

What is heat? Is heat temperature? **No!**

Which would be worse, putting your finger in a jigger of boiling water or putting your finger in a pot of boiling water?

The pot, right?

Why, they are both at the same temperature?

Since there is more water in the pot the pot has the ability to transfer more heat to your finger, and thus do more damage.

The jigger of boiling water releases heat to your finger. Since your finger contains a good deal of water all the energy is absorbed, and the temperature of your finger goes up.

The pot of boiling water releases heat to your finger. If you can keep your finger in the water long enough all the energy is absorbed by your poor tiny little finger, and the temperature of your finger goes up a lot!

Heat is a flow of energy that changes the temperature of an object.

$$q = (\text{something}) \times \Delta T$$

So, a certain amount of heat can change the temperature of an object, but the size of the temperature change does not simply depend on the amount of heat; the size of the temperature change also depends on the object.

The “something” is the heat capacity (C) of the object.

$$q = C \Delta T$$

C is the amount of heat (q) required to raise the temperature of an object by 1 K ( $\Delta T$ ). The units would be  $\text{J} \cdot \text{K}^{-1}$

The heat capacity of an object depends on (1) the size of the object (Simply put it takes more heat to boil a 12 qt. pot of water than it does to boil a 2 qt. pot of water), and (2) the identity of the object.

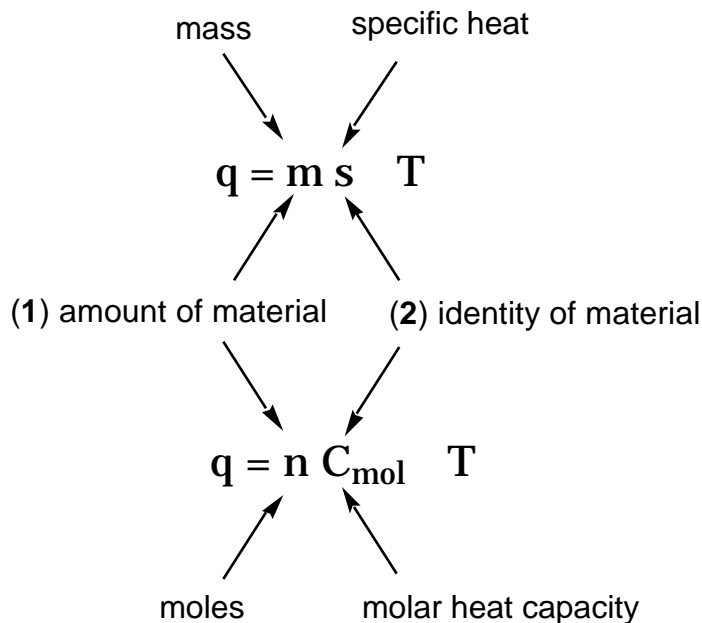
My 5lb frying pan gets hotter faster than 1/2 a gallon (1 gal of water weighs ca 8 lbs) water with the same amount of heat.

For object that have a uniform composition; i.e., objects that are the same all the way through. It is convenient to break down C into its components, size and **identity**.

**Identity**, what do you mean?

Water is water, and it will always take the same amount of heat to increase the temperature of 1 g of water by 1 K.

The same thing can be said from a molar point of view; it will always take the same amount of heat to increase the temperature of 1 mol of water by 1 K.



The amount and identity statement can be expressed in terms of grams and specific heat, or moles and molar heat capacity.

$$q = m s \Delta T$$

When discussing heat in terms of grams of material, the **specific heat**,  $s$ , is the quantity that is used. Specific heat is the flow of heat required to change the temperature of 1 g of a substance by 1 K.

The unit for  $s$  (specific heat) is  $\text{J} \cdot \text{g}^{-1} \cdot \text{K}^{-1}$ .

$$q = n C_{\text{mol}} \Delta T$$

When discussing heat in terms of moles of material, the **molar heat capacity**,  $C_{\text{mol}}$ , is the quantity that is used. The molar heat capacity is similar to specific heat, but instead of heat flow per g the unit is heat flow per mole. The molar heat capacity is the heat flow required to change the temperature of one mole of a substance by 1 K.

