

CHAPTER 16

Energy & Chemical Change

16.1 Energy

- Energy is the ability to do work or to produce heat.
- Light, heat, electricity

2 Types:

1. **Kinetic:** Energy of motion
2. **Potential:** due to chemical composition or physical position

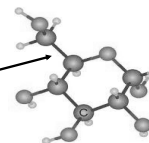
Law of Conservation of Energy

- Energy is never created or destroyed
- Energy is converted from one form to another
- When wood burns...
 - > *Energy stored in the bonds of the wood is liberated as heat.*
- When hot metal is put in cold water...

Chemical Potential Energy

- Energy stored in substances due to their chemical compositions, which is released in chemical reactions
- Octane is C_8H_{18} (gasoline)
- Glucose $C_6H_{12}O_6$

Energy is stored in the bonds.
When our bodies burn glucose, the bonds are broken and energy is released.



Heat

- Represented by "q"
- form of energy that always flows from a warmer object to a cooler object.
- Measured by the flow of energy and the temperature change

Units that measure Heat

- **Calorie:** amount of heat required to raise 1-gram of pure water 1°C. (cal)
- **SI unit of heat & energy is the joule. (J)**
 - > 1 calorie = 4.184 joules
- **Nutritional calories are kilocalories**

Dietary Calories

- 1 "C" = 1 kilocalorie = 1000 calories
- 10 grams of sugar has 41 calories of heat
- Means that when your body burns 10 grams of sugar, it releases 41,000 calories of energy.

SPECIFIC HEAT

The specific heat capacity, or simply the specific heat, represented by "C", of a substance is the amount of heat it takes to raise the temperature of 1-g of the substance 1°C and is measured in J/(g x °C).

- Unique for each substance
- See table 17.1 on page 508.

SPECIFIC HEAT CAPACITY

- 100-g of water has the same specific heat capacity as 5-g of water specific heat relates to 1-gram only
- 500-g of aluminum has the same specific heat capacity as 2-g of aluminum.
- Mass does not matter!!!!

Heat Capacity

The amount of heat it takes to change an object's temperature by exactly 1°C.

Depends upon mass and chemical composition.

How much heat can a given mass hold...

Kinetic Energy increases as temperature increases!

Heat Capacity

100-g of water has twice the heat capacity of 50-g of water.

20-g of lead has twice the heat capacity of 10-g of lead.

Mass is important with regard to heat capacity!!!!

Heat Capacity vs. Specific Heat

Water has a greater "C" than iron, regardless of the mass. Specific heat only refers to 1-g.

$$(C_{\text{iron}} = 0.45 \text{ J/g} \times \text{°C})$$
$$(C_{\text{water}} = 4.18 \text{ J/g} \times \text{°C})$$

Does 100-g of Fe have a greater heat capacity than 5-g of water?

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Heat Capacity vs. Specific Heat

1000-g of iron has a greater heat capacity than 5-g of water. ($C_{\text{iron}} = 0.45 \text{ J/g} \times ^\circ\text{C}$)
($C_{\text{water}} = 4.18 \text{ J/g} \times ^\circ\text{C}$)

1000-g Fe

x 0.45 J

460 - J

5-g Water

x 4.18 J

20.9 - J

Calculating Heat Change

$$q = m \times C \times \Delta T$$

$$\Delta H = m \times C \times \Delta T$$

Δ = change in

CALCULATING SPECIFIC HEAT

- Specific Heat = "C"
- Heat = "q"
- Change in temperature = ΔT

heat (Joules or calories)
mass (g) x change in temp. ($^\circ\text{C}$)

Example #1

A piece of copper with a mass of 95.4 grams changes temperature from 25°C to 48°C . The piece of copper absorbs 849 joules during the change. What is the specific heat of copper in $\text{J/g} \times ^\circ\text{C}$? Convert your answer into $\text{cal/g} \times ^\circ\text{C}$.

