Chem 115 Class Notes 3/13/08 KF

THERMOCHEMISTRY (Ch.5)

- **thermochemistry** is the portion of thermodynamics that deals with the relationship between chemical reactions and the exchange of energy in the form of **heat (q)**. *Another form of energy is *work*, but we're focusing on heat for now. (Take a look at the Lecture 13 slides, top of page 2 to see a general synopsis of thermochemistry)
-energy is used in the making and breaking of bonds between atoms. When bonds break, energy in the form of heat is absorbed (ie, it takes energy input to break a bond). In contrast (the opposite), when a bond is formed, energy is released (given off) by molecules.

Kinetic Energy vs. Potential Energy

- kinetic energy is the energy of motion. Anything that is moving has kinetic energy. The higher the object's velocity (or speed), the more kinetic energy it has. Particle level energy includes:

-thermal energy (a.k.a. heat)
-electrical energy: electrons moving through a conductor
-sound energy: particles vibrate in a material where they expand and compress
-light energy: photons are used as energy in photosynthesis

-potential energy is stored energy that can be transformed into kinetic energy. For example, a bike standing motionless at the top of a hill has lots of potential energy but no kinetic energy. Once the bike begins to roll downhill, its potential energy decreases as its kinetic energy increases. (See figure 5.2 in book, bottom of page 168)

Laws of Thermodynamics

1) Energy is neither created nor destroyed; it is just transferred from one place to another.
2) Energy transfer happens in such a way that energy becomes spread out or more chaotic.

Check out these websites on heat transfer:
http://www.colorodo.edu/physics/2000/bec/temperature.html
http://jersey.uoregon.edu/vlab/Thermodynamics/therm1a.html
http://jersey.uoregon.edu/vlab/Thermodynamics/index.html

Systems vs. Surroundings

- a **system** is the defined portion of matter (imagine the matter is in a 3-D box where heat can be transferred but not matter) under study in thermodynamics. In contrast, the **surroundings** consist of everything that is not the system, or everything outside the system.

Ex: **system**: human body
surroundings: air, clothes, chair, pen, etc.

Ex: **system**: a thermos containing coffee
surrounding: air, counter top, etc.

- energy, or heat, can be transferred from a system to its surroundings and the surroundings can also transfer heat to a system. The terms **endothermic** and **exothermic** are used to describe the direction of the heat exchange during a reaction.

- from the perspective of a system:
  - in an **endothermic** reaction--------heat is *received by the system* by the surroundings
  - in an **exothermic** reaction----------heat is *given by the system* to the surroundings

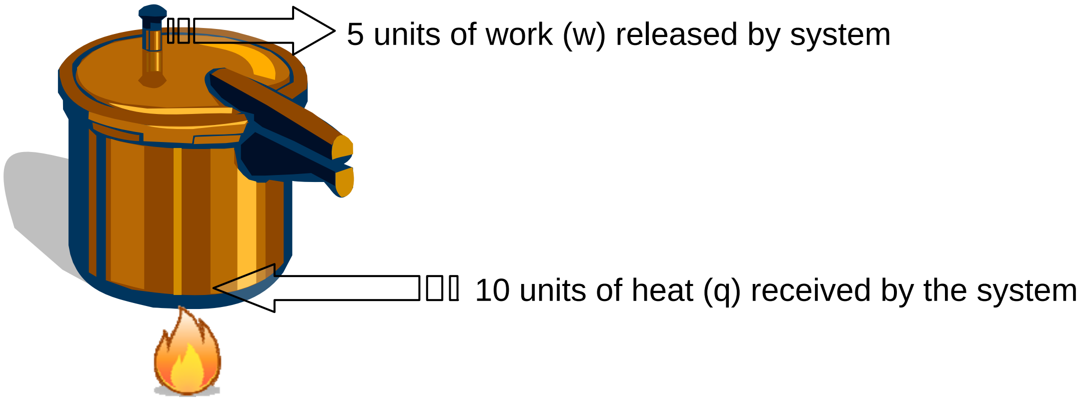
Internal Energy (E)

- the **internal energy**, symbolized as **E**, of a system is the sum of kinetic energy and potential energy of all the system's particles. The internal energy can change depending on heat and work. A change of internal energy is represented as delta E, or Δ**E**.

