

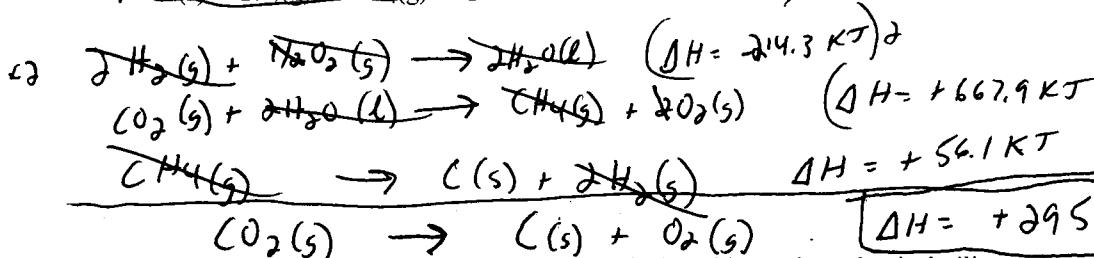
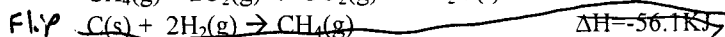
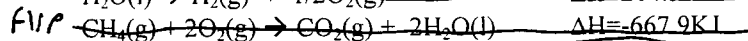
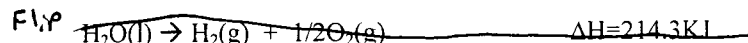
Section 16.1 Worksheet

Directions: Solve the following problems. You will need the reference charts we have been using throughout the chapter to solve some problems.

1. Calculate the amount of heat produced when 11.94 grams of nitric oxide, NO, burn in air to produce nitrogen dioxide according to the following equation: $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2 + 113.04 \text{ kJ}$

$$\frac{11.94 \text{ g NO}}{1} \times \frac{1 \text{ mol NO}}{30.0 \text{ g NO}} \times \frac{113.04 \text{ kJ}}{2 \text{ mol NO}} = \boxed{22.49 \text{ kJ}}$$

2. Calculate the heat change involved in the reaction $\text{CO}_2(\text{g}) \rightarrow \text{C}(\text{s}) + \text{O}_2(\text{g})$ using the information below.



3. Calculate the amount of heat released when 16.8 grams of ethanol is condensed at its boiling point.

$$q = n \Delta H_{\text{vap}}$$

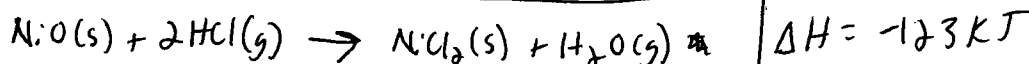
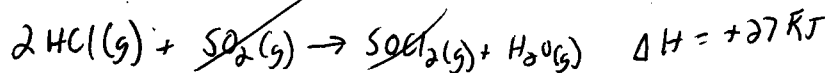
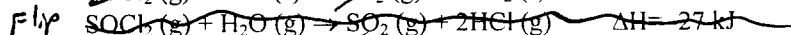
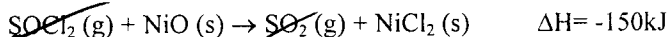
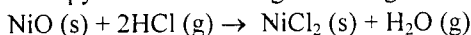
$$q = (0.3647 \text{ mol})(38.6 \text{ kJ/mol})$$

ethanol ($\text{C}_2\text{H}_5\text{OH}$)
 $\Delta H_{\text{vap}} = 38.6 \text{ kJ/mol}$

$$\frac{16.8 \text{ g C}_2\text{H}_5\text{OH}}{1} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.07 \text{ g C}_2\text{H}_5\text{OH}} = 0.3647 \text{ mol}$$

$$q = \boxed{14.08 \text{ kJ}}$$

4. Calculate the change in enthalpy for the following reaction given the information below.



5. Calculate the mass of water produced when methane is burned in excess air producing 1546 kJ of heat according to the following equation: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890.2 \text{ kJ}$

$$\frac{1546 \text{ kJ}}{1} \times \frac{1 \text{ mol CH}_4}{890.2 \text{ kJ}} \times \frac{16.043 \text{ g CH}_4}{1 \text{ mol CH}_4} = \boxed{27.86 \text{ g CH}_4}$$

I don't want + methane though

$$\frac{1546 \text{ kJ}}{1} \times \frac{2 \text{ mol H}_2\text{O}}{890.2 \text{ kJ}} \times \frac{18.015 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{62.57 \text{ g H}_2\text{O}}$$

This is the Answer.

