

## Heat and $q=mc\Delta T$

1. How much heat is required to warm 400. g of ethanol from 25.0°C to 40.0°C?
2. What mass of water can be heated from 0.00°C to 25.0°C with 90,000. J of energy?
3. If 7,500. J of energy is added to 1.43 kg of iron (originally at 25.0°C), what temperature will it rise to?
4. What is the temperature change for 500. g of each metal below, when 1,000. J of energy is added to each?
  - a) Aluminum
  - b) Lead

5. How much heat is released to the surroundings when a cloud containing 100. kg of air cools from 34.0°C to 28.0°C?

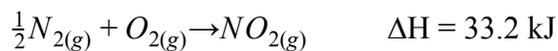
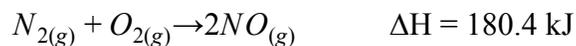
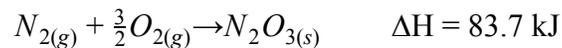
Substance	Heat Capacity (J / g°C)
H <sub>2</sub> O <sub>(l)</sub>	4.184
H <sub>2</sub> O <sub>(s)</sub>	2.029
Pb <sub>(s)</sub>	0.130
Fe <sub>(s)</sub>	0.452
Al <sub>(s)</sub>	0.891
Ethanol <sub>(l)</sub>	2.400
Air <sub>(g)</sub>	0.718

### Calculating Enthalpy Change ( $\Delta H$ ) given heat change

6. Find  $\Delta H$  for the dissolution of potassium chlorate ( $\text{KClO}_3$ ) if dissolving 1.2 g in 100 mL of water decreases the temperature from 25.00°C to 24.03°C.
  
7. Mixing 25.00 mL of 0.1000 M NaOH with 20.00 mL of 0.8000 M HCl causes the temperature of the mixture (initially 25.00°C) to increase to 29.92°C. What is the enthalpy of neutralization for this reaction?
  
8. Suppose you have 50.0 mL of 0.2 M  $\text{NaOH}_{(\text{aq})}$  and mix it with 50.0 mL of 0.2 M  $\text{HCl}_{(\text{aq})}$ . Both mixtures start at 25.00°C. The temperature change is measured and you calculate  $\Delta H_{\text{rxn}}$ . Answer the following with “doubles”, “halves” or “stays the same”:
  - a) What happens to Q if we use 100 mL of each solution?
  - b) What happens to Q if we use 100 mL of  $\text{NaOH}_{(\text{aq})}$  but still only 50 mL of  $\text{HCl}_{(\text{aq})}$ ?
  - c) What happens to  $\Delta T$  if we use 100 mL of each solution?
  - d) What happens to  $\Delta T$  if we use 150 mL of  $\text{NaOH}_{(\text{aq})}$  but still only 50 mL of  $\text{HCl}_{(\text{aq})}$ ?
  - e) What happens to  $\Delta H_{\text{rxn}}$  if we use 100 mL of each solution?
  - f) What happens to  $\Delta H_{\text{rxn}}$  if we use 150 mL of  $\text{NaOH}_{(\text{aq})}$  but still only 50 mL of  $\text{HCl}_{(\text{aq})}$ ?
  - g) What happens to  $\Delta H_{\text{rxn}}$  if we use 1850 mL of  $\text{NaOH}_{(\text{aq})}$  but only 0.2 mL of  $\text{HCl}_{(\text{aq})}$ ?

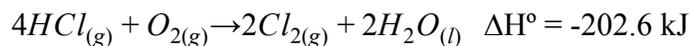
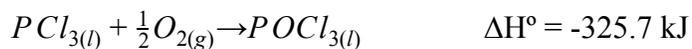
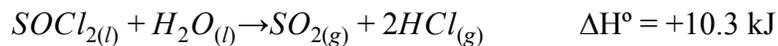
## Hess' Law

9. Given the following thermochemical equations:



What is  $\Delta H$  for the reaction  $N_2O_{3(s)} \rightarrow NO_{(g)} + NO_{2(g)}$ ?

