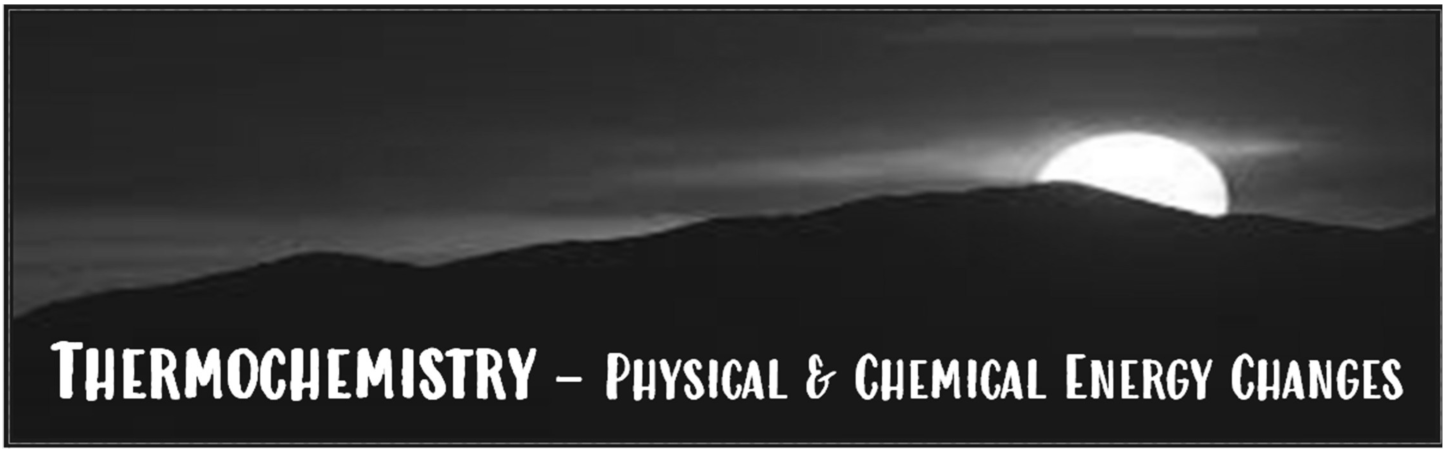
name



Heat	$q = mC\Delta T$ $q = mH_f$ $q = mH_v$ q = heat m = mass C = specific heat capacity ΔT = change in temperature H_f = heat of fusion H_v = heat of vaporization

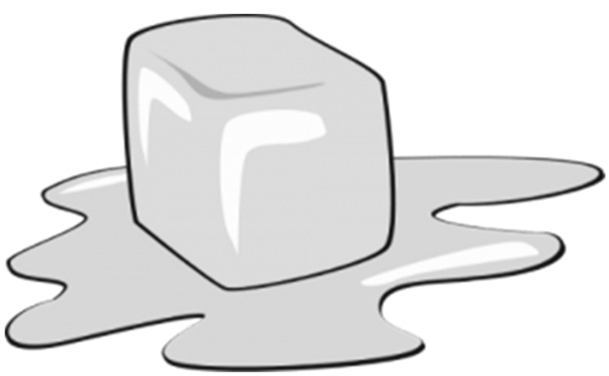


Table B
Physical Constants for Water

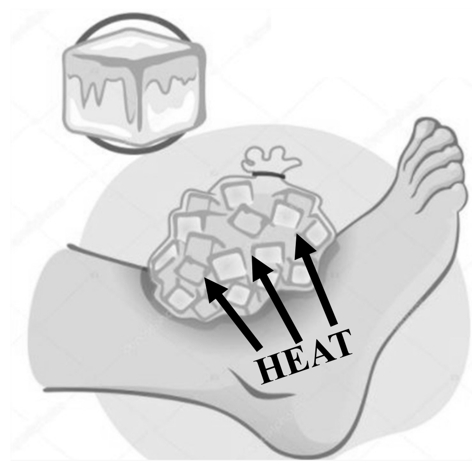
Heat of Fusion	334 J/g
Heat of Vaporization	2260 J/g
Specific Heat Capacity of $H_2O(\ell)$	4.18 J/g•K

Thermochem Basics

Thermochemistry is the part of our course that connects chemistry with the changes of heat energy, in physical processes like phase changes, and in chemical reactions that are exothermic or endothermic. It's about how much energy is needed to melt an ice cube, or warm up some water — or cool it down in the fridge, or to vaporize it into steam, or HOW exothermic or how endothermic a chemical reaction is. We will be able to convert this energy in several units.

In the best movie of all time, the Wizard of Oz, Dorothy is confused when the Yellow Brick Road splits into two paths. The Scarecrow tells Dorothy that "some people go this way", then he says that "some people do go the other way", and finally he says, "of course, some people do go both ways".

Thermochemistry is exactly that: it takes the same amount of energy to melt one gram of ice as it does to freeze one gram of water into ice. It's only a matter of either adding this energy to the ice, or removing it from the water to make this phase change occur. To warm up one gram of water by one degree the water needs to absorb a specific amount of energy. To cool water in the fridge, you must remove the same amount of energy per gram to cool it by one degree.



Thermochem is a 2 way process… Adding energy makes stuff hotter, removing the same amount of energy makes them colder, by the same amount.

It's the energy that is "something", while the cold is nothing. Cold is just the lack of heat energy. Heat energy can move, the cold doesn't.

When you get into a hot sauna, you feel hot because the heat in the air moves into your body. If you go to the mailbox in your PJ's in the winter, you would say that you got cold, but scientifically speaking, you got less hot, heat left your body. Heat (energy) moves.

If you fall and twist your ankle, you should ICE your ankle.
Your ankle gets colder, which reduces the swelling. You ankle does not absorb this cold from the ice, rather the heat from your ankle moves into the ice. Your ankle really gets less hot, but we don't talk like that. You say your ankle is cold but it is because heat is leaving your body.

Thermochemistry will allow us to calculate how much heat is released in a chemical reaction, or how much is absorbed. We will use several units, which can be converted to other units. Most of the units you never heard of, so let's get the funky names out now, so you can start relaxing about them.

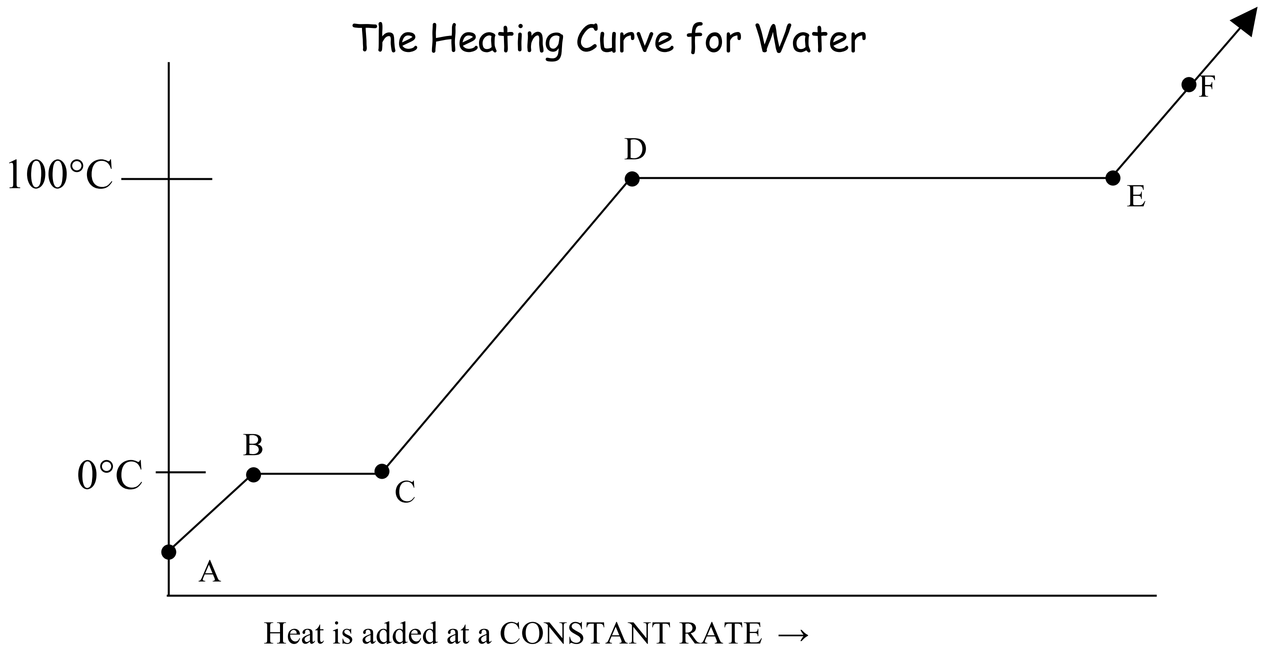
We will measure energy in joules (named after a chemist). There also units called kilo-joules (1000 joules). We will also use "calories", which are scientific (and lower case), and "Calories" (capital "C") that we eat as FOOD CALORIES or kilocalories.

