SPECIFIC GRAVITY

Specific Gravity is the density of a given substance divided by the density of a reference substance. Specific Gravity is used mainly for liquids.

- In most cases the reference substance used is liquid water at 4°C which has D = 1.00 g/ml. Thus.

$$\text{Specific Gravity} = \text{S.G.} = \frac{\text{D Liquid}}{\text{D } H_2O} = \frac{\text{D Liquid}}{1.00 \text{ g/ml}}$$

- Since for water at 4°C, we are merely dividing the D substance by 1.00 g/ml, the specific gravity of a substance is the same as its density—with the exception that specific gravity has no units . . .

$$\text{S.G.} = \frac{\text{D liquid}}{\text{D } H_2O} = \frac{\cancel{g/ml}}{\cancel{g/ml}}$$

- Thus, specific gravity problems are very similar to density problems.
- For very precise scientific work, the temperature of the liquid is measured, and the value for the density of water is used at that same temperature. The density of water varies from 1.00 at 4°C to 0.96 at 100°C.

EXAMPLE 1

Find the S.G. of a liquid which has a 15 g mass and a 14 ml volume.

$$\begin{aligned}\text{S.G.} &= \frac{\text{D liquid}}{\text{D } H_2O} = \frac{(M/V)\text{ liquid}}{\text{D } H_2O} = \frac{15\text{ g}/14\text{ ml}}{1.00\text{ g/ml}}\\ &= \frac{1.07}{1.00}\\ &= 1.07\end{aligned}$$

EXAMPLE 2

Given that acetic acid has a S.G. of 1.05, find the mass of 250 ml of acetic acid. (H_2O is the reference substance.)

$$\text{S.G.} = \frac{\text{D acet}}{D_{H_2O}} = \frac{(M/V)\text{ acet}}{D_{H_2O}}$$

$$\therefore\ M/V = (\text{S.G.})(D_{H_2O}),\text{ and,}$$

$$M = (\text{S.G.})(D_{H_2O})(V) = (1.05)(1.00\text{ g/ml})(250\text{ ml}) = 262.5\text{ g}$$

EXAMPLE 3

Find the volume of chloroform (CCl_4) in ml that weighs 14.9 g. S.G. of chloroform is 1.49. (The reference substance is water.)

$$\text{S.G.} = \frac{D_{CCl_4}}{D_{H_2O}} = \frac{(M/V)_{CCl_4}}{D_{H_2O}}$$

$$\therefore (M/V)_{CCl_4} = (\text{S.G.})(D_{H_2O})$$

$$\therefore\ V = \frac{M}{(SG)(D_{H_2O})}$$

$$V = \frac{14.9\ g}{(1.49)(1.00\ g/ml)} = 10.0\ ml$$

On occasion a specific gravity may be found using a substance other than H_2O as a reference . . .

EXAMPLE 4

Find the S.G. of a liquid if 25 ml weighs 27 g. Use benzene as a reference substance, D = 0.88 g/ml (Note: since the reference liquid is not water, the S.G. will be different from the density.)

$$\text{S.G.} = \frac{D\ \text{substance}}{D\ \text{benzene}} = \frac{(M/V)\ \text{substance}}{D\ \text{benzene}} = \frac{\frac{27\ g}{25\ ml}}{0.88\ g/ml}$$

$$\text{S.G.} = \frac{1.08}{0.88} = 1.23$$

EXERCISES

1. Using H_2O as a reference substance, find the S.G. of a liquid with a mass of 50. g and a volu of 45 ml.
2. 125 ml of a liquid with a S.G. of 0.708 would have what mass? (Water is the reference substanc
3. 75 g of a liquid with a S.G. of 1.49 would have what volume in milliliters? (Use water as reference.)
4. Using carbon tetrachloride as a reference (CD = 1.60 g/mL), find the S.G. of water if 100 g water occupies 100 ml.

SPECIFIC HEAT

I. DEFINITION AND USE OF SPECIFIC HEAT VALUES

Before we define specific heat we must first define a unit used for the measurement of heat . . .

