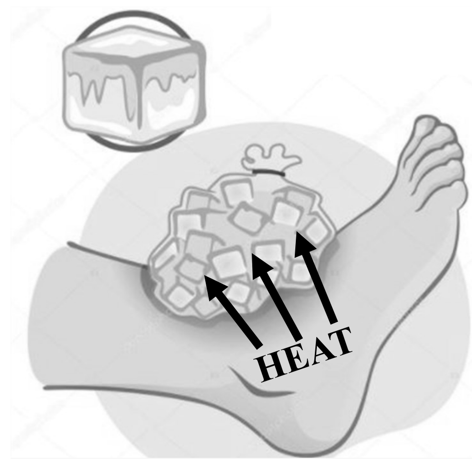
Thermochem Basics

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.

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, heat left your body. Heat (energy) moves, the cold is nothing.

If you fall and twist your ankle, you should 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 into the ice.
You get a colder ankle because your body heat exits your ankle as it moves into the ice.

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.

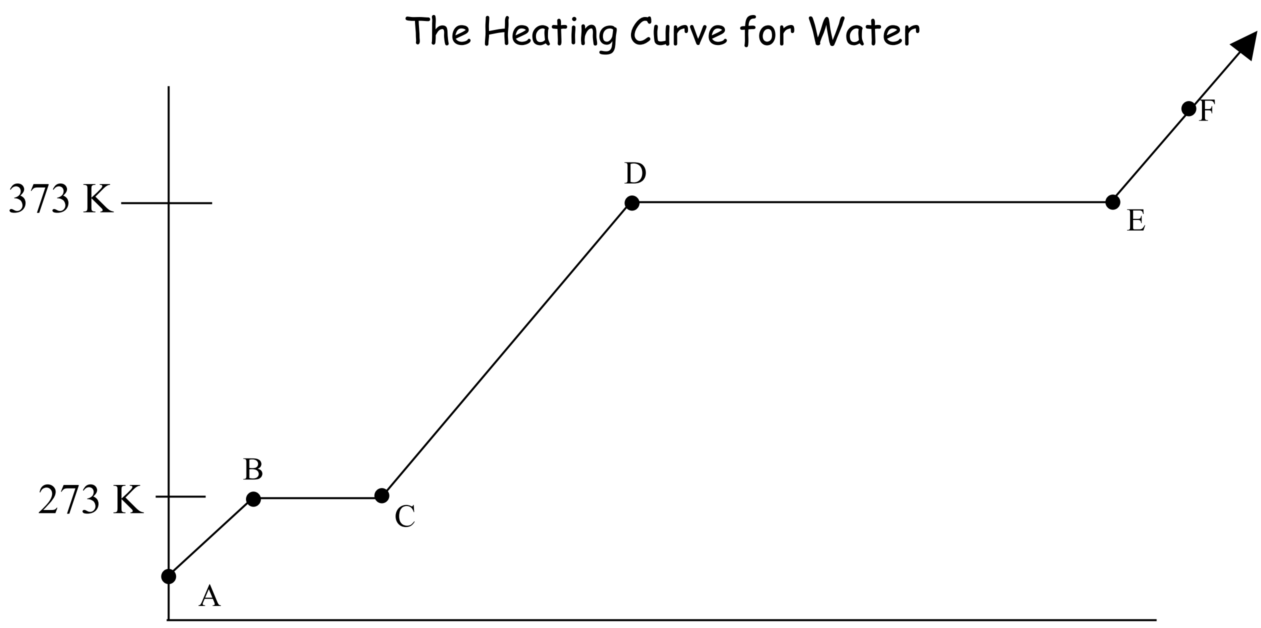
We will measure energy in joules (named after a chemist). There also units called kilo-joules (1000 joules).
We will also use "calories", which are scientific (and lower case), and
"Calories" (capital "C") that we eat as FOOD CALORIES or kilocalories.

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. Write these three equalities into your reference tables now, underneath table B. (now).

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

Melting or Freezing

All sorts of stuff freezes and melts. Let's start by talking about some stuff you 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 see this in the heating curve for water below from point B to C.



