

Chapter 12

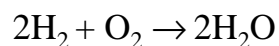
Stoichiometry

Stoichiometry

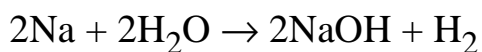
- ◆ Greek for “measuring elements”
- ◆ The calculations of quantities in chemical reactions based on a balanced equation.
- ◆ We can interpret balanced chemical equations several ways.

In terms of Particles

- ◆ Atom - Element
- ◆ Molecule
 - Molecular compound (non- metals)
 - or diatomic (O_2 etc.)
- ◆ Formula unit
 - Ionic Compounds (Metal and non-metal)



- ◆ Two molecules of hydrogen and one molecule of oxygen form two molecules of water.
- ◆ $2 Al_2O_3 \rightarrow 4Al + 3O_2$
2 formula units Al_2O_3 form 4 atoms Al
and 3 molecules O_2

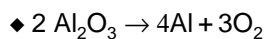


2 atoms Na and 2 molecules H_2O form
2 formula units NaOH and 1 molecule H_2

Look at it differently

- ◆ $2H_2 + O_2 \rightarrow 2H_2O$
- ◆ 2 dozen molecules of hydrogen and 1 dozen molecules of oxygen form 2 dozen molecules of water.
- ◆ $2 \times (6.02 \times 10^{23})$ molecules of hydrogen and $1 \times (6.02 \times 10^{23})$ molecules of oxygen form $2 \times (6.02 \times 10^{23})$ molecules of water.
- ◆ 2 moles of hydrogen and 1 mole of oxygen form 2 moles of water.

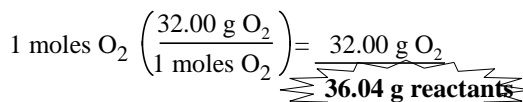
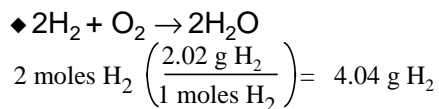
In terms of Moles



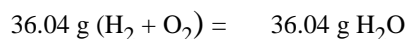
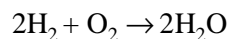
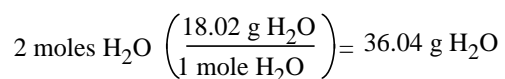
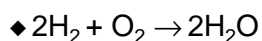
- ◆ The coefficients tell us how many moles of each kind

In terms of mass

- ◆ The law of conservation of mass applies
◆ We can check using moles

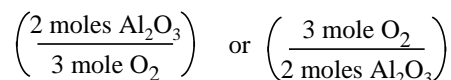


In terms of mass



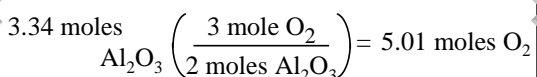
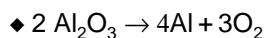
Mole to mole conversions

- ◆ $2 \text{Al}_2\text{O}_3 \rightarrow 4\text{Al} + 3\text{O}_2$
◆ every time we use 2 moles of Al_2O_3 we make 3 moles of O_2



Mole to Mole conversions

- ◆ How many moles of O_2 are produced when 3.34 moles of Al_2O_3 decompose?



Practice

- ◆ $2\text{C}_2\text{H}_2 + 5 \text{O}_2 \rightarrow 4\text{CO}_2 + 2 \text{H}_2\text{O}$
◆ If 3.84 moles of C_2H_2 are burned, how many moles of O_2 are needed?

Practice

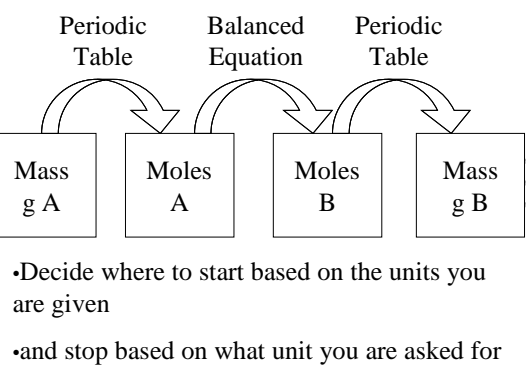
- ◆ $2\text{C}_2\text{H}_2 + 5\text{O}_2 \rightarrow 4\text{CO}_2 + 2\text{H}_2\text{O}$
- ◆ How many moles of C_2H_2 are needed to produce 8.95 mole of H_2O ?

Mass in Chemical Reactions

How much do you make?
How much do you need?

