Introduction

The concept at the heart of chemistry is that of the **chemical bond**. The chemical reactions that we've studied before involve the breaking and formation of these **bonds**. If we can describe and predict how these bonds form and what causes them to break, then we can predict and describe how chemical reactions work and use this to our advantage. We could create advanced new drugs. We could create polymers (plastics) with enhanced strength. In short, the possibilities are limitless.

Over the years, we've learned a lot about chemical bonding and **have** done some of the things I've listed above. We still, of course, have more to learn, but this note pack will introduce you to the basic terminology and concepts of chemical bonding.

The bond - a definition

To talk about chemical bonds, we need a working definition of a bond. A **chemical bond** is a strong attractive force existing between atoms in a substance. There are three major types of chemical bond:

Type	***Description***	***Example***
ionic bond	held together by attractive forces between oppositely charged ions	NaCl: held together by attraction between Na^+ and Cl^-
covalent bond	held together by the sharing of valence electrons between two atoms	H_2O: held together by sharing of electrons between hydrogen and oxygen.
metallic bond	held together by the sharing of valence electrons with all atoms in the metal's structure - make the metal conduct electricity	Any metal.

We will discuss the **ionic bond** and the **covalent bond** in more detail. You don't need to know anything about the metallic bond except for what is in the table above.

Ionic bonds

We've already talked about **ionic compounds** and **molecular compounds** when we discussed chemical nomenclature, or the naming of compounds. **Ionic compounds** are held together by **ionic bonds**.

Ionic bonds form when **electrons are transferred** from one atom to another. This transfer causes **ions** (charged species) to form. Let's look at an example reaction - the reaction between a neutral atom of lithium and a neutral atom of fluorine:

$$Li + F \rightarrow LiF$$

If you were naming this compound, you'd see it's made of a metal combined with a nonmetal and is likely **ionic**. You'd name it **lithium fluoride**. Lithium fluoride contains lithium (Li^+) ions and fluoride (F^-) ions. How and why do these ions form?

Let's look at the electron configuration of each element and ion. I've indicated the valence electrons in bold.

Li: $1s^2\mathbf{2s^1} \rightarrow Li^+$: $1s^2$ (one electron has been **lost**)
F: $1s^2\mathbf{2s^2 2p^5} \rightarrow F^-$: $1s^2\mathbf{2s^2 2p^6}$ (one electron has been **gained**)

Each atom now has the electron configuration of one of the **noble gases**. Lithium's electron configuration is just like helium's. Fluorine's electron configuration is just like neon's.

Remember that the noble gases are stable and unreactive. They're stable and unreactive because they have **filled valence shells**. So, **atoms** (when forming ions) will generally lose or gain enough electrons to form ions with noble gas electron configurations - **filled valence shells**.

Look at the main-group **metals** on the periodic table. You can see that most of them have few electrons in their outer shells (especially groups IA and IIA). The easiest way for them to get filled valence shells is to **lose electrons**, so elements like lithium and magnesium easily lose electrons to form **cations** (positively charged ions).

Look at the main-group **nonmetals**. You can see that most of them only need a few electrons to get filled valence shells (especially groups VIA and VIIA). The easiest way for them to get filled valence shells is to **gain electrons**, so elements like oxygen and fluorine easily gain electrons to form **anions** (negatively charged ions).

If you put something that easily loses electrons together with something that easily gains electrons, you make the conditions right for **electron transfer**. This is why when you react a **metal with a nonmetal**, you normally form an **ionic compound**!

[We generally ignore the d and f subshells when we talk about valence shells. The d and f subshells are high-energy, so they do not get filled until the next shell starts to fill. For example, the 3d shell doesn't fill until after 4s does. That said, the d subshells are very important for the chemistry of some elements – particularly the transition metals.]

<u>A graphical representation of bonding: Dot structures</u>

We can represent both ionic and covalent bonds with a graphical notation known as the **Lewis dot structure**. The Lewis dot structure (abbreviated from now on as the **Lewis structure**) shows certain of the valence electrons as **dots** around each atom involved in the reaction. The way the dots are drawn around the structure is important - it indicates how many bonds the compound is likely to form!

So how do we draw these? Let's look at the simplest type of Lewis structure first - that of a single-element species. These can be lone atoms of an element or they can be **monatomic ions**. To draw the Lewis structure for these, first count up all the **s and p** valence electrons, giving you a number ranging from zero to eight. Then, draw a dot around the symbol of the species for each valence electron, as illustrated below.

