

Slide 1

**Enthalpy**

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Calorimetry of Chemistry

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Slide 2

**Reaction Energies**

In our earlier discussions of calorimetry, we used **physical** sources of heat (hot metal slug). It is also possible to use **chemical** sources of heat (like hot packs and cold packs).

The energy change associated with a chemical reaction is called the **enthalpy of reaction** and abbreviated  $\Delta H$ .

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Slide 3

**Enthalpy of Reactions**

There are actually a number of different types of enthalpies because enthalpy depends on conditions. THEY ARE ALL JUST SPECIFIC TYPES OF A GENERAL CONCEPT CALLED "ENTHALPY".

$$\Delta H = H_{\text{final}} - H_{\text{initial}}$$

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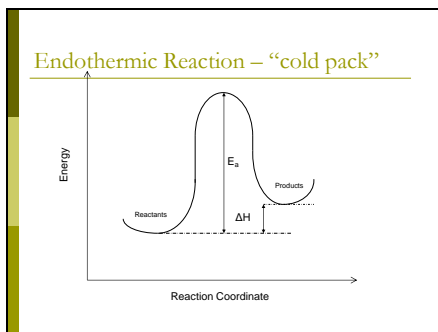
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Slide 8

Where does the Energy go?

In the case of a chemical reaction, you need to keep the different types of energy separate in your mind:

- Bond energy – energy INSIDE the molecules
- Thermal energy (heat) – kinetic energy of the molecules
- Energy of the “bath” – kinetic energy of solvent or other molecules in the system

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Energy changes

$\Delta H$  represents the change in INTERNAL MOLECULAR ENERGY.

$\Delta H = H_{\text{final}} - H_{\text{initial}}$

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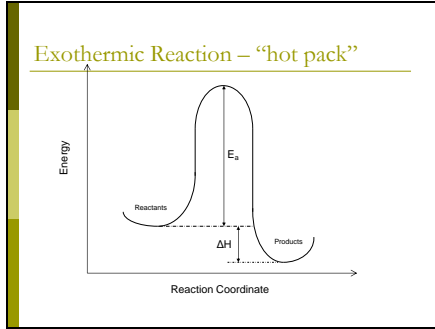
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Slide 11

Exothermic energy changes

$\Delta H = H_{\text{final}} - H_{\text{initial}} < 0$

$H_{\text{initial}} > H_{\text{final}}$

This energy is internal to the molecule. The excess gets absorbed by the rest of the system as heat causing the molecules to move faster (more kinetic energy) and the temperature to increase.

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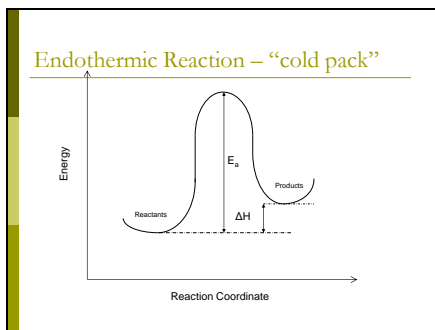
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Slide 13

**Endothermic energy changes**

$$\Delta H = H_{\text{final}} - H_{\text{initial}} > 0$$
$$H_{\text{initial}} < H_{\text{final}}$$

This energy is internal to the molecule and must come from somewhere.  
The additional energy required by the system gets absorbed from the rest of the system as heat causing the molecules to move slower (less kinetic energy) and the temperature to decrease.

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**Clicker Question**

Consider the following reaction:  
 $2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \rightarrow 2 \text{H}_2\text{O} (\text{g})$

If  $\Delta H_{\text{rxn}} < 0$ , it means:

- A) The products have less energy than the reactants – you could make a hot pack.
- B) The reactants have less energy than the products – you could make a cold pack.
- C) The products have less energy than the reactants – you could make a cold pack.
- D) The reactants have less energy than the products – you could make a hot pack.

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Slide 15

**The hard part is getting over the hump.**

Energy

Reactants

Products

$E_a$

$\Delta H$

Reaction Coordinate

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### $E_a = \text{Activation Energy}$

The tale of a reaction is not limited strictly to the identity and energetics of the products and reactants, there is a path (reaction coordinate) that must get followed.

The "hump" represents a hurdle that must be overcome to go from reactants to products.