- Calorie—the amount of heat needed to change the temperature of 1 g of H_2O liquid by 1°C.
  - Notice that the definition specifies the substance chosen as the standard (H_2O), the amount of the substance (1 g) and the change in temperature (1°C).
- Specific Heat—the amount of heat (# calories) needed to raise the temperature of 1 g of a substance by 1°C.
  - Note how similar this definition is to the definition of the calorie.
  - The units of specific heat (S.H.) are . . .

$$\text{S.H.} = \frac{\text{cals}}{(\text{g})(°\text{C})}$$

(i.e.) the # cals *per* gram needed to raise the substance in temperature *by* 1°C.

per ⇒ divided by
by ⇒ multiplied by

- It should be obvious that from the definition of the calorie it would hold that the specific heat of water is 1.0 cal./(g)(°C).
- To see the need for a quantity such as specific heat answer this question . . .

> . . . If we take 2 different substances such as sand and water both at the same initial temperatures, both of the same mass, and add the same amount of heat (# cals) to each, will the sand and the water both end up at the same final temperature? The answer is NO—as anyone who has been to the beach on a hot day can tell you! The sand gets much hotter. Why? Because of the inherent differences in molecular structure between sand and water, water is able to store more heat without much of a temperature rise.

- Specific heat values can be used to tell at a glance which substances require more heat to bring about a given rise in temperature and which require less.
- Look at the table of specific heats on the next page and answer this question . . .
  . . . If we have 10 g of each substance, each initially at a temperature of 25°C, and we add 10 calories of heat to each . . .
  (a) which reaches the highest temperature?
  (b) which the lowest temperature?

Substance	Specific Heat cals (g) (°C)
H_2O (*l*)	1.0
H_2O (*l*)	0.75
Aluminum	0.21
Sand	0.18
Iron	0.108

Answer . . .

(a) the highest T—Iron

(b) the lowest T—H_2O liquid

Why? Reread the definition of specific heat. Specific heat is the number of calories needed the temperature of 1 g of a substance by 1°C. So, those numbers on the table represent a of calories. For example, H_2O (*l*) requires 1 calorie to raise 1 g by 1°C while iron requi 0.108 cal for the same mass to change in temperature by the same amount. Clearly, H_2O more energy to raise in temperature than does iron. Since the smallest number on the table it will be the easiest to raise in temperature and thus will reach the highest temperature. Sir has the biggest number it will be the most difficult to raise in temperature and thus will least in temperature.

- In general, when comparing two substances, the one with the bigger specific heat will ris perature the least.
  - Specific Heat and change in temperature are inversely related.

$$\text{as} \quad \text{S.H.} \uparrow, \quad \Delta T \downarrow$$

  - Note that for a given substance the S.H. value is a constant and thus:

$$\text{Cals} = \text{SH} \cdot \text{g} \cdot \Delta T = \text{Constant} \cdot \text{g} \cdot \Delta T$$

Thus, the mass (g) and the change in temperature (ΔT) are related by a D.P. to the a heat (calories). So, as g $\uparrow$, cals. $\uparrow$; as $\Delta T\uparrow$, cals. $\uparrow$.

II. PROBLEMS INVOLVING SPECIFIC HEAT

In problems involving specific heat, the equation should be used in the following form . . .

$$\text{S.H.} = \frac{\#\ \text{cals.}}{(\text{g})(\Delta T)}$$

Where ΔT is the change in temperature in °C brought about by the addition of a given I calories to the substance.

EXAMPLE 1

Calculate the specific heat of a metal if 25 g of the metal rises in temperature from 25°C to 57°C when 200 calories are added.

$$\begin{aligned}\text{S.H.} &= \frac{\text{cals.}}{(\text{g})(\Delta\text{T})} = \frac{200\text{ cals.}}{(25\text{ g})(57°\text{C} - 25°\text{C})} = \frac{200\text{ cals.}}{(25\text{ g})(32°\text{C})} \\ &= \frac{200\text{ cals.}}{800(\text{g})(°\text{C})} = 0.25\frac{\text{cal}}{(\text{g})(°\text{C})}\end{aligned}$$

There are other types of problems which require you to manipulate the specific heat equation algebraically . . .

$$\begin{aligned}\text{Cals.} &= (\text{S.H.})(\text{g})(\Delta\text{T}) \\ \text{g} &= \frac{\text{cals}}{(\text{S.H.})(\Delta\text{T})} \\ \Delta\text{T} &= \frac{\text{cals}}{(\text{g})(\text{S.H.})}\end{aligned}$$

EXAMPLE 2

How many calories are needed to raise the temperature of 10. g of a metal from 30°C to 70°C if the metal has a S.H. of 0.15 cal/(g) (°C).

