Energy & Chemical Reactions

Chapter 6

The Nature of Energy

Chemical reactions involve energy changes

- Kinetic Energy - energy of motion
  - macroscale - mechanical energy
  - nanoscale - thermal energy
  - movement of electrons through conductor - electrical energy
  - $E_k = (1/2)\ mv^2$
- Potential Energy - stored energy
  - object held above surface of earth - gravitational energy
  - energy of charged particles - electrostatic energy
  - energy of attraction or repulsion among electrons and nuclei - chemical potential energy

The Nature of Energy

Energy Units

- SI unit - joule (J)
  - amount of energy required to move 2 kg mass at speed of 1 m/s
  - often use kilojoule (kJ) - 1 kJ = 1000 J
- calorie (cal)
  - amount of energy required to raise the temperature of 1 g of water 1°C
  - 1 cal = 4.184 J (exactly)
  - nutritional values given in kilocalories (kcal)

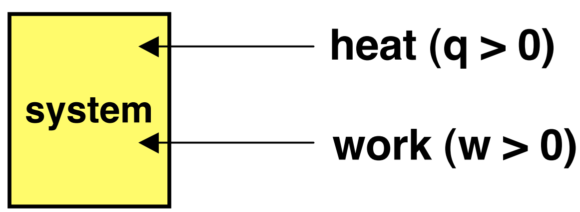
First Law of Thermodynamics

The First Law of Thermodynamics says that energy is conserved. So energy that is lost by the system must be gained by the surrounding and vice versa.

- Internal Energy
  - sum of all kinetic and potential energy components of the system
  - $E_{initial} \rightarrow E_{final}$
  - $\Delta E = E_{final} - E_{initial}$
  - not possible to know the energy of a system but easy to measure energy changes

First Law of Thermodynamics

- ΔE and heat and work
  - internal energy changes if system loses or gains heat or if it does work or has work done on it
  - $\Delta E = q + w$
  - sign convention to keep track of



  - both heat added to system and work done on system increase internal energy

First Law of Thermodynamics

Sign Conventions Used and the Relationship Among q, w, and ΔE

Sign Convention for q:	**Sign of $\Delta E = q + w$**
$q > 0$: Heat is transferred from the surroundings to the system	$q > 0$ and $w > 0$: ΔE . 0
$q < 0$: Heat is transferred from the system to the surroundings	$q > 0$ and $w < 0$: The sign of ΔE depends on the magnitudes of q and w
Sign Convention for w:	$q < 0$ and $w > 0$: The sign of ΔE depends on the magnitudes of q and w
$w > 0$: Work is done by the surroundings on the system	$q < 0$ and $w < 0$: ΔE , 0
$w < 0$: Work is done by the system on the surroundings	

First Law of Thermodynamics

Divide the universe into two parts

- system - what we are studying
- surroundings - everything else

$\Delta E > 0$	system gained energy from surroundings
$\Delta E < 0$	system lost energy to surroundings

Work and Heat

Energy can be transferred in the form of work, heat or a combination of the two.

Work is done against a force over a certain distance:

$$w = F \times d$$

Heat is energy transferred from hotter object to colder one.

Energy - capacity to do work or to transfer heat

Example 1

Calculate the change in the internal energy of the system for a process in which the system absorbs 140 J or heat from the surroundings and does 85 J of work on the surroundings.

$q = + 140$ J

$w = -85$ J

$\Delta E = q + w$

$\Delta E = 140 \text{ J} - 85 \text{ J} = + 55 \text{ J}$

Heat Capacity

- determines the temperature change experienced by an object when it absorbs energy
- amount of heat required to raise temperature by 1 degree
- the larger the heat capacity the greater the amount of heat needed to raise T

Heat Capacity

- molar heat capacity - heat capacity of one mole of substance
- specific heat capacity - heat capacity of one gram of substance

$$c = \frac{\textbf{quantity of heat transferred}}{\textbf{(grams of substance) x (temp. change)}} = \frac{q}{m\,\Delta T}$$

Example 2

How much heat is needed to warm 300 g of water from 20ºC to 95ºC? The specific heat of water is 4.18 J/g ºC.

$q = (4.18\ \mathrm{J/g\ ^\circ C})(300\ \mathrm{g})(95^\circ\mathrm{C} - 20^\circ\mathrm{C})$

$q = 94050\ \mathrm{J} = 94\ \mathrm{kJ}$

Calorimetry

- assume not heat loss or gain through calorimeter
- any heat lost by rxn (or substance) is gained by solution
- any heat gained by rxn (or substance) is lost by solution
- $q_{sol'n} = -q_{rxn}$ OR $q_{sol'n} = -q_{metal}$

$$c_{sol'n} \times m_{sol'n} \times \Delta T_{sol'n} = -c_{metal} \times m_{metal} \times \Delta T_{metal}$$

Example 3

When 50.0 mL of 0.100 M $AgNO_3$ and 50.0 mL of 0.100 M HCl are mixed in a calorimeter, the temperature of the mixture increases from 22.30ºC to 23.11 ºC. The temperature increase is caused by this reaction:

$$AgNO_3\ (aq) + HCl\ (aq) \rightarrow AgCl\ (s) + HNO_3\ (aq)$$

Calculate the heat transferred in this reaction, assuming that the combined solution has a mass of 100.0g and a specific heat of 4.18 J/g ºC.

Example 3

$q_{sol'n}$ = (4.18 J/g ºC)(100.0 g)(23.11 ºC - 22.30 ºC)
$q_{sol'n}$ = 338.58 J

q_{rxn} = –$q_{sol'n}$ = –338.58 J = –339 J