10. Find the enthalpy change for this reaction:  $2SO_{2(g)} + 2P_{(s)} + 5Cl_{2(g)} \rightarrow 2SOCl_{2(l)} + 2POCl_{3(l)}$   
given:



## Enthalpies of Formation

11. Find the enthalpy change for this reaction:  $2NH_{3(g)} + \frac{7}{2}O_{2(g)} \rightarrow N_2O_{4(g)} + 3H_2O_{(g)}$  given the  $\Delta H_f^\circ$  values in the table below.
12. Find  $\Delta H_f^\circ$  for  $C_8H_{18}$  if the following reaction:  $C_8H_{18(l)} + \frac{25}{2}O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(l)}$  has  $\Delta H = -5430 \text{ kJ}$
13. Write equations for the chemical reactions that  $\Delta H_f$  represents for:
- a)  $COCl_{2(g)}$
- b)  $PF_{5(g)}$
- c)  $Ni(IO_3)_{3(s)}$
14. Why is  $\Delta H_f$  for liquid water  $-286 \text{ kJ/mol}$ , while  $\Delta H_f$  for water vapour is  $-242 \text{ kJ/mol}$ ?

$\Delta_f H^\circ / \text{kJ mol}^{-1}$					
$H_2O(l)$	- 286	$NO_2(g)$	+ 33	$CH_4(g)$	- 75
$H_2O(g)$	- 242	$N_2O_4(g)$	+ 9	$C_2H_6(g)$	- 85
$CO(g)$	- 111	$CaO(s)$	- 635	$C_2H_4(g)$	+ 52
$CO_2(g)$	- 394	$CaCO_3(s)$	- 1207	$C_2H_2(g)$	+ 227
$HCl(g)$	- 92	$Fe_2O_3(s)$	- 824	$C_2H_5OH(l)$	- 278
$NH_3(g)$	- 46	$NaCl(s)$	- 411	$C_6H_6(l)$	+ 49
$NO(g)$	+ 90	$BaSO_4(s)$	- 1473	$CH_2O(g)$	- 109

## Bond Enthalpies

15. Use the bond enthalpies in the table below to estimate  $\Delta H$  for  $H_2CO_{(g)} \rightarrow H_{2(g)} + CO_{(g)}$ .  
(Remember: You'll have to draw the Lewis Structures to know what types of bonds are there!)

16. Use bond enthalpies to estimate the  $\Delta H$  for  $C_2H_{2(g)} + 2F_{2(g)} \rightarrow C_2H_2F_{4(g)}$

17. Why are the  $\Delta H$  values from bond enthalpies called *estimates* whereas the values from Hess' Law and  $\Delta H_f^\circ$  are more accurate?

**Average Bond Enthalpies (kJ/mol)**

**Single Bonds**

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146	Cl—F	253
C—N	293	N—O	201	O—F	190	Cl—Cl	242
C—O	358	N—F	272	O—Cl	203		
C—F	485	N—Cl	200	O—I	234	Br—F	237
C—Cl	328	N—Br	243			Br—Cl	218
C—Br	276			S—H	339	Br—Br	193
C—I	240	H—H	436	S—F	327		
C—S	259	H—F	567	S—Cl	253	I—Cl	208
		H—Cl	431	S—Br	218	I—Br	175
Si—H	323	H—Br	366	S—S	266	I—I	151
Si—Si	226	H—I	299				
Si—C	301						
Si—O	368						

**Multiple Bonds**

C=C	614	N=N	418	O <sub>2</sub>	495
C≡C	839	N≡N	941		
C=N	615			S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

## Changes of State

18. What mass of water can be **melted** at  $0^{\circ}\text{C}$  ( $\Delta H_{\text{fus}} = +6.03 \text{ kJ/mol}$ ) with the energy released when 1 kg of ethanol ( $\Delta H_{\text{c}} = -2020 \text{ kJ/mol}$ ) is burned?

