Chapter 4 – Properties of Matter

4.1 Properties of Substances

Recall from Chapter 3 that there are two basic classes of properties: physical and chemical. Physical properties are ones in which the identity of the material remains unchanged, while chemical properties are usually referred to as chemical reactivity because they describe the tendency of the material to change.

Classic physical properties would include color, melting and boiling temperatures, hardness (of a solid), viscosity (flow rate) of a liquid, among many others. Chemical properties/reactivity are almost always discussed in the context of a second substance. For example, sodium has high chemical reactivity with respect to chlorine gas, with which it reacts violently to produce sodium chloride (table salt). In contrast, neon gas does not react with chlorine under any circumstances. In the middle, most transition metals react slowly with chlorine at ambient temperature to form a MCl_2 product. Functionally, every material has an enormous number of chemical properties, but as chemists learn reasonably quickly there are a lot of patterns to chemical reactivity that simplify learning about them.

One point the book makes requires refining. No two substances have identical physical properties nor do they have the same chemical properties. (It's a point of grammar, but the book's sentence, p. 70, suggests that a substance could have identical physical properties or chemical properties, but not both.)

4.2 Physical and Chemical Changes

Normally, measuring a physical property does not change the substance in any way. If determining the color of a solid one may look at it for a qualitative assessment or measure the wavelengths of light either absorbed or reflected by the substance. However, in either case the substance is the same before and after examining it. Measuring a few physical properties does result in a change in the substance, but not its identity. The most common example of this is melting and boiling. When measuring the melting/freezing temperature of a substance, its phase changes, so that water becomes ice, but it remains H_2O.

In contrast, chemical changes always result in the change in the identity of a substance. The very words, chemical reaction, say that the original materials are changed. The major point of this section is the introduction of the chemical equation. In chemical equations species that react and the products from them are written out in a format very similar to that of a mathematical equation. For example, the burning of methane looks like this:

$$CH_{4\,(g)} \; + \; 3\,O_{2\,(g)} \; \rightarrow \; CO_{2\,(g)} + \; 2\,H_2O_{(\ell)}$$

There is a lot of information in that equation. First, if it were read out loud, the equation would be: one molecule of methane gas reacts with 3 molecules of oxygen gas to produce one molecule of gaseous carbon dioxide and two molecules of liquid water. From this equation, we obtain three pieces of information:

1) The number of molecules (particles) of each species either reacting or produced,
2) the physical state of each of the materials participating in the reaction, and
3) the number and type of each atom present. (Notice that the number and type of each atom is the same on either side of the arrow. This will be important later in this course.)

Sometimes symbols are placed above or below the arrow to provide additional information. For example, the Greek letter capital delta, Δ, represents heat and hν represents light. Seeing a Δ above or below the arrow indicates that the reaction mixture was heated and hν means that the reaction required light to proceed (e.g. photosynthesis). Occasionally the solvent in which the reaction is conducted is placed on the arrow. For example, when hydrogen and oxygen react to form water, heat must be added to get the reaction started, so

$$2\,H_{2\,(g)} \; + \; O_{2\,(g)} \xrightarrow{\Delta} 2\,H_2O_{(\ell)}$$

Most textbooks do not discuss the following point, but it is an important one. As noted at the beginning of this section, there is a similarity between math and chemistry equations. As you will see in CHM 211/212, they can be added just like in algebra. For example, consider the following reactions where everything is a gas:

$$2\,NO_2 \rightarrow NO_3 + NO$$

$$\underline{NO_3 \rightarrow NO + O_2}$$

$$2\,NO_2 \rightarrow 2\,NO + O_2$$

If the chemicals were replaced with letters and the arrow with an equals symbol, the result would look just like a traditional algebra problem.

$$2x = y + z$$

$$\underline{y = z + w}$$

$$2x = 2z + w$$

Recognizing the similarity between chemical reactions as expressed as chemical equations and mathematical equations can be very helpful.

4.3 Learning to Solve Problems

This was already covered in detail in the Chapter 2, Section 6: Dimensional Analysis: A Problem-Solving Method notes. You should review it.

4.4 Energy

To understand why reactions occur or don't, understanding energy is extremely important. Energy is the capacity of matter to do work or transfer heat. (The second part of the definition is **not** in your book, but is an important part of energy.) Energy exists in many forms, the most important of which in chemistry are heat, light, and electricity.

