SUBJECT and GRADE	Physical Sciences Grade 11
TERM 1	Week 7
TOPIC	Matter and materials: Atomic combinations
AIM OF LESSON	**Summary of content, which you should be familiar with at the end of the series of lessons:** Chemical bond • Definition of: chemical bond; covalent bond; bonding pair; and lone pair. • Drawing of Lewis diagrams. • Determining valence electrons in an atom of an element. • Electrostatic forces between protons and electrons of atoms. • Rules for bond formation. • Formation of dative covalent *(or coordinate covalent)* bonds. Molecular shape: Valence shell electron pair repulsion (VSEPR) theory • Principles used in the VSEPR and usage thereof to classify given molecules as one of the five ideal molecular shapes. • Define / describing: electronegativity; non-polar covalent bond; and polar covalent bond. • Show polarity of bonds using partial charges. • Using polarity and the table of electronegativities to compare chemical bonds. Bond energy and bond length • Definition of: bond energy; and bond length. • Relationship between bond energy and bond length. • Relationship between the strength of a chemical bond and the: length of the bond; size of the bonded atoms; and number of bonds *(single, double, or triple)* between the atoms.

RESOURCES	**Paper-based / physical resources** • Prescribed CAPS Physical Sciences textbook, as well as Siyavula Grade 11 Physical Sciences resource (learner book, pg. 140 - 164); Physical Sciences CAPS document (pg. 67 - 71); and Grade 11 Physical Sciences Examination Guideline (pg. 13 - 15). *(Additional subject-related material, e.g. Mind the Gap, Science Clinic, Answer Series, etc.)*. • Scientific calculator, ruler, pen and pencil.
	Digital resources • Technological devices such as a cell phone, tablet, laptop, etc. and sufficient data would be very useful. WCED ePortal – Website links to access recommended platforms: *http://wcedeportal.co.za/; https://wcedonline.westerncape.gov.za/elearning; https://wcedeportal.co.za/curriculum-support; https://wcedeportal.co.za/partners* • Siyavula links *(atomic combinations)*: *https://intl.siyavula.com/read/science/grade-11/atomic-combinations* • Youtube videos: *https://www.youtube.com/watch?v=Q9-JjyAEqnU* • Mind the Gap: *https://www.education.gov.za/Curriculum/LearningandTeachingSupportMaterials(LTSM)/MindtheGapStudyGuides.aspx*
INTRODUCTION	In grade 10 we started exploring chemical bonding. In grade 11 we will continue studying chemical bonding to gain some understanding why chemical bonding occurs. We looked at the three types of bonding: covalent, ionic and metallic. In the following series of lessons, we will focus mainly on covalent bonding and on the molecules that form as a result of covalent bonding. From previous discussions we came to realise that a model is a representation of what is happening in reality. According to the model of the atom, the atom is made up of a central nucleus, surrounded by electrons that are arranged in fixed energy levels. Within each energy level, electrons move in orbitals of different shapes. The electrons in the outermost energy level of an atom are called the valence electrons. This model of the atom is useful in trying to understand how different types of bonding take place between atoms.

	It is important that you take note of the following aspects discussed previously concerning electrons and energy levels: • Electrons always try to occupy the lowest possible energy level. • The noble gases have a full valence electron orbital. • Atoms form bonds to try to achieve the same electron configuration as the noble gases. • Atoms with a full valence electron orbital are less reactive. In addition to the above mentioned, you also need to revise Lewis diagrams covered in grade 10. This will enable you to predict which atoms will bond with one another and the shape of the molecules being formed. You need to be familiar with the spectroscopic notation when drawing Lewis diagrams. For example, the spectroscopic notation for the configuration of chlorine is: $1s^2s^2p^5$. The condensed form of the last-mentioned would be: $[He]2s^22p^5$. Please note that the spectroscopic notation is an indication of the arrangement of the valence electrons of an element.
CONCEPTS AND SKILLS	***This section must be read in conjunction with the CAPS, pg. 67 - 71.*** **1. LEWIS DIAGRAMS AND BONDS** We can easily draw Lewis diagrams for any element by making use of the number of valence electrons. Too refresh your memory, please refer to the following table: *Table: Lewis diagrams for the elements in period 2.*

