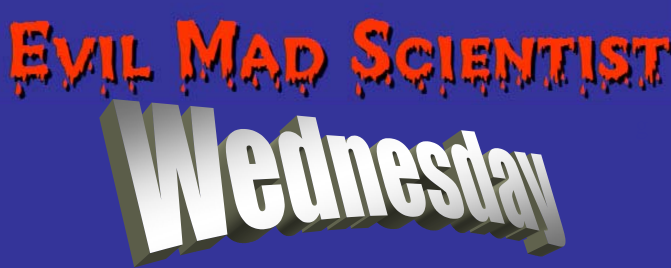
Chemistry 6A F2007



10/31/07

Exam 2: Friday 11/2/07 (here in lecture)

What will be covered on the exam?

- Chapter 4: (4.6-4.9 and 4.11)
- Chapter 5: All
- Chapter 6: (6.1-6.8)
- Any thing from lab as well

What do I need to bring?

Bring a Pencil, Eraser, Calculator and scamtron form 882

YOU NEED TO KNOW YOUR LAB SECTION NUMBER!

Heat and the Specific Heat Capacity:

✧ When heat is absorbed or lost by a body, the temperature must change as long as the phase (*s, g or l*) remains constant.

✧ The amount of heat (q) transfer is related to the mass and temperature by:

q = heat lost or gained (J)

m = mass of substance (g)

$$q = m \times C \times \Delta T$$

C = the Specific Heat Capacity of a compound $\left(\frac{J}{g \cdot {}^\circ C}\right)$

ΔT is the temperature change in degrees Celsius or Kelvins

The Specific Heat Capacity :

✧ "C" reflects the temperature change for a specific substance that absorbs or loses heat (*q*) per gram.

✧ When a Substance with a high heat capacity absorbs a given amount of heat it undergoes a small temperature change.

$$\Delta T = \frac{q}{m \times C}$$ ⟵ **big "C", little ΔT** / **little "C", big ΔT**

Examples: Water Glass Space Shuttle tiles } ***insulators***

✧ The opposite holds true for "C"

Examples: metals water vapor } ***heat conductors***

Lecture problem :Calculate the amount of heat that must be removed form 46.8g of water at 56°C to bring the temperature down to 22°C.

$$q = m \times C \times \Delta T \qquad \Delta T = T_f - T_i$$

$$C_{H_2O} = 4.184 \frac{J}{g \cdot {}^{o}C}$$

$$q = 46.8g \times 4.184 \frac{J}{g \cdot {}^{o}C} \times (22^{o}C - 56^{o}C)$$

$$= -6.7 \times 10^3 \text{ J} \quad \text{or} -6.7 \text{ kJ}$$

The negative sign means heat was lost.

The specific heat of lead is: $0.128 \frac{J}{g\,{}^{o}C}$

The specific heat of sulfur is: $0.706 \frac{J}{g\,{}^{o}C}$

If 10.0 g of each initially at 25.0°C absorbs 25 cal of heat, which will have the highest final temperature?

Recall: $q = m \times C \times \Delta T$

Rearranging: $\Delta T = \frac{q}{m \times C}$

By definition: $\Delta T = T_f - T_i$

Solving: $T_f = \frac{q}{m \times C} + T_i$

$$T_f = \frac{q}{m \times C} + T_i$$

For lead: $$T_f = \frac{25 \text{ cal} \times \frac{4.184 \text{ J}}{\text{cal}}}{10.0g \times 0.128 \frac{J}{g \cdot {}^{o}C}} + 25.0^{o}C = 110\ {}^{o}C$$

For sulfur: $$T_f = \frac{\left(25 \text{ cal} \times \frac{4.184 \text{ J}}{\text{cal}}\right)}{10.0g \times 0.706 \frac{J}{g \cdot {}^{o}C}} + 25.0^{o}C = 40.\ {}^{o}C$$

The smaller the specific heat: The greater the temperature change.

Changes of State:

Energy Changes Accompanying Phase Changes

The sequence:

heat solid → form liquid → heat liquid → form gas (vapor) → heat gas

(melt) *(boil or vaporize)*

is ***endothermic***. *(energy must be supplied)*

The sequence:

cool gas → form liquid → cool liquid → form solid → cool solid

(condense) *(freeze)*

is ***exothermic***. *(energy is released)*