19. What amount of energy is required to heat 200 g of ice at  $-160^{\circ}\text{C}$  to steam at  $125^{\circ}\text{C}$  ?

$$c_{\text{ice}} = 2.01 \text{ J/g}^{\circ}\text{C}$$

$$\Delta H_{\text{fus}} = +6.03 \text{ kJ/mol}$$

$$c_{\text{water}} = 4.184 \text{ J/g}^{\circ}\text{C}$$

$$\Delta H_{\text{vap}} = +40.2 \text{ kJ/mol}$$

$$c_{\text{steam}} = 2.01 \text{ J/g}^{\circ}\text{C}$$

20. What are the types of thermal energy present in:

Solids

Liquids

Gases

### Potential Energy Diagrams

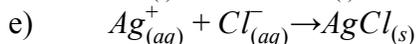
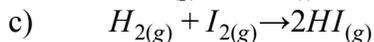
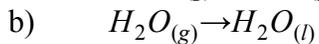
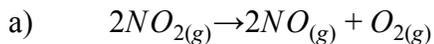
21. Draw a potential energy diagram for each of these reactions, and label  $\Delta H$  on each:
- The **exothermic** reaction  $\text{C}_2\text{H}_2 + 2 \text{H}_2 \rightarrow \text{C}_2\text{H}_6 + 312 \text{ kJ}$
  - The **endothermic** reaction  $3 \text{O}_2 + 285.4 \text{ kJ} \rightarrow 2 \text{O}_3$

### Working with Unit Conversions

22. Given that  $\text{NO}_2$  has an enthalpy of formation of  $+33 \text{ kJ/mol}$  :
- What is  $\Delta H_f$  in  $\text{kJ/g}$  ?
  - What is  $\Delta H_f$  in  $\text{J/molecule}$ ?
  - How many molecules can be “formed” with  $1 \text{ J}$  of energy?

### $\Delta S$ (entropy) and $\Delta G$ (Gibbs Free Energy and Spontaneity)

23. For each of these reactions, is  $\Delta S$  positive or negative?



24. Assuming  $\Delta H$  and  $\Delta S$  do not vary with temperature, at what temperature will this reaction become spontaneous?  $C_{(s)} + H_2O_{(g)} \rightarrow CO_{(g)} + H_{2(g)}$ , with  $\Delta H = 131.3$  kJ and  $\Delta S = 133.6$  J/K

*Hint: Use  $\Delta G = \Delta H - T\Delta S$*

25. Mercury freezes at  $-38.9^\circ C$ , and its enthalpy of fusion is 2.33 kJ/mol. Calculate the value of  $\Delta S$  for the freezing of 1.00 mol of  $Hg_{(l)}$  at  $-38.9^\circ C$ .

*Hint: Freezing at the "freezing point" is spontaneous, so  $\Delta G = 0$*

## Answers

A1. How much heat is required to warm 400. g of ethanol from 25.0°C to 40.0°C?

$$\begin{aligned}q &= mc\Delta T \\ &= (400 \text{ g})(2.400 \text{ J/g}^\circ\text{C})(40.0^\circ\text{C} - 25^\circ\text{C}) \\ &= 14,400 \text{ J} \\ &= 14.4 \text{ kJ}\end{aligned}$$

A2. What mass of water can be heated from 0.00°C to 25.0°C with 90,000. J of energy?

$$\begin{aligned}q &= mc\Delta T \\ m &= q / c\Delta T \\ &= (90,000 \text{ J}) / [(4.184 \text{ J/g}^\circ\text{C})(25.0^\circ\text{C})] \\ &= 860. \text{ g}\end{aligned}$$

A3. If 7,500. J of energy is added to 1.43 kg of iron (originally at 25.0°C), what temperature will it rise to?

$$\begin{aligned}q &= mc\Delta T \\ \Delta T &= q / mc \\ &= (7500 \text{ J}) / [(1430 \text{ g})(0.452 \text{ J/g}^\circ\text{C})] \\ &= 11.6^\circ\text{C}\end{aligned}$$

A4. What is the temperature change for 500. g of each metal below, when 1,000. J of energy is added to each?

a) Aluminum

$$\begin{aligned}\Delta T &= q / mc \\ &= (1000 \text{ J}) / [(500 \text{ g})(0.891 \text{ J/g}^\circ\text{C})] \\ &= 2.24^\circ\text{C}\end{aligned}$$

b) Lead

$$\begin{aligned}\Delta T &= q / mc \\ &= (1000 \text{ J}) / [(500 \text{ g})(0.130 \text{ J/g}^\circ\text{C})] \\ &= 15.4^\circ\text{C}\end{aligned}$$

