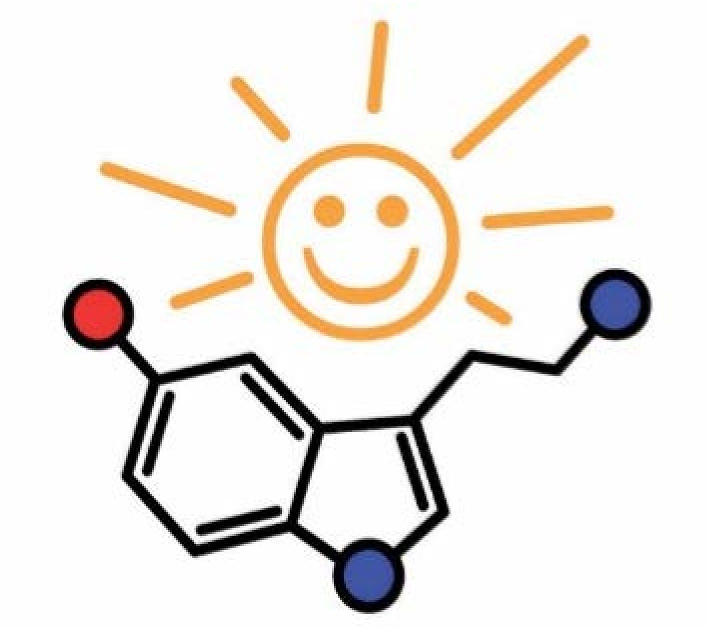
Student Name: ____

Class period: ____

AP Chemistry FTW!



Unit 4: Thermochemistry

&

Unit 5: Thermodynamics

Table of Contents: Unit 4 & 5

Class Info

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2nd Marking Period: October-December 2022

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
9 Oct.	**10** Student Holiday/ Staff Development	**11** **A**	**12** **PSAT**	**13** **B**	**14** **C** End of 1st MP	15
16	**17** **A**	**18** **B**	**19** **A**	**20** **B**	**21** **A**	22
23	**24** Student Holiday/ Staff Development	**25** **B**	**26** **A**	**27** **B**	**28** **A**	29
30	**31** **B**	**1 Nov.** **A**	**2** **B**	**3** **A**	**4** **B**	5
6	**7** **A**	**8** **B**	**9** **A**	**10** **B**	**11** **A**	12
13	**14** **B**	**15** **A**	**16** **B**	**17** **A**	**18** **B**	19

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
20 Nov	**21** FALL BREAK	**22** FALL BREAK	**23** FALL BREAK	**24** FALL BREAK	**25** FALL BREAK	26
27	**28** **A**	**29** **B**	**30** **A**	**1 Dec.** **B**	**2** **A**	3
4	**5** **B**	**6** **A**	**7** **B**	**8** **A**	**9** **B**	10
11	**12** Exams 2, 1	**13** Exams 6, 5	**14** Exams 3, 4	**15** Exams 7, 8	**16** **BREAK BEGINS!**	17
18	**19** **WINTER**	**20** **BREAK**	**21** **WINTER**	**22** **BREAK**	**23** 😊	24
25	**26** 😊	**27** **WINTER**	**28** **BREAK**	**29** **WINTER**	**30** **BREAK**	31

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL = liter(s), milliliter(s)
g = gram(s)
nm = nanometer(s)
atm = atmosphere(s)

mm Hg = millimeters of mercury
J, kJ = joule(s), kilojoule(s)
V = volt(s)
mol = mole(s)

