Bond Energies

$$\Delta H = \sum \text{Bond Energies}_{\text{broken}} - \sum \text{Bond Energies}_{\text{formed}}$$

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	485	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

Table 8.5 | Bond Lengths and Bond Energies for Selected Bonds

Bond	Bond Type	Bond Length (pm)	Bond Energy (kJ/mol)
C—C	Single	154	347
C=C	Double	134	614
C≡C	Triple	120	839
C—O	Single	143	358
C=O	Double	123	745
C—N	Single	143	305
C=N	Double	138	615
C≡N	Triple	116	891

1. Calculate ΔH for each reaction below.

a) $H_2(g) + F_2(g) \rightarrow 2\ HF(g)$

ΔH = (432 kJ/mol + 154 kJ/mol) − (2 × 565 kJ/mol)

ΔH = (586 kJ/mol) − (1130 kJ/mol) = **−544 kJ/mol**

b) $H_2 + Cl_2 \rightarrow 2\ HCl$

ΔH = (432 kJ/mol + 239 kJ/mol) − (2 × 427 kJ/mol)

ΔH = (671 kJ/mol) − (854 kJ/mol) = **−183 kJ/mol**

c) $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

ΔH = (941 kJ/mol + 3(432 kJ/mol)) − (2(3 × 391 kJ/mol))

ΔH = (2237 kJ/mol) − (2346 kJ/mol) = **−109 kJ/mol**

Heat of Formation

$$\Delta H = \sum \Delta H_f \text{ (products)} - \sum \Delta H_f \text{ (reactants)}$$

2. Find the ΔH_{rxn} for the following reactions using standard enthalpy of formation values. (Appendix 4)

a) $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(l)$

$\Delta H = [(4)(90\ kJ/mol) + (6)(-286\ kJ/mol)] - [(4)(-46\ kJ/mol) + (5)(0\ kJ/mol)]$

$\Delta H = [-1356\ kJ/mol] - [-184\ kJ/mol] = \boxed{-1172\ kJ/mol}$

b) $2\ Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2\ Fe(s)$

$\Delta H = [(1)(-1676\ kJ/mol) + (2)(0\ kJ/mol)] - [(2)(0\ kJ/mol) + (1)(-826\ kJ/mol)]$

$\Delta H = [-1676\ kJ/mol] - [-826\ kJ/mol] = \boxed{-850\ kJ/mol}$

c) $2\ CH_3OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(l)$

$\Delta H = [(2)(-393.5\ kJ/mol) + (4)(-286\ kJ/mol)] - [(2)(-239\ kJ/mol) + 3(0\ kJ/mol$

$\Delta H = [-787\ kJ/mol + (-1144\ kJ/mol)] - [-478\ kJ/mol] = \boxed{-1453\ kJ/mol}$

Thermochemistry Stoichiometry

3. Using the balanced equations above, answer the following questions.

a) How many grams of NO(g) are made for the above reaction letter (a) at 25°C when 200. kJ of heat energy is released?

$$\frac{200.\ kJ}{} \left| \frac{4\ mol\ NO}{1172\ kJ} \right| \frac{30.01\ g\ NO}{1\ mol\ NO} = \boxed{20.5\ g\ NO}$$

b) What mass of Fe_2O_3 is needed to produce 2560 kJ of heat energy for the reaction in letter (b) above?

$$\frac{2560\ kJ}{} \left| \frac{1\ mol\ Fe_2O_3}{850\ kJ} \right| \frac{159.7\ g\ Fe_2O_3}{1\ mol\ Fe_2O_3} = \boxed{481\ g\ Fe_2O_3}$$

c) How many kJ of energy are released when 200.0 g of , methanol, CH_3OH are combusted in reaction letter (c)?

$$\frac{200.0\ g\ CH_3OH}{} \left| \frac{1\ mol\ CH_3OH}{32.05\ g\ CH_3OH} \right| \frac{1453\ kJ}{2\ mol\ CH_3OH} = \boxed{4534\ kJ}$$