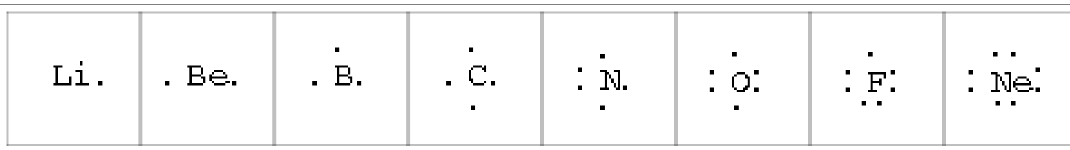


Illustration 1 - The Lewis structures of Period 2 atoms

Compare the dot structures above to the electron configurations below.

Element	***Electron Configuration***	***Number of s and p valence electrons***
Lithium	Li: $1s^2\mathbf{2s^1}$	1
Beryllium	Be: $1s^2\mathbf{2s^2}$	2
Boron	B: $1s^2\mathbf{2s^23p^1}$	3
Carbon	C: $1s^2\mathbf{2s^22p^2}$	4
Nitrogen	N: $1s^2\mathbf{2s^22p^3}$	5
Oxygen	O: $1s^2\mathbf{2s^22p^4}$	6
Fluorine	F: $1s^2\mathbf{2s^22p^5}$	7
Neon	Ne: $1s^2\mathbf{2s^22p^6}$	8

The order you put the dots around the atom isn't important. For example, drawing the dot on the left-hand side of lithium instead of the right-hand side doesn't make the structure any different. The pairing of electrons, though, is important. Draw the first four electrons around the atom as single dots. Start paring when you add the fifth electron.

Electrons represented as **single dots** tend to be involved in bonding (both ionic and covalent). This is **not** a hard-and-fast rule, but you can make a good guess as to how an element might bond by looking at the Lewis formula.

What about ions? Ions are formed by gaining or losing valence electrons. This is reflected in the Lewis structure. Looking back to our example of lithium reacting with fluorine, let's remember that lithium loses a valence electron while fluorine gains one. This is reflected in the Lewis structures for each element.

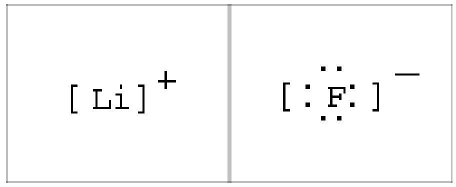


Illustration 2 - Lewis structures of lithium ion and fluoride ion

Lithium started off with only one valence electron, and it lost that electron when forming the ion. Lithium ion has **no** dots in its Lewis structure. Fluorine gains an electron to form fluoride ion by filling up its 2p subshell, so its structure is drawn with **eight** dots.

So what does the Lewis structure of these ions show us? It shows us graphically how the reaction proceeded - a transfer of an electron from one atom to another. *[What kind of chemical reaction is this, using the classification schemes we learned earlier?]*

Illustration 3 - Graphical representation of the reaction of elemental lithium with elemental fluorine

[Why only "s" and "p" electrons? The "d" electrons of elements do not typically get involved in bonding, since they are usually not on the outside of the atom. This isn't true for many transition metal compounds and some molecules, but we won't concern ourselves much with these now. A few examples of the chemistry of d orbitals will be covered later.]

<u>Covalent bonds: The formation of molecular compounds</u>

Molecular compounds are held together by **covalent bonds**. These bonds don't involve a transfer of electrons between atoms. One atom does not lose an electron to another to form ions when a covalent bond forms. Instead, the atoms share electrons - effectively letting the shared electrons be in the subshells of both the sharing elements at the same time. To do this, the elements come together close enough for the orbitals on the elements to overlap each other - this overlap is a chemical bond. We call a collection of atoms held together by this overlapping a **molecule**.

Why do atoms share electrons? Basically, for the same reason other atoms transfer electrons - they want **filled valence shells**. Most of the atoms that form covalent bonds end up with a noble gas configuration, which has filled "s" and "p" subshells. The observation that atoms bond to fill their "s" and "p" subshells is called the **octet rule** - because that's the total number of electrons in a filled "s" and "p" orbital (2+6=8). The octet rule is a useful **rule of thumb** in determining how atoms in a molecular compound bond together!

Atoms share electrons in pairs, and two atoms may share one, two, or three pairs. We call the number of pairs shared the **bond order.**

If we share [...] pair(s)...	*... the bond is called a ...*	*... and the bond order is ...*
one	single bond	1
two	double bond	2
three	triple bond	3

So, when we speak of a **"double bond"** or a **"second-order bond"**, then we mean that the atoms in question are sharing **two pairs** of electrons.

Let's look at a simple example - the bonding in oxygen.

Looking back in the first illustration, recall that the Lewis structure of oxygen has **two unpaired electrons**. If one oxygen atom encounters another oxygen atom, they react together by sharing these unpaired electrons. A total of **four electrons** are shared between the oxygen atoms, and a **double bond** forms.

