ENTHALPY

Thermochemistry

Thermochemistry is the study of the energy changes in a chemical reaction.

Basically, the heat given off or absorbed when a chemical reaction happens.

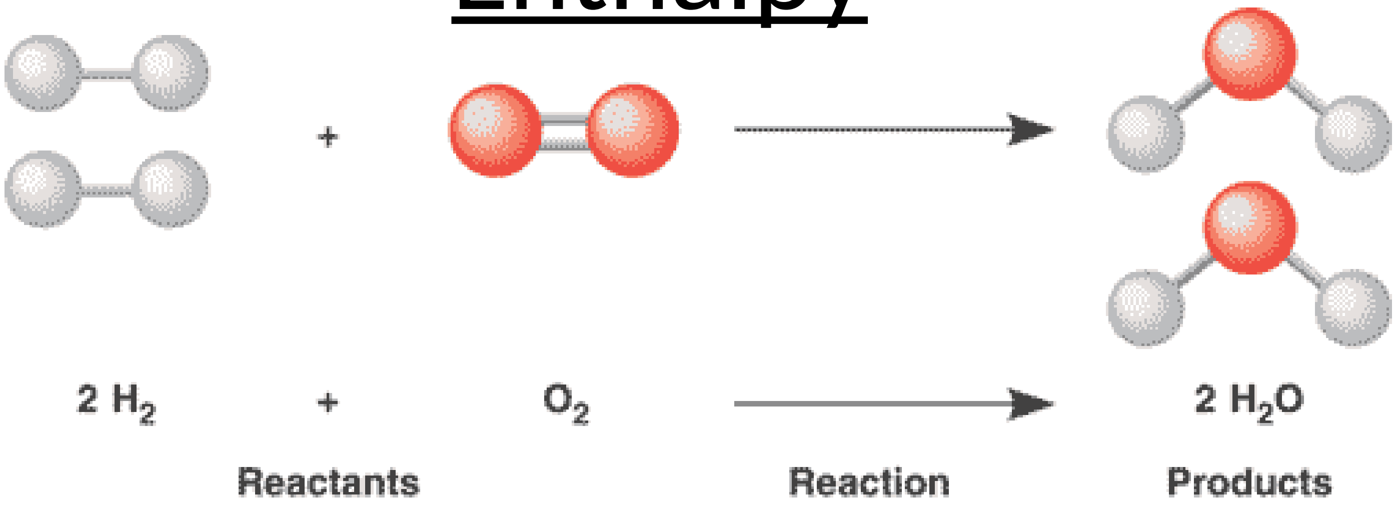
Enthalpy

Enthalpy (given the symbol ‘H’) is the measure of kinetic and potential energy in a chemical species.

This energy exists in the bonds of the molecules.

And the energy given off or absorbed by a chemical reaction is a result of the breaking and forming of these bonds.

Enthalpy



In this reaction the bonds between the H atoms and the bonds between the O atoms have to break apart.

They then have to make new bonds to make the water molecules.

This happens in all chemical reactions.

All of this requires energy

Enthalpy

The **change in Enthalpy (ΔH)** is the difference in energy between products and reactants in a chemical reaction.

It is measured in kilojoules (kJ) or in kilojoules per mole (kJ/mol)

Copy this down

$\Delta H = H_{products} - H_{Reactants}$

Enthalpy

The **change in Enthalpy (ΔH)** is the difference in energy between products and reactants in a chemical reaction.

It is measured in kilojoules (kJ) or in kilojoules per mole (kJ/mol)

Copy this down

$$\Delta H = H_{products} - H_{Reactants}$$

Little side note here:
Heat is a form of energy. That is what we'll be talking about in this topic

Temperature is just a measure of the average kinetic energy. If particles are heated up, they move more and thus have higher kinetic energy. Temperature just measures this.

Enthalpy

Because we’re talking about the **change** in enthalpy (ΔH), were talking about the difference in energy between the products and the reactants.

We’re talking about the difference in the energy in the bonds between the products and reactants.

Enthalpy

That is why using the formula $\Delta H = \Delta H_{Products} - \Delta H_{Reactants}$

When the products have more energy than the reactants,

ΔH = +ve it's called Endothermic

And, when the reactants have more energy than the products

ΔH = -ve it's called Exothermic

Enthalpy

ΔH	**Type of Reaction**	**Absorbs/ Releases**
+ΔH	Endothermic	Absorbs energy from environment
-ΔH	Exothermic	Releases energy to environment

Copy this down

Enthalpy

$HCl + NaOH \longrightarrow NaCl + H_2O$

The reaction of an acid and a base (as above) is a vigorous reaction that gives off heat (energy)

Is the reaction Exothermic or Endothermic?