Point A represents a temperature below the freezing point but above absolute zero. 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, and the temperature gets "stuck" at the melting point until all of the ice melts at point C.

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 (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.

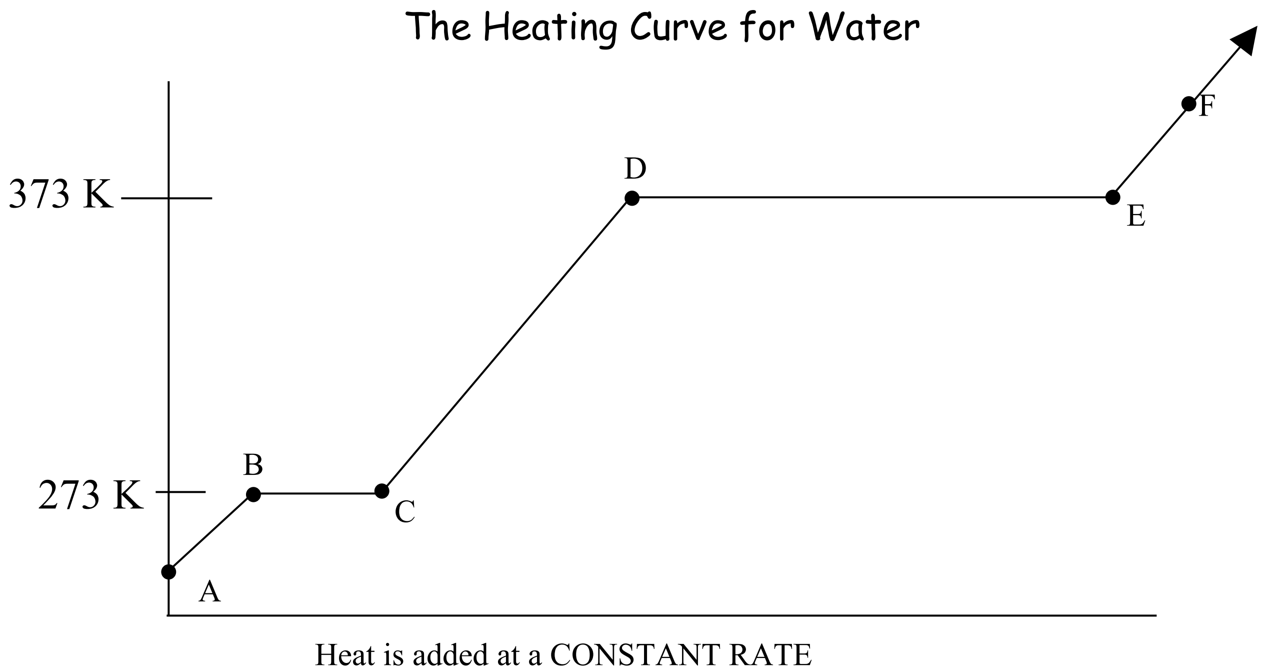
The energy unit we need to introduce before this formula is the Joule. The H_F constant for water is 334 Joules per gram. Written in symbols, it is this: $H_F = 334$ J/g

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.

To move from point B to C, or reverse on a cooling curve from C to B, requires the H_F energy.

Vaporizing or Condensing

All sorts of stuff boils into a gas, or condenses back into a liquid. Let's talk about water again. The process of vaporizing hot water into steam happens at the boiling point of 373 Kelvin or 100°C. There is no temperature change, we see this in the heating curve for water below from point D to E.



Point D represents the HOT PHASE change temperature. Below 373 Kelvin water is always liquid. Liquid water can't get hotter than that temperature, the boiling point. From D to E the amount of liquid and gas changes. There is NO GAS at D, and by the time the graph glides over towards E, there is less liquid and more gas, until at point E there is ALL steam and no liquid left. It takes a certain amount of energy PER GRAM to vaporize water, and the temperature gets "stuck" at the boiling point until all of the water vaporizes into steam at point E.

Every substance has its own energy requirement to vaporize. This amount of energy is called the HEAT OF VAPORIZATION constant. It takes that amount of energy, the heat of vaporization, to either boil off a gram of water into steam (NO TEMPERATURE CHANGE), or the reverse, how much energy would be released when one gram of steam condensed into one gram liquid water.