Two units with the same name (ugh!). To keep track we will use "cal" for the smaller scientific calories, and the word "Calories" and use a capital "C" for food Calories. Put these equalities under table B now:

4.18 Joules = 1 cal   1000 cal = 1 Calorie   1000 Joules = 1 kilojoule

Heating Curves

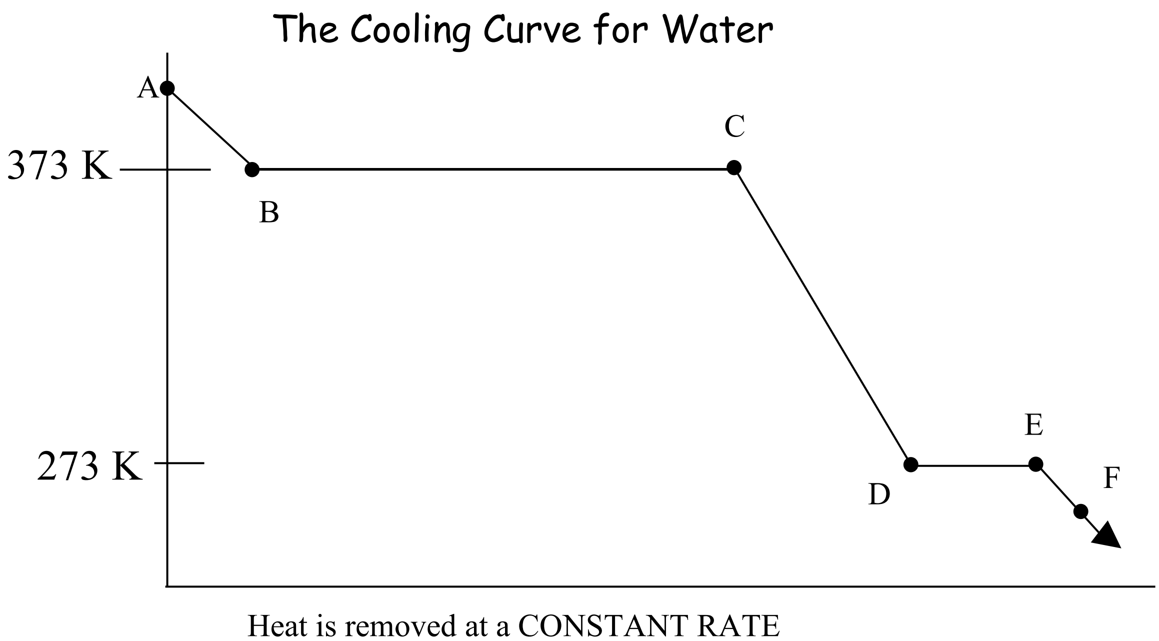
A heating curve shows how a substance reacts to different amounts of energy. We will look at the heating curve for water because you know what to expect. On all heating curves the temperature starts BELOW the freezing point, and goes well above the boiling point. Look at the graph and follow the boxes below, make sure you "get" what happens at each segment of the heating curve graph.



Segment	Phase or phases	What's going on in this segment?
AB	Solid only	Point A is colder than the freezing point (but above absolute zero). As heat is added, the ice warms up. Point B is the hottest ice can be and still hold solid.
BC	Solid + Liquid	Ice melts, note that the temperature is constant. All heat energy added is being used to shake the solid apart into a liquid, no increase in temperature.
CD	Liquid only	As heat is added the temperature rises, all the way to the hottest that a liquid can be before it blows apart into steam (at the boiling point)
DE	Liquid + Gas	Water vaporizes at the boiling point, note that the temperature is constant. All heat energy being added is shaking the water molecules apart into a gas.
EF	Gas only	Steam can be hotter than the boiling point, and you can heat it up until at some point the energy is so high it begins to decompose into H_2 and O_2. (over 3000°C)
Melting	In order to melt, ice absorbs a certain amount of energy per gram (if it's at 0°C) This amount of energy is called the HEAT OF FUSION, symbolically it is H_F For water the heat of fusion constant is written this way: $H_F = 334$ J/g Joules are small amounts of energy, g is per gram.	
Boiling	In order to boil, hot water absorbs a certain amount of energy per gram (if it's at 100°C) This amount of energy is called the HEAT OF VAPORIZATION, symbolically it is H_V For water, the heat of vaporization is written as $H_V = 2260$ J/g	

Cooling Curves

A cooling curve is the exact opposite of a heating curve, and the temperatures are the same. They start at a very hot temperature, above the condensing point (boiling point) and heat is removed at a constant rate, until the temperature of the substance is so cold, it freezes into a solid, but never reaches absolute zero. This graph is in Kelvin, but the same temperaures for condensing/boiling, and for freezing/melting exist for water, they are constants in any unit.



Segment	Phase or phases	What's going on in this segment?
AB	Gas only	Point A hot, above the condensing point, and as heat energy is removed, the steam is cooling but still a gas. It's too hot yet to condense.
BC	Gas Liquid	Steam condense into hot water, note that the temperature is constant. All of the heat energy that is emitted is causing condensing. It is "un-vaporizing"
CD	Liquid only	As heat is removed the temperature decreases, all the way to the coldest a liquid can be before it freezes into a solid (at the freezing point)
DE	Liquid + Solid	Water freezes at the freezing point, note that the temperature is constant. All heat energy being removed allows the water molecules to lock together.
EF	Solid only	Solids can be as cold as you can make them (above 0 Kelvin).
Condensing	In order to condense, steam emits a certain amount of energy per gram (if it's at 100°C) This amount of energy is called the HEAT OF VAPORIZATION, symbolically it is H_V For water, the heat of vaporization is written as $H_V = 2260$ J/g	
Freezing	In order to freeze, water emits a certain amount of energy per gram (if it's at 0°C) This amount of energy is called the HEAT OF FUSION, symbolically it is H_F For water the heat of fusion constant is written this way: $H_F = 334$ J/g Joules are small amounts of energy, g is per gram.	

Warming or Cooling of Water

You can change the temperature of a liquid (hotter or colder) in between the melting point and the boiling point, let's only discuss water because you understand that already. You can heat up some water on the stove to make it warmer, or you can put some tap water into your fridge to make it cooler. If you add energy, the water temperature increases, if you remove energy, the water cools.

