*Unit 11: Thermochemistry*



*Important Vocabulary for Thermochemistry*

Energy

Average Kinetic Energy (KEavg)

Exothermic

Endothermic

System

Surrounding

Universe

Heat

Energy

Temperature

Enthalpy / Heat of Reaction (ΔH)

Calorimeter

Calorimetry

Reaction Coordinate or Reaction Pathway

First Law of Thermodynamics

Phase Diagram

Phase Change

Melting/Freezing

Evaporating/Condensing

Sublimation/Deposition

Heat of Vaporization (Hv)

Heat of Fusion (Hf)

*Thermochemistry*

Most chemical and physical changes involve a measureable change in energy but what does this value of energy tell us? Often we will use a thermometer to measure the temperature change of a system, but *temperature is not the same thing as energy*. So what is energy?! For now let us settle on a description of energy that we can put to use – ***Energy is the ability to do something useful***.

Energy can be absorbed or released during a physical or chemical change. The ***First Law of Thermodynamics*** states that the total energy content of the universe is constant – energy may not be created or destroyed; only rearranged. A useful diagram for illustrating this point is shown below.

*Surroundings*

The Universe

*System*

Exothermic

Endothermic

For us, the ***system*** will typically be a chemical reaction or physical object. The ***surroundings*** is everything else that is not the system! Heat flow out of the system is an ***exothermic*** process. Heat flow into the system is an ***endothermic*** process. All energy exchanges take place inside the ***universe*** – energy is conserved, rest easy. Remember this, all energy changes take place from the point of view of the *system.*

This chapter will address 6 big questions about Thermochemistry:

1. How is the average kinetic energy of a substance related to its phase? (KMT)
2. What can we infer about a substance from its phase diagram? (Phase Diagrams and phase changes)
3. How can the energy change of a physical process be measured and quantified? (q=Cat, phase change energy, heating curve? What can we infer from a graphical representation of the energy change of a physical process?
4. How can we measure the energy changes of physical and chemical processes?
5. How can the energy change of a chemical reaction be quantified? Big mamma, heats of formation
6. What can we infer from a graphical representation of the energy change of a chemical reaction? (rxn coordinate)

*Notes: Energetics of Phase Changes*

Kinetic energy is the energy a particle has due to its motion: (1/2 *mv2*). Particles that move faster will have a greater kinetic energy. Temperature is a measurement of the average kinetic energy of a substance. With these basic principles we can outline a mechanism for how a substance such as water is able to transform into a solid when the temperature is lowered and why it rapidly vaporizes when its temperature is raised. We will start with the gas phase.

* Like all gases, water vapor is assumed to obey the ***Kinetic Molecular Theory (KMT)***. This theory presumes that gases consist of constantly moving particles that have negligible volumes and do not attract or repel one another. When collisions happen between particles, the resulting trajectories and speeds are predictable, similar to billiard balls. Physicists call these collisions “perfectly elastic.” If gases truly behaved in this ‘ideal’ fashion, solids and liquids could never exist! How then does any gas ever transition into a condensed state of matter (liquid and solid)? We must overlook a tenant of the KMT in order to answer this…colliding gas particles do attract one another.
	+ Imagine you are walking down the hall way and spot your friend walking towards you. You intend to shake their hand but you don’t want to stop because you will be late for chemistry! Would it be easier to shake hands if you sped up of slowed down? Imagine both of you running towards each other and trying to shake hands!
* Condensed states of matter happen when gas molecules are moving slow enough to “shake hands.” The forces of attraction (hand shakeness?) that are responsible for holding solids and liquids together are the intermolecular forces (IMFs): hydrogen bonds, dipole interactions and dispersion forces. When the temperature of a gas is lowered, the average kinetic energy is lowered. Since kinetic energy is directly proportional to mass and velocity, the velocity of the gas particle must be lowered as well.
* The freezing point and condensing point of a substance are the temperatures at which molecular motion is slow enough for intermolecular forces to overcome the tendency for a molecule to bounce off of another molecule. When water vapor condenses into liquid, hydrogen bonds hold the water molecules close together but not so rigidly as to prevent molecules from sliding past one another. The ability of the molecules to slide past one another is what characterizes water as a liquid. If the molecular motion becomes slow enough that the IMF no longer allows molecular sliding, the substance freezes…and that’s where solids come from.
* Pressure also determines the phase (solid, liquid, gas) of a substance. We will examine the relationship between temperature, pressure and phase in the next section.

**Questions to consider:**

1. How is the phase of a substance related to average kinetic energy?
2. What has to happen to molecules in the liquid state in order for them to become gases?

*Notes: Phase Diagrams*

1. A ***phase diagram*** is a graphical representation of the relationship between the temperature, pressure and phase of a substance. It depicts the conditions necessary for any phase change a substance can undergo. We will begin this section by examining the phase diagram of water in order to tease out as much information as we can.



