What is the difference between "q" and ΔH?

energy ↗ q is the energy considering mass, substance, and Δ temp

$$q = m * C * \Delta T$$

Enthalpy — ΔH is the amount of energy from a reaction

$$\Delta H = \sum \Delta H_{f\,(products)} - \sum \Delta H_{f\,(reactants)}$$

relationship:

$$\Delta H = \frac{q}{mole}$$

$\frac{energy}{mole}$

two ways to determine:

physical collection (in lab) Calorimetry

theoretical (data table) final - initial bond energies

use data table in the front of your workbook

ΔH_f ← formation

Aluminium	H kJ/mol	S	G

ΔH_f – energy to form compound from element

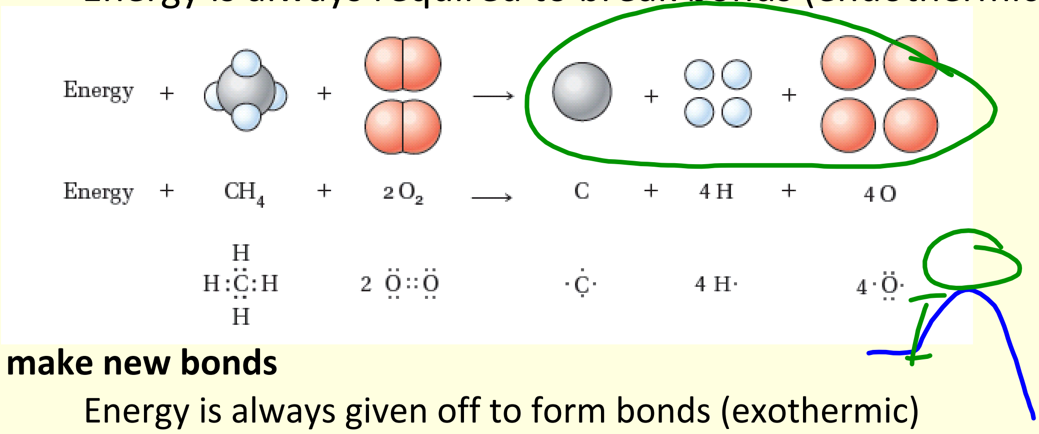
$$C + O_2 \rightarrow CO_{2(g)}$$

0 0 -393.5

Chemical Reaction--think of this as 2 steps

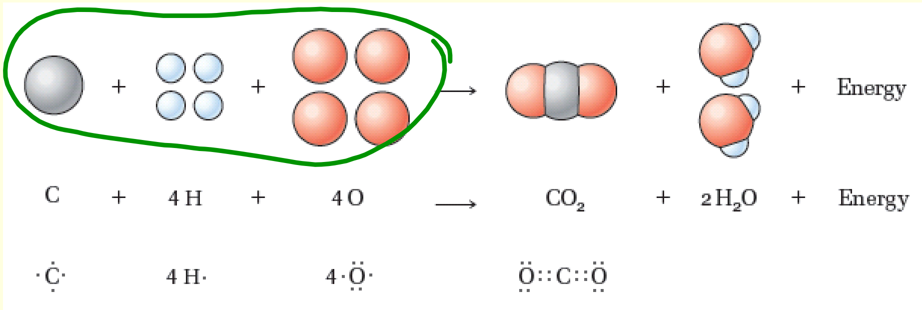
1. **break bonds** $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Energy is always required to break bonds (endothermic)



2. **make new bonds**

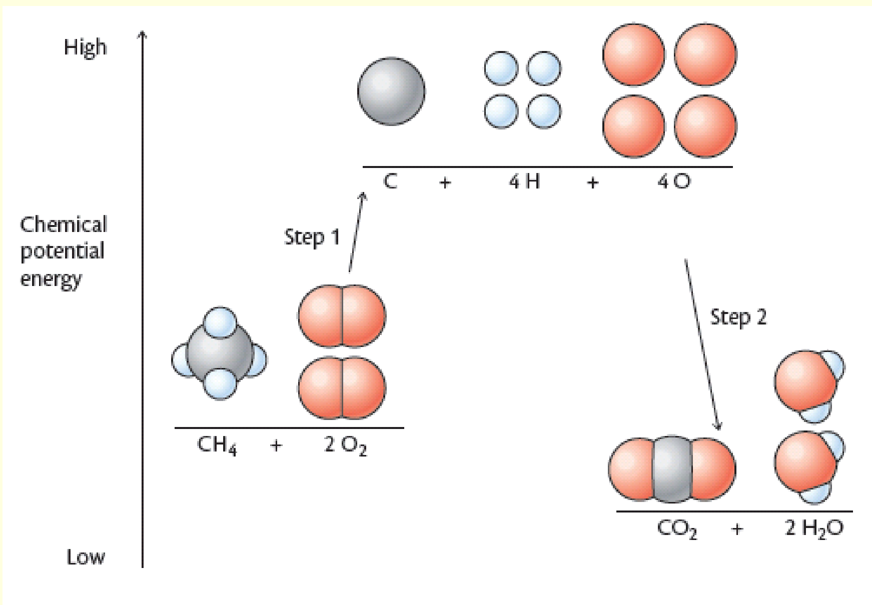
Energy is always given off to form bonds (exothermic)



