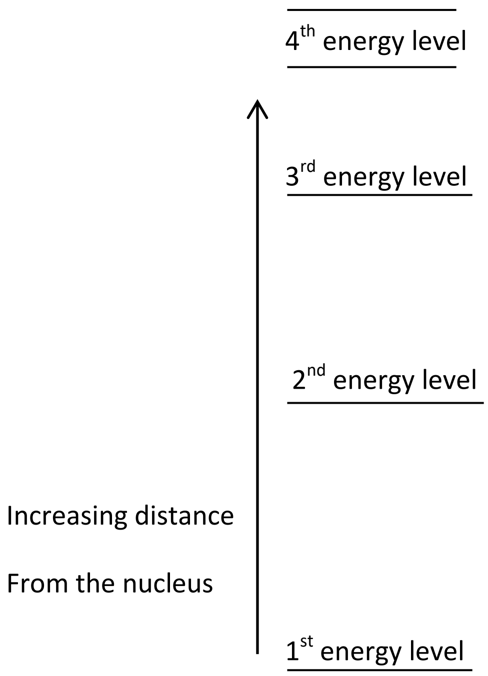
Week 3 Notes

ELECTRON ARRANGEMENT

- The energy of electron is **quantised,** i.e they can only have certain energies with no energy values in between.
- Electrons occupy certain **energy levels** around the nucleus.

Electron energy levels



- Electrons in atoms **ALWAYS** occupy lower energy levels before they occupy higher energy levels.

- There is a limit to the number of electrons which can occupy any one level and the following octet rules apply to the first 20 elements.

Octet Rule

1. The first energy level can hold a maximum of only 2 electrons.
2. The second level holds a maximum of 8 electrons.
3. The third level holds a maximum of 8 electrons.
4. The fourth level holds the remaining electrons.

- The way electrons are arranged is called the **electron configuration (electron arrangement)**

- Electron configuration can be written in brackets after the element symbol with the number of electrons in each level separated by commas.

Electron Configurations

Example 1:

Carbon

The atomic number of carbon is 6. Carbon has 6 six electrons and six protons:

- The first two electrons are placed in the first energy level.
- The remaining four are in the second energy level.
- The electron configuration for Carbon is written C (2, 4).

Example 2:

Argon

The atomic number of argon is 18. Argon has 18 electrons.

- The first two electrons go into the first energy level.
- Eight go into the second energy level.
- The last eight go into the third energy level.
- The electron configuration for argon is written Ar (2, 8, 8).

S P D F ORBITALS

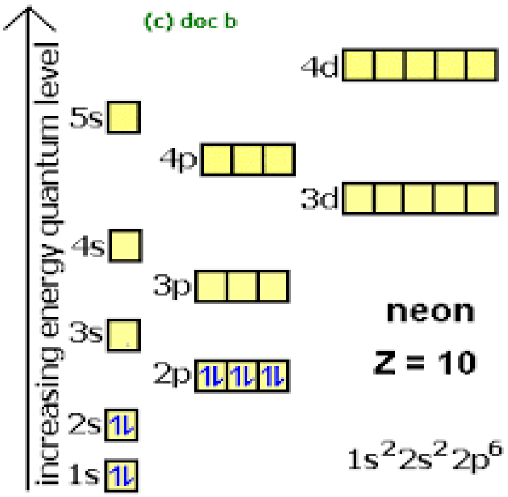
- s, p, d, and f are names given to orbitals that hold the electrons in atoms.
- These orbitals have different shapes and energies.

Value of S P D F

- S subshell can hold a total of two electrons.
- P subshell can hold a total of six electrons.
- D subshell can hold a total of 10 electrons.
- F subshell can hold a total of 14 electrons.

NOTE: To write electron configuration of an atom, identify the energy level and write the number of electrons in the energy level as its superscript.

Example 3:



IONS OF ATOMS IN THE PERIODIC TABLE

IONS

The gain or loss of electrons by atoms produces **ions**. When atoms gain or lose electrons they produce ions.

Ions are either cations or anions:

- **Cations (**positively charge ions) form when atoms or groups of (covalently bonded) atoms lose electrons. Cations have few electrons than protons,

Example 4:

(i) Magnesium ions, Mg^{2+}, have 10 electrons and 12 protons whereas the atoms of the element, Mg, have 12 elements and 12 protons,

(ii)Aluminium ions, Al^{3+}, have 10 electrons and 13 protons, whereas atoms of the element, Al, have 13 electrons and 13 protons.

- **Anions** (negatively charged ions) form when atoms or groups of (covalently bonded) atoms gain electrons. Anions have more electrons than protons.

Example 5:

(i) Chloride ions, Cl^-, have 18 electrons, whereas atoms of the element, Cl, have 17 electrons and 17 electrons.

(ii) (ii) Oxide ions, O^{2-}, have 10 electrons, whereas atoms of the element, O, have 8 electrons.