Energy can be categorized as either potential or kinetic. Potential energy is frequently called the "energy of position," and represents stored energy. An object on a shelf has more potential energy than the same object on the floor. i.e. the object on the shelf has the 'potential' to do more work than the one on the shelf as a function of its position further from the center of the earth. Kinetic energy is the energy is the energy of motion. When the object on the shelf falls to the floor, the difference in their potential energies becomes (equals) the kinetic energy of the object. Consider the following cycle. An object on the floor has a certain potential energy and no kinetic energy because it is stationary. Lift the object to the shelf. Lifting the object requires the expenditure of energy. That energy is kinetic energy and is transferred into the object. When the object stops moving upward it has a new potential energy, one that is higher than when it was on the floor, and one that is equal to the original potential energy plus the kinetic energy used to raise the object. When that object falls back to the floor, the kinetic energy added to the object is released to do work and the object returns to the original potential energy at the outset of this cycle.

This thought experiment demonstrates that energy can be interconverted, but it cannot be created or destroyed. The text in blue is one of the most important laws in science, the first law of thermodynamics. Heat is the most important form of energy to chemists. We frequently warm reactions to speed them and cool to slow them. Reactions that happen naturally, almost always release heat into the environment. For that reason, as your book notes, sometimes heat is written as a product in a chemical reaction, but that is very uncommon. If a lot of heat is produced and the reaction occurs rapidly, the reaction mixture will increase in temperature. If small amounts of heat are produced or the reaction is very slow (e.g. rusting), the energy disperses so slowly that it is almost impossible to detect warming.

4.5 Heat: Quantitative Measurement

To discuss heat energy, we must have a unit to quantify it. In chemistry, that unit is the joule (J), while in everyday life it is the Calorie. That the "C" in Calorie is capitalized is very important because there is also the related unit "calorie."

The joule is used because it is a metric unit and actually has subunits within it. Formally,

$$1\ \mathrm{J} = \frac{1\ \mathrm{kg{\bullet}m^2}}{\mathrm{s}^2}$$

If this looks strange, consider Einstein's famous mass-energy equation, $E = mc^2$, where "c" is the speed of light. Do you see how that units are the same as in the joule? Using $\mathrm{kg{\bullet}m^2/s^2}$ is cumbersome, so renaming it to a combined unit is much easier to work with.

The calorie is the old chemical energy unit and is related to the joule by the equality

$$4.184\ \ \mathrm{J} = 1\ \mathrm{cal}\ (\text{exactly})$$

with the calorie based on English units. You probably won't ever work with the calorie, but should be familiar with it. As your book notes, joules and calories are very small quantities of energy and working in the macroscopic world a larger unit would be helpful. Rather than make a new, larger unit, kilojoules (kJ) and kilocalories (kcal) are the usual units used in chemistry.

You are familiar with energy in food being provided in Calories (cap C).

1 Calorie = 1000 calories = 1 kcal

Thus, the reason for introducing a unit that won't be used is that there is a unit for you to relate the units of the world you live in, to the units of the scientific world. In short, a bowl of breakfast cereal is about 200 Calories, which is about 850 kJ. This provides a frame of reference to calculations and whether the calculated answer is reasonable.

Of interest to us is how much energy is used or released when a substance changes temperature. As you know intuitively, different substances retain heat differently. For example, if a 1-ounce steel ball and a one-ounce plastic ball were put in boiling water and you were forced to have one placed in your hand, you would certainly pick the plastic ball. Both balls would come out of the water at the same temperature, but the plastic would cool down much faster. Likewise, you know that a metal pot heats more rapidly than the water in it, even if they weigh the same. The innate ability of a substance to retain heat is called its <u>specific heat</u> and is formally defined as the energy required to raise the temperature of one gram of the substance by one degree Celsius. The larger the specific heat, the more heat the substance retains at a particular temperature. The result is that substances with higher specific heats require <u>**more**</u> energy to increase in temperature (i.e. it is more difficult to warm them and they cool more slowly too). Each substance has a specific heat with units J/g•ºC. .

In general,

heat (J) = (mass)(specific heat)(temperature change)

Example 4.4 in the book is: Calculate the specific heat of a solid in J/g•ºC if 1638 J raises the temperature of 125 g of it from 25.0 ºC to 52.6 ºC.

<u>Dimensional Analysis – Part II</u>

The original rules for doing dimensional analysis won't exactly work here because you now have 3 pieces of given information, instead of 1. How do you proceed in this situation?

1) As originally, collect the given information.
2) Write the 'to be calculated' value with an equal sign to its right.
3) Select one of the pieces of given information and place it next to the equal sign. The way to

choose which datum is to look at the unit on the left of the equal sign and see if one of the given units is the same as one of the given. If so, place it to the left of the equal sign. If it is the inverted value (e.g. m^{-1} vs. m), then place it to the right in the form $\frac{1}{m}$. Do not spend a lot of time on this decision because there is an easy fix if you choose incorrectly. If you put the wrong unit in the numerator, then your answer will be inverted. All you have to do is put the entire equation to the right of the equals sign in square brackets and raise it to the -1 power. Then use the 1/x button on your calculator. Another common mistake is to forget that there is a difference or sum. For example, in this problem the temperature changes so you must plug in the temperature change, rather than either starting or ending temperature.