Compare 2 steps to determine overall energy exchange

$$\Delta H = \sum \Delta H_{f\,(products)} - \sum \Delta H_{f\,(reactants)}$$

http://employees.oneonta.edu/viningwj/sims/bond_energy_dh_reaction_s.html



Calculate enthalpy of combustion (in kJ/mol) of CH_4

Write out the balanced combustion rxn

__$CH_4(g)$ + _2_$O_2(g)$ → ___$CO_2(g)$ + _2_$H_2O(g)$

(look on data chart)

ΔH^0_f $CH_4(g)$ = -74.8 kJ/mol
ΔH^0_f $H_2O(g)$ = -241.8 kJ/mol
ΔH^0_f $CO_2(g)$ = -393.5 kJ/mol

$CH_4(l)$ + $2O_2(g)$ → $CO_2(g)$ + $2H_2O(g)$

(-74.8) 2(0) (-393.5) 2(-241.8)

$$\Delta H = \sum \Delta H_{f\ (products)} - \sum \Delta H_{f\ (reactants)}$$

ΔH_{rxn} =[(-393.5kJ)+ 2(-241.8kJ)] - [(-74.8.0kJ)) +2(~~0~~ kJ)]

= -877.1 - (-74.8) = 802.3 KJ
per reaction cycle

This is the energy released when one mole of methane combusts.

$$\frac{802.3\text{ kJ}}{1\text{ mol } CH_4}$$

If I have 266 grams of water produced, how much energy will be produced?

266 grams |

$$\Delta H = \sum \Delta H_{f\,(products)} - \sum \Delta H_{f\,(reactants)}$$

State function: Final - Initial

this is relative (a comparison), not absolute

"snapshot" of energy in bonds
found by

H of formation of products - H of formation of products

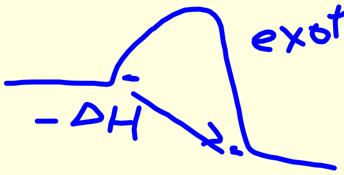
$2H_2 + O_2 \rightarrow 2H_2O_{(g)}$

2(0) 1(0) 2(-241.8)

0

ΔHf formation from element

1. Balance
2. Data
3. (x)·coefr.
3. $\Delta H = \sum \Delta H_{f\,prod} - \sum \Delta H_{f\,React.}$



$\Delta H = -483.6 - 0 = -483.6$

Calculate enthalpy of combustion (in kJ/mol) of methane, C_6H_6

Write out the balanced combustion rxn of Benzene (C_6H_6)

2 $C_6H_6(l)$ + 15 $O_2(g) \rightarrow$ 12 $CO_2(g)$ + 6 $H_2O(l)$

0

(look on data chart)

ΔH^0_f $H_2O(l)$ =

ΔH^0_f $C_6H_6(l)$ =

ΔH^0_f $CO_2(g)$ =

$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l)$

2(+49.0) + 15(0)   12(-393.5) + 6(-285.8)

$$\Delta H = \sum \Delta H_{f\ (products)} - \sum \Delta H_{f\ (reactants)}$$

Prod   Reactants

ΔH_{rxn} = (-4722 + -1714.8) − (+98)

−6436.8 − 98 = −6534.8 KJ ~~/mol~~

per reaction cycle

per 2 mole C_6H_6

This is the energy released when two moles of benzene combusts. We want the energy per one mole of benzene:

$$\frac{-6534.8\ \text{kJ}}{2\ \text{mol}\ C_6H_6} = 3267.4\ \text{kJ/mol}$$

C − 12×6 → 72
H 1×6 → 6
} 78 g/mol

If I have 429 grams of benzene, how much energy will be produced?

$$\frac{429\text{g}\ C_6H_6}{} \left| \frac{1\ \text{mol}\ C_6H_6}{78\ \text{g}\ C_6H_6} \right| \frac{-3267.4\ \text{KJ}}{1\ \text{mol}} = -17{,}971\ \text{kJ}$$