Ions are written with the charge shown on the *right* and *above* the symbol for the atom or group of atoms.

Example 6: Ions

The sodium ion is written Na^+, which indicates that a sodium atom loses one electron when it forms a sodium cation.

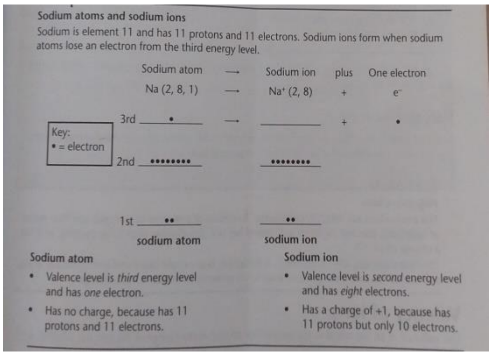
The oxygen ion is written O^{2-}, which indicates that an oxygen atom gains two electrons when it forms an oxide anion.

Atoms gain or lose electrons to produce stable (full) valence shells.

<u>Cations</u>

Metals are on the left of the periodic table in groups 1, 2 and 13, and tend to lose electrons to form **monatomic** (‘one atom’) cations.

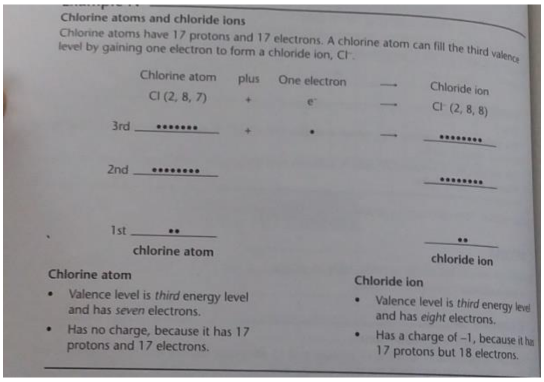
Example 7:



<u>Anions</u>

Non-metals, which exist in groups 15, 16 and 17, tend to gain electrons to form monatomic anions. These ions are named using the suffix -*ide*, eg. Chloride

Example 8:



NOTE: The group 18 elements such as neon and argon are atoms with stable valence levels so they are unlikely to form ions.

Polyatomic Ions

- Ions that consist of a stable group of atoms with an overall charge.

Example 9: Ammonium ion, $NH4^+$

It consist of one atom of Nitrogen and four atoms of hydrogen, bonded together and overall has one less electron than it has proton, so it has a charge of +1.

Diatomic Ions

- A polyatomic ion with two atoms

Example 10: Hydroxide ion, OH^-

It comprises one oxygen atom and one hydrogen atom and overall has one more electron than it has protons, so it has a charge of -1.

Valency of Ions

- Refers to the magnitude or size of the charge of the ion, not whether it is positive or negative.

Example 11: Valency

The sodium ion, $Na^{+,}$ chloride ion, Cl^- and ammonium ion, $NH4^+$, all have a valency of 1. The aluminium ion, Al^{3+} has a valency of 3 and the oxide ion, O^{2-} has a valency of 2.

The **transition metals** such as copper, manganese, and iron form cations which can have more than one valency.

A roman numeral is used to indicate the valency of a transition metal ion.

Exercise 5: Ions.

1. Write the symbols for each of the following ions:

a. Lithium ion
b. Bromide ion
c. Sulfide ion
d. Hydroxide ion
e. Copper ion

2. For each of the ions:

a. K^+ b. O^{-2} c. F^-

i. Name the ion and write down its electron arrangement.
ii. State whether the iron formed from the corresponding atom by gain or loss of electrons.

3. Write the :

a. Electron arrangement for calcium atom.

b. Electron arrangement for calcium ion.

c. Electron arrangement for sulphide ion.

4. Write the names and give the valency for each of the following ions:

a. Mg^{2+}

b. Fe^{3+}

c. Cu^{2+}

d. F-

e. O^{2-}

BONDS, FORMATION AND NATURE

i. Define:	
b. Chemical bond	1
c. Electrostatic attraction	1
d. Bonding electrons	1
e. Octet rule	1
iii. Describe the octet rule.	2
iv. Explain why atoms form a chemical bond.	3
v. Identify the types of electrons involved in chemical bonding between two atoms.	1
vi. Define:	
a. Ionic bond	1
b. Covalent bond	1
vii. Give examples of substances that have:	
a. Ionic bond	1
b. Covalent bond	1
viii. Describe how ionic bond forms using Lewis structures.	2
ix. Describe how covalent bond forms using Lewis structures.	2
x. Differentiate between covalent and ionic bonds using examples.	3
xi. Describe the convention used in drawing Lewis Structures of any given compound.	2
xii. Draw the Lewis structure (or electron dot diagram) of any given compound. *(including double and triple bonds. Limit to 4 valence electron pairs)*	3

Chemical Bonding

A chemical bonds are forces(electrostatic force) that hold atoms together to make compounds or molecules.

