Thermochem Basics

Thermochemistry is the part of our course that connects chemistry with the changes of heat energy, in processes like melting or freezing, boiling (vaporization) and condensation, and with the times that energy is released, or absorbed when a reaction happens. As with most of our topics, this part is highly measured, extremely exact, and allows us to calculate energy in several units.

This topic is mathematical, and has two excellent lab experiments. The C of Cu Lab will give us near perfect results (a small percent error) while the Dorito's Lab will give us comically huge error (which is fine) but shows us exactly how this sort of science and our real life connects in the food we eat every day.

In the best movie of all time, the Wizard of Oz, the Scarecrow tells Dorothy that "of course, some people do go both ways". It's the same with the thermochem; it takes the same amount of energy to melt ice as it does to freeze it, 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. Same for cooling or warming stuff up by the same number of degrees, or for condensing or vaporizing.

The energy is a 2 way process, equal, but being added to make things "hotter", or removed to make them "colder".

It's the energy that is "something", while the 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 get the mail in the winter time, in your pajamas, the cold DOES NOT move in to you, the heat leaves our body. Heat (or energy) moves, the cold is nothing.

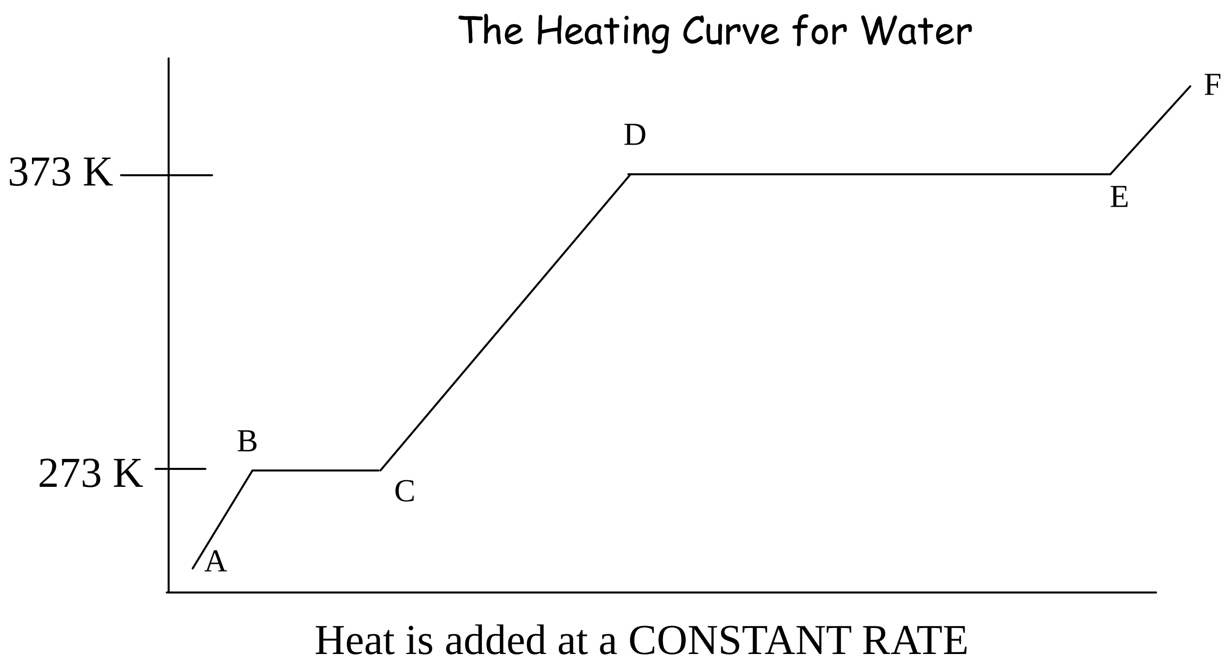
If you fall and twist your ankle, you sometimes ICE your ankle. Your ankle gets cold, which reduces the swelling. You ankle does not absorb this cold from the ice, rather the heat from your ankle moves towards the ice (to warm it up). You get a cold ankle because your body heat exits into the ice. Heat is a form of energy, heat moves. Cold is nothing.

Thermochemistry will allow us to grasp how much heat is produced in some reactions, or how much is absorbed. It let's us use some standard energy changes and do proportional math on the energy units (stoichiometry with energy, like we did with moles, grams, liters, and particles). It explains in more detail how phases changes occur, and separates multistep processes into parts that we can measure separately, them combine with addition into the whole (think melting an ice cube in your hand, from below the freezing point, up to the freezing point, then melting it, and only then - warming it up now that it's a liquid). Or turning water in a pot into steam by heating it all up to the boiling point with a measure of energy, then actually vaporizing this liquid water into a gas. Or the reverse: steam to cooler water, or water into cold water, then freezing it solid.