We can't measure moles!!

- ◆ What can we do?
- ◆ We can convert grams to moles.
 - Periodic Table
- ◆ Then use moles to change chemicals
 - Balanced equation
- ◆ Then turn the moles back to grams.
 - Periodic table



Conversions

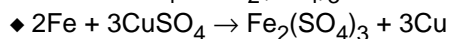
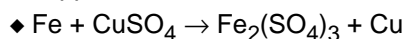
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Conversions

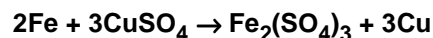
- ◆ $2\text{C}_2\text{H}_2 + 5\text{O}_2 \rightarrow 4\text{CO}_2 + 2\text{H}_2\text{O}$
- ◆ If 2.47 moles of C_2H_2 are burned, how many g of CO_2 are formed?

For example...

- ◆ If 10.1 g of Fe are added to a solution of Copper (II) Sulfate, how much solid copper would form?



$$10.1 \text{ g Fe} \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) \left(\frac{3 \text{ mol Cu}}{2 \text{ mol Fe}} \right) \left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} \right) = 17.2 \text{ g Cu}$$



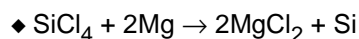
$$0.181 \text{ mol Fe} \left(\frac{3 \text{ mol Cu}}{2 \text{ mol Fe}} \right) = 0.272 \text{ mol Cu}$$

$$0.272 \text{ mol Cu} \left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} \right) = 17.3 \text{ g Cu}$$

Could have done it

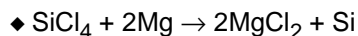
$$10.1 \text{ g Fe} \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) \left(\frac{3 \text{ mol Cu}}{2 \text{ mol Fe}} \right) \left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} \right) = 17.3 \text{ g Cu}$$

- ◆ To make silicon for computer chips they use this reaction



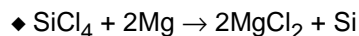
- ◆ How many moles of Mg are needed to make 9.3 g of Si?

- ◆ To make silicon for computer chips they use this reaction



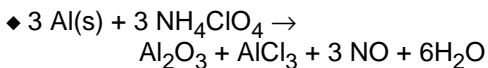
- ◆ 3.74 g of Mg would make how many moles of Si?

- ◆ To make silicon for computer chips they use this reaction



- ◆ How many grams of MgCl_2 are produced along with 9.3 g of silicon?

◆ The U. S. Space Shuttle boosters use this reaction



- ◆ How much Al must be used to react with 652 g of NH_4ClO_4 ?
- ◆ How much water is produced?
- ◆ How much AlCl_3 ?

How do you get good at this?

Practice!!

Gases and Reactions

We can also change

- ◆ Liters of a gas to moles
- ◆ At STP
 - 0°C and 1 atmosphere pressure
- ◆ At STP 22.4 L of a gas = 1 mole

For Example

◆ If 6.45 grams of water are decomposed, how many liters of oxygen will be produced at STP?

- ◆ $\text{H}_2\text{O} \rightarrow \text{H}_2 + \text{O}_2$
- ◆ $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

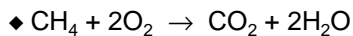
$$6.45 \text{ g H}_2\text{O} \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}} \right) \left(\frac{22.4 \text{ L O}_2}{1 \text{ mol O}_2} \right)$$

Your Turn

◆ How many liters of CO_2 at STP will be produced from the complete combustion of 23.2 g C_4H_{10} ?

Example

- ◆ How many liters of CH₄ at STP are required to completely react with 17.5 L of O₂ ?



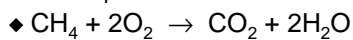
$$17.5 \text{ L O}_2 \left(\frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right) \left(\frac{1 \text{ mol CH}_4}{2 \text{ mol O}_2} \right) \left(\frac{22.4 \text{ L CH}_4}{1 \text{ mol CH}_4} \right) = 8.75 \text{ L CH}_4$$

Avogadro told us

- ◆ Equal volumes of gas, at the same temperature and pressure contain the same number of particles.
- ◆ Moles are numbers of particles
- ◆ You can treat reactions as if they happen liters at a time, as long as you keep the temperature and pressure the same.

Example

- ◆ How many liters of CO₂ at STP are produced by completely burning 17.5 L of CH₄ ?



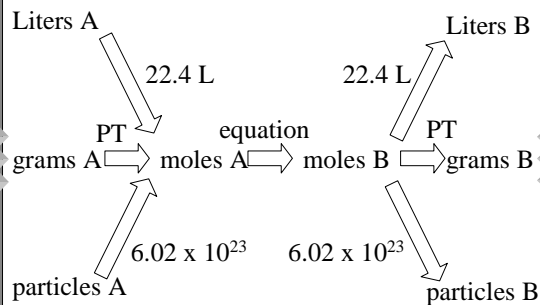
$$17.5 \text{ L CH}_4 \left(\frac{1 \text{ L CO}_2}{1 \text{ L CH}_4} \right) = 17.5 \text{ L CO}_2$$

Particles

- ◆ We can also change between particles and moles.
- ◆ 6.02×10^{23}
 - Molecules
 - Atoms
 - Formula units

Example

- ◆ If 2.8 g of C₄H₁₀ are burned completely, how many water molecules will be made?



Limiting Reagent

- ◆ If you are given one dozen loaves of bread, a gallon of mustard and three pieces of salami, how many salami sandwiches can you make?
- ◆ The limiting reagent is the reactant you run out of first.
- ◆ The excess reagent is the one you have left over.
- ◆ The limiting reagent determines how much product you can make

How do you find out?

- ◆ Do two stoichiometry problems.
- ◆ The one that makes the least product is the limiting reagent.
- ◆ For example
- ◆ Copper reacts with sulfur to form copper (I) sulfide. If 10.6 g of copper reacts with 3.83 g S how much product will be formed?

◆ If 10.6 g of copper reacts with 3.83 g S. How many grams of Cu_2S will be formed?

◆ $2\text{Cu} + \text{S} \rightarrow \text{Cu}_2\text{S}$

$10.6 \text{ g Cu} \left(\frac{1 \text{ mol}}{63.55 \text{ g Cu}} \right) \left(\frac{1 \text{ mol Cu}_2\text{S}}{2 \text{ mol Cu}} \right) \left(\frac{159.16 \text{ g Cu}_2\text{S}}{1 \text{ mol Cu}_2\text{S}} \right) = 13.3 \text{ g Cu}_2\text{S}$

$3.83 \text{ g S} \left(\frac{1 \text{ mol S}}{32.06 \text{ g S}} \right) \left(\frac{1 \text{ mol Cu}_2\text{S}}{1 \text{ mol S}} \right) \left(\frac{159.16 \text{ g Cu}_2\text{S}}{1 \text{ mol Cu}_2\text{S}} \right) = 19.0 \text{ g Cu}_2\text{S}$

Cu is Limiting Reagent

How much excess reagent?

- ◆ Use the limiting reagent to find out how much excess reagent you used
- ◆ Subtract that from the amount of excess you started with

- ◆ $\text{Mg}(s) + 2 \text{HCl}(g) \rightarrow \text{MgCl}_2(s) + \text{H}_2(g)$
- ◆ If 4.87 mol of magnesium and 9.84 mol of HCl gas are reacted, how many moles of gas will be produced?
 - ◆ What is the limiting reagent?

- ◆ $\text{Mg}(s) + 2 \text{HCl}(g) \rightarrow \text{MgCl}_2(s) + \text{H}_2(g)$
- ◆ If 4.87 mol of magnesium and 9.84 mol of HCl gas are reacted, how many moles of gas will be produced?
 - ◆ How much excess reagent remains?

- ◆ If 10.3 g of aluminum are reacted with 51.7 g of CuSO_4 how much copper will be produced?
- ◆ How much excess reagent will remain?

YIELD

Yield

- ◆ The amount of product made in a chemical reaction.
- ◆ There are three types
- ◆ Actual yield- what you get in the lab when the chemicals are mixed
- ◆ Theoretical yield- what the balanced equation tells you you should make.
- ◆ Percent yield = $\frac{\text{Actual}}{\text{Theoretical}} \times 100\%$

Example

- ◆ 6.78 g of copper is produced when 3.92 g of Al are reacted with excess copper (II) sulfate.
- ◆ $2\text{Al} + 3\text{CuSO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{Cu}$
- ◆ What is the actual yield?
- ◆ What is the theoretical yield?
- ◆ What is the percent yield?
- ◆ If you had started with 9.73 g of Al, how much copper would you expect?

Details

- ◆ Percent yield tells us how "efficient" a reaction is.
- ◆ Percent yield can not be bigger than 100 %.

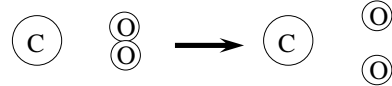
Energy in Chemical Reactions

How Much?
In or Out?

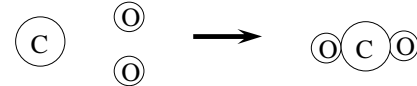
Energy

- ◆ Energy is measured in Joules or calories
- ◆ Every reaction has an energy change associated with it
- ◆ Exothermic reactions release energy, usually in the form of heat.
- ◆ Endothermic reactions absorb energy
- ◆ Energy is stored in bonds between atoms
- ◆ Making bonds gives energy
- ◆ Breaking bonds takes energy

In terms of bonds



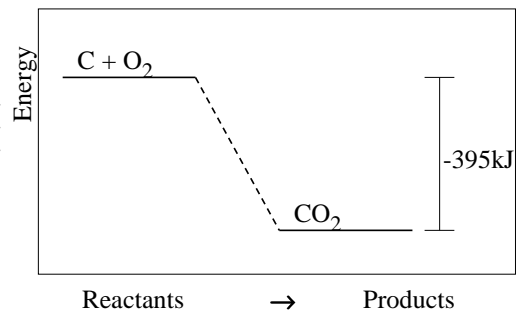
Breaking this bond will require energy



Making these bonds gives you energy
In this case making the bonds gives you more energy than breaking them

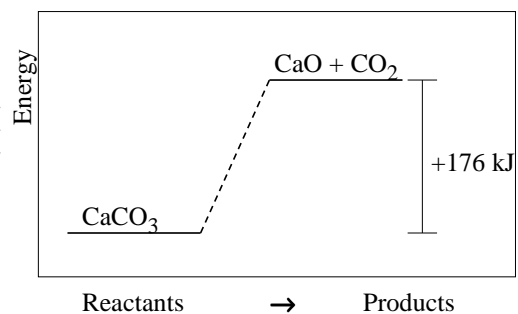
Exothermic

- ◆ The products are lower in energy than the reactants
- ◆ Releases energy
- ◆ Often release heat



Endothermic

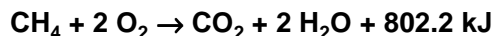
- ◆ The products are higher in energy than the reactants
- ◆ Absorbs energy
- ◆ Absorb heat



Chemistry Happens in

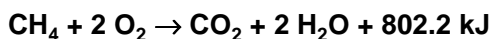
◆ MOLES

- ◆ An equation that includes energy is called a thermochemical equation
- ◆ $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} + 802.2 \text{ kJ}$
- ◆ 1 mole of CH_4 makes 802.2 kJ of energy.
- ◆ When you make 802.2 kJ you make 2 moles of water



- ◆ If 10.3 grams of CH_4 are burned completely, how much heat will be produced?

$$10.3 \text{ g CH}_4 \left(\frac{1 \text{ mol CH}_4}{16.05 \text{ g CH}_4} \right) \left(\frac{802.2 \text{ kJ}}{1 \text{ mol CH}_4} \right) = 515 \text{ kJ}$$



- ◆ How many liters of O_2 at STP would be required to produce 23 kJ of heat?
- ◆ How many grams of water would be produced with 506 kJ of heat?

Calorimetry

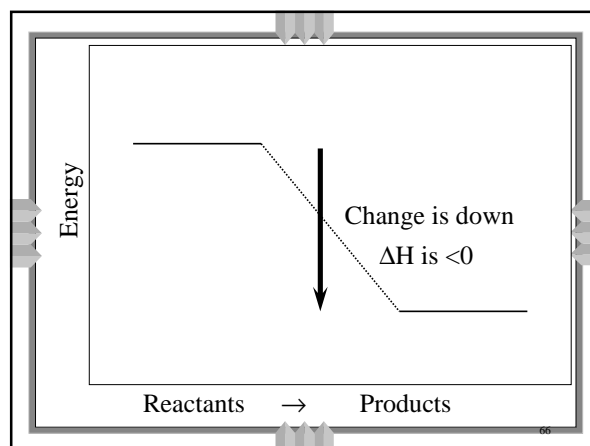
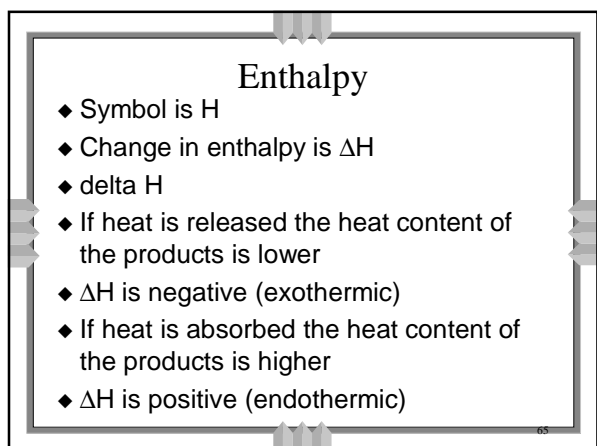
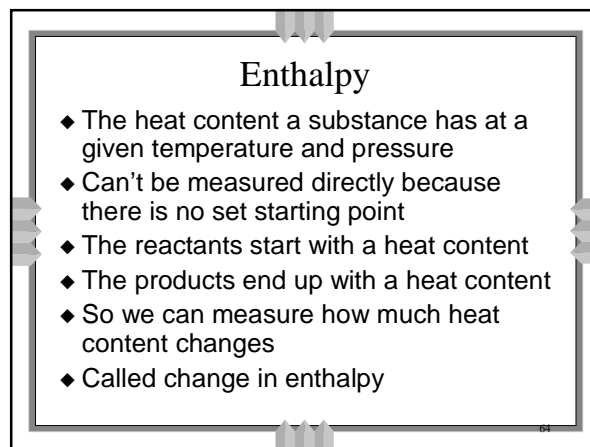
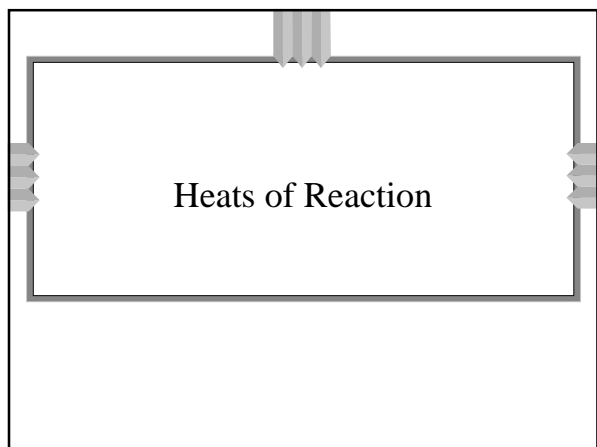
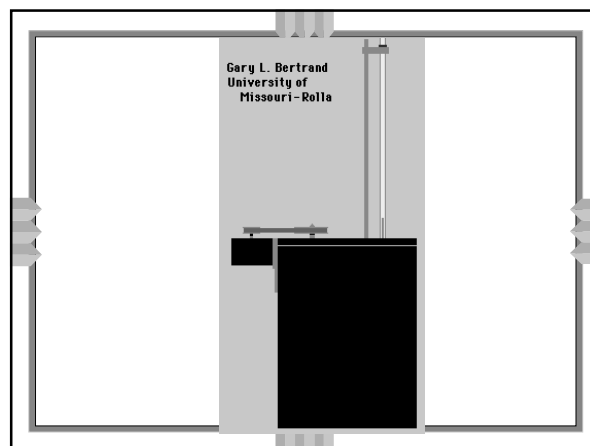
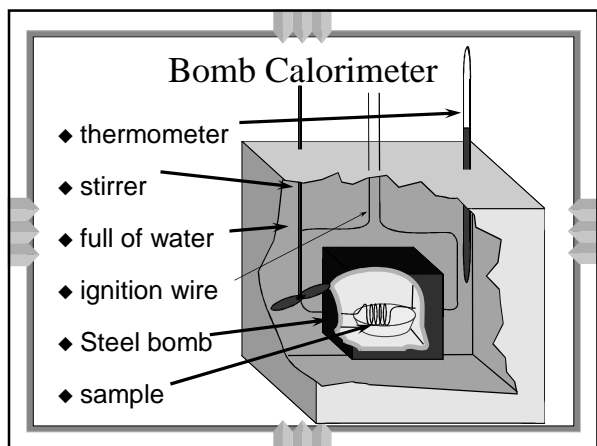
- ◆ Measuring heat.
- ◆ Use a calorimeter.
- ◆ Two kinds
- ◆ Constant pressure calorimeter (called a coffee cup calorimeter)
- ◆ An insulated cup, full of water.
- ◆ The specific heat of water is $1 \text{ cal/g}^\circ\text{C}$
- ◆ $\text{heat} = \text{specific heat} \times m \times \Delta T$

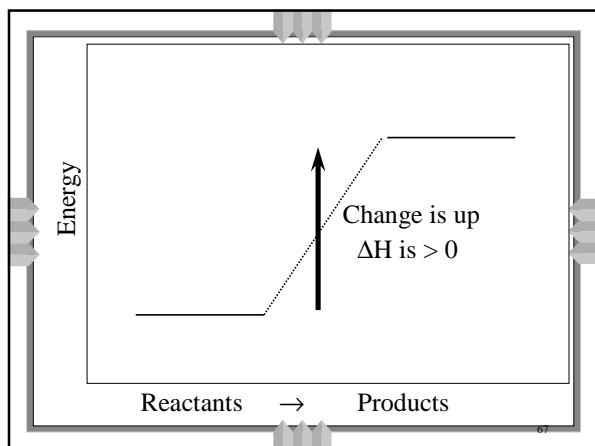
Example

- ◆ A chemical reaction is carried out in a coffee cup calorimeter. There are 75.8 g of water in the cup, and the temperature rises from 16.8°C to 34.3°C . How much heat was released?

Calorimetry

- ◆ Second type is called a bomb calorimeter.
- ◆ Material is put in a container with pure oxygen. Wires are used to start the combustion. The container is put into a container of water.
- ◆ The heat capacity of the calorimeter is known and/or tested. ($\text{cal}/^\circ\text{C}$)
- ◆ Multiply temperature change by the heat capacity to find heat





Heat of Reaction

- ◆ The heat that is released or absorbed in a chemical reaction
- ◆ Equivalent to ΔH
- ◆ $C + O_2(g) \rightarrow CO_2(g) + 393.5 \text{ kJ}$
- ◆ $C + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5 \text{ kJ}$
- ◆ In thermochemical equation it is important to say what state
- ◆ $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g) \quad \Delta H = -241.8 \text{ kJ}$
- ◆ $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8 \text{ kJ}$

Heat of Combustion

- ◆ The heat from the reaction that completely burns 1 mole of a substance
- ◆ $C_2H_4 + 3 O_2 \rightarrow 2 CO_2 + 2 H_2O$
- ◆ $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
- ◆ $2 C_2H_6 + 7 O_2 \rightarrow 2 CO_2 + 6 H_2O$
- ◆ $C_2H_6 + (7/2) O_2 \rightarrow CO_2 + 3 H_2O$

Standard Heat of Formation

- ◆ The ΔH for a reaction that produces 1 mol of a compound from its elements at standard conditions
- ◆ Standard conditions 25°C and 1 atm.
- ◆ Symbol is ΔH_f°
- ◆ The standard heat of formation of an element is 0
- ◆ This includes the diatomics

What good are they?

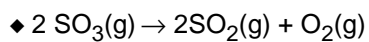
- ◆ There are tables (pg. 190) of heats of formations
- ◆ For most compounds it is negative
 - Because you are making bonds
 - Making bonds is exothermic
- ◆ The heat of a reaction can be calculated by subtracting the heats of formation of the reactants from the products

$$\Delta H = \Delta H_f^\circ(\text{products}) - \Delta H_f^\circ(\text{reactants})$$

Examples

- ◆ $CH_4(g) + 2 O_2(g) \rightarrow CO_2(g) + 2 H_2O(g)$
- ◆ $\Delta H_f^\circ \text{ CH}_4(g) = -74.86 \text{ kJ}$
- ◆ $\Delta H_f^\circ \text{ O}_2(g) = 0 \text{ kJ}$
- ◆ $\Delta H_f^\circ \text{ CO}_2(g) = -393.5 \text{ kJ}$
- ◆ $\Delta H_f^\circ \text{ H}_2O(g) = -241.8 \text{ kJ}$
- ◆ $\Delta H = [-393.5 + 2(-241.8)] - [-74.86 + 2(0)]$
- ◆ $\Delta H = -802.2 \text{ kJ}$

Examples



Why Does It Work?

- ◆ If $\text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \Delta\text{H} = -285.5 \text{ kJ}$
- ◆ then $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{l}) \Delta\text{H} = +285.5 \text{ kJ}$
- ◆ If you turn an equation around, you change the sign
- ◆ $2 \text{H}_2\text{O}(\text{g}) \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{l}) \Delta\text{H} = +571.0 \text{ kJ}$
- ◆ If you multiply the equation by a number, you multiply the heat by that number.

Why does it work?

- ◆ You make the products, so you need their heats of formation
- ◆ You “unmake” the reactants so you have to subtract their heats.

