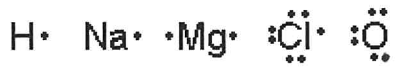
## CHEMISTRY Electron Dot Diagrams

Electron dot diagrams are a simple and useful way of showing the electron arrangements about atoms and the changes that occur as bonds are formed. Electron dot diagrams for an atom use a dot or some other marker for each valence electron. For example, the electron dot diagrams for hydrogen, sodium, magnesium, chlorine and oxygen atoms are as follows.



Electron dot diagrams can be used to represent the formation of sodium ions and chloride ions from the neutral atoms.

*sodium atom* *chlorine atom* *sodium ion* *chloride ion*

Although the sodium and chloride electrons are identical they are shown differently in the diagram so that the transfer can be followed easily. Remember, also, that dots are only a means of counting electrons, they do not show the location of electrons.

The group II element magnesium also forms an ionic compound with chlorine. In this instance the metal achieves a noble gas electron configuration by the loss of two electrons.

*magnesium atom* *2 chlorine atom* *magnesium ion* *2 chloride ions*

Similarly, magnesium forms an ionic compound with oxygen.

*magnesium atom* *oxygen atom* *magnesium ion* *oxide ion*

The most typical ionic compounds are formed when a metal element from group I or group II of the periodic table combines with a non-metal element form group VI or VII. When the reaction occurs, electrons are **transferred** from the metal to the non-metal until the valence electron configurations of the resulting ions are identical to those of the nearest noble gas. The metal elements tend to lose electrons readily, since they have low ionisation energies and form positive ions while the non-metal elements with high ionisation energies tend to accept electrons and form negative ions.

Electron dot diagrams can also be used to illustrate **covalent bond** formation. For hydrogen this can be represented as follows.

$$H\bullet \quad + \quad H^{\times} \longrightarrow H\overset{\bullet}{\underset{\times}{}}H$$

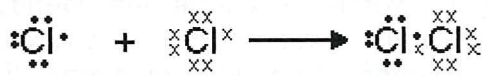
*hydrogen atom* *hydrogen atom* *hydrogen molecule*

Or simply,

$$H\bullet \quad + \quad H^{\times} \longrightarrow H\overset{\bullet}{\underset{\times}{}}H$$

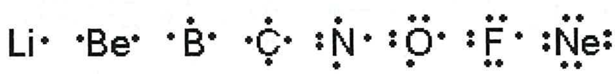
A covalent bond in which two electrons are **shared** is called a single covalent bond and can be represented by a line drawn between the atoms. Hence a hydrogen molecule is represented as H—H.

When two chlorine atoms combine to form a chlorine molecule, the chlorine atoms share a pair of electrons. In this way, each chlorine atom obtains a share in eight valence electrons and acquires the electron configuration of the noble gas argon. As with ionic compounds the octet rule is obeyed in most covalent molecular substances.



*chlorine atom* *chlorine atom* *chlorine molecule*

The number of covalent bonds formed by an atom of an element depends on the number of electrons in its valence or outer energy level. The electron dot diagrams for elements in the second period of the periodic table are shown as follows.

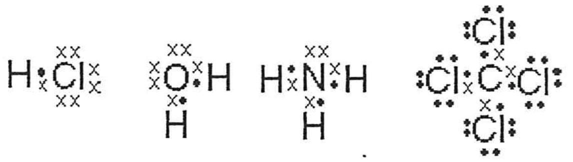


Fluorine will tend to form one single covalent bond in molecules as it will then have a share in eight valence electrons. Oxygen, with six valence electrons, requires two more electrons to achieve a complete octet. Oxygen will therefore tend to form two covalent bonds when it forms covalent molecular substances. Similarly nitrogen and carbon will tend to form three and four covalent bonds respectively.

Elements in the same group of the periodic table have the same valence electron configuration. Consequently, they will tend to form the same numbers of covalent bonds. For example, the group VII elements tend to form one covalent bond in molecular substances. Carbon and silicon, in group VI, form four covalent bonds.

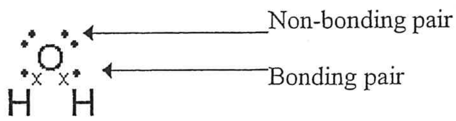
Below are the electron dot diagrams for some simple molecules. Notice that in all cases each atom has a share in a noble gas valence electron configuration.

For most atoms this is eight valence electrons but for hydrogen it is only two valence electrons. Therefore each of the atoms has a stable valence electron configuration.



*Hydrogen chloride* *water* *ammonia* *carbon tetrachloride*

The electron pairs forming the covalent bonds in molecules are called bonding electron pairs. The remaining electron pairs, if any, are called non-bonding electron pairs or lone pairs. In the water molecule, for example, there are two bonding electron pairs and two non-bonding pairs.