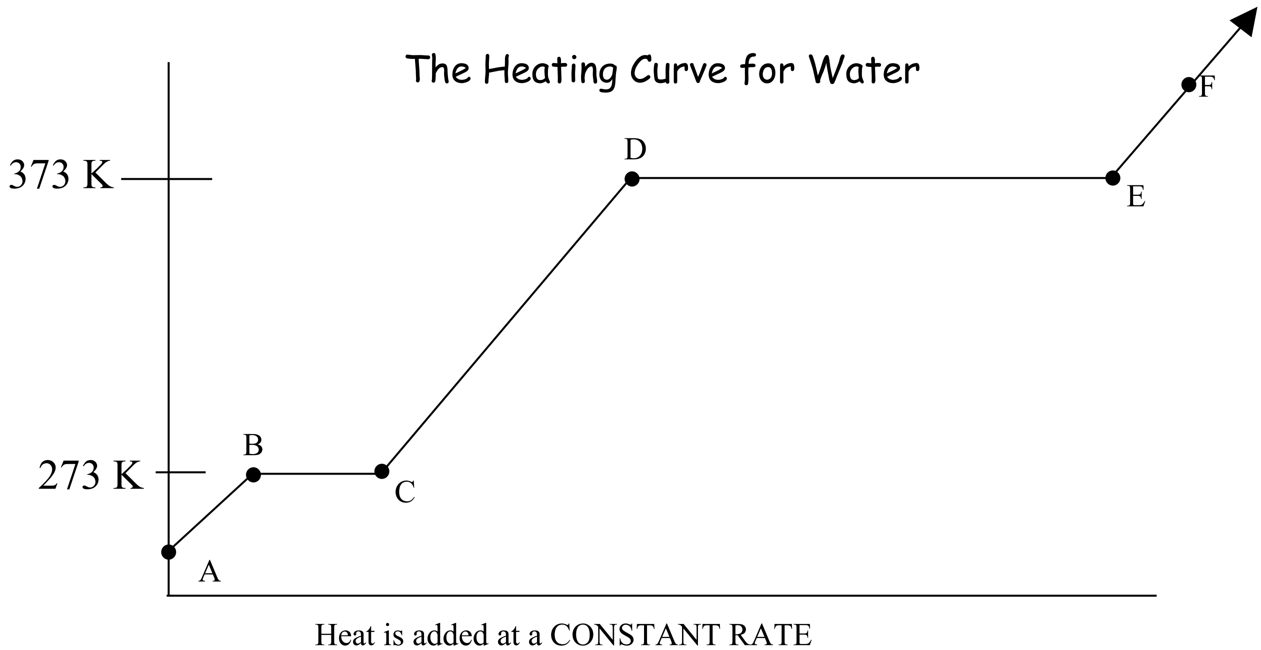
The energy unit we need for this formula is the Joule. The H_V constant for water is 2260 Joules per gram. Written in symbols, it is this: $H_V = 2260$ J/g

To vaporize one gram of liquid water at the boiling point takes 2260 Joules of energy. To condense one gram of steam at 373 Kelvin into the liquid phase, you must remove that same amount of energy, 2260Joules.

To move from point D to E, or reverse on a cooling curve from E to D, requires the H_V energy.

Warming or Cooling of Liquid Water

You can change the temperature of liquid water (hotter or colder) in btween the melting point and the boiling point too. 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 for after your exercising. This changing temperature of the liquid phase takes a different amount of energy, and we have to take into consideration HOW MUCH temperature change is happening. More temperature means adding, or removing more energy.



On the heating curve this all happens from point C to point D. Liquid stays liquid, but gets hotter or colder.

To change the temperature of liquid water it requires a certain amount of energy per gram per degree. This is called the SPECIFIC HEAT CAPACITY CONSTANT. For liquid water it takes the addition of 4.18 Joules per gram per degree to make the water warmer. To cool water's temperature lower, you 'd have to remove this same amount of energy. The unit is weird, but also way cool. The symbol for specific heat capacity constant is a capital "C". So, for liquid water, C = 4.18 J/g·K which reads "C" equals 4.18 Joules per gram per Kelvin or "C" equals 4.18 Joules per gram dot Kelvin.

