BOND ENERGIES (5.8)

exo: $E_{out} > E_{in}$

endo: $E_{in} > E_{out}$

"Steps" in a chemical rxn:

a) breaking existing bonds (req. energy) → E_{in}

b) forming new bonds (releases energy) → E_{out}

BOND ENERGY (D)

- the average energy needed to break a particular bond.
- also equals the E released when that same bond is formed.

eg/ C–C bond energy is 347 kJ/mol

→ 347 kJ req'd to break 1 mol of C–C bonds

→ 347 kJ released when 1 mol C–C bonds is formed

* see Table 1 (p. 307)

BOND ENERGIES CAN BE USED TO APPROXIMATE ENTHALPY CHANGES (ΔH)

$$\Delta H = E_{in} - E_{out}$$

$$\Delta H = \underbrace{\sum n \times D}_{\text{Bonds BROKEN}} - \underbrace{\sum n \times D}_{\text{Bonds FORMED}}$$

Example 1.

Use bond energies to find ΔH for the rxn:

$$2H_2 + O_2 \rightarrow 2H_2O$$

H-H　　　O=O　　　H-O-H

H-H　　　　　　　　H-O-H

① Bonds Broken

2 mol H-H　　$E_{in} = 2\,mol \times D_{H-H}$

$= 2\,mol\,(432\,\frac{kJ}{mol})$

$E_{in} = 864\,kJ$

1 mol O=O　　$E_{in} = 1\,mol\,(495\,\frac{kJ}{mol})$

$E_{in} = 495\,kJ$

Total E_{in}
$= 864 + 495$
$= 1359\,kJ$

② Bonds Formed

4 mol O-H　　$E_{out} = (4\,mol)(467\,\frac{kJ}{mol})$

$E_{out} = 1868\,kJ$

Total E_{out}
$= 1868\,kJ$

③ $\Delta H = E_{in} - E_{out}$

$= 1359 - 1868$

$\Delta H = -509\,kJ$ → ⊖ value ∴ Exothermic. E is released

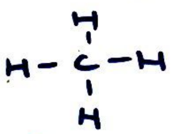
$$2H_2 + O_2 \rightarrow 2H_2O + 509\,kJ$$

Problem Solving Strategy: steps

1. write the balanced equation for the rxn
2. Draw Lewis structures for each substance
3. ID the bonds present, and the # of each
4. Look up values for D
5. Calculate ΔH

Example 2. Use bond enthalpies to find ΔH for the combustion of CH_4.

$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$$



O=O   O=C=O   H-O-H

	CH_4	O_2	CO_2	H_2O
D (kJ/mol)	413	495	799	467

① $E_{in} = 4(413) + 2(495) = 2642$ kJ

② $E_{out} = 2(799) + 4(467) = 3466$ kJ

③ $\Delta H = E_{in} - E_{out}$

$= 2642 - 3466 = -824$ kJ

$$\boxed{\Delta H = -824 \text{ kJ}}$$

∴ exothermic.

HW

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