P C1 (pg 1 of 6) Calorimetry

Name______________________________Per_____

Please sketch a *temperature diagram* as shown to the right for each problem before starting the problem.
The most important issue is to be able to set up the equation. Then dig in and work the algebra.

Temperature Diagram	
T_M ______	
	ΔT_M _________
T_f ______	
	ΔT_W _________
T_w ______	

1. Determine the equilibrium temperature
   a. when 95 g of water at 88ºC is mixed with 67 g of water at 33ºC

   b. when 125 g of water at 68ºC is mixed with 138 g of water at 17ºC.

2. Determine the volume of the cold water
   a. when 345 ml of water at 78ºC is added to water at 12ºC and the thermal equilibrium temperature is 43ºC.

   b. when 99 ml of water at 89ºC is added to water at 31ºC and the thermal equilibrium temperature is 43ºC.

3. Determine the starting temperature of the hot water.
   a. when 128 g of water at 11ºC is mixed with 95 g of water and the final temperature is 28ºC.

   b. when 86 g of water at 98ºC is mixed with 56 g of water and the final temperature is 68ºC.

4. Determine the equilibrium temperature
   a. when 95 g of aluminum at 120ºC is mixed with 167 g of water at 33ºC

   b. when 55 g of lead at −15ºC is mixed with 55 g of water at 65ºC.

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5. How many kilojoules of energy are required to change the temperature of 36 g of water from 5ºC to 65ºC?

Specific Heat Capacity	(J/gºC)
Liquid water	4.18
Liquid gasoline	2.22

   a. Compared to water, would more or less energy be required to change the temperature of 36 g of gasoline 60ºC? Justify your response with a calculation.

6. If 125 g of water at 68ºC is mixed with 138 g of gasoline at 17ºC, set up an equation and then solve that equation to find the equilibrium temperature. (The specific heat capacity of gasoline is 2.22 J/gºC)

7. Equal masses of two different substances absorb the same amount of heat. The temperature of substance B increases twice as much as the temperature of substance A. Which substance has the higher specific heat capacity? EXPLAIN and/or SHOW CLEARLY.

8. Calculate the specific heat of a piece of metal with a mass of 125 g. The metal was heated in boiling water for an hour and was put into 250.0 ml of water that started at 20.5ºC and reached equilibrium at 26.0ºC. Fill in the temperature diagram to the right, show your problem set up and then solve. Circle your final answer.

Temperature Diagram

T_M _______

ΔT_M __________

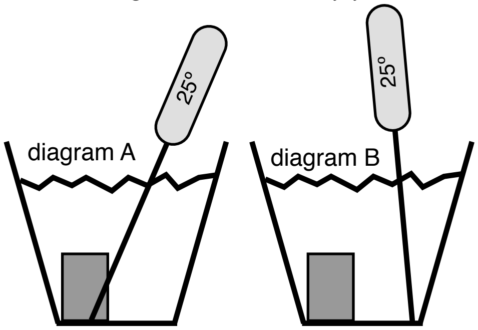
T_f _______

ΔT_W __________

T_W _______

   a. From your list of specific heat capacities, select the two metals that are closest to the value you calculated. Propose which metal it is most likely to be the one tested in the procedure above. Justify your choice.

   b. After having made your choice in part a, and having been told that the students making the measurements did not stir their water metal mixture, is the arrangement of their thermometer most likely to be similar to diagram A or B. Justify your choice.



Specific Heat Capacity Values	J/gºC
sodium	1.22
aluminum	0.9
potassium	0.748
chromium	0.448
iron	0.447
nickel	0.443
copper	0.386
zinc	0.386
gallium	0.374
silver	0.236
cadmium	0.232
tin	0.22
barium	0.179
mercury	0.139
lead	0.138
platinum	0.134
gold	0.128

Still More Calorimetry

Remember to get your work on paper before you even begin to touch a calculator.
Remember you may need to access reference pages for specific heat capacity or density values.