Each gram of water requires 4.18 Joules of energy to change temperature by 1°C or by 1 Kelvin. If you add 4.18 joules to a gram of water it gets hotter by 1°C or by 1 Kelvin. If you remove 4.18 joules to a gram of water it gets colder by 1°C or by 1 Kelvin.

This constant is called the specific heat capacity constant, symbolically it is "C". For water, the specific heat capacity constant is C = 4.18 Joules/gram x Kelvin or C = 4.18 J/g·K

This allows us to use any number of grams and any sized temperature change as long as the liquid stays between the melting and boiling point.

Every substance has a different specific heat capacity constant. Water has an unusually high "C" value.	
Substance	Specific heat capacity constant "C"
Liquid water	4.18 J/g·K
Steam	1.90 J/g·K
Ice	2.10 J/g·K
Copper	0.391 J/g·K
Iron	0.45 J/g·K
Mercury	0.140 J/g·K

All substances have constants, some you are used to, like density, or boiling point or melting point. Every substances has a heat of fusion constant, a heat of vaporization constant, and a specific heat capacity constant. Only one formula works for any substance that is melting or freezing; One formula works for any substance that is boiling or condensing, and one formula works for warming up or cooling down any substance.		
What's happening?	Constant to use in your math problem	For water these values are...
Melting or freezing	Heat of Fusion constant H_F	334 J/g
Boiling or condensing	Heat of Vaporization H_V	2260 J/g
Changing temperatures	Specific Heat Capacity constant C	4.18 J/g·K

Thermochemistry Math

q = heat in joules m = mass in grams ΔT is the "delta" T, the change in temperature

	Formula	Math and answer
How much energy is needed to melt an ice cube of 48.6 grams into liquid water at the melting point? (no ΔT)	$q = mH_F$	(48.6 g)(334 J/g) = 16,232.4 Joules = 16,200 J 3 SF
How much energy is needed to freeze 48.6 grams of liquid water at the melting point into ice? (no ΔT)	$q = mH_F$	(48.6 g)(334 J/g) = 16,232.4 Joules = 16,200 J 3 SF
How much energy is needed to vaporize 2.29 grams of water (at the boiling point) into steam? (no ΔT)	$q = mH_V$	(2.29 g)(2260 J/g) = 5175.4 Joules = 5180 J 3 SF
How much energy is released when 2.29 grams of steam condenses (at condensing point) into water? (no ΔT)	$q = mH_V$	(2.29 g)(2260 J/g) = 5175.4 Joules = 5180 J 3 SF
SPECIAL NOTE →	With temperature changes, the ΔT Kelvin = the ΔT Centigrade There are 100 units of temperature from melting to boiling in both scales, (0 to 100 in centigrade, and 273 to 373 in Kelvin). ΔT K = the ΔT°C	
How much energy must be removed to cool a bottle of water (500 mL) from room temperature of 24.0 C to cool at 17.5 C?	$q = mC\Delta T$	(500.0 g)(4.18 J/g·K)(6.50 K) = 13585 J = 13600 J 3 SF
How much energy must be added to warm a bottle of water (500 mL) from 24.0 C to cool at 30.5 C?	$q = mC\Delta T$	(500.0 g)(4.18 J/g·K)(6.50 K) = 13585 J = 13600 J 3 SF

Thermochemistry is "the same" in both directions. It takes the same amount of energy in either direction,

the difference is to make things melt or get hotter or boil you must add energy (endothermic)

To make thigs freeze, get cooler, or condense you must remove energy (exothermic)

To convert joules into other units will require a few equalities.	
1000 cal = 1 Calorie (UPPER CASE C)	4.18 Joules = 1 cal (lower case c)
1000 Joules = 1 kilo-Joule	*1 Calorie = 1 Food Calorie* 1 Calorie = 1 kilo-calorie

Convert 16,200 Joules into kilo joules.

$$\frac{16{,}200\ \cancel{\text{Joules}}}{1} \quad \text{X} \quad \frac{1\ \text{kilo-Joule}}{1000\ \cancel{\text{Joules}}} \quad = \quad 16.2\ \text{kilo-Joules} \qquad \textit{(3 SF)}$$

It is important to point out that we both know that you do not really grasp the amount of energy that a Joule represents. It is okay that you are confused a bit, you're normal. Let it flow through you, you will "get" these units as we progress through this unit. For now, just let yourself do this math, do it properly, and trust that the units are going to work themselves out. Breathe, do not get worried.

Convert these 16200 Joules into scientific calories.

$$\frac{16{,}200\ \cancel{\text{Joules}}}{1} \quad \text{X} \quad \frac{1\ \text{cal}}{4.18\ \cancel{\text{Joules}}} \quad = \quad 3880\ \text{cal} \qquad \textit{(3 SF)}$$

Convert these 3880 cal into Calories (AKA FOOD calories or kilocalories)

$$\frac{3880\ \cancel{\text{cal}}}{1} \quad \text{X} \quad \frac{1\ C}{1000\ \cancel{\text{cal}}} \quad = \quad 3.88\ \text{Calories or } 3.88\ \text{kCal} \qquad \textit{(3 SF)}$$

Food Energy

Food you eat is energy. You use this food (energy) to stay alive, and to move. The more food you eat, the more energy you eat. If you don't use up this energy, your body stores it as fat for another day. Eating too much energy makes you store this energy. If you are not eating enough for a short time, this stored energy gets used up, and you lose some weight as your fat is converted back into usable energy.

If you stop eating for good, you will slow down from the lack of energy until you run out of energy and you stop living. Food is energy, energy is required to stay alive. One must maintain a balance between how much energy you take in and how much you use to stay the same mass. Out of balance means you begin to gain mass or lose mass.

The food you eat is measured in Calories. The FOOD CALORIE is a kilo-calorie, or 1000 cal. That amount of energy also has an equivalent in Joules, and kilo-Joules. All energy units can be converted back and forth with the equalities given on the previous page.

Measuring energy in food can't be done directly. The way scientists have devised a way to measure energy in food is to burn it up in a machine called a calorimeter (calorie—meter). The food burning creates heat, which goes into a very well measured amount of water, mass and starting temperature.

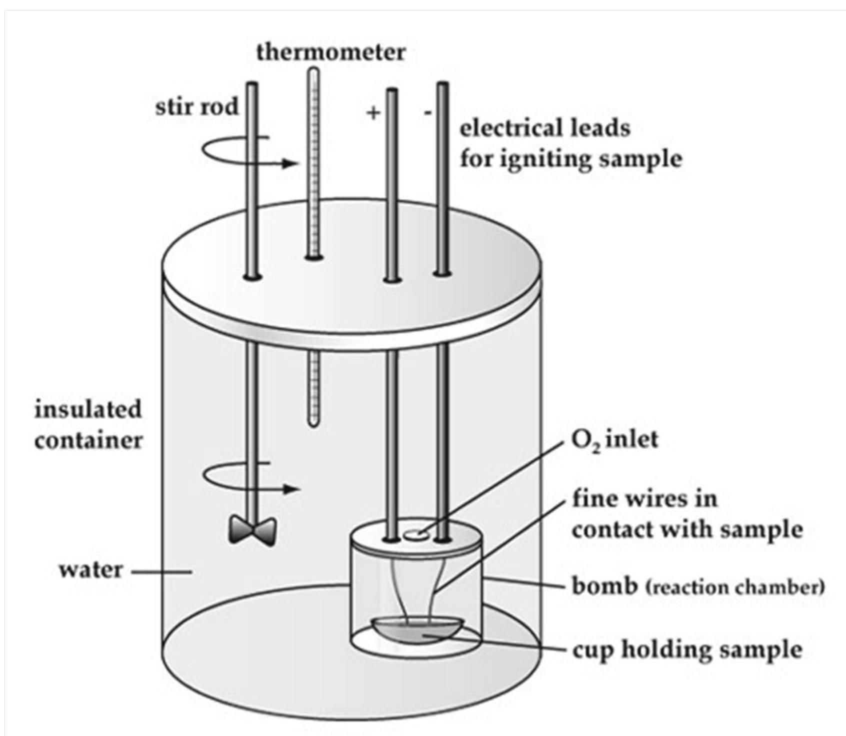
This heat energy makes the water hotter. Since we have this amazing formula: $q = mC\Delta T$ we can measure the mass of the water, and the starting and ending temperature of the water, and use the "C" value for water, to determine how much energy it took to make this temperature change happen.

