Chem 115 Class Notes
3/25/08

Calometry and Hess ‘s law

“Calor” means heat and “metry” means measure, therefore calorimetry is the measurement of heat that get absorbed or released.
Examples of calometry include Heat transferred (which can be represented as q) from a hot metal block placed in a known amount of cool water.

See problem solving example in last lecture’s class notes.

Heat Changes Generally

The graph on page 4 of the lecture slide show how heat energy input affects the temperature of a substance from the solid state to the vapor state.

- Heating up the solid: temperature increases. Kinetic energy increases when temperature increases, particles move faster, still in crystal lattice structure. Potential energy essentially stays the same (actually increases a tiny bit).
- Changing phase from solid to liquid: temperature remains constant. Kinetic energy remains constant, particles are not moving faster. Potential energy increases. Once fully a liquid, particles can move around past each other, no longer restricted to lattice structure. Amount of energy required (endothermic) per mole is called heat of melting. Same amount as in the other direction (liquid to solid) except then it’s called heat of fusion and it is exothermic.
- Heating up the liquid: temperature increases. Kinetic energy increases when temperature increases, particles move faster. Potential energy remains essentially the same (actually increases a tiny bit).
- Changing phase from liquid to vapor: temperature remains constant. Kinetic energy remains constant, particles are not moving faster. Potential energy increases, this is easy to see because the particles get a lot further apart. Amount of energy required (endothermic) per mole is called heat of vaporization. Same amount as in the other direction (vapor/gas to liquid) except then it’s called heat of condensation and it is exothermic.
- Heating up the gas: temperature increases. Kinetic energy increases when temperature increases, particles move faster.

As of now, we have learned that:
- System can either lose or gain heat during a change
- Exothermic is the heat that is released while Endothermic is the heat that is absorbed . -
- We use calorimetry to measure heat changes. when liquid water absorbs heat from the system, the temperature of the water increases but when the water gives heat to the system, the temperature of the water decreases
- Enthalpy is state of function and because of this, an enthalpy change can be calculated by (1) Hess’s law, or (2) standard heat of formation. There is a third way to calculate

enthalpy change that is an estimate: (3) Bond dissociation energies. This third way is not based on Hess's law.

Calorimeter Example: (question is on page 6 of the lecture slide)

Our goal is to find ΔH in kJ/mol of $AgNO_3$ that reacts.
T initial = 22.20°C
T final = 23.11
Q water = mCΔT
Where m= 100g, C= 4.184 J/g°C, T= T final- T initial= 23.11- 22.20= 0.91°C
q_{water}= (100.0g) (4.184 J/g°C) (0.91°C) = 380.74 J
Convert to 380.74 to kJ =0.38074 kJ
Where does mol of $AgNO_3$ comes from?
You know how much $AgNO_3$ reacted from the initial solution used

$$\text{Molarity} = \frac{\underline{\text{Moles Solute}}}{\text{L of solution}}$$

Moles solute ($AgNO_3$) = (Molarity of $AgNO_3$)(l of $AgNO_3$)

$$(0.100\text{m})(50.0\text{ml})\frac{(\underline{1\text{L}})}{(1000\text{ml})}$$

= 0.005000mol $AgNO_3$ reacted
ΔH= q/ moles reacted
= 0.38074 kJ / 0.00500 mol
= 76.148 kJ/mol
= 76 kJ/mol, because there are 2 sig figs

It is exothermic because water (surrounding) temperature increased so system must have released heat.