

Reading Guide

for the Heat and Temperature unit

Chemistry Chapter 5: Heat of Reaction

This chapter is fairly long, dealing with a number of rather thorny topics. It is written to lay a foundation for important concepts in thermal chemistry, which is not really our focus in this unit (we'll return to it later in the course). The part that is of interest to us now is the explanation of the nature of heat and how thermal energy is shared between different objects in contact with each other.

pp. 85–87. This gives the very briefest explanation of what work is (we fortunately spent more time on that earlier), and the concepts of kinetic and potential energy. I hope you already understand this, but if it isn't perfectly clear to you, read these pages carefully.

pp. 88–89. This doesn't really explain what temperature is, but it does define **heat** well. Note that heat and temperature are related, but not the same thing!

pp. 90–91. Now we get a better explanation of what **temperature** is, along with a confusing quantity called **internal energy**. The important lesson here is that temperature is related to the kinetic energy of molecules and atoms moving about. It's kinetic energy on a microscopic scale.

(Interestingly, the explanation of thermal processes in terms of molecular-scale motions was developed *after* the thermal properties of directly-observable processes was worked out. Understanding thermal physics was really difficult then!)

p. 92. Explains how J. P. Joule determined the conversion between energy and heat (the mechanical equivalent of heat). That's interesting for a History of Science perspective, but the important concepts introduced on this page are

- **heat capacity**
- **specific heat**

Try to define these terms.

pp. 93–95. This discusses how to measure heat. Recall that heat is a molecular-scale kinetic energy transfer, so it can't be measured or calculated by the energy formulas we already know. It must be measured indirectly, from temperature changes and known specific heats. These pages explain the principles of how it is done.

pp. 96–103. This does not really apply to what we'll be studying in this unit, so skip over it.

p. 104 This is a quick lead-in to chapter 6.

Chemistry Chapter 6: Matter in a State

You already read part of this chapter last week for the “fluids” unit. This week, though, we have a different focus, so I’ll ask you to go back and thoroughly consider parts that last week I only had you skim. The important concept to us this week is that molecules in different states of matter (solid, liquid, gas) have different potential energies, and this affects their thermal behavior.

pp. 105–108. Skip.

p. 109. Skim. You read this last week, but it makes sense to refresh it in your mind. This explains why it takes energy, and a temperature increase, to melt solids and boil liquids.

pp. 110–114. Skip.

p. 115. Briefly explains some of the properties of liquids, including thermal expansion.

pp. 116–117. Explains how evaporation adsorbs heat over and above anything needed to simply raise the temperature of water. Don’t worry about the “ $\Delta H > 0$ ” on p. 116; that just means that a liquid absorbs heat when it evaporates.

- Why does a liquid absorb heat when it evaporates?

pp. 118–119. Hang on through this. It isn’t terribly important for this unit, but the concept of **equilibrium** will be critical when we begin studying chemistry in a few weeks.

- What is **vapor pressure**?
- What happens when a liquid boils?
- How is boiling different from evaporation?

pp. 120–125. Skip this. We don’t have the time to do it justice.

pp. 126–127. More on the heat transfers associated with phase changes. Read this carefully, and make sure that you understand what is going on in the graph at the top of page 127!

- Why does the temperature of constantly-heated water stall at 0 and 100 °C?

p. 128. You don’t need anything here for this class.

Chemistry Chapter 10: Chemical Thermodynamics

pp. 191–193. Read this carefully; it's deceptively simple. Answer for yourself: what is the main message of these three pages?

194–195. The explanation in the book here gets tricky. The main point is that, just as natural processes tend to spread matter out, they also tend to spread energy out, and for the same reason.

This book uses a model to describe how energy flows in physical and chemical (that is to say, *all*) processes. It takes a little thinking to get through this, so if you don't have time for heavy thinking right now, do something else and come back when you do.

Set aside a block of time, remove all distractions, take a deep breath, and carry on.