Element	Group number	Valence electrons	Spectroscopic notation	Lewis diagram
Lithium	1	1	$[He]2s^1$	Li •
Beryllium	2	2	$[He]2s^2$	Be ••
Boron	13	3	$[He]2s^22p^1$	B •••
Carbon	14	4	$[He]2s^22p^2$	C ••••
Nitrogen	15	5	$[He]2s^22p^3$	N •••••
Oxygen	16	6	$[He]2s^22p^4$	O ••••••
Fluorine	17	7	$[He]2s^22p^5$	F •••••••
Neon	18	8	$[He]2s^22p^6$	Ne ••••••••

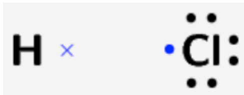
THINGS TO NOTE: Please familiarize yourself with the following content:

- A chemical bond is the physical process that causes atoms to be attracted to one another and to be bound in new compounds.
- Noble gases have a full valence shell. Atoms of other elements bond to try fill their outer valence shell in a similar way.
- There are three forces that act between atoms, which include:
  - attractive forces between the positive nucleus of one atom and the negative electrons of another;
  - repulsive forces between like-charged electrons; and
  - repulsion between like-charged nuclei.
- The energy of a system of two atoms is at a minimum when the attractive and repulsive forces are balanced.

- Lewis diagrams are one way of representing molecular structure. In a Lewis diagram, dots or crosses are used to represent the valence electrons around the central atom.
- A covalent bond is a form of chemical bond where pairs of electrons are shared between two atoms.
- A single bond occurs if there is one electron pair that is shared between the same two atoms.
- A double bond occurs if there are two electron pairs that are shared between the same two atoms.
- A triple bond occurs if there are three electron pairs that are shared between the same two atoms.

Worked out example 1: Lewis diagrams and simple molecules.
QUESTION: Represent hydrogen chloride (HCl) using a Lewis diagram.
SOLUTION
Step 1: For each atom, determine the number of valence electrons in the atom, and represent these using dots and crosses.
The electron configuration of hydrogen is $1s^1$ and the electron configuration for chlorine is $[He]2s^2p^5$. The hydrogen atom has 1 valence electron and the chlorine atom has 7 valence electrons.

The Lewis diagrams for hydrogen and chlorine are:



Step 2: Arrange the electrons so that the outermost energy level of each atom is full.

Hydrogen chloride is represented as follows:

Notice how the two unpaired electrons (one from each atom) form the covalent bond.

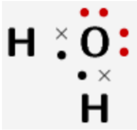
Worked out example 2: Lewis diagrams and simple molecules.
QUESTION: Represent water (H_2O) using a Lewis diagram.

SOLUTION
Step 1: For each atom, determine the number of valence electrons in the atom, and represent these using dots and crosses.

The electron configuration of hydrogen is $1s^1$ and the electron configuration for oxygen is $[He]2s^2p^4$. Each hydrogen atom has 1 valence electron and the oxygen atom has 6 valence electrons.

Step 2: Arrange the electrons so that the outermost  energy level of each atom is full.

The water molecule is represented below:



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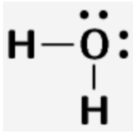


We see that oxygen forms two bonds, one with each hydrogen atom. Oxygen however keeps its electron pairs and does not share them. We can generalize this to any atom. If an atom has an electron pair it will normally not share that electron pair.

Note that a lone pair is an unshared electron pair. A lone pair stays on the atom that it belongs to.

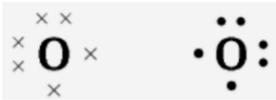
Worked out example 3: Lewis diagram and molecules with double bond.

QUESTION: Represent oxygen (O2) using a Lewis diagram.

SOLUTION

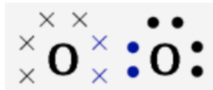
Step 1: For each atom, determine the number of valence electrons that the atom has from its electron configuration.

The electron configuration of oxygen is $[He]2s^2p^4$. Oxygen has 6 valence electrons.



Step 2: Arrange the electrons in the O_2 molecule so that the outermost energy level in each atom is full.

The O_2 molecule is represented below. Notice the two electron pairs between the two oxygen atoms. Because these two covalent bonds are between the same two atoms, this is a double bond.



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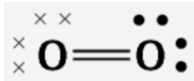


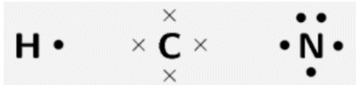
Worked out example 4: Lewis diagram and molecules with triple bond.

QUESTION: Represent hydrogen cyanide (HCN) using a Lewis diagram.

SOLUTION

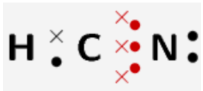
Step 1: For each atom, determine the number of valence electrons that the atom has from its electron configuration. The electron configuration of hydrogen is $1s^1$, the electron configuration of nitrogen is $[He]2s^2p^3$ and for carbon is $[He]2s^2p^2$.

Hydrogen has 1 valence electron, carbon has 4 valence electrons and nitrogen has 5 valence electrons.



Step 2: Arrange the electrons in the HCN molecule so that the outermost energy level in each atom is full.

The HCN molecule is represented below. Notice the three electron pairs between the nitrogen and carbon atom. Because these three covalent bonds are between the same two atoms, this is a triple bond.



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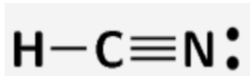
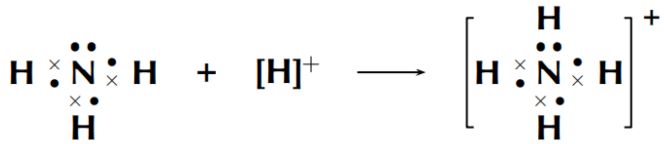


Dative covalent bond

A dative covalent bond is a description of covalent bonding that occurs between two atoms in which both electrons shared in the bond come from the same atom. *(Dative covalent bonds occur between atoms of elements with a lone pair and atoms of elements with no electrons)*.

One example of a molecule that contains a dative covalent bond is the ammonium ion (NH_4^+) shown in the figure below. The hydrogen ion H+ does not contain any electrons, and therefore the electrons that are in the bond that forms between this ion and the nitrogen atom, come only from the nitrogen.



2. VALENCE SHELL ELECTRON PAIR REPULSION (VSEPR) THEORY

Valence shell electron pair repulsion (VSEPR) theory is a model in chemistry, which is used to predict the shape of individual molecules. *(VSEPR is based upon minimising the extent of the electron-pair repulsion around the central atom being considered).*

To predict the shape of a covalent molecule, follow these steps:

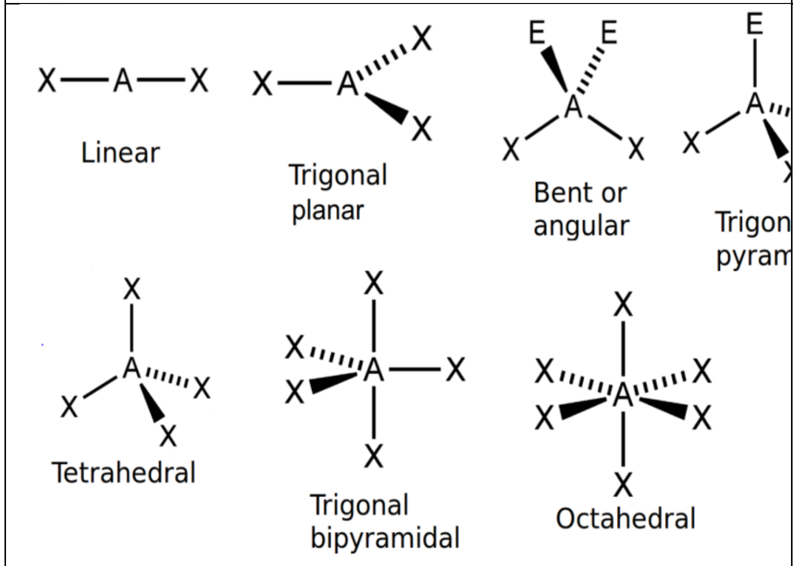
- Draw the molecule using a Lewis diagram. Make sure that you draw all the valence electrons around the molecule's central atom.
- Count the number of electron pairs around the central atom.
- Determine the basic geometry of the molecule using the table below. For example, a molecule with two electron pairs *(and no lone pairs)* around the central atom has a linear shape, and one with four electron pairs *(and no lone pairs)* around the central atom would have a tetrahedral shape.

The table below gives the common molecular shapes. In this table we use A to represent the central atom, X to represent the terminal atoms (i.e. the atoms around the central atom) and E to represent any lone pairs.

