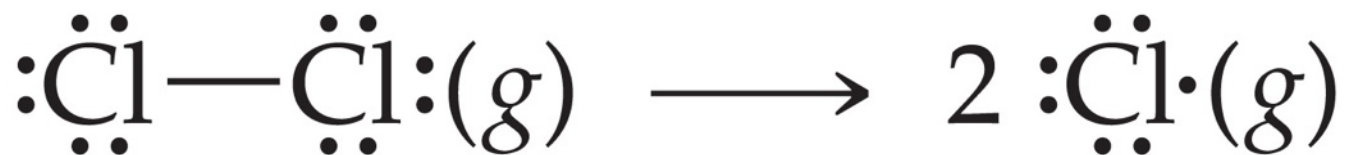


# **Strength of the Covalent Bond**

# Covalent Bond Strength



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- Most simply, the strength of a bond is measured by determining how much energy is required to break the bond.
- This is the **bond enthalpy**.
- The bond enthalpy for a Cl—Cl bond,  $D$  (Cl—Cl), is measured to be 242 kJ/mol.

# Average Bond Enthalpies

- Table 8.4 lists the average bond enthalpies for many different types of bonds.
- Average bond enthalpies are positive (+), because bond breaking is an endothermic process.

TABLE 8.4 • Average Bond Enthalpies (kJ/mol)

## Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						

## Multiple Bonds

C=C	614	N=N	418	O <sub>2</sub>	495
C≡C	839	N≡N	941		
C=N	615	N=O	607	S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

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# Average Bond Enthalpies

*Note:* These are *average* bond enthalpies, not absolute bond enthalpies; the C—H bonds in methane, CH<sub>4</sub>, will be a bit different than the C—H bond in chloroform, CHCl<sub>3</sub>.

TABLE 8.4 • Average Bond Enthalpies (kJ/mol)

## Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
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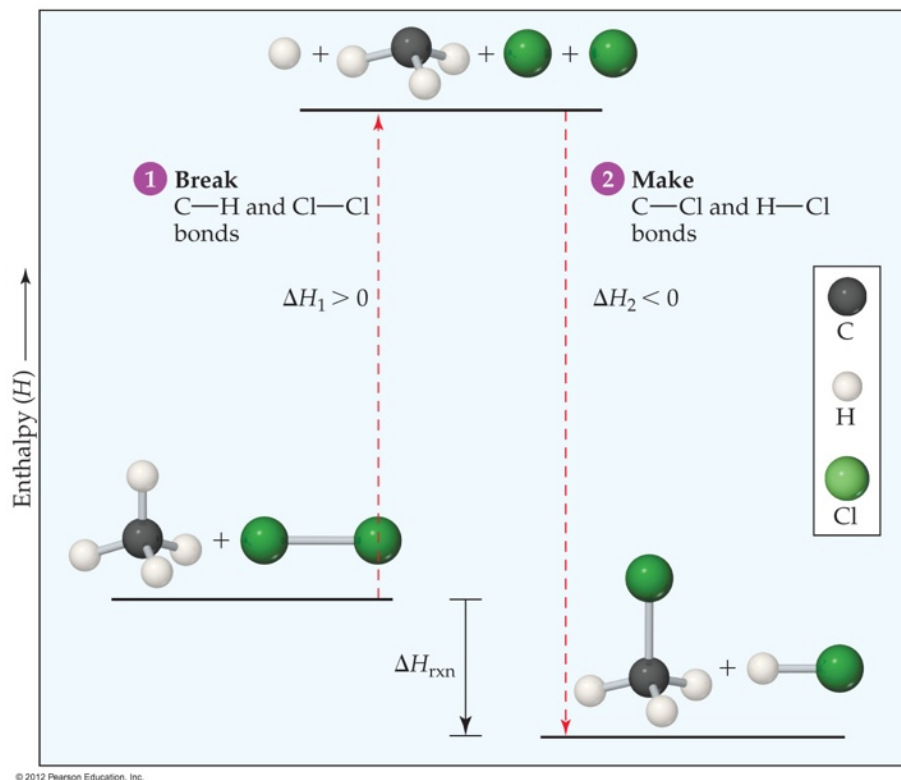
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# Enthalpies of Reaction

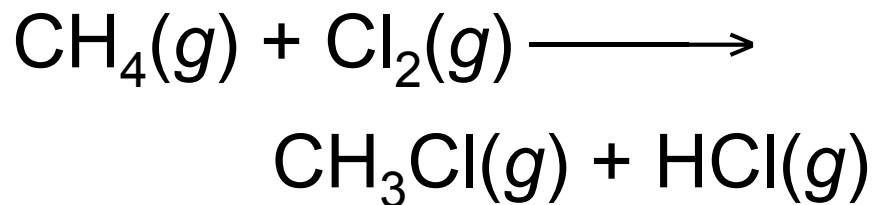
- Yet another way to estimate  $\Delta H$  for a reaction is to compare the bond enthalpies of bonds broken to the bond enthalpies of the new bonds formed.



