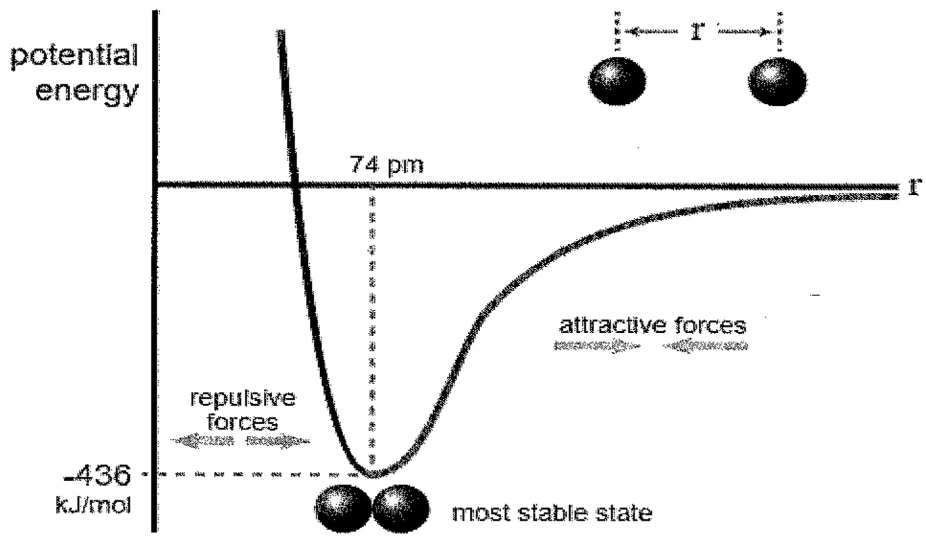
Bond Energy and Enthalpy

A chemical bond forms when a system of bonded atoms is lower in potential energy than that of independent atoms. Since bond-making is an electrostatic process, it is governed by Coulomb's Law.

Coulomb's Law tells us:

- The negative electrons of one atom and the positive nucleus of another atom attract each other.
- If the nuclei of two atoms get too close together, their like charges repel each other.

The optimum distance between two atoms is the bond length, which represents the lowest energy state. The bond length is a balance between the attractive electrostatic forces between the nucleus of one atom and the electrons of another, and the repulsive forces between the positively charged nuclei and the negatively charged electrons.



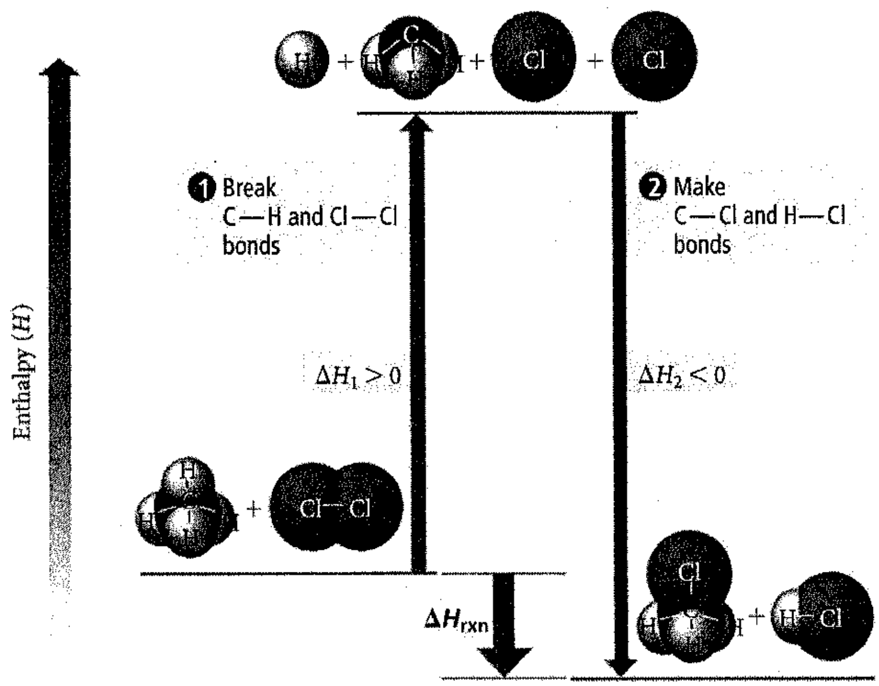
Breaking chemical bonds requires an input of energy to overcome the attractive forces. When new bonds are formed, energy is released. The difference between input and output determines whether a process is endothermic or exothermic.