Since it was the food burning up that caused this change, the amount of energy gained by the water is the amount of energy that came out of the food.

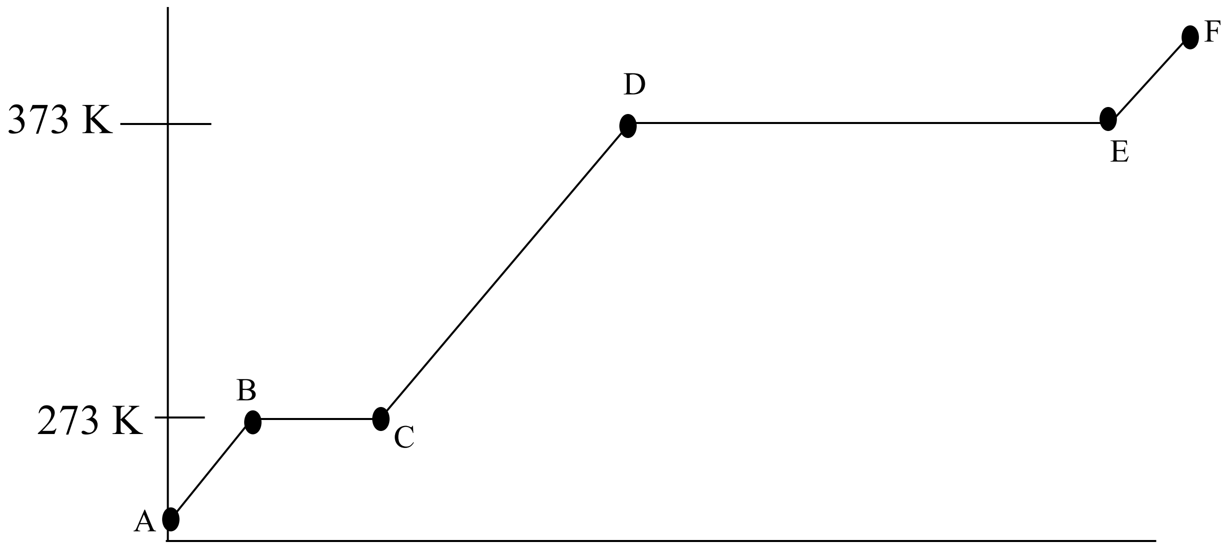
BOMB CALORIMETER is the device to measure energy in food. The calorimeter is filled with pure water and the mass is very carefully measured (the "m") . A great thermometer can measure the exact temperature before and after the food is burned (the ΔT). Pure water has a specific heat constant (4.18 $J/g{\cdot}K$).

A measured sample of food goes into the "BOMB" and that is pumped up with oxygen to insure complete combustion, releasing all of the energy stored in the food. The food is sparked by electric charge, setting it on fire, releasing the energy.

Once the q value (energy in joules) is calculated, it gets converted to cal, and then to Calories.



Let's look over this heating curve for water again. See how the formulas connect with the graph. Use the guide below.



Segment	Phase or phases present	Formula	Constant	Temp	KE	PE
AB	Solid only	$q = mc\Delta T$	$C = 2.10$ J/g·K	Inc	Inc	steady
BC	Solid → Liquid	$q = mH_F$	$H_F = 334$ J/g	Steady	Steady	Inc
CD	Liquid only	$q = mc\Delta T$	$C = 4.18$ J/g·K	Inc	Inc	steady
DE	Liquid → Gas	$q = mH_V$	$H_V = 334$ J/g	Steady	Steady	Inc
EF	Gas only	$q = mc\Delta T$	$C = 1.90$ J/g·K	Inc	Inc	steady

SPECIAL NOTE HERE: the "C" value in table B says

Specific Heat Capacity of $H_2O_{(\ell)}$ = 4.18 J/g·K   that is a "script L for liquid!

Liquid water takes 4.18 J of energy per gram to change temp (hotter or colder) by 1°C = 1 K.

Ice and steam are chemically identical (all are H_2O) but physically different.
It takes different amounts of energy to cool or warm ice and steam than it does for water.

$C_{STEAM} = 1.90$ J/g·K $C_{ICE} = 1.90$ J/g·K

In real life things often get “complicated”. For instance, you might melt an ice cube starting at the freezing point, up to body temperature, with your mouth. Before you can warm the water you first have to melt it. That’s 2 thermochem math problems, summed up to one amount of energy in joules.

Step 1: energy is needed to melt the ice into liquid. Step 2: warm up this cold water to body temperature.

Or you might put water into a pot to make macaroni and cheese, and some of that water will vaporize away at the boiling point. To measure the energy required, first calculate the energy required to warm up ALL of the water up to 373 Kelvin, then measure how much energy to boil away just PART of that water.

Finally, you might even have multistep problems, and you SUM UP THE JOULES for the total energy.

Imagine this: How much energy does it take to warm up really cold ice (-11°C) into hot steam (115°C)

1. warm the ice to the melting point $q = mc\Delta T$ (ΔT here is 11 K)
2. melting it at the melting point (at melting point, no ΔT)
3. warming up the water to the boiling point (ΔT here is 100 K)
4. vaporizing it (at boiling point, no ΔT)
5. then heating up the steam to the final hottest temperature. (ΔT here is 15 K)

A five step problem is the longest thermochem problem possible.
You have to use the right formulas, and the right constants at the right time.

Practice problems…

657 grams of ice at 0°C is warmed to body temperature when you sit on a big ice cube until is warms up to your skin temperature 37.0°C. How much energy is required to do this? *(reminder: ΔT °C = ΔT K)*		
Melt the ice this way:	$q = mH_F$	= (657 g)(334 J/g) = 219,000 Joules
Warm the water this way:	$q = mc\Delta T$	= (657 g)(4.18 J/g·K)(37.0 K) = 102,000 Joules
Sum up the total energy:		219,000 J + 102,000 J = 321,000 Joules 3 SF
When 13.0 grams of steam condenses on your finger (it’s bad luck when that happens, hence the thirteen!) and then the hot water cools to body temperature of 37.0°C, how much energy is absorbed by your body? (assume no energy is lost to the air)		
condensation	$q = mH_V$	= (13.0 g)(2260 J/g) = 29,400 Joules
Cool the water down	$q = mc\Delta T$	= (13.0 g)(4.18 J/g·K)(63.0 K) = 3420 Joules
Sum up the total energy		29,400 Joules + 3420 Joules = 32,800 Joules

If 25.0 grams of ice at 273 Kelvin, and you vaporize it into steam, how much energy is required to do this?		
Melt the ice this way:	$q = mH_F$	$= (25.0\ g)(334\ J/g) = 8350$ Joules
Warm the water this way:	$q = mc\Delta T$	$= (25.0\ g)(4.18\ J/g{\cdot}K)(100\ K) = 10{,}450$ Joules
Vaporize water into steam	$q = mH_V$	$= (25.0\ g)(2260\ J/g) = 56{,}500$ Joules
Sum up the total energy:		8350 J + 10,450 J + 56,500 J = 75,300 Joules

We can solve for q, m, C, ΔT, H_F, or H_V given other information.
Put the data in the proper place in the proper formula, then DO the math correctly, watch units and SF.