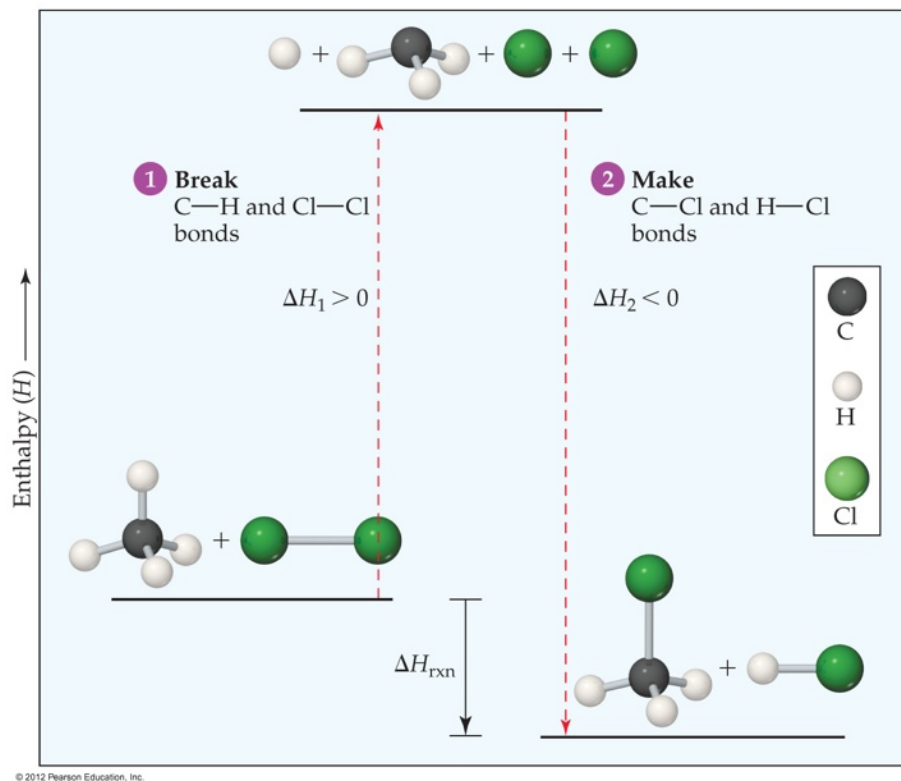
- In other words,

$$\Delta H_{rxn} = \Sigma(\text{bond enthalpies of bonds broken}) - \Sigma(\text{bond enthalpies of bonds formed})$$

# Enthalpies of Reaction



In this example, one C—H bond and one Cl—Cl bond are broken; one C—Cl and one H—Cl bond are formed.



# Enthalpies of Reaction

So,

$$\begin{aligned}\Delta H &= [D(\text{C—H}) + D(\text{Cl—Cl})] - [D(\text{C—Cl}) + \\ &\quad D(\text{H—Cl})] \\ &= [(413 \text{ kJ}) + (242 \text{ kJ})] - [(328 \text{ kJ}) + (431 \text{ kJ})] \\ &= (655 \text{ kJ}) - (759 \text{ kJ}) \\ &= -104 \text{ kJ}\end{aligned}$$

# Bond Enthalpy and Bond Length

**TABLE 8.5 • Average Bond Lengths for Some Single, Double, and Triple Bonds**

Bond	Bond Length (Å)	Bond	Bond Length (Å)
C—C	1.54	N—N	1.47
C=C	1.34	N=N	1.24
C≡C	1.20	N≡N	1.10
C—N	1.43	N—O	1.36
C=N	1.38	N=O	1.22
C≡N	1.16		
		O—O	1.48
C—O	1.43	O=O	1.21
C=O	1.23		
C≡O	1.13		

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- We can also measure an average bond length for different bond types.
- As the number of bonds between two atoms increases, the bond length decreases.