TWO WAYS TO FORM COMPOUNDS

1. Electrons are shared between atoms (covalent bond).

2. Electrons are transferred between atoms (ionic bond).

WHY DO ATOMS BOND TOGETHER?

- Atoms combine to form molecules because molecules are energetically more stable than the individual atoms which form them.
- All other atoms bond together to become more stable.
- Nobel gases have very stable electron configuration, because their outermost shells are full.

OCTET RULE:

- Atoms other than hydrogen tend to lose, gain, or share electrons until they are surrounded by 8 valence electrons.
- The filled valence shell (8electrons) of all of the inert gas atoms except **helium** is commonly called **'an octet of electrons'.**

NOTE: When an atom of one element combines chemically with an atom of another element both atoms usually attain a stable outer shell having a noble gas configuration. This kind of electronic structure has chemical stability.

TYPES OF CHEMICAL BONDS

- Chemical bonds include both strong intramolecular forces (covalent and ionic bonds) and weak intermolecular forces (London dispersion forces, dipole-dipole interactions and hydrogen bonding).
- **Intramolecular forces** are forces of attraction that exist **within** molecules. **Intermolecular forces** are forces of attraction exist **between** molecules.

Types of Intramolecular Forces

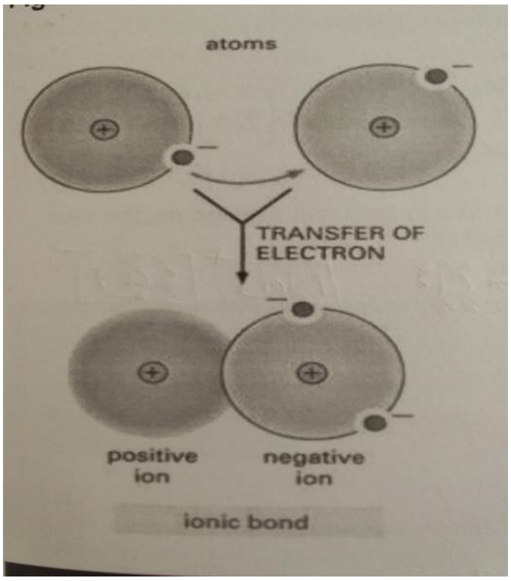
- Ionic Bonding
- Covalent Bonding
- Metallic Bonding

Note: In this unit we will mainly focus on ionic and covalent bonding. You will learn more about metallic bonding crystals later in the course.

IONIC BONDING

- Where an atom of a metallic element transfers one or more valence electron(s) to an atom of a non-metallic element. The metal becomes a positive ion, the non-metal a negative ion. The unlike charges attract the different ions together to form a crystal lattice.
- **Occurs** when electrons **transferred** from the outer shell of one atom to the outer shell of another atom.
- Strong electrostatic force of attraction between oppositely charges ions.
- The atom losing electrons forms positive ions. Refer to Figure 1 below.

Figure 1: Ionic Bond



Example 1

A metal from group 1, 2, or 3 can combine with a non-metal from group 6 or 7 (halogens).

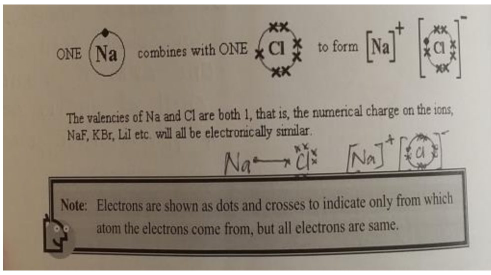
Sodium and chloride combine to form salt sodium chloride (NaCl). **How is this possible?**

Explanation:

Sodium atom with an electron configuration (2, 8, 1) has a valence electron of 1. Chlorine atom with an electron configuration (2, 8, 7) has a valence electron of 7. Thus the sodium atom can lose this one electron and become sodium ion (Na^+). The chlorine atom can gain one electron and become chloride ion (Cl^-). Thus a strong electrostatic attraction between Na^+ and Cl^- ions (opposite charged) will lead to chemical bonding of the ions thus the formation of sodium chloride. Na^+ has a noble gas configuration, that is, after losing one electron, the outermost shell will have 8 electrons (fully filled), similar to neon. Cl^- has a noble gas configuration, that is after gaining one electron, the outermost shell will have 8 electrons (fully filled) similar to argon.

