0. How much heat is required to raise the temperature of 250.0 g of Mercury from 20°C to 57°C? The specific heat of mercury is .14 J/g °C.

$.14(250)(57-20) = q$

$q = \boxed{1295\ J}$

10. When 435 J of heat is added to 3.4 g of olive oil at 21 °C, the temperature increases to 85 °C. What is the specific heat of olive oil?

$435\ J = 3.4(c)(85-21)$

$\boxed{c = 1.99}$

11. A sample of silver with a mass of 63.3 g is heated to a temperature of 101.3°C and placed in a container of water at 25.85°C. The final temperature of the silver and water is 27.25°C. Assuming no heat loss, what mass of water was in the container? The specific heat of silver is 0.235 (J/g·°C).

$63.3g(.235)(27.25-101.3) = -1101.5\ J$

$\frac{1101.5\ J}{4.18} = \frac{4.18(m)(27.25-25.85)}{4.18}$

$263.5 = m(27.25-25.85)$

$H_2O = 18g/mol$ $\boxed{m = 188.23g}$

11. A 13.5 g sample of gold is heated, then placed in a calorimeter containing 60.0 g of water. Temperature of water increases from 19.00 °C to 20.00 °C. The specific heat of gold is 0.130 J/g°C. What was the initial temperature of the gold metal sample?

$\frac{60(4.18)(1)}{(13.5 \cdot .13)} = \frac{13.5 \cdot .13(20 - T_i)}{(13.5 \cdot .13)}$