We will use several units, which can of course be switched around with the proper conversion factors. 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 with a mustache). There also are kilo-joules (1000 joules). We will also use calories, which are too small to eat, and Calories with a capital C, which are 1000 calories. The unfortunate calorie vs. Calorie is a pain in the neck when we talk, but we'll keep track of them this way: the "Big C" Calories are the ones you eat. We will only use the "small letter c" for calories, which we will call "cal", okay?

Melting or Freezing

All sorts of stuff freezes and melts. Everything almost. Let's start by talking about some stuff we know pretty well, like water. The process of melting solid ice into liquid water happens at the melting point of 273 Kelvin or 0°C. There is no temperature change, we saw this in our heating curve for water.
There is one of those graphs below to follow along on. Melting happens from point B to C.
The temperature does not change at all.



Point A represents "all" temperatures below the freezing point but above absolute zero (of course). The graph "starts" with very, very cold ice. Ice can be different temperatures, some colder or warmer than other ice. It's solid until it melts, which happens only at 273 Kelvin. If heat is being added, once the ice reaches the melting point, it melts. It takes a certain amount of energy PER GRAM to melt it.

Every substance (ice, iron, copper, sodium chloride, etc.) has its own energy requirement to melt. This amount of energy is called the HEAT OF FUSION constant. It takes that amount of energy, the heat of fusion, to either melt a gram of ice into water (with NO TEMPERATURE CHANGE), or the reverse, how much energy it would take to remove from one gram of liquid water (at point C) to move to the solid phase (at point B).

The energy unit we need to introduce before this formula is the Joule. The heat of fusion constant for water is 334 Joules per gram. Written in symbols, it is this: $H_F = 334$ J/g

To melt one gram of ice into liquid water at the freezing point takes 334 Joules of energy. To freeze one gram of ice at 273 Kelvin into the solid phase, you must remove that same amount of energy, 334 Joules.

If you have an ice cube that has a mass of 48.6 grams, relatively normal sized, the amount of energy required to turn it into 48.6 grams of cold water (same temperature: 273 Kelvin) would be calculated with this formula:

The "q" stands for energy in JOULES. They equal the mass of the substance in grams multiplied by the constant called the heat of fusion constant for the substance. Water's H_F = 334 Joules/gram. So, to melt an ice cube of 48.6 grams into the liquid phase, with no temperature change, use this formula, watch your significant figures:

$$q = mH_F$$

$$q = (48.6\ g)(334\ J/g)$$

$$q = 16{,}232.4 \text{ Joules} = 16{,}200 \text{ Joules with 3 SF}$$

During a phase change (melting or freezing) or (vaporization or condensing) there is NO TEMPERATURE CHANGE. It happens at THE freezing point, or THE melting point, which is the same temperature.

Vaporization of water occurs at 373 Kelvin, and condensing happens at the same 373 Kelvin.
NO TEMPERATURE CHANGES occur during phase changes.

A symbol for change in temperature is ΔT, which reads "delta T", and stands for change in temperature. A phase change has $\Delta T = 0$, there is NO ΔT in a phase change.

To convert these joules into other units will require a few equalities. Copy these now, into your reference table, below Table B (which has the three constants for water for thermochemistry).

1000 Joules = 1 kilo-Joule
4.18 Joules = 1 cal (small c)
1000 cal = 1 Calorie *(1 Calorie = 1 Food Calorie)*
1 Calorie = 1 kilo-calorie

Let's convert those 16,200 Joules into some other units now. First, into kilo joules this way:

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

It is important to point out that you do not really grasp the amount of energy that a Joule represents. That is okay, you are confused a bit, that's alright. 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.

The "opposite" of melting is freezing. Say you have 3500. milliliters (4SF) of pure water at the freezing point but still in the liquid phase. It won't freeze until you remove enough energy to slow the molecules down enough to let the attraction they have exceed their kinetic energy to move around. How much energy must be removed from 3500. mL of water at 273 Kelvin to freeze it solid? Use the same formula:

$$q = mH_F$$

$$q = (3500.\ g)(334\ J/g)$$

$$q = 1{,}169{,}000\ \text{Joules}$$

Now, convert those Joules into kilo-Joules this way:

$$\frac{1{,}169{,}000\ \cancel{\text{Joules}}}{1} \times \frac{1\ \text{kilo-Joule}}{1000\ \cancel{\text{Joules}}} = 1{,}169\ \text{kilo-Joules} \qquad \textit{(4 SF)}$$

To convert these Joules into cal, or Calories, do this:

$$\frac{1{,}169{,}000\ \cancel{\text{Joules}}}{1} \times \frac{1\ \text{cal}}{4.18\ \cancel{\text{Joules}}} = 279{,}700\ \text{cal} \qquad \textit{(4 SF)}$$

Then:

$$\frac{279{,}900\ \cancel{\text{cal}}}{1} \times \frac{1\ \text{Calorie}}{1000\ \cancel{\text{cal}}} = 2{,}797\ \text{cal} \qquad \textit{(4 SF)}$$

Food you eat is energy. You use this food (energy) to stay alive, to move, to be. 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. Too much energy in makes too much of you. If you are not eating much, this stored energy is used up, and you lose some weight as your fat is converted back into energy.