6. Determine the energy needed when 55.6 grams of aluminum at 43.2°C is heated to 78.1°C .

$$Q = m \times \Delta T \times C$$

$$Q = (55.6 \text{ g})(34.9^\circ\text{C})(0.897 \text{ J/g}^\circ\text{C}) = \boxed{1740.6 \text{ J}}$$

$$\Delta H_{fus} = 5.65 \text{ kJ/mol}$$

7. If a sample of ammonia, NH_3 , required 18.2 kJ of energy to melt it, what is the mass of the sample?

$$q = n \Delta H_{\text{melt}}$$

$$\frac{18.2 \text{ kJ}}{5.65 \text{ kJ/mol}} = n \left(\frac{5.65 \text{ kJ/mol}}{5.65 \text{ kJ/mol}} \right)$$

$$n = \frac{3.22 \text{ mol NH}_3}{1} \times \frac{17.01 \text{ g NH}_3}{1 \text{ mol NH}_3} = \boxed{54.77 \text{ g NH}_3}$$

8. A student places 42.3 grams of ice at 0.0°C in an insulated bottle. The student adds 255.8 grams of water at 90.0°C . Determine the final temperature of the mixture. *Harder

<p>ice</p> <p>$m = 42.3 \text{ g}$</p> <p>$T_i = 0^\circ\text{C}$</p> <p>$T_f = ?$</p>	<p>water</p> <p>$m = 255.8 \text{ g}$</p> <p>$T_i = 90^\circ\text{C}$</p> <p>$T_f = ?$</p>	<p>$q_{\text{ice}} = q_{\text{water}}$</p> <p>two steps</p> <p>$q [n \Delta H_{\text{fus}}] + [m \Delta T \times C] = m \Delta T \times C$</p> <p>$[2.35 \text{ mol} (6.01 \text{ kJ/mol})] + [42.3 \text{ g} (x) (4.18 \text{ J/g}^\circ\text{C})] = [(255.8 \text{ g})(90^\circ\text{C} - x)(4.18 \text{ J/g}^\circ\text{C})]$</p> <p>$14,115 \text{ J} + 176.8 \text{ J/g}^\circ\text{C} x = 1069.2 \text{ J/g}^\circ\text{C} (90^\circ\text{C} - x)$</p> <p>$14,115 \text{ J} + 176.8 \text{ J/g}^\circ\text{C} x = 96232 \text{ J} - 1069.2 \text{ J/g}^\circ\text{C} x$</p>	<p>$\boxed{65.9^\circ\text{C}}$</p>
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9. Calculate the amount of energy released when converting 34.0 grams of steam at 100.0°C to liquid water at 20.0°C .

① Steam $100^\circ\text{C} \rightarrow$ water 100°C

$q = n \Delta H_{\text{vap}}$

$q = (1.888 \text{ mol})(40.67 \text{ kJ/mol})$

$q = 76.78 \text{ kJ}$

② water $100^\circ\text{C} \rightarrow$ water 20°C

$q = m \Delta T \times C$

$q = (34 \text{ g})(80^\circ\text{C})(4.18 \text{ J/g}^\circ\text{C})$

$q = 11369.6 \text{ J}$

$q = 11.369 \text{ kJ}$

$\frac{34.0 \text{ g}}{18.015 \text{ g/mol}} = 1.888 \text{ mol}$

$76.78 \text{ kJ} + 11.369 \text{ kJ} = \boxed{88.15 \text{ kJ}}$

10. It takes 565 Joules of heat to raise a sample of iron from 21.0°C to 75.4°C . What is the mass of the sample?

$q = 565 \text{ J}$

$\Delta T = 54.4^\circ\text{C}$

$m = ?$

$C = 0.449 \text{ J/g}^\circ\text{C}$

$q = m \Delta T \times C$

$565 \text{ J} = m (54.4^\circ\text{C})(0.449 \text{ J/g}^\circ\text{C})$

$m = \frac{565 \text{ J}}{(54.4^\circ\text{C})(0.449 \text{ J/g}^\circ\text{C})} = \boxed{m = 23.13 \text{ g}}$

11. A piece of stainless steel weighing 1.55 g absorbs 141 J of heat energy when its temperature increases from 22°C to 200°C . What is the specific heat of stainless steel?

$m = 1.55 \text{ g}$

$q = 141 \text{ J}$

$\Delta T = 178^\circ\text{C}$

$C = ?$

$C = \frac{q}{m \Delta T}$

$C = \frac{141 \text{ J}}{(1.55 \text{ g})(178^\circ\text{C})} = \boxed{0.511 \text{ J/g}^\circ\text{C}}$

12. Calculate the change in enthalpy of $\text{PCl}_3(\text{l}) + \text{Cl}_2(\text{g}) \rightarrow \text{PCl}_5(\text{s})$ given the following reactions:

