Unit 6 – Thermodynamics
Ch 5, 11

Endothermic and exothermic processes, *ch 5.1, 5.2*

New vocab

Energy

Work

Heat

Kinetic energy

Potential energy

Chemical energy

Thermal energy

Units for energy: Conversions for energy units: ________J = _______cal

During endothermic processes, the reactants have higher/lower (*circle one*) energy than the products. What would you notice about an endothermic reaction?

During exothermic processes, the reactants have higher/lower (*circle one*) energy than the products. What would you notice about an exothermic reaction?

List five everyday processes and label if they are endothermic or exothermic. How can you tell?

Heating a substance is endothermic/exothermic (*circle one*). Why?

Cooling a substance is endothermic/exothermic (*circle one*). Why?

Are the phase changes s → l → g endothermic or exothermic? What is happening to the heat? What is happening to the particles?

Are the phase changes g → l → s endothermic or exothermic? What is happening to the heat? What is happening to the particles?

Dissolution means…

Dissolution can be endo- or exo- depending on…

What is the system? Give an example.

What are the surroundings? Give an example.

Where is the energy going during an endothermic process? How is work involved?

Where is the energy going during an exothermic process? How is work involved?

Energy diagrams, *ch 14.5*

(Review from kinetics)

Draw an energy diagram for endo- and exo- processes. Label the axes.

Endo-

Exo-

Heat transfer

Particles of a warmer body have a higher ______________________________ than a cooler body.

How can energy transfer from one particle to another?

What is thermal equilibrium?

What are the particles doing once a process reaches thermal equilibrium?

Heat capacity and calorimetry, *ch 5.5*

How can heat transfer be measured?

Write down your observations and questions for the water and isopropanol demo.

Observations | Questions?

$\boxed{q = mC\Delta T}$

q = | m = | C = | ΔT =

Sample problem

How much heat is needed to warm 250 g of water from 15˚C (the previous temperature in our classroom) to 95˚C (the approximate boiling point in Denver)?

Using the same amount of heat, what would be the final temperature of a 250 g piece of iron?

Practice 1

1. The specific heat of graphite is 0.71 J/gºC. Calculate the energy needed to raise the temperature of 75 g of graphite from 294 K to 348 K.

2. Aluminum has a specific heat capacity of 0.902 J/g ^{o}C. How much energy is release when 1.0 kg of aluminum cools from 35 ^{o}C to 20 ^{o}C?

What is the first law of thermodynamics? Give some examples that you see in everyday life.

Applying 10 J of heat energy to a gram of water will raise the temperature by 2.4^{o}C. Applying the same amount of heat to a gram of gold will raise the temperature by 78^{o}C. Explain.

Explain what $q_{rxn} = -q_{surroundings}$ and $q_{rxn} = - q_{soln}$ mean.

Explain how a coffee cup calorimeter works.

Sample problem

A 46.2 g sample of copper is heated to 95.4ºC and then placed in a calorimeter containing 75.0 g of water at 19.6ºC. The final temperature of both the water and the copper is 21.8ºC. What is the specific heat of copper?

Practice 2

1. A piece of metal weighing 59.047 g was heated to 100.0 °C and then put it into 100.0 mL of water (initially at 23.7 °C). The metal and water were allowed to come to an equilibrium temperature, determined to be 27.8 °C. Assuming no heat lost to the environment, calculate the specific heat of the metal.

2. What is the final temperature when 0.032 kg of milk at 11˚C is added to 0.16 kg of coffee at 91˚C? Assume the specific heat capacities of the two liquids are the same as water, and disregard any energy transfer to the surroundings.

Other terms

Heat capacity

Molar heat capacity

Practice 3

1. Determine the heat needed to increase the temperature of 10.0 g of mercury by 7.5˚C. The molar heat capacity for mercury is 27.8 J/mol˚C.

2. The specific heat of iron is 0.451 J/g˚C. What is the *molar* specific heat of iron?

When hot water is added to cold water inside a calorimeter, what happens to the heat?

Sample problems

When 40.0 mL of water at 60.0˚C is added to 40.0 mL water at 25.0˚C already in a calorimeter, the temperature rises to 15.0˚C. What is the heat capacity of the calorimeter?

When 50.0 mL of 0.10 M HCl(*aq*) and 50.0 mL of 0.10 M NaOH(*aq*), both at 22.0 °C, are added to a coffee cup calorimeter, the temperature of the mixture reaches a maximum of 28.9 °C. What is the approximate amount of heat produced by this reaction?

