Week 5/Th: Heat, Lecture Units ‘11 & 12’

Unit 11: Thermodynamics
-- definitions, heat, work, system, $\Delta E$
-- 1st law
-- energy and temperature
-- specific heat vs. heat capacity
-- another example of 1st Law

Unit 12: Heat capacity
-- heating curves, change of state
-- “PV” work, enthalpy, state function

Unit 13: Chemical Energy
-- Hess’s Law (aka conservation of energy)
-- bond energies

Issues: ?

NO Homework this week due to exam

“Hold it, sweetie, you forgot your heat.”

New Yorker Cartoon by Lee Lorenz

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Week 5/Th: Exam Results

CEM 141 Fall 2012 Exam 1

Count | 2344
Mean  | 109.0
Median| 110
Mode  | 110
SD    | 31.7
Range | 170
Minimum| 10
Maximum| 180

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Q: What is the specific heat of a metal based on the fact that immersion of a 20 g sample at 75°C into 100 g of water initially at 25°C will warm the system to only 27 °C?

\[
\Delta E = q + w, \text{ but no work (w) and no external heat flow (q)}
\]

\[
\Delta E = 0
\]

Energy lost cooling metal + energy gain heating water = 0

Energy loss = 20g \times X J/g-\circ C \times (27-75) \circ C

= 20 \times X \times (-48) J

Energy gain = 100g \times 4.18J/g-\circ C \times (27-25) \circ C

= 418 \times (2) J

-960 X +836 = 0

\[\Rightarrow \quad \frac{836}{960} = X\]

\[\Rightarrow \quad 0.871 = X \quad [J/g-\circ C]\]
Some chemical processes involve changing the “state” of the components – let’s consider just state changes for a pure material; define the words.
Some chemical processes involve changing the “state” of the components – let’s consider just state changes for a pure material that everyone is familiar with: an ice cube (10 g).

Slopes give Specific Heats
Steam: 2.0 J/g-K
Water: 4.2
Ice: 2.1

Flat Regions give “Latent Heats”
Melt: 333 J/g  Vaporize: 2260 J/g

Calculated Curve for 10 g ice at 10 J/s
Q: What is the final temperature for 128 mL cup of water at 76.5°C if an ice cube (20.5 g) at 0°C is put into the water?

Another case with no external heat or work .. \( \Delta E = 0 \)
One of the important types of work in a chemical system is that of physical expansion or compression. For example, the combustion of octane (organic molecule) in air is central to modern life. \[ 2 \text{C}_8\text{H}_{18} (l) + 25 \text{O}_2 (g) \rightarrow 16 \text{CO}_2 (g) + 18 \text{H}_2\text{O} (g) \]

Notice that the piston is open to atmosphere

A) If the piston is seized, then no change in volume, no work, only heat
\[ \Delta E = q + 0 \]
\[ \Delta E = q_V \text{ “the heat at constant volume”} \]

B) The piston moves, work comes out that is volume increase times the pressure
\[ \Delta E = q + w = q_p - P \Delta V \]
\[ \Delta E + P \Delta V = q_p \text{ “heat at constant pressure”} \]
It is clearly easier to carry out chemical reactions under a constant pressure (open to the atmosphere) than it is to carry them out at constant volume. Thus, $q_P$ is more “useful” to describe chemical processes than $q_V$ – thus, it gets a special name: enthalpy and the symbol $H$. \[ q_V = \Delta E \quad . \quad q_P = \Delta H \]

$\Delta E$ represents the change in internal energy by the process

$\Delta H$ represents the observed change in energy in atmosphere

Both of these quantities represent internal chemical potential energy and are called state functions because they only depend on what the chemical substance is and its ‘state’ or condition.
You all have experience with another state function: gravitational energy.

Notice that these functions are relative and not absolute (no “zero”).
Enthalpy

**Exothermic** process: heat OUT, \( \Delta H \) is a negative number

\[ 2 \left[ \text{H}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{H}_2\text{O} + 242 \text{ kJ} \right] \]
heat or energy is a product
\( \Delta H = -242 \text{ kJ per mole of water} \)

**Endothermic** process: heat IN, \( \Delta H \) is a positive number

\[ 2 \left[ 90\text{kJ} + \frac{1}{2} \text{N}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{NO} \right] \]
heat or energy is a reactant
\( \Delta H = +90 \text{ kJ per mole of NO} \)