q:

m:

C:

ΔT :

Example #2

How many kilojoules are absorbed by water when 32.0 grams are heated from 25.0°C to 80.0°C ? ($C_{\text{water}} = 4.18 \text{ J/g} \times ^\circ\text{C}$)

q:

m:

C:

ΔT :

Example #3

A 181 gram chunk of silver has a heat capacity of $42.8 \text{ J/}^\circ\text{C}$. Calculate the specific heat of silver using the information provided?

q:

m:

C:

ΔT :

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Example #4

A 237 gram piece of iron was at 17 °C. We added 1246 joules of heat to the iron. What is the new temperature? ($C_{\text{iron}} = 0.45 \text{ J/g} \times ^\circ\text{C}$)

q:

m:

C:

ΔT :

Changes in heat & temperature

You cannot convert ΔT into Fahrenheit.

Is there such a thing as a $-\Delta T$?

If temperature goes down, it is a $-\Delta H$, but it is a $+\Delta H$ if the temperature goes up.

16.2 Heat In Chemical Reactions

Calorimeters are devices used to measure the amount of heat absorbed or released during chemical or physical processes.

Calorimetry is the accurate and precise measurement of the heat change for chemical and physical processes.

One System vs. Two System

The heat change in one substance, you determine **q**

1 mass
1 ΔT
1 specific heat

The transfer of heat from one substance to another. You do not determine **q**. One temperature increases, the other decreases.

2 masses
2 ΔT
2 specific heats

Example Problem #5 (One System)

How much heat (in joules) is needed to raise 27.0 grams of water from 10.0 °C to 90.0 °C?

q:

m:

C:

ΔT :

$$\text{Heat (q)} = \text{mass} \times C_{\text{water}} \times \Delta T$$

Example Problem #6 (Two different Systems)

A 100 gram drinking glass decreased from 30.0 °C to 18.0 °C as 150 grams of water was added. What was the original temperature of the water?

($C_{\text{glass}} = 0.5 \text{ J/g} \times ^\circ\text{C}$)

$$q_{\text{glass}} = q_{\text{water}}$$

$$m \times C_{\text{glass}} \times \Delta T = m \times C_{\text{water}} \times \Delta T$$

	Glass	Water
q		
mass		
C		
ΔT		

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Example Problem #7 (Two different Systems)

A 130 gram sample of iron is added to 100-mL of water at 19.0 °C. The final temperature of the water/iron mixture is to 25.0 °C. What was the original temperature of the iron? ($C_{\text{iron}} = 0.46 \text{ J/g} \times ^\circ\text{C}$) (1-mL = 1-g)

	Iron	Water
q		
mass		
C		
ΔT		

$$q_{\text{iron}} = q_{\text{water}}$$

$$m \times C_{\text{water}} \times \Delta T = m \times C_{\text{iron}} \times \Delta T$$

Thermochemistry

The study of the heat changes that occur during chemical reactions and physical changes of state.

Chemical Change: new substances created during chemical reaction

Physical Change: same substances, different form.

The Flow of Heat Energy

- A **system** is the specific part of the universe on which you focus your attention. (*Always the Reactants!*)
- The **surroundings** include everything outside the system. (*Products*)
- Together, the system and surroundings constitute the **universe**.
- This determines whether a reaction is exothermic or endothermic.

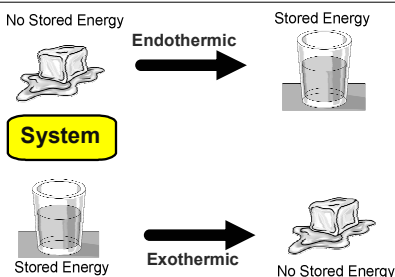
Physical Changes (NO CHEM RXN)

Endothermic: surroundings cool down

- Snow melting or Steam forming
 - Ice \rightarrow Liquid \rightarrow Vapor
- $+\Delta H =$ gaining heat (*Heat into system.*)

Exothermic: surroundings heat up

- Steam condensing or Ice forming
 - Vapor \rightarrow Liquid \rightarrow Ice
- $-\Delta H =$ losing heat (*Heat out of system*)



Chemical Reaction

Endothermic: $A + \text{Energy} \rightarrow B + C$

- Bonds in substance B + C have more stored energy than the bonds in substance A.