4) Add the remaining units similarly, so that units that should be in the numerator of the final unit are multiplied and units that should be in the denominator of the final unit are divided. Look for additions and subtractions, as in the case for temperature changes.

With this as background: Calculate the specific heat of a solid in J/g•ºC if 1638 J raises the temperature of 125 g of it from 25.0 ºC to 52.6 ºC.

Find: specific heat (J/g•ºC)

Given: 125 g, 1638 J, T_i = 25.0 ºC, T_f = 52.6 ºC

$$\text{spec. heat (J/g•ºC)} = 1638\ \text{J}$$

$$\text{spec. heat (J/g•ºC)} = (1638\ \text{J})\left(\frac{1}{125\ \text{g}}\right)$$

$$\text{spec. heat (J/g•ºC)} = (1638\ \text{J})\left(\frac{1}{125\ \text{g}}\right)\left(\frac{1}{52.6\ \text{ºC} - 25.0\ \text{ºC}}\right) = 0.475\ \text{J/g•ºC}$$

A similar calculation would be: "What is the final temperature of a 225 g piece of copper (spec. heat = 0.385 J/g•ºC) at 85.0 ºC which is added to 325 g of water at 25.0 ºC (spec. heat = 4.184 J/g•ºC)?"

This problem is as complicated as any get in CHM 211/212, so if you can master this, you can handle anything you'll see. Let's start as always:

Find: Final temperature (ºC)

Given: copper: 225 g_{Cu}, spec. heat = 0.385 J/g•ºC, T_i = 85.0 ºC water: 325 g_{H2O}, spec. heat = 4.184 J/g•ºC, T_i = 25.0 ºC

What else do we know? We know that the final temperature will be between 85.0 ºC and 25.0 ºC, right? We know that the heat released by the copper as it cools will warm the water, resulting in the temperature between 85.0 ºC and 25.0 ºC. A good guess here is that the heat released by the copper will exactly equal absorbed by the water. (In CHM 211, this will be a given.) So

$$heat_{Cu} = heat_{H2O}$$

From a specific heat calculation using the formula introduced earlier

$$\text{heat (J)} = (\text{mass})(\text{specific heat})(\text{temperature change})$$

each of these heats can be calculated, but there is one final problem. Mass and specific heat are given, as is the starting temperature of each material, but the final temperature isn't and that is necessary to do the calculation. How is this resolved?

This is fundamentally an applied algebra problem, so all of the information is provided except one thing, that is the variable in the equation. In this case it is the final temperature, T_f.

$$heat_{Cu} = heat_{H2O}$$

$$(mass_{Cu})(\text{specific } heat_{Cu})(\text{temperature change}) = (mass_{H2O})(\text{specific } heat_{H2O})(\text{temperature change})$$

$$(225\text{ g})(0.385\text{ J/g•ºC})(85.0\text{ ºC} - T_f) = (325\text{ g})(4.184\text{ J/g•ºC})(T_f - 25.0\text{ ºC})$$

$$(86.625\text{ J/ºC})(85.0\text{ ºC} - T_f) = (1{,}359.8\text{ J/ºC})(T_f - 25.0\text{ ºC})$$

$$7{,}363\text{ J} - (86.625\text{ J/ºC})T_f = (1{,}359.8\text{ J/ºC})T_f - 33{,}995\text{ J}$$

$$41{,}358\text{ J} = (1{,}446.4\text{ J/ºC})T_f$$

$$T_f\ 41{,}358\text{ J} = (1{,}446.4\text{ J/ºC})$$

$$T_f = 28.6\text{ ºC}$$

Is this a reasonable answer? A reasonable, initial reaction to this result is that it seems too low, even if it is between 25 ºC and 85 ºC. The copper weighs less (about three-quarters that of the water), so closer to 25 ºC than 85 ºC is reasonable, but so close to the initial water probably seems wrong. Is it? Remember that water has a specific heat of 4.184 J/g•ºC, which is more than 10 times larger than that of the copper. That means one gram of water holds more than 10 times the heat of a gram of copper. Doing a quick estimate: the copper starts 60 ºC warmer than the water, but holds one-tenth the heat, so equal masses of water and copper would equate to a roughly 6 ºC temperature increase in the water. However, you have only about two-thirds (225/325) of the mass of copper, relative to water and 2/3 of 6 ºC is 4 ºC. What is found in the calculation is that the water warms 3.6 ºC, which is close to the estimate, so the answer is plausible.

4.6 Energy in the Real World – Read this section on your own

August 21, 2023