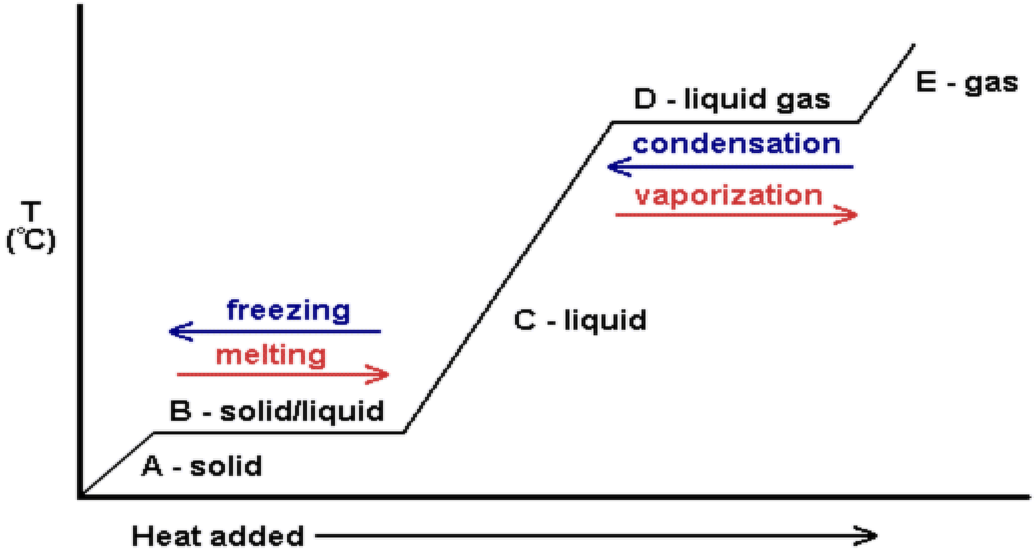
Chapter 11: Heat and Energy Diary

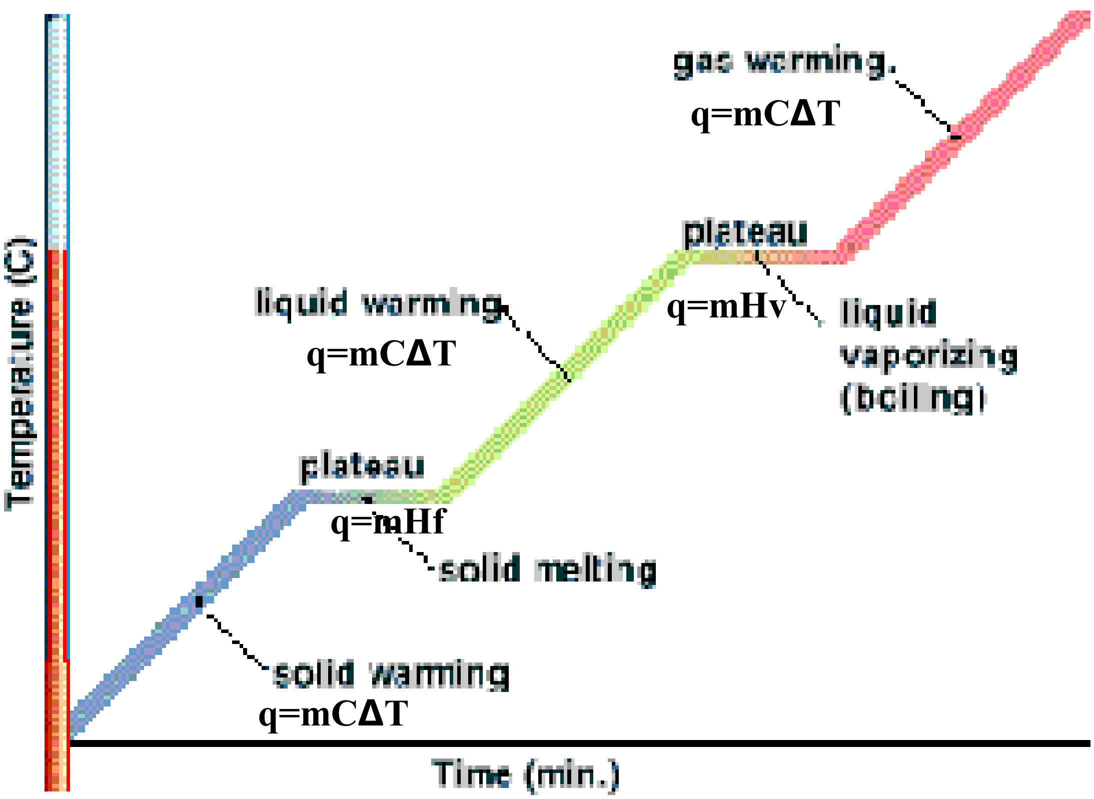
Thermo-chemistry is the study of how heat (energy) plays a role in chemical reactions and physical changes (phase changes). Heat can be released or absorbed by chemical reactions (exothermic or endothermic). All substances can go through different phase changes. The exact amounts of energy required to be absorbed or released is measurable and predictable. There are the three heat formulas listed on your reference table to become familiar with, and Table B, which are the three physical constants for water.