Practice 4

1. 50.0 mL of water at 40.5 °C is added to a calorimeter containing 50.0 mL of water at 17.4 °C. After waiting for the system to equilibrate, the final temperature reached is 28.3 °C. Calculate the heat capacity of the calorimeter (just the calorimeter without water).

2. When 100. mL of 0.200 M NaCl(*aq*) and 100. mL of 0.200 M $AgNO_3$(*aq*), both at 21.9 °C, are mixed in a coffee cup calorimeter, the temperature increases to 23.5 °C. Write a balanced equation and net ionic equation. How much heat is produced by this precipitation reaction? What assumptions did you make to determine your value?

3. When 3.12 g of glucose, $C_6H_{12}O_6$, is burned in a bomb calorimeter, the temperature of the calorimeter increases from 23.8 °C to 35.6 °C. The calorimeter contains 775 g of water, and the bomb itself has a heat capacity of 893 J/°C. How much heat was produced by the combustion of the glucose sample?

Phase changes, *ch 11.4*

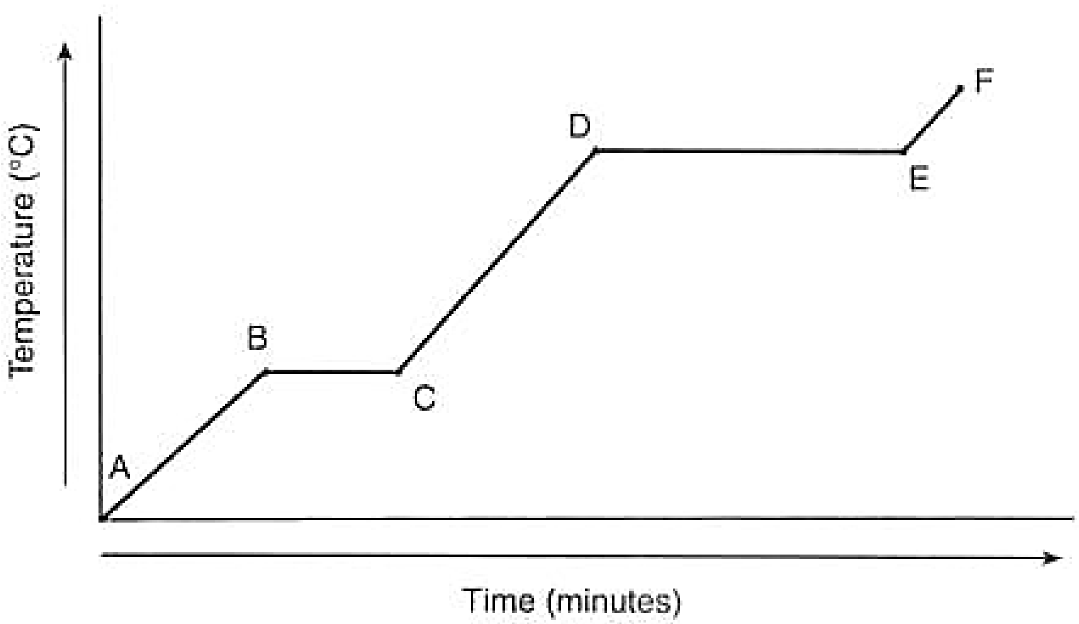
Changing a substance from s → l → g requires/releases (*circle one)* heat; the energy of the system ________________

Changing a substance from g → l → s requires/releases (*circle one)* heat; the energy of the system ________________

During a phase change, the temperature of a __________ substance remains ______________

What do you notice on the diagram? (slide 40)

Label the following heating curve with the states and transitions. What is the energy used for? Include the melting point and boiling point of the substance.



Practice 5

1. Calculate the heat required to change 9.00 g of solid H_2O at -25 ˚C to vapor at 125 ˚C. The specific heats of ice, liquid water, and steam are 2.03 J/g·K, 4.184 J/g·K, and 1.84 J/g·K, respectively. ΔH_{fus} = 6.01 kJ/mol and ΔH_{vap} = 40.67 kJ/mol.

2. Calculate the heat released when 9.00 g of H_2O vapor at 125 ˚C is cooled to a solid at -25 ˚C. The specific heats of ice, liquid water, and steam are 2.03 J/g·K, 4.184 J/g·K, and 1.84 J/g·K, respectively. ΔH_{solid} = -6.01 kJ/mol and $\Delta H_{consense}$ = -40.67 kJ/mol.

