Specific Heat Capacities of some common substances

Substance	Specific Heat Capacity (J/g°C)
Water (l)	4.184
Water (s)	2.03
Water (g)	2.0
Aluminum(s)	0.89
Iron(s)	0.45
Mercury(l)	0.14
C[illegible](s)	0.71
Si[illegible])	0.24
Gold(s)	0.13

$$Q = Cp\ m\ \Delta T$$

1. Calculate the joules of energy required to heat 454 g of water from 5.4 °C to 98.6 °C.

$\Delta T = 98.6 - 5.4$

$93.2°C$

$$\frac{4.184\ J}{g\ °C} \Big/ \frac{454\ g}{} \Big/ \frac{93.2°C}{} = \boxed{177000\ J}$$

2. a) What quantity of energy (in joules) is required to heat a piece of iron weighing 1.3 g from 25 °C to 46 °C?

$\Delta T\ \ 46 - 25 = 21°C$

$$\frac{.45\ J}{g\ °C} \Big/ \frac{1.3\ g}{} \Big/ \frac{21°C}{} = \boxed{12\ J}$$

b) What is the answer in calories?

$1\ cal = 4.184\ J$

3. A 1.6-g sample of metal that has the appearance of gold requires 5.8 J of energy to change its temperature from 23 °C to 41 °C. is the metal pure gold?

$$Cp = \frac{Q}{m\Delta T}$$

$\Delta T = 41 - 23 = 18$

$$\frac{5.8\ J}{1.6\ g \ \big/\ 18°C} = .20\ \frac{J}{g\ °C}$$

$\boxed{\text{No!}}$

Not Au

4. If 21.5-g of iron are heated to 100. °C and then dropped into an insulated container of water (mass of the water is 132 g and its Temperature is 20°C), what will be the final temperature of the water? *Hint:* heat lost by iron = heat gained by water

$$Q_{iron} = \frac{.45 J}{g °C} \Big| 21.5 g \Big| (100 - x) °C$$

$$Q_{water} = \frac{4.184 J}{g °C} \Big| 132 g \Big| (x - 20) °C$$

$$(x - 20)\, 4.184 J\, (132) = (.45)\, 21.5\, (100 - x)$$

$$x - 20 = .017518\, (100 - x)$$

$$x - 20 = 1.7518 - .0175 x$$

$$1.0175 x = 21.7518$$

$$x = 21.377$$

$$x = 21.4$$