**Instructor's Note: This section is a bit of a derivation, which we try to avoid. However, in this case it is worth it. Spend some time trying to follow along; get out a piece of paper and work it for yourself. The result will be a “woah” moment that explains numbers you had to look up in a table in chemistry!**

If you have taken chemistry before, then the context of the previous problem may seem familiar to you. Problems about how much energy you need to raise or lower the temperature of a substance are often described in chemistry books as being calorimetry problems *(David provide a link to OpenStax Chemistry section on calorimetry)* to be solved using the expression

.

In this expression, is the mass of substance, is the change in temperature, and is a quantity called *specific heat* with units . If you have done these types of problems, then you have probably either solve for the specific heat or looked it up in a table like the one below for the substance in which you were interested.

|  |  |  |
| --- | --- | --- |
| Substance | State | Specific Heat at constant volume (J/gK) |
| Helium | He(g) | 3.12 |
| Water | H2O(l) | 4.19 |
| Ethanol | C2H6O(l) | 2.3 |
| Nitrogen | N2(g) | 0.743 |
| Aluminum | Al(s) | 0.87 |
| Argon | Ar(g) | 0.315 |
| Iron | Fe(s) | 0.46 |
| Copper | Cu(s) | 0.39 |

***Table:*** *A table of specific heats measured while holding the volume constant taken from* [*http://www.engineeringtoolbox.com/specific-heat-capacity-gases-d\_159.html*](http://www.engineeringtoolbox.com/specific-heat-capacity-gases-d_159.html)

It turns out that, with what you know now, you can actually *predict* these specific heats to a high degree of accuracy based upon the properties of the molecule!

Let’s compare the chemistry calorimetry formula

with the equation you were asked to note in the last example

which after some rearranging becomes

.

In both expressions we have a quantity telling us how much material we have (mass or number of moles), multiplied by some number, and then multiplied again by the change in temperature to get the amount of heat.

To further see the comparison, let’s multiply our result by a clever form of the number 1, where is the molar mass of the substance in g/mol (I know this seems dumb, but bear with me for just a few lines):

This is one of a physicist's favorite tricks and is totally legit: I did the same thing to both sides. One the right-hand side of our equation, where the heat is, nothing happens, I multiplied by 1 after all. On the left-hand side, however, I can rearrange things in an interesting way:

.

All I did is take the top molar mass and put it next to the number of moles , and took the bottom molar mass and put it with the 2. Now, is the mass of the substance by definition; has units g/mol and is the number of moles leaving us with grams. Making this substitution we get

which when we compare with the calorimetry formula from chemistry

we see that the only way this can work is if

# Example: Let’s test this result for two materials

We shall test this result for two different materials (a) argon and (b) copper and compare with the table above

## Solution for argon

Argon is an ideal gas so . The molar mass for argon is 39.948 g/mol. Substituting into our expression we get

which matches the table value exactly!

## Solution for copper

Copper is a solid with and a molar mass of 63.546 g/mol. Thus, we expect the specific heat of copper to be

Again, a perfect match!

# Takeaway

This is our first example where microscopic properties of atoms can be used to get macroscopically measured phenomena: we went from the microscopic structure of atoms to predicting the macroscopic specific heat. We will do more of connecting these two worlds in class throughout this unit. The next unit on entropy continues this pattern.