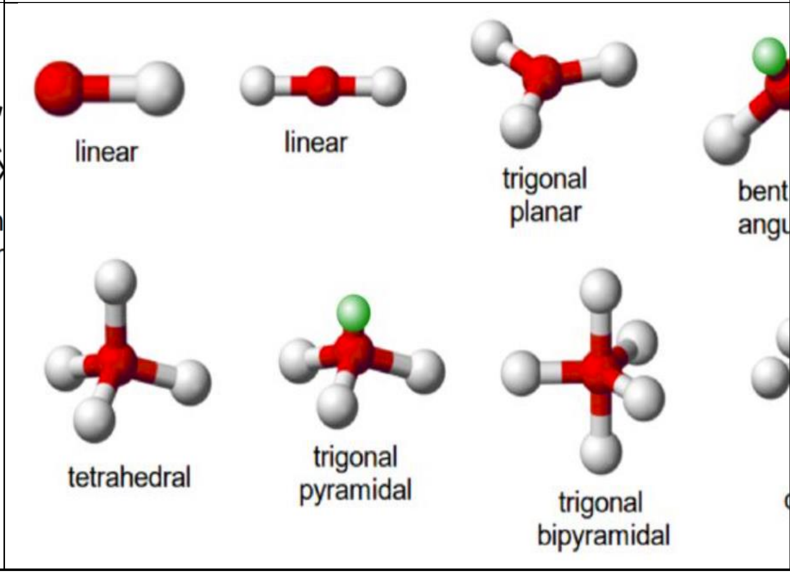
Number of bonding electron pairs	Number of lone pairs	Geometry	General formula
1 or 2	0	linear	AX or AX_2
2	2	*bent or angular*	AX_2E_2
3	0	trigonal planar	AX_3
3	1	*trigonal pyramidal*	AX_3E
4	0	tetrahedral	AX_4
5	0	trigonal bipyramidal	AX_5
6	0	octahedral	AX_6

Table: Molecular shapes

The common molecular shapes.



The common molecular shapes in 3-D.



3. ELECTRONEGTIVITY

Electronegativity is a chemical property which describes the power of an atom to attract electrons towards itself in a chemical.

The greater the electronegativity of an atom of an element, the stronger its attractive pull on electrons. For example, in a molecule of hydrogen bromide (HBr), the electronegativity of bromine (2,8) is higher than that of hydrogen (2,1), and so the shared electrons will spend more of their time closer to the bromine atom. Bromine will have a slightly negative charge, and hydrogen will have a slightly positive charge. In a molecule like hydrogen (H_2) where the electronegativities of the atoms in the molecule are the same, both atoms have a neutral charge.

Table of electronegativities for selected elements.

Element	Electronegativity	Element	Electronegativity
Hydrogen (H)	2,1	Lithium (Li)	1,0
Beryllium (Be)	1,5	Boron (B)	2,0
Carbon (C)	2,5	Nitrogen (N)	3,0
Oxygen (O)	3,5	Fluorine (F)	4,0
Sodium (Na)	0,9	Magnesium (Mg)	1,2
Aluminium (Al)	1,5	Silicon (Si)	1,8
Phosphorous (P)	2,1	Sulfur (S)	2,5
Chlorine (Cl)	3,0	Potassium (K)	0,8
Calcium (Ca)	1,0	Bromine (Br)	2,8

Electronegativity and bonding

Electronegativity can be used to explain the difference between two types of covalent bonds, which include: polar covalent bonds *(between non-identical atoms)*; and non-polar covalent bonds *(between identical atoms or atoms with the same electronegativity)*.

Polar molecule

A polar molecule is one that has one end with a slightly positive charge, and one end with a slightly negative charge. *Examples include water, ammonia and hydrogen chloride*.

Non-polar molecule

A non-polar molecule is one where the charge is equally spread across the molecule or a symmetrical molecule with polar bonds.

The table below lists the approximate values. Although we have given ranges here bonding is more like a spectrum than a set of boxes.

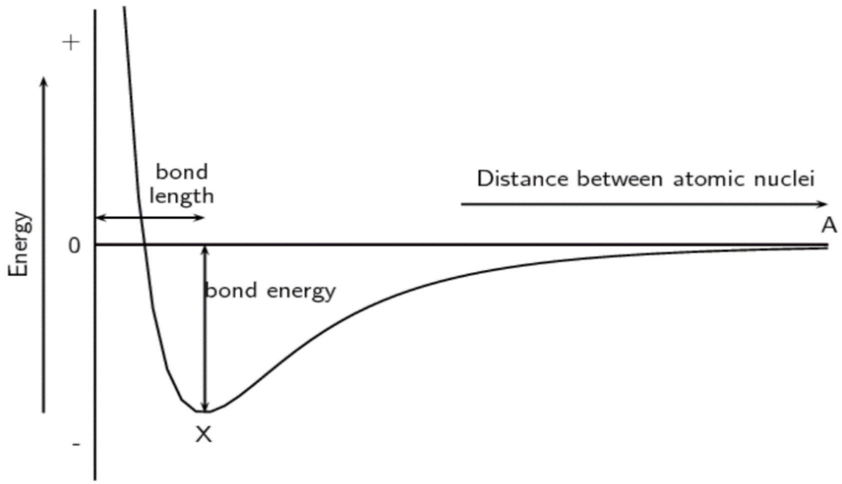
Electronegativity difference	Type of bond
0	Non-polar covalent
0 - 1	Weak polar covalent
1,1 - 2	Strong polar covalent
$> 2,1$	Ionic