9. Determine the thermal equilibrium temperature (the specific heat capacity of gasoline is 2.22 J/gºC)
   a. If 100.0 g of water at 77ºC is mixed with 80.0 g of water at 23ºC
   b. If 125 g of water at 68ºC is mixed with 138 g of gasoline at 17ºC, determine the thermal equilibrium temperature.

10. Determine the (Assume the density of alcohol = 2.7 g/ml)
    a. volume of 11ºC cold water that is combined with 145 ml of water at 78ºC and thermal equilibrium is reached at 33ºC.
    b. size of the hot piece of aluminum If 99 ml of water at 21ºC has a piece of hot aluminum at 106ºC dropped into it, and the thermal equilibrium temperature is 29ºC, determine the.

11. Determine the starting temperature (the specific heat capacity of alcohol = 2.46 J/gºC)
    a. of 205 g hot water, when mixed with 325 g of water at 13ºC and the final temperature is 28ºC.
    b. of 100.0 g of alcohol, when mixed with 100.0 g of water at 45ºC and the final temperature is 33ºC.

12. How many joules of energy are required to change the temperature
    a. of 1 g of water from 5ºC to 10ºC?
    b. of 23 g of water from 5ºC to 6ºC?

13. What is the specific heat of (The specific heat capacity of olive oil is 2.17 J/gºC)
    a. olive oil if 435 J of heat is added to 3.4 g of 21.0ºC olive oil that heats up to 85ºC?
    b. some liquid when 785 J of heat is added to 13.4 g of this 21.0ºC liquid that heats up to 75ºC?

14. How many kilojoules of heat is absorbed
    a. when 1.00 L of water is heated from 18ºC to 85ºC?
    b. when 15.0 L of water is heated from 18ºC to 55ºC?

15. What is the specific heat of
    a. A piece of stainless steel with a mass of 1.55 g which absorbs 141 J of heat energy when its temperature increases by 178ºC.
    b. A piece of some metal with a mass of 8.67 g which absorbs 967 J of heat energy when its temperature increases by 157ºC.

16. Determine the mass of a piece of lead if it drops from 100.0ºC to 16.5ºC in 120.0 g of water which started at 10.5ºC. (The specific heat capacity of lead is 0.138 J/gºC)

17. Determine the volume of water starting at 15.00ºC that 200.0 g of aluminum starting at 100.0ºC can heat to 27.0ºC. (The specific heat capacity of aluminum is 0.900 J/gºC)

18. What is the mass of a copper block if it was heated to 100.0ºC and reached equilibrium with 200.0 ml of water at 33.5ºC and the water started at 30.5ºC. (The specific heat capacity of copper is 0.386 J/gºC)

19. Determine the specific heat
    a. for a 37.360 g piece of cold −16.0ºC aluminum that cooled 155 ml of water from 63.0ºC to 59.0ºC.
    b. Compare this experimental value to the theoretical value (listed on page 1) and determine the % error.

20. Calculate the specific heat of cadmium if a piece with a mass of 106 g heated to 100.0ºC was put into 250.0 ml of water that started at 30.7ºC and reached equilibrium at 32.0ºC.
    a. Compare this experimental value to the theoretical value and determine the % error.
    b. Is it more likely that the error was caused by the cadmium not being heated to 100.0ºC, or was it heated higher than 100.0ºC by accidentally passing the metal through the Bunsen flame before dropping into the calorimeter?

21. A bowl of Rice Krispees provides 627 kJ of energy. What volume of water could the same amount of energy raise from 22.0ºC to boiling temp?

22. A 65.2 g lump of aluminum that was put into 150.0 ml of water which rose from 12.3ºC to 15.3ºC. (The specific heat capacity of aluminum is 0.900 J/gºC)
    a. Determine the temperature change for this piece of metal
    b. What temp did the aluminum start at?