Vocabulary

calorie	The amount of energy required to increase the temperature of exactly 1.0 grams of pure water by exactly 1.0°C
Calorie	This is a “food calorie”, equal to 1000 calories (also called kilo-calorie)
Joule	The metric unit of energy. 4.18 Joules = 1.0 calorie
Kilo-Joule	1000 Joules
Specific Heat Capacity	It is the energy required to make 1.0 grams of a substance change temperature by 1.0°C. This is a constant for any substance. (“C” value)
Specific Heat	This is the energy change required to change a substance by 1.0°C. It is not a constant, because it concerns a sample of stuff (for example, NOT 1.0 grams of water, but a whole pot of water). It’s tied to the specific heat capacity of the substance AND how much of it you have.
Specific Heat for Water *from table B*	“abbreviated” from specific heat capacity= 4.18 J/g·C° which means it takes 4.18 Joule of energy added to make one gram of water 1.0°C hotter (or that many joules need to be removed to make one gram of pure water 1.0°C colder). Don’t forget that 4.18 J = 1.0 calories.
Heat of Fusion for Water (H_F) *from table B*	The amount heat needed to turn 1.0 grams of solid ice water from 0°C to a liquid at 0°C. NO TEMPERATURE CHANGE. All substances have a H_F Also the reverse, the same amount of energy would be removed to freeze 1.0 grams of liquid water into solid ice. The constant for water H_F = 334 Joules/gram
Heat of Vaporization for Water (H_V) *from table B*	The amount heat needed to turn 1.0 grams of liquid water from 100°C to a gas at 100°C. NO TEMPERATURE CHANGE. All substances have a H_V Also the reverse, the same amount of energy would be removed to condense 1.0 grams of gas vapor into liquid water. The constant for water H_V =2240 Joules/gram
Calorimeter	A device used to measure heat loss by a sample of matter (or food). Also called a bomb calorimeter. These are highly sophisticated and exact, unlike our high school “calorimeters” made up of 2 styro-foam cups.

