Additional notes on Hess's Law

1. ΔH_f° (element, most stable form) = 0 kJ/mol

<u>Examples</u>:
C(s, graphite), O_2(g), H_2(g) => $\Delta H_f^\circ = 0$.

Not C(s, diamond), O_3(g), H(g).

2. $\Delta H_{rxn} = \Sigma\Delta H_f^\circ(\text{products}) - \Sigma\Delta H_f^\circ(\text{reactants})$

Can be asked to solve for ΔH_{rxn}, ΔH_f° of a product, or ΔH_f° of a reactant.

3. $\Delta H_{rxn} = \Delta H_1 + \Delta H_2 + \Delta H_3 \ldots$

Can be asked to solve for ΔH_{rxn}, ΔH_1, etc.

Light and radiation

Light is a type of electromagnetic (EM) radiation, and light has energy. Many kinds of light exist.

Ultraviolet (UV) light causes skin to tan or burn.

Infrared (IR) light is used in heat lamps and remote controls.

Visible light has energy between UV and IR.

Some types of light and EM radiation have more energy and cause more damage than other types.

Examples: UV light/radiation vs. IR light,
X-rays vs. radio waves

Light and EM radiation have features of waves. (Figs. 7-1, 7-2)

Frequency (ν): "nu", the number of wave crests (peaks) that pass a fixed point in space in one second. Units are "per second" = 1/s = 1 s^{-1}. Also, known as Hertz (Hz), 1 Hz = 1/s.

Wavelength (λ): "lambda", the distance between wave peaks. Units are in length, such as meters or nm (nanometers).

Amplitude: the height of a wave, often measured from peak to trough, sometimes measured from peak to baseline. Also known as *intensity*, the brightness of a light source.

For all waves,
wavelength x frequency = speed of wave
$\lambda \quad x \quad \nu = \text{speed}.$

All EM radiation types behave like waves with wavelength (λ) and frequency (ν).

All EM radiation travels at the speed of light (c).

So $\lambda\nu = c$, where c = speed of light
$= 2.998 \times 10^8$ m/s

<u>Example</u>: A radio station transmits at 106.7 MHz. What is the wavelength of the radio signal? (reminder: mega = 10^6, Hz = 1/s)

Wave-particle concept of light and radiation

EM radiation (including light) also behaves like a collection of discrete particles.

Photon: a particle of light or radiation, a discrete packet or bundle of light or radiation. Photons have no mass.

Einstein came up with the idea of photons and experimentally found how to calculate the energy of a photon based on its frequency:

$E_{photon} = h\nu_{EM\ radiation}$

h = Planck's constant = 6.626×10^{-34} J-s

(Planck also developed some of our explanations of light and EM radiation, and Einstein saw how to apply Planck's work to his own work.)

So we use the frequency of the EM wave to calculate the energy of one photon of EM radiation with that frequency.

<u>Example</u>: Given that λ_{red} = 656.3 nm and λ_{blue} = 434.1 nm, what are the energies of a photon of red light and a photon of blue light? (Qualitative estimate first! Which should have more energy?)

Atoms can absorb photons and emit electrons
(photoelectric effect)

A photon transfers enough energy to an atom that the electrons gain enough energy to escape. Simple picture of escaping electrons

$M(s) \rightarrow M^{+} + e^{-}$

<u>Binding energy</u>: the energy needed by a photon to eject an electron from an atom, the energy needed by an electron to escape from the atom, called E_{min} in textbook.

If the photon transfers more energy than just the minimum (binding) energy, then the ejected electron can carry the extra energy as kinetic energy.

$E_{photon} = E_{min} + E_{\text{kinetic of ejected electron}}$

Atoms can absorb photons and emit photons
(emission spectrum and the Bohr model)

Simple simulation of photon absorption, energy loss, and photon emission:

http://www.avogadro.co.uk/light/bohr/spectra.htm

Key ideas:

1. Know ground state vs. excited state.

2. For visible light and UV radiation, the electron is actually the part of the atom that is changing in energy. The e^- can move to an excited state, and then emit a photon as it moves to lower energy levels.

3. The energy values for the electrons are restricted to specific values! (Fig. 7-8) Energy levels are not continuous; they cannot have any value. In other words, the energy values (levels) are <u>quantized</u>.

<u>Example</u>: Ball on a staircase model. The ball can only be at specific heights, not any height.

The planetary (Bohr) model and the stair model are oversimplifications of reality. The e^- does not literally jump across space to a new orbit or stair. The e^- moves to a different energy level, which is an abstract concept that does not fit a concrete model.

Describing the energy levels:

1. Bound electrons (electrons in atoms) have quantized energy levels.

2. Due to its wave characteristics, an electron in an energy level is delocalized over space without an exact location or exact motion. The electron is a three-dimensional wave-particle that occupies an energy level.

3. The periodic table lists these energy levels in order from lowest to highest energy. These energy levels are called atomic orbitals. The electrons occupy the atomic orbitals.

Bound electrons are delocalized in quantized energy levels that are three-dimensional and called atomic orbitals.

Mathematically, we can calculate the energy of the energy levels (i.e., the atomic orbitals):

$$E_n = -\frac{2\pi^2 m e^4 Z^2}{(4\pi\varepsilon_o)^2 h^2} \times \frac{1}{n^2}$$

n = 1 is the first level (the ground state)

n = 2 is the second level (the first excited state, 1 level above the ground state)

n = 3 is the third level (the second excited state, 2 levels above the ground state)

This equation and variations of this equation are only correct for hydrogen and hydrogen-like ions (ions with only 1 electron).
Examples: H, He^+, Li^{2+}, etc.

More complicated equations are needed for other situations.

Most often the equation is used to calculate ΔE for the electronic transitions in H-like atoms.

$$\Delta E = -2.179 \times 10^{-18}\ \mathrm{J} \left(\frac{1}{n^2_f} - \frac{1}{n^2_i} \right) = h\nu$$

<u>Example</u>:
Calculate the energy for the n = 3 to n = 2 transition in a hydrogen atom.

Some variations of this equation replace ΔE with ν_{photon} or $1/\lambda_{photon}$.