23. A lump of aluminum (density = 2.7 g/ml) raised 150.0 ml of water from 23.0ºC to 30.0ºC. (The specific heat capacity of aluminum is 0.900 J/gºC)
    a. If the piece of aluminum displaces a volume of 23.2 ml of water, what was the temp change of the Al?
    b. What was the starting temp of the Al?

P C.1 (pg 4 of 6) **Calorimetry** *If you are having trouble, try NS C1 for more help.* ***ANSWERS***

In all of these problems, we must assume that the heat lost equals the heat gained.
heat lost = heat gained
$-(m \times c \times \Delta T) = m \times c \times \Delta T$ *(Since the c is the same, all the problems are about water, you can cancel it out on both sides.)*

1. *For these problems, when the final equilibrium temp is not known you must give it the value of x. Remember that both the hot and the cold substance end at the same final temp which you do not know, so you must call it x. Using the x you can write the ΔT in terms of x. Refer to the temp diagram below. Since both substances are water, c has cancelled out.*

   a. $-(95)(x - 88) = (67)(x - 33)$
   How is your algebra??? x = **65ºC**

   b. $-(125)(x - 68) = (138)(x - 17)$
   How is your algebra??? x = **41ºC**

Temperature Diagram
Problem 1a

T_m 88
ΔT_m $x - 88$
T_f x
ΔT_w $x - 33$
T_w 33

Temperature Diagram
Problem 1b

T_m 68_
ΔT_m $x - 68$
T_f x
ΔT_w $x - 17$
T_w 17

2. *In these problems you know the ΔT values, so you simply solve for the volume of the cold water.*
   Since both substances are water, c has cancelled out.

   a. $(345)(43 - 78) = -(x)(43 - 12)$
   How is your algebra??? x = 390 g
   which of course equals **390 ml** since the density of water is 1 g/ml

   b. $(99)(43 - 89) = -(x)(43 - 31)$
   How is your algebra??? x = 380 g
   which of course equals **380 ml** since the density of water is 1 g/ml

Temperature Diagram
Problem 2a

T_m _78_
ΔT_m −35
T_f _43_
ΔT_w 31
T_w _12_

Temperature Diagram
Problem 2b

T_m _89_
ΔT_m −46
T_f _43_
ΔT_w 12
T_w _31_

3. *In these problems, the algebra is a little bit faster if you set x equal to the ΔT, however this means you have one more step to finish the problem after solving for x. Since both substances are water, c has cancelled out.*

   a. $(95)(x) = (128)(28 - 11)$
   How is your algebra??? x = −23ºC
   but x was defined as the $\Delta T = T_f - T_i$ thus $-23ºC = 28ºC - T_i$ solve: $T_{i\ hot}$ = **51ºC**
   OR you may have set it up this way $-(95)(28 - x) = (128)(28 - 11)$, and x = **51º**

   b. $-(86)(68 - 98) = (56)(x)$
   How is your algebra??? x = 46ºC
   but x was defined as the $\Delta T = T_f - T_i$, thus $68 - T_i = 46ºC$, thus T_i = **22ºC**
   OR you may have set it up this way $-(86)(68 - 98) = (56)(68 - x)$ and x = **22º**

Temperature Diagram
Problem 3a

T_m ____
ΔT_m x
T_f 28
ΔT_w 17
T_w 11

Temperature Diagram
Problem 3b

T_m 98
ΔT_m −30
T_f 68
ΔT_w x
T_w ____

4. *For these problems, when the final equilibrium temp is not known you must give it the value of x. Remember that both the hot and the cold substance end at the same final temp which you do not know, so you must call it x. Using the x you can write the ΔT in terms of x. Refer to the temp diagram below. In these problems the materials are different, so you must include the SHC values. Watch out for your T value which is quite different because the temp passes through 0ºC. ΔT is the absolute value of the total number of degrees changed)*

   a. $-(95)(0.90)(x - 120) = (167)(4.18)(x - 33)$
   How is your algebra??? x = **42.5ºC**

   b. $-(4.18)(55)(x - 65) = (0.138)(55)(x + 15)$
   How is your algebra??? x = **62.4ºC**