*(From: PJ Garnett (Ed.), 1989, <u>Foundations of Chemistry</u>, Longman Cheshire, Melbourne, pg 95, 102)*

# ELECTRON-DOT STRUCTURES ASSIGNMENT

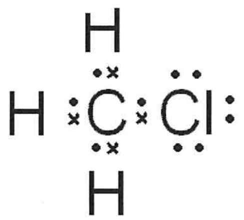
► Draw electron dot structures for the following Elements, Molecules, Ions or Ionic Compounds:

1. $CH_3Cl$
2. $Br_2$
3. HF
4. $Cl_2O$
5. $NF_3$
6. $PCl_3$
7. $SF_2$
8. Calcium Ion
9. Lithium Ion
10. Bromide Ion
11. Potassium Chloride
12. Aluminium
13. Oxygen
14. Nitrogen
15. Magnesium Iodide
16. Sodium Sulphide
17. Argon
18. Rubidium
19. Caesium Fluoride
20. Ammonia ($NH_3$)

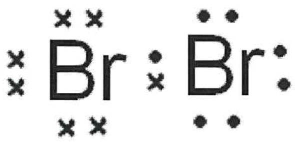
# ELECTRON-DOT STRUCTURES ASSIGNMENT: **SOLUTIONS**

► Draw electron dot structures for the following Elements, Molecules, Ions or Ionic Compounds:

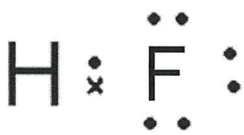
1. $CH_3Cl$ ➡



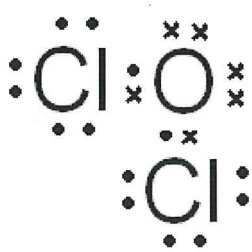
2. $Br_2$ ➡



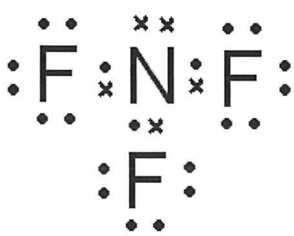
3. HF ➡



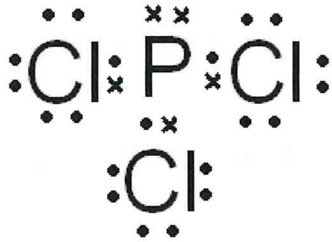
4. $Cl_2O$ ➡



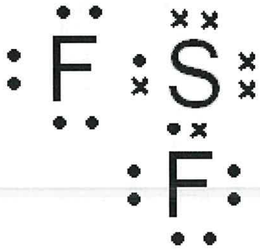
5. $NF_3$ ➡



6. $PCl_3$ ➡



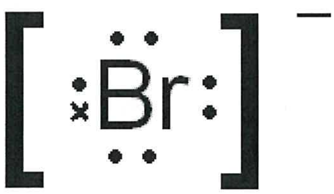
7. $SF_2$ ➡



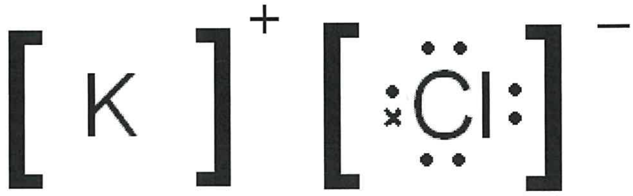
8. Calcium Ion ➡ $[Ca]^{2+}$

9. Lithium Ion ➡ $[Li]^{+}$

10. Bromide Ion ➡

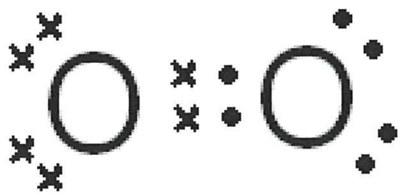


11. Potassium Chloride ⇨

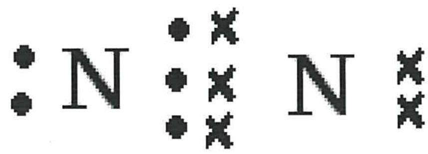


12. Aluminium ⇨

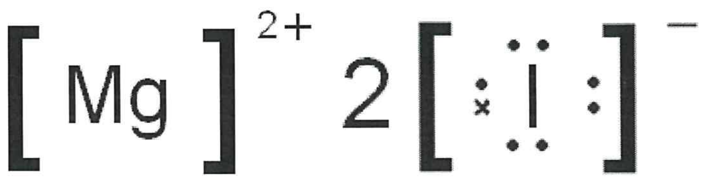
13. Oxygen ⇨

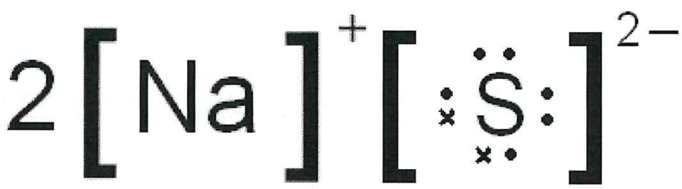


14. Nitrogen ⇨



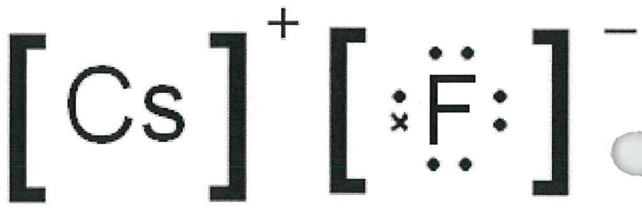
15. Magnesium Iodide ⇨



16. Sodium Sulphide ➡ 

17. Argon ➡ 

18. Rubidium ➡ 

19. Caesium Fluoride ➡ 

20. Ammonia ($NH_3$) ➡