General Heating Curves





General Cooling Curve

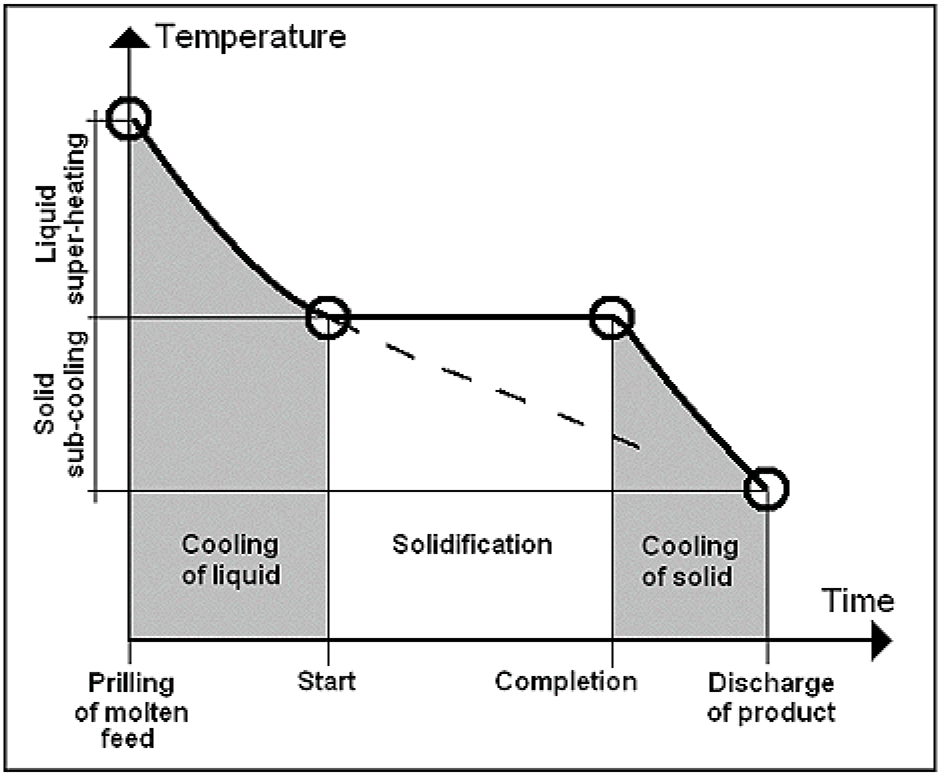


Fig.3: Solidification Curve

The **BASIC HEAT FORMULA**, is $q = mC\Delta T$ (Reference Table T)

Where q is the amount of heat in joules, m is the mass in grams of the substance you are examining, C is the specific heat capacity, which is a constant for every substance, with the unusual unit of J/g·°C
ΔT is read as DELTA-T, and means the change in temperature, in C°. Recall that temperature=Average kinetic energy(movement of molecules)

This formula is used when ever there is a temperature change (NOT for phase changes). Please refer to the diagram on the prior pages and note the location of the formula on the curve. You can solve for any part of the equation, energy, mass, specific heat capacity or temperature change, if you know the rest of the numbers.

Example: How many grams of water can change temperature from 23.4°C to 18.2°C with the removal of exactly 499 Joules of energy?

$q = mC\Delta T$ Substitute in what you have, and solve for the unknown.

499 J = (m) (4.18 J/g·°C) (5.20°C) ***[now, solve for m]***

$\frac{499 \text{ Joules}}{21.736 \text{ J/g}}$ = m = 22.957... grams of water

m = 23.0 grams of water (with 3 sf)

The COLD PHASE CHANGE HEAT FORMULA,

$$q = mH_F$$

This is used only during the cold phase change, when $\Delta T = 0$. The cold phase change occurs during melting or freezing. Using the $q = mC\Delta T$ formula when there is no temperature change is silly, it would always work out to zero. Please refer to the diagram on the prior pages and locate the formula on the phase change plateau. During this plateau both solid and liquid states are in equilibrium. All energy added is affecting the potential energy of the system and the making or breaking of intermolecular attractive forces between them.

The H_F or heat of fusion is the energy associated with "fusing" of a liquid into a solid, or "unfusing" a solid into a liquid. For our purposes we will be referring mainly to the heating curve of water.

The formula has q again, the amount of heat in Joules, equal to m again, the mass in grams, times, a new constant, H_F. This constant is 334 Joules/gram, which is how much energy it takes to melt 1.00 grams of water from solid to liquid with NO TEMPERATURE CHANGE, from 0°C ice to 0°C liquid.

Example :How much energy is needed to melt 722 grams of ice at 0°C to liquid water at the same temperature?

$q = mH_F$ so, fill in what you know,

q = (722 grams) (334 J/g) = 241,000 Joules (with 3 sf)

Example: How much energy is needed to freeze 722 grams of liquid water at 0°C into solid ice at 0°C?

$q = mH_F$ so, fill in what you know,

q = (722 grams) (334 J/g) = 241,000 Joules (with 3 sf)

In thermo-chemistry, the EXACT same amounts of energy affect water in BOTH TEMPERATURE DIRECTIONS.

It doesn't matter if water is cooling or heating, the CONSTANTS are CONSTANT.

The HOT PHASE CHANGE HEAT FORMULA,

$$q = mH_V$$

This formula is used only during the hot phase change, when $\Delta T = 0$. Using the $q = mC\Delta T$ formula if there is no temperature change is silly: it would always work out mathematically to zero. Please refer to the diagram on the prior pages and locate the formula on the phase change plateau. During this plateau both liquid and gaseous states are in equilibrium. All energy added is affecting the potential energy of the system and the making or breaking of intermolecular attractive forces between them.

The H_V or heat of vaporization is the energy associated with "vaporizing" of water into a gas, or the reverse, un-vaporizing it "condensing" a gas into a liquid.

The formula has q again, the amount of heat in Joules, equal to m again, the mass in grams, times, a new constant, H_V. This constant is 2240 Joules/gram, which means it takes 2240 Joules of energy to vaporize 1.0 grams of water from liquid to gas with NO TEMPERATURE CHANGE, from 100°C ice to 100°C liquid.