How are vapor pressure and temperature related? Why?

Enthalpy of reaction, *ch 5.4*

$\Delta H = q$ at <u>constant pressure</u>

Sample problems (just stoichiometry)

1. $CO_2(g) + 2\ H_2O(l) + 890\ kJ \rightarrow CH_4(g) + 2\ O_2(g)$

How much heat is needed for 10.0 g of carbon dioxide to react?

2. $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H = -890\ kJ$

How much heat is released with 10.0 g of carbon dioxide is produced?

Practice 6

1. $2\ H_2O_2(l) \rightarrow 2\ H_2O(l) + O_2(g) + 196\ kJ$

Calculate the quantity of heat released when 5.00 g of hydrogen peroxide decomposes.

2. $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H = -890\ kJ$

How much heat is released with 4.50 g of methane gas is burned?

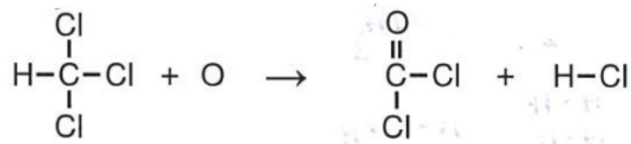
Bond enthalpies

Breaking bonds ______________ energy

Forming bonds _________________ energy

How do you estimate the enthalpy of the reaction using the reactants and products?

Sample problem



Enthalpy of formation, *ch 5.7*

For ΔH_f° ° indicates ________________________________and $_f$ indicates ______________________

Practice 7

Write the equation for the formation of the following compounds from its elements in their standard states:

1. Sodium oxide

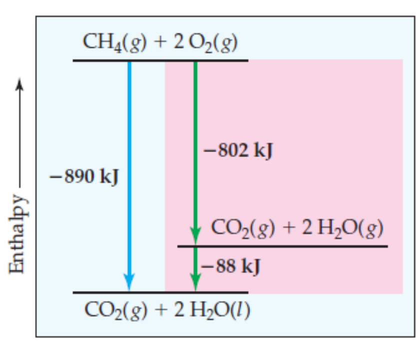
2. Potassium chloride

3. Glucose ($C_6H_{12}O_6$)

Hess's law, *ch 5.6*

What is a state function? Give an example from your life.

What is this diagram showing?



What is Hess's law? How can you manipulate the equations?

Sample problem

(1) $C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_1 = -393.5 \text{ kJ}$

(2) $CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_2 = -283.0 \text{ kJ}$

Using these data, calculate the enthalpy for the combustion of C to CO:

(3) $C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g) \qquad \Delta H_3 = ?$

Practice 8

1. Calculate ΔH for the reaction 2 C(s) + H_2(g) → C_2H_2(g) given the following equations and their respective enthalpy changes

$$C_2H_2(g) + \tfrac{5}{2}O_2(g) \longrightarrow 2\,CO_2(g) + H_2O(l) \qquad \Delta H = -1299.6 \text{ kJ}$$
$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5 \text{ kJ}$$
$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l) \qquad \Delta H = -285.8 \text{ kJ}$$

2. Calculate ΔH for the reaction NO(g) + O(g) → NO_2(g) given the following equations and their respective enthalpy changes

$$NO(g) + O_3(g) \longrightarrow NO_2(g) + O_2(g) \qquad \Delta H = -198.9 \text{ kJ}$$
$$O_3(g) \longrightarrow \tfrac{3}{2}\, O_2(g) \qquad \Delta H = -142.3 \text{ kJ}$$
$$O_2(g) \longrightarrow 2\, O(g) \qquad \Delta H = 495.0 \text{ kJ}$$

Describe how dissolving a solute in a solvent can be an endo- or exothermic process

$$\boxed{\Delta H^\circ_{rxn} = \Sigma n \Delta H^\circ_f(\text{products}) - \Sigma m \Delta H^\circ_f(\text{reactants})}$$

Sample problem

Calculate the standard enthalpy change for the combustion of 1 mole of benzene, C_6H_6(l) to its products (H_2O is liquid).

Practice 9

1. Calculate the enthalpy change for the decomposition of calcium carbonate into calcium oxide and carbon dioxide.

2. Calculate the enthalpy change for the combustion of 1 mol of C_2H_5OH (the water is liquid).

3. Calculate the standard enthalpy of formation of CuO(s).

What happens if the products of a reaction are at a different temperature from the surroundings?