4. ENERGY AND BONDING

The energy changes that occur as atoms come together, can be graphically represented. The accompanying graph shows the change in energy that takes place as atoms move closer together.

Bond length is the distance between the nuclei of two atoms when they bond.

Bond energy is the amount of energy that must be added to the system to break the bond that has formed.



	Bond strength means how strongly one atom attracts and is held to another atom. Bond strength depends on the length of the bond, the size of the atoms and the number of bonds between the two atoms. Thus: - The shorter the bond length, stronger the bond. - The smaller the atoms involved, the stronger the bond. - The more bonds that exist between the same atoms, the stronger the bond. *Extracted and summarized from: Siyavula Grade 11 Physical Sciences resource (learner book, pg. 140 - 164); Physical Sciences CAPS document (pg. 67 - 71); and Grade 11 Physical Sciences Examination Guideline (pg. 13 - 15).*

EXERCISES FOR CONSOLIDATION PURPOSES

Please use ample time to complete the following activities (questions 1 to 3), which will aid in preparing you for tests / examinations in future.

QUESTION 1

Four options are provided as possible answers to the following questions. Each question has only ONE correct answer. Choose the answer and write only the letter (A–D) next to the question number (1.1 – 1.3).

1.1 The type of bond formed between a H^+ ion and H_2O is called a/an …

A hydrogen bond.
B dative covalent bond.
C ionic bond.
D covalent bond.

1.2 The shape of the molecule in which the central atom is surrounded by two lone pairs and two bonding pairs is …

A linear.
B trigonal planar.
C tetrahedral.
D bent.

1.3 The bond energy of a C–Cℓ bond is 338 kJ.mol^{-1} whereas the bond energy of a C–I bond is 238 kJ.mol^{-1}. The difference in bond energy exists because …

A the bond length of the C–Cℓ bond is greater than that of the C–I bond.
B chlorine is more electronegative than iodine.
C the bond length of the C–I bond is greater than that of the C–Cℓ bond.
D the chlorine atom is bigger than the iodine atom. (2)

[3 x 2 = 6]

QUESTION 2

Electronegativity of atoms may be used to explain the polarity of bonds.

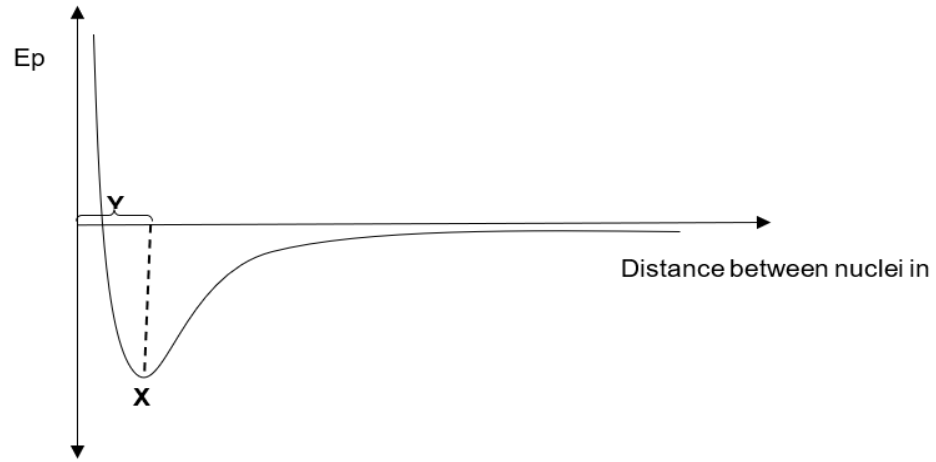
2.1 Define the term electronegativity. (2)

2.2 Draw the Lewis diagram of an oxygen difluoride molecule. (2)

2.3 Calculate the electronegativity difference between O and F in oxygen difluoride and predict the polarity of the bond. (2)

2.4 A polar bond does not always lead to a polar molecule. Explain the statement by referring to OF_2 and CO_2 molecules. In your explanation, include the polarity of the bonds and the shape of the molecules. (4)

2.5 The diagram below shows the energy change that takes place when two atoms move towards each other.



2.5.1 What does X and Y represent? (2)

2.5.2 Define the concept represented by X. (2)

2.5.3 Explain the relationship between bond order, bond length and bond energy. (3) **[17]**

QUESTION 3

Molecules such as CO_2 and H_2O are formed through covalent bonding.

