

Worksheet 13 - Chemical Bonding

The concept of **electron configurations** allowed chemists to explain **why** chemical molecules are formed from the elements. In 1916 the American chemist Gilbert Lewis proposed that atoms can achieve a **noble gas electronic configuration** by gaining, losing or sharing electrons with other atoms. Since the noble gases (except He) have **8 valence electrons** his proposal is known as the **octet rule**.

The rule states that, except for hydrogen, an atom combines with other atoms to form bonds in order to have 8 electrons in its valence shell. Hydrogen shares electrons with other atoms to achieve the He electronic configuration.

Lewis dot symbols are representations of the elements which give a dot (·) for each **valence** electron on the atom.

The Lewis dot symbols for the period 2 elements are:



Notice that Hund's rule is followed and that the electrons are left unpaired if possible. Transition metals (and the lanthanides and actinides) can't be represented by simple Lewis Dot symbols.

1. Fill in the Lewis dot symbols for:

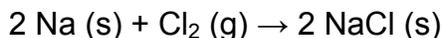


Lewis dot structures allow us to understand two types of bonding, **ionic** and **covalent**.

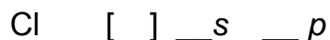
Ionic Bonds

Ionic bonds are usually formed by the reaction of **metals** with **non-metals**.

Sodium reacts explosively with chlorine gas to form sodium chloride:



2. Fill in the Lewis dot symbols for Na and Cl, below, and complete the shorthand electron configuration for each:



3. Now allow each atom to complete its octet, by **losing** or **gaining** electrons. Think about the **Periodic Trends** when deciding which atom will lose electrons and which will gain electrons, and how many electrons will be exchanged. Then write the shorthand electron configuration for each **ion** and its **charge**.

Na

Cl

Sodium is oxidized (loses an e^-) because metals have relatively low **ionization energies**. Chlorine is reduced (gains an e^-) because the non-metals have relatively high **electron affinities**.

These **ions** are then held together in a crystalline structure by the interaction of their charges (lattice energy). **No chemical bonds are formed** between the anions and cations.

4. Use Lewis dot symbols to show the transfer of electrons between the following atoms to form cations and anions:
- a) K and S
 - b) O and Ba
5. Balance the following reactions and write the Lewis dot symbols of the reactants and products.
- a) $\text{Sr} + \text{Se} \rightarrow$
 - b) $\text{Al} + \text{S} \rightarrow$
 - c) $\text{Mg} + \text{F} \rightarrow$
6. Name an **ionic** compound that contains only **non-metallic elements**. (Hint: Think of the polyatomic ions)

Covalent Bonds

These are bonds formed between **non-metals**. The non-metals all have relatively high ionization energies so octets of valence electrons are obtained by **sharing** electrons, instead of gaining or losing electrons. An example of covalent bonding is F_2 , fluorine gas.

7. Complete the Lewis dot symbols for the fluorine atoms below



8. Determine the number of **valence electrons** available in the F_2 molecule.

total valence electrons = _____

Use two of these electrons to make a **bond** between the 2 fluorine atoms. A chemical bond is a **pair of shared electrons**.



9. Determine the number of **unused valence electrons**. _____

10. Place the remaining electrons as **pairs of electrons** around the F_2 structure.



Does each of the F atoms have an **octet** of electrons?

The **shared electrons** form the **chemical bond** between the F atoms. The other 3 pairs of electrons on each of the F atoms are called **lone pair electrons** and are not involved in bonding.

We represent F_2 as **F-F**. Each of the F atoms will have 3 lone pairs of electrons and an equal shared in the bonding electrons.

11. Draw the Lewis structure of water, H_2O .
- Total # of valence electrons available = _____
 - Using **pairs** of electrons, bond the central O to the 2 H atoms.
 - Complete the octet on O using the remaining valence electrons.

The H - O bond is formed by a shared pair of electrons. However, the electrons will **not** be shared equally between the H and the O. This is due to another periodic property, **electronegativity**. This is a concept introduced by Linus Pauling who assigned a value to each element, describing its ability to attract shared electrons to itself. Electronegativity increases going from metals to non-metals, and going up the periods. Fluorine is the most electronegative element.

		H 2.1														
IA	IIA											IIIA	IVA	VA	VIA	VIIA
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	IIIB	IVB	VB	VIB	VII B	VIII B			IB	IIB	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2

Oxygen (3.5) is more electronegative than hydrogen (2.1) so the shared electron pair will be more strongly attracted to the oxygen which will result in a **partial negative charge** (δ^-) on the oxygen and a **partial positive charge** (δ^+) on each of the hydrogens.

The bigger the **difference in electronegativity**, the larger the partial charges, the **more polar** the bond.

12. Using **only** the Periodic Table, arrange the following elements in order of **increasing** electronegativity.

- Rb, Na, Li
- P, S, Al
- Fe, As, Br

13. Label "i" (ionic) or "c" (covalent) the bonds in the following compounds.



14. Rank the following bonds in order of **increasing** ionic character.

a) C-Cl NaCl Cl-Cl

b) N-F O-F C-F

c) C-F N-O Si-F

d) H-Cl B-Cl S-Cl

15. Consider the following bonds:

Mg-O C-O O-O Si-O N-O

a) Which would be the most polar, covalent bond?

b) Which would be the least polar (but still polar) covalent bond?

c) Describe the other three bonds.

16. How many **lone pair electrons** exist in the following compounds?
(Hint: Draw the Lewis structures)

a) HCl

b) H₂S

c) CH₄

17. Which of the compounds in problem 16 has the most polar bonds?

Which has the least polar bonds?