Introduction to Thermodynamics:

Energy -The capacity to do work (w) or transfer heat (q).

<u>Kinetic Energy - energy of motion:</u> Molecular kinetic energy can be translation, rotational and vibrational.

<u>Potential Energy - stored energy:</u> Molecular potential energy can be the either nuclear or columbic. There are two basic types of columbic energy, those within an atom, and those between atoms. The first are the attractions of electrons to the nucleus of an atom, the second are the covalent and ionic bond energies associated with compounds

<u>Law of Conservation of Energy:</u> (First Law of Thermodynamics): The energy of the universe is constant and can not be created or destroyed, only transformed from one form to another.

Exothermic and Endothermic Changes: If the potential energy of the bonds is not the same between reactants and products the first law of thermodynamics state that energy must be released or absorbed during a chemical reaction.

Exothermic changes release energy. These types of reactions cause an increase in the temperature or a phase change like melting or evaporation.

Endothermic changes absorb energy. These types of reactions cause a decrease in the temperature or a phase change like freezing or condensation.

Units of Energy: As energy can be descried as the capacity to do work or transfer heat, there are two basic units, the Joule (work based) and Calorie (heat based).

Joule: $1J = 1N \cdot m = \left(\frac{1kg \cdot m}{s^2} \right)$, note how this definition can be expressed in terms of SI base units.

Calorie: 1 calorie is the heat energy required to raise 1 g of water from 14.5 to 15.5^{o}C.

1cal = 4.184J

Temperature: A measure of something's "hotness". Adding/removing energy to/from a substance causes its temperature to change unless it undergoes a phase transition.

Heat Capacity: [C (j/^{o}C] The amount of heat an object absorbs or releases as it undergoes a temperature change.

Specific Heat Capacity: [C (j/g$\cdot$ ^{o}C] The heat capacity of a substance on a per gram basis. That is, the heat capacity of an object depends on its mass, so for a given object of mass m, C = mc. Or: $q=mc\Delta T = mc(T_F - T_I)$

Problem Sets:

1. Fill in the blanks for the following table of specific heat capacities. Use this table for the rest of this exercise

Substance	c (cal/g ^oC)	c (j/g ^oC)
Water	1.000	
Ice		2.06
Lead		0.127
Iron	0.444	
Tungsten	0.134	
Gold		0.129

2. 35.0 kJ raises the temperature of a gold object from 22^oC to its melting point of 1064.18^oC. What is the mass of the gold object?

3. What is the specific heat capacity in cal/g^oC of a substance if it takes 5.09 kJ to raise a 30.1 g mass sample of the material from 25.0 to 400.0^oC?

4. How much energy (in Joules) is absorbed when 100.0g of two different substances are raised from a temperature of 10^oC to 100^oC if the substances are:

a. Gold

b. Water

c. How much more heat did it take to raise the water to 100^oC than the gold?

d. What are the ratio of the heat required for the water to the gold?

e. What are the ratio of the heat capacities of water to gold?

5. Find the final temperature when 7.25kJ of heat is added to a 132.0 g object if at 10.0^{o}C if:

a. The object is made of iron

b. The object is made of water

Heat Transfer from a hot object to a cold object in an isolated system. Since energy is conserved, the energy lost by a hot object ($-Q_H$) is gained by the cold object (Q_C)

$$Heat\ Lost = Heat\ Gained$$

so

$$-q_{Hot} = q_{Cold}$$
$$-\left(mc\Delta T\right)_H = \left(mc\Delta T\right)_C$$
$$-m_H c_H \Delta T_H = m_C c_C \Delta T_C$$

6. What is the specific heat capacity of an unknown material if 69.07 g sample of the material at 100.0^{o}C is dropped into 100.0 ml of water at 20.0^{o}C in an isolated system and cools to a final equilibrium temperature of 23.0^{o}C.

Do Chapter 3 ext Problems: 63,64,66,68-70,72,74,76,80,81,91