Chem HOMEWORK
Unit 10 Heat Reiview

Name Key Period _____ Date _________

Heat Measurement …

Heat (or energy) can be measured in units of calories or joules. When there is a temperature change (ΔT), heat (Q) can be calculated using this formula:

$$Q = \text{mass} \times \text{specific heat capacity} \times \Delta T$$

$$\mathbf{Q = mCp\Delta T}$$

<u>Important Values</u>

1.000 cal = 4.184 J

For water: $Cp = 1\ cal/g^{\circ}C = 4.184\ J/g^{\circ}C$

For other Cp values … see table on p. 297!

Examples:

(1) How many calories of energy are needed to heat 250.0 g of water from 25.00 °C to 78.50 °C? How many joules does this equal?

$$Q = mCp\Delta T$$

$$Q = (250.0\ g)(1\ cal/g^{\circ}C)(78.50 - 25.00)^{\circ}C$$

$$Q = 13{,}375.\ cal = \underline{13{,}380\ cal}$$

$$\frac{250.0\,g}{} \cdot \frac{1\,cal}{g^{\circ}C} \cdot \frac{53.50\,^{\circ}C}{}$$

13,380 cal (4.184 J/1.000 cal) = <u>55,980 J</u>

(2) A 2.80 g sample requires 10.1 J of energy to heat it from 21.0 to 36.0 °C. What is the metal?

$Q = mCp\Delta T$ $\quad Cp = \frac{Q}{m\Delta T}$ $\quad \boxed{\text{SILVER}}$

$$Cp = \frac{10.1\,J}{2.80\,g \cdot 15.0\,^{\circ}C} = \boxed{.240\,\frac{J}{g^{\circ}C}}$$

$\Delta T = 36.0 - 21.0 = 15.0$

(3) A 1.60 g sample of gold at 25.0 °C absorbs 5.8 J of energy. What is the final temperature of the gold?

$$\Delta T = \frac{Q}{mCp} = \frac{5.8\,J}{1.60\,g} \cdot \frac{g\,^{\circ}C}{.13\,J} = 28^{\circ}C$$

$+ 25.0$

$\boxed{53\ ^{\circ}C}$

Solve each of the following problems: show all equations and work, pay attention to sig. figs.

(1) 3,568 calories = 14930 J

$$\frac{3568\text{ cal}}{} \times \frac{4.184\text{ J}}{1.000\text{ cal}}$$

(2) 731.5 J = 174.8 cal

$$\frac{731.5\text{ J}}{} \times \frac{1.000\text{ cal}}{4.184\text{ J}}$$

$Q = mc_p\Delta T$

(3) How many calories are needed to warm 75.00 g of water from 22.00 °C to 38.00 °C?

$$38.00 - 22.00 = 16.00\ °C$$

$$\frac{75.00\text{g}}{} \times \frac{1\text{ cal}}{\text{g }°C} \times \frac{16.00\ °C}{} = \boxed{1200.\text{ cal}}$$

$Q = mc_p\Delta T$

(4) How many joules of energy are needed to warm 353.0 g of water from 18.00 °C to 95.00 °C?

$$\Delta T \quad 95.00 - 18.00 = 77.00\ °C$$

$$\frac{353.0\text{g}}{} \times \frac{4.184\text{ J}}{\text{g }°C} \times \frac{77.00\ °C}{} = \boxed{113700\text{ J}}$$

$Q = mc_p\Delta T$

(5) How many calories are released as 53.00 g of water cools from 82.00 °C to 15.00 °C?

$$\Delta T \quad 82.00 - 15.00 = 67.00\ °C$$

$$\frac{53.00\text{g}}{} \times \frac{1\text{ cal}}{\text{g }°C} \times \frac{67.00\ °C}{} = \boxed{3551\text{ cal}}$$

$\Delta T = \frac{Q}{mc_p}$

(6) How hot can you heat 137.0 g of water at 51.00 °C with 3050. calories of heat energy?

$$\frac{3050.\text{ cal}}{137.0\text{ g}} \times \frac{\text{g }°C}{1\text{ cal}} = 22.26\ °C$$

$$51.00 + 22.26 = \boxed{73.26\ °C}$$

$\Delta T = \frac{Q}{mc_p}$

(7) What will the final temperature of water be if 45.00 g are heated with 3.500 X 10^2 joules of energy? The water is initially at 25.00 °C.

$$\frac{3.500\times10^2\text{ J}}{45.00\text{g}} \times \frac{\text{g }°C}{4.184\text{ J}} = 1.859$$

$$25.00 + 1.859 = \boxed{26.86\ °C}$$