1. What mass of ice can be melted with 7,543 Joules of heat?
   $q = mH_F$ 7543 J = (m)(334 J/g) m = 22.58 g (4 SF) 334 J/g has unlimited SF

2. Calculate the heat of fusion constant for an unknown metal if it takes 85,600 Joules to melt exactly 198.10 grams of this metal.
   $q = mH_F$ 85,600 J = (198.10)(H_F) H_F = 432 J/g (3 SF) SF limited to 3 in the Joules

3. When 7,399 Joules are zapped into 123.4 grams of iron, the metal changes temperature from 265.2 K to 338.4 Kelvin. What is the C constant for iron?
   $q = mC\Delta T$ 7399 J = (123.4 g)(C)(133.2 K) C = 7399 J/16,436.88 g·K
   q = 0.4501 J/g·K 4 SF note: funky units for "C" that match table B, as they should

4. Calculate the temperature change when 12.5 grams of water is heated up with 836.0 Joules.
   $q = mC\Delta T$ 836.0 J = (12.5 g)(4.18 J/g·K)(ΔT) solve for change in temp

$$\Delta T = \frac{836.0\ J}{52.25 J/K} = 16.0\ K$$ 3 SF from mass

5. Calculate the H_V for water if 11,706.8 J are released when 5.18 grams of steam condenses.
   $q = mH_V$ 11,706.8 J = (5.18 g)(H_V) 11,706.8 J/5.18 g = H_V H_V = 2260 J/g

6. What is the final temperature of 355.0 grams of copper at 288 Kelvin when you it absorbs 5,604 Joules of energy. Round your answer here to nearest whole number Kelvin temperature.

   This requires you to realize that you must solve for the ΔT, then ADD that ΔT to the starting temperature.
   The question is WHAT IS THE FINAL TEMP, not what is the ΔT?

   $q = mC\Delta T$ 5604 J = (355.0 g)(0.391 J/g·K)(ΔT)

$$\Delta T = \frac{5604\ J}{138.8\ J/K)}$$ ΔT = 40.37 Kelvin

Start temp = 288 K
Ad in the T = + 40.37 K

FINAL TEMP = 328.37 K = 328 K 3 SF

On the first day of school we synthesized water. Remember how loud that was (and the FIREBALL?). That reminds us of an important chemistry adage: When bonds form, energy is released.	
When the hydrogen and oxygen bonded together, the bonds formed and released energy exothermically.	$2H_2 + O_2 \rightarrow 2H_2O$
The regents exam will give you reactions like this to consider	
When bonds form, energy is released. (exothermic)	$F_2 \rightarrow F + F$
When bonds break, energy is absorbed. (endothermic)	$Cl + Cl \rightarrow Cl_2$
Water freezing (exothermic)	$2H_2O_{(L)} \rightarrow 2H_2O_{(S)}$
Ice melting (endothermic)	$2H_2O_{(S)} \rightarrow 2H_2O_{(L)}$
Steam condensing at the condensing point (exothermic)	$2H_2O_{(G)} \rightarrow 2H_2O_{(L)}$
Water vaporizing at the boiling point (endothermic)	$2H_2O_{(L)} \rightarrow 2H_2O_{(G)}$

Table I

In our reference tables is table I, which shows 25 chemical processes, and the energy associated with each one. The first six are combustion reactions, then there are a batch of synthesis reactions, and the bottom is a bunch of phase changes from solid to aqueous phase as ionic compounds dissolve into water. The last one we'll leave be for a few months.

The first reaction, the combustion of methane is balanced and has a ΔH. This "delta H" stands for the "change in heat". This reaction is exothermic, because –ΔH values are exothermic. The energy is just energy (in kJ). It is NOT negative or positive energy.

Just remember this: energy is like money. You might find a ten dollar bill & have ten extra dollars. You don't have a "positive ten dollars". If you lost a twenty dollar bill, you lost $20, you don't "have –$20". Money can't be negative or positive, and neither can energy. The sign only indicates if the energy is a product in an exothermic reaction, or a reactant in an endothermic reaction.

Take out Table I now, the first reaction is the combustion of methane:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ the $\Delta H = 890.4$ kJ

This means that when one mole of methane combusts, 890.4 kilojoules of energy is released. We can write a *balanced thermochemical reaction*, and make the energy a product this way:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4\ kJ$

Written either way,

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ the $\Delta H = 890.4$ kJ or $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4$ kJ

The mole ratio of this reaction would be 1:2:1:2:890.4 kJ energy is in ratio with the chemistry

The MOLE RATIO will now include the energy, and we can use in Mole Ratio type problems.

If you combust 11.4 moles of propane (C_3H_8) with sufficient oxygen, how many kilo-Joules of energy will be released? (the second balanced thermochemical reaction on table I)

$$\text{MR} \quad \frac{\text{Propane}}{\text{energy}} \quad \frac{1 \text{ mole}}{2219.2 \text{ kJ}} = \frac{11.4 \text{ moles}}{X \text{ kJ}}$$

Solve for X $X = 2219.2 \text{ kJ} \times 11.4$

$X = 25{,}298.88 \text{ kJ} = 25{,}300 \text{ kJ}$ *with 3 SF*

When 2.44 moles of aluminum form into aluminum oxide, how much energy is released?

The balanced thermochemical reaction from table I is $4Al + 3O_2 \rightarrow 2Al_2O_3 + 3351$ kJ

$$\text{MR} \quad \frac{\text{aluminum}}{\text{energy}} \quad \frac{4 \text{ moles}}{3351 \text{ kJ}} = \frac{2.44 \text{ moles}}{X \text{ kJ}}$$

Solve for X $4X = 2219.2 \text{ kJ} \times 11.4$

$X = 8176.44 \text{ kJ} = 8180 \text{ kJ}$ *with 3 SF*

When ammonium chloride dissolves into water table I shows us that this is a phase change from solid → aqueous   It has a ΔH = +14.78 kJ

This can be written properly as: $NH_4Cl_{(S)} + 14.78\ kJ \xrightarrow{\text{water}} NH_4^{+1}{}_{(AQ)} + Cl^{-1}{}_{(AQ)}$

Here the +ΔH indicates to you that the energy is a REACTANT since this is an endothermic reaction. When this dissolves, the solution will feel cooler and the energy required to let this dissolve will be absorbed from the water.

One last note on this, fine the formation of ammonia gas, NH_3

The ΔH here is a –91.0 kJ, it's exothermic.

$$N_2 + 3H_2 \rightarrow 2NH_3 + 91.8\ kJ \qquad (\Delta H \text{ is } -91.8\ kJ)$$

If you REVERSE THIS IN YOUR MIND, and decompose ammonia into nitrogen gas and hydrogen gas, you also can just reverse the ΔH. So,

$$2NH_3 + 91.8\ kJ \rightarrow N_2 + 3H_2 \qquad \textbf{(here the } \boldsymbol{\Delta H} \textbf{ is +91.8 kJ)}$$

Decomposition of ammonia and the synthesis of ammonia take the same amount of energy. One is "endothermic" and the other is "exothermic" to the same value.

