens is that the electrons are no longer shared equally – the pair of electrons is located closer to the more electro-negative atom in the bond, as shown in the following diagram:

Insert diagram

Cloud type pics for HCl

M&M U2.16

# Notation for polar bonds

A polar bond has a “negative” end and a “positive” end and is said to have a **dipole** (meaning two poles). The “negative” end is indicated by the symbol δ- (delta minus) and the “positive” end by the symbol δ+ (delta plus). These charges are small, much smaller than the charge found on an ion, for example. The dipole bond is represented by an arrow pointing towards the negative pole and having a cross (+ sign) attached at the positive pole, as shown in the diagram below.

Got it: Polar covalent bonds carry *partial* charges due to the unequal sharing of electrons.

δ- (delta minus) indicates a partial negative charge while δ+ (delta plus) indicates a partial positive charge.

 δ- δ+ ------------represents polarity

 X --- Y

 +----→
Note: The + sign is placed at the atom that is δ+ (delta plus) with an arrowhead directed towards the atom that is δ- (delta minus)

Activity 9: Test your understanding of ΔEN values and how they can be used.

1. Calculate the ΔEN values for the following pairs of atoms and indicate whether the elements would be involved in covalent or ionic bonds.

|  |  |
| --- | --- |
| H and F | Na and S |
| P and O | N and H |
| K and Cl | Li and F |
| Ca and O | C and H |
| H and H | C and Cl |
| N and O | N and N |

Answers: Assign one mark for each ΔEN value and one mark for each “type of bond”

|  |  |  |
| --- | --- | --- |
| **Element pairs** | **ΔEN** | **Type of bond** |
| H and F | 4.0 - 2.1 = 1.9 | Polar covalent |
| P and O | * 1. 2.1 = 1.4
 | Polar covalent |
| K and Cl | 3.0 - 0.8 = 2.2 | Ionic |
| Ca and O | 3.5 -1.0 = 2.5 | Ionic |
| H and H | 2.1 - 2.1 = 0 | Non-polar covalent |
| N and O | 3.5 - 3.0 = 0.5 | Polar covalent |
| Na and S | 2.5 - 0.9 = 1.4 | Ionic |
| N and H | 3.0 - 2.1 = 0.9 | Polar covalent |
| Li and F | 4.0 - 1.0 = 3.0 | Ionic |
| C and H | 2.5 - 2.1 = 0.4 | Polar covalent |
| C and Cl | 3.0 - 2.5 = 0.5 | Polar covalent |
| N and N | 3.0 - 3.0 = 0 | Non-polar covalent |

Ask yourselves: What is the difference between a polar covalent bond and an ionic bond? Use the information that you have just calculated to support your answer.

Answer: In an ionic bond one atom has lost an electron entirely and its bond partner has gained the electron completely whereas in a polar covalent bond one atom has greater attraction for the electrons than the other but does not have complete control over them. The ΔEN values for non-polar bonds are zero, for polar covalent bonds relatively small (between 0.4 and 1.9 in the examples) and high for ionic bonds (above 2 in the examples).

Got it: Ionic, polar covalent and non-polar covalent bonds are closely related to one another. They differ only in the extent to which the electro-negativities of the bonded atoms differ from one another. The general rule is that a ΔEN value of 0 – 0.4 usually indicates a non-polar bond, ΔEN values greater than 0.4 and smaller than usually 2 indicate polar covalent bonds and ΔEN values greater than 2 indicate ionic bonds.

Bonding can therefore be described as a **continuum** and is shown in the following diagram:

Insert diagram – continuum

M&M U2.17

# Multiple bonds

Up until now you have been considering simple, single covalent bonds, i.e. covalent bonds where a single bond (sharing of one pair of electrons) exists between two atoms. This is not always the case, however. Sometimes more than one pair of electrons is shared between two atoms. If two electron pairs are shared a double bond forms, and if three pairs of electrons are shared a triple bond forms.

Activity 10 : Covalent bonds and the Periodic Table.

1. Draw part of the Periodic Table according to the following sketch. Each block should be large enough to accommodate 3 pieces of information.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Group number** | **14** | **15** | **16** | **17** | **18** |
| Row 2 |  |  |  |  |  |
| Row 3 |  |  |  |  |
| Row 4 |  |  |  |
| Row 5 |  |  |
| Row 6 |  |
|  |

1. Insert the relevant non-metal element symbols into each block.
2. Insert the number of valence electrons available to each element
3. Insert the number of shared electrons required by each atom for stability – this is an indication of the number of bonds that need to be formed.
4. What do you notice about these numbers for elements occurring in the same Group?
5. Using the Lewis Dot Notation, take a representative element from each Group and draw the kind of compound that it might form with hydrogen, bearing in mind that the hydrogen atom has one electron to share.
Note: If an atom needs more than one electron from somewhere else it will have to bond with more than one hydrogen atom. Here is an example for you:

 e.g PH3 P + 3H → H P H insert dots
 H

Remember that H (and He) obeys the “duet” rule, so is stable when surrounded by 2 electrons while most of the other elements obey the “octet” rule