To warm up one gram of liquid water it takes just 4.18 Joules of energy. To cool off one gram of liquid water remove that same amount of energy, 4.18 Joules.

To move from point D to E, or reverse on a cooling curve from E to D, requires the specific heat capacity constant.

Every substance has a different specific heat capacity constant. For example…

Substance	Specific heat capacity constant		Substance	Specific heat capacity constant
Liquid water	4.18 J/g·K		Mercury	0.140 J/g·K
Copper metal	0.45 J/g·K		Iron	0/45 J/g·K
Ice	2.10 J/g·K		Steam	1.90 J/g·K

Thermochemistry Math

Formulas on back of reference table in the "HEAT" section.

How much energy is needed to melt a normal sized ice cube with mass of 48.6 grams into liquid water at the melting point (no temp change)?

This is calculated with this formula:

$q = mH_F$

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, (no temperature change) do this:

$$q = mH_F = (48.6\text{ g})(334\text{ J/g}) = 16{,}232.4\text{ Joules} = 16{,}200\text{ Joules with 3 SF}$$

How much energy is needed to vaporize 2.29 grams of hot water (at the boiling point) into steam at the same temperature?

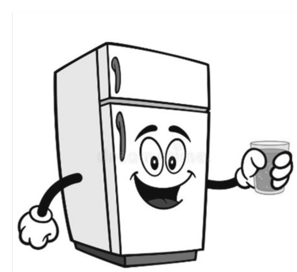
This is calculated with this formula:

$q = mH_V$

The "q" stands for energy in JOULES. They equal the mass of the substance in grams multiplied by the constant called the heat of vaporization constant for the substance.

Water's H_V = 2260 Joules/gram. So, to vaporize this hot water into steam, do this:

$$q = mH_V = (2.29\text{ g})(2260\text{ J/g}) = 5175.4\text{ Joules} = 5180\text{ Joules with 3 SF}$$



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?

This is calculated with this formula:

$q = mC\Delta T$

The "q" stands for energy in JOULES. They equal the mass of the substance in grams multiplied by the constant called the specific heat capacity constant for the substance. Water's C = 4.18 J/g·K.

The temperature here changes by 6.50 C, the ΔT is 6.50 C. NOTE: the change in temp is the same in centrigrade as it is in Kelvin. The ΔT C = ΔT K.

So, to cool this warm water into colder water, do this:

$$q = mC\Delta T = (500.0\text{ g})(4.18\text{ J/g·K})(6.50\text{ K}) = 13585\text{ J} = 13600\text{ J with 3 SF}$$

During a phase change (melting or freezing) or (vaporization or condensing) there is NO TEMPERATURE CHANGE. This 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 temperature too 373 Kelvin. NO TEMPERATURE CHANGE, This occurs at THE boiling point or THE condensing point (phase change)

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 joules into other units will require a few equalities.

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 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 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 the same 16200 Joules into scientific calories

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

Convert these cal into FOOD calories — AKA kilocalories

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

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 enough for a short time, this stored energy is used up, and you lose some weight as your fat is converted back into energy. If you stop eating for good, you will slow down from the lack of energy and then you will stop living too. Food is energy, energy is required to stay alive. A balance between your life (your energy use) and your energy input is needed to maintain your size.

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 sort of a measure of the kinetic energy the stuff has, how fast or slow the particles are moving or vibrating. Potential energy is something altogether different. Energy can be thermal (heat), electrical, or in many other forms. Energy can be stored in chemical bonds, 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.