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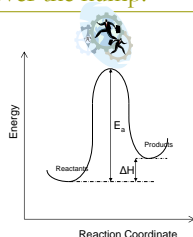
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### How do you get over the hump?

If you are at the top, it is easy to fall down into the valley (on either side), but how do you get to the top?



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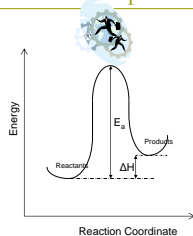
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### How do you get over the hump?

The molecules acquire or lose energy the same way: by colliding with each other!

The energy comes from the "bath", the rest of the system.



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Types of  $\Delta H$

- $\Delta H$  – generic version
- $\Delta H_{\text{rxn}}$  – generic version
- $\Delta H^\circ$  – enthalpy change under Standard Temperature and Pressure (298 K, 1 atm)
- $\Delta H_f$  – enthalpy of formation, refers to a specific reaction type
- $\Delta H_{\text{comb}}$  – enthalpy change of combustion
- $\Delta H_f^\circ$  – enthalpy of formation at STP

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Slide 20

Start here

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Slide 21

$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$

The enthalpy change involved in this reaction depends, to some extent, on conditions.

At STP,  $\Delta H^\circ$  will be known if it's ever been measured:

$\Delta H^\circ = -483.66 \text{ kJ}$

The enthalpy change must include some accounting for the amount (moles of the substance)

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$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$

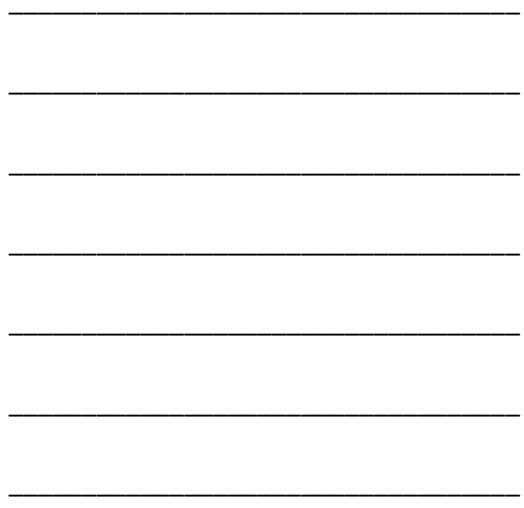
This reaction is a very special type of reaction.

This is a **reaction of formation**.

A formation reaction creates a molecule from the most common elemental form of its constituent atoms

$\Delta H_f^\circ = -241.83 \text{ kJ/mol}$

We form 2 moles of  $\text{H}_2\text{O}$  in this reaction, so:  
 $\Delta H_{rxn} = 2 \text{ mol} \cdot (-241.83 \text{ kJ/mol}) = -483.66 \text{ kJ}$



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**Determining the  $\Delta H$**

Where did I get that number ( $\Delta H_f^\circ = -241.83 \text{ kJ/mol}$ )?

From the table in the appendix of your book which lists enthalpies of formation for a number of different molecules!