* There are three main sections of the graph: solid, liquid, and gas.
* Curve B represents the conditions necessary for water to ***freeze*** or ***melt***.
* Curve C represents the conditions necessary for water to ***condense*** or ***vaporize***.
* Curve A represents the conditions necessary for water to ***sublimate*** or ***deposit***.
* Point D is called the ***triple point***. This point represents the pressure and temperature at which all three phases are in equilibrium.
* The normal boiling point is the temperature at which a substance boils at 1 ATM. Label it as X on your graph. 1 ATM = 760 mm Hg = 101.3 kPa
* The normal freezing point is the temperature at which a substance freezes at 1 ATM. Label it as Y on your graph 1 ATM = 760 mm Hg = 101.3 kPa
* Point E is referred to as the ***critical point***. This is the minimum temperature at which a gas cannot be condensed into a liquid, regardless of how high the pressure goes. The state of the substance beyond this point is called a *supercritical fluid*.
* Curve B shows a negative slope (it leans to the left) which is rare. This indicates that solid water (Ice) can be converted into liquid by increasing the pressure. This is known as regelation*. A negative solid/liquid boundary slope indicates that the density of the solid state is less than the density of the liquid state*. In other words, the solid form of the substance floats in the liquid form. Ice floats in water.

**Questions to consider:**

1. Do you know which sections correspond to solid, liquid and gas?
2. Do you know how to find the normal boiling and freezing point at *standard pressure?*
3. Do you know what a negative slope for solid/liquid boundary represents

*Notes: Specific Heat*

1. The energy (heat) required to change the temperature of a substance and to change the phase of a substance can both be quantified with relatively simple but slightly different equations. This section will be divided into two main parts: 1) heat changes within one phase and 2) heat changes of phase changes. The discussion will conclude by combining both concepts to evaluate heating curve data.
* We will start by looking at how an intensive property known as ***specific heat capacity (C)*** affects a substances ability to change temperature when heat energy is added or taken away. Temperature is what determines the direction of heat flow between two substances. Heat always flows from high temperature to low temperature. *When heat flows into or out of a substance, its temperature changes*. The specific heat capacity of a substance describes how much the temperature changes when a given amount of heat is absorbed.
	+ Imagine a bucket of soapy water sitting on the ground next to a black car on a hot July afternoon. Not a cloud in the sky. The sun has been shining on the bucket of water and the car for the same amount of time. What might you notice about the water when you reach in to grab the rag? What might you notice about the black metal car when you accidentally lean on the hood?

The density and intermolecular forces exhibited by a substance determine the magnitude of specific heat. Metals tend to have low specific heat while non-metals and compounds tend to be higher. Water has a relatively high specific heat value. It takes a lot of energy to overcome the extensive hydrogen bonding in water in order to increase the average kinetic energy of the molecules. In plain terms, water takes a long time to heat up. Below is a table of specific heat values. Note the differences between metals and non-metals.

|  |
| --- |
| **SPECIFIC HEAT VALUES OF COMMON SUBSTANCES AT 20°c (J/g°C)** |
| **Substance** | **Sp. Ht** | **Substance** | **Sp. Ht.** | **Substance** | **Sp. Ht.** |
| Air | 1.000 | Ethyl Alcohol,C2H5OH | 2.405 | Mercury | 0.139 |
| Aluminum | 0.895 | Fluorine | 0.824 | Nickel | 0.443 |
| Asbestos (Mg3Si2O5(OH)4) | 2.660 | Gold | 0.129 | Paraffin(C31H64) | 2.13 |
| Calcium | 0.652 | Granite | 0.803 | Platinum | 0.134 |
| Carbon (diamond) | 0.502 | Hydrogen | 14.03 | Silver | 0.233 |
| Carbon (graphite) | 0.711 | Ice | 2.060 | Stainless Steel | 0.510 |
| Carbon dioxide | 0.832 | Iron | 0.448 | Steam | 2.020 |
| Chlorine | 0.477 | Lead | 0.128 | Tin | 0.220 |
| Copper | 0.387 | Magnesium | 1.010 | Water | 4.184 |

*Notes: Calorimetry of single phase temperature changes*

The process of calculating temperature changes and specific heats is known as calorimetry. Knowing the specific heat of a substance allows us to calculate the energy required to change its temperature. The specific heat formula is:

Where

q = Heat energy in joules

m = mass in grams

Cs = specific heat in J / g °C

ΔT = change in temperature in °C. Change is always final minus initial.

Please note that this formula is used only for calculating heat when the temperature *of one phase* changes. Sometimes you’ll hear this equation as “q equals M cat.”

Some common units for energy or heat units are

calories (cal); kilocalories (kcal) or food calories (Cal.); Joules (J)

Conversion factor: 1000 calorie = 1 kilocalorie = 1 **C**alorie = 4.184 joules

Ex. 1) How much heat energy is required to raise the temperature of 44.0 grams of lead from room temperature, 21oC, to its melting point, 327oC?