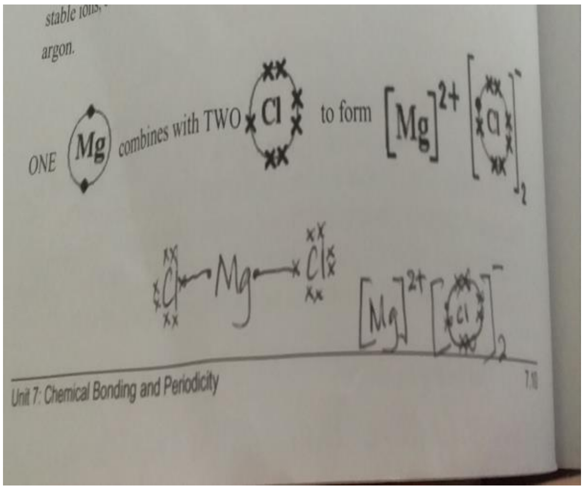
Summary: **Word Equation**

1. Sodium + chlorine $\rightarrow$sodium chloride
   Ionic Equation
   Na^++ Cl^- $\longrightarrow$ Na+Cl-
   Summarised electronically
   Na (2,8,1) + Cl (2,8,7) Na^+(2,8) Cl^- (2,8,8)
2. Electronic structure diagram can be used to show the same. Only outer electrons will be shown. Thus:

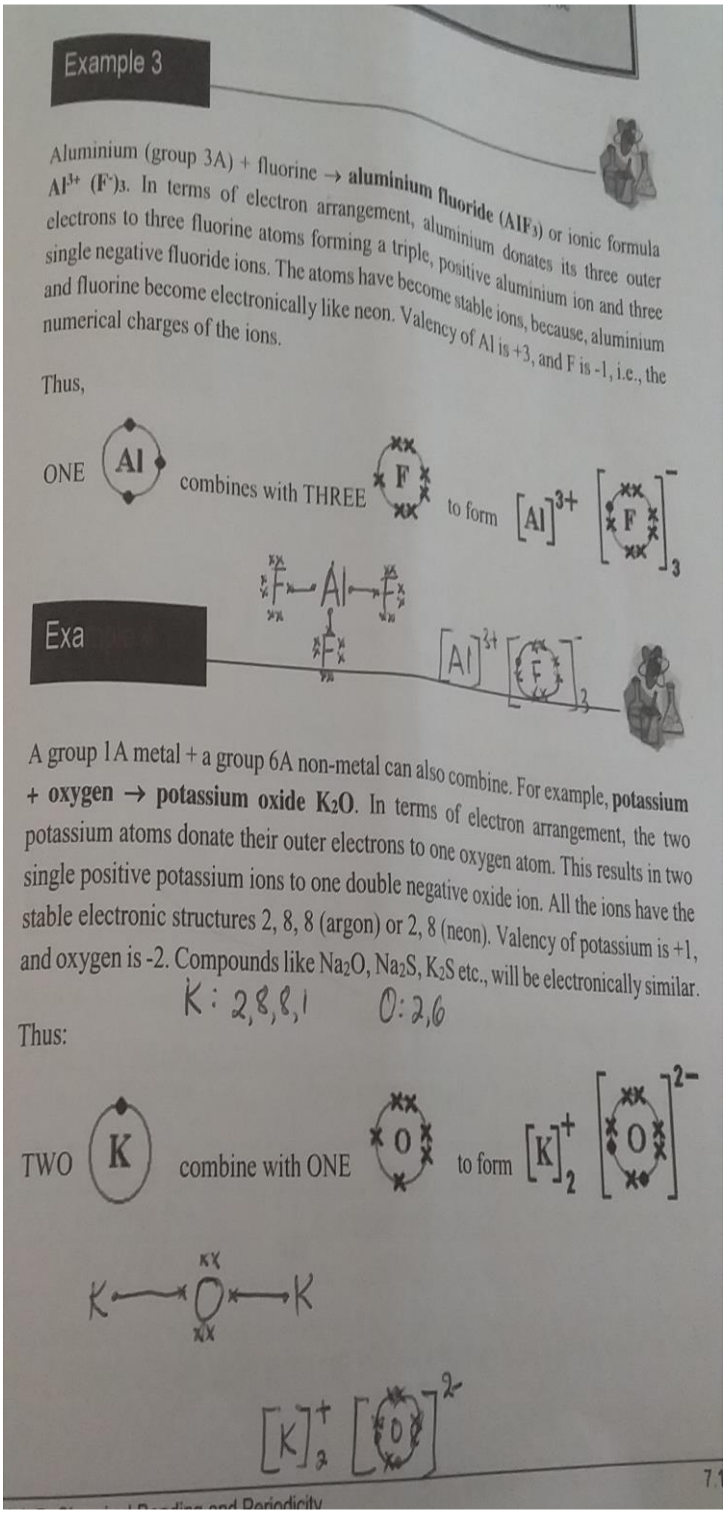


Example 2

Magnesium reacts with chlorine to form magnesium chloride (MgCl2). In terms of electron arrangement, the magnesium donates it two outer electrons to two chlorine atoms(one atom gains 1 electron each) forming a double positive magnesium ion and two single negative chloride ions. The atoms have become stable ions, because electronically magnesium becomes like neon and chlorine like argon.