3.1 Define the term covalent bonding. (2)

3.2 ONE of the above molecules has lone pairs of electrons on the central atom. Draw the Lewis diagram for this molecule. (2)

3.3 H_3O+ is formed when H_2O forms a dative covalent bond with an H^+ ion.

3.3.1 Draw the Lewis diagram for H_3O^+. (1)
3.3.2 State TWO conditions for the formation of such a bond. (2)

3.4 The polarity of molecules depends on the DIFFERENCE IN ELECTRONEGATIVITY and the MOLECULAR SHAPE.

3.4.1 Define the term electronegativity. (2)

3.4.2 Calculate the difference in electronegativity between:
(a) C and O in CO_2 (1)
(b) H and O in H_2O (1)

3.5 Explain the difference in polarity between CO_2 and H_2O by referring to the polarity of the bonds and the shape of the molecules. (6) **[17]**

TOTAL = 40

Link to the answers: https://drive.google.com/file/d/1AfUOKWPsR4ZGFCwCpQqjn5qQwtL9PfJd/view?usp=sharing

CONSOLIDATION	**Summary of lesson content, which you should be familiar with at this stage:**
	Chemical bond
	• Define a chemical bond as a mutual attraction between two atoms resulting from the simultaneous attraction between their nuclei and the outer electrons.
	• Determine the number of valence electrons in an atom of an element.
	• Explain, in terms of electrostatic forces between protons and electrons, and in terms of energy considerations, why:
	  - Two H atoms form an H_2 molecule
	  - He does not form He_2
	  - Interpret the graph of potential energy versus the distance between nuclei for two approaching hydrogen atoms.
	• Define a covalent bond as the sharing of electrons between two atoms to form a molecule.
	• Draw Lewis-diagrams, given the formula and using electron configurations, for simple molecules.
	• Describe rules for bond formation.
	• Define a bonding pair as a pair of electrons that is shared between two atoms in a covalent bond.
	• Define a lone pair as a pair of electrons in the valence shell of an atom that is not shared with another atom.
	• Describe the formation of the dative covalent *(or coordinate covalent)* bond by means of electron diagrams using NH_4^+ and H_3O^+ as examples.
	
	Molecular shape: Valence shell electron pair repulsion (VSEPR) theory
	• State the major principles used in the VSEPR.
	• Use the VSEPR theory to classify given molecules as one of the five ideal molecular shapes by finding the number of atoms bonded to the central atom in molecules where there are NO lone pairs on the central atom.
	• Use the VSEPR theory to determine the shapes of molecules with lone pairs on the central atom (H_2O, NH_3, SO_2) and that CANNOT have one of the ideal shapes.
	• Define electronegativity as a measure of the tendency of an atom in a molecule to attract bonding electrons.
	• Describe a non-polar covalent bond as a bond in which the electron density is shared equally between the two atoms.
	• Describe a polar covalent bond as a bond in which the electron density is shared unequally between the two atoms.
	• Show polarity of chemical bonds using partial charges.

	• Compare the polarity of chemical bonds using a table of electronegativities. Bond energy and bond length • Define bond energy of a compound as the energy needed to break one mole of its molecules into separate atoms. • Define bond length as the average distance between the nuclei of two bonded atoms. • Explain the relationship between bond energy and bond length, i.e. bonds with a shorter bond length require more energy to break than bonds with a longer bond length. • Explain the relationship between the strength of a chemical bond between two atoms and the: - Length of the bond between them. If the force of attraction between two atoms is strong, the nuclei come very close together resulting in a short bond length. - Size of the bonded atoms. The bond length between larger atoms is longer than the bond length between smaller atoms. - Number of bonds *(single, double, triple)* between the atoms. Bond strength increases as the number of bonds between atoms increases, i.e. triple bonds are stronger than double bonds, which are stronger than single bonds.
VALUES / APPLICATIONS IN PRACTICE	Visit the following weblink and watch the video illustrating the value of using the VSEPR-theory in modern molecular geometrics: https://www.youtube.com/watch?v=Q9- JjyAEqnU