Ex. 2) What mass of silver is present if it requires 225 Joules of heat to increase the temperature from 13.5°C to 45.0°C?

Ex. 3) To what temperature will 117 grams of nickel be heated if it starts at 18.3oC and absorbs 2.46x103  joules? The specific heat of nickel is 0.444 J/g\*oC.

# Questions to consider:

# How do you calculate change in temperature?

# What is the physical significance of specific heat?

# Can you use q=mCAT to determine how much energy is required to melt ice?

1. What does the title at the top of this page mean?

*Notes: Calorimetry of phase changes*

Next we will examine the quantitative relationship between energy and phase changes. ***The energy required to change the phase of a substance does not change its temperature***. During a phase change, the added energy disrupts intermolecular forces instead of increasing the average kinetic energy of the molecules.

* + The energy required to change a solid to liquid at its equilibrium temperature is called the ***Latent Heat of Fusion (Hf)***. The added energy overcomes enough of the intermolecular forces to allow the molecules to slip past one another and flow. This value represents the number of joules of heat energy required to change one gram of a substance from a solid to liquid. To calculate the energy required to melt on object we only to know its mass and Hf.

Where

q = Heat energy in joules

m = mass in grams

Hf = latent heat of fusion in J/g

Notice that temperature is not a variable in this equation. Why not?

* + The energy required to change a liquid to a gas at its equilibrium temperature is called the ***Latent Heat of Vaporization (Hv)***. Since all intermolecular forces must be overcome in order to completely separate the particles into the gas phase, the magnitude of the Heat of Vaporization is much higher than that of the Heat of fusion. To calculate the energy required to boil a substance, substitute Hv into the equation above for Hf

Below are values of Heat of Fusion and Heat of Vaporization for some common substances

|  |  |  |
| --- | --- | --- |
| Substance | Heat of Fusion (J/g) | Heat of Vaporization (J/g) |
| Copper | 205 | 4726 |
| Ethyl alcohol, C2H5OH | 109 | 879 |
| Gold | 64.5 | 1578 |
| Silver | 88 | 2300 |
| Water | 334 | 2260 |

Ex. 1) How much heat must be added to 75 g of water at 100°C to boil it?

Ex. 2) If 1500 J of heat energy is added to a block of ice at 0oC, how many grams of ice will melt?

**Questions to Consider:**

1. Why is temperature not included in the equation above?
2. Why is the magnitude of Hv larger than Hf?
3. What does the title at the top of this page mean?

*Notes: Calorimetry and Heating Curves.*

The last part of this section will explain how to calculate the heat required to change the temperature and phase of a substance. The main focus for you the student is to ensure that you break apart the problem into separate steps…temperature changes and phase changes.

* + The graph below shows the relationship between temperature and heat energy during phase changes of water. This is graph is called a ***Heating Curve***. Locate the appropriate specific heats and heats of phase changes for water in the previous two tables then write them on the graph below.

Temperature oC

*Melt*

*Freeze*

*Boil*

*Condense*

Heat Energy

BP

FP

***Solid***

*Csolid­*

A

***Gas***

***Liquid***

*Hf*

B

*Cliquid*

C

*Hv*

D

*Cgas*

E

* + The heating curve for most substances has the same basic form as shown above. The boiling point and freezing point will be different, but the same five areas exist in all cases. ***In order to calculate the heat energy for all or part of this process, the correct equation must be used for each area***.
1. Warming the solid q = mCsΔT Cs = specific heat capacity of solid (J/goC)
2. Melting the solid q = mHf Hf = heat of fusion (J/g)
3. Warming the liquid q = mClΔT Cl = specific heat capacity of liquid (J/goC)
4. Boiling the liquid q = mHv Hv = heat of vaporization (J/g)
5. Warming the gas q = mCvΔT Cv = specific heat capacity of gas (J/goC)

**Questions to consider:**

1. How does the temperature change during a phase transition?
2. How does the energy change during a phase transition?
3. Is it possible for separate phases to coexist at the same temperature?
4. What is significant about the slope each segment on the heating curve?
5. What is significant about the lengths of the phase transitions on the heating curve?

Heating curve example problems

Ex 1.) Calculate the amount of heat required to convert 5 grams of ice at -25ºC to steam at 120ºC.



BP = 100ºC Csolid = 2.06 J/gºC

FP = 0ºC Cliquid = 4.18 J/gºC

Hf = 334 J/g Cvapor = 2.02 J/gºC

Hv = 2260 J/g

1. Heat the solid to the melting point (change the temp)
2. Melt the solid (change the phase):
3. Heat the liquid to the boiling point (change the temp):
4. Boil the liquid (change the phase):
5. Heat the gas to the required temperature (change the temp):



Ex. 2) Calculate the amount of heat needed to cool 25.0 g of the vapor at 100ºC to a solid at -50ºC.