You can reverse the ΔH for ALL Table I REACTIONS by turning around the arrow direction.

On Table I:

A. The first 6 of equations are combustion reactions
B. The next 11 are synthesis reactions.
C. Then there's 6 phase changes of solid salts dissolving into water (S→AQ). These are not technically "reactions" but still have energy associated with them.
D. We will avoid that last, strange reaction until later in the school term.

For every one of these you can "turn the arrow around and reverse signs on ΔH. For example:

Synthesis of Al_2O_3 (on table I) ΔH is –3351 kJ	$4Al_{(S)} + 3O_{2(G)} \rightarrow 2\ Al_2O_{3(S)} + 3351\ kJ$ (exo)
Decomp of Al_2O_3 (reverse of Table I) ΔH is +3351 kJ	$2Al_2O_{3(S)} + 3351\ kJ \rightarrow 4Al_{(S)} + 3O_{2(G)}$ (endo)

Thermochemistry Class Notes

Thermochemistry concerns itself with how much energy is absorbed or released in a chemical reaction, or in a phase change. Since we just learned about phases, we'll start with them first, then move to chemical reactions. Physical changes, or Phase changes can be either exo or endothermic, depending if they either release or emit heat energy, or if they absorb it.

1. Exothermic means to ________________________ heat energy. Things feel hot to you.

2. Endothermic means to ___________________________ heat energy. Things feel cold to you.

3. To melt a solid into a liquid, solids _________________________ heat energy.

4. To freeze a liquid into asolid, liquids ____________________________ heat energy.

5. To vaporize a liquid into a gas, liquids ________________________________ heat energy.

6. To condense a gas into a liquid, gases __________________________ heat energy.

7. For sublimation, when a solid turns directly into a gas, solids ______________________ heat energy.

8. For deposition, when a gas turns directly into a solid, gases _________________ heat energy.

9. When substances are hotter, their particles are moving ________________________

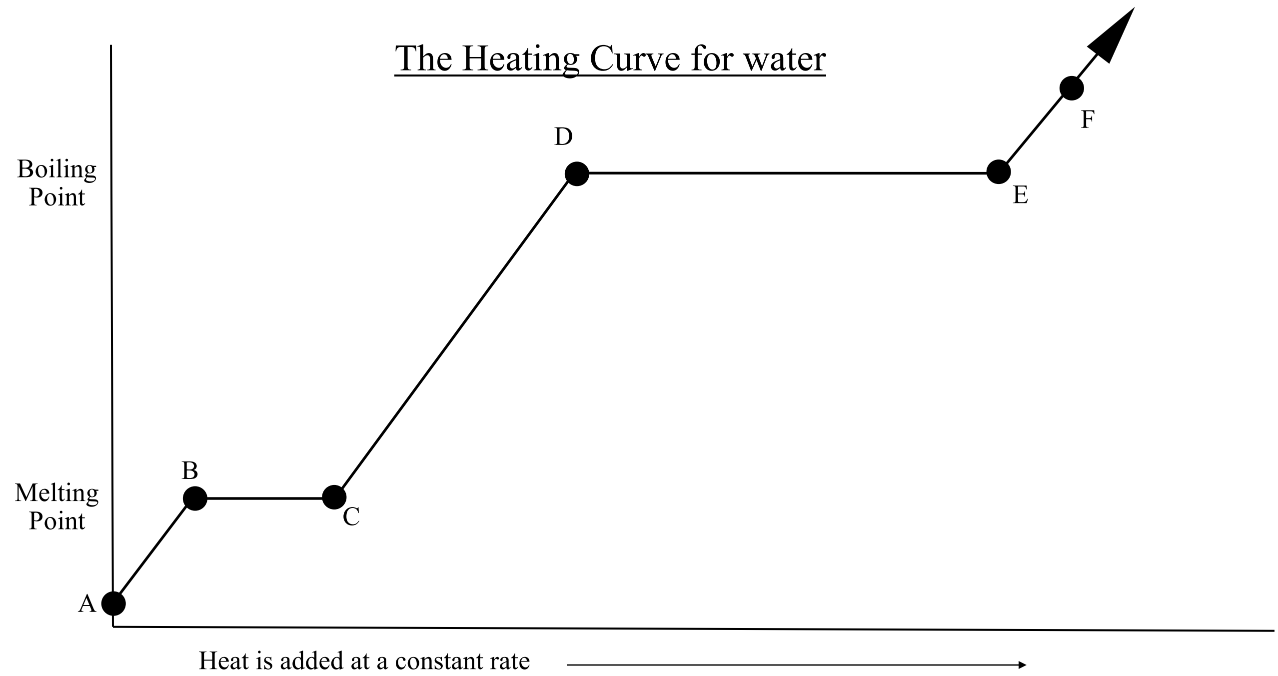
10. Hot substances have _____________________ Kinetic Energy than colder ones.

11. When substances are colder, their particles are moving _______________________

12. Colder substances have _____________________ Kinetic Energy than hotter ones.

13. Skip this one, okay? (Kinetic energy is energy of motion. Potential energy is energy of phase.)

14. Kinetic energy is the energy of motion. Faster particles have higher kinetic energy.
Kinetic energy and temperature change together, they are directly proportional.



15. Segment	Temperature and Kinetic Energy INC or DEC or STEADY?	Phase or Phases Present	Potential Energy INC or DEC or STEADY?
AB			
BC			
CD			
DE			
EF			

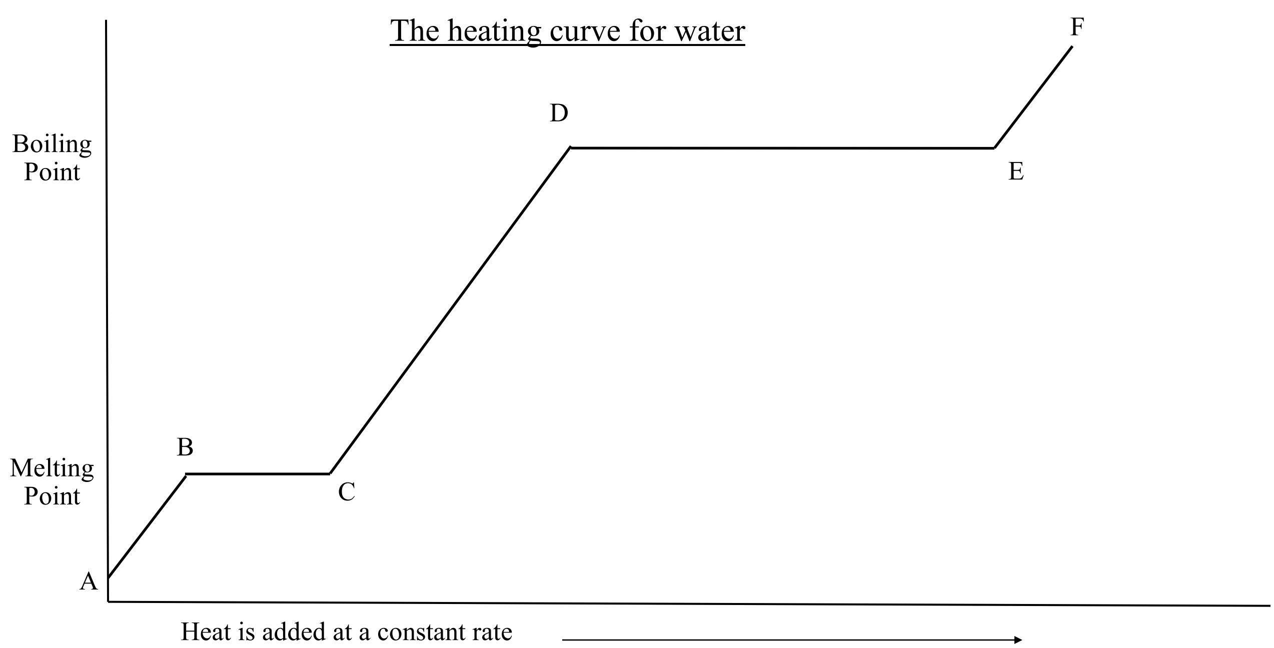
16. Why is BC shorter than DE?