Example: How much energy is needed to vaporize 722 grams of water at 100°C to steam at the same temperature?

$q = mH_V$ so, fill in what you know:

q = (722 grams) (2240 J/g) = 1,617,280 Joules

1,620,000 J (with 3 sf)

Example: How much energy is needed to condense 722 grams of steam at 100°C to liquid water at the same temperature?

$q = mH_V$ so, fill in what you know:

$q = (722 \text{ grams})(2240 \text{ J/g}) = 1{,}617{,}280$ Joules

1,620,000 J (with 3 sf)

In thermo-chemistry, the EXACT same amounts of energy affect water in BOTH PHASE CHANGE DIRECTIONS - increasing or decreasing energy. It doesn't matter if H_2O is freezing or melting, or vaporizing or condensing, the CONSTANTS are CONSTANT.

Let's look at the answers to the examples again.
Whether you freeze or melt, during the cold phase change, the amount of energy to get a certain amount of water through the phase change is the same (except that energy is removed to freeze and it's added to melt).

with a constant mass of water 722 grams of water	energy needed to be removed to 722 grams of H_2O	energy needed to be added to 722 grams of H_2O
the COLD phase change	241,000 Joules (freeze)	241,000 Joules (melt)
Hot phase change	1,620,000 J (condense)	1,620,000 J (vaporize)

During the hot phase change, it also takes the SAME AMOUNT of energy in either temperature direction, condensing removes the same energy it took to vaporize the 722 grams of water.

FINALLY NOTE: There is a major difference between the hot and cold phase changes here concerning energy. The same 722 grams of H_2O: Melting took ONLY 241,000 Joules compared to vaporizing which took a whopping 1,620,00 Joules.

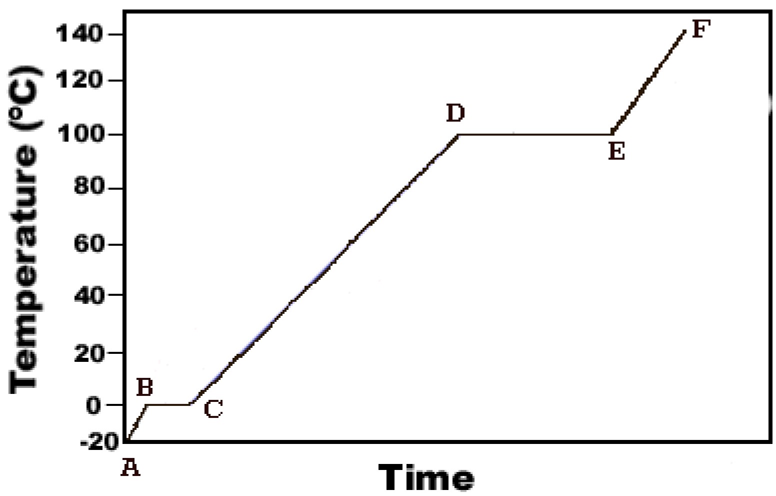
It's almost a 7X difference in energy. That is clear on TABLE B, the heat of fusion is only 334 J/g while the heat of vaporization is 2240 Joules/gram. It takes much more energy to overcome the intermolecular forces of attraction holding the liquid together to convert it into a gas.

Heat Formulas and Cooling Curves or Heating Curves

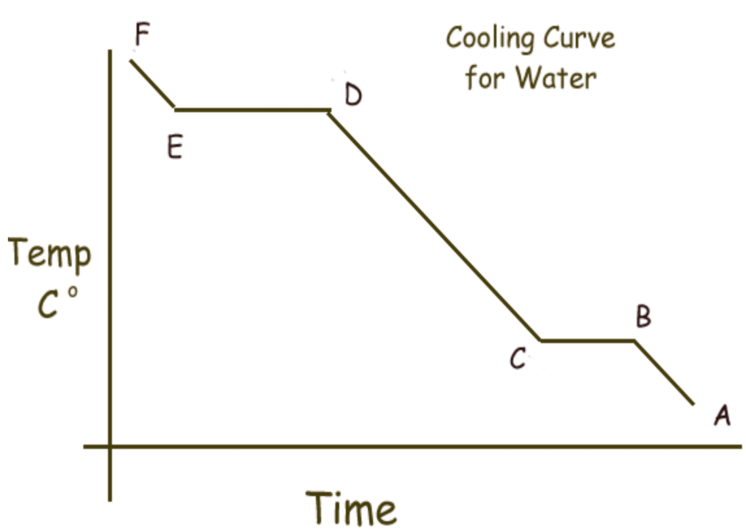
Below are diagrams of both the heating curve and cooling curves for water. The cooling curve is the REVERSE, with the **same corresponding points at ABCDEF**. The points BC represent the cold phase change while the points EF represent the hot phase change. CD is the liquid water phase. AB is solid ice only, and EF is totally the gas phase. This chart outlines all values for Temp, KE, PE, and formulas needed to do any energy calculations for that section of each graph.