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.

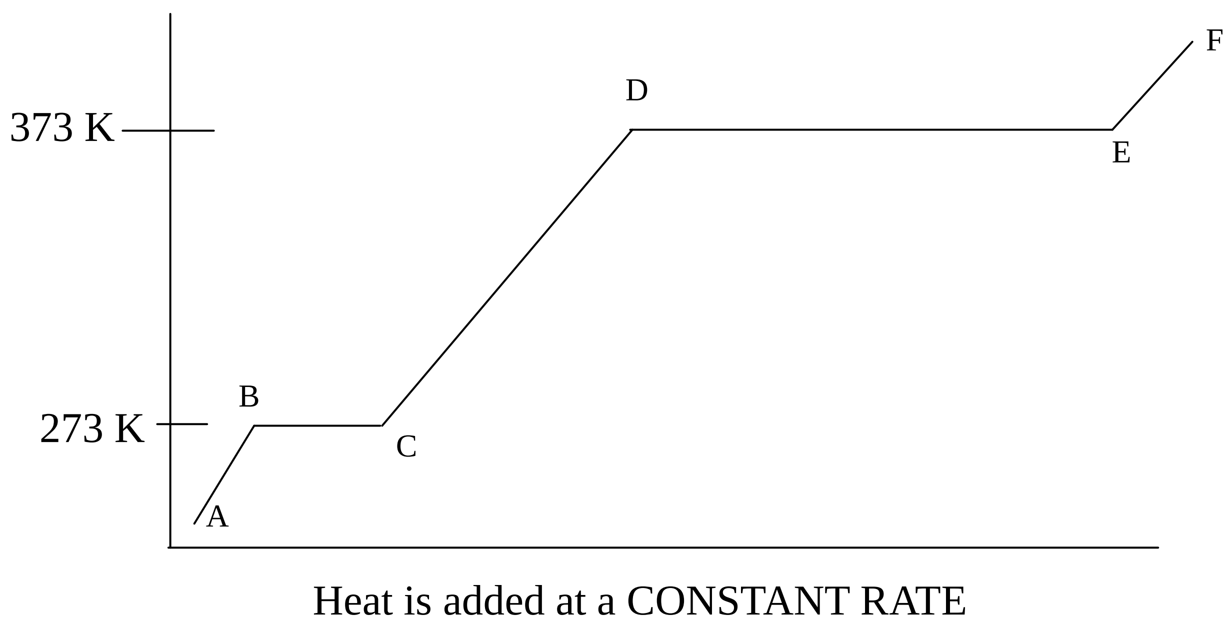
Energy is not temperature, temperature is a measure of the kinetic energy the stuff has. Potential energy is something altogether different. Energy can be thermal (hot), electrical, or in other forms. Energy is stored in bonds in chemistry, or released from bonds.

I want to say that Energy is Cool, but really, energy is HOT, the lack of energy is cool. Don't forget that.

Vaporizing or Condensing

All sorts of stuff boils off or condenses back into a liquid. Let's start again by talking about some stuff we know pretty well, like water. The process of vaporizing liquid water into a gas called steam happens at the boiling point of 373 Kelvin or 100°C. There is no temperature change, we saw this in our heating curve for water. The graph will help you to follow along on. Vaporizing happens from point D to E. The temperature does not change at all.

The Heating Curve for Water



Point D represents the hottest temperature that liquid water can exist. At that point, the kinetic energy is so strong (meaning the temperature is so hot), that the molecules of water a at the edge of not being able to hold together anymore. They are vibrating so fast that if they vibrate any faster, if they gain any more energy, they will blow apart into the gas phase. The amount of energy required to push these hot liquid molecules into the gas phase is called the HEAT OF VAPORIZATION constant.

For water this constant is 2260 Joules/gram. In symbols, H_V=2260 J/g.

To phase change from liquid to solid, to move H_2O from point D to E on this graph, happens at one temperature, THE boiling point, which is 373 Kelvin. No temperature change at all.

The reverse, the condensation of steam into the liquid phase, without temperature change requires the steam to release that heat of vaporization, releases 2260 J/g to go from point E to point D on this graph.

Water has a heat of vaporization constant of 2260 J/g.
Every other substance has a different H_V a different constant for that substance.