**Moral of this question: Lower heat capacity = Easier to heat up**

A5. How much heat is released to the surroundings when a cloud containing 100. kg of air cools from 34.0°C to 28.0°C?

$$\begin{aligned}q &= mc\Delta T \\ &= (100,000 \text{ g})(0.718 \text{ J/g}^\circ\text{C})(6^\circ\text{C}) \\ &= 430,800 \text{ J} \\ &= 431 \text{ kJ}\end{aligned}$$



A6. Find  $\Delta H$  for the dissolution of potassium chlorate ( $\text{KClO}_3$ ) if dissolving 1.2 g in 100 mL of water decreases the temperature from 25.00°C to 24.03°C.

$$\begin{aligned}q &= mc\Delta T \\ &= (100 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(-0.97^\circ\text{C}) \\ &= -405.8 \text{ J}\end{aligned}$$

$$\begin{aligned}n &= m/M \\ &= (1.2 \text{ g})/(122.55 \text{ g/mol}) \\ &= 0.00979 \text{ mol}\end{aligned}$$

$$\begin{aligned}\Delta H &= -q/n \\ &= (405.8 \text{ J}) / (0.00979 \text{ mol}) \\ &= 41442 \text{ J/mol} \\ &= \mathbf{+41.4 \text{ kJ/mol}}\end{aligned}$$

A7. Mixing 25.00 mL of 0.1000 M NaOH with 20.00 mL of 0.0800 M HCl causes the temperature of the mixture (initially 25.00°C) to increase to 25.49°C. What is the enthalpy of neutralization for this reaction?

$$\text{Moles of NaOH: } n=CV=(0.025 \text{ L})(0.1000 \text{ M}) = 0.0025 \text{ mol}$$

$$\text{Moles of HCl: } n=CV=(0.020 \text{ L})(0.0800 \text{ M}) = 0.0016 \text{ mol}$$

HCl is limiting and  $n=0.0016 \text{ mol}$

$$\begin{aligned}q &= mc\Delta T \\ &= (45 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(0.49^\circ\text{C}) \\ &= 92.26 \text{ J}\end{aligned}$$

$$\begin{aligned}\Delta H &= -q/n \\ &= -92.26 \text{ J} / 0.0016 \text{ mol} \\ &= -57660 \text{ J/mol} \\ &= \mathbf{-57.6 \text{ kJ/mol}}\end{aligned}$$

A8. Suppose you have 50.0 mL of 0.2 M  $\text{NaOH}_{(\text{aq})}$  and mix it with 50.0 mL of 0.2 M  $\text{HCl}_{(\text{aq})}$ . Both mixtures start at 25.00°C. The temperature change is measured and you calculate  $\Delta H_{\text{rxn}}$ .

Answer the following with “doubles”, “halves” or “stays the same”:

a) What happens to Q if we use 100 mL of each solution?

**DOUBLES.**

Total amount of both reactants double, so you have double the number of bonds to break/form

b) What happens to Q if we use 100 mL of  $\text{NaOH}_{(\text{aq})}$  but still only 50 mL of  $\text{HCl}_{(\text{aq})}$ ?

**STAYS THE SAME.**

HCl is now the limiting reactant, so there's no more bonds broken/formed than before.

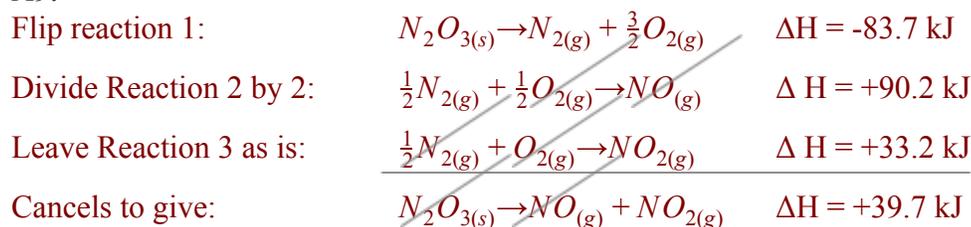
c) What happens to  $\Delta T$  if we use 100 mL of each solution?