Temperature Diagram
Problem 4a

T_m 120
ΔT_m $x - 120$
T_f x
ΔT_w $x - 33$
T_w 33

Temperature Diagram
Problem 4b

T_m 65
ΔT_m $x - 65$
T_f x
ΔT_w $x - (-15)$
T_w −15

5. *In this problem there is only one substance. This means you should use the heat formula* $q = m \times c \times \Delta T$

$$36g \times \frac{4.187J}{g^\circ C} \times (65^\circ - 5^\circ C) = 9028.8J \times \frac{1kJ}{1000J} = 9.03kJ$$

   a. The gasoline would require less energy because it has a smaller specific heat capacity and thus less energy is needed to move its temperature.

6. In this problem, there are both a hot and cold substances going through temperature changes. this requires using this equation: $-(m \times c \times \Delta T) = m \times c \times \Delta T$

$$-\left(125g \times \frac{4.187J}{g^\circ C} \times (x^\circ - 68^\circ C)\right) = 138g \times \frac{2.22J}{g^\circ C} \times (x^\circ - 17^\circ C)$$ solve and **x = 49ºC**

7. Substance A must have a higher specific heat capacity because an equal amount of energy causes it to change fewer degrees.

   You could choose values, choose simple values $1g \times \frac{\#_A J}{g^\circ C} \times 1^\circ C = 1g \times \frac{\#_B J}{g^\circ C} \times 1^\circ C$ thus $\#_A > \#_B$

   Plug in some simple (absolute) values for ΔT, making ΔT of B twice as large as ΔT, and choose equal mass values.

   If you wish, you can pick a simple value for the specific heat capacity, c of A, and solve for B
   $1 \times 1 J/g^\circ C \times 1^\circ = 1 \times c_B \times 2^\circ$ so, $c_B = 0.5$, thus **A > B**

8. In this problem you can assume that boiling water is ºC and since the metal was in this water warming up for over an hour, the metal must also at 100ºC. Solve this equation: $-(m \times c \times \Delta T) = m \times c \times \Delta T$ for c of the metal $c_m - \left(\frac{m_w \times c_w \times \Delta T_w}{m_m \times \Delta T_m}\right)$

   then substitute the values and solve $c_m - \left(\frac{250 \times 4.187 \times (26 - 20.5)}{125 \times (26 - 100)}\right)$ Solve for **c = 0.62 J/gºC**

   a. Check out the values on the specific heat capacity list and the two closest values are potassium = 0.748 J/gºC and chromium = 0.448 J/gºC (or iron which is so close, 0.447 J/gºC). Given these two choices, the metal must be **chromium or iron**. Even thought the experimental value is closer to the potassium value, it can not be potassium because if you placed potassium in water it would react and burst into flames, thus potassium would never be able to be heated up in boiling water.

   b. Since the experimental specific heat capacity is higher (0.62 J/gºC > 0.448 J/gºC) than the theoretical value, we need to select an error source that would have resulted in a higher calculation. If the thermometer were arranged as in diagram A, and the water were not stirred, the final temperature would be erroneously higher than if this error did not occur. An erroneously higher final temp would cause ΔT_M to be smaller and ΔT_W to be larger.

   Transfer these problem into the formula and you will see that
   the erroneous experimental value turns up to be too large.

$$c_m \uparrow = -\left(\frac{m_w \times c_w \times \Delta T_w \uparrow}{m_m \times \Delta T_m \downarrow}\right)$$

 Calorimetry ***ANSWERS***

*Using the formula Heat lost/gained = $c \times M \times \Delta T$ if only **one** substance is involved,*
or $-(c \times m \times \Delta T) = c \times m \times \Delta T$ = if there is both a hot and a cold substance involved.