To vaporize 5.60 grams of water at 373 Kelvin into the gas phase requires how much energy in Joules? Do this problem use this formula. This looks like the other one, but with a different constant H_V instead of H_F.

$$q = mH_V$$
$$q = (5.60\ g)(2260\ J/g)$$
$$q = 12{,}656\ \text{Joules} = 13{,}700\ \text{Joules} \quad \text{(with 3 SF)}$$

To convert these 13,700 Joules into cal, Calories, or kilo-Joules do this math:

$$\frac{13{,}700\ \cancel{\text{Joules}}}{1} \times \frac{1\ \text{kilo-Joule}}{1000\ \cancel{\text{Joules}}} = 13.7\ \text{kilo-Joules} \quad \textit{(3 SF)}$$

Then:

$$\frac{13{,}700\ \cancel{\text{Joules}}}{1} \times \frac{1\ \text{cal}}{4.18\ \cancel{\text{Joules}}} = 3280\ \text{cal} \quad \textit{(3 SF)}$$

Then:

$$\frac{3280\ \cancel{\text{cal}}}{1} \times \frac{1\ \text{Calorie}}{1000\ \cancel{\text{cal}}} = 3.280\ \text{Calories} \quad \textit{(3 SF)}$$

If steam hits your finger, you get quite burned. The steam condenses into a liquid, and has to release the same 2260 Joules/gram that it took to make that steam earlier. Making steam requires fire (a stove, or a Bunsen burner). A lot of energy is required to make steam, the same "lot of energy" is released when it phase changes back into a liquid with NO TEMPERATURE change.

For water,

Freezing or Melting, use: $q = mH_F$ with $H_F = 334\ J/g$

Boiling (Vaporizing) or Condensing, use: $q = mH_v$ with $H_V = 2260\ J/g$

These constants (with one more) are on Table B of your reference tables.
These formulas are on Table T, in the HEAT SECTION.

The q is the "heat in joules", m is the "mass in grams". The two constants are H_V AND H_F
Just don't mix them up! Remember that $H_V > H_F$.

Changing Temperatures

There are times when heat is added or removed from H_2O that do not change it's phase. Sometimes the temperature just increases or decreases, but the phase is not affected. When there is a temperature change, we need a formula with a ΔT so we can include that temperature change into the math. We will use Kelvin for temperature. If you have centigrade temperatures, change them to Kelvin for our formulas.

We will start with water first, then move on to other substances.

It takes a certain amount of energy to change the temperature of water. To change the temperature of one gram of pure water by one Kelvin, it takes 4.18 Joules. Since 4.18 J = 1 cal, you might also say, it takes 1 cal to make 1 gram H_2O change by one Kelvin.

This energy is added to HEAT water, or to remove energy to COOL water. It is the same amount of energy going in or out of water to heat or cool it. It takes 4.18 Joules per gram per Kelvin to change temperature in either direction. This energy has a name too, it's called the SPECIFIC HEAT CAPACITY CONSTANT, or the amount of energy it takes to change one gram of water by one Kelvin.

In symbols it is called the "C" of water (this C is different from the capital C of Calories, I am sorry that I wasn't around in the beginning to straighten all those mustaches out).

$C = 4.18\ J/g{\cdot}K$ this is the coolest (hardest), or oddest unit of the year.

It reads: the specific heat capacity constant for pure water is 4.18 Joules PER gram PER Kelvin.
That dot in the formula → · means multiplication. That formula is on table T.

Our first problem with this will be, how many joules of energy will it take to warm up 134 mL of water from 275 Kelvin to 297 Kelvin (cold water becomes warmer)?

We start with a formula, then fill in the parts. It is $q = mC\Delta T$, and the letters are all the same as before, q is the heat in joules, m is the mass in grams, C is the specific heat capacity constant just above for water, and the ΔT is the change in temperature in Kelvin.

Let's do this problem.
134 mL of pure water = 134 grams because the density of pure water is 1 gram/mL,
but you remembered that, right?

$$q = mC\Delta T$$
$$q = (134\ g)(4.18\ J/g{\cdot}K)(22.0\ K)$$
$$q = 12{,}322.64\ \text{Joules} = 12{,}300\ \text{Joules} \quad \textit{with 3 SF}$$

We will see later that you can solve for any of the variables, solving for "q" being the "easiest" since there is no algebra.

To cool water in a fridge it's the same formula, but this time, to cool water, the energy is removed, not added. Joules come out and water cools down since the kinetic energy decreases. Putting Joules in makes the particles move faster, which is measured as increased temperature.

*(Michael Jackson's hand and glove are temperature and kinetic energy,
what ever one does, so does the other)*

How much energy is removed from 3.25 L of water at 306 K when it drops to 275 K in the fridge?

$q = mC\Delta T$
$q = (3250\ g)(4.18\ J/g{\cdot}K)(31.0\ K)$
$q = 421{,}135$ Joules = 421,000 Joules *with 3 SF*

Time to talk about Kelvin and Centigrade temperatures. Two different guys with mustaches, Mr. Celsius and Lord Kelvin came up with these different ways to measure temperature. They both measure temperature but on different scales. We know that K = C +273, that these temperatures are convertible with each other. Water freezes at 273 K or 0°C. Kelvin does not use degrees, just Kelvins, but water freezes when the molecules get locked together, and it really doesn't matter what scale you are using. Boiling water boils at 100°C or 373 Kelvin. Note here that the difference is 100 units of temperature for this in either scale.