$142.9 = 20 - T_i$

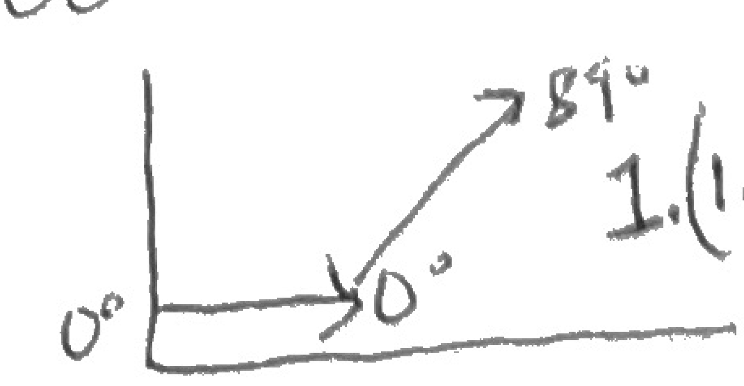
$\boxed{T_i = -122.94}$

12. How much energy is needed to convert 24 g ice at 0°C to water at 89 °C?

$6.01\ kJ \cdot (24/18) = 8.01\ kJ$

$(4.18)(24)(89) = + \frac{8,928.48\ J}{1000} = 8.93\ kJ$

$\boxed{16.94\ kJ}$



12. How much energy is released when 46 g of steam at 420 K is converted to water at 360 K?

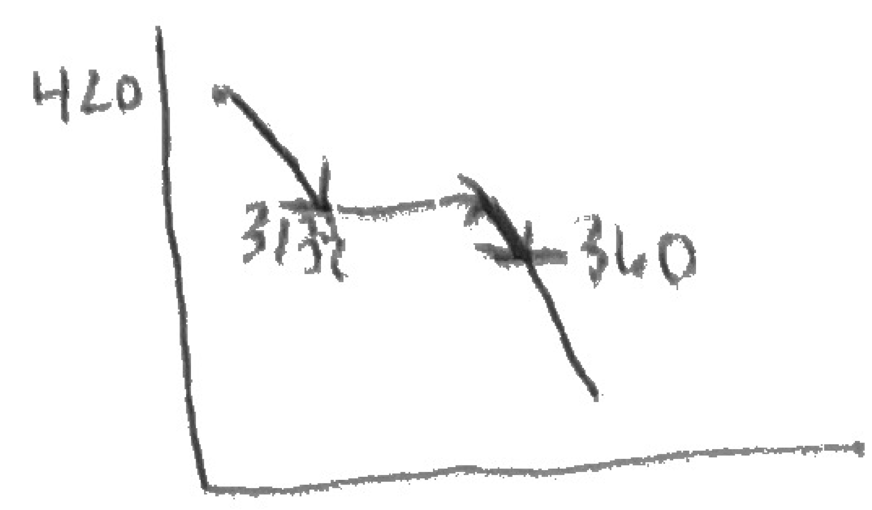
g → l

2. $-40.6\ kJ(46/18) = -103,755.6\ kJ$

1. $(1.84)(46)(373-420) = -\frac{3978.1\ J}{1000} = -3.98\ kJ$

3. $4.18(46)(360-373) = -\frac{2499.6\ J}{1000} = 2.50\ kJ$

$\boxed{-110.23\ kJ}$



Specific Heat Values	Molar Heat of Fusion
$H_2O_{(s)}$ = 2.09J/g°C	H_2O = 6.01 kJ/mol
$H_2O_{(l)}$ = 4.18 J/g°C	Molar Heat of Vaporization
$H_2O_{(g)}$ = 1.84 J/g°C	H_2O = 40.6 kJ/mol

13. Calculate the standard enthalpy change, ΔH°, for the formation of 1 mol of strontium carbonate (the material that gives the red color in fireworks) from its elements.

$Sr(s) + C(graphite) + \frac{3}{2}O_2(g) \rightarrow SrCO_3(s)$

The information available is

(1) $Sr(s) + \frac{1}{2}O_2(g) \rightarrow SrO(s)$ ΔH° = -592 kJ

(2) $SrO(s) + CO_2(g) \rightarrow SrCO_3(s)$ ΔH° = -234 kJ

(3) $C(graphite) + O_2(g) \rightarrow CO_2(g)$ ΔH° = -394 kJ

$\boxed{-1220 kJ}$

13. Calculate the standard enthalpy change, ΔH°, for the following reaction:

$2B(s) + 3H_2(g) \rightarrow B_2H_6(g)$

a) $4 B(s) + 3 O_2(g) \rightarrow 2 B_2O_3(s)$ ΔH° = -2509.1 kJ

b) $2 H_2(g) + O_2(g) \rightarrow 2 H_2O(l)$ $\frac{3}{2}$ ΔH° = -571.7 kJ

c) $B_2H_6(g) + 3 O_2(g) \rightarrow B_2O_3(s) + 3 H_2O(l)$ ΔH° = -2147.5 kJ

$B_2O_3 + 3H_2O \rightarrow B_2H_6 + 3O_2$ + 2147.5 kJ

-1,254.5
-857.55
+ 2147.5

$\boxed{\Delta H\ 35.45}$

Work this problem together.

Consider the following skeletal reaction which occurs under standard thermodynamic conditions:

4 NH_3 (g) + 1 O_2 (g) → 2 N_2H_4 (g) + 2 H_2O (l)

a) Use standard enthalpy tables to determine ΔH° for this reaction.

$[2(95.4) + 2(-285.8)] - [4(-46.1)] = \boxed{-196.4 kJ}$

-380.8 - (-184.4)

b) Use standard entropy values to determine ΔS° for this reaction.

$[2(238.4) + 2(70)] - [4(192.3) + 2(205.0)]$

616.8 - 974.2 $\Delta S = \boxed{-357.4 J}$

c) Determine ΔG for this reaction at 300K. Is the reaction spontaneous or not at this temperature?

$\Delta G = \Delta H - T\Delta S$ $\Delta G = (-196.4)_{kJ} - \left[\frac{300(-357.4)_J}{1000}\right]$ $\boxed{-89.2}$

$(-196.4) - (-107.2)$

d) Will your answer to C be true at every temperature? Explain your reasoning.

No, because ΔG = ΔH - TΔS and depends on the temperature the subtracted value could be increased and eventually flip the sign.

Formula	State of Matter	ΔH Enthalpy (kJ/mol)	ΔS Entropy (J mol/K)	ΔG Gibbs Free Energy (kJ/mol)
N_2H_4	(g)	95.3952	238.36248	159.28488
O_2	(g)	0	205.028552	0
NH_3	(g)	-46.10768	192.33848	-16.48496
H_2O	(l)	-285.82996	69.91464	-237.178408