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Substance	$\Delta H_f^\circ$ (kJ/mol)	$\Delta G_f^\circ$ (kJ/mol)	$S^\circ$ (J/mol · K)	Substance	$\Delta H_f^\circ$ (kJ/mol)	$\Delta G_f^\circ$ (kJ/mol)	$S^\circ$
$\text{NH}_3(g)$	-46.1	-16.4	192.8	<b>Potassium</b>			
$\text{NH}_2\text{OH}(l)$	-50.5	-18.9	131.1	K(s)	0	0	
$\text{NH}_4\text{NO}_3(s)$	-381.9	-204.8	259.8	K(g)	89.0	86.5	
$\text{HNO}_3(l)$	-134.4	-73.5	266.9	$\text{K}^+(aq)$	-252.14	-203.3	
$\text{HNO}_3(aq)$	-207	-110.9	146	$\text{KNO}_3$	-393.8	-362.7	
$\text{NO}_2(g)$	33.2	51.3	240.1	$\text{KNO}_2$	-113.0	-101.9	
$\text{NO}(g)$	91.3	87.6	210.8	$\text{KCN}(s)$	-436.5	-406.5	
$\text{NO}_2^-(aq)$	-206.85	-136.2	186.78	$\text{K}_2\text{S}(s)$	-395.7	-286.3	
$\text{NO}(g)$	82.2	85.4	273.7	$\text{K}_2\text{SO}_4$	-432.8	-303.1	
$\text{NO}_2(g)$	51.7	66.1	261.7	$\text{KHF}_2$	-367.3	-337.8	
$\text{N}_2\text{H}_4(l)$	50.4	149.3	121.2	$\text{K}_2\text{O}$	-425.9	-324.9	
$\text{N}_2\text{H}_4(g)$	95.4	194.4	218.5	$\text{KNO}_3(l)$	-494.6	-391.9	
$\text{N}_2\text{H}_2\text{O}$	81.6	180.7	210.6	$\text{KNO}_2$	-424.6	-319.4	
$\text{N}_2\text{H}_4\text{O}$	-19.3	97.5	209.2	$\text{KOH}(aq)$	-482.4	-440.5	
$\text{N}_2\text{H}_4\text{O}$	11.1	99.9	206.6	$\text{K}_2\text{O}_2$	-284.9	-230.4	
$\text{N}_2\text{H}_4\text{O}$	-63.1	113.9	176.2	$\text{K}_2\text{CO}_3$	-1151.8	-1063.5	
$\text{N}_2\text{H}_4\text{O}$	15.9	117.1	165.7	$\text{K}_2\text{O}(l)$	-361.5	-322.1	
<b>Oxygen</b>				$\text{K}_2\text{S}_2\text{O}_8$	-494.1	-423.1	
$\text{O}_2(g)$	0	0	205.2	$\text{K}_2\text{SO}_4$	-1437.9	-1321.4	
$\text{O}_3(g)$	142.7	163.2	238.9	<b>Rubidium</b>			
$\text{O}_2^-(aq)$	-230.02	-157.3	-10.99	$\text{Rb}(s)$	0	0	
$\text{H}_2\text{O}(l)$	-285.8	-173.1	70.0	$\text{Rb}(g)$	81.9	51.1	
$\text{H}_2\text{O}(g)$	-241.8	-228.4	188.8	$\text{Rb}^+(aq)$	-275.12	-201.1	
$\text{H}_2\text{O}_2(l)$	-187.8	-120.4	109.6	$\text{RbNO}_3$	-394.6	-311.8	
$\text{H}_2\text{O}_2(g)$	-136.3	-105.6	232.7	$\text{Rb}_2\text{O}$	-456.4	-402.8	
				$\text{Rb}_2\text{CO}_3$	-932.4	-791.8	





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Determining the  $\Delta H$   
Suppose the molecule you care about isn't in that table?  
Find a different table! ☺  
Half-kidding! You could also determine the value from the bond energies involved.

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Enthalpy is a "State Function"  
What's a "state function"?  
A "state function" is a value that is a function only of the initial and final states of the system, not the path you take to get there!

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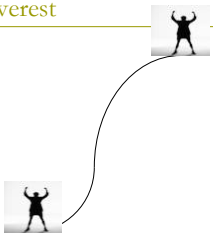
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Climbing Mt. Everest  
Suppose you start at Himalayan Base Camp #1, climb to the summit of Everest over the course of 3 weeks, then return to Himalayan Base Camp #1.



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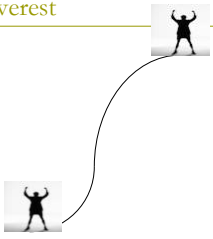
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### Climbing Mt. Everest

Back at base camp, I figure out my altitude change. What is it?

ZERO - I'm back where I started



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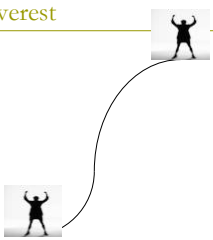
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### Climbing Mt. Everest

I did a lot of work along the way, but all that matters is I'm back where I started. The net change in altitude is NADA, ZERO, ZILCH!



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### Enthalpy as a State Function

Enthalpy is like that. It doesn't care how you got where you are going, it simply looks at the difference from where you started.