Releases energy to environment

Absorbs energy from environment

Enthalpy

$HCl + NaOH \longrightarrow NaCl + H_2O$

The reaction of an acid and a base (as above) is a vigorous reaction that gives off heat (energy)

Is the reaction Exothermic or Endothermic?

Releases energy to environment

Absorbs energy from environment

Would these reactions be exothermic or endothermic?

The combustion of butane.

A hand warmer pouch.

Melting ice cubes.

Would these reactions be exothermic or endothermic?

The combustion of butane. Exothermic **ΔH= -ve**

A hand warmer pouch. Exothermic **ΔH = -ve**

Melting ice cubes. Endothermic **ΔH = +ve**

Enthalpy and Spontaneity

Endothermic reactions absorb energy from the environment. They usually need some form of energy to be able to break/form the chemical bonds to make the reaction happen.

Like baking a cake. It requires heat to make the chemical reactions take place to cause the cake to rise.

Enthalpy and Spontaneity

Exothermic reactions on the other hand, release energy to the environment.

They don't require energy to make the reaction happen. It's much easier for the reaction to go to completion so it pretty much will.

You don't have to keep feeding energy to thermite for it to burn, the chemical reaction will just happen.

Enthalpy and Spontaneity

Therefore, exothermic reactions are more spontaneous than endothermic reactions.

This means they don’t really need anything to start happening.

Another thing that affects spontaneity is something called entropy, but we’ll get to that next term.

Important thing though is, even if something’s more spontaneous, it doesn’t mean the reaction will happen faster. That’ll come up again later….

Bond Enthalpies

As stated before, the energy of a reaction has to do with bonds breaking in molecules and reforming to make other molecules.

Overtime, through a lot of experiments, chemists have calculated the average energy of various bonds between different elements.

Note about these bond enthalpies, they are averages, they are calculated for gases, and they don't take into account other forces inside the molecules (e.g. polarities, etc.). But for us they'll do nicely.

Bond Enthalpies

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

Multiple Bonds

C=C	614	N=N	418	O_2	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				

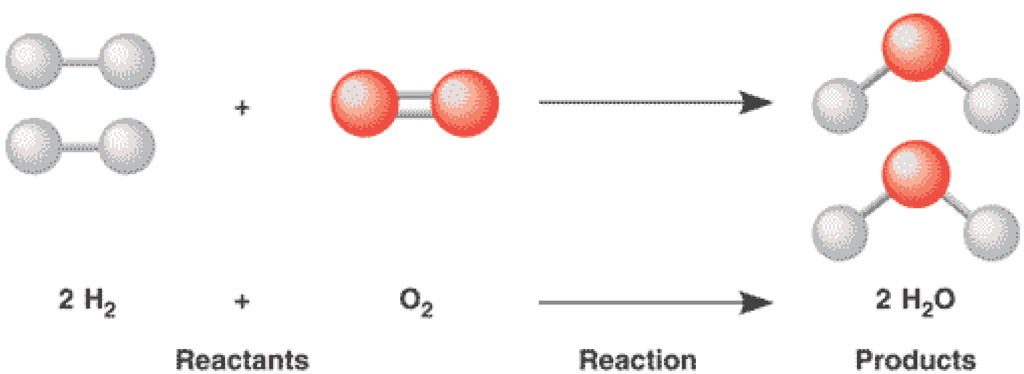
Bond Enthalpies

These are the energies that are inside the bonds between these atoms. This is the energy required to break these bonds **and** the energy released when they're formed again.

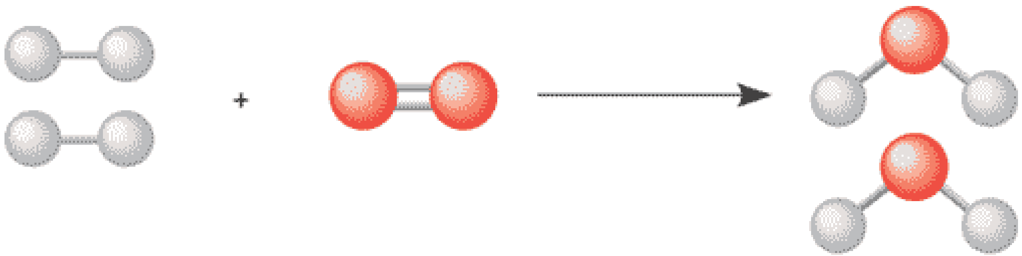
We can use these to calculate the ΔH of a reaction.