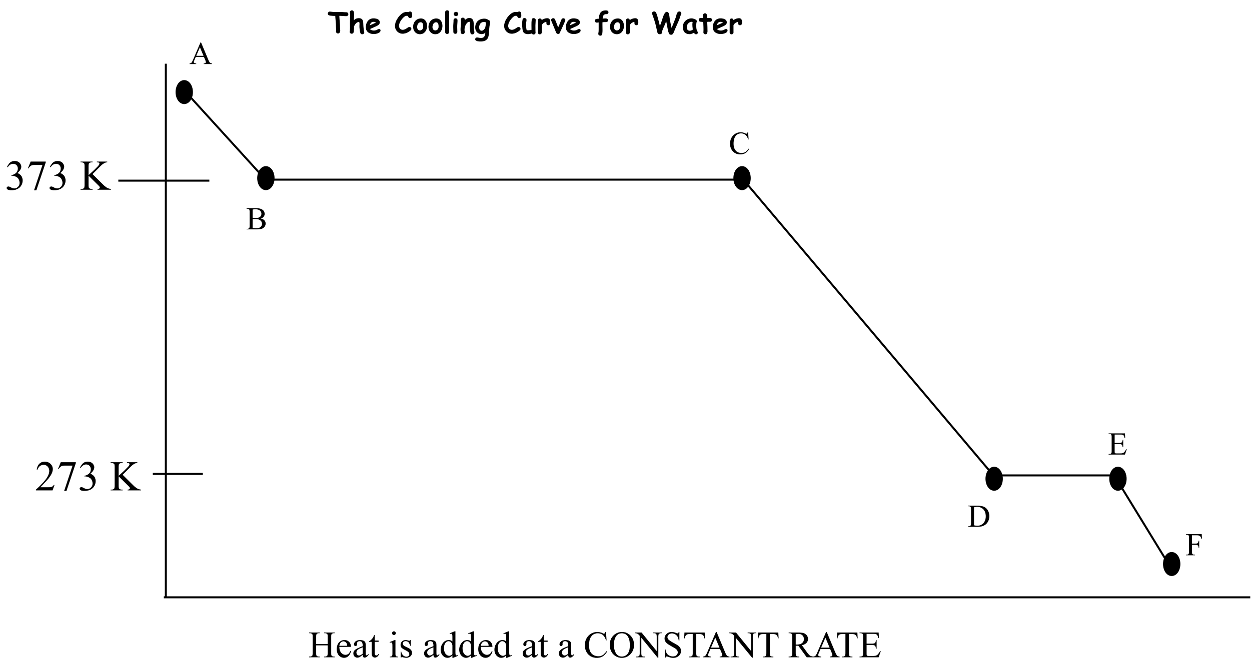
When you want to be cute, you can tell people that you have really low kinetic energy. That means that you are really COOL. They may not understand that, or believe you!

You could tell someone that you have very high kinetic energy too, but they probably with think you are being narcissistic, or that you need new eyeglasses!

It's always more polite to say things like this about others than to describe yourself. Just saying.

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. Freezing happens D to E. The temperature does not change at all.



Point B represents the coldest temperature that steam gas can exist. At that point, the kinetic energy is so low that the gas is almost ready to collapse into a liquid. Any energy removed at that point and the steam begins to condense into liquid. There is no temperature change when this happens. To make one gram of steam condense into 1 gram of liquid water (no ΔT), we must remove the HEAT OF VAPORIZATION energy constant.

For water, the Heat of Vaporization 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 freezing point, which is 273 Kelvin. No temperature change at all.

In reverse, the vaporization of liquid water into steam (no temp change) requires the addition of the H_V energy to every gram of liquid water to make a gram of gas steam. That is, remove 2260 J/g to go from point B to C, or add 2260 J/g to go from C to B, 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 particular 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.
Use the right formula here! H_V not H_F.

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

Now, convert these 12,700 Joules into cal, Calories, or kilo-Joules do this math:

$$\frac{12{,}700\ \cancel{\text{Joules}}}{1} \times \frac{1\ \text{kilo-Joule}}{1000\ \cancel{\text{Joules}}} = 12.7\ \text{kilo-Joules} \quad \textit{(3 SF)}$$

Then, convert the 12,700 joules into scientific (small "c") calories:

$$\frac{12{,}700\ \cancel{\text{Joules}}}{1} \times \frac{1\ \text{cal}}{4.18\ \cancel{\text{Joules}}} = 3{,}038.277\ \text{cal} = 3040\ \text{cal} \quad \textit{(3 SF)}$$

Then, convert these cal to Calories (food Calories):

$$\frac{3040\ \cancel{\text{cal}}}{1} \times \frac{1\ \text{Calorie}}{1000\ \cancel{\text{cal}}} = 3.04\ \text{Calories} \quad \textit{(3 SF)}$$

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. Making steam requires heat (a stove, or a Bunsen burner). A lot of energy is required to make steam, the same "lot of energy" is released when steam phase changes back into a liquid (NO ΔT).

