Specific Heat and Calorimetry

substance	c (J/g°C)
water	4.184
ethanol	2.452
graphite	0.720
diamond	0.502
iron	0.444
copper	0.385
silver	0.237
gold	0.129
ice	2.092

$m = 654 g$

1. How much heat is required to raise the temp of 654 g of water from 34.5°C to 89.7°C? ↗ q is +

$\Delta T = T_F - T_i$
$= 89.7 - 34.5$
$= 55.2°C$

$q = mC\Delta T$
$= 654 \cdot 4.184 \cdot 55.2°C$
$= 1.51 \times 10^5 J$

Ans $1.51 \times 10^5 J$

2. How much heat is required to raise the temp of 654 g of silver from 34.5°C to 89.7°C? → q = +

$\Delta T = T_F - T_i$
$= 89.7 - 34.5$
$= 55.2°C$

$q = mC\Delta T$
$= 654 \cdot 0.237 \cdot 55.2$
$= 8.56 \times 10^3 J$

Ans $8.56 \times 10^3 J$

3. If 7350 J were added to 152 g of ethanol, its temp would go up by how much? ΔT?

$q = +7350 J$
$m = 152 g$
$C = 2.452 J/g \cdot °C$

$q = mC\Delta T$ or $\Delta T = \frac{q}{(m \cdot C)}$

$= \frac{7350 J}{(152 \cdot 2.452)} = 19.7°C$

Ans $19.7°C$

4. 16.25 g of water at 54.0°C releases 402.7 J. What will be its final temp? ↳ q is −

$m = 16.25 g$
$C = 4.184 J/g \cdot °C$
$T_i = 54.0°C$
$q = -402.7 J$

$\Delta T = T_F - T_i$ so $\Delta T + T_i = T_F$

$\Delta T = \frac{q}{(m \cdot C)} = \frac{-402.7}{(16.25 \cdot 4.184)}$

$\Delta T = -5.9229... °C$

$T_F = -5.9229 + 54.0 = 48.1°C$

Ans $48.1°C$

5. 697 J (q+) are added to a 36.8 g of kerosene and the temp increases from 22.5°C to 34.7°C. Determine kerosene's specific heat.

$q = 697 J$
$m = 36.8 g$
$\Delta T = T_F - T_i = 34.7 - 22.5$
$\Delta T = 12.2°C$

$C = \frac{q}{(m \cdot \Delta T)}$

$= \frac{697}{(36.8 \cdot 12.2)} = 1.55 J/g \cdot °C$

Ans $1.55 J/g \cdot °C$

6. 25 copper pennies (each weighing 3.12 g) are placed in 36.0 g of ethanol at room temp (22.1°C). How much heat will it take to raise the temperature up to 65.8°C? Both Cu and ethanol warm up @ same starting T

$\Delta T = T_F - T_i$
$= 65.8 - 22.1$
$= 43.7°C$

$q_{total} = q_{\text{to heat copper}} + q_{\text{to heat ethanol}}$
$= m_{Cu} \cdot C_{Cu} \cdot \Delta T + m_{eth} \cdot C_{eth} \cdot \Delta T$
$= 78 \cdot 0.385 \cdot 43.7 + 36.0 \cdot 2.452 \cdot 43.7$
$= 5.17 \times 10^3 J$

Ans $5.17 \times 10^3 J$

7. What mass of 54.0°C water must be added to 468 g of 21.0°C water to make the final temp of both come out to be 29.0°C?