The substance is known to have to following properties:

BP = 72.5ºC Csolid = 1.54 J/gºC

FP = -43.5ºC Cliquid = 1.86 J/gºC

Hf = 8.32 J/g Cvapor = 1.15 J/gºC

Hv = 26.3 J/g

1. Cool the gas to the required temperature:
2. Condense the liquid (change the phase):
3. Cool the liquid to the freezing point:
4. Freeze the liquid (change the phase):
5. Cool the solid to the required temperature:

Ex. 3) How much heat is required to convert a 50.0 g block of silver from 90ºC to 3100ºC? Is this endothermic or exothermic?



BP = 2212ºC Csolid = 0.187 J/gºC

FP = 961ºC Cliquid = 0.233 J/gºC

Hf = 88 J/g Cvapor = 0.192 J/gºC

Hv = 2300 J/g

A. Heat the solid to the melting point (change the temp)

B. Melt the solid (change the phase):

C. Heat the liquid to the boiling point (change the temp):

D. Boil the liquid (change the phase):

E. Heat the gas to the required temperature (change the temp):

*Notes: Energy Coordinate Diagram*

1. **Reaction Pathways** are often represented as energy coordinate diagrams that illustrate important characteristics about the energetics of a reaction. Reactions only take place when the reactant molecules collide with enough energy to cause the reactants to recombine into products. The effective collisions of molecules transmute kinetic energy into ***potential energy (PE)***. An energy diagram like the one below shows how energy (y-axis) changes throughout the course of a chemical reaction (x-axis). Notice that the energy level of the reactants must first be increased before it can be lowered to that of the products. When the reactants collide at the proper angle with energy equal to the ***activation energy (Ea)*** they create a ***transition state*** in which they form an ***activated complex (AC)****.* The ***heat of reaction (ΔH)*** is the difference in energy between reactants and products. If the products have less energy than the reactants, then heat has been released and the reaction is ***exothermic*** (shown below.) *A* ***catalyst*** *provides a different pathway for the reactants to take*. The alternate pathway has a lower activation energy. Alternately, an ***inhibitor*** raises the activation energy.
* Energy Coordinate diagram for an ***exothermic* *reaction*** **(**- **ΔH)**

AC

Energy

Exothermic Reaction Pathway

(By sign convention, ΔH is negative)

Ea (reverse reaction)

Energy of Products

ΔH

Ea

Reaction Pathway with catalyst

Energy of Reactants

Course of Reaction

* ***Endothermic reactions (+ΔH)***: energy is absorbed so the reactants have a lower energy than the products. Draw an energy coordinate diagram for this type of reaction in space below. Use appropriate labels. A table that compares endothermic and exothermic reactions is shown below:

|  |
| --- |
| **Exothermic vs. Endothermic reactions** |
| **Exothermic (- ΔH)** | **Endothermic (+ ΔH)** |
| Reactants | Products | Reactants | Products |
| Higher P.E. | Lower P.E. | Lower P.E. | Higher P.E. |
| Less stable | More stable | More stable | Less stable |
| Weaker bonds | Stronger bonds | Weaker bonds | Stronger bonds |

*Notes: Stoichiometry of Thermochemistry*

The energy released or absorbed during a chemical reaction is referred to as the ***Enthalpy of Reaction (ΔH)*.**

Other names for enthalpy of reaction include: heat of reaction, heat of combustion, molar heat of reaction, molar heat of combustion , enthalpy of combustion.

Energy lost or gained is typically measured as kilojoules and can be treated as if it were a reactant or product.

Example: How much ethane must be combusted to produce 1250 kJ of heat?

**2C2H6 + 7O2 → 4CO2 + 6H2O + 345 kJ**

Example: How much heat is absorbed when 0.722 g of cobalt (II) carbonate decomposes into cobalt (II) oxide and carbon dioxide?

**CoCO3 + 81.6 kJ → CoO + CO2**

*Notes: Hess’ Law, a.k.a. the Big Mamma Equation.*

ΔH°rxn = Σ ΔH°f (products) minus Σ ΔH°f (reactants)

* ΔH° = standard heat of reaction or *enthalpy* of reaction. The “°” symbol means standard conditions…25oC, 1 ATM pressure and 1 molar for all solutions.
* Σ = “sum of” as in add up all of the values.
* ΔH°f = standard heat of formation or enthalpy of formation. This is the energy change associated with forming a compound from its component elements. *The standard heat of formation of all elements in their natural state is 0. Nada. Zip.*
* Other common names for heat of reaction (ΔHrxn­) are: heat of combustion (ΔHcomb), heat of solution (ΔHsoln), enthalpy of reaction (ΔHcomb), enthalpy of solution (ΔHsoln). All of them have ΔH.
* In order to solve these problems, you must have a table of Standard Heats of Formation OR the values must be given to you in the problem.
* Having fractions as coefficients in the balanced equation is okay!

|  |
| --- |
| Standard Enthalpies of Formation ΔH°f kJ/mol |
| C2H5OH  (l) | -278 |
| O2 (g) | zero |
| CO2 (g) | -393.5 |
| H2O (l) | -285.8 |

Ex. 1) Calculate the standard enthalpy of combustion for the following reaction:

C2H5OH (l) + O2 (g) → 2 CO2 (g) + 3 H2O (l)

ΔH°rxn = Σ ΔH°f (products) - Σ ΔH°f (reactants)

ΔH°comb = [ **2** (-393.5) + **3** (-286) ] - [ (-278) + **(7/2)** (0) ] =

**Question to consider:**

How do you differentiate enthalpies of reaction from enthalpies of formation?