$$\begin{aligned}\text{cals} &= (\text{S.H.})(\text{g})(\Delta\text{T}) \\ &= \left[.15\frac{\text{cal.}}{(\cancel{\text{g}})(°\cancel{\text{C}})}\right][10\,\cancel{\text{g}}][70°\cancel{\text{C}} - 30°\cancel{\text{C}}] \\ &= 60.\text{ cal.}\end{aligned}$$

EXAMPLE 3

450 calories were required to raise the temperature of a piece of metal from 25°C to 70°C. If the S.H. of the metal is 0.31 cal/(g) (°C), find the number of grams of metal used.

$$\begin{aligned}\#\,\text{g} &= \frac{\#\text{ cals.}}{(\text{S.H.})(\Delta\text{T})} = \frac{450\,\cancel{\text{cal.}}}{\left(0.31\frac{\cancel{\text{cal.}}}{(\text{g})(°\cancel{\text{C}})}\right)(70°\cancel{\text{C}} - 25°\cancel{\text{C}})} \\ \#\,\text{g} &= \frac{450}{\left(0.31\frac{1}{\text{g}}\right)(45)} = 32.3\text{ g}\end{aligned}$$

EXAMPLE 4

When 20. cal are added to 100. g of a liquid having a S.H. of 0.15 cal/(g) (°C), what temperature change will be affected?

$$\Delta T = \frac{\#\text{ cals.}}{(\text{g})(\text{S.H.})} = \frac{20\ \cancel{\text{cals.}}}{(100\ \cancel{\text{g}})(0.15\ \cancel{\text{cals}}\ (\cancel{\text{g}})\ (°\text{C}))}$$

$$\Delta T = \frac{20}{15 \cdot \frac{1}{°\text{C}}} = 1.3°\text{C}$$

EXERCISES

1. Calculate the S.H. of a metal if 10 g of the metal rises in temperature from 25°C to 40°C when 100 calories are added.
2. How many grams of a solid substance were used if 150 calories raised the temperature from 30°C to 80°C? The S.H. of the solid was 0.23 cal/(g) (°C).
3. How many calories would be used in raising the temperature of 5 g of a metal from 50 to 100°C? The S.H. of the metal is 0.28 cal/(g) (°C).
4. What change in temperature would occur if 10 g of a metal with a S.H. of 0.17 cal/(g) (°C) absorbed 250 calories?
5. What is the specific heat of a metal if 10 g of the metal requires 90 calories to raise the temperature by 35°C?

III. OBTAINING SPECIFIC HEAT VALUES—CALORIMETRY

In order to determine the specific heat value for a given substance, one would determine the number of calories needed to raise the temperature of a given weight of the substance by a given number of degrees centigrade. In theory this sounds simple. However, adding heat to a substance while trying to prevent the substance from losing calories to the surroundings is difficult. In the method typically employed, the investigator heats a known weight of a substance to a given temperature, drops the heated substance into a known weight of water at a given temperature, and measures the temperature rise of the water. The specific heat of the substance is calculated algebraically. The method will be shown later.

The basic assumption in this method is that the heat lost by the unknown substance is completely transferred to the water—no heat is lost to the surroundings. In order to insure that no heat loss occurs, the operation is carried out within a vessel which has insulated walls called a calorimeter. (literally: calorie meter—see the figure on next page).

Suppose you are asked to determine the specific heat of an unknown metal. The following are the steps which would be followed:

1. Weigh the metal.
2. Weigh the water for the calorimeter.
3. Boil some water and add the metal to it. Let it sit a few minutes. This ensures that the initial metal temperature is 100°C. This is not the same water which is used in the calorimeter.