Covalent Bond Energies and Enthalpy

- The enthalpy change for a reaction, ΔH_{rxn}, can also be understood in terms of bonds breaking (endothermic) and bond forming (exothermic) during a chemical reaction.
- The total enthalpy change can be negative (exothermic) or positive (endothermic) depending on the relative magnitude of two (breaking and making) processes.

<u>**Endo**thermic</u>: $+\Delta H^\circ$ (system gains energy)	<u>**Exo**thermic</u>: $-\Delta H^\circ$ (system loses energy)
➢ Breaking a chemical bond is always endothermic ➢ Hint: "**end**-ing a bond is **end**othermic"	➢ Forming a chemical bond is always exothermic

Estimating the Enthalpy Change of a Reaction from Bond Energies



The following expression can be written to express this relationship:

$$\Delta H^\circ_{rxn} = \Sigma H^\circ_{(bonds\ broken)} - \Sigma H^\circ_{(bonds\ formed)}$$

MEMORIZE! OR **MEMORIZE!**

$$\boxed{\Delta H^\circ_{rxn} = \Sigma \mathrm{BE}_{(reactants)} - \Sigma \mathrm{BE}_{(products)}}$$

Not on F.C.!

IMPORTANT: Draw the Lewis Dot structures of the compounds involved in the reaction to determine the bonds being broken or formed.

But how will I know what bond energies to use?

You will be provided with the bond energies in the problem or be asked to look at a table.

Bond Energy Trends:

- Double and triple bonds are Stronger than single bonds, since double and triple bonds contain a greater number of electrons than single bonds, so Coulombic attractions to the nuclei are stronger!
- Shorter bonds (with the atoms closer to each other) tend to be stronger than longer bonds with the atoms further apart, since the distance between the charges is smaller.

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

Table 8.4 ▶ Average Bond Energies (kJ/mol)

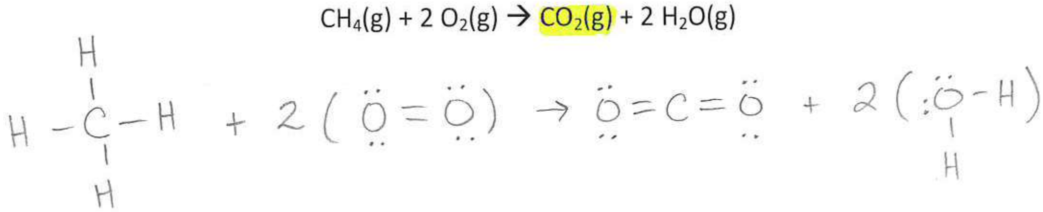
Single Bonds						Multiple Bonds	
H—H	432	N—H	391	I—I	149	C=C	614
H—F	565	N—N	160	I—Cl	208	C≡C	839
H—Cl	427	N—F	272	I—Br	175	O=O	495
H—Br	363	N—Cl	200			C=O*	745
H—I	295	N—Br	243	S—H	347	C≡O	1072
		N—O	201	S—F	327	N=O	607
C—H	413	O—H	467	S—Cl	253	N=N	418
C—C	347	O—O	146	S—Br	218	N≡N	941
C—N	305	O—F	190	S—S	266	C≡N	891
C—O	358	O—Cl	203			C=N	615
C—F	?	O—I	234	Si—Si	340		
C—Cl	339			Si—H	393		
C—Br	276	F—F	154	Si—C	360		
C—I	240	F—Cl	253	Si—O	452		
C—S	259	F—Br	237				
		Cl—Cl	239				
		Cl—Br	218				
		Br—Br	193				

*C=O(CO_2) = 799

Let's Practice!

1. Using the bond energies from the reference chart above, estimate ΔH_{comb} for the following reaction. Does the reaction illustrate an endothermic or exothermic process?

$$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$$



$$\Delta H^\circ_{comb} = \Sigma BE(react) - \Sigma BE(prod)$$
$$= [4(C\text{-}H) + 2(O{=}O)] - [2(C{=}O) + 4(O\text{-}H)]$$
$$= [4(413) + 2(495)] - [2(799) + 4(467)]$$
$$= 2642 - 3466 = \boxed{-824\ \mathrm{kJ/mol_{rxn}}}$$