What does the chemical equation for the standard heat of formation of CO2 look like? C2H5OH?

Ex. 2) Calculate the standard enthalpy of combustion for the following reaction using the data table below.

ΔH°f, glucose = - 1275.0 kJ mol-1
ΔH°f, CO2  = -393.5 kJ mol-1
ΔH°f, H2O  = -285.8 kJ mol-1

 C6H12O6 (s) + 6 O2 (g) → 6 CO2 (g) + 6 H2O (l)

* 1. Show the reaction that represents the standard heat of formation of glucose.

Ex. 3) Calculate the standard enthalpy of formation for glucose, given the following values:

ΔH°comb, glucose  = -2800.8 kJ/mol
ΔH°f, CO2  = -393.5 kJ/mol
ΔH°f, H2O  = -285.8 kJ/mol

*Homework #1: Energy and States of Matter*

**Use your notes and vocabulary to answer the following.**

1. Describe the motion of particles in gases based on the kinetic theory of matter.
2. The process by which wet laundry dries on an outdoor clothesline in winter is called what? Explain how this process is different than vaporization.
3. Is melting an endothermic or exothermic process? Explain.
4. What information can you infer from a phase diagram?
5. What is the triple point of a substance?
6. What does the word normal refer to when used in a phase diagram?

***Use the phase diagram below to answer the questions that follow.***

1. Label the phases that exist for the substance on the diagram.

**Pressure (kPa)**

 (atm((

120

100

80

60

40

20

0

-30 -10 10 30 50 70 90

 **Temperature oC**

1. Estimate the normal boiling point:
2. Estimate the normal melting point:
3. Estimate the conditions necessary for all three phases to

 coexist. What is the name for this special condition?

1. What phase is the substance at 50 kPa and 30 oC?
2. What phase is the substance at 50 kPa and 80 oC?
3. What phase is the substance below 18 kPa?
4. Will the solid phase of the substance sink or float in

 the liquid phase? How can you tell?

# *Homework #2 Energy Change in one phase physical systems*

1. How much heat is required to raise the temperature of 40.0 g of aluminum from 50.0°C to 90.0°C?
2. How much heat is required to raise the temperature of 1.248 kg of gold from 22.0°C to 45.0°C?
3. How much heat is required to raise the temperature of 100.0 mL of water from 42.0°C to 44.0°C?
4. Calculate the specific heat for air if 1500.0 Joules are required to raise he temperature of 50.0 g from 0.0°C to 40.0°C.
5. Calculate the specific heat for sea water if 13,650 Joules are required to raise the temperature of 500.0 mL from 28.0°C to 35.0°C.
6. 2660 Joules are needed to raise the temperature of 50.0 g of asbestos from 80.0°C to 100.0°C. What is the specific heat of asbestos?
7. Calculate the specific heat for a substance if 58.0 joules are required to raise the temperature of 10.0 g from 25.0°C 40.0°C. What is the substance?
8. 100.0 g of water was heated in a calorimeter and 1674.5 Joules were gained by the water. If the initial temperature was 18.0°C, what was the final temperature of the water?
9. Find the mass of nickel present if 575 Joules raises the temperature from 19.5°C to 32.4°C.
10. What was the initial temperature of a piece of gold that has a mass of 75.0 g and absorbed 1275 Joules to have a final temperature of 65.0°C.

*Homework #3: Energy and Phase Changes.*

 1. How much heat is needed to boil 542 mg of ethyl alcohol?

 2. How much heat is released when 100 grams of steam is condensed to water?

 3. How much heat is released when 100 grams of water is frozen?

 4. What is the latent heat of fusion for gold if 5483 J of heat are necessary to melt 0.085 kg?

 5. How much heat energy is required to melt 85.0 g of silver metal?

 6. What mass of copper can be melted by the addition of 2500 J of heat energy?

 7. What mass of ethyl alcohol can be boiled if 3500 J of heat energy is added?

 8. How much energy would be released when 675 g of steam is condensed to water?

 9. A mystery metal requires 10643 J to melt 165 g. Is the metal copper, gold, or silver?

10. Which is more effective at cooling a drink: ten grams of water at 0°C or ten grams of ice at 0°C?

*Homework #4: Heating Curve Calculations*



1. Calculate the amount of heat required to change 20.0 g of steam at 130ºC to –40.0ºC. Is this endothermic or exothermic?