1. What do you notice about the electrons surrounding N (and other Group 15 elements)?
2. Pairs of electrons that are not involved in a covalent bond are called **lone pairs**, or *non-bonded electrons* while a shared pair of electrons is called a **bonding pair**. How many lone pairs are present in your compound containing N and H?
3. What do you notice about the electrons surrounding O (and other Group 16 elements)? How many lone pairs and bonding pairs are present?
4. What do you notice about the electrons surrounding C ? How many lone pairs and bonding pairs are present?
5. What do you notice about the electrons surrounding F (and other Group 17 elements)? How many lone pairs and bonding pairs are present?
6. What do you notice about the electrons surrounding Ne (and other Group 18 elements)? How many lone pairs and bonding pairs are present?

Formative assessment: Self/Group

|  |  |  |
| --- | --- | --- |
| At the end of this activity I was able to | Yes | No |
| Write down the number of valence electrons for the non-metals |  |  |
| Write down the number of shared electrons required for a stable octet of electrons |  |  |
| Determine the number of bonds that need to form between a particular element and H |  |  |
| Draw Lewis Dot formula for covalent compounds between representative elements from Groups 14 –17 with H atoms |  |  |
| Recognise lone pairs of electrons |  |  |
| Recognise bonding pairs of electrons |  |  |
| Understand why the noble gases are usually inert |  |  |

 Having trouble understanding? Talk to your teacher

Answers

Challenge: Can you draw Lewis dot formulae for O and N and then try to draw the Lewis dot formulae for O2 and N2?

Answers: The learners may have problems with getting this correct. If that is the case you will need to carefully explain how multiple bond form.

 N + N → N N 3 shared electron pairs = triple bond insert dots

 O + O → O O 2 shared electron pairs = double bond

Got it: For some elements in certain bonds there are not enough electrons to go around in such a way that both atoms have a full octet. In these cases additional sharing of electrons has to take place, and this is where **multiple bonds** are formed.

Ask yourselves: How do you think multiple bonds compare with single bonds in terms of bond strength and bond length?

Answer: Multiple bonds are stronger (as shown in the bond dissociation table) and therefore shorter than single bonds.

Got it: The length of a bond(s) is related to the strength of a bond(s): the stronger the bond, the shorter it is. This is summarised on the following table where bond lengths and bond strengths for four molecules are compared. In one pair a comparison can be made between an O-O single bond and an O=O double bond, and in the other between an N-N single bond and an N N triple bond.

Insert Table: insert Lewis dots and bonds

|  |  |  |  |
| --- | --- | --- | --- |
| Molecule | Name | Bond length (pm) | Bond strength (kJ/mol) |
| O=O | Oxygen | 121 | 498 |
| H-O-O-H | Hydrogen peroxide | 148 | 213 |
| N N | Nitrogen | 110 | 945 |
|  H HH- N- N- H | Hydrazine | 145 | 275 |

1. Count the number of electron charge clouds surrounding the atom of interest ie count the

**Metallic bonding**

## Metals are giant structures with free electrons

Metals form these giant structures in which electrons in the outer shells of the metal atoms are free to move. The metallic bond is the force of attraction between these free electrons and metal ions. Metallic bonds are strong, so metals can maintain a regular structure and usually have high melting and boiling points.



Atomic structure of a metal

Metals are good conductors of electricity and heat, because the free electrons carry a charge or heat energy through the metal.

In addition, the free electrons allow metal atoms to slide over each other, so metals are malleable and ductile.

Malleable means that metals can be shaped into new things by pressing or hammering e.g. the shapes from which motor cars are built.

Ductile means that metals can be pulled into stands e.g. wire.

Activity 12 : Assessment

 50 Marks

1. Describe how an ionic bond forms in terms of what happens to the valence electrons (2)
2. Describe how a covalent bond forms in terms of what happens to the valence electrons (2)
3. State whether the atoms in the following compounds are held together by ionic or covalent bonds:
	1. Carbon dioxide, CO2
	2. Zinc oxide, ZnO
	3. Lithium chloride, LiCl
	4. Iodine heptafluoride, IF7
	5. Sulfur dioxide, SO2 (5)
4. Write electron dot formulae for the following examples:
	1. HBr
	2. N2
	3. NaCl
	4. O2
	5. F2 (5)
5. Using the electro-negativity information supplied, determine whether the following bonds are polar covalent or non-polar covalent. Using the delta notation label the atoms that are involved in polar covalent bonds.
	1. C-O
	2. N-N
	3. P-I
	4. S-F
	5. H-Br (10)
6. What is the general trend in electro-negativity down a group in the Periodic Table? (1)
7. What is the general trend in electro-negativity across a period (row) in the Periodic Table? (1)
8. Select a key term from the list that corresponds to the following definitions:
	1. A bond composed of one shared electron pair between two atoms …..
	2. The valence electrons in a molecule that are not shared ……
	3. A bond composed of two shared electron pairs between two atoms …
	4. A bond composed of three shared electron pairs between two atoms …
	5. The ability of an atom to attract a shared pair of electrons….
	6. A bond in which the electron pair is shared equally between two atoms…
	7. A bond in which the electron pair is shared unequally between two atoms….
	8. A method used to indicate partial positive and partial negative charges in a covalent bond…
	9. The statement that an atom must be surrounded by eight valence electrons to be stable….
	10. The distance between the nuclei of two atoms joined by a covalent bond…