For water,

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

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

COLD phase change (melting or freezing)	$q = mH_F$	$H_F = 334$ J/g
HOT phase change (boiling or condensing)	$q = mH_V$	$H_F = 2260$ J/g

Changing Temperatures

There are times when heat is added or removed from H_2O and it does not change phase. Sometimes water gets colder or warmer but stays liquid water. Temperature can just increase or decrease, but the phase is constant.

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. Remember that the ΔT°C = ΔT K

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

To change the temperature of one gram of pure water by one Kelvin, it takes 4.18 Joules. Add 4.18 Joules to one gram of water, it gets 1 K warmer. Remove 4.18 Joules from one gram of water it gets 1 Kelvin cooler.

Since 4.18 J = 1 cal, you might also say, you must add 1 cal to one gram of water to make it 1 Kelvin hotter. Or, you must remove 1 cal from one gram of water to make it 1 Kelvin cooler.

This amount of energy has a name, it's called the SPECIFIC HEAT CAPACITY CONSTANT, which is 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 mustached folks out of being this confusing!).

$C = 4.18$ J/g·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. GRAMS X KELVIN That formula is on table T.

How many joules of energy will it take to warm up 134 mL of water from 275 K to 297 K?
(cold water becomes warmer)

Let's do this problem. 134 mL of pure water = 134 grams because the density of pure water is 1 gram/mL

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

We will see later that you can solve for any one 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.

How much energy is removed from 475 mL of water if it changes from 306 K to 275 K ?

$q = mC\Delta T$
$q = (475\ g)(4.18\ J/g\cdot K)(31.0\ K)$
$q = 61{,}550.5$ Joules = 61,600 Joules *with 3 SF*

$\Delta TK = \Delta T°C$

A reminder here. Temperature is most often measured in chemistry in centigrade.
Some formulas require Kelvin because it has no negative numbers and no zero either.

The two temperature scales have different "starting points", centigrade starts at the freezing point of water (0°C) and Kelvin starts are the real zero, Absolute Zero (0 K). The temperatures can never be the same numbers.

Ice melts at one temperature, EITHER 273 Kelvin or 0°C.
Same temperature but different numbers on different scales.

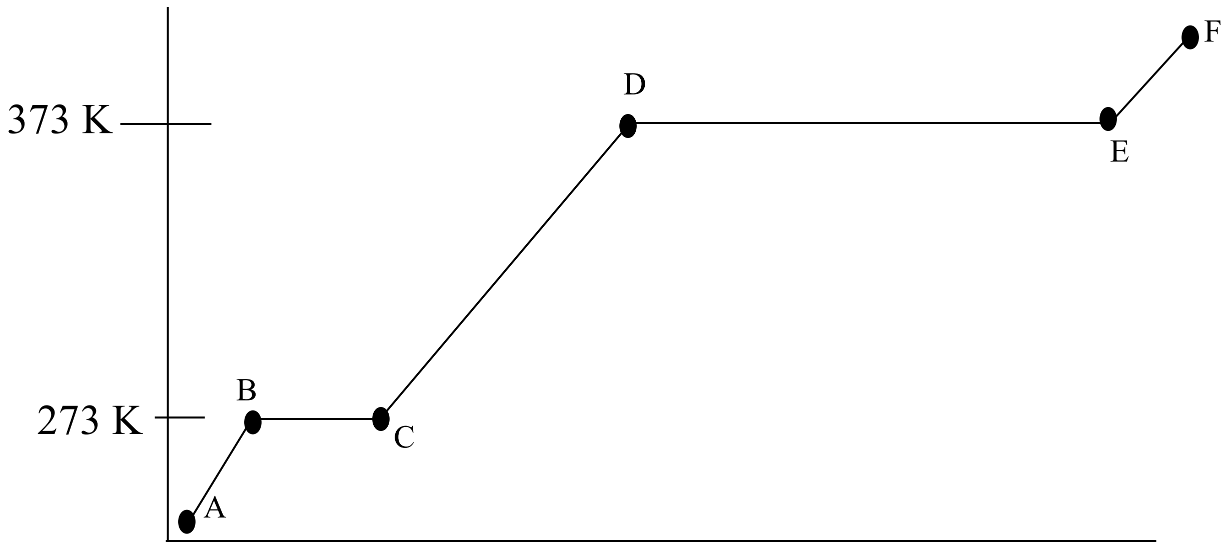
Water boils at EITHER 373 Kelvin or 100°C
Both of these numbers are 100 times bigger than freezing point.