Exothermic: $A + B \rightarrow C + \text{Energy}$

- Bonds in substance A + B have more stored energy than the bonds in substance C.

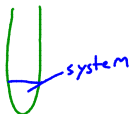
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During a chemical change...

- Process that absorbs heat from the surroundings is called endothermic.

(Positive heat change)

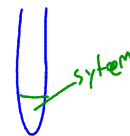
- Test tube feels colder



- A process that loses heat to the surroundings is called exothermic.

(Negative heat change)

- Test tube feels warmer



Endothermic

Nothing would happen to "A" if it didn't absorb energy.

Exothermic

A & B react because they have stored energy, therefore, the energy is released into the surroundings.

Chemical Reaction:

- System is the reactants
- Surroundings are the products

Physical Change:

- The system must be defined
- Hot iron in cold water...
- Iron is system=exothermic
- Water is system=endothermic

Enthalpy (H)

- The heat content of a system at constant pressure.
- You cannot measure the energy of a system, however, you can measure enthalpy changes.
- The change in enthalpy for a reaction is called the enthalpy of reaction (ΔH_{rxn})
- $q = \Delta H_{rxn}$

Enthalpy of Reaction

- The $\Delta H_{rxn} = H_{final} - H_{initial}$
- The $\Delta H_{rxn} = H_{products} - H_{reactants}$
- Exothermic = $-\Delta H_{rxn}$
 $4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) + 1625\text{-kJ}$
- Endothermic = $+\Delta H_{rxn}$
 $27\text{-kJ} + \text{NH}_4\text{NO}_3(s) \rightarrow \text{NH}_4^+(aq) + \text{NO}_3^-(aq)$

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Endothermic: surroundings \rightarrow system



- Positive ΔH (Heat Change)

Exothermic: system \rightarrow surroundings



- Negative ΔH (Heat Change)

16.3 THERMOCHEMICAL EQUATIONS

- Balanced equations that include physical state and the energy change
- Reactants are system, products are surroundings
- You can interpret whether the reaction is exothermic or endothermic by looking at the equation.



Enthalpy of Combustion

- The **heat of combustion** is the enthalpy change for the complete burning of one mole of a substance.
- ΔH_{comb} expressed in kJ/mol
- ΔH^0 = standard conditions (1 atm & 25°C)

Example Problem #8

How much heat is released when combining 25.0-mL of 0.5 molar HCl (aq) with 25.0-mL of 0.5 molar NaOH (aq)? During the reaction, there was a temperature change of 7 °C.

(Hint: aqueous is water) (1-mL = 1-g)



q:

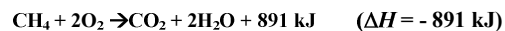
m:

C:

ΔT :

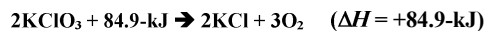
Example Problem #9

Using the equation below, calculate the amount of heat (in kJ) released by burning 24.0-g of CH₄.



Example Problem #10

Using the equation below, calculate the amount of heat (in kJ) required to break down 152.0-g of KClO₃.



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CHANGES OF STATE

- The energy required for a phase change is different than the energy required to raise the temperature of a substance 1°C.
- How much energy (in kJ) is absorbed when 25.0-grams of water increases from water at 40 °C to steam at 120 °C?

Molar Enthalpy of Vaporization

- The heat required to vaporize 1 mole of a liquid
- Endothermic process, hence ΔH is positive
- The reverse of vaporization is condensation and is an endothermic process.
- $\Delta H_{\text{vap}} = -\Delta H_{\text{cond}}$

Molar Enthalpy of Fusion

- The heat required to melt 1 mole of a liquid
- Endothermic process, hence, ΔH is positive
- The reverse of fusion is solidification, which is an exothermic process.
- $\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$

Standard enthalpies of Vaporization and Fusion

Substance	Formula	ΔH_{vap} (kJ/mol)	ΔH_{fus} (kJ/mol)
Water	H ₂ O	40.7	6.01
Ethanol	C ₂ H ₅ OH	38.6	4.94
Methanol	CH ₃ OH	35.2	3.22
Ammonia	NH ₃	23.3	5.66

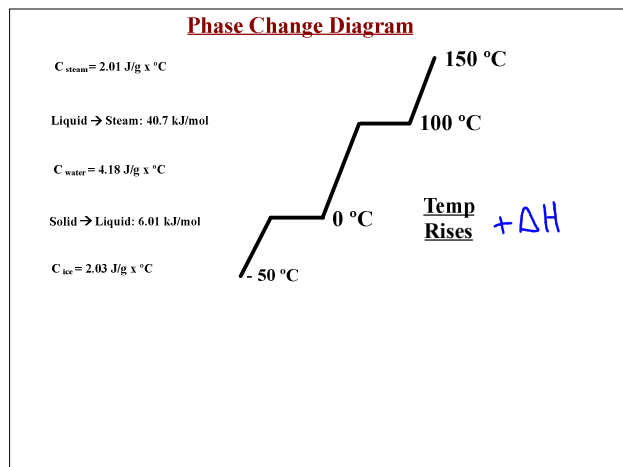
Example Problem

A 56.8-g piece of ice is exposed to 2.0 kJ of heat. Calculate the amount of ice (*in grams*) that should melt? How many grams of ice are left over?

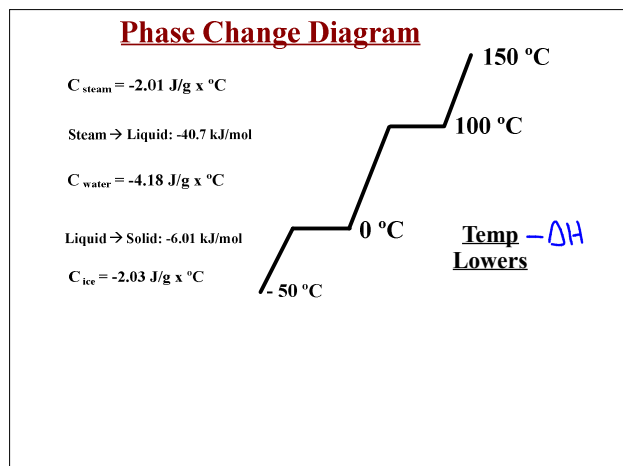
Phase Change Example Problem

A 50.0 gram piece of ice at -125 °C changes into steam at 125 °C. How much energy (in kJ) is absorbed during this process?

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1. Ice @ -125°C \longrightarrow ice @ 0°C
2. Ice @ 0°C \longrightarrow water @ 0°C
3. Water @ 0°C \longrightarrow water @ 100°C
4. Water @ 100°C \longrightarrow steam @ 100°C
5. Steam @ 100°C \longrightarrow steam @ 125°C



Phase Change Example Problem

A 39.0 gram sample of steam at 112°C changes into ice at -7°C . How much energy (in kJ) is released during this process?

1. Steam @ 112°C \longrightarrow Steam @ 100°C
2. Steam @ 100°C \longrightarrow water @ 100°C
3. Water @ 100°C \longrightarrow water @ 0°C
4. Water @ 0°C \longrightarrow ice @ 0°C
5. ice @ 0°C \longrightarrow ice @ -7°C

Molar Heat of Solution

- The heat change caused by dissolution (dissolving) of one mole of substance is the molar heat of solution.
- Grams \rightarrow Moles \rightarrow KJ
- $\Delta H_{\text{solution}}$

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Example Problem

When 5-g of solid NaOH is dissolved in water, how much heat is released? $\Delta H_{\text{solution}} = -445.1 \text{ kJ/mol}$

16.4 Calculating Enthalpy change

Hess's Law: two or more thermochemical equations can be added to produce a final equation

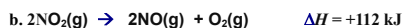
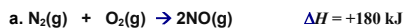
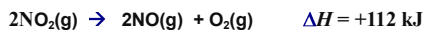
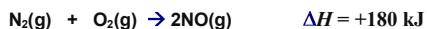
- Allows you to calculate enthalpy changes for hard to simulate reactions by imagining that they occur through steps
- Chemical equations with the desired substances are needed, plus their known enthalpy changes.

Applying Hess's Law

Use the following 2 equations to determine ΔH

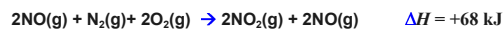


The following equations are used to determine ΔH



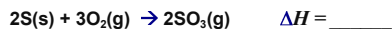
b is reversed
to make c

a & c are
added
together

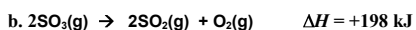
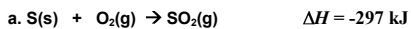
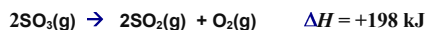
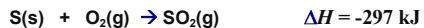


Applying Hess's Law

Use the following 2 equations to determine ΔH

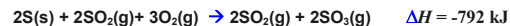


The following equations are used to determine ΔH



a & b were
used to make
c & d

c & d are
added
together



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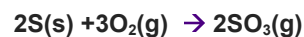
Standard Enthalpy of Formation

- The change in enthalpy that accompanies the formation of one mole of a compound in its standard state form. (See table of ΔH_f° in Appendix C)

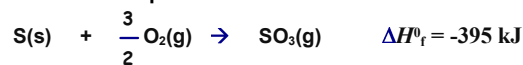
Where do standard enthalpies come from?

- Standard state is 25°C (298 K) & 1-atm of pressure
- They were measured experimentally
- Free elements & diatomic elements have a ΔH_f° of zero.
 - > Free Elements: S, Cu, Fe, Al,

What is the ΔH_f° for the following equation?



Standard enthalpies are for 1 mole of a substance



$\Delta H_f^\circ = 0 \text{ kJ}$ $\Delta H_f^\circ = 0 \text{ kJ}$ $\Delta H_f^\circ = -395 \text{ kJ}$ From Table of ΔH_f°

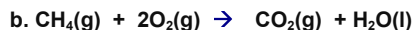
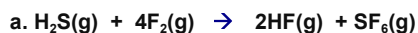
Calculating Enthalpy Change using Standard Enthalpy of Formation

- $\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$
- Subtract the sum of heats of formation of the reactants from the sum of heats of formation of the products.

Standard Enthalpies of Formation

Compound	ΔH_f° (kJ/mol)	Compound	ΔH_f° (kJ/mol)	Compound	ΔH_f° (kJ/mol)
H ₂ S(g)	-21	CCl ₄ (l)	-128	C ₂ H ₅ OH(l)	-278
SO ₂ (g)	-297	CaO	-635	glucose(s)	-1274
SO ₃ (g)	-396	H ₂ SO ₄ (l)	-814	O ₂ (g)	-396
HF(g)	-273	CO(g)	-111	N ₂ O(g)	+82
SF ₆ (g)	-1220	H ₂ O(l)	-285	PCl ₅ (g)	-375
Fe ₂ O ₃ (s)	-824	H ₂ O(g)	-242	PCl ₃ (g)	-287
CO ₂ (g)	-394	NaHCO ₃ (s)	-951	Al ₂ O ₃ (s)	-1680
CaCO ₃ (s)	-1207	Na ₂ CO ₃ (s)	-1131	NaOH(s)	-427
NO(g)	+90	FeS(s)	-100	Hl(g)	+27
NH ₃ (g)	-46	H ₂ O ₂ (l)	-188	C ₂ H ₆ (g)	-85
CH ₄ (g)	-75	I ₂ (g)	+62	KCl(s)	-436
NO ₂ (g)	+33	I ₂ (s)	0	KClO ₄ (s)	-391

Using standard enthalpies of formation, calculate the $\Delta H_{\text{rxn}}^\circ$



16.5 Reaction Spontaneity

- Physical or chemical processes that occur with no outside intervention, however, some energy needs to be applied to get the process started.
 - > Lighting a burner or a sparkler
- Enthalpy and Entropy play a role in whether a reaction is spontaneous or nonspontaneous
- Most spontaneous reactions are exothermic, but not all

Enthalpy vs. Entropy

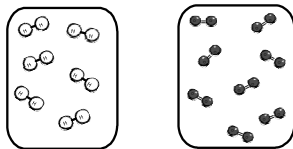
- The heat change (enthalpy) is not the only factor in determining whether a reaction will be spontaneous.
- Entropy is the measure of the disorder or randomness that the particles of a system possess.
- The Law of Disorder states that particles tend to be arranged randomly and that spontaneous process increase the entropy in the universe.

Predicting Entropy Change ΔS

1. Solids > Liquids > Gases
2. Dissolving gases in a solvent decreases entropy
3. When products have more gas particles than the reactants
4. A solid or liquid is dissolved in a solution increases entropy
5. Increasing temperature increases entropy

Reactants with High Entropy are less likely to react because they collide less due to their randomness.

- Reactants with High Entropy forming solids with low entropy is less likely to occur



Reactants with Low Entropy are more likely to react because they collide more due to their fixed positions.

- Reactants with low Entropy forming gases with high entropy is more likely to occur



Entropy of phases:

- Gases > Liquids > Solids



Entropy increases when:

- Solid reactants give gas or liquid products
- Liquid reactants give gas products

Changing the number of parts

N_2 or $2N$ (Which has more entropy?)

$N \equiv N$ — particles

Entropy increases when a substance is divided into parts

$N \ N$ — particles

Dissolved substances

Ionic compounds are dissolved in water and become more separated.



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Enthalpy = ΔH

- **Exothermic:** energy stored in reactant
 - > $A + B \rightarrow C + \text{Heat}$
 - > Spontaneous $-\Delta H$ (Enthalpy Decrease)
- **Endothermic:** reactants have no energy
 - > $A + \text{Heat} \rightarrow B + C$
 - > Nonspontaneous $+\Delta H$ (Enthalpy increase)

$$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surrounding}}$$

The $\Delta S_{\text{universe}}$ is positive, then it is a spontaneous process.

- The entropy of the surroundings increases during an exothermic reaction because heat is released, making the $\Delta S_{\text{surrounding}}$ positive.
- The entropy of a system increases, causing the ΔS_{system} to be positive.

Entropy Increase

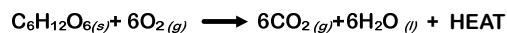
- Fewer particles to more particles
- Solid \rightarrow Gas
- Liquid \rightarrow Gas
- More disorder and randomness on products side

Entropy Decrease

- More particles to fewer particles
- Gas \rightarrow Solid
- Gas \rightarrow Liquid
- More disorder and randomness on reactants side

Ideal Spontaneous Reaction

1. **Exothermic:** releasing energy
 - Enthalpy Decrease, $-\Delta H$
2. **Entropy Increase:** high disorder & randomness = many collisions
 - $N_2 \rightarrow 2N + \text{heat}$



- 1) Exothermic $-\Delta H$
Enthalpy Decrease

- 2.) Entropy Increase
Solid + gas \rightarrow liquid + Gas

Ideal NonSpontaneous Reaction

- Endothermic:** absorb energy
 - Enthalpy Increase, + ΔH
- Entropy Decrease:** low disorder & randomness = no collisions
 - $2N + \text{heat} \rightarrow N_2$



- Endothermic**
Enthalpy Increase
- Entropy Decrease**
liquids \rightarrow solid

Ideal Spontaneous

Enthalpy Decrease
- ΔH
Entropy Increase
Solid \rightarrow Gas

Ideal NonSpontaneous

Enthalpy Increase
+ ΔH
Entropy Decrease
Gas \rightarrow Solid

Everything in between is not ideal. The magnitude of the enthalpy and entropy change determines whether it is spontaneous or nonspontaneous.

Reaction Scenarios

- The following diagrams are "hypothetical" scenarios that will occur if a reaction takes place.
- Look at the scenarios and determine whether or not the reaction will be spontaneous or nonspontaneous.

Spontaneous or Nonspontaneous

- Enthalpy:** (- ΔH) Large Release, **Entropy:** Large Increase
- Enthalpy:** (+ ΔH) Large Gain, **Entropy:** Large Decrease
- Enthalpy:** (+ ΔH) Small Gain, **Entropy:** Large Increase
- Enthalpy:** (- ΔH) Small Loss, **Entropy:** Large Decrease
- Enthalpy:** (- ΔH) Large Loss, **Entropy:** Small Decrease
- Enthalpy:** (+ ΔH) Large Gain, **Entropy:** Small Increase

Ideal Spontaneous

Exothermic
- ΔH
Enthalpy Decrease
Stored energy in Reactants
Low Entropy Reactants
Solid Reactants
Solids form Gases
Entropy Increase

Ideal NonSpontaneous

Endothermic
+ ΔH
Enthalpy Increase
No energy stored in Reactants
High Entropy Reactants
Gaseous Reactants
Gases form Solids
Entropy Decrease

Free Energy

- Useful energy that is liberated in chemical reactions that can be harnessed to do work.
- Just because it is available, it will not necessarily be used.
- 30% of free energy is used to propel a car.
- Exergonic = release of free energy
- This energy is called Gibbs Free Energy
- $\Delta G_{\text{system}} = \Delta H - (T \Delta S)$

Gibbs Free-Energy Summary

- ΔG = Free Energy Change
 - > Negative if spontaneous
 - > **Positive if nonspontaneous**
- ΔH = Enthalpy change
- T = Temperature in Kelvin
- ΔS = Entropy change

Calculating Free Energy Change

- $\Delta H = -91.8 \text{ kJ}$
- T = 298 Kelvin (*Standard Conditions*)
- $\Delta S = -197 \text{ J/K}$
- $\Delta G = -91,800 \text{ J} - (298 \text{ K} \times -197 \text{ J/K})$
= $-91,800 \text{ J} - 58,700 \text{ J}$
= $-33,100 \text{ J}$ of Free Energy
Spontaneous Reaction

Calculating Free Energy Change

- $\Delta H = +2000 \text{ kJ}$
- T = 100 Kelvin
- $\Delta S = -10 \text{ kJ}$
- $\Delta G = +2000 \text{ kJ} - (100 \times -10 \text{ kJ})$
= $+2000 \text{ kJ} - (-1000 \text{ kJ})$
= $+3000 \text{ kJ}$ (No free energy)
Nonspontaneous

Calculating Free Energy Change

- $\Delta H = +2000 \text{ kJ}$
- T = 100 Kelvin
- $\Delta S = +10 \text{ kJ}$
- $\Delta G = 2000 \text{ kJ} - (100 \times 10 \text{ kJ})$
= $2000 \text{ kJ} - 1000 \text{ kJ}$
= $+1000 \text{ kJ}$ (No free energy)
Nonspontaneous