- since **E** is the total amount of energy *in* a system, a negative value of Δ**E** would mean that the system lost energy and a positive value for Δ**E** means that the system gained energy.

$$\Delta \mathbf{E} = E_{Final} - E_{Initial}$$

Ex: **system = inside the pot**



So, $\Delta \mathbf{E} = q + w = (+10 \text{ units}) + (-5 \text{ units}) = +5$

If the initial energy inside the pot was $E_{Initial} = 100$ units, then

$\Delta E = E_{Final} - E_{Initial}$ so $E_{Final} = E_{Initial} + \Delta E = (100 \text{ units}) + (5 \text{ units}) = 105$ units

Enthalpy

- the term **enthalpy** refers to the heat flow processes happening at a constant pressure without forms of work except **pressure-volume work**. Enthalpy, also written as H, equals the internal energy plus the product of pressure and volume:

$$H = E + PV$$

- most chemistry systems are concerned with enthalpy (H) as opposed to internal energy (E).
- also, at constant pressures, $\Delta H = q$ (heat). A negative value for ΔH means there was a loss of heat (exothermic). A positive value for ΔH means there was a gain of heat (endothermic).

-enthalpy is a **state function**, meaning that a system's value is determined by its present state regardless of the steps taken to get to this present state. The steps are not state functions and include heat and work. (See book, page 175) You can also think of altitude when thinking of a state function. You can take many different paths to travel from sea level to 3,000 ft. above sea level but your ending point is the same no matter which path you choose. (See Lecture 13 slides, top of page 10)

Relative vs. Absolute Temperature Scales

-Relative temperature scales = Fahrenheit and Celsius measure temperature of something relative to two different points.

-Absolute temp. scales = Kelvin measures the absolute motion of particles in something

-the size of 1 degree C = to the size of 1 K

State Functions In Rxns

- Consider the following rxn:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_20_{(g)} \qquad \Delta H = -571.66kJ$$

- when gaseous (*g*) water is formed from hydrogen and oxygen gases, 571.66 kJ of heat are released/lost…this is why the value of ΔH is negative.

- If the reaction is reversed:

$$2H_20_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)} \qquad \Delta H = 571.66kJ$$

notice that the ΔH value is exactly the same amount as before but positive, meaning a gain of heat. This is consistent with the First Law of Thermodynamics.

Calorimetry

-**calorimetry** is the measuring of heat flow from one thing to another.
-**heat capacity** is the amount of heat needed to raise the temperature of an object by 1 K (or 1 degree C). Water has a relatively high heat capacity: 4.184 J/gK. This means that one gram of water requires 4.184 J to raise it 1 K.
-this equation is helpful in calorimetry problems:

$$q \text{ (heat)} = m \text{ (mass)} \times C \text{ (heat capacity)} \times \Delta T \text{ (change of temperature)}$$

-this equation can help you find one out one of these four values if you are given the values of three of them.

Sample Problem

An aluminum block is placed into a beaker of water under certain conditions…

Given Info:
Initial temperature of an aluminum block = 90.00°C
Mass of Al block = 5.00g
Mass of water = 100.0g
Temperature of water before = 23.00 °C
Temperature of both after = 23.71°C
C of water = 4.184 J/gK

Question:
What is the heat capacity of Al?

1) We need to find C of Al. Before this, we need to find the ΔT of Al:

$$\Delta T_{Al} = 23.71^{o}C - 90.00^{o}C = -66.29\ K$$

2) So far, we have:

$$q_{Al} = 5.00g \times C_{al} \times (-66.29K)$$
$$q_{Al} = C_{al} \times (-331.45gK)$$

3) Since we still can't get the value of C_{al} or q_{Al}, we can turn to the given information for water for help:

$$q_w = 1000.0g \times 4.184J/gK \times .71K$$
$$q_w = 297.06J$$

4) Following the First Law of Thermodynamics, we can state that:

$$q_{Al} = -q_w$$
$$q_{Al} = -297.06J$$

5) Put the value of q_{Al} over -331.45gK to get C_{Al}

$$\frac{-297.06\ J}{-331.45\ g{\cdot}K} = 0.90\ J/gK \quad \text{(2 sig. figs. because of 0.71 K)}$$

General problem solving strategy with calorimetry problems:

1) Figure out what the system is. Usually the surroundings will be water (or the whole solution). Usually you will have to figure out the heat change of the surroundings, and then the heat change of the system is opposite that.

2) Calculate the heat change using $q = m\,C\,\Delta T$. Note that in a system where you have an aqueous solution, the mass (m) will be the mass of the whole solution (solvent + solute), so if it's a solid solute that you're adding to water, you'll have to include the mass of the solute in it.

3) If the problem requires you to calculate ΔH as J/mol, then you next have to find out how many moles reacted so you can put it in the denominator, using $\Delta H = q/\text{mol}$

4) Make sure the sign of ΔH is correct (+ if the system had an endothermic change, - if the system had an exothermic change)

5) Check that the sig figs are correct (ΔT is usually what limits these calculations since in the subtraction to calculate ΔT you often lose sig figs)