$\Delta T = T_F - T_i$

$-q_{H_2O @ 54.0°C} = q_{H_2O @ 21.0°C}$

$-(m_{H_2O} \cdot C_{H_2O} \cdot \Delta T) = (m_{H_2O} \cdot C_{H_2O} \cdot \Delta T)$

$m_{H_2O @ 54.0°C} = \frac{m_{H_2O} \cdot C_{H_2O} \cdot \Delta T}{-C_{H_2O} \cdot \Delta T} = \frac{468 \cdot 4.184 \cdot (29.0 - 21.0)}{-4.184 \cdot (29.0 - 54.0)}$

Ans $150.\ g$

8. What mass of 54.0°C gold must be added to 468 g of 21.0°C water to make the final temp of both come out to be 29.0°C?

$-q_{gold} = q_{H_2O}$

$-(m_{gold} \cdot C_{gold} \cdot \Delta T_{gold}) = (m_{H_2O} \cdot C_{H_2O} \cdot \Delta T_{H_2O})$

$m_{gold} = \frac{m_{H_2O} \cdot C_{H_2O} \cdot \Delta T_{H_2O}}{-(C_{gold} \cdot \Delta T_{gold})} = \frac{468 \cdot 4.184 \cdot (29.0 - 21.0)}{-(0.129 \cdot (29.0 - 54.0))}$

$= \boxed{4860 g}$

Ans 486 g

9. A 325 g brass rod at 100.0°C is placed in a cup containing 162 g of 24.3°C water. The final temp comes out to be 37.4°C. Determine brass's specific heat.

$$-q_{brass} = q_{H_2O}$$

$$-(m_{brass} \cdot C_{brass} \cdot \Delta T_{brass}) = (m_{H_2O} \cdot C_{H_2O} \cdot \Delta T_{H_2O})$$

$$C_{brass} = \frac{m_{H_2O} \cdot C_{H_2O} \cdot \Delta T_{H_2O}}{-(m_{brass} \cdot \Delta T_{brass})} = \frac{162 \cdot 4.184 \cdot (37.4-24.3)}{-(325 \cdot (37.4-100.0))}$$

Ans 0.436 J/g·°C

10. 100.0 g of water at 20.0°C are mixed with 200.0 g of copper at 40.0°C. What will the final temp come out to be? $T_F = ?$ (same in both cases)

$$-q_{Cu} = q_{H_2O}$$

$$-(m_{Cu} \cdot C_{Cu} \cdot \Delta T_{Cu}) = (m_{H_2O} \cdot C_{H_2O} \cdot \Delta T_{H_2O})$$

$$-(200.0 \cdot 0.385 \cdot (T_F - 40.0)) = (100.0 \cdot 4.184 \cdot (T_F - 20.0))$$

$$-77(T_F - 40.0) = 418.4(T_F - 20.0)$$

$$-77T_F + 3080 = 418.4T_F - 8368$$

$$11448 = 495.4T_F$$

$$23.1°C = T_F$$

Ans 23.1°C

Ans (IRO+1): 0.436 1.55 19.7 23.1 29.5 48.1 150 4860 5170 8560 151,000 units (IRO+1): J J J °C °C °C °C g g J/g°C J/g°C

3.9: Enthalpy of Phase Changes

All matter is made of uncountable numbers of particles. All particles are in constant motion. Even solid objects are in motion, even though we cannot detect motion. With two bodies in contact, energy will transfer from the body with the higher energy to the body with the lower amount of energy until the temperatures of each body are the same. Melting of ice is an example of this transfer. Heat enters melting ice, but as long as both solid and liquid are present the temperature does not change. The energy is being used to rearrange the molecules. The potential energy of the liquid water is higher than the ice, therefore the melting of ice is endothermic. Evaporation of sweat is another example.

More Thermodynamics:

Remember: If $q > 0$ we say that the object gained heat - heat flowed into our sample (endothermic), the system absorbed heat. Likewise if $q < 0$ we say that our object lost heat, heat flowed out of our sample (exothermic). The system lost heat.

The temperature change, ΔT, is also related to heat. It can be positive or negative. It is obvious that if $\Delta T > 0$ then the temperature increased and if $\Delta T < 0$ then the temperature decreased.

Clearly whenever $\Delta T > 0$ then $q > 0$, and vice versa.

Changes of State: If you have a glass of ice water (the glass contains both ice and liquid water) and you add some heat to the ice water the temperature does not change. The heat melts some the ice rather than changing the temperature of the system. During a change of state even though energy is being