373 K is not 100°C even though they mean the same thing, they are NOT the same numbers. To get from this boiling point to the freezing point, the numbers shift to 273 K and 0°C. Both scales drop by 100 units of temperature.

This is very important, the change in temperature in centigrade is the same as the change in temperature Kelvin. The ΔT for boiling to freezing points in Kelvin is 100 K. If we measured that in centigrade it would be 100°C. The change is 100 units either way.

Since all of our thermometers measure in centigrade and not Kelvin, remember this: the change in any temperature in centigrade can be switched to Kelvin by a switch of the unit.

The TEMPERATURE NUMBERS are not the same, but the CHANGE IN TEMPERATURE is.

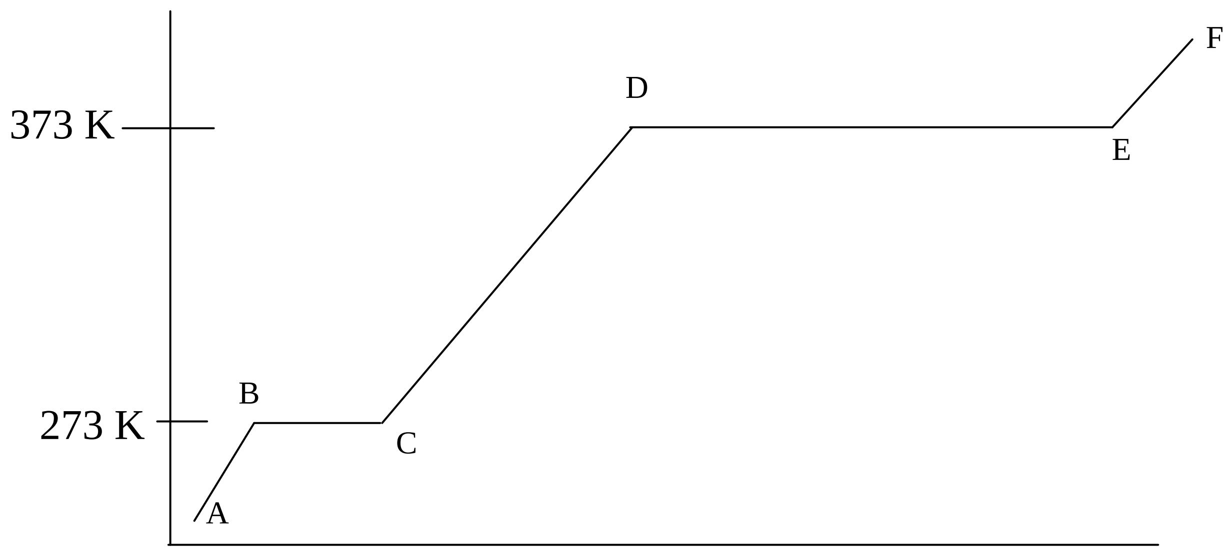
Let's disregard SF for a moment, look at this scale:

If your water changes temperature like this in centigrade, your ΔT is (84.7 - 25.6 =) 59.1°C
Or you might say the ΔT is (357.7 - 289.6 =) 59.1 Kelvin. The change is the same, and easily adjusted from centigrade to Kelvin by just switching the units.

In symbols, the °C ≠ K, but the ΔT°C = ΔT Kelvin.

That is important to grasp, or else you will be converting temperatures from C to K and back all of the time. The temperatures are not the same, but the change in temperatures is.

Let's look over this **heating curve for water** again, but this time, let's have YOU put the formulas into the graph where they belong. Use the guide below.



Heat is added at a CONSTANT RATE

To change from point B to C OR from C to B, we use this formula: $q = mH_F$

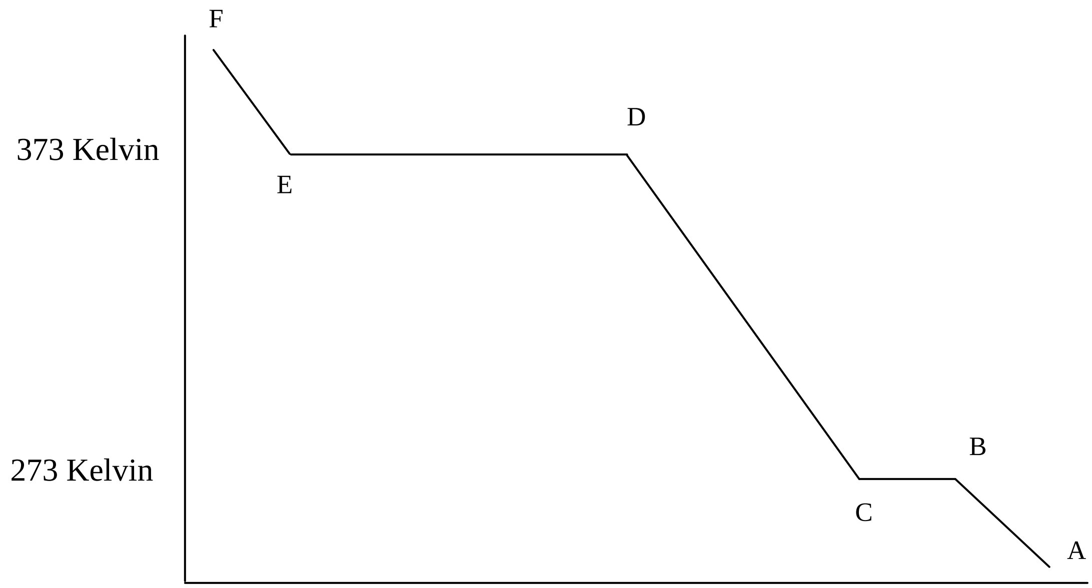
To change from C to D, or anywhere along between that, heating liquid water up, or cooling it down without phase change, we use this formula: $q = mC\Delta T$