17. To melt ONE GRAM of SOLID ICE into one gram of liquid water requires

18. To Boil Away (vaporize) one gram of hot liquid water into steam requires the

25. Table B has three important physical constants for water. Fill in this chart now.

Constant	Value with units	Will make one gram of H_2O...
Heat of Fusion		
Heat of Vaporization		
Specific Heat Capacity		



26. To move from B to C on this graph, we would need to ADD ______________________ of energy.

27. To move from D to E on this graph, we would need to ADD ______________________ of energy.

28. Moving from B to C is __________thermic.

29. Moving from D to E is ____________ thermic.

30. How much energy does it take to melt a real sized ice cube of 73.5 grams? (no ΔT)

31. How much energy does it take to freeze 125 grams of pure water into ice if the water starts at 0°C? (no ΔT)

32. How much energy does it take to vaporize 73.5 grams of water into steam? (no ΔT)

33. If 2.75 grams of steam condenses into water onto your finger, how much energy do you absorb? (no ΔT)

34. Which phase change takes more energy, the COLD one, or the HOT one?

35. If heat is added AT A CONSTANT RATE,
why is BC shorter than DE on the heating curve for water? ____________ ____________

36. fill in this chart		
1 gram ice		1 gram water liquid
1 gram water liquid		1 gram ice
1 gram water liquid		1 gram steam gas
1 gram steam gas		1 gram water liquid
To make 1 gram of cold water 1 Kelvin or 1°C hotter		1 gram of warmer water
1 gram of cooler water		To make 1 gram of hot water 1 Kelvin or 1°C cooler

37. To heat one gram of pure water from 275 K → 276 K you need to add ____________________ Joules.

38. To heat 1.0 gram of H_2O from 365 K → 366 K, you need to add ______________ Joules

39. If one gram of hot water, at 368 K cools to 367 K how much heat is radiated out of it? _________ Joules

40. If you put 375 grams of tap water at 294 Kelvin into the fridge to cool it all down to 275 Kelvin, how much energy needed to be removed from this water?

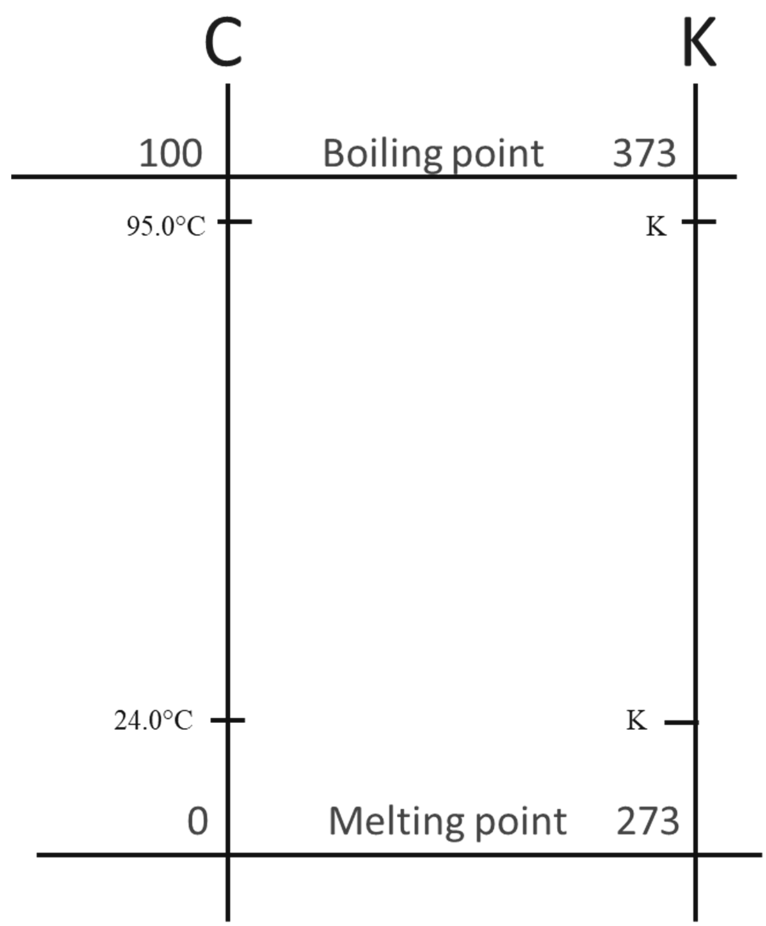
41. Let’s say you have a centigrade thermometer but need to do a temperature change thermochem problem. What do you have to do? Assume the temp change is 24.0°C to 95.0°C. First, convert these to Kelvin.

24.0°C → Kelvin K = C + 273	95.0°C → Kelvin K = C + 273

42. The T in centigrade is ______________

The T in Kelvin is _______________

43. ____________ = ______________



44. A pot has 650. grams of water at room temp, or 24.0°C. You think to make some mac & cheese and turn on the stove to heat the water. Your BFF shows up with pizza, so you turn off the stove. The water is at 95.0°C. How much energy would your Dad say you wasted by heating up this water for nothing?

45. You want some mac & cheese, so you put 650. grams of water at 24.0°C onto the stove top. You begin to text your BFF and "forget" about the pot. Your Mom yells "Who left this pot boiling on the stove?!" You realize that you foolishly vaporized 35.0 grams of your water while being distracted. You turn off the stove and decide to eat out instead. How much total energy did you waste?

Copy this here in the notes, and on the back page of your reference tables.

The most important thing in doing phase change thermomchem problems it to always use the right formula, always use the right constants, and to always use units. Challenge time…

46. How many joules of energy are required to freeze <u>355 mL</u> of water at 273K?

47. How many joules required to melt a snowball of 415 g?

48. How many joules does it take to warm up 375 grams of water at 285.0 K to 292.5 K?

Sometimes in thermochem problems you need to solve for H_F or H_V or C or even ΔT. To do this you must put the numbers in their proper place and remember it's not always water we're talking about.

49. What is the specific heat capacity constant for GOLD if it takes 271 joules to warm up a ring with mass of 34.2 g from 294.0 K to a "too hot to wear" 355.5 Kelvin?

50. What is the specific heat capacity constant of copper, if it takes 951 joules to warm up 41.63 grams of copper from 294.5 K up to 352.9 Kelvin?

51. What is the heat of fusion constant for an unknown metal if it takes 6750 Joules to melt 28.0 grams of it?

52. How many grams of ice can be melted when you add 87,500 joules of heat to it?

53. If 92.0 grams of a substance absorbs 27,496 Joules and the temperature increased from 12.3°C to 83.8°C, what is the specific heat capacity constant of this unknown?