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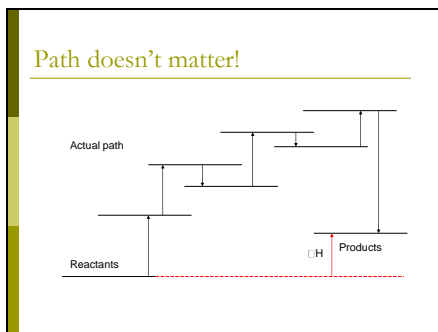
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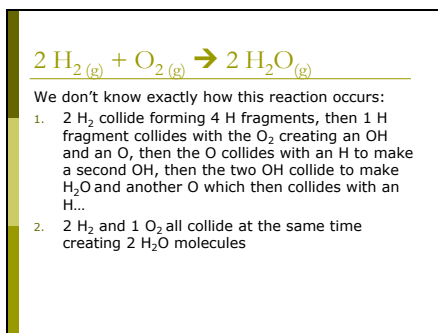
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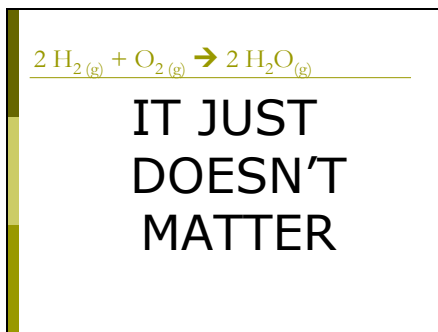
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Slide 33



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$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

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You can pick whatever path you want that makes it easy to calculate.

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$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

H-H            H H            H H  
H-H            \ /            \ /  
O=O            O                O

→

One common path: add up all the broken bonds and the made bonds and see what the net difference is!

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Slide 36

$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

H-H            H H            H H  
H-H            \ /            \ /  
O=O            O                O

→

I need to break 2 H-H bonds and 1 O=O bond  
I need to make 4 O-H bonds

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$$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$$

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This compares well to the known value given earlier:  $\Delta H^\circ = -483.66 \text{ kJ}$

Calculated from Bond Dissociation Energies,  $\Delta H_{\text{rxn}} = -482 \text{ kJ}$

Why the slight difference?  
Bond energies are not identical, depending on who their neighbors are: an O-H next to another O-H isn't exactly the same as an O-H next to an O-N, for example.

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Ways to determine  $\Delta H$

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1. Find  $\Delta H^\circ$  in a table
2. Find  $\Delta H_f^\circ$  in a table
3. Calculate from Bond Energies

And...

4. Calculate from  $\Delta H_f^\circ$
5. Calculate from other  $\Delta H$  that you already know (Hess's Law)

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Slide 45

$$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$$

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This is a **reaction of formation** so we simply found the value for it in the table.

$\Delta H_f^\circ = -241.83 \text{ kJ/mol}$

But even if the reaction of interest isn't itself a formation reaction, we can still use the enthalpies of formation to get the  $\Delta H_{\text{rxn}}$

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Slide 46

$$\text{CH}_4(g) + 4 \text{Cl}_2(g) \rightarrow \text{CCl}_4(l) + 4 \text{HCl}(g)$$

This is not, itself, a formation reaction.

BUT remember  $\Delta H$  is a STATE FUNCTION

What does that mean?

$\Delta H$  doesn't depend on the path, just the start and the end.

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Slide 47

$$\text{CH}_4(g) + 4 \text{Cl}_2(g) \rightarrow \text{CCl}_4(l) + 4 \text{HCl}(g)$$

All the elements known to mankind!

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Slide 48

$$\text{CH}_4(g) + 4 \text{Cl}_2(g) \rightarrow \text{CCl}_4(l) + 4 \text{HCl}(g)$$

How does that help?

You can take the long road. Don't do the reaction as written, take a convenient path that you know the value of each step.

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$$\text{CH}_4(\text{g}) + 4 \text{Cl}_2(\text{g}) \rightarrow \text{CCl}_4(\text{l}) + 4 \text{HCl}(\text{g})$$
$$\Delta H_{\text{rxn}}^0 = \sum \Delta H_f, \text{prod} - \sum \Delta H_f, \text{react}$$
$$\Delta H_{\text{rxn}}^0 = [\Delta H_f(\text{CCl}_4) + 4 * \Delta H_f(\text{HCl})] - [\Delta H_f(\text{CH}_4) + 4 * \Delta H_f(\text{Cl}_2)]$$
$$\Delta H_{\text{rxn}}^0 = [-139.3 \text{ kJ/mol} + 4 * (-92.3 \text{ kJ/mol})] - [-74.8 \text{ kJ/mol} + 4 * 0 \text{ kJ/mol}]$$
$$\Delta H_{\text{rxn}}^0 = -433.7 \text{ kJ}$$