STAYS THE SAME.

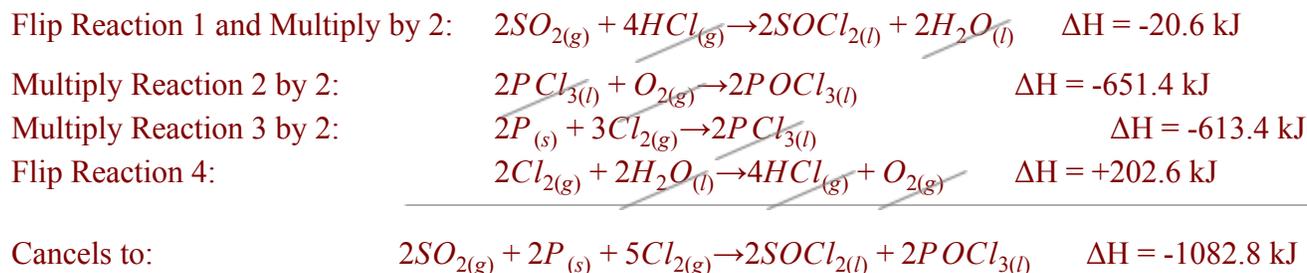
Even though Q doubles, so does the amount of water that's being heated. So  $\Delta T$  is the same.

- d) What happens to  $\Delta T$  if we use 150 mL of  $\text{NaOH}_{(aq)}$  but still only 50 mL of  $\text{HCl}_{(aq)}$ ?  
HALVES.  
We now have the same number of molecules reacting (since HCl is now limiting), so Q is the same, but the amount of water to be heated doubles, so the  $\Delta T$  is only half what it was.
- e) What happens to  $\Delta H_{\text{rxn}}$  if we use 100 mL of each solution?  
STAYS THE SAME.  
 $\Delta H$  is a constant for the reaction. It doesn't matter how much of each you use, if you measure the  $\Delta T$  and amounts correctly it should turn out the same.
- f) What happens to  $\Delta H_{\text{rxn}}$  if we use 150 mL of  $\text{NaOH}_{(aq)}$  but still only 50 mL of  $\text{HCl}_{(aq)}$ ?  
STAYS THE SAME  
 $\Delta H$  is constant. So none of the values matter when you're asked if it changes.
- g) What happens to  $\Delta H_{\text{rxn}}$  if we use 1850 mL of  $\text{NaOH}_{(aq)}$  but only 0.2 mL of  $\text{HCl}_{(aq)}$ ?  
STAYS THE SAME  
See? So little heat is produced here (since there's only 0.2 mL of HCl) but the moles of HCl is low as well so the calculated  $\Delta H$  should be the same as in (e) and (f).

A9.



A10.



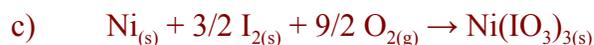
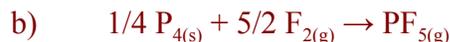
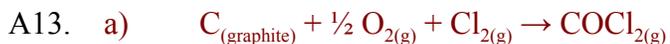
A11. Find the enthalpy change for this reaction:  $2NH_{3(g)} + \frac{7}{2}O_{2(g)} \rightarrow N_2O_{4(g)} + 3H_2O_{(g)}$  given the  $\Delta H_f^\circ$  values in the table below.

$$\begin{aligned}\Delta H &= (\text{Sum of } n\Delta H_f^\circ \text{'s for products}) - (\text{Sum of } n\Delta H_f^\circ \text{'s for reactants}) \\ &= [(1) \times (9 \text{ kJ/mol}) + (3) \times (-242 \text{ kJ/mol})] - [(2) \times (-46 \text{ kJ/mol}) + (7/2) \times (0 \text{ kJ/mol})] \\ &= \mathbf{-625 \text{ kJ}}\end{aligned}$$

