Heat Problems I

Standards:

- 3.1.10 B Describe the concepts of models as a way to predict and understand science and technology.
- 3.4.10 B Analyze energy sources and transfers of heat.

Specific Heat Capacities of some Common Substances

Substance	Cp (J/g°C)		Substance	Cp (J/g°C)
Water (l)	4.184		Iron (s)	0.45
Water (s)	2.03		Mercury (l)	0.14
Water (g)	2.0		Carbon (s)	0.71
Aluminum (s)	0.89		Silver (s)	0.24
			Gold (s)	0.13

1Cal = 4.184Joules　　　1kJ = 1000J

Problems:

1. Convert the following number of calories into joules. (Use D.A. and sig figs.)
   a. 100.0 cal = 418.4 J

   (4)
   $$\frac{100.0\,\text{cal}}{} \left| \frac{4.184\,\text{J}}{1\,\text{cal}} \right.$$

   b. 1.00 X 10^3 cal = 4.18×10^3 J

   (3)
   $$\frac{1.00 \times 10^3\,\text{cal}}{} \left| \frac{4.184\,\text{J}}{\text{cal}} \right.$$

2. Convert the following numbers of joules (J) into kilojoules (kJ). (Use D.A. and sig figs.)
   a. 243,000 J = 243 kJ

   (3)
   $$\frac{243{,}000\,\text{J}}{} \left| \frac{1\,\text{kJ}}{1000\,\text{J}} \right.$$

   b. 0.251 J = .000251 kJ or 2.51×10^{-4} kJ

   (3)
   $$\frac{.251\,\text{J}}{} \left| \frac{1\,\text{kJ}}{1000\,\text{J}} \right.$$

3. Convert the following number of joules (J) into calories (cal). (Use D.A. and sig figs.)

   a. 2450 J = 586 cal or 5.86×10^{2} cal

   $$\frac{\overset{(3)}{2450\,\text{J}}}{} \left| \frac{1\,\text{cal}}{4.184\,\text{J}} \right.$$

   b. 3.864×10^{-4} J = 9.235×10^{-5} cal

   $$\frac{\overset{(4)}{3.864 \times 10^{-4}\,\text{J}}}{} \left| \frac{1\,\text{cal}}{4.184\,\text{J}} \right.$$

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

   $\Delta Q = m\,C_p\,\Delta T$ <- Formula

   $98.6 - 5.4 = 93.2\,°C \leftarrow \Delta T$

   $$\frac{\overset{(3)}{454\,\text{g}}}{} \left| \frac{\overset{(4)}{4.184\,\text{J}}}{\text{g}\,°C} \right| \frac{\overset{(3)}{93.2\,°C}}{}$$ <- D.A.

   <- Answer (Sig Figs and Units)

   177036.7552 Round to 3 Sig Figs

   $\boxed{177{,}000\,\text{J} \text{ or } 1.77 \times 10^{5}\,\text{J}}$

5. What quantity of energy (in joules) is required to heat a piece of iron weighing 1.30 g from 25.0°C to 46.0°C?

   $\Delta Q = m\,C_p\,\Delta T$ <- Formula

   $46.0 - 25.0 = 21.0\,°C \leftarrow \Delta T$

   $$\frac{\overset{(3)}{1.30\,\text{g}}}{} \left| \frac{\overset{(4)}{4.184\,\text{J}}}{\text{g}\,°C} \right| \frac{\overset{(3)}{21.0\,°C}}{}$$ <- D.A.

   <- Answer (Sig Figs and Units)

   114.2232 J Round to 3 Sig Figs

   $\boxed{114\,\text{J}}$

6. A 1.60 g sample of metal that has the appearance of gold requires 5.80 J of energy to change its temperature from 23.0°C to 41.0°C. What is the specific heat of the sample?

$\Delta Q = m C_p \Delta T$ $\quad C_p = \frac{\Delta Q}{m \Delta T}$ <- Formula

$41.0 - 23.0 = 18.0\,°C$ <- Change in Temp

$$\frac{\overset{(3)}{5.80\,J}}{\underset{(3)}{1.60\,g} \;\; \underset{(3)}{18.0\,°C}}$$ <- D.A.

.201388889 Round to 3 sig Figs <- Answer (Sig Figs and Units)

$\boxed{.201\ J/g\,°C}$

Is the metal pure gold and why? (Yes or no, followed by explanation)

No – Probably gold mixed with silver since the value of .201 J/g°C is between Gold .13 and silver .24 J/g°C.

7. Calculate the amount of energy required (in calories) to heat 145.0 g of water from 22.3°C to 75.0°C.

$\Delta Q = m C_p \Delta T$ <- Formula

$75.0 - 22.3 = 52.7\,°C$ <- ΔT

$$\frac{\overset{(4)}{145.0\,g}}{} \;\; \frac{\overset{(4)}{4.184\,J}}{g\,°C} \;\; \frac{\overset{(3)}{52.7\,°C}}{} \;\; \frac{1\,cal}{4.184\,J}$$ <- D.A.

<- Answer (Sig Figs and Units)

7641.5 cal 3 sig Figs

$\boxed{7640\ cal \text{ or } 7.64 \times 10^3\ cal}$

8. Calculate the energy required (in joules) to heat 25.0 g of gold from 120. °C to 155 °C. $\Delta Q = m C_p \Delta T$

<- Formula

$120.0 - 25.0 = 35.\ ^\circ C \leftarrow \Delta T$

(3) 25.0 g	(2) .13 J / g °C	(2) 35. °C

<- D.A.

<- Answer (Sig Figs and Units)

113.75 J → 2 sig Figs

$\boxed{110\ J \text{ or } 1.1 \times 10^2\ J}$

9. A 22.5 g sample of metal X requires 540. J of energy to heat it from 10.0 °C to 92.0 °C. What is the specific heat capacity of metal X?

$\Delta Q = m C_p \Delta T \qquad C_p = \dfrac{\Delta Q}{m \Delta T}$

<- Formula

92.0 °C − 10.0 °C = 82.0 °C

<- Change in Temp

3 540. J		
	22.5 g (3)	82.0 °C (3)

<- D.A.

<- Answer (Sig Figs and Units)

.292682927 J/g °C

Round to 3 sig figs

$\boxed{.293\ J/g\ ^\circ C}$