Enthalpy and Chemical Reactions

Combustion Reactions

1 mole

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} + 802.3\ kJ\ (heat)$$

(*Exothermic*)

$$\Delta H_{combustion} = -802.3\ \frac{kJ}{mol_{CH4}}$$

(*...per mole of substance burned*)

Formation Reactions

$$C_{(s)} + 2H_{2(g)} + \tfrac{1}{2}\ O_{2(g)} \rightarrow CH_3OH_{(g)} + 239\ kJ\ (heat)$$

1 mole

(*Exothermic*)

$$\Delta H_{formation} = -239\ \frac{kJ}{mol_{CH3OH}}$$

(*...per mole of substance formed*)

$\Delta H_{combustion}$ Problem

How much heat is produced when 10.0 grams of methane (CH_4) burn in the presence of excess oxygen gas?

16.04246

$CH_{4(g)}$ + $2O_{2(g)}$ → $CO_{2(g)}$ + $2H_2O_{(g)}$ + 802.3 kJ (heat) *(Exothermic)*

10.0 grams Excess

$$\frac{10.0\ g_{CH4}}{1} \times \frac{1\ mole_{CH4}}{16.04246\ g_{CH4}} \times \left(-802.3 \frac{kJ}{mol_{CH4}}\right) = -500.\ kJ$$

Negative sign tells us heat is released (exothermic) and...

...that the system's energy has decreased.

What has happened to the energy of the surroundings? ...the universe?

Enthalpy and Chemical Reactions

Heats of fusion: Melting a solid into a liquid

(Endothermic)

$$6.02\ kJ + 1\ mole\ H_2O_{(s)} \rightarrow H_2O_{(l)}$$

$$\Delta H_{fusion} = +6.02\ \frac{kJ}{mol_{H2O}}$$ *(...per mole of substance melted)*

Heat of vaporization: Vaporizing a liquid into a gas.

(Endothermic)

$$2260\ kJ + 1\ mole\ H_2O_{(l)} \rightarrow H_2O_{(g)}$$

$$\Delta H_{vaporization} = +2260\ \frac{kJ}{mol_{H2O}}$$ *(...per mole of substance vaporized)*

Why should it require so much more energy to vaporize water in comparison to melt it?

Because it requires much more energy to separate molecules to great distances found in gases.