More Calorie Calculation Practice   Name: KEY   Period:

For each of the situations below:

- Think about what is happening to the ice/water/steam.
- Decide which equation makes the most sense to use → write the original equation!
- Identify what each of the numbers means.
- Plug in numbers with units and solve, answer with the correct number of sig figs.

1. How much energy is required to raise 125 g of liquid water from 20.°C to 100.°C?

$m = 125g$
$\Delta T = 100°C - 20°C = 80°C$
$Q = ?$

$Q = mc\Delta T$
$Q = 125g \times 1.0 \frac{cal}{g°C} \times 80.°C = \boxed{10000.\,cal \text{ or } 1.0 \times 10^4\,cal}$

2. Once that liquid water (from #1) is at 100.°C, how much energy is required to vaporize all of it?

$Q = mH_v$
$Q = ?$
$H_v = 539.4 \frac{cal}{g}$

$Q = 125g \times 539.4 \frac{cal}{g}$
$Q = 67425\,cal = \boxed{67400\,cal}$

3. If you add 3750. cal of energy to solid ice at 0.0°C, how many grams (mass) of ice can be melted?

$Q = mH_f$
$Q = 3750.\,cal$
$m = ?$

$3750.\,cal = m \times 79.72 \frac{cal}{g}$
$m = 47.0396g = \boxed{47.04g}$

4. Once the ice in #3 is melted (at 0.0°C), if you add another 3750. cal of energy to it, what temperature will the liquid water be? $Q = mc\Delta T$

$Q = 3750\,cal$
$m = 47.04g$
$\Delta T = ?$

$3750.\,cal = 47.04g \times 1 \frac{cal}{g°C} \times \Delta T$
$\Delta T = 79.72°C$

$\Delta T = T_f - T_i$
$79.72°C = T_f - 0.0°C$
$\boxed{T_f = 79.72°C}$

5. How many calories are needed to raise 38 g of liquid water by 40.°C?

$Q = ?$
$m = 38g$
$\Delta T = 40.°C$

$Q = mc\Delta T$
$Q = 38g \times 1 \frac{cal}{g°C} \times 40.°C$ $Q = 1520\,cal = \boxed{1500\,cal}$

6. If you add 4500 cal of energy to liquid water that is already at 100.°C, how many grams of water can be vaporized?

$Q = mH_v$
$Q = 4500\,cal$
$m = ?$

$4500\,cal = m \times 539.4 \frac{cal}{g}$
$m = 8.34260g = \boxed{8.3g}$

7. How much energy does it take to melt 1 g of ice at 0°C? How much energy does it take to vaporize 1 g of water at 100°C? What is the name of each of those constants (the number you just found)?

melt: $79.72 \frac{cal}{g}$ → heat of fusion   vaporize: $539.4 \frac{cal}{g}$ → heat of vaporization

8. Propose one reason why the two numbers in #7 are so different from one another. (hint: think about why one phase change might take more energy)

It takes much more energy to vaporize 1 g of a material (pull the particles completely apart from one another) than to melt 1 g of a material (let the particles slide past one another)

9. If you remove 3200 cal from 58 g of liquid water, how much will the temperature decrease?

$Q = 3200\,cal$
$m = 58g$

$Q = mc\Delta T$
$3200\,cal = 58g \times 1 \frac{cal}{g°C} \times \Delta T$

$\Delta T = 55.17°C$
$\Delta T = -55°C$ (a decrease)

10. If the water in #9 started at 95˚C, what temperature will it end at?

$\Delta T = T_f - T_i$ $\quad -55°C = T_f - 95°C$

$\boxed{T_f = 40°C}$

11. How much energy needs to be removed to condense (the opposite of vaporize) 85 g of water vapor at 100.˚C?

$m = 85g$
$Q = ?$
$H_c = -539.4 \frac{cal}{g}$

$Q = mH_c$
$Q = 85g \times -539.4 \frac{cal}{g}$

$Q = -45849 cal = \boxed{-46000 cal \text{ ↑removed}}$

12. How much energy needs to be removed to freeze (the opposite of melt) 85 g of liquid water at 0.0˚C?

$Q = mH_s$
$Q = 85g \times -79.72 \frac{cal}{g}$ $\quad Q = -6776.2 cal = \boxed{-6800 cal \text{ ↑removed}}$

13. How much energy needs to be removed to cool 46 g of liquid water from 75˚C to 25˚C?

$Q = mc\Delta T$
$m = 46g$
$\Delta T = -50°C$

$Q = 46g \times 1 \frac{cal}{g°C} \times -50°C$
$Q = \boxed{-2300 cal}$

14. If you have 57 g of solid ice at 0.0˚C and you want to melt all of it, raise it all to 100.˚C, and then vaporize all of it, how much energy must you add? (hint: think about what is happening to the ice/water/steam)

$Q = mH_f = 57g \times 79.72 \frac{cal}{g} = 4544.04 cal$
$Q = mc\Delta T = 57g \times 1 \frac{cal}{g°C} \times 100°C = 5700 cal$
$Q = mH_v = 57g \times 539.4 \frac{cal}{g} = 30745.8 cal$

$= 40989.84 cal$
$= \boxed{41000 cal}$

15. If you have 28 g of water vapor at 100.˚C and you want to condense all of it, cool it all to 0.0˚C, and freeze all of it, how much energy must be removed? (hint: think about what is happening to the ice/water/steam)

$Q = mH_c = 28g \times -539.4 \frac{cal}{g} = -15103.2 cal$
$Q = mc\Delta T = 28g \times 1 \frac{cal}{g°C} \times -100°C = -2800 cal$
$Q = mH_s = 28g \times -79.72 \frac{cal}{g} = -2232.16 cal$

$= -20135.36 cal$
$\boxed{= -20000 cal}$

16. Find the Physical Constants for Water in your data book. <u>There is a different "c" value (specific heat capacity) for solid water (ice).</u> What is that value? If you have 15 g of solid ice at -14˚C and you want to raise the temperature to 0.0˚C how much energy will you have to add?

$Q = mc\Delta T$
$m = 15g$
$c = 0.502 \frac{cal}{g}$
$\Delta T = 14°C$

$Q = 15g \times 0.502 \frac{cal}{g°C} \times 14°C$
$Q = 105.42 cal \approx \boxed{110 cal}$

17. Challenge ☺: You have 43 g of solid ice at 0.0˚C. You add 4300 cal of energy to it. Describe the status of the ice/water/steam at the end. Is it all the same state? What state is that? What temperature is it?

$Q = mH_f = 43g \times 79.72 \frac{cal}{g} = 3427.96 cal$ needed to melt all the ice
$4300 - 3427.96 = 872.04 cal$ left to add after it has melted
$Q = mc\Delta T$ $\quad 872.04 cal = 43g \times 1 \frac{cal}{g°C} \times \Delta T$ $\quad \Delta T = 20.28°C$ *started at 0°C so final is $\boxed{20.°C}$