BP = 100ºC Csolid = 2.06 J/gºC

FP = 0ºC Cliquid = 4.18 J/gºC

Hf = 334 J/g Cvapor = 2.02 J/gºC

Hv = 2260 J/g

1. Calculate the amount of heat required to convert 45.0 g of ethyl alcohol at 90.0ºC to -125ºC. Is this endothermic or exothermic?

BP = 78.5ºC Csolid = 1.05 J/gºC

FP = -115ºC Cliquid = 2.46 J/gºC

Hf = 5.98 J/g Cvapor = 0.954 J/gºC

Hv = 48.8 J/g

1. How much heat is required to convert a 75.0 g block of silver from 100ºC to 3000ºC? Is this endothermic or exothermic?

BP = 2212ºC Csolid = 0.187 J/gºC

FP = 961ºC Cliquid = 0.233 J/gºC

Hf = 88 J/g Cvapor = 0.192 J/gºC

Hv = 2300 J/g

1. How much heat is required to convert 65.0g of diethyl alcohol at 50.0ºC to -150ºC? Is this endothermic or exothermic?

BP = 34.6ºC Csolid = 2.20 J/gºC

FP = -116.3ºC Cliquid = 3.74 J/gºC

Hf = 5.50 J/g Cvapor = 2.35 J/gºC

Hv = 20.6 J/g

*Homework #5: Heating Curve Data*

**More Heating Curve Calculations**

|  |
| --- |
| Thermodynamic Data for some Imaginary Substances |
| Substance | Csolid  J/gºC | MP ºC | Hf J/g | Cliquid J/gºC | BP ºC | Hv J/g | Cvapor  J/gºC |
| Bacolit | 0.829 | 9.25 | 472 | 2.56 | 114 | 1897 | 0.798 |
| Fizzlefub | 0.103 | -73.4 | 16.0 | 0.489 | 47.8 | 146 | 0.154 |
| Glunk | 0.078 | 356 | 25.0 | 0.086 | 2587 | 567 | 0.054 |
| Soofie | 1.56 | 24.7 | 216 | 3.87 | 172 | 1249 | 1.32 |
| Jezger | 3.05 | 176 | 739 | 6.28 | 417 | 1723 | 2.91 |

***Calculate the total amount of heat needed for each of the following changes. Pay very close attention to the numbers in the questions. Draw your own heating curve for each question, being sure to label the melting point and boiling point correctly. Also label your start and end points on the diagram that are given in each question. You may not need all five steps!!!***

1. Calculate the total amount of heat needed to raise 25.0g of bacolit from a solid at 5.00ºC to a liquid at 100ºC
2. How much heat will be released when 75.3 g of fizzlefub vapor at 75.0ºC is condensed to a liquid and cooled to –73.4ºC?
3. How much energy is released when 115g of glunk vapor at 3000ºC is cooled to a solid at 250ºC?
4. How many joules would you have to add to 57.6g of solid soofie at –3.50ºC to change it all to a vapor at 255ºC?
5. How many joules would be released when 125g of liquid jezger at 315ºC is cooled until it becomes a solid at room temperature (22.0ºC)?

*Homework #6: Thermochemical Stoichiometry*

***Show all work as directed in the space below. Don’t forget units and sig figs.***

1. How much heat is emitted when 10.0 g of Al is reacted with excess iron (III) oxide?

 2Al + Fe2O3 → Al2O3 + 2Fe + 850. kJ

2. How much heat will be released in the formation of 193 g of ammonium bromide from ammonia and hydrogen bromide? NH3 + HBr → NH4Br + 188.32 kJ

3. How much heat will be produced in the displacement of 0.0663 g of bromine from sodium bromide by chlorine? Cl2 + 2NaBr → 2NaCl + Br2 + 100.18 kJ

4. What mass of CrO3 reacted if the following reaction produced 6.18 kJ of heat?

 CrO3 + H2O → H2CrO4 + 5.4 kJ

5. What volume of oxygen at STP will be produced if the following reaction absorbed 275 kJ of heat?

 2SO3 + 198 kJ → 2SO2 + O2

6. What amount of heat will be produced when 85.0 g of chlorine react with excess tin in the following reaction?

 Sn + 2Cl2 → SnCl4 + 511 kJ

7. What amount of heat will be absorbed when 25.0 g of sodium bicarbonate is decomposed according to the following reaction? 2NaHCO3 +129 kJ → Na2CO3 + CO2 + H2O