To change from point D to E OR from E to D, we use this formula: $q = mH_V$

For water
H_V = 2260 J/g H_F = 334 J/g C = 4.18 J/g·K

Since heat is being added at a constant rate, the reason segment BC is so short is that it ONLY takes the H_F or 334 Joules/gram to melt ice, while it takes the H_V or 2260 Joule/gram to vaporize that same one gram mass of water. The H_F is about 7x smaller than the H_V.

The Cooling Curve for Water



Energy removed at a constant rate →

To condense, move from point E to point D. There is NO TEMPERATURE CHANGE.

To freeze, move from point C to point B. There is NO TEMPERATURE CHANGE.

To cool water from hot liquid to cool liquid, move down D to C, or part way in between.

EF is still nearly 7x longer than CB because releasing energy to condense requires releasing 2260 Joules/gram for water, and freezing only takes 334 Joules/gram to be removed.
The H_V is STILL much bigger than the H_F in the colder direction on this graph.

In real life things often get "messy". For instance, you might melt an ice cube at the freezing point to body temperature in your mouth. Before you can make the temperature rise from freezing to about 37°C. It takes two steps: 1st… energy to melt ice into liquid, 2nd… energy to warm the water up. You must calculate both energy requirements, and add them together.

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 to 373 Kelvin, then measure how much energy to boil away PART of that water.

Sometimes you 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 into hot steam:
1. warm the ice to the melting point
2. melting it,
3. warming up the water to the boiling point
4. vaporizing it
5. then heating up the steam to the final hottest temperature.

5 steps is the longest thermochem problem in the world!

Here are some two 2-step problems, then one 3-step problem. Get psyched now.

A. 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°C (it could happen). How much energy is required to do this?
(note: when ΔT is 37°C, then the ΔT is 37 K)

First, melt it, then warm it, finally, sum up all the joules. We will use 3 SF for this problem.

1. $q = mH_F$ = (657 g)(334 J/g) = 219,000 Joules
2. $q = mc\Delta T$ = (657 g)(4.18 J/g·K)(37 K) = 102,000 Joules
3. The total energy is 219,000 J + 102,000 J = 321,000 Joules

B. 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°C, how much energy is absorbed by your body, assuming no energy is lost to the air.

First, condense it, then cool it, finally, sum up all the joules. We will use 3 SF for this problem.

1. $q = mH_V$ = (13.0 g)(2260 J/g) = 29,400 Joules
2. $q = mc\Delta T$ = (13.0 g)(4.18 J/g·K)(63 K) = 3420 Joules

The total energy is 29,400 J + 3420 J = 32,800 Joules

C. When you have 25.0 grams of ice at the freezing point, and vaporize it, how much energy is required to do this?

First, melt it, then heat it up to the boiling point, then vaporize it, in 3 steps.

1. $q = mH_F$ = (25.0 g)(334 J/g) = 8350 Joules
2. $q = mc\Delta T$ = (25.0 g)(4.18 J/g·K)(100 K) = 10,450 Joules
3. $q = mH_V$ = (25.0 g)(2260 J/g) = 56,500 Joules

The total energy is 8350 J + 10,450 J + 56,500 J = 75,300 Joules

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All substances have their own constants. Water constants are in table B.
The H_F constant for aluminum is 403 J/g. It takes 403 Joules per gram to melt aluminum at the melting point of aluminum (look up the melting point if you want to on table S).

The C of Cu, the specific heat capacity constant for copper is 0.39 J/g·K. The constants for other substances will be given to you, or you can solve for them using our 3 thermochemistry formulas when given other information. It's just doing math.

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 is unlimted

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 q

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 the funky units that come out exactly as the "C" units 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)

   ΔT = 836.0 J/52.25J/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) H_V = 11,706.8 J/5.18 g H_V = 2260 J.g

6. What is the final temperature of 355.0 grams of copper at 29.5°C when you make it absorb 5,604 Joules of energy. Round answer here to nearest whole number Kelvin

   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?