Note: You can draw two separate chloride ions, but in this example a number subscript has been used, as in ordinary chemical formula. (BeF_2, $MgBr_2$, $CaCl_2$ or CaI_2 will all be electronically similar).



Example 3

Aluminium (group 3A) + fluorine → **aluminium fluoride (AlF_3)** or ionic formula $Al^{3+}(F^-)_3$. In terms of electron arrangement, aluminium donates its three outer electrons to three fluorine atoms forming a triple, positive aluminium ion and three single negative fluoride ions. The atoms have become stable ions, because, aluminium and fluorine become electronically like neon. Valency of Al is +3, and F is -1, i.e., the numerical charges of the ions.

Thus,

ONE Al combines with THREE F to form $[Al]^{3+}[F]^-_3$

Exa

A group 1A metal + a group 6A non-metal can also combine. For example, **potassium + oxygen → potassium oxide K_2O**. In terms of electron arrangement, the two potassium atoms donate their outer electrons to one oxygen atom. This results in two single positive potassium ions to one double negative oxide ion. All the ions have the stable electronic structures 2, 8, 8 (argon) or 2, 8 (neon). Valency of potassium is +1, and oxygen is -2. Compounds like Na_2O, Na_2S, K_2S etc., will be electronically similar.

K: 2,8,8,1 O: 2,6

Thus:

TWO K combine with ONE O to form $[K]^+_2[O]^{2-}$

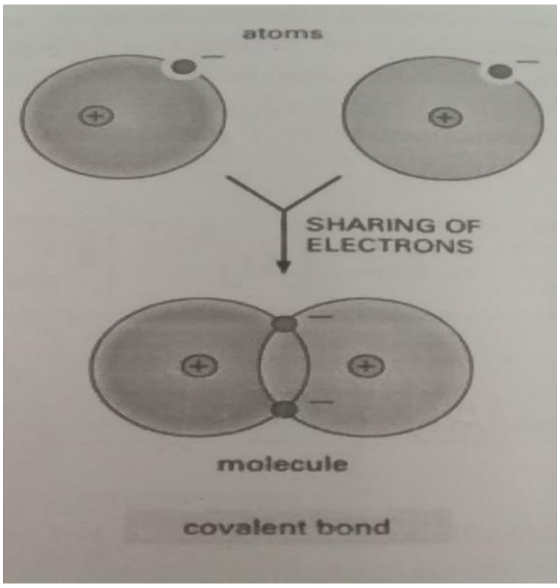
K — O — K

$[K]^+_2 [O]^{2-}$

Covalent Bonding

- Where two or more atoms (usually from non-metals) form bonds by **sharing electrons** between them. Because atoms have similar affinity for electrons and neither has a tendency to donate them, they share electrons in order to achieve octet configuration and become more stable.
- This type of bond is usually formed between **two non-metallic** elements.
- Caused by mutual electrical attraction between the two positive nuclei of the two atoms of the bond, and electrons between them. Refer to Figure 2 below.

Figure 2: Covalent Bond



Note: The shared electrons, called **bonding electrons**. It comes from the valence shells of the atoms sharing electrons and are found between the nuclei.

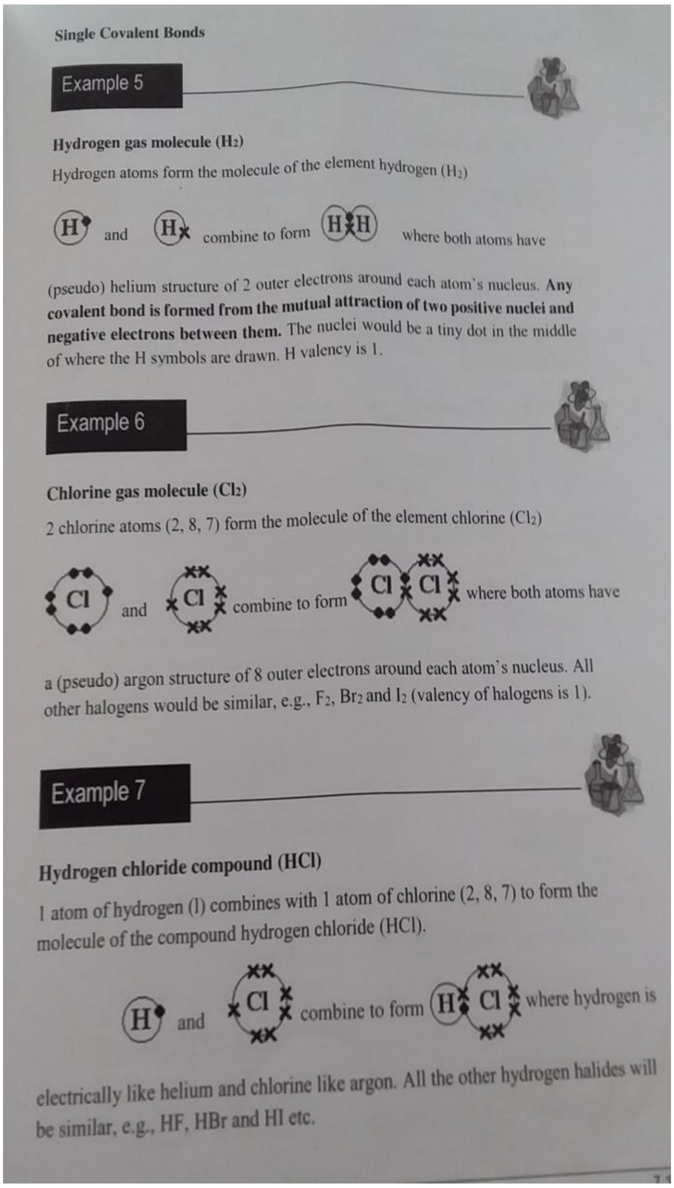
When groups of atoms bond covalently, **molecules** are formed such as hydrogen gas H_2, methane CH_4 and carbon dioxide CO_2.