Let's do an example,

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$



Bond Enthalpies



$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

Here, 2 molecules of hydrogen gas and one molecule of oxygen gas break apart and reform bonds to make 2 water molecules.

We look at the table to find how much energy is required to break the bonds of the reactants, and how much energy is released when the products are formed

Bond Enthalpies

$2H_{2(g)} + O_{2(g)}$ $2H_2O_{(g)}$

$\mathbf{H}_{\mathbf{break\ bonds}}$

From the table,

H-H = 436kJ/mol

O=O (or O_2) = 495kJ/mol

Bond Enthalpies

$2H_{2(g)} + O_{2(g)}$ $2H_2O_{(g)}$

$H_{\text{break bonds}}$

From the table,

H-H = 436kJ/mol but there's $2H_2$ so, this becomes 2 x 436 = 872kJ/mol

O=O (or O_2) = 495kJ/mol

Therefore, total $H_{\text{reactants}}$ = 872 + 495 = 1367kJ/mol.

Bond Enthalpies

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

$H_{\text{bonds formed}}$

From the table,

O-H = 463kJ/mol.

How many of these O-H bonds are formed in the products?

Bond Enthalpies

$2H_{2(g)} + O_{2(g)}$ $2H_2O_{(g)}$

$H_{\text{bonds formed}}$

From the table,

O-H = 463kJ/mol.

How many of these O-H bonds are formed in the products?

Well, there's 2 in each H_2O molecule and the reaction produces two water molecules so….

Bond Enthalpies

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

$\mathbf{H_{bonds\ formed}}$

From the table,

O-H = 463kJ/mol.

$H_{products}$ = 463 x 2 x 2 = 1852kJ/mol.

Bond Enthalpies

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

So now we've got the Enthalpies for breaking the bonds and how much is released when they're reformed. We'll use this formula

$\Delta H = H_{break\ bonds} - H_{bonds\ formed}$

Bond Enthalpies

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

So now we've got the Enthalpies for breaking the bonds and how much is released when they're reformed. We'll use this formula

$\Delta H = H_{break\ bonds} - H_{bonds\ formed}$

$= 1367 - 1852$

$= -485 kJ/mol.$

Bond Enthalpies

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

So this reaction has a ΔH of -485kJ/mol. and we write that on the side of the reaction like this.

Bond Enthalpies

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$ $\Delta H = -485 kJ/mol.$

That’s a negative ΔH which means that it’s exothermic.

This states that energy is released when this reaction happens. Which makes sense, you don’t really want to do this reaction in real life because it’s what rockets do to take off.

You combust hydrogen gas with oxygen and you’ll end up with water.

Bond Enthalpies

Alright, you try one.

Use the Bond Enthalpy table to determine if the following reaction is endothermic or exothermic.

$2N_{2(g)} + 3H_{2(g)} \rightarrow NH_{3(g)}$

Using this equation

$\Delta H = H_{break\ bonds} - H_{bonds\ formed}$

Btw, N_2 is two nitrogen atoms bonded together with a triple bond.

Bond Enthalpies

Alright, you try one.

Use the Bond Enthalpy table to determine if the following reaction is endothermic or exothermic.

$2N_{2(g)} + 3H_{2(g)} \rightarrow NH_{3(g)}$

ANSWER: +844kJ/mol.
It is endothermic

2 Formulas? WTF?

Ok, so you might be getting a bit confused with the two formulas for ΔH.

There's this one, $\Delta H = H_{products} - H_{Reactants}$

And this one,
$\Delta H = H_{break\ bonds} - H_{bonds\ formed}$

Which really end up giving you opposite answers right? They're like the opposite of each other right?

2 Formulas? WTF?

Here's what we'll do,

If you're ever using the **Bond enthalpies** to calculate something use this one,

$\Delta H = H_{break\ bonds} - H_{bonds\ formed}$

For everything else, you'll use this one

$\Delta H = H_{products} - H_{reactants}$ *Hess's Law*

I'll try and make it easier for you as we go. Next powerpoint looks at enthalpy diagrams and calorimetry.