   $q = mC\Delta T$ 5604 J = (355.0 g)(0.39 J/g·K)(ΔT) ΔT = 5604 J/138.45 J/K)

   ΔT = 40.48 Kelvin. So, since 29.5°C is really 302.5 K, the final temp is Start + ΔT, or

   302.5 K + 40.48 K = 342.98 K = 343 Kelvin

To measure the energy content in food, in Calories, an indirect method is used. The food is completely burned inside a device that will capture all of the heat given off, in a measured amount of pure water. Since the heat is gained by the water is FROM the food, the heat gained by the water can be measured using the $q = mC\Delta T$ formula. That heat is calculated in joules, then converted in two steps to Calories.

The device is called a bomb calorimeter. The "bomb" part is the box that the burning happens in. This bomb is put inside a carefully measured mass of pure water, and the start temperature of the water is measured. The hottest temperature of the water after the burning lets you measure the ΔT. The C for water is 4.18 J/g·K, and the mass is the mass in grams. Solve for q in Joules, then convert.

For example, if you burn a Hershey bar in a bomb calorimeter and measure the heat gained by the water surrounding it, that same amount of heat gain is the energy FROM the candy.

When you burn a candy bar in a calorimeter with 6345.0 grams of pure water at 274.00 Kelvin, the water temperature rises to 325.22 Kelvin, how many Calories are in the candy bar?

$$q = mC\Delta T$$

$$q = (6345.0 \text{ g})(4.18 \text{ J/g·K})(325.22 - 274.00 \text{ K})$$

$$q = (6345.0 \text{ g})(4.18 \text{ J/g·K})(51.22 \text{ K})$$

$$q = 1{,}358{,}461.96 \text{ Joules} = 1{,}358{,}500 \text{ Joules} \quad \text{with 5 SF}$$

Then:

$$\frac{1{,}358{,}500 \cancel{\text{ Joules}}}{1} \times \frac{1 \text{ cal}}{4.18 \cancel{\text{ Joules}}} = 325{,}000 \text{ cal} \qquad \textit{(5 SF)}$$

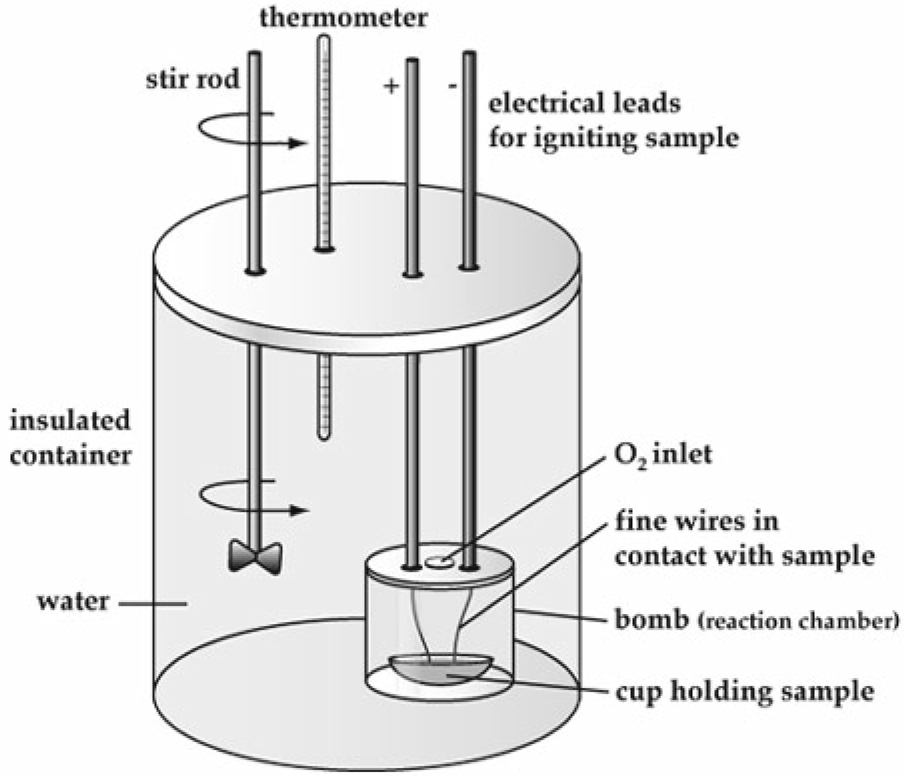
$$\frac{325{,}000 \cancel{\text{ cal}}}{1} \times \frac{1 \text{ C}}{1000 \cancel{\text{ cal}}} = 325.00 \text{ Calories} \qquad \textit{(5 SF)}$$

A **bomb calorimeter** looks like this diagram below. We will "make" a calorimeter in the Dorito's Lab, out of a can and a piece of aluminum foil. We will measure the energy gain in the water from the burning chip under it. Most (80% or more) of the heat will be lost, but we will go though the SAME process, and math, to calculate how many Calories are in our Dorito's Chip. I will shake your hand if you get better than -80% Error, but not if you do worse!