This is important: the CHANGE in temperature in both scales is the same.

Changing in centigrade from 0°C to 100°C means $\Delta T = 100°C$
In Kelvin, the change is from 273 K to 373 K, or the $\Delta T = 100K$

The change in temp can just be switched to K from C, no math required. Just change the units.

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



Heat is added at a CONSTANT RATE

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(script "L" for liquid) 4.18 J/g·K

Liquid water takes that much energy go change each gram hotter or colder by 1 Kelvin.

Ice is chemically identical, but physically different than liquid water, and takes a DIFFERENT amount of energy. So does steam, slightly different too.

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.

This takes two steps: 1st… energy to melt ice into liquid
2nd… energy to warm the water up.

You must calculate both energy requirements, and then 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 up to 373 Kelvin, then measure how much energy to boil away just 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 (say 262 Kelvin)
2. melting it at the melting point (at 273 Kelvin)
3. warming up the water to the boiling point (up to 373 Kelvin)
4. vaporizing it (at 373 Kelvin)
5. then heating up the steam to the final hottest temperature. (maybe 388 Kelvin)

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.

Problem…

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 (this could happen). How much energy is required to do this?
(reminder: when ΔT is 37.0°C, then the ΔT is 37.0 K)

First, melt it, then warm it, finally, sum up all the joules.

Step	Formula	Math
1 Melt the ice this way:	$q = mH_F$	= (657 g)(334 J/g) = 219,000 Joules
2 Warm the water this way:	$q = mc\Delta T$	= (657 g)(4.18 J/g·K)(37.0 K) = 102,000 Joules
3 Sum up the total energy:		219,000 J + 102,000 J = 321,000 Joules

Problem...

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)

1. Condense the steam to liquid water — $q = mH_V = (13.0\text{ g})(2260\text{ J/g}) = 29{,}400$ Joules
2. Cool the liquid water to body temp — $q = mc\Delta T = (13.0\text{ g})(4.18\text{ J/g}\cdot\text{K})(63.0\text{ K}) = 3420$ Joules
3. Sum the total energy: — 29,400 Joules + 3420 Joules = 32,800 Joules

Problem...

If 25.0 grams of ice at 273 Kelvin, and you vaporize it into steam, how much energy is required to do this?

1. First melt the ice to liquid water at 273 K: — $q = mH_F = (25.0\text{ g})(334\text{ J/g}) = 8350$ Joules
2. Then warm the water up to the BP at 373 K: — $q = mc\Delta T = (25.0\text{ g})(4.18\text{ J/g}\cdot\text{K})(100\text{ K}) = 10{,}450$ Joules
3. Now vaporize it at the boiling point: — $q = mH_V = (25.0\text{ g})(2260\text{ J/g}) = 56{,}500$ Joules
4. Finally, sum up all those joules to an answer: — 8350 J + 10,450 J + 56,500 J = 75,300 Joules

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 (The melting point of aluminum is 933 K)

You can melt ice in your hand, producing 334 J/g is not impossible at all. Producing 403 J/g is likely easy for a human to produce (but you can't get close to that 933 K temperature to do it).

The C of Cu, the specific heat capacity constant for copper is 0.391 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 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 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

   $$\Delta T = \frac{836.0\text{ J}}{52.25\text{J/K}} = 16.0\text{ K}$$ 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 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)

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

.

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

FINAL
TEMP = 328.37 K

Rounds to……… 328 K

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 into food 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 to cal, then to Calories.