The model the book uses to describe the process of apportioning energy is a stack of horizontal lines. The lines represent different energy levels of the system, that is, basically different places where energy can be. The height of a line denotes the temperature that results if the level is filled. Lower levels fill before higher levels, just as water goes to the bottom of a glass if it isn't already occupied. Some systems are shown with energy levels spaced far apart, and others with energy levels close together. This spacing indicated how many options are available: lines close together mean that there are many options within a given energy range; lines far apart mean that there are very few places for energy to go in a given range. Gray shading between lines indicates that the energy levels are occupied. Energy will only move from one system to another if it has more freedom after the move, just as matter spreads out when it moves. Also, energy will not move to a system in which it would have to occupy higher levels than it already does, any more than water will. (In plain English, heat flows from areas of high temperature to areas with low temperature, but not the other way around.) So, back to the book!

p. 195. Don't worry about the formula $\Delta S = q/T$. Pay attention instead to the stuff on the right. Adding heat to a system may simply mean that energy goes into higher levels; or, in special situations, it may mean that more levels open up without a change in temperature! The example given in the book is for ice to melt. Does it make sense that energy has more options in liquid water than in solid ice?

pp. 196–197. Skip this. It costs more work to follow than any benefit you can get from following it.

p. 198. Read this. Although it does not really explain why the changes described increase entropy, it should be reasonable enough to you.

p. 199. The key idea on this page is the stuff in the middle in the box, the second law of thermodynamics. Ignore the rest of the stuff about entropy tables.

pp. 200–201. Introduces the concept of the Gibbs function. Skip this. I'll try to walk you through what comes next.

p. 202. In the diagrams on this page, the book breaks down the two different ways that entropy can be changed in a physical or chemical process. It refers to them as ΔH and $-T\Delta S$, which are specific thermodynamic terms. ΔH refers to how much the process allows energy to spread out, and $-T\Delta S$ to how much it allows matter to spread out.

The first diagram represents what happens if $\Delta H > 0$ for a process: the lowest available energy level becomes higher. By itself, this will not favor the process. (Water doesn't flow uphill!)

The second diagram represents what happens if $-T\Delta S < 0$: the energy levels become closer together, so that there are more levels available at a given temperature.

p. 203. The stuff on this page is key, if only you can make sense of it. It shows four possible combinations of ΔH and $-T\Delta S$ changes that can occur for any process, in which either quantity can increase or decrease.

Top left: Matter spreads out and energy spreads out. This is a no-brainer. This sort of process is favorable, like marbles rolling down the steps and under the refrigerator.

Top right: Matter becomes more restricted and energy becomes more restricted. This is also a no-brainer. This sort of process is never favorable, and never, ever happens. Ever.

Bottom left: Matter spreads out, but energy becomes more restricted. What happens now? As you might expect, if matter spreads out enough, it will outweigh the restrictions on energy and the process will occur. On the other hand, if energy becomes very restricted and matter is freed up only a little, the process will be unfavorable. It turns out that the temperature affects whether such a process occurs! The diagram shows two such situations, one at low temperature (not much gray in the stack), and one at high temperature (lots of gray in the stack). When there is a lot of energy in the system (high temperature), energy can spread out by occupying the more numerous levels, even though the bottom is higher. At lower temperatures, this benefit does not exist.

Bottom right: Energy spreads out and matter becomes more restricted. This is the reverse of the process on the bottom left, and not surprisingly is favored by low temperatures and disfavored by high temperatures. Can you make sense of the diagrams to see why?

pp. 204–208. Skip this. You have worked hard enough. It will be a month before we will worry about the concepts here.

Wrap-Up

Tolstoy famously wrote, in the opening of *Anna Karenina*, “Happy families are alike; every unhappy family is unhappy in its own way.” (Translated by Constance Garnett.) I claim that this is an analogy for chemical thermodynamics. Why?