- Exam 2 solutions are on my website.
- I've included working for multiple-choice problems too.

Next exam: Ch 5-7

Heat capacity : C

C is the heat req'd to increase temp by 1°c.

$$q = C \times \Delta t$$

Δt = Change in temp : $\Delta t = t_{FINAL} - t_{INITIAL}$

ex: if $t_{Final} = 21°c$, $t_{INIT} = 22°c$

$$\Delta t = 21°c - 22°c = -1°c$$

Specific heat capacity, s

s = heat required to raise temp of 1g of an object by 1°c.

$$q = m \cdot s \cdot \Delta t$$

q: heat (J)

m: mass (g)

s: $\left(\frac{J}{g \cdot °c}\right)$

Δt: change in temp (°c)

ex: Au : $s = 0.129\ J/g \cdot °c$

H_2O : $s = 4.184\ J/g \cdot °c$

ex: 12g of H_2O.

What will its temp. increase be if it absorbs 480J of heat?

ex: 12g of Au

what will its temp increase be if it absorbs 480J of heat?

Table 6.2

The Specific Heats of Some Common Substances

Substance	Specific Heat (J/g · °C)
Al	0.900
Au	0.129
C (graphite)	0.720
C (diamond)	0.502
Cu	0.385
Fe	0.444
Hg	0.139
H_2O	4.184
C_2H_5OH (ethanol)	2.46

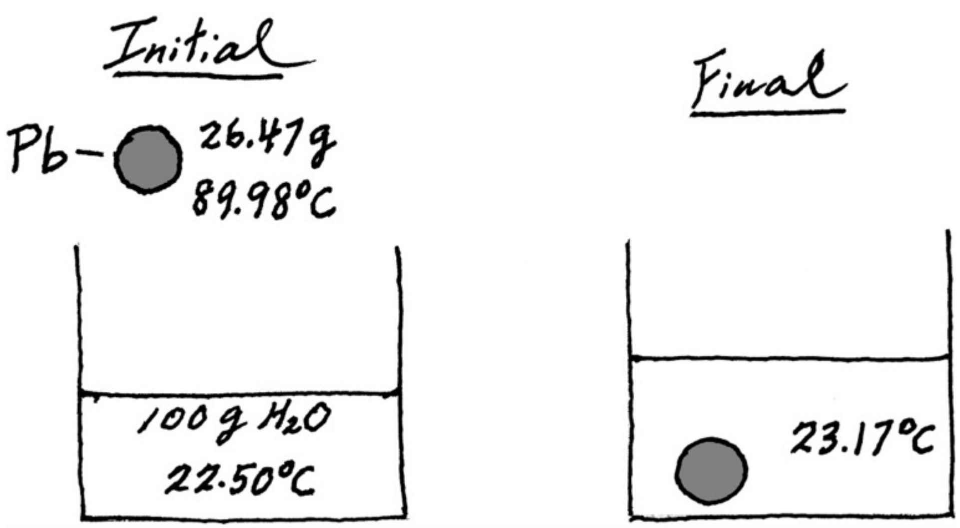
$\underline{H_2O}$ $q = m \cdot s \cdot \Delta t \Rightarrow \Delta t = \dfrac{q}{m \cdot s}$

$\Rightarrow \Delta t = \dfrac{+480\cancel{J}}{12\cancel{g} \cdot 4.184\ \cancel{J}/\cancel{g}\cdot°C} = +9.6\ °C$

$\underline{Au}$ $\Delta t = \dfrac{q}{m \cdot s} = \dfrac{+480J}{12g \cdot 0.129\ J/g\cdot°C} = +310\ °C$

ex: A lead BB with a mass of 26.47g at a temp. of 89.98°C was placed into 100.0mL of water (in an insulated container). The final temp. of the water + lead is 23.17°C. If the initial water temp was 22.50°C, then what must the specific heat capacity of Pb be?

$S_{H_2O} = 4.184\ J/g\cdot°C$



1st Law of thermo: Energy cannot be created nor destroyed.