# Example: Using BDEs

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break bonds

make bonds



349

495

380

431

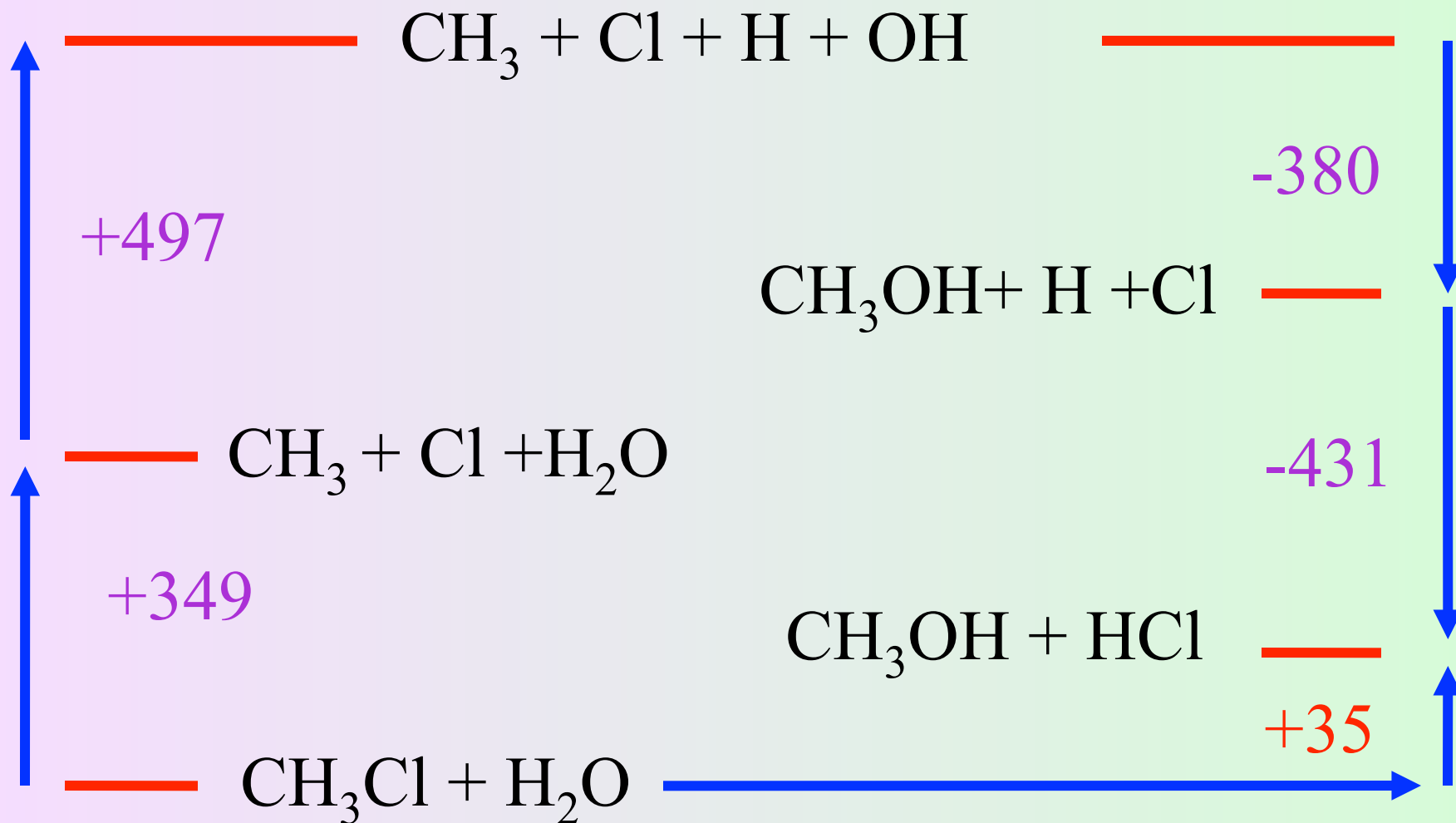
846

811

$$\Delta H^\circ = \Sigma(\text{BDEs broken}) - \Sigma(\text{BDEs formed})$$

$$\Delta H^\circ = 846 - 811 = +35\text{kJ}$$

# Hess's Law, BDEs, and $\Delta H^\circ$





# Bond Dissociation Energy

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The energy required to break a bond



Endothermic for stable compounds :

$\Delta H^\circ$  is always positive ( + )

# Use of Bond Energies in

## Thermochemistry (The heat energy released or absorbed)

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Bond-dissociation energies can be used to calculate  $\Delta H^\circ_{\text{reaction}}$

