1. Calculate the standard enthalpy change, $\Delta H°$, 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) $Sr(s) + \frac{1}{2}O_2(g) \rightarrow SrO(s)$ $\Delta H° = -592$ kJ
(2) $SrO(s) + CO_2(g) \rightarrow SrCO_3(s)$ $\Delta H° = -234$ kJ
(3) $C(graphite) + O_2(g) \rightarrow CO_2(g)$ $\Delta H° = -394$ kJ

Keep all the same

$Sr + \frac{1}{2}O_2 + SrO + CO_2 + C_{(graphite)} + O_2 \rightarrow SrO + SrCO_3 + CO_2$ $\Delta H = -592 + -234 + -394$

$3/2 O_2$

$Sr + 3/2 O_2 + C_{(graphite)} \rightarrow SrCO_3$ $\boxed{\Delta H = -1220 \text{ kJ}}$

2. The combination of coke and steam produces a mixture called coal gas, which can be used as a fuel or as a starting material for other reactions. If we assume coke can be represented by graphite, the equation for the production of coal gas is
$2C(s) + 2H_2O(g) \rightarrow CH_4(g) + CO_2(g)$

Determine the standard enthalpy change for this reaction from the following standard enthalpies of reaction :

(1) $C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)$ $\Delta H° = 131.3$ kJ X2
(2) $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$ $\Delta H° = -41.2$ kJ Don't change
(3) $CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$ $\Delta H° = 206.1$ kJ flip

① $2C + 2H_2O \rightarrow 2CO + 2H_2$ $\Delta H = 131.3 \times 2 = 262.6$
② $CO + H_2O \rightarrow CO_2 + H_2$ $\Delta H = -41.2$
③ $3H_2 + CO \rightarrow CH_4 + H_2O$ $\Delta H = -206.1$

$2C + 2H_2O + \cancel{CO} + \cancel{H_2O} + \cancel{3H_2} + \cancel{CO} \rightarrow \cancel{2CO} + \cancel{2H_2} + CO_2 + \cancel{H_2} + CH_4 + \cancel{H_2O}$

$2C + 2H_2O \rightarrow CO_2 + CH_4$ $\Delta H = 427.1$ kJ

3. One reaction involved in the conversion of iron ore to the metal is
$FeO(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$

Calculate the standard enthalpy change for this reaction from these reactions of iron oxides with CO :

(1) $3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$ $\Delta H° = -47.0$ kJ flip and ?
(2) $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$ $\Delta H° = -25.0$ kJ ÷2
(3) $Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$ $\Delta H° = 19.0$ kJ flip and ÷3

①
② $(1/2)Fe_2O_3 + (3/2)CO \rightarrow Fe + (3/2)CO_2$ $\Delta H = 25 \div 2 = 12.5$
③ $FeO + (1/3)CO_2 \rightarrow (1/3)Fe_3O_4 + (1/3)CO$ $\Delta H = -19 \div 3 = -6.3$

4. Calculate the enthalpy of reaction for the double replacement reaction that occurs between barium chloride and silver nitrate. Is this reaction endothermic or exothermic?

$BaCl_2 + 2AgNO_3 \rightarrow Ba(NO_3)_2 + 2\,AgCl$

$\Delta H = (Ba(NO_3)_2 + 2 \cdot AgCl) - (BaCl_2 + 2 \cdot AgNO_3)$

$\Delta H = (-873.2 + 2(-100.7)) - (-951.4 + 2(-127.1))$

$\Delta H = \boxed{-130.8\ kJ}$ $\boxed{\text{exothermic}}$

5. Make the following conversions using the factor label method:

a. 39.2 kJ = ______calories

$39.2\ kJ \times \frac{1000\ J}{1\ kJ} \times \frac{1\ cal}{4.184\ J} = \boxed{9370\ cal}$

b. 42 Calories = ______ joules

$42\ Cal \times \frac{1000\ cal}{1\ Cal} \times \frac{4.184\ J}{1\ cal} = \boxed{1.8 \times 10^5\ J}$

6. If all 20.7 kJ of energy is added to 11.50g of ice at -10.0°C, what would be the final temperature of the $H_2O$? What phase would it be in?

Heat the ice to 0° = $Q = Cm\Delta T = (2.06\ J/g°C)(11.5)(0 - -10) = 237\ J$

Melt the ice = $(\Delta H_{fus})(mass) = (334\ J/g)(11.5) = 3841\ J$

Heat the water to 100° → $Q = Cm\Delta T = (4.184)(11.5)(100-0) = 4812\ J$

Total energy used so far:

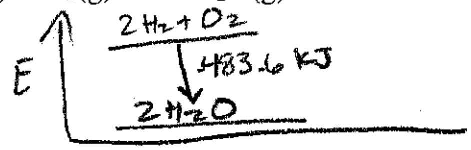
$237 + 3841 + 4812 = 8890\ kJ$

Energy remaining = $20700 - 8890 = 11810\ kJ$

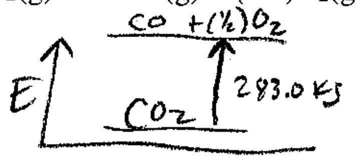
Boil the water completely = $(\Delta H_{vap})(mass) = 2260 \times 11.5 = 25990$

We don't have enough energy to completely vaporize the water, so after 20.7 kJ is added, the water will be $\boxed{\text{boiling at } 100°C}$

7. Draw an energy diagram for each of these two reactions:
   a. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$ $\Delta H = -483.6$ kJ




   b. $CO_2(g) \longrightarrow CO(g) + (1/2)O_2(g)$ $\Delta H = 283.0$ kJ




8. A piece of copper (C=0.386J/g°C) is heated to 215.00°C and dropped into 42.8mL of water at 5.00°C. If the final temperature of the system is 7.50°C, what is the mass of the piece of copper?

$$\underbrace{C m \Delta T}_{\text{water}} = -\underbrace{C m \Delta T}_{\text{metal}}$$

$$(4.184)(42.8g)(7.5-5.0) = -(0.386 J/g°C)(mass_{copper})(7.5-215)$$

mass copper = $\boxed{5.59 g}$

9. Is entropy increasing or decreasing in these reactions:
   a. $2C(s) + O_2(g) \longrightarrow 2CO(g)$ Increasing
      1 gas  2 gas
   b. $2H_2(g) + O_2(g) \longrightarrow 2H_2(g)$ Decreasing
      3 gas  2 gas
   c. $H_2O(s) \longrightarrow H_2O(l)$ Increasing
      solid  liquid
   d. $NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$ Increasing
      dissolving

10. $2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$ $\Delta H = -1271.0$ kJ
    How much energy is released when 55.7 grams of Calcium is converted to calcium oxide?

$$55.7 g\,Ca \times \frac{1\ mol\ Ca}{40.08 g\,Ca} \times \frac{-1271.0 kJ}{2 mol\,Ca} = \boxed{-883 kJ}$$