8. What mass of iron (III) oxide will be produced if 955 kJ of energy are released in the following reaction?

 4FeS + 7O2 → 2Fe2O3 + 4SO2 + 1352 kJ

9. What mass of iron will be produced if 1265 kJ are absorbed according to the following reaction?

 Fe3O4 + 4H2 + 149.8 kJ → 3Fe + 4H2O

10. How much heat is absorbed when 67.4 g of nitrogen dioxide is formed from its elements?

 N2 + 2O2 + 68 kJ→ 2NO2

*Problem Set #7: Hess’ Law (Big Mamma)*

1. Complete combustion of 1.00 mol of acetone (C3H6O) liberates 1790 kJ:

C3H6O (l) + 4 O2 (g) → 3 CO2 (g) + 3 H2O (l); ΔH°comb, acetone = -1790 kJ

Using this information together with the data below (values in kJ/mol), calculate the enthalpy of formation of acetone. ΔH°f, O2: 0; ΔH°f, CO2: -393.5; ΔH°f, H2O: -285.83

|  |  |
| --- | --- |
| C6H14(l) | -4163.0 |
| CO2(g) | -393.5 |
| H20(g) | -285.8 |

1. Calculate the standard enthalpy of combustion of hexane using the enthalpies of formation (in kJ/mol) given to the right. You must write a balanced equation first. There will be a fraction as a coefficient and that is okay. Hexane (C6H14) must have a 1 for a coefficient.

1. Ammonia reacts with oxygen to form nitrogen dioxide and steam, as follows:

4 NH3 (g) + 7 O2 (g) → 4 NO2 (g) + 6H2O (l)

Given the following standard enthalpies of formation (kJ/mol) calculate the heat of the reaction.

ΔH°f NH3(g): -45.90 ΔH°f NO2(g): +33.1; ΔH°f H2O(l):-241.83

1. The standard enthalpy change, ΔHrxn°, for the thermal decomposition of silver nitrate according to the following equation is +78.67 kJ/mol:

AgNO3(s) → AgNO2(s) + (1/2)O2(g)

The standard enthalpy of formation of AgNO3(s) is -123.02 kJ/mol. Calculate the standard enthalpy of formation of AgNO2.

|  |  |
| --- | --- |
|  | ΔHfkJ/mol |
| C2H6(g) | -84.68 |
| O2 (g) | zero |
| CO2 (g) | -393.5 |
| H2O (g) | -241.8 |

1. Using standard enthalpies of formation, calculate the heat of combustion per mole of gaseous water formed during the complete combustion of ethane gas.
2. The ΔHrxn for the following reaction equals - 89 kJ/mol:

IF7 + I2 → IF5 + 2 IF

In addition, two standard enthalpies of formation are known: IF7 = -941 kJ/mol; IF5 = -840 kJ/mol

Determine the ΔHf° for IF.

1. What is the enthalpy change for the following reaction?

SiCl4(l) + 2H2(g) → Si(s) + 4HCl(g)

Use the following standard enthalpies of formation: SiCl4(l): -687 kJ mol¯1 ; HCl(g): -92 kJ mol¯1

*Review: Multiple Choice*

1. Which of the following is true when a sample of water is heated gently?

1. The average velocity of the particles increases.
2. The temperature increases.
3. The total neat energy of the sample increases.
4. All of the answers are correct.

2. If the same amount of energy is absorbed by equal masses of the following materials, which one will have the least temperature change? (Specific heats are given in parenthesis.)

 a. Ca (0.653) b. Mg (0.250) c. Fe (0.451) d. Zn (0.380)

3. How many grams of ice at 0°C would be melted if the ice absorbed 4.50 kJ of heat?

 a. 0.0742 b. 0.135 c. 74.2 d. 13.5

4. The heat of fusion is

1. the energy required to boil a liquid into a vapor with an increase in temperature.
2. the energy required to melt a solid into a liquid with an increase in temperature.
3. the energy required to boil a liquid into a vapor at constant temperature.
4. the energy required to melt a solid into a liquid at constant temperature.

5. When the temperature of a sample increases, this indicates that the sample.

1. no energy change is occurring, but the molecules are slowing down in
2. no energy change is occurring, but the molecules are speeding up in
3. cold is leaving
4. heat energy is entering.

6. A 75 mL sample of freezing water and a 150 mL sample of freezing water will have the same
 a. mass b. temperature c. heat and temperature d. heat

7. The energy required to change on gram of a substance from a liquid to a vapor is the
 a. specific heat b. heat of fusion c. heat of vaporization d. heat energy

1. What is the specific heat of nickel if 475 g requires 7545 J to increase the temperature from 50.3°C to 86.4°C?

 a. 0.00365 J/g°C b. 0.316 J/g°C c. 0.183 J/g°C d. 0.440 J/g°C

1. What amount of heat energy is needed to change 35 g of ice at -25°C to steam at 115°C?

a. 20482 J b. 108284 J c. 111273 J d. 108270 J

1. How many joules are released when 275 g of a vapor are condensed into a liquid? (Heat of fusion = 0.87 J/g, Heat of vaporization = 1.43 J/g)

 a. 239 J b. 316 J c. 192 J d. 393 J

11. Which of the following is a factor in determining the amount of heat transferred to an object?

a. mass of the object b. temp of the object

c. specific heat of the object d. all of the above

12. Boiling point is the same as the

a. melting point b. freezing point

c. condensation point d. sublimation point

13. Melting point is the same as the

a. sublimation point b. freezing point

c. condensation point d. boiling point

14. When a hot object is placed in contact with a cooler object, which of the following occurs?

1. The hot object gains heat energy and the cooler object loses heat energy
2. The hot objects loses heat energy and the cooler object gains heat energy.
3. Both objects lose heat energy.
4. Both objects gain heat energy.