The unit for  $C_{\text{mol}}$  is  $\text{J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ .

How much heat is required to raise the temperature of your tea water from room temperature to  $100^\circ\text{C}$ . Assume your tea cup holds 10.0 mol of water (It actually does! 10 mol water is approximately 6 fl. oz.)? The specific heat of water is  $4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1}$ .

If you are going to work in moles you need to know the molar heat capacity

$$\frac{4.184 \text{ J}}{1 \text{ g H}_2\text{O} \cdot 1 \text{ K}} \times \frac{18 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 75.318 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$$

$$q = n C_{\text{mol}} \Delta T$$

$$q = (10.0 \text{ mol})(75.318 \text{ J} \cdot \text{mol}^{-1} \cdot ^\circ\text{C}^{-1})(100 - 23^\circ\text{C})$$

$$q = 57,990 \text{ J}$$

## Calorimetry to determine $\Delta H$ or measure heat transfers

In lab we determined the heat capacity of a metal; in other words, we determined how much energy is required to raise the temperature of 1 g of metal by 1 °C.

To determine the specific heat of metal we construct a calorimeter, and measure the heat transferred from the hot metal to the calorimeter. All the heat released by the metal is absorbed by the calorimeter. Since we can measure the change in temperature for the calorimeter and the metal we can determine the amount of heat transferred. Thus, we can determine the specific heat of the metal

We know that all the energy in the metal is transferred to the calorimeter.

That is

$$q_{\text{tot calorimeter}} = - q_{\text{metal}}$$

If the thing is releasing heat then the calorimeter must be absorbing heat. Did you notice in English we say release and absorb? Release and absorb indicate that the quantities are opposite. I give you money (unlikely) and you take the money, give and take are opposites.

If I give you \$10 I am -\$10 and you are +\$10.

Clearly,

$$-10 \quad +10,$$

but

$$-(-10) = (+10).$$

Same thing with energy...

rather

$$-q_{\text{absorbed}} = q_{\text{released}}$$

Since we know the heat capacity of water we can determine the heat capacity of the metal.

$$-q_{\text{metal}} = T_{\text{water}} (C_{\text{water}})$$

and

$$-[ T_{\text{metal}} (C_{\text{metal}})] = T_{\text{water}} (C_{\text{water}})$$

in terms of specific heat

$$-q_{\text{metal}} = q_{\text{water}}$$

$$-[ T_{\text{metal}} \cdot m_{\text{metal}} \cdot S_{\text{metal}}] = T_{\text{water}} \cdot m_{\text{water}} \cdot S_{\text{water}}$$

Since we would know the mass of the metal, the  $T$  for metal, the mass of water, the  $T$  for water, and the specific heat we could solve for the specific heat of the metal.

**But** there is a small problem. Water is not the only thing absorbing heat. The Styrofoam cup, and the thermometer are also absorbing heat. Which means that heat flows from the metal to the water and the calorimeter (the Styrofoam cup).

or...

$$-q_{\text{metal}} = q_{\text{water}} + q_{\text{calorimeter}}$$

which means that we must account for the amount of heat that the calorimeter absorbs; otherwise, we will not know how much heat the metal is actually releasing.

It may be obvious that calorimeters do not have specific heats (calorimeters are not universally constant in their design and construction), but nevertheless a calorimeter does have a *heat capacity*. If we find the heat capacity for our calorimeter then we know how much heat it absorbs when it increases in temperature.

We find the *heat capacity* of the calorimeter by supplying a **known** amount of heat and determining the amount of heat that went into the calorimeter.