$\Delta \Delta \text{anything} = \text{final anything} - \text{initial anything}$

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Slide 53

$$\Delta H_f(\text{HCl})$$
$$\frac{1}{2} \text{H}_2(\text{g}) + \frac{1}{2} \text{Cl}_2(\text{g}) = \text{HCl}(\text{g})$$

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Slide 54

$$\text{CH}_4(\text{g}) + 4 \text{Cl}_2(\text{g}) \rightarrow \text{CCl}_4(\text{l}) + 4 \text{HCl}(\text{g})$$
$$\Delta H_{\text{rxn}}^0 = \sum \Delta H_f, \text{prod} - \sum \Delta H_f, \text{react}$$
$$\Delta H_{\text{rxn}}^0 = [\Delta H_f(\text{CCl}_4) + 4 * \Delta H_f(\text{HCl})] - [\Delta H_f(\text{CH}_4) + 4 * \Delta H_f(\text{Cl}_2)]$$
$$\Delta H_{\text{rxn}}^0 = [-139.3 \text{ kJ/mol} + 4 * (-92.3 \text{ kJ/mol})] - [-74.8 \text{ kJ/mol} + 4 * 0 \text{ kJ/mol}]$$
$$\Delta H_{\text{rxn}}^0 = -433.7 \text{ kJ}$$

-433.7 kJ/mol CH<sub>4</sub>  
-433.7 kJ/4 mol Cl<sub>2</sub>

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$\Delta H_f(\text{CCl}_4) = -139.3 \text{ kJ/mol}$

If I make 1 mol  $\text{CCl}_4$ , I get/need 139.3 kJ?

A. Get  
B. Need

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Substance	$\Delta H_f^\circ$ (kJ/mol)	$\Delta G_f^\circ$ (kJ/mol)	$S^\circ$ (J/mol · K)	Substance	$\Delta H_f^\circ$ (kJ/mol)	$\Delta G_f^\circ$ (kJ/mol)	$S^\circ$
$\text{Ar(g)}$	-161.9	110.9		<b>Potassium</b>			
$\text{Ni}_2\text{S}_3(\text{s})$	-308.4	-183.9	151.1	K(s)	0	0	
$\text{Ni}_2\text{S}_3(\text{aq})$	-139.9	-196.6	239.8	K(g)	89.0	88.3	
$\text{HBr(g)}$	-118.7	-73.5	266.9	$\text{K}^+(\text{aq})$	-351.34	-351.3	
$\text{HBr(aq)}$	-307	-118.9	186	KBr(s)	-291.8	-186.7	
$\text{NO}_2(\text{g})$	95.3	87.6	210.8	KNO(s)	-119.0	-104.9	
$\text{NO}_2(\text{aq})$	33.2	25.3	281.1	KNO <sub>2</sub> (s)	-436.3	-406.3	
$\text{NO(g)}$	-206.92	-132.2	188.70	KNO <sub>3</sub> (s)	-495.7	-296.3	
$\text{NO(aq)}$	82.2	82.4	275.7	K <sub>2</sub> SO <sub>4</sub> (s)	-432.8	-301.1	
$\text{NO}_2(\text{aq})$	51.7	96.1	281.7	K <sub>2</sub> (g)	-861.3	-151.4	
$\text{N}_2\text{H}_4(\text{g})$	50.8	149.3	131.2	K <sub>2</sub> (l)	-327.9	-324.9	
$\text{N}_2\text{H}_4(\text{aq})$	61.8	276.4	238.3	KNO <sub>2</sub> (l)	-494.8	-394.9	
$\text{N}_2\text{H}_6(\text{g})$	61.8	185.7	238.0	KNO <sub>2</sub> (s)	-426.4	-179.4	
$\text{N}_2\text{H}_6(\text{aq})$	-119.3	97.5	206.7	KNO <sub>3</sub> (aq)	-462.4	-446.3	
$\text{N}_2\text{H}_8(\text{g})$	11.1	99.8	306.4	KNO <sub>3</sub> (s)	-246.9	-234.4	
$\text{N}_2\text{H}_8(\text{aq})$	-43.1	133.9	178.2	K <sub>2</sub> CO <sub>3</sub> (s)	-1151.8	-1063.3	
$\text{N}_2\text{H}_8(\text{aq})$	15.3	112.1	315.7	K <sub>2</sub> (l)	-961.3	-122.1	
<b>Oxygen</b>				$\text{K}_2\text{SO}_4$ (s)	-494.1	-421.1	
O(g)	242.2	231.7	161.1	$\text{K}_2\text{SO}_4$ (l)	-1457.8	-1321.4	
O(l)	0	6	205.2				
O(g)	142.7	163.2	238.9	<b>Argon</b>			
$\text{OH}^-(\text{aq})$	-230.2	-157.3	-109.0	Ar(g)	0	0	
$\text{H}_2\text{O}(l)$	-285.8	-237.1	70.0	Br(g)	86.4	58.1	
$\text{H}_2\text{O}(g)$	-241.8	-228.6	188.8	Br <sub>2</sub> (l)	-291.22	-263.1	
$\text{H}_2\text{O}_2(l)$	-187.8	-120.4	109.6	Br <sub>2</sub> (g)	-30.9	-31.8	
$\text{H}_2\text{O}_2(g)$	-136.3	-105.6	232.7	BrCl(s)	-435.4	-421.8	
				BrCl <sub>3</sub> (s)	-392.4	-291.0	