15. What amount of heat is needed to completely boil 45.0 g of ethyl alcohol to a vapor at its boiling point? The heat of fusion is 105 J/g and the heat of vaporization is 853 J/g

a. 4725 J b. 2.33 J c. 38385 J d. 19.0 J

16. Water is sometimes used as a heat storage medium because it requires a great deal of heat to change the temperature of water. This is best explained by

a. magic b. its high heat of fusion

c. its high heat of vaporization d. its high specific heat

17. Which of the following states of matter can be easily compressed?

 a. solids b. liquids c. gases d. all states

18. During boiling, the temperature of a liquid

 a. remains constant b. increases

 c. decreases d. approaches the standard boiling point

19. When the molecules at the surface of a liquid below its boiling point have enough energy to overcome the intermolecular forces and escape we call this

 a. evaporation b. boiling c. sublimation d. condensing

20. Ice cubes get smaller and smaller if they are left in the freezer for a long period of time. This is an example of

 a. freezing b. melting c. sublimation d. deposition

***Solve the following problems. Show all work!***

1. What is the mass of water present if 2500 J were released when the temperature decreased 35°C?
2. How much heat is needed to melt 65.6 g of lead? (Hf for lead = 24.7 J/g)
3. What amount of heat will be absorbed according to the following equation when 25.0 g of calcium oxide reacts with excess water? CaO + H2O + 65.2 kJ 🡪 Ca(OH)2
4. What is the change in temperature if 1325 J of heat is added to 55 g of water?
5. How much heat is needed to change 65 g of ice at -13°C to steam at 143°C?

26. How much heat is released by cooling 325 g of water from 45°C to -15°C?

Name:\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Period:\_\_\_\_\_\_\_\_\_

Unit 11: Thermochemistry. Big Ideas: The energy changes of physical and chemical processes can be depicted graphically and quantitatively evaluated to yield important information about heat flow.

|  |  |  |
| --- | --- | --- |
|  |  | Assignment Name and Date |
| K | Learning ObjectivesLearning objectives are scored as minor and major grades. You will have ***multiple opportunities*** during the unit to demonstrate mastery of each learning objective. |  |  |  |  |  |  |  |  |  |  | **average** |
| K1 | I can describe and use the important aspects of a phase diagram (triple point, critical point, normal boiling, freezing point and phase boundaries, slope) to evaluate physical changes under specified conditions. I can describe a physical change as endothermic or exothermic. |  |  |  |  |  |  |  |  |  |  |  |
| K2 | I can calculate heat, mass, specific heat and temperature for individual phases (q  = mC∆T) AND I can calculate heat change associated with phase changes (q = mHf,v). I can interpret heating curve diagrams. |  |  |  |  |  |  |  |  |  |  |  |
| K3 | I can describe, draw, and use an Energy Coordinate diagram; I can describe a chemical process as endothermic or exothermic. |  |  |  |  |  |  |  |  |  |  |  |
| K4 | I can use mole-energy ratios from balanced thermochemical equations to calculate heats of reaction as well as mass amounts of products and reactants (stoichiometry) |  |  |  |  |  |  |  |  |  |  |  |
| K5 | I can use Hess’s law to calculate the heat of reaction or heat of formation and describe a reaction as endothermic or exothermic. |  |  |  |  |  |  |  |  |  |  |  |
| I1 | I can convert between mass, particles, volume of gases and moles. |  |  |  |  |  |  |  |  |  |  |  |
| A0 | I can round answer to the correct number of significant figures; I know how to use the calculator correctly; I can record the answer with appropriate units. |  |  |  |  |  |  |  |  |  |  |  |
| A1 | I can perform appropriate unit/metric conversions |  |  |  |  |  |  |  |  |  |  |  |
| 00 | I did my homework like I should! |  |  |  |  |  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |  |  |  |  |  |  |
|  | **Example Score** | 1 | 2 | 3 |  |  |  |  |  |  |  |  |
|  | A ‘1’ is a 50, a ‘2’ is two 75’s and a ‘3’ is three 100’s | 50 | 75+75 | 100+100+100 |  |  |  |  |  |  |  | 83 |
|  | Average = (50 + 150 + 300) / 6 = 83 |  |  |  |  |  |  |  |  |