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?

$$q = mC\Delta T$$

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

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

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

Then:

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

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

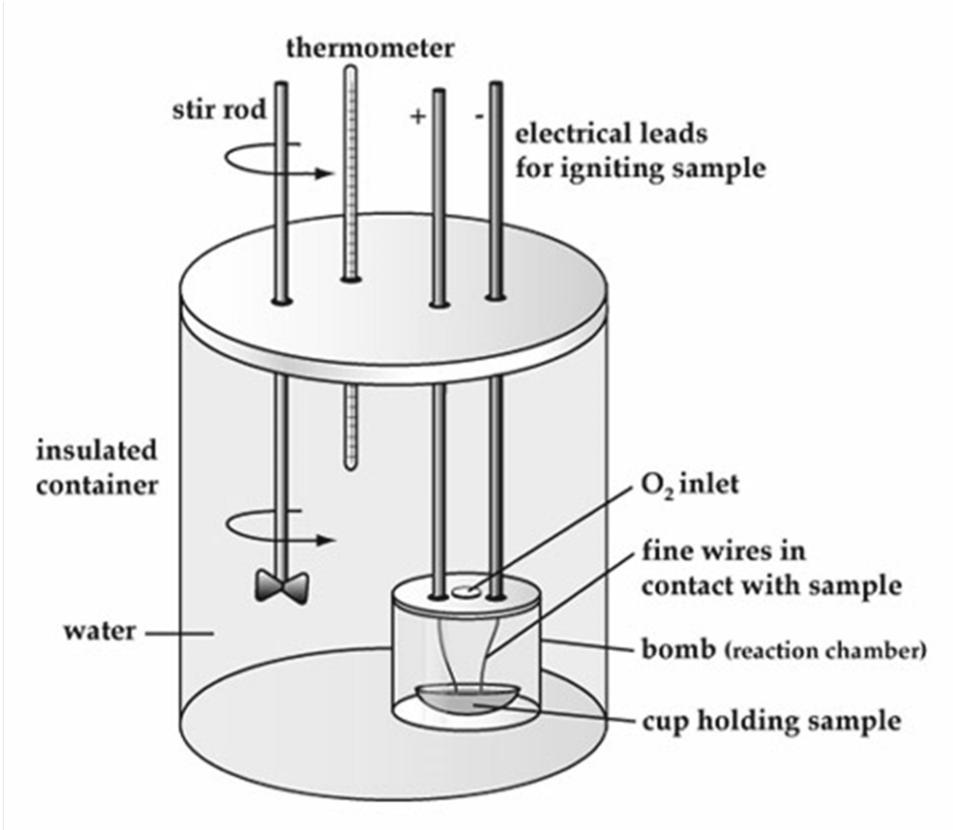
A Bomb Calorimeter looks like this diagram below. We will "make" a calorimeter in the Dorito's Lab, out of a can and a piece of aluminum foil. We will measure the energy gain in the water from the burning chip under it. Most (80% or more) of the heat will be lost, but we will go though the SAME process, and math, to calculate how many Calories are in our Dorito's Chip. I'll shake your hand if you get a better percent error than -80% 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, then 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 energy must be absorbed. You have to add energy to the molecule to break it apart. This is endothermic and has a +ΔH.

The reverse, $Cl + Cl \rightarrow Cl_2$ is the opposite, it releases energy exothermically, it has a –ΔH

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

Ice + ENERGY → Liquid water	Energy breaks the bonds of water molecules apart. Energy is a REACTANT, this is endothermic, +ΔH.
Water → Ice + ENERGY	Energy is emitted to freeze (bond) water into solid ice. Energy is a PRODUCT, this is exothermic, –ΔH.

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". This reaction is exothermic, because $-\Delta H$ 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 and energy can't be negative or positive. The sign only indicates if the energy is a product in an exothermic reaction, or a reactant in an endothermic one.

The proper balanced THERMOCHEMICAL reaction would be written this way:

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

The 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)

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

Solve for X

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

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

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

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

Solve for X

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

$X = 8176.44 \text{ kJ} = 8180 \text{ kJ}$ *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 energy can be converted to other units.

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

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 exothermic.

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

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

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

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.

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