So, we add 50.0 g 100.0 °C water to our calorimeter that contains 50.0 g of water at 23.0 °C. The final temperature is 59.0 °C.

$$-(50.0 \text{ g})(4.184 \text{ J}\cdot\text{g}^{-1}\cdot\text{°C}^{-1})(59 \text{ °C} - 100 \text{ °C}) = C_{\text{cal}}(59 \text{ °C} - 23 \text{ °C}) + (50.0 \text{ g})(4.184 \text{ J}\cdot\text{g}^{-1}\cdot\text{°C}^{-1})(59 \text{ °C} - 23 \text{ °C})$$

$$8577.2 \text{ J} = C_{\text{cal}}(36.0) + 7531.2 \text{ J}$$

$$1046 \text{ J} = C_{\text{cal}}(36.0)$$

$$C_{\text{cal}} = 29.1 \text{ J}\cdot\text{K}^{-1}$$

Now that we know the heat capacity of our calorimeter we can use our calorimeter to determine the amount of transferred between the water and some metal.

So let's now determine the specific heat of a metal.

$$-q_{\text{metal}} = T_{\text{water}} (C_{\text{water}})$$

and 
$$-[ T_{\text{metal}} (C_{\text{metal}})] = T_{\text{water}} (C_{\text{water}})$$

$$-q_{\text{metal}} = q_{\text{water}} + q_{\text{cal}}$$

in terms of specific heat

$$-[ T_{\text{metal}} \cdot m_{\text{metal}} \cdot s_{\text{metal}}] = T_{\text{water}} \cdot m_{\text{water}} \cdot s_{\text{water}} + C_{\text{cal}} T_{\text{water}}$$

100.0 g of Fe at 100.0 °C are added to 50.0 g of water which start at 23.0 °C. The final temperature of the calorimeter is 35.2 °C.

$$-[(35.2-100 \text{ }^\circ\text{C}) \cdot 100 \text{ g} \cdot s_{\text{metal}}] = (35.2-23.0) \cdot 50 \text{ g} \cdot 4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1} + 21.9 \text{ J} \cdot \text{K}^{-1} \cdot (35.2-23.0 \text{ }^\circ\text{C})$$

$$-[-64.8 \text{ }^\circ\text{C}] \cdot 100 \text{ g} \cdot s_{\text{metal}} = (12.2 \text{ }^\circ\text{C}) \cdot 209.2 \text{ J} \cdot \text{K}^{-1} + 21.9 \text{ J} \cdot \text{K}^{-1} \cdot (12.2 \text{ }^\circ\text{C})$$

$$6480 \text{ g} \cdot \text{ }^\circ\text{C} \cdot s_{\text{metal}} = 2552.24 \text{ J} + 267.18 \text{ J}$$

$$6480 \text{ g} \cdot \text{ }^\circ\text{C} \cdot s_{\text{metal}} = 2819.42 \text{ J}$$

$$s_{\text{metal}} = 0.435 \text{ J} \cdot \text{g}^{-1} \cdot \text{ }^\circ\text{C}^{-1}$$

Calorimeters can be used to measure the transfer of heat between objects or they can be used to measure the heat transfers involved in reactions.

Design and perform an experiment that will determine the  $q_p$  ( $H_{rxn}$ ) for the dissolution of  $H_2SO_4$  in water.

How can we determine the heat of a reaction?

When we discussed the heat transfer to water we measured the temperature change of the water.

Can we measure the temperature change of a reaction? **No**, we can only measure the temperature of objects. A reaction is not an object; it is a process.

The key to calorimetry is that all of the heat released by the thing being studied

in this case the dissolution of  $H_2SO_4$

is absorbed by the calorimeter and its contents which are usually water.

3.20 g of  $H_2SO_4$  were dissolved in 100.0 g of water, and the temperature went from 23 °C to 29.8 °C. Determine  $q$  for the dissolution of  $H_2SO_4$  in water.

$$-q_{rxn} = q_{cal} + q_{water}$$

$$-q_{rxn} = C_{cal}(29.8-23) + ms(29.8-23)$$

Since the  $H_2SO_4$  solution is actually the thing that absorbs the heat, we have to use the mass of the solution. Here is where we are going to make a small assumption. The specific heat capacity of the solution is the same as the specific heat capacity of water.

$$-q_{rxn} = (29.1)(29.8-23) + (103.2)(4.184)(29.8-23)$$



$$-q_{\text{rxn}} = 176.8 + 2930.5$$

$$-q_{\text{rxn}} = 3134.0 \text{ J}$$

$$q_{\text{rxn}} = -3134.0 \text{ J}$$

Find the amount of heat released per mole of  $\text{H}_2\text{SO}_4$ .