The way it "works" is this, the food goes into the "bomb". The bomb is pumped up with oxygen so it will allow the food to burn up completely. Two wires run into the bomb to set the explosion off electrically.

The bomb is put into a carefully measured mass of water. The temperature is carefully measured as well. A propeller spins to make sure that the water temperature is measured right. Once the bomb goes off, the food burns, the heat generated changes the water temperature.

The change of water temperature is put into the q = mCΔT formula. Solve for the q JOULES, convert to Calories.



In our Dorito's Lab we will "make" a calorimeter out of an empty seltzer can, and some aluminum foil to help direct the heat into the can. We will lose most of the heat, but that's just okay.

When you look at this chemistry process: $F_2 \rightarrow F + F$ it shows how a fluorine molecule is broken into two fluorine atoms. To do this it takes energy. You have to add energy to the molecule to break it apart. The reverse, $F + F \rightarrow F_2$ is the opposite, it releases energy.

Most importantly: **When BONDS FORM, energy is RELEASED**.
To Break Bonds, Energy is ADDED.

Ice + ENERGY → Liquid water   to break the bonds of water molecules apart, add energy.
Energy is a REACTANT

Water → Ice + ENERGY   In this case energy is removed to lower the kinetic energy enough to freeze solid. ENERGY is a PRODUCT

Energy can be pulled into a chemical system if the reaction requires it, or it can be given off. Energy is given off when we burn methane gas with the Bunsen burners. When energy is given off as a product of the reaction, it's said to be an EXOTHERMIC reaction. The energy is "out" or exo.

If a reaction requires more energy input than it gives off, then it's called an ENDOTHERMIC reaction. Endo means "in".

People have ENDO skeletons, lobsters and crabs and bugs have EXO skeletons.

If a chemical process is exothermic, we feel heat given off. If the process is endothermic and energy is absorbed by the system, it feels cold (as the energy is absorbed into the bonding).

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 until we get to acids + bases.

The first reaction, the combustion of methane is balanced and has a ΔH. This "delta H" stands for the "change in heat". It tells us that the reaction is exothermic, because all negative ΔH values are exothermic. The energy is just energy (in kilo-Joules to be exact). It is NOT negative or positive energy. It is Exothermic when a negative ΔH and it is Endothermic when the ΔH is +.

That positive or negative value comes later when we look at certain potential energy graphs in kinetics. For now, 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 negative $20". Money is. Energy is, too. The sign only indicates if the energy is a product in an exothermic reaction, or a reactant in an endothermic one.

The proper balanced reaction WITH the heat included would be written this way:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4 \text{ kJ}$$

The combustion of one mole of methane will produce 890.4 kilo-Joules of energy. The energy is not "negative", but this is an EXOTHERMIC reaction, so energy is written with the products.

If you remember stoichiometry (you better), you remember if you combust 3 moles of methane, you would need 6 moles of oxygen, you'd end up with 3 moles of carbon dioxide, and 6 moles of water. You would also end up with 3 x 890.4 kJ of energy as well.

Same as if you only combusted ½ mole of methane, you'd end up with half the energy being given off.

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)

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

Solve for X

X = 2219.2 kJ x 11.4

X = 25,298.88 kJ = 25,300 kJ *with 3 SF*

The moles:moles ratio now includes the energy in kJ.

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

The balanced thermochemical reaction from table I is

$$4Al + 3O_2 \rightarrow 2Al_2O_3 + 3351\ \text{kJ}$$

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

Solve for X

4X = 2219.2 kJ x 11.4

X = 8176.44 kJ = 8180 kJ *with 3 SF*

NOTE: the energy is NOT NEGATIVE. The $-\Delta H$ means exothermic.

If the reaction had a $+\Delta H$ that means it's endothermic.

Moles are in proportion to the energy, which is expressed in kJ, but can be converted at any time.

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

This can be written properly as:

$$NH_4Cl_{(S)} + 14.78\text{ kJ} \xrightarrow{\text{water}} NH_4^{+1}{}_{(AQ)} + Cl^{-1}{}_{(AQ)}$$

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 an exothermic reaction.

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

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

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

The decomposition of ammonia is "as endothermic" , as how exothermic the synthesis of ammonia is. The change in energy is equal in value, but opposite.

You can do this for ALL REACTIONS on table I.

NOTE for table I:
The first 6 of equations are for combustion reactions
The next 11 are synthesis reactions.
Then there are 6 phase changes of solid salts dissolving into water (S→AQ phase changes).
Those are not technically "reactions" but still have energy associated with them.

The last reaction we'll avoid until later in the year.

For every one of these you can "turn the arrow around and change the sign for ΔH

The synthesis of Al_2O_3 is exothermic, it gives off heat energy. (ΔH is –3351 kJ)

The decomposition of Al_2O_3 is endothermic, it requires heat energy. (ΔH is +3351 kJ)