Writing "Thermochemical" Equations

A thermochemical equation is a chemical equation that includes the value of ΔH

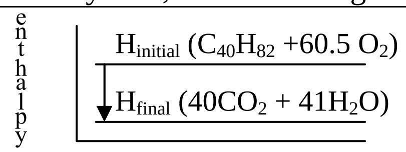
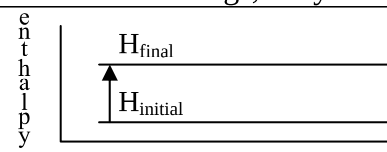
Example 1: Burning one mole of wax releases 20,000 kJ of heat energy

This could be written as: $C_{40}H_{82} + 60.5\ O_2 \longrightarrow 40\ CO_2 + 41\ H_2O + 20{,}000\ kJ$

Instead we usually write: $C_{40}H_{82} + 60.5\ O_2 \longrightarrow 40\ CO_2 + 41\ H_2O \quad \Delta H = -20{,}000\ kJ$

Why is ΔH negative? Recall that, by definition, $\Delta H = H_{products} - H_{reactants}$

In example 1: $H_{products} < H_{reactants}$, therefore ΔH is negative

	Exothermic	Endothermic
ΔH (+, –)	– (energy decreased/released)	+ (energy increased/absorbed)
Movement of Heat	From system, to surroundings	From surroundings, to system
"Enthalpy diagram"		

Notice in the above equation that the chemicals must be thought of as moles and not atoms/molecules. This is because a mole of molecules will absorb/release far more energy than a single molecule. In fact, if twice as many chemicals are involved, the reaction should produce twice as much energy:

$$2C_{40}H_{82} + 121\ O_2 \longrightarrow 80\ CO_2 + 82\ H_2O \quad \Delta H = -40{,}000\ kJ$$

Also, according to the law of conservation of energy, the reverse of a reaction should require the same amount of energy as the given reaction (but the energy will be absorbed instead of being released):

$$40\ CO_2 + 41\ H_2O \longrightarrow C_{40}H_{82} + 60.5\ O_2 \quad \Delta H = +20{,}000\ kJ$$

When you multiply a thermochemical equation, ΔH is also multiplied.
When a thermochemical equation is reversed, the sign of ΔH changes (law of conservation of energy)

Q – Write thermochemical equations for 1-3, sketch enthalpy diagrams for 1 and 2:

1. $C + O_2 \rightarrow CO_2 + 393.5\ kJ$
2. H_2 and I_2 react to form 2HI. This reaction requires 53.2 kJ
3. The reaction that forms one mole of HI

Standard Heats of Reaction

To compare the ΔH for different reactions, the same conditions must apply. Thus standard values for temperature and pressure are required. These have been set at 25°C and 1 atm. A " ° " symbol is added to ΔH to indicate that it refers to standard conditions. $\Delta H°$ is called the Standard Heat of Reaction.

$\Delta H° \equiv$ Standard Heat of Reaction $\equiv \Delta H$ at 25°C and 1 atm of pressure

Example 2 – Write the following as a thermochemical equation: 2 mol H_2 gas reacts with 1 mol O_2 gas to produce 2 mol H_2O(l). At 25°C and 1 atm of pressure this reaction releases 571.8 kJ of heat.

We write: $2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (l) \quad \Delta H° = -571.8\ kJ$

Notice that the state of H_2, O_2, and H_2O must be indicated. This is because $\Delta H°$ depends on the states of products and reactants (g = gas, l = liquid, s = solid):

$H_2\ (g) + ½\ O_2\ (g) \rightarrow H_2O$ (**l**) $\quad \Delta H° = -285.9\ kJ$

$H_2\ (g) + ½\ O_2\ (g) \rightarrow H_2O$ (**g**) $\quad \Delta H° = -241.8\ kJ$

Q – Do RE 5.45 – 5.50 on pg. 175