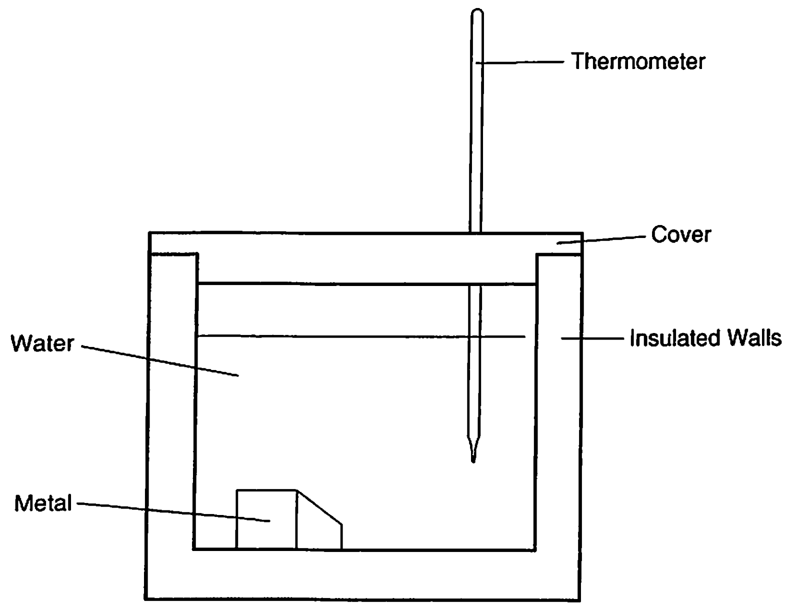

FIGURE 7.1—A CALORIMETER.

4. Take the initial water temperature in the calorimeter.
5. Using forceps (metal pincers), quickly remove the metal from the boiling water and add it to the calorimeter. Close the cover quickly.
6. Take the temperature of the water each 15 seconds until the temperature peaks and falls. Use the maximum temperature as the final temperature in the calculation below.
   Suppose the following data were obtained . . .

	Metal	H_2O
Mass	10. g	100 g
T_i^*	100°C	25°C
T_F^{**}	26°C	26°C
S.H.	X	$1.0 \frac{\text{cal.}}{(\text{g})(°\text{C})}$

$^*T_i$ = initial temperature
$^{**}T_F$ = final temperature

As stated earlier, the basic assumption is . . .

$$\{\text{Heat (Cals.) Lost By Metal}\} = \{\text{Heat (Cals.) Gained By } H_2O\}$$

From the previous section recall that . . .

$$\text{Cals.} = (\text{g})\ (\text{S.H.})(\Delta T)$$

Replacing calories in the above equation . . .

$$(g_M)(S.H._M)(\Delta T_M) = (g\ H_2O)(S.H._{H_2O})(\Delta T_{H_2O})$$

Solving for $S.H._M$ . . .

$$S.H._M = \frac{(g_{H_2O})(S.H._{H_2O})(\Delta T_{H_2O})}{(g_M)(\Delta T_M)}$$

Now simply plug in the values and solve for the specific heat of the metal . . .

$$S.H._M = \frac{(100\ \cancel{g})\left(1.0\frac{\text{cal}}{(\cancel{g})(°\cancel{C})}\right)(26°\cancel{C} - 25°\cancel{C})}{(10\ g)(100°C - 26°C)}$$

$$= \frac{100\ \text{cal.}}{(10\ g)(74°C)}$$

$$= \frac{100\ \text{cal.}}{740(g)(°C)} = 0.14\frac{\text{cal.}}{(g)(°C)}$$

Notice that the ΔT values must be set up to be positive . . .

$$\Delta T\ H_2O = T_F - T_i = 26°C - 25°C = 1°C$$
$$\Delta T_M = T_i - T_F = 100°C - 26°C = 74°C$$

In the ΔT for H_2O, notice that the final temperature comes first, while in the ΔT for the metal, initial temperature comes first. (See Figure 6.2)

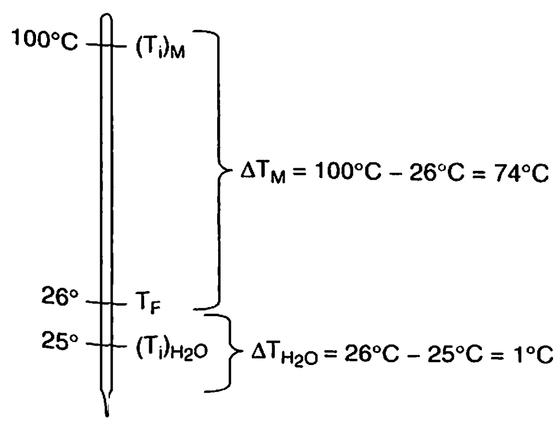

FIGURE 7.2—ΔT VALUES MUST ALWAYS BE POSITIVE.

EXERCISES

6. 15 g of a metal at an initial temperature of 100°C is placed in a calorimeter. The calorimeter c tains 75 g of H_2O at 20°C ($S.H._{H_2O}$ = 1.0 cal/(g)(°C)). The final temperature is 23°C. I S.H. of metal.
7. 10 g of iron (S.H. = 0.108 cal/(g)(°C)) at 100°C is placed in 100 g of H_2O at 25°C ($S.H._{H_2C}$ 1.0 cal/(g)(°C)). Find the final temperature.