9 a $-(4.18)(100)(x-77) = (4.18)(80)(x-23)$ x = **53ºC**
b $-(4.18)(125)(x-68) = (2.22)(138)(x-17)$ x = **49ºC**

10 a $-(4.18)(145)(33-78) = (4.18)(m)(33-11)$ m = **297 g which is 297 ml** since the density of water = 1 g/ml
b $(4.18)(99)(29-21) = -(0.900)(m)(29-106))$ m = 47.8 g and since D = m/v so m/D = V 47.8/2.7 = **17.6 ml**

11 a $(4.18)(325)(28–13) = -(4.18)(205)(\Delta T)$ $\Delta T = -23.8$ºC but this is the ΔT, so $28 - (-23.8) =$ **51.8ºC**
OR you could put x in for the final temp: $(4.18)(325)(28–13) = (4.18)(205)(28-x)$ x = **51.8ºC**

b $-(4.18)(100)(33-45) = (2.17)(100)(\Delta T)$ ΔT = 23ºC but this is the ΔT, so 33 − 23 = **10ºC**
OR you could put x in for the final temp: $-(4.18)(100)(33-45) = (2.46)(100)(33-x)$ x = **12.6ºC**

12 a (4.18)(1)(5) = **21 J**
b (4.18)(23)(1) = **96 J**

13 a (x)(3.4)(64) = 435 x = **2.0 J/gºC**
b (x)(13.4)(54) = 785 x = **1.08 J/gºC**

14 a (4.18)(1000)(67) = **280 kJ**
b (4.18)(15,000)(37) = 2,319,900 J = **2300 kJ**

15 a (c)(1.55)(178) = 141J c = **0.511 J/gºC**
b (c)(8.67)(157) = **0.710 J/gºC**

16 $-(0.138)(m)(16.5-100) = (4.18)(120)(6)$ m = **261 g**

17 $-(0.900)(200)(27-100) = (4.18)(m)(27-15)$ m = 261.6 g and since the density of water is 1 g/ml 261.6 g = **262 ml**

18 $-(0.386)(m)(35.5-100) = (4.18)(200)(33.5-30.5)$ m = **97.7 g**

19 $(c)(37.360)[59-(-16)] = -(4.18)(155)(59-63)$ c = **0.924 J/gºC**
b (0.924 – 0.900)/0.900 = **approx 2.8 %**

20 a $-(c)(106)(32-100) = (4.18)(250)(32-30.7)$ c = **0.188 J/gºC**
b (0.232−0.188)/0.232 = **19.0 %**
c $c_m \downarrow = -\left(\dfrac{m_w \times c_w \times (\downarrow T_f - T_i)}{m_m \times \Delta T_m}\right)$ The experimental value is smaller than theoretical.

The error options given both affect the measured final temp.
Not heating the cadmium all the way through would cause the measured equilibrium temp to appear too low which would cause the calculated ΔT_m to be too large (which is on the bottom of the quotient) and would also cause the calculated ΔT_w to be too small (which is on the top of the quotient), and all other values being the same (c_w and mass values) would result in a smaller specific heat capacity for cadmium.

21 627 kJ >> 627,000 J = $(4.18)(m)(100-22)$ m = 1920 g which is of course equal to **1920 ml**

22 a $(0.900)(65.2)(\Delta T) = (4.18)(150)(15.3-12.3)$ ΔT = **−32.1ºC**
b Since the aluminum must have been hotter to cause the water to raise in temp, 15.3 − (−32.1) = **47.4ºC**

23 a Small density twist: for the aluminum, D × V = m 2.7 g/ml × 23.2 ml = 62.64 g
$-(0.900)(62.64)(\Delta T) = (4.18)(150)(30-23)$ ΔT = **−78.0ºC**

b Since the aluminum must have been hotter to cause the water's temp to raise, 30 − (−78) = **108.0ºC**