A12. Find  $\Delta H_f^\circ$  for  $C_8H_{18}$  if the following reaction:  $C_8H_{18(l)} + \frac{25}{2}O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(l)}$  has  $\Delta H = -5430 \text{ kJ}$

In this equation,  $x$  is the unknown  $\Delta H_f^\circ$

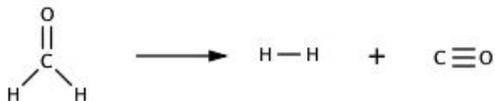
$$\begin{aligned}\Delta H &= (\text{Sum of } n\Delta H_f^\circ \text{'s for products}) - (\text{Sum of } n\Delta H_f^\circ \text{'s for reactants}) \\ -5430 \text{ kJ} &= [(8) \times (-394 \text{ kJ/mol}) + (9) \times (-286 \text{ kJ/mol})] - [(1) \times (x) + (25/2) \times (0 \text{ kJ/mol})] \\ -5430 &= -3152 + -2574 - x \\ x &= -3152 + -2574 + 5430 \\ x &= \mathbf{-296 \text{ kJ}}\end{aligned}$$



A14. Why is  $\Delta H_f$  for liquid water  $-286 \text{ kJ/mol}$ , while  $\Delta H_f$  for water vapour is  $-242 \text{ kJ/mol}$ ?

The difference is the energy released when the water vapour condenses down into its liquid form.

- A15. Use the bond enthalpies in the table below to estimate  $\Delta H$  for  $H_2CO_{(g)} \rightarrow H_{2(g)} + CO_{(g)}$ .  
 (Remember: You'll have to draw the Lewis Structures to know what types of bonds are there!)



Bonds of reactants:  $2 \times 413 + 1 \times 799 = 1625$  kJ required to break all of these bonds

Bonds of products:  $1 \times 436 + 1 \times 1072 = 1508$  kJ released when these bonds form

Therefore, there is a net of 117 kJ *required* and so  $\Delta H = +117$  kJ

- A16. Use bond enthalpies to estimate the  $\Delta H$  for  $C_2H_{2(g)} + 2F_{2(g)} \rightarrow C_2H_2F_{4(g)}$



Bonds of reactants:  $2 \times 413 + 1 \times 839 + 2 \times 155 = 1975$  kJ required to break all these bonds

Bonds of products:  $2 \times 413 + 1 \times 348 + 4 \times 485 = 3114$  kJ released when these bonds form

This is a net of 1139 kJ released, and so  $\Delta H = -1139$  kJ

- A17. Why are the  $\Delta H$  values from bond enthalpies called *estimates* whereas the values from Hess' Law and  $\Delta H_p$  are more accurate?

Bond enthalpies are *average* values over many different molecules in the gas phase

Not all C-H bonds require exactly 413 kJ to break, for example.

Also, if the molecules are in a different *state*, i.e. Liquid, different amounts of energy will be needed.

A18. What mass of water can be **melted** at 0°C ( $\Delta H_{\text{fus}} = +6.03 \text{ kJ/mol}$ ) with the energy released when 1 kg of ethanol ( $\Delta H_c = -2020 \text{ kJ/mol}$ ) is burned?

Calculate amount of energy released by the ethanol:

$$1 \text{ kg} = 1000 \text{ g}$$

$$n = m/M = 1000 \text{ g} / (46.068 \text{ g/mol}) = 21.71 \text{ mol}$$

$$Q = -n\Delta H_c = -(21.71 \text{ mol})(-2020 \text{ kJ/mol}) = 43847.8 \text{ kJ}$$

If Q for this process is positive, then Q for the other process is negative, so  $Q = -43847.8$

Amount of ice that can be melted by this:

$$n = -Q/\Delta H_{\text{fus}} = -(-43847.8 \text{ kJ}) / (6.03 \text{ kJ/mol}) = 7271.6 \text{ mol} \quad (\text{Negatives cancel})$$

$$m = nM = (7271.6 \text{ mol}) \times (18.02 \text{ g/mol}) = 131034 \text{ g} = \mathbf{131 \text{ kg}}$$

A19. What amount of energy is required to heat 200 g of ice at -160°C to steam at 125°C ?