How many moles of acid were dissolved?

$$3.20 \text{ g } \cancel{\text{H}_2\text{SO}_4} \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.07 \text{ g } \cancel{\text{H}_2\text{SO}_4}} = 0.0326 \text{ mol H}_2\text{SO}_4$$

So,

$$\frac{3.134 \text{ kJ}}{0.0326 \text{ mol H}_2\text{SO}_4} = -96.1 \text{ kJ/mol}$$

We have found  $H_{\text{rxn}}$  but what is  $H$ ?

**Enthalpy is defined as follows**

$$H = E + PV$$

Essentially, enthalpy is defined to eliminate the work term in the  $E$  equation.

**We cannot measure the enthalpy of a system, but we can look at changes in enthalpy.**

$$H = E + P V$$

to make life easier we will make certain that Pressure is held constant...

$$H = E + P V$$

what is  $E + P V$ ?

**We have change in internal energy, and PV work. Lets look at the internal energy expression.**

$$E = q + w$$

or... because we are holding P constant

$$E = q_p - P \Delta V$$

or....rearrange

$$q_p = E + P \Delta V$$

substitute for  $E + P \Delta V$

$$E = q + w$$

and

$$E = q_p - P \Delta V$$

substitute for  $E$

$$H = q_p - P \Delta V + P \Delta V$$

subtract

So,

$$H = q_p$$

$$H = q_p$$

**Under certain conditions, at constant pressure with only PV work being done,  $H$  is the change in *heat*. The difference between the enthalpy of the products and the enthalpy of the reactants is the heat released by the reaction!**

$$H = H_{\text{products}} - H_{\text{reactants}} = q_p$$

$H < 0$  then  $q_p$  is negative which means heat is leaving the system: we call this kind of process exothermic.

$H > 0$  then  $q_p$  is positive which means heat is being absorbed by the system: we call this kind of process endothermic.

Remember this does not represent the change in the total energy content (internal energy,  $E$ ) it just tells us how much heat energy is released during a change occurring under very specific conditions. The conditions are that any work done is simple expansion at constant pressure.

If the  $H$  of a reaction is known, then heat can be treated as though it were a reactant or a product; that is, the amount of heat absorbed or released can be related to the amount of material produced or consumed by using the stoichiometry of the reaction.

How much methane would have to be burned to heat the water?

Since the water absorbs 57.990 J of heat, the reaction must release 57,990 J of heat. So,

$$q_{\text{rxn}} = - 57,990 \text{ J}$$

You use factor label and stoichiometry to do the rest!

$$-57,990 \text{ J} \times \frac{1 \text{ mol CH}_4}{-890,000 \text{ J}} \times \frac{16.04 \text{ g CH}_4}{1 \text{ mol CH}_4} = 1.04 \text{ g CH}_4$$

Some one determined that a mole of  $\text{CH}_4$  releases 890 kJ of heat energy when it is burned;  $H_{\text{combustion}} = -890 \text{ kJ/mol CH}_4$ .

How much heat energy is released when 5.8 g methane are burned?

$$5.8 \text{ g CH}_4 \times \frac{1 \text{ mol}}{16.0 \text{ g CH}_4} \times \frac{-890 \text{ kJ}}{1 \text{ mol}} = -320 \text{ kJ}$$

This is the energy released as heat. Remember the reaction creates gases, and the gases expand into the room/container, so there is work being done by the reaction. That is, some of the energy of the reaction goes into expanding the gases that are produced, and when we measure or discuss  $\Delta H$  we are ignoring the energy which is used to expand the gas.