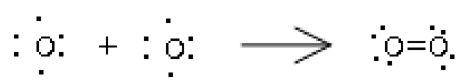


Illustration 4 - Atomic oxygen combines to form an oxygen molecule

Things to note:

- The lines between the oxygen atoms each represent two electrons. It's common to abbreviate bonds by drawing a line between the atoms for each pair of electrons shared.
- Each oxygen now sees eight electrons (four belonging to the oxygen atom and four involved in the bond), so the oxygen atoms each have a **noble gas configuration**.

[How many bonds would you expect a ***nitrogen*** *atom to form if it reacted with another nitrogen atom? Three.]*

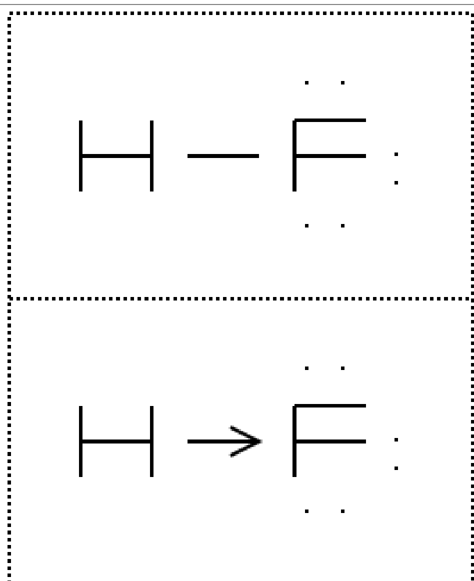
<u>Electron sharing: Some animals are more equal than others</u>

We've discussed so far only covalent bonds between identical atoms. In these bonds, it's obvious that electrons are shared equally between the two bonding atoms. In oxygen, for example, the two oxygen atoms are identical. There's no way for one of the oxygen atoms to hoard the shared electrons.

But what if the atoms **aren't** identical? The electrons might not be evenly shared. The extreme case is the **ionic bond**, where electrons are completely moved from one atom to another!

We call this concept of uneven sharing **polarity** (like the poles of a magnet, or the poles of a battery). We say that a bond is a **polar bond** if the electrons in the bond spend more time near one atom than they do the other. A bond is **nonpolar** if the electrons are (approximately) equally shared between the electrons.

We've already looked at a nonpolar bond in this note pack - the double bond between the oxygen atoms in the oxygen atom. Let's look at an example of a polar bond. Consider the bond in a chemical compound like hydrogen fluoride (HF). Both hydrogen and fluorine need to gain one electron in order to have filled valence shells, so they share electrons and form the hydrogen fluoride molecule.

	• The upper picture is the Lewis structure for HF. • The lower picture is another way the bond can be drawn - to represent that electrons are held more closely to fluorine than hydrogen.
Illustration 5 - Bonding in hydrogen fluoride	

How can we rationalize that the fluorine **does** hold the bonding electrons close to itself? From experiment, we can detect the **dipole moment** (a charge separation across the molecule) of the hydrogen fluoride molecule. One end of the molecule has a slight positive charge and the other end has a slight negative charge, though overall the molecule is neutral. The fluorine end of the molecule takes on a slightly negative charge, while the hydrogen end takes on a slight positive charge.

If we don't have experimental data handy, we could remember our discussion of periodic trends - in particular **electron affinity**. Fluorine loses more energy than hydrogen does when it acquires an electron, so it stands to reason that the fluorine would hold the bonding electron more tightly – so fluorine holds its bonded electrons closer.

Electronegativity

The **electronegativity** of an atom describes how closely that atom holds electrons in a covalent bond. A **larger** value of electronegativity indicates that the electron is held **closer** to the atom. You can look up electronegativities of various atoms in books, or you can remember the rough trend that electronegativity increases as you get closer to fluorine on the periodic table (ignoring the noble gases).

You can determine what type of bond you're dealing with by looking at the **electronegativity difference** between the two atoms involved in the bond.

Electronegativity Difference	***Type of Bond***
Little or no difference	nonpolar covalent bond: C-C, C-H, etc.
Large difference	polar covalent bond: H-F, C-F, etc.
Very large difference	ionic bond: NaCl, KCl, etc.

Summary

In this note pack, we've covered the basics of chemical bonding as found in the first few sections of your book. We discussed the three types of bonds: ionic, covalent, and metallic. We discussed ionic and covalent bonds in some detail, and introduced the notion of polarity - uneven electron sharing.

We also discussed how to draw very simple Lewis structures that show us graphically how ionic and molecular compounds are put together.