Chapter 17
Energy

Energy

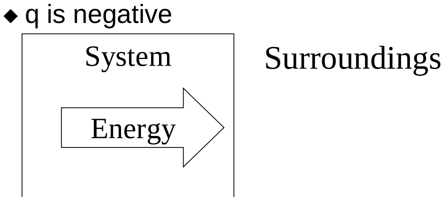
- The ability to do work or cause a change
- Work- using force to move something
  - Symbol is w
- Thermochemistry- studies energy changes in reactions
- q is heat
- Heat flows from high temperature to low temperature

The Universe

- Can be divided into 2 pieces
- System- the part you are investigating
- Surroundings- the rest of the universe
- Law of conservation of energy-
  - Energy can't be created or destroyed
  - The energy of the universe is constant
  - Energy change of System + Energy change of surroundings = 0

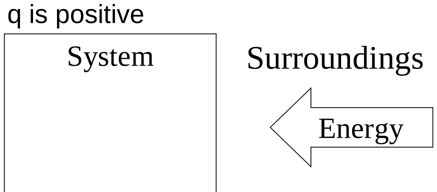
Exothermic

- System releases energy
- Heat flows out
- Surroundings get hotter
- q is negative



Endothermic

- System absorbs energy
- Heat flows in
- Surroundings get cooler
- q is positive



Units of Energy

- Energy is measured in Joules or calories
- calorie is amount of heat to change 1 g of water by 1 °C
- Food Calories are kilocalories
  - 1Calorie = 1000 calories
- 1 cal = 4.184 J

6

Heat capacity

- How much heat it takes to heat an object by 1°C
- Affected by two things
  - What the substance is
  - Mass of the object
- Specific heat is the amount needed to heat 1 g by 1°C
- Only depends on the substance
- Table pg 17.1 Pg. 508

7

Heat capacity

- The higher the specific heat the more energy it takes to change its temperature.
- Pizza burning the roof of your mouth
- The same amount of heat is released when an object cools down

8

Heat capacity

- Equation $q = m\,\Delta T\,C$
- Heat = mass x temp change x specific heat
- Same as Chapter 15

9

- How much heat is needed to change the temperature of 12 g of silver with a specific heat of 0.057 cal/g°C from 25°C to 83 °C?

- If you put 6500 J of heat into a 15 g piece of Al at 25 °C , what will the final temperature be? (C = 0.90 J/g°C)

Calorimetry

- Measuring heat.
- Use a calorimeter.
- Two kinds
- Constant pressure calorimeter (called a coffee cup calorimeter)
- An insulated cup, full of water.
- $q = m\,\Delta T\,C$
- For water C is 1 cal/gºC
- Dissolve chemicals, measure temp before and after

Calorimetry

- Enthalpy (H) – heat content at constant pressure
- Coffee cup calorimeter measure how much heat content changes
- ΔH
- $\Delta H = q$
- We will use heat and change in enthalpy interchangeably
- If temperature goes up exothermic

Example

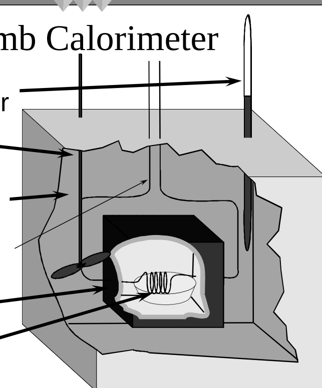
- A chemical reaction is carried out in a coffee cup calorimeter. There are 75.8 g of water in the cup, and the temperature rises from 16.8 ºC to 34.3 ºC. How much heat was released?

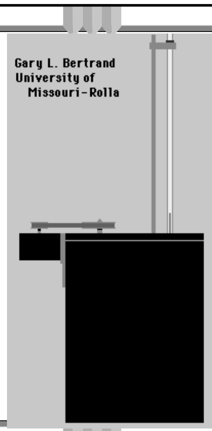
Calorimetry

- Second type is called a bomb calorimeter. (constant volume)
- Material is put in a container with pure oxygen.
- The container is put into a container of water.
- Wires are used to start the combustion.

Bomb Calorimeter

- thermometer
- stirrer
- full of water
- ignition wire
- Steel bomb
- sample





Calorimetry

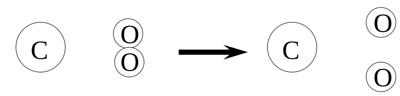
- Run first with a known amount of heat to find the heat capacity of the calorimeter (cal/ ºC)
- Put in your unknown and run a second time
- Multiply temperature change by the heat capacity to find heat of unknown

Thermochemistry

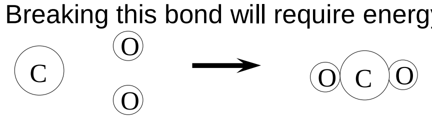
- Every reaction has an energy change associated with it
- Energy is stored in bonds between atoms
- Making bonds gives energy
- Breaking bonds takes energy

19

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

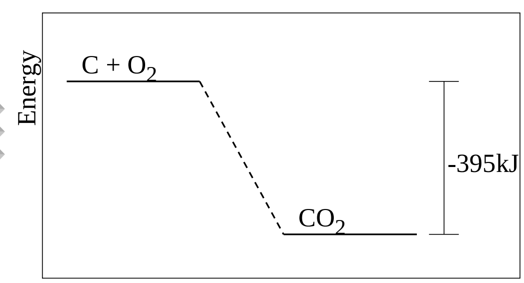
20

Exothermic

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

21

$$C + O_2 \rightarrow CO_2 + 395 \text{ kJ}$$



Reactants → Products

22

When will a reaction be exothermic

A) When breaking the bonds of the reactants takes more energy than making the bonds of the products.

B) When breaking the bonds of the reactants takes less energy than making the bonds of the products

C) When you put in energy to break the bonds

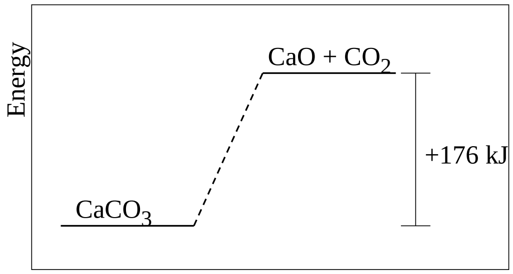
D) When you get energy by breaking bonds

Endothermic

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

24

$CaCO_3 + 176\ kJ \rightarrow CaO + CO_2$



Reactants → Products

25

Chemistry Happens in

- MOLES
- An equation that includes energy is called a thermochemical equation
- $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$
- Energy is a product in this example
- 1 mole of CH_4 makes 802.2 kJ of energy.
- When you make 802.2 kJ you make 2 moles of water

26

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$

- If 10. 3 grams of CH_4 are burned completely, how much heat will be produced?

$$10.\ 3\ g\ CH_4 \left(\frac{1\ mol\ CH_4}{16.05\ g\ CH_4}\right)\left(\frac{802.2\ kJ}{1\ mol\ CH_4}\right)$$

=515 kJ

27

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$

- How many liters of O_2 at STP would be required to produce 23 kJ of heat?

28

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$

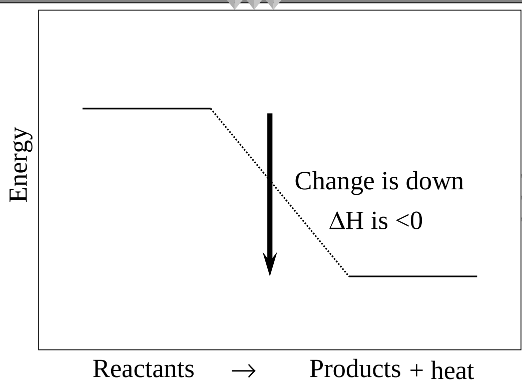
- How many grams of water would be produced with 506 kJ of heat?

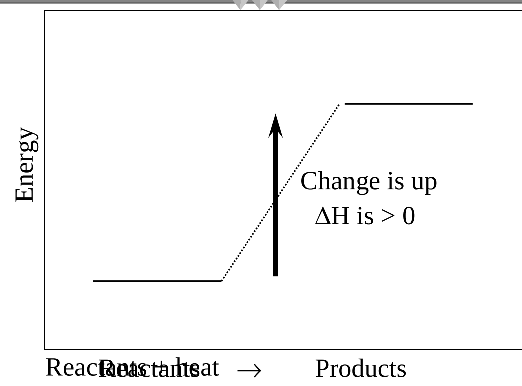
29

Heat of Reaction

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

30





Choose all that apply...

$C(s) + 2\ S(g) \rightarrow CS_2(l)$ $\Delta H = 89.3$ kJ

Which of the following are true?

A) This reaction is exothermic

B) It could also be written

$C(s) + 2\ S(g) + 89.3\text{ kJ} \rightarrow CS_2(l)$

C) The products have higher energy than the reactants

D) It would make the water in the calorimeter colder

Heat of Combustion

- The heat from the reaction that completely burns 1 mole of a substance at 25°C and 1 atm
- $C_2H_4 + 3\ O_2 \rightarrow 2\ CO_2 + 2\ H_2O$
- $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
- $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$
- $C_2H_6 + (7/2)\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$
- Always exothermic

34

Heat and phase change

- Melting and vaporizing are endothermic
  - Breaking things apart
- Freezing and condensing are exothermic
  - Forming connections

Heat of Fusion

- Heat of fusion-ΔH_{fus}- heat to melt one gram
- $q = \Delta H_{fus} \times m$
- For water 80 cal/g or 334 J/g
- Same as heat of solidification
- Book uses molar heat of fusion- heat to melt one mole of solid
- $q = \Delta H_{fus} \times n$

Heat of Vaporization

- ΔH_{vap}- heat to change one gram of liquid to gas
- $q = \Delta H_{vap} \times m$
- For water 540 cal/g or 2260 J/g
- Same as heat of condensation

Calculating Heat

- If there is a temperature change
  - $q = m \Delta T C$
- If there is a phase change
  - $q = \Delta H_{fus} \times m$ or $q = \Delta H_{solid} \times m$
  - $q = \Delta H_{vap} \times m$ or $q = \Delta H_{cond} \times m$
- If there is both, do them separately and add.

Example

- Ammonia has a heat of fusion of 332 cal/g. How much heat to melt 15 g of ammonia?

Example

- Methanol has a heat of vaporization of 1100 J/g. How much heat will be absorbed by 23 g of ethanol vaporizing?

Example

- Butane, C_4H_{10} , absorbs energy as it vaporizes. If 25.3 g of butane absorb 1630 cal by vaporizing, what is the heat of vaporization of butane?

Example

- How much heat does it take to turn 25 g of water at 22°C into steam at 100 °C ?

Heat of Solution

- ΔH_{soln}- heat change when one mole of solute is dissolved.
- $q = \Delta H_{soln} \times n$
- Sometimes endothermic
  - Ammonium nitrate for cold packs
- Sometimes exothermic
  - Acids and bases

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℃ and 1 atm.
- Symbol is ΔH_f^o

◆The standard heat of formation of an element is 0

◆This includes the diatomics

44

What good are they?

- There are tables (pg. 530) 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

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

45

Example

- $CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$

ΔH_f^o $CH_4(g) = -74.86$ kJ

ΔH_f^o $O_2(g) = 0$ kJ

ΔH_f^o $CO_2(g) = -393.5$ kJ

ΔH_f^o $H_2O(g) = -241.8$ kJ

- $\Delta H = [-393.5 \text{ kJ} + 2(-241.8 \text{ kJ})] - [-74.86 \text{ kJ} + 2(0 \text{ kJ})]$
- $\Delta H = -802.2$ kJ

46

Examples

- $2\,SO_3(g) \rightarrow 2SO_2(g) + O_2(g)$

47

Why Does It Work?

- If $H_2(g) + 1/2\,O_2(g) \rightarrow H_2O(l)$ $\Delta H = -285.5$ kJ
- then
  $H_2O(l) \rightarrow H_2(g) + 1/2\,O_2(g)$ $\Delta H = +285.5$ kJ
- If you turn an equation around, you change the sign
- $2\,H_2O(l) \rightarrow 2\,H_2(g) + O_2(g)$ $\Delta H = +571.0$ kJ
- If you multiply the equation by a number, you multiply the heat by that number.
  - Twice the moles, twice the heat

48

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.

49