Pearson

5.6: Bond Energies and Bond Lengths

In the Lewis model, a bond is a shared electron pair; when we draw Lewis structures for molecular compounds, all bonds appear identical. However, from experiments we know that they are not identical—they can vary both in their energy (how strong the bond is) and their length. In this section, we examine the concepts of bond energy and bond length for a number of commonly encountered bonds. In Chapter 9 we will learn how to use these bond energies to calculate energy changes occurring in chemical reactions.

> Bond energy is also called bond enthalpy or bond dissociation energy.

Bond Energy

The **bond energy** of a chemical bond is the energy required to break 1 mole of the bond in the gas phase. For example, the bond energy of the Cl—Cl bond in Cl_2 is 243 kJ/mol. Bond energies are positive because energy must be put into a molecule to break a bond (the process is endothermic, which, as discussed in Chapter E, absorbs heat and carries a positive sign).

$$\mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{Cl}(g) \quad \text{Bond energy} = 243\ \mathrm{kJ}$$

The bond energy of HCl is 431 kJ/mol.

$$\mathrm{HCl}(g) \rightarrow \mathrm{H}(g) + \mathrm{Cl}(g) \quad \text{Bond energy} = 431\ \mathrm{kJ}$$

We say that the HCl bond is *stronger* than the Cl_2 bond because it requires more energy to break it. In general, compounds with stronger bonds tend to be more chemically stable, and therefore less chemically reactive, than compounds with weaker bonds. The triple bond in N_2 has a bond energy of 946 kJ/mol.

$$\mathrm{N_2}(g) \rightarrow \mathrm{N}(g) + \mathrm{N}(g) \quad \text{Bond energy} = 946\ \mathrm{kJ}$$

It is a very strong and stable bond, which explains nitrogen's relative inertness.

The bond energy of a particular bond in a polyatomic molecule is a little more difficult to determine because a particular type of bond can have different bond energies in different molecules. For example, consider the C—H bond. In CH_4, the energy required to break one C—H bond is 438 kJ/mol.

$$\mathrm{H_3C{-}H}(g) \rightarrow \mathrm{H_3C}(g) + \mathrm{H}(g) \quad \text{Bond energy} = 438\ \mathrm{kJ}$$

However, the energy required to break a C—H bond in other molecules varies slightly, as shown here:

$$\begin{array}{lcll} \mathrm{F_3C{-}H}(g) & \rightarrow & \mathrm{F_3C}(g) + \mathrm{H}(g) & \text{Bond energy} = 446\ \mathrm{kJ} \\ \mathrm{Br_3C{-}H}(g) & \rightarrow & \mathrm{Br_3C}(g) + \mathrm{H}(g) & \text{Bond energy} = 402\ \mathrm{kJ} \\ \mathrm{Cl_3C{-}H}(g) & \rightarrow & \mathrm{Cl_3C}(g) + \mathrm{H}(g) & \text{Bond energy} = 401\ \mathrm{kJ} \end{array}$$

We can calculate an *average bond energy* for a chemical bond, which is an average of the bond energies for that bond in a large number of compounds. For the limited number of compounds we just listed, we calculate an average C—H bond energy of 422 kJ/mol.

Table 5.3 lists average bond energies for a number of common chemical bonds averaged over a large number of compounds. Notice that the C—H bond energy listed is 414 kJ/mol, which is not too different from the value we calculated from our limited number of compounds. Notice also that bond energies depend not only on the kind of atoms involved in the bond, but also on the type of bond: single, double, or triple. In general, for a given pair of atoms, triple bonds are stronger than double bonds, which are, in turn, stronger than single bonds. For pair of atoms, triple bonds are stronger than double bonds, which are, in turn, stronger than single bonds. For example in Table 5.3, for carbon–carbon bonds, notice the increase in bond energy in going from a single to a double and then to a triple bond.

Table 5.3 Average Bond Energies

Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)
H—H	436	C—C	347	N≡N	946
H—C	414	C═C	611	O—O	142
H—N	389	C≡C	837	O═O	498
H—O	464	C—O	360	F—F	159
H—F	565	C═O	736*	Cl—Cl	243
H—Cl	431	C—Cl	339	Br—Br	193
H—Br	364	N—N	163	I—I	151
H—I	297	N═N	418		

*799 in CO_2

Bond Length

Just as we can tabulate average bond energies, which represent the average energy of a bond between two particular atoms in a large number of compounds, we can tabulate average bond lengths (Table 5.4). The average **bond length** represents the average length of a bond between two particular atoms in a large number of compounds. Like bond energies, bond lengths depend not only on the kind of atoms involved in the bond, but also on the type of bond: single, double, or triple. In general, for a particular pair of atoms, triple bonds are shorter than double bonds, which are in turn shorter than single bonds. For example, consider the bond lengths (shown here with bond energies, repeated from earlier in this section) of carbon–carbon triple, double, and single bonds:

Bond	Bond Length (pm)	Bond Energy (kJ/mol)
C≡C	120 pm	837 kJ/mol
C═C	134 pm	611 kJ/mol
C—C	154 pm	347 kJ/mol

Table 5.4 Average Bond Lengths

Bond	Bond Length (pm)	Bond	Bond Length (pm)	Bond	Bond Length (pm)
H—H	74	C—C	154	N≡N	110
H—C	110	C═C	134	O—O	145
H—N	100	C≡C	120	O═O	121
H—O	97	C—O	143	F—F	143
H—F	92	C═O	120	Cl—Cl	199
H—Cl	127	C—Cl	178	Br—Br	228
H—Br	141	N—N	145	I—I	266

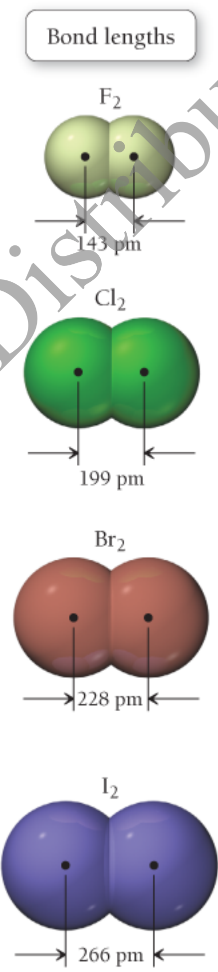
H—I	161	N═N	123	

Notice that, as the bond gets longer, it also becomes weaker. This relationship between the length of a bond and the strength of a bond does not necessarily hold true for all bonds. Consider the following series of nitrogen–halogen single bonds:

Bond	Bond Length (pm)	Bond Energy (kJ/mol)
N—F	139	272
N—Cl	191	200
N—Br	214	243
N—I	222	159

Although the bonds generally get weaker as they get longer, the trend is not a smooth one.

In this chapter, we look at ways to predict and account for the shapes of molecules. The molecules we examine are much smaller than the molecules we discussed in Section 5.1, but the same principles apply to both. The simple model we examine to account for molecular shape is VSEPR theory, and we use it in conjunction with the Lewis model. In Chapter 6 we will explore two additional bonding theories: valence bond theory and molecular orbital theory. These bonding theories are more complex, but also more powerful, than the Lewis model. They predict and account for molecular shape as well as other properties of molecules.



Bond lengths in the diatomic halogen molecules.