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Slide 57

**Clicker Question**

$\text{CH}_4(\text{g}) + 4 \text{Cl}_2(\text{g}) \rightarrow \text{CCl}_4(\text{l}) + 4 \text{HCl}(\text{g})$

$\Delta H_{\text{rxn}}^0 = \frac{-433.7 \text{ kJ}}{4 \text{ mol Cl}_2}$

$\Delta C_{\text{H}_2\text{O}} = 4.18 \text{ J/g}^\circ\text{C}$

If I want to heat 1000 g of water from 25 °C to boiling. How much chlorine would I need to react (assuming I have infinite  $\text{CH}_4$ )?

A. 51.26 g  
B. 0.723 g  
C. 205.04 g  
D. Don't tell me it's over ☹

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Slide 58

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$$H^{\circ}_{\text{rxn}} = \frac{-433.7 \text{ kJ}}{4 \text{ mol Cl}_2}$$
$$Q = mc\Delta T = (1000 \text{ g}) (4.18 \text{ J/gK}) (373-298)$$

Q=313,500 J energy needed

$$Q_{\text{water}} + Q_{\text{rxn}} = 0$$
$$Q_{\text{water}} = -Q_{\text{rxn}} = -\Delta H_{\text{rxn}}$$
$$Q = 313,500 \text{ J} \frac{4 \text{ mol Cl}_2 \text{ react } 70.9 \text{ g Cl}_2}{433,700 \text{ J/mol Cl}_2} = 205 \text{ g}$$

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Slide 59

### Ways to determine $\Delta H$

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1. Find  $\Delta H^{\circ}$  in a table
2. Find  $\Delta H_f^{\circ}$  in a table
3. Calculate from Bond Energies
4. Calculate from  $\Delta H_f^{\circ}$
5. Calculate from other  $\Delta H$  that you already know (Hess's Law)

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Slide 60

### Hess's Law

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We already basically used Hess's Law when we added together the heats of formation

Hess's Law is simply the tools that go with enthalpy being a state function.

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Slide 61

### Hess's Law

1. If you add two reactions together,  $\Delta H$  adds together.
2. If you subtract two reactions,  $\Delta H$  gets subtracted.
3. If you reverse a reaction,  $\Delta H$  changes sign.
4. If you multiply or divide a reaction,  $\Delta H$  gets multiplied or divided.

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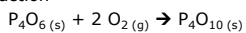
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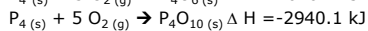
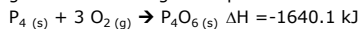
Slide 62

### Sample Hess's Law Problem

Calculate the enthalpy change for the reaction



given the following enthalpies of reaction:



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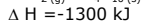
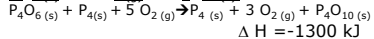
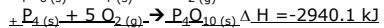
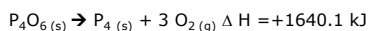
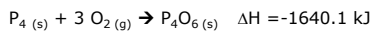
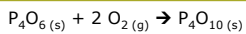
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Slide 63

### Sample Hess's Law Problem



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Slide 64

Enthalpy & Calorimetry

You can combine the enthalpies of reaction with the calorimetry we discussed earlier, using the reactions to generate the heat.