|  |  |
| --- | --- |
| Double bond | Non-polar covalent bond |
| Electro-negativity | Single bond |
| Nonbonding electrons | Octet Rule |
| Polar covalent bond | Bond length |
| Triple bond | Ionic bond |
| Delta (δ) notation | Bonding electrons |

 (10)

Answers:

1. An ionic bond forms when a metal atom donates its valence electron(s) to a non-metal atom. The metal becomes positively charged and the non-metal negatively charged. The two ions then bond by means of electrostatic attraction.
2. A covalent bond forms between two non-metal elements which share a pair of electrons to form a bond.
3. Bond types:
	1. Covalent
	2. Ionic
	3. Ionic
	4. Covalent
	5. Covalent
4. insert dot formulae
5. Using electro-negativity data:
	1. C-O ΔEN = 3.5 – 2.5 = 1.0 Bond is polar covalent
	2. N-N ΔEN = 3.0 – 3.0 = 0 Bond is non-polar covalent
	3. P-I ΔEN = 2.5 – 2.1 = 0.4 Bond is non-polar covalent
	4. S-F ΔEN = 4.0 – 2.5 = 1.5 Bond is polar covalent
	5. H-Br ΔEN = 2.8 – 2.1 = 0.7 Bond is polar covalent
6. Electro-negativity decreases down a group
7. Electro-negativity increases across a period
8. Answers in bold:
	1. A bond composed of one shared electron pair between two atoms..**single bond**
	2. The valence electrons in a molecule that are not shared…**on-bonding electrons**
	3. A bond composed of two shared electron pairs between two atoms.. **double bond**
	4. A bond composed of three shared electron pairs between two atoms…**triple bond**
	5. The ability of an atom to attract a shared pair of electrons…**electro-negativity**.
	6. A bond in which the electron pair is shared equally between two atoms…**non-polar covalent bond**
	7. A bond in which the electron pair is shared unequally between two atoms…**polar covalent bond**.
	8. A method used to indicate partial positive and partial negative charges in a covalent bond…**delta (δ) notation**
	9. The statement that an atom must be surrounded by eight valence electrons to be stable….**octet rule**
	10. The distance between the nuclei of two atoms joined by a covalent bond…**bond length**

REFLECTION

|  |  |  |
| --- | --- | --- |
| Now that I have finished this unit I understand: | Yes | Need help |
| That ionic bonds form between ions and that ions are charged particles that form when one atom (metal) donates electron(s) to another (non-metal) atom difference between ionic and covalent bonds |  |  |
| That covalent bonds form between non-metal atoms which share electrons to obey the octet (or duet) rule. |  |  |
| That electronegativity is a measure of the power of attraction that a particular nucleus has for the electrons surrounding it |  |  |
| That ionic bonds form when the ∆ EN between the elements forming the bond is greater than 2 |  |  |
| That non-polar covalent bonds form when the ∆ EN between the elements forming the bond is less than 0.5 |  |  |
| That polar covalent bonds form when the ∆ EN between the elements forming the bond is between 1.5 and 2 |  |  |
| The difference between polar and non-polar covalent bonds |  |  |
| How to determine whether bonds are polar or non-polar using the Electro-negativity Table. |  |  |
| That bond length is a measure of the distance between the two nuclei involved in a particular bond |  |  |
| The concept of bond strength and bond dissociation energy |  |  |
| That the Lewis dot notation is a shorthand method of representing the valence electrons of an atom or ion |  |  |
| That dipole notation is used to indicate a polar covalent bond – δ+ means delta positive and shows the part of the molecule that has positive character and δ- means delta negativeand shows the part of the molecule that has negative character  |  |  |
| That oxidation is the loss of electrons and reduction is the gain of electrons |  |  |
| That oxidation numbers allow us to determine which atoms are being oxidised in a reaction and which are being reduced |  |  |
| That oxidation numbers for atoms can be calculated using a set of simple rules |  |  |
| That molecules are not necessarily 2-dimensional, but can have a 3-dimensional shape |  |  |
| That the VSEPR Theory for molecular shape allows us to predict the shape that a molecule will have |  |  |
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Needing help? Speak to your teacher.