NOTE: TEMP & KINETIC ENERGY ALWAYS CHANGE TOGETHER, POTENTIAL ENERGY IS DIFFERENT

Heating Curve for Water



Cooling Curve for Water



Choosing the right formulas... If your problem concerns just one part of the heating or cooling curve, just use the one correct formula...

AB	ice changes temp, getting hotter or colder	$q = mC\Delta T$
BC	Cold phase change Liquid to solid or solid to liquid	$q = mH_F$
CD	water temp changes, getting hotter or colder	$q = mC\Delta T$
DE	Hot phase change Liquid to gas or gas to liquid	$q = mH_V$
EF	steam changing temp, getting hotter or colder	$q = mC\Delta T$

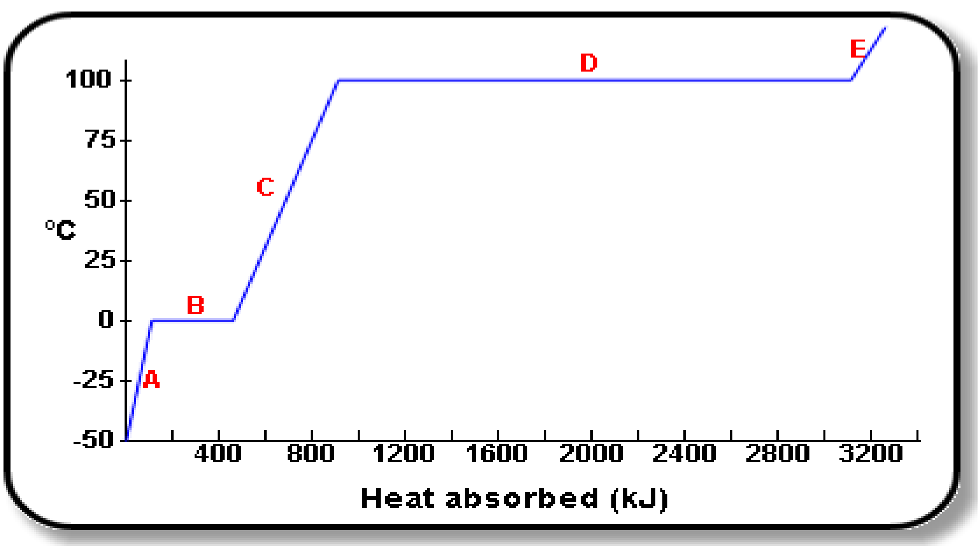
If you have a problem that phase change and a temperature change, combine more than one step, and combine the total number of joules together to get one answer.

Multiple Phase Change problems

Calculate the amount of heat required to completely convert 50 g of ice at -10 ºC to steam at 120 ºC.

The heat taken up in the complete process is the sum of the heat taken up in each stage.

1.Heat taken up heating the ice from -10 °C to the melting point, 0 °C.	mass of water x specific heat x temperature change = 50g x 2.09 (J g^{-1} °C^{-1})x 10 °C = 1,045 J
2. Heat taken up for converting ice at °C to water at °C.	mass of water x latent heat of fusion = 50g x 334 (J g^{-1}) = 16,700J
3. Heat taken up heating the water from 0 °C to the boiling point, 100 °C.	mass of water x specific heat x temperature change = 50g x 4.18 (J g^{-1} °C^{-1})x 100 °C = 20,900 J
4. Heat taken up vapourizing the water.	mass of water x heat of vaporization 50g x 2260 J g^{-1} = 113,000 J
5.Heat taken up heating the steam from 100 °C to 120 °C.	mass of water x specific heat x temperature change = 50g x 2.09 (J g^{-1} °C^{-1})x 20 °C = 2,090 J
The sum of these is:	1,045 J +16,700J + 20,900 J+ 113,000 J +2,090 J = 153,735 roughly 154 kJ



The diagram above shows the uptake of heat by 1 kg of water, as it passes from ice at -50 ºC to steam at temperatures above 100 ºC, affects the temperature of the sample.