This is nothing new, just a combination of the two concepts we already discussed and a few things we knew from before.

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Slide 65

$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$

A bomb calorimeter is a water calorimeter with a small chamber inside in which combustion reactions can be executed. If I put 1 mole of hydrogen and 1 mole of oxygen in a bomb calorimeter containing 1 L of water at 25 °C, what will the temperature of the water be after ignition? The empty (no water) calorimeter has a specific heat capacity of 145.1 J/°C.

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Slide 66

Q=mc (delta T)

$S_h = 145.1 \text{ J/}^\circ\text{C}$   
 $c = 4.18 \text{ J/g }^\circ\text{C}$

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Slide 67

$$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$$

A bomb calorimeter is a water calorimeter with a small chamber inside in which combustion reactions can be executed. If I put 1 mole of hydrogen and 1 mole of oxygen in a bomb calorimeter containing 1 L of water at 25 °C, what will the temperature of the water be after ignition? The empty (no water) calorimeter has a specific heat capacity of 145.1 J/°C.

What the heck does that mean?

UNITS! UNITS! UNITS

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Slide 68

$$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$$

The empty (no water) calorimeter has a specific heat capacity of 145.1 J/°C.

$$\frac{145.1 \text{ J}}{^\circ\text{C}}$$

If you have an object (like a calorimeter) that absorbs heat and is always the same, you can calibrate it. It takes 145.1 Joules to make the calorimeter itself go up 1 °C.

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Slide 69

$$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$$

A bomb calorimeter is a water calorimeter with a small chamber inside in which combustion reactions can be executed. If I put 1 mole of hydrogen and 1 mole of oxygen in a bomb calorimeter containing 1 L of water at 25 °C, what will the temperature of the water be after ignition? The empty (no water) calorimeter has a specific heat capacity of 145.1 J/°C.

$q_{\text{H}_2\text{O}} = m c \Delta T$   
 $q_{\text{calorimeter}} = m c \Delta T$   
 $Q = S_h \Delta T$   $S_h$  is a combination of m's and c's.

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Slide 70

$$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$$

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$q_{\text{rxn}} = -(q_{\text{bomb}} + q_{\text{water}})$

$$1 \text{ mol O}_2 \xrightarrow{2 \text{ mol H}_2\text{O}} \frac{1 \text{ mol O}_2}{1 \text{ mol O}_2} = 2 \text{ mol H}_2\text{O}$$
$$1 \text{ mol H}_2 \xrightarrow{2 \text{ mol H}_2\text{O}} \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} = 1 \text{ mol H}_2\text{O}$$

Hydrogen is the limiting reagent.

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Slide 71

$$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$$

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How much energy do I get out of the reaction? Depends on how much stuff reacts!

$$\Delta H_{\text{rxn}}^{\circ} = 2 \Delta H_f^{\circ}(\text{H}_2\text{O}) - (0+0) =$$
$$= 2(-241.83 \text{ kJ/mol}) = -483.66 \text{ kJ}$$

So every 2 moles of H<sub>2</sub> that reacts gives you -483.66 kJ

$$1 \text{ mol H}_2 \xrightarrow{-483.66 \text{ kJ}} \frac{-483.66 \text{ kJ}}{2 \text{ mol H}_2} = -241.83 \text{ kJ released}$$
$$1 \text{ L H}_2\text{O} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ g H}_2\text{O}}{1 \text{ mL H}_2\text{O}} = 1000 \text{ g H}_2\text{O}$$

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Slide 72

$$2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g)$$

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$q_{\text{rxn}} = -(q_{\text{bomb}} + q_{\text{water}})$

$$-241.83 \text{ kJ} = -(S_{\text{bomb}} \Delta T + m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}} \Delta T)$$
$$-241.83 \text{ kJ} = -(145.1 \text{ J/}^\circ\text{C} \Delta T + 1000 \text{ g} \cdot 4.18 \text{ J/g}^\circ\text{C} \Delta T)$$
$$-241.83 \text{ kJ} = -(4325.1 \text{ J/}^\circ\text{C} \Delta T)$$
$$-241.83 \times 10^3 \text{ J} = -4325.1 \text{ J/}^\circ\text{C} \Delta T$$
$$55.9 = \Delta T = T_f - T_i = T_f - 25^\circ\text{C}$$
$$T_f = 80.9^\circ\text{C}$$

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Slide 79



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