$$c_{\text{ice}} = 2.01 \text{ J/g}^\circ\text{C}$$

$$\Delta H_{\text{fus}} = +6.03 \text{ kJ/mol}$$

$$c_{\text{water}} = 4.184 \text{ J/g}^\circ\text{C}$$

$$\Delta H_{\text{vap}} = +40.2 \text{ kJ/mol}$$

$$c_{\text{steam}} = 2.01 \text{ J/g}^\circ\text{C}$$

There are five processes here:

$$\text{ice at } -160 \rightarrow \text{ice at } 0 \quad q = mc\Delta T = (200\text{g})(2.01 \frac{\text{J}}{\text{g}}^\circ\text{C})(160^\circ\text{C}) = 64320\text{J} = 64.32\text{kJ}$$

$$\text{ice at } 0 \rightarrow \text{water at } 0 \quad q = n\Delta H_{\text{fus}} = (\frac{200\text{g}}{18.02 \frac{\text{g}}{\text{mol}}})(6.03 \frac{\text{kJ}}{\text{mol}}) = 66.93\text{kJ}$$

$$\text{water at } 0 \rightarrow \text{water at } 100 \quad q = mc\Delta T = (200\text{g})(4.184 \frac{\text{J}}{\text{g}}^\circ\text{C})(100^\circ\text{C}) = 83680\text{J} = 83.68\text{kJ}$$

$$\text{water at } 100 \rightarrow \text{steam at } 100 \quad q = n\Delta H_{\text{vap}} = (\frac{200\text{g}}{18.02 \frac{\text{g}}{\text{mol}}})(40.2\text{kJ}) = 446.17\text{kJ}$$

$$\text{steam at } 100 \rightarrow \text{steam at } 125 \quad q = mc\Delta T = (200\text{g})(2.01 \frac{\text{J}}{\text{g}}^\circ\text{C})(25^\circ\text{C}) = 10050\text{J} = 10.05\text{kJ}$$

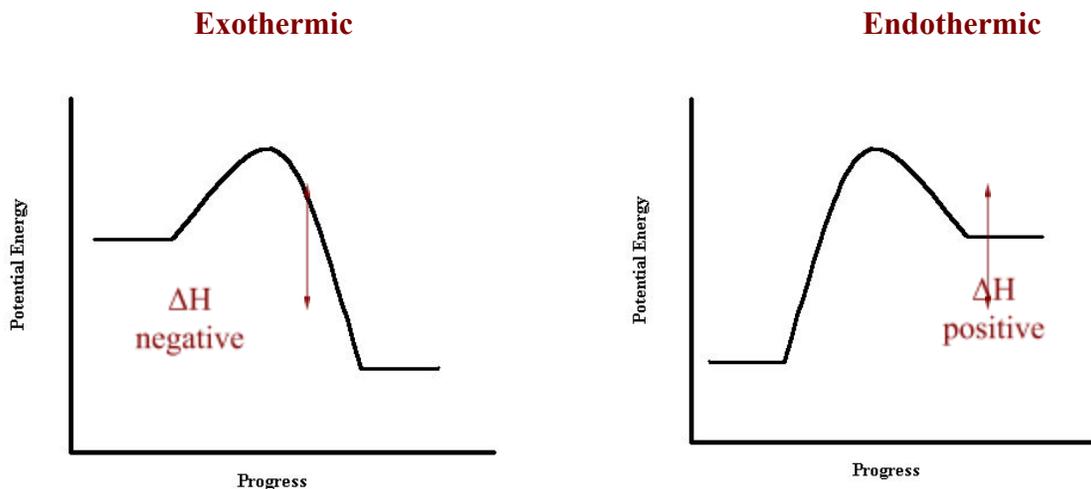
$$\text{Sum of these energy changes: } 64.32 + 66.93 + 83.68 + 446.17 + 10.05 = \mathbf{671.15 \text{ kJ}}$$

A20. What are the types of thermal energy present in:

Solids	Liquids	Gases
Vibrational	Vibrational Rotational	Vibrational Rotational Translational

A21. Draw a potential energy diagram for:

- The **exothermic** reaction  $C_2H_2 + 2 H_2 \rightarrow C_2H_6 + 312 \text{ kJ}$
- The **endothermic** reaction  $3 O_2 + 285.4 \text{ kJ} \rightarrow 2 O_3$
- Label the  $\Delta H$  on each of these curves.