A: Rise in temperature as ice absorbs heat.
B: Absorption of heat of fusion.
C: Rise in temperature as liquid water absorbs heat.
D: Water boils and absorbs heat of vaporization.
E: Steam absorbs heat and thus increases its temperature.

Rather than doing more calculating, <u>let's do a lot of thinking</u>. Don't do the math, just tell what formulas are required to determine the answer. Write, cold phase change, basic heat, or hot phase change formulas. Answers at the bottom in red.

Example 1: Water at 2.543°C is warmed to an even 100°C but it remains liquid.
How many joules are needed to be absorbed?

Example 2: Steam at 100°C is cooled all the way to 0°C liquid.
How many joules of energy are released?

Example 3: Steam condenses from 100°C until ice forms at 0°C.
How many joules are released?

Example 4: Ice at 0°C is melted then warmed to 22.5°C which is room temperature.
How many joules are absorbed?

Answers:
Ex 1: just the basic heat formula.
Ex 2: Hot phase change, basic heat formula
Ex 3: hot phase change, basic heat, and cold phase change formulas.
Ex 4: cold phase change, basic heat formula.

Table I: The Heats of Reaction

Table I shows us 25 exothermic and endothermic reactions, and the exact ΔH associated with each one, at standard pressure and room temperature 298K. "ΔH' is another way of saying "change in heat (q)." As you may recall, during an exothermic process heat I energy is released and during an endothermic process heat energy is absorbed.

The balanced chemical reactions have mole ratios associated with them. We also can include the energy gained or released in kilo-Joules for these chemical reactions

Look at the first reaction (here), the combustion of methane:

$$CH_4 \; + \; 2O_2 \; \text{—->} \; CO_2 \; + \; 2H_2O \; + \; 890.4 \text{ kJ}$$

This means that one mole of methane needs 2 moles oxygen to form one mole carbon dioxide plus 2 moles of water, and it releases 890.4 kJ of energy in an Exothermic reaction. A negative "delta H" (-ΔH) means heat is given off.

Recalling from Stoichiometry, the mole ratios work mole:mole only. If you have grams, liters, or numbers of particles, you must FIRST convert to moles to do the Stoichiometry. NOW the energy is included into this ratio.

The WHOLE Mole Ratio for this reaction is

1mole:2mole:1mole:2mole:*890.4 kJ*

The energy is now included in the mole ratio.

Table I lets use see exothermic reactions with their negative ΔH, and also the endothermic reactions with their positive ΔH. Though this is just a list of 25, there are thousands of thermo-chem reactions in the universe.

Note that from a stoich point of view, in reaction 3 on table I, as the octane combusts, it starts with 2 moles of octane to produce the 10943 kJ of energy release. The mole ratio there is 2:25:16:18:*10943kJ*

The **most endothermic** reaction on table I is
$C_2H_{4(G)}$ forms, and absorbs 227.4 kJ of energy.

The **most exothermic** reaction by far is when octane combusts and releases energy. It takes 2 moles of octane produce 10943 kJ; one mole octane would release half that amount of energy or 5471.5 kJ is released.

Imagine a heating pad put onto your achy elbow. The pad undergoes a chemical reaction that is exothermic (-ΔH), releasing heat. The heat, or kinetic energy goes from the pad into your elbow. You feel good.

If you would rather use a cold pack, then those reactions undergo endothermic reactions (+ΔH), and the pack feels cold. Place it on your twisted ankle. The pad does not "send" cold to your ankle, rather the heat from your ankle is transferred to the pack.

Example: How much heat is absorbed when 2.50 moles of NaOH is dissolved into water?

The ΔH for this reaction on Table I for NaOH dissolving into water is +44.51 kJ. That is for ONE MOLE of sodium hydroxide. The mole ratio is when one mole NaOH is dissolved into water it absorbs that many kilojoules of energy. You have 2.50 moles here. The ratio would be this way:

1 $\frac{\text{1 mole NaOH}}{\text{44.51 kJ}}$	2 $\frac{\text{2.50 moles NaOH}}{\text{X kJ}}$	3 Cross multiply and solve for X kilojoules of energy
4 X = (44.51kJ)(2.50)	5 X = 111.275 kJ	6 With 3 SF: 111 kJ energy absorbed