PREDICTING THE FORMULA OF COVALENT COMPOUNDS

Predicting the formulae of covalently–bonded binary compounds (compounds made from two different elements) is a very simple process. In fact, the process is almost identical to the process used to predict the formulae of ionic compounds.

Instead of using the charge on an ion to decide the formula, the "combining capacity" is used. The combining capacity is the number of bonds which an atom is expected to form when involved in covalent bonding, and is equal to the ***valence*** of the atom.

Column	1	2	13	14	15	16	17	18
Valence	1	2	3	4	3	2	1	0

EXAMPLES: Predict the formula of the compound formed from N and F.

N has a valence of 3 and F has a valence of 1. "Criss–cross" the valences.

$\overset{3}{N}_1 \times \overset{1}{F}_3$ gives the formula N_1F_3 or simply **NF_3**

Predict the formula of the compound formed from C and O.

C has a valence of 4 and O has a valence of 2. "Criss–cross" the valences.

$\overset{4}{C}_2 \times \overset{2}{O}_4$ gives the formula C_2O_4 or more simply **CO_2**

EXERCISE:

72. Predict the formula of the compound formed by bonding together the following.

(a) P and Cl	(d) P and O	(g) H and O	(j) C and Cl
(b) B and O	(e) H and Se	(h) N and I	(k) Si and P
(c) C and S	(f) F and O	(i) B and C	(l) Si and S

(c) LONDON FORCES

Individual molecules are held together by covalent bonds between the atoms in the molecule. Such bonds are STRONG and are called **INTRAMOLECULAR FORCES** ("intra" means "within"). In addition to the bonds holding atoms together into molecular units, there are weak forces which hold one complete, neutrally–charged molecule next to another such molecule. These **INTERMOLECULAR FORCES** ("inter" means "between") are called **van der Waals forces.**

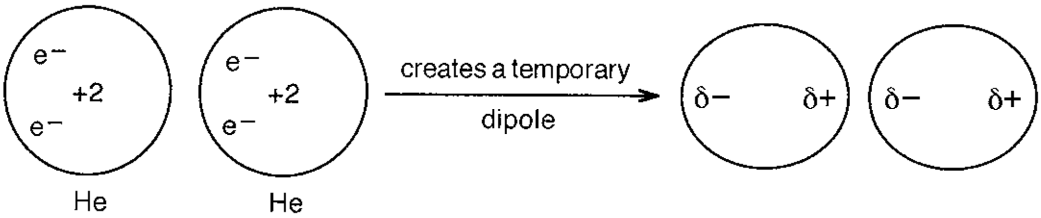
There are two main types of van der Waals forces. This section only deals with a special type of van der Waals force called the **London force** (also called the Heitler–London force or London dispersion force). The topic of van der Waals forces will come up again in the next unit, which examines the role that van der Waals forces play in solubility.

Definitions: A **DIPOLE** is a partial separation of charge which exists when one end of a molecule (or bond) has a slight excess of positive charge and the other end of the molecule (or bond) has a slight excess of negative charge.

LONDON FORCES are weak attractive forces which arise as a result of temporary dipolar attractions between neighbouring atoms. The atoms may exist individually or as parts of molecules.

The Origin of London Forces

The electrons around an atom tend to avoid each other (negative charges repel) and at the same time are attracted to positively charged regions. When two atoms are close to each other, the electrons on one atom repel the electrons on the other atom and also experience an attraction to the nucleus of the other atom (in addition to the attraction to their own nucleus). This process of repelling the electrons on a nearby atom is called "***polarization***" and atoms which can be polarized easily are said to have a "***high polarizability***". The polarization process sets up a very short–lived dipole. In the diagram, two adjacent helium atoms find themselves with a greater electron "density" on one side than the other, creating a slight excess of negative charge, $\delta-$, on one side and a slight excess of positive charge, $\delta+$, on the other side. ("δ" is the Greek symbol for "d" and stands for "slightly" or "a little bit.") These dipoles are extremely short–lived because the overall distribution of the charges of the electrons within the atom is constantly changing; in one instant there may be more negative charge on one side of an atom while in the next instant the distribution of charge may be more to another side. As a result of these temporary dipoles, a weak attractive force exists between the atoms.



electron charge distribution is temporarily more to one side of the atoms than the other

In small atoms such as helium, the electrons are tightly held by the nucleus and are not easily polarized so that helium experiences very weak London forces and melts below $-272^{o}C$. Similarly, the hydrogen molecule (H_2) has only two tightly–held electrons engaged in a covalent bond. The low polarizability of these electrons again gives rise to very weak London forces between adjacent H_2 molecules so that $H_2(s)$ melts at $-259^{o}C$.

LONDON FORCES ARE THE WEAKEST TYPE OF BONDING FORCE KNOWN. The rule governing the strength of London forces is simple and without any exceptions.

The **more** electrons an atom or molecule has altogether, the **stronger** the London forces existing between it and a neighbouring atom or molecule. In other words: **the greater the atomic number of an atom, the stronger the London forces it experiences.**

SPECIAL NOTE: London forces are ALWAYS present, even in species which have covalent or ionic bonding.

EXERCISE:

73. What happens to the strength of the London forces between two identical atoms going
    (a) down a column in the periodic table? (b) left to right across the periodic table?

HOW CAN I TELL WHEN LONDON FORCES ARE IMPORTANT?

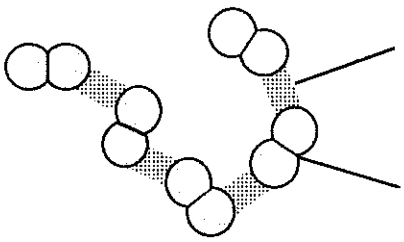
London forces are always present, but are much weaker than covalent and ionic bonds. Hence, London forces are important when they are the ***ONLY*** force of attraction existing between two species. That is, London forces are important between the following closed–shell species:

a) adjacent noble gas atoms, and
b) adjacent covalently–bonded molecules (made up of atoms having a full shell after bonding).

EXAMPLE: London forces hold the following species together in the solid or liquid phase.

Kr(s), H_2(l), H_2(s), Cl_2(s), O_2(l), He(l), CH_4(l)

Notice that in each case the species are either noble gases (He & Kr), or molecules made up of covalently–bonded atoms having filled shells (H_2, Cl_2, O_2 & CH_4). In the case of liquid hydrogen, the bonding present resembles the figure below.



EXERCISES:

74. Are London forces INTERMOLECULAR or INTRAMOLECULAR forces?

75. The F atoms in a single molecule of F_2 are held together by covalent bonding. A sample of F_2(l) consists of F_2 molecules held next to each other by London forces. When a sample of F_2(l) boils and becomes F_2(g), are the covalent bonds between F atoms or the London forces between F_2 molecules broken? Why does this occur?

76. What should happen to the melting and boiling temperatures of atoms held together by London forces when the atomic number of the atoms increases?

77. The following substances rely on London forces to hold them in the liquid phase. Which substance in each pair should have a higher boiling temperature?
(a) Ne or Ar (b) Br_2 or Cl_2 (c) CH_4 or CF_4 (d) CBr_4 or CCl_4

78. In response to a request to "write an equation showing what happens when H_2(s) melts", a student writes the following:

$$H_2(s) \longrightarrow 2\,H(l).$$

What does this equation incorrectly imply about the bonds and forces in a sample of H_2(s)?