A22. Given that  $NO_2$  has an enthalpy of formation of  $+33 \text{ kJ/mol}$  :

- What is  $\Delta H_f$  in  $\text{kJ/g}$  ?

$$\left(\frac{33 \text{ kJ}}{\text{mol}}\right)\left(\frac{1 \text{ mol}}{78.01 \text{ g}}\right) = 0.423 \frac{\text{kJ}}{\text{g}}$$

- What is  $\Delta H_f$  in  $\text{J/molecule}$ ?

$$\left(\frac{33 \text{ kJ}}{\text{mol}}\right)\left(\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molecules}}\right)\left(\frac{1000 \text{ J}}{1 \text{ kJ}}\right) = 5.482 \times 10^{-20} \frac{\text{J}}{\text{molecule}}$$

- How many molecules can be “formed” with 1 J of energy?

Set up a ratio:  $\frac{5.482 \times 10^{-20} \text{ J}}{1 \text{ molecule}} = \frac{1 \text{ J}}{x \text{ molecules}}$

Cross multiply:  $(5.482 \times 10^{-20})(x) = (1)(1)$

Solve for x:  $x = \frac{1}{5.482 \times 10^{-20}} = 1.824 \times 10^{19} \text{ molecules}$

A23. For each of these reactions, is  $\Delta S$  positive or negative?

- a)  $2NO_{2(g)} \rightarrow 2NO_{(g)} + O_{2(g)}$  positive, for two reasons: 1. More molecules created  
2. pure substance  $\rightarrow$  mixture.
- b)  $H_2O_{(g)} \rightarrow H_2O_{(l)}$  negative. Gases have more entropy (disorder) than liquids
- c)  $H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$  negative. Pure substances have less disorder than mixtures
- d)  $Hg_{(l)} (20C) \rightarrow Hg_{(l)} (25C)$  positive. Higher temps  $\rightarrow$  More disorder
- e)  $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$  negative. Solids have very low entropy.

A24. Assuming  $\Delta H$  and  $\Delta S$  do not vary with temperature, at what temperature will this reaction become spontaneous?  $C_{(s)} + H_2O_{(g)} \rightarrow CO_{(g)} + H_{2(g)}$ , with  $\Delta H = 131.3$  kJ and  $\Delta S = 133.6$  J/K

$\Delta G = \Delta H - T\Delta S$ , and we want  $\Delta G < 0$  for this to be spontaneous. Therefore,  $\Delta H - T\Delta S < 0$ .

$$(131300 \text{ J}) - (T)(133.6 \text{ J/K}) < 0$$
$$131300 \text{ J} < (T)(133.6 \text{ J/K})$$
$$\frac{131300 \text{ J}}{133.6 \frac{\text{J}}{\text{K}}} < T$$
$$\mathbf{983 \text{ K} < T}$$

(You could convert this to  $^{\circ}\text{C}$  if you want: it is  $709^{\circ}\text{C}$ )

A25. Mercury freezes at  $-38.9^{\circ}\text{C}$ , and its enthalpy of fusion is  $2.33$  kJ/mol. Calculate the value of  $\Delta S$  for the freezing of  $1.00$  mol of  $Hg_{(l)}$  at  $-38.9^{\circ}\text{C}$ .

The freezing point is the point when the  $Hg_{(l)} \rightarrow Hg_{(s)}$  process *becomes* spontaneous. Therefore, it is where  $\Delta G$  crosses from positive to negative ... which means  $\Delta G = 0$  at this exact temperature.

$$0 = \Delta H - T\Delta S$$
$$0 = (2330 \text{ J}) - (234.25 \text{ K})(\Delta S)$$
$$(234.25 \text{ K})(\Delta S) = 2330 \text{ J}$$
$$\Delta S = (2330 \text{ J}) / (234.25 \text{ K})$$
$$\Delta S = \mathbf{+9.95 \text{ J/K}}$$