ATOMIC STRUCTURE

$E = h\nu$
$c = \lambda\nu$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}, \text{ where } a\,\mathrm{A} + b\,\mathrm{B} \rightleftarrows c\,\mathrm{C} + d\,\mathrm{D}$$

$$K_p = \frac{(P_\mathrm{C})^c(P_\mathrm{D})^d}{(P_\mathrm{A})^a(P_\mathrm{B})^b}$$

$$K_a = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$K_b = \frac{[\mathrm{OH}^-][\mathrm{HB}^+]}{[\mathrm{B}]}$$

$$K_w = [\mathrm{H}^+][\mathrm{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\mathrm{C}$$
$$= K_a \times K_b$$

$$\mathrm{pH} = -\log[\mathrm{H}^+],\ \mathrm{pOH} = -\log[\mathrm{OH}^-]$$

$$14 = \mathrm{pH} + \mathrm{pOH}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$\mathrm{p}K_a = -\log K_a,\ \mathrm{p}K_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$\ln[\mathrm{A}]_t - \ln[\mathrm{A}]_0 = -kt$$

$$\frac{1}{[\mathrm{A}]_t} - \frac{1}{[\mathrm{A}]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^{\circ}\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$
$= 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$
$= 62.36 \text{ L torr mol}^{-1} \text{ K}^{-1}$
1 atm = 760 mm Hg = 760 torr
STP = 273.15 K and 1.0 atm
Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^{\circ} = \sum S^{\circ} \text{ products} - \sum S^{\circ} \text{ reactants}$$

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

$$\Delta G^{\circ} = \sum \Delta G_f^{\circ} \text{ products} - \sum \Delta G_f^{\circ} \text{ reactants}$$

$$\begin{aligned} \Delta G^{\circ} &= \Delta H^{\circ} - T\Delta S^{\circ} \\ &= -RT \ln K \\ &= -nFE^{\circ} \end{aligned}$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

AP Chem: Effective Study Skills Tips and Tricks!

Study smarter, not harder. ☺

What to Do	What **NOT** to Do
Be **ACTIVE** in while learning/studying: • Close your booklet and try problems on your own with just a periodic table and formula chart! Only check your answer/work when you've finished, or you can't go any farther. • Use flashcards (physical or digital) • Struggle with challenging problems and keep trying, even if you're stuck initially (or convinced you're doing it wrong)	Be passive while learning/studying: • Re-read over your booklet and practice problems you've already completed • Ask your friend or look up the answer if you don't immediately know how to do the problem
Focus when studying • Decrease distractions while studying; don't read texts, check social media, or watch Netflix while studying. Put your phone out of sight/hearing.	Multitask • Study while checking/writing texts, checking social media, and/or watching Netflix. • Keep your computer or tv on in the background
Use **Intensity** when studying • You control the effort that you apply in your work! 30 minutes of high focus, high intensity study can be better than 2 hours of unfocused, low energy multi-tasking.	Low intensity/low effort • Look over problems and try them "in your head" but then just look up the answer • Use flashcards but don't try to recall the info on the other side before looking at the answer
Space out studying over time • Study a little bit of chemistry most days • Start long-term homework (like Mastering Chem or lab reports) the day they're assigned, and work a little bit every day or two • Less is more! Spaced practice studying is more effective than LONG hours of "studying" with multitasking and little focus.	Cram • Only study for quizzes/tests the night before • Start Mastering Chem or your lab report only 1-2 days before it's due • Study for many hours at a time all at once
Interleave your Studying • Study more than just one type of problem; mix it up and jump between different concepts • Review and practice old units while studying (especially important since AP Chem assessments are cumulative!)	One Concept Studying • Study only one type of problem, and practice those problems over and over • Don't review older content or units while studying
Test Yourself! • The best way to prepare for a test is to take a test! ○ Time yourself while trying practice problems ○ Access only the AP Periodic Table and Formula Chart when practicing problems	Open Notes Practice • Use your notes, friends, and/or the internet while trying practice problems • Give yourself unlimited time for each problem

AP Chemistry Unit 4 Objectives

BIG IDEA 2 - Chemical and physical properties of materials can be explained by the structure and arrangement of atoms, ions, molecules and the forces between them.

- Enduring Understanding 2.B: