Energy transfer for physical processes

For physical processes, we will focus on kinetic energy and thermal energy. Nanoscale pictures relating kinetic molecular theory, kinetic energy, and thermal energy: Figures 6.5, 6.6, and p. 189.

For chemical processes, we will additionally study how chemical potential energy relates to thermal energy. In CHEM 103, chemical potential energy will also be related to electrical energy.

The two main concepts you need to apply are
1. conservation of energy
2. mole ratio

The main units and conversions you need are
1. 1 calorie = 4.184 Joules, 1 cal = 4.184 J
2. kilo = 1000, 1 kilojoule = 1 kJ = 1000 J

Conservation of Energy

Energy can neither be created nor destroyed. The total energy of the universe is constant. (Law of Conservation of Energy, First Law of Thermodynamics)

Example:
A hot bar of copper at 100 °C is placed into cold water at 4.0 °C. What is the final temperature of the bar and water?

Internal energy of each substance
(kinetic and thermal energy)

Heating (heat): the transfer of thermal energy between substances with different temperatures when they are in contact with each other.

Need to know how much energy is needed to change the temperature of each substance (the copper and the water).

Heat capacity: the amount of thermal energy gained or lost when the temperature of a substance changes by one degree.

For water, 4.184 J of energy for every 1°C change for each gram of water that changes temperature. For copper, 0.385 J per 1 °C per 1 gram.

Written as *specific heat capacity, c*:

$$c_{water} = 4.184 \frac{J}{g\ {}^{\circ}C} = 4.184\ J\ g^{-1}\ {}^{\circ}C^{-1}$$

$$c_{copper} = 0.385 \frac{J}{g\ {}^{\circ}C} = 0.385\ J\ g^{-1}\ {}^{\circ}C^{-1}$$

Example:

A hot bar of copper at 100.0 °C is placed into 75.0 mL of 4.0 °C cold water. The final temperature of the bar and water reaches 6.3 °C. Determine the mass of the copper bar.

$q = c \times m \times \Delta T$ = thermal energy transferred

Substance 1: $q_1 = c_1 \times m_1 \times \Delta T_1$
Substance 2: $q_2 = c_2 \times m_2 \times \Delta T_2$

<u>Conservation of energy</u>

1. Energy transfer between substances is equal.
2. Direction of energy transfer is opposite.

$$q_1 = -q_2 \quad \text{(same as } q_2 = -q_1\text{)}$$

$$c_1 \times m_1 \times \Delta T_1 = -(c_2 \times m_2 \times \Delta T_2)$$

Or $m_1 c_1 \Delta T_1 = -(m_2 c_2 \Delta T_2)$

Can be asked to solve for any of the 8 variables:
m_1, c_1, $T_{i,1}$, $T_{f,1}$, m_2, c_2, $T_{i,2}$, $T_{f,2}$

For amounts in grams (mass = m), we must use specific heat capacity = c.

$$c = \frac{J}{g\text{-}^\circ C} \quad \text{or} \quad \frac{cal}{g\text{-}^\circ C} \qquad (\text{sometimes } c_p,\ s)$$

$$q_1 = -q_2$$

$$m_1 c_1 \Delta T_1 = -(m_2 c_2 \Delta T_2)$$

For amounts in moles (number of particles = n), we must use molar heat capacity = C.

$$C = \frac{J}{mol\text{-}^\circ C} \quad \text{or} \quad \frac{cal}{mol\text{-}^\circ C} \qquad (\text{sometimes } c_{mol},\ C_{mol})$$

$$q_1 = -q_2$$

$$n_1 C_1 \Delta T_1 = -(n_2 C_2 \Delta T_2)$$

Practice:

A 400.0 g piece of iron is heated in a flame and then immersed in 1000.0 g of water at 20.0 °C. Afterward, the water and iron are at 32.8 °C. What was the initial temperature of the iron? c_{iron} = 0.451 J/g-°C. (hint: solve for ΔT)

Energy transfer that changes the temperature of a substance (Sec. 6.3)

ΔE = amount of substance

x

energy to change T of 1 gram of substance by 1°

x

total change in T.

$\Delta E = \Delta q = q = mc\Delta T$

Note: Δq is usually written simply as q, with the assumption that the reader knows q means Δq.

Energy transfer that changes the phase of a substance (Sec. 6.4)

ΔE = amount of substance

x

energy to change the phase of 1 gram of substance.

$\Delta E = \Delta q = q = m\Delta H$.

ΔH = enthalpy = amount of energy transfer.

ΔH_{fus} = enthalpy of fusion = energy transfer to change between liquid and solid = freezing or melting. (333 J/g for H_2O)

ΔH_{vap} = enthalpy of vaporization = energy transfer to change between gas and liquid = boiling or condensing. (2260 J/g for H_2O)

During a phase change, the substance does not change temperature. No ΔT occurs until the phase change is done, e.g., all the ice melts.

<u>Example</u>:
Ice left at 0 °C or warmer will melt into water. The energy to change water between solid and liquid is 333 J/g (ΔH_{fus} = 333 J/g).

10.0 grams of water will require energy transfer of:

$$\Delta E = \Delta q = m \times \Delta H$$
$$= 10 \text{ g} \times 333 \text{ J/g} = 3330 \text{ J} = 3.33 \text{ kJ}.$$

The ice gains energy from the surrounding air.

We use the definition that gaining energy is a positive change:

Δq_{ice} = ice gains energy = + 3.33 kJ (+ sign).

Losing energy is defined as a negative change,
Δq_{air} = air loses energy = - 3.33 kJ (- sign).

$\Delta q_{ice} = -\Delta q_{air}$ (conservation of energy)

The ice is considered the "system" of interest because chemists are most interested in the substance undergoing the change (melting in this case).

The air is considered the "surroundings", that is, what interacts with the system to exchange energy.

<u>endothermic</u>: a process in which the system gains energy (+ sign).
<u>exothermic</u>: a process in which the system loses energy (- sign).

Ice undergoes an endothermic process to melt.

Practice:

Water cooled to 0 °C will freeze and become ice. How much energy (in J) must be transferred to freeze 2.00 moles of water? Is the change of state an endothermic or exothermic process?