$\Rightarrow$ heat lost by the lead = heat gained by water!

$$\underset{\ominus}{q_{lead}} + \underset{\oplus}{q_{water}} = 0$$

$$q = m \cdot s \cdot \Delta t$$

$$q_{H_2O} = m_{H_2O} \cdot s_{H_2O} \cdot \Delta t_{H_2O}$$

$$= 100.0g \times 4.184 \frac{J}{g \cdot °C} \times \overbrace{(23.17°C - 22.50°C)}^{0.67°C}$$

water: 1g ≈ 1mL

$d_{H_2O} \approx 1 g/mL$

$$\Rightarrow q_{H_2O} = +280 \text{ J}$$

$$\Rightarrow q_{Pb} = -280 \text{J}$$

$$q_{Pb} = m_{Pb} \times s_{Pb} \times \Delta t_{Pb}$$

$$\Rightarrow -280\,J = 26.47\,g \times S_{Pb} \times \underbrace{(23.17^\circ C - 89.98^\circ C)}_{-66.81^\circ C}$$

$$\Rightarrow S_{Pb} = \frac{\cancel{-}280\,J}{26.47\,g \times \cancel{-}66.81^\circ C}$$

$$= 0.16 \ \frac{J}{g \cdot {}^\circ C}$$

Enthalpy of chemical rxns

H (state function)

Changes in Enthalpy: ΔH

changes in state functions are independent on how the process is carried out!

$$\Delta H = q_P \neq q_V$$

const. V.

heat

under conditions of constant pressure.

Thermochemical equations...

ex: $H_2O(s) \rightarrow H_2O(l)$; $\Delta H = +6.01 \frac{kJ}{mol}$

endothermic

coefficients ARE the # moles, not just the ratios!

1 mol $H_2O(s)$
OR
1 mol $H_2O(l)$ $\longrightarrow$ per mole of reaction

ex: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$;

exothermic

$\Delta H = -890.4 \frac{kJ}{mol}$

per 1 mol $CH_4(g)$
OR 2 mol $O_2(g)$
OR 1 mol $CO_2(g)$
OR 2 mol $H_2O(l)$ $\longrightarrow$ per mole of rxn

ΔH is a conversion factor that links mol $\leftrightarrow$ energy!

ex: $-890.4\ kJ = 1\ mol\ CH_4$

$-890.4\ kJ = 2\ mol\ O_2$

$= 1\ mol\ CO_2$

$= 2\ mol\ H_2O$.

ex: What's q is 3.8 mol H_2O is made?

$$\frac{3.8\ mol\ \cancel{H_2O}}{} \Big| \frac{-890.4\ kJ}{2\ mol\ \cancel{H_2O}} = \underbrace{-1700\ kJ}_{}$$

1700 kJ of heat is released.

ex: Consider the rxn:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3 \ ; \ \Delta H = +89\ \frac{kJ}{mol}$$

Q. What's q_p is 1.7 mol H_2 is reacted?

Q. What's q_p if 1.7 mol NH_3 is formed?

$$\frac{1.7\ mol\ H_2}{} \Big| \frac{+89\ kJ}{3\ mol\ H_2} = +50.\ kJ$$

$$\frac{1.7\ mol\ NH_3}{} \Big| \frac{89\ kJ}{2\ mol\ NH_3} = +76\ kJ.$$

Hints when using thermochemical eqs.

(1) Must show state symbols!

ex: $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$

Diff't ΔH!

(2) Multiplying the chemical equation by a constant, multiplies ΔH by same amount!

ex: $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \ ; \ \Delta H = -20\ \frac{kJ}{mol}$

$3H_2(g) + \frac{3}{2}O_2(g) \longrightarrow 3H_2O(g) \ ; \ \Delta H = -60\ \frac{kJ}{mol}$

(3) Reversing the chemical equation, reverses the sign of ΔH.

ex : $H_2O(g) \longrightarrow H_2(g) + \frac{1}{2}O_2(g)$; $\Delta H = +20 \frac{kJ}{mol}$