54. You put a 155. gram snowball at –4.00°C into the back of your friend's jacket. It ultimately warms up to his body temperature of 37.0°C. How much energy did that take? (think) Look back in notes 2 pages.

step	called	formula	Do the math
1		$q = mC\Delta T$	
2		$q = mH_F$	
3		$q = mC\Delta T$	
		Total Joules required →	

55. You allow 5.75 grams of steam to condense onto your kitchen window, then it cools to room temperature of 23.5°C. How much energy is emitted in this process?			
step	called	formula	Do the math
1			
2			
Total Joules required →			

56. Table I is titled

____

57. There are _____ chemical reactions, and ___ PHYSICAL CHANGES (phase changes), and 1 weirdo thing at the bottom that we can't talk about for about 7 weeks, but we will.

LABEL YOUR REFERENCE TABLE NOW…

Let's look at the first reaction on the table… the Combustion of Methane gas (Bunsen burners)

Reaction	ΔH (kJ)

58. when the ΔH is negative, that means that the reaction is exothermic.

The mole ratio for the chemistry is ____________________________________

The thermochemical mole ratio is ____________________________________
because the energy is ALSO in ratio with the reaction.

59. What is the first endothermic reaction?

59. First endothermic reaction	ΔH (kJ)
60. Demo reaction	

61. Can energy be positive or negative? _______
It's sort of like money. You might have a $1 but you don't have +$1

If you have no money but owe me a dollar, you don't have -$1 !

62. The +ΔH means that the reaction is ABSORBING energy, it is ________________________________

The –ΔH means that the reaction is EMITTING energy, it is ________________________________

63. The 6th reaction on table I	$C_6H_{12}O_{6(S)} + 6O_{2(G)} \rightarrow 6CO_{2(G)} + 6H_2O_{(G)}$	
The ΔH is	— 2804 kJ	Heat is emitted exothermically Heat is a PRODUCT
Rewrite this		
Heat is a product in an exothermic reaction, it can be written WITH the products. The ΔH is negative, but this is not changed to a + here. The negative sign MEANS exothermic, not negative!		
Kilojoule	Is a big amount of energy	1000 Joules = 1 kilojoule
Kilogram	Is a big amount of mass	1000 grams = 1 kilogram
Kilometer	Is a big amount of length	1000 meters = 1 kilometer

64. How many joules are released when one mole of NaOH dissolves into water?

65. How many joules are absorbed when one mole of NaCl dissolves into water?

66. How many joules are absorbed when 9.75 moles of NaCl dissolves into water?

67. When you dissolve 11.9 moles of LiBr into water, how many kJ are released?

68. How much energy is JOULES is absorbed when 52.61 moles of KNO_3 is dissolved into water?

69. Convert 309 Calories into the 3 other units of energy.

70. Copy the conversion factors here too….

The specific heat capacity constant for water is 4.18 J/g·K

It takes 4.18 joules to warm up 1 gram of water by 1 K

4.18 joules = 1 cal

It takes 1 cal to warm up 1 gram of water by 1 K

It would take 354 cals to warm up 354 g H_2O at 7°C → 8°C

71. How many joules is 354 cal?

72. Convert 3429 cal into kilojoules.

73. When two mole of water gas synthesize (like the first day of school) the amount of energy released is 483.6 kJ. Convert that amount of energy into joules, cals, and Calories.

74. Methane combusts, what is the ΔH? Convert these kJ → joules, cals, and Calories.

75. If you combust 8.44 moles of methane, how many kilojoules are released?

76. If you combust 40.8 moles of glucose, how many kJ are released?

77. Liquid octane (reaction #3) combusts in car engines. If you burn up 12.890 moles of C_8H_{18}, how many kilojoules are released?

78. Table I shows us this: $4Al + 3O_2 \rightarrow 2Al_2O_3 \qquad \Delta H = -3351\ kJ$ It can be rewritten as this…
Heat energy is a product. The ΔH is negative, but we add this energy to the product side. No mix up on signs.

79. When 7.50 moles of aluminum oxide form, how many kJ are released?

80. How many calories are removed when 45.8 grams of ice forms from water at 273K?

81. What is the HEAT OF FUSION for candle wax if it takes 3388 Joules to melt a whole birthday candle with mass of 23.04 grams?

82. When a 355 mL can of seltzer, is warmed from a temperature of 293 K by adding 64,000 Joules of energy to it, what is the final temperature? Assume the seltzer is just water.

83. How many grams of water can be heated by 25.5 K when it absorbs 17,500 joules?

84. When 51.1 g of copper cools, it emits 1788 Joules. If it was at 381.5 Kelvin, what is its final temperature?
the C_{Cu} = 0.391 J/g·K ?

95. A 246.4 gram snowball at 273.0 K first melts, then warms to 26.55°. How many total joules did it take?

86. We will draw a ____.

87. Let's assume that there is exactly 2120. mL of water in our bomb calorimeter and it's at exactly 295.0 K.
After burning up our food sample, the temperature of the water rises to 354.5 K.
How many Calories of energy are in this food sample?

88. Copy the simple Cooling Curve for Chromium, labels, titles, axis labels, etc.

A. What temps are 1 + 2?

B. What's PE doing BC and CD?

C. What's KE doing AB and DE

D. Why is BC longer than DE?

E. Which thermochem formula do you use for BC?

F. Which thermochem formula do you use for EF?

89. Propane gas, C_3H_8 combusts according to Table I. How much energy (in kJ) is released when 5.75 moles C_3H_8 combusts? (find this - Table I, it's the 2nd reaction)

90. How much energy is absorbed by the reaction of 9.20 moles of $HI_{(G)}$ forming?

91. When 45.0 g of an unknown metal absorbs 1.51 kJ of heat. The temperature changes from 268 K to 345 K. What is the specific heat capacity constant for this metal? Note: KILOJOULES!

92. Copy these…

The Law of Conservation of Matter (or mass)

The Law of Conservation of Energy

93. Which takes more energy? Melting 50.0 g of ice or vaporizing 50.0 g of water into steam?

94. Which takes more energy? Heating 23.0 mL of water from 274 K to 299 K

or

Heating 23.0 mL of water from 299 K to 323 K

95. Which takes more energy? Vaporizing 21.0 g water from 373 K liquid to gas

or

Changing the temp of 100.0 g H_2O by 97.0 K?

96. Which has the LOWEST AVERAGE KINETIC ENERGY?
A. 100 mL water at 51.0°C
B. 100 mL water at 50.0°C
C. 175 mL water at 49.0°C
D. 175 mL water at 23.0°C

97. This horse has a boo boo. That's an ICE PACK on his leg. Describe the thermochemistry.
A. Heat flows from ice pack → leg
B. Cold flows from ice pack → leg
C. Heat flows from leg → ice pack
D. Cold flows from leg → ice pack

98. This is important and has not been covered yet. Thermochemistry is a 2 way street. The amount of energy is used to melt one gram of ice into water is 4.18 Joules. That same amount of energy REMOVED from a gram of cold water at the freezing point will freeze one gram of ice.

Table I shows us this: the synthesis of ammonia is exothermic

$$N_{2(G)} + 3H_{2(G)} \rightarrow 2NH_{3(G)} \qquad \Delta H = -91.8\ kJ$$

That means that the decomposition of ammonia looks like this:

Table I shows us this: the synthesis of aluminum oxide is exothermic

$$4Al_{(S)} + 3O_{2(G)} \rightarrow 2Al_2O_{3(S)} \qquad \Delta H = -3351 \text{ kJ}$$

99. How many kilojoules are absorbed when 4.58 moles of aluminum oxide decompose?

$$2Al_2O_{3(S)} \rightarrow 4Al_{(S)} + 3O_{2(G)} \qquad \Delta H = +3351 \text{ kJ}$$

You can turn around any reaction or phase change on table I, just reverse the sign and you are golden (#79)

100. You must remember these…	
The old chemistry adage goes like this…	
A good example of this is…	
Never forget…	