The number of pairs of electrons shared between two atoms determines the type/number of the covalent bond formed between them, refer to Table 1.

Table 1: Single and Multiple Covalent Bonds.

Number of electron pairs shared	Type of covalent bond formed	Examples
1	Single	H – H or H – Cl
2	Double	C = C or O = O
3	Triple	C ≡ C or C ≡ N

Now let's study a couple of examples and the type of covalent bonds formed.



Single Covalent Bonds

Example 5

Hydrogen gas molecule (H_2)

Hydrogen atoms form the molecule of the element hydrogen (H_2)

H and H combine to form HH where both atoms have

(pseudo) helium structure of 2 outer electrons around each atom's nucleus. **Any covalent bond is formed from the mutual attraction of two positive nuclei and negative electrons between them.** The nuclei would be a tiny dot in the middle of where the H symbols are drawn. H valency is 1.

Example 6

Chlorine gas molecule (Cl_2)

2 chlorine atoms (2, 8, 7) form the molecule of the element chlorine (Cl_2)

Cl and Cl combine to form Cl Cl where both atoms have

a (pseudo) argon structure of 8 outer electrons around each atom's nucleus. All other halogens would be similar, e.g., F_2, Br_2 and I_2 (valency of halogens is 1).

Example 7

Hydrogen chloride compound (HCl)

1 atom of hydrogen (1) combines with 1 atom of chlorine (2, 8, 7) to form the molecule of the compound hydrogen chloride (HCl).

H and Cl combine to form H Cl where hydrogen is

electrically like helium and chlorine like argon. All the other hydrogen halides will be similar, e.g., HF, HBr and HI etc.

Double Covalent Bonds

Example 8

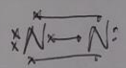
Oxygen gas (O_2) molecule

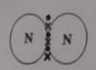
Two atoms of oxygen (2, 6) combine to form the molecules of the element oxygen O_2. The molecule has double covalent bond as O=O. Oxygen valency is 2.

Triple Covalent Bonds

Example 9

Nitrogen gas (N_2) molecule





$\longrightarrow$ N ≡ N

N≡N

6 electrons (3 pairs) shared between the two nitrogen atoms.

IMPORTANT NOTES:

- Electronic structure diagrams or electron dot/cross structure can be referred to as Lewis Structure.
- Covalent bonds or ionic bonds can be represented by a dash (-). A dash indicates a shared pair of electron (**structural formula**).
- Formulae written showing actual composition of atoms is molecular formula, **e.g. H2 or HCl.**
- **Atomicity** is the number of atoms present in one molecule of an element.

Atomicity	**Number of atoms**	**Examples of elements**
Monoatomic	1	He or Na
Diatomic	2	H_2 or O_2
Triatomic	3	O_3
Polyatomic	3 or more	S_8

- Sharing of electrons between bonding atoms can either be equal or unequal. It depends on the type of bonding atoms and their electronegativity.

Properties of Ionic and Covalent Compounds

Substances with ionic bonds can be called **ionic compounds** (have ions) and substances with covalent bonds are called **covalent compounds**. Covalent compounds could either be simple or giant covalent compounds.

Types of Covalent Molecules/Compounds

1. Small covalent molecules like hydrogen (H_2) and chlorine (Cl_2) gases or compounds like hydrogen chloride (HCl), water (H_2O), ammonia (NH_3), methane (CH_4) and carbon dioxide (CO_2) gases.
2. Large covalent molecules like diamond, graphite (made of carbons) or silicon dioxide (SiO_2).

Table 7.2 summarises and compares the properties of ionic and covalent compounds.

Table 7.2: Properties of Ionic and Covalent Compounds

Ionic Compounds	Covalent Compounds
1. Have high melting and boiling points (due to strong electrostatic forces of attraction between ions).	1. Simple covalent have relatively low melting and boiling points (due to weak forces between molecules). Giant covalent usually have high melting and boiling points. The covalent bonds need to be broken.
2. Conduct heat and electricity when in molten or in solution state but not when in solid state. (Ions are free to move in molten state only).	2. Simple covalent do not conduct electricity as the overall charge of molecule is zero. Giant covalent (graphite) conducts electricity due to mobile electrons.
3. Soluble in polar solvent only. (e.g., in water). Due to attraction of the atoms on the water molecules for the positive or negative ions.	3. Soluble in non-polar solvents only. (e.g., organic solvents).

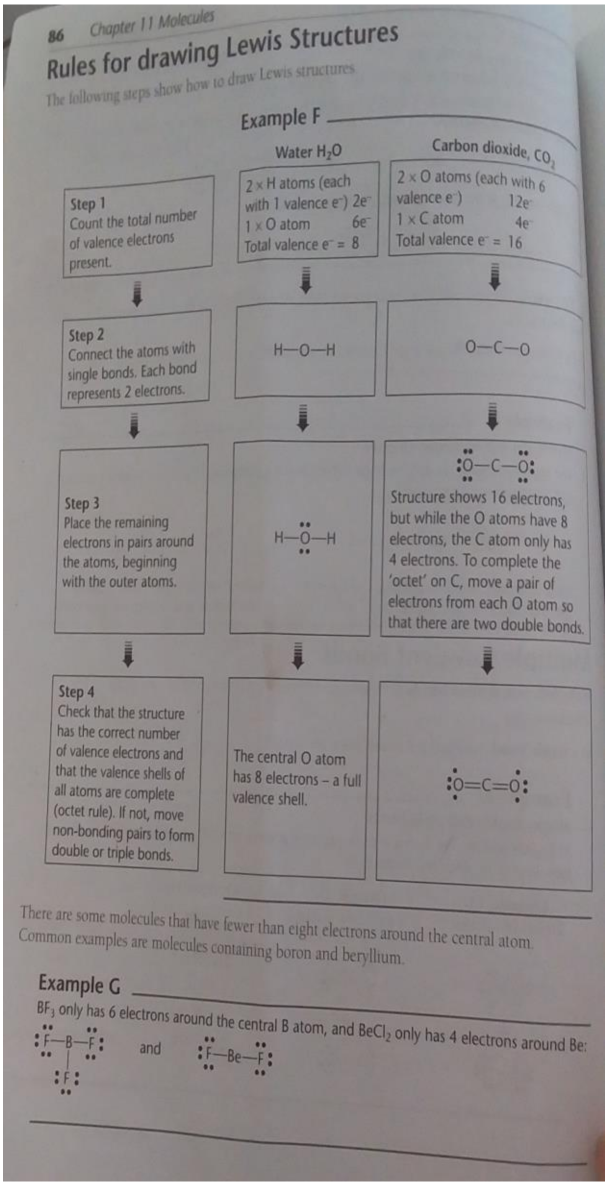
Exercise 6:

1. Why do atoms need to bond with other atoms?
2. Explain what does 'noble gas' configuration mean.
3. Define terms ionic bonding and covalent bonding using suitable examples.
4. Using electronic structure diagrams explain how chemical bonding occurs in the following molecules and also state the type of bonding.

a. Potassium chloride (KCl)

b. Hydrogen chloride (HCl)

c. Ocygen gas (O_2)



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Rules for drawing Lewis Structures

The following steps show how to draw Lewis structures

Example F

	Water H_2O	Carbon dioxide, CO_2
Step 1 Count the total number of valence electrons present.	2 × H atoms (each with 1 valence e^-) $2e^-$; 1 × O atom $6e^-$; Total valence e^- = 8	2 × O atoms (each with 6 valence e^-) $12e^-$; 1 × C atom $4e^-$; Total valence e^- = 16
Step 2 Connect the atoms with single bonds. Each bond represents 2 electrons.	H—O—H	O—C—O
Step 3 Place the remaining electrons in pairs around the atoms, beginning with the outer atoms.	H—Ö—H	:Ö—C—Ö: Structure shows 16 electrons, but while the O atoms have 8 electrons, the C atom only has 4 electrons. To complete the 'octet' on C, move a pair of electrons from each O atom so that there are two double bonds.
Step 4 Check that the structure has the correct number of valence electrons and that the valence shells of all atoms are complete (octet rule). If not, move non-bonding pairs to form double or triple bonds.	The central O atom has 8 electrons – a full valence shell.	:O=C=O:

There are some molecules that have fewer than eight electrons around the central atom. Common examples are molecules containing boron and beryllium.

Example G

BF_3 only has 6 electrons around the central B atom, and $BeCl_2$ only has 4 electrons around Be:

:F—B—F: and :F—Be—F:

Exercise 7: Lewis Structures

1. Draw Lewis structure for each of the following molecules.

a. i. H_2 ii. N_2 iii. HOCl

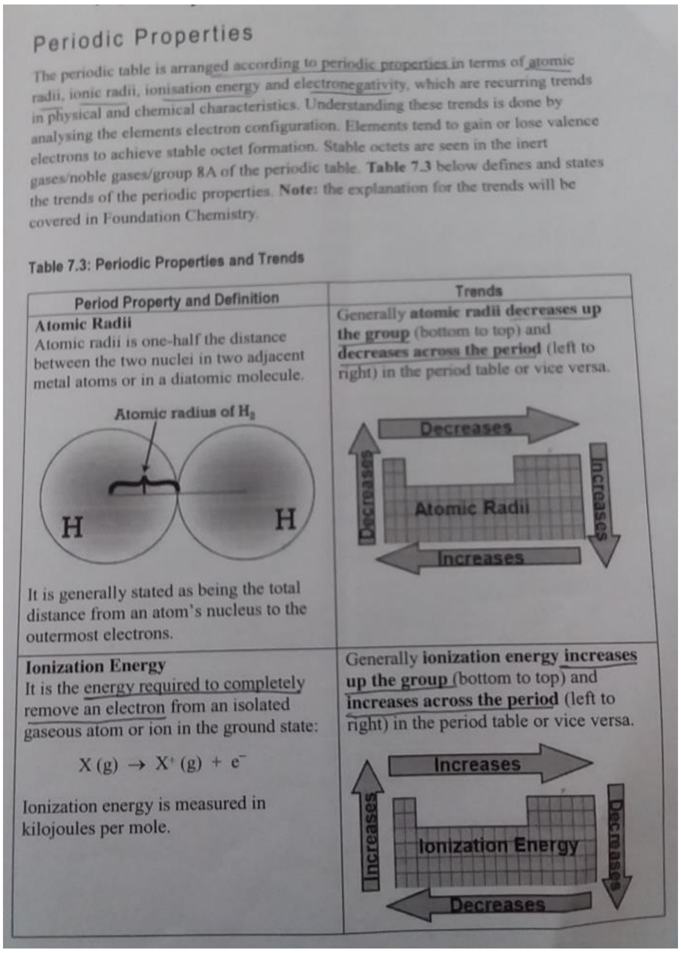
b. i. HCl ii. CO_2 iii. C_2H_6

c. i. CH_4 ii. O_2 iii. H_2O_2

d. i. CCl_4 ii. H_2S iii. CH_3Cl

2. Draw Lewis structures for.

PH_3, CH_2Cl_2, H_2CO, F_2O, $COCl_2$ (C is the central atom).

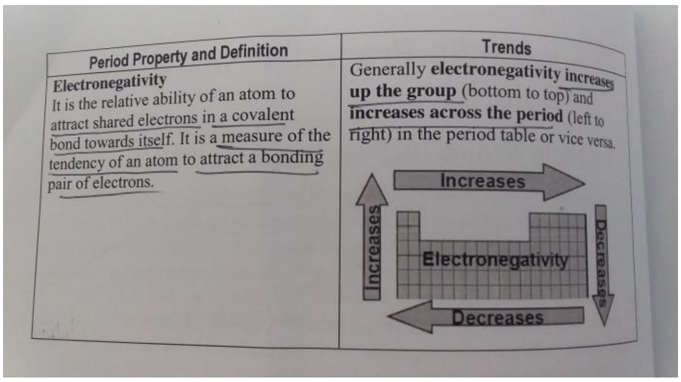


Periodic Properties

The periodic table is arranged according to periodic properties in terms of atomic radii, ionic radii, ionisation energy and electronegativity, which are recurring trends in physical and chemical characteristics. Understanding these trends is done by analysing the elements electron configuration. Elements tend to gain or lose valence electrons to achieve stable octet formation. Stable octets are seen in the inert gases/noble gases/group 8A of the periodic table. **Table 7.3** below defines and states the trends of the periodic properties. **Note:** the explanation for the trends will be covered in Foundation Chemistry.

Table 7.3: Periodic Properties and Trends

Period Property and Definition	Trends
Atomic Radii Atomic radii is one-half the distance between the two nuclei in two adjacent metal atoms or in a diatomic molecule. It is generally stated as being the total distance from an atom's nucleus to the outermost electrons.	Generally **atomic radii decreases up the group** (bottom to top) and **decreases across the period** (left to right) in the period table or vice versa.
Ionization Energy It is the energy required to completely remove an electron from an isolated gaseous atom or ion in the ground state: $X(g) \rightarrow X^+(g) + e^-$. Ionization energy is measured in kilojoules per mole.	Generally **ionization energy increases up the group** (bottom to top) and **increases across the period** (left to right) in the period table or vice versa.



Period Property and Definition	Trends
Electronegativity It is the relative ability of an atom to attract shared electrons in a covalent bond towards itself. It is a measure of the tendency of an atom to attract a bonding pair of electrons.	Generally **electronegativity increases up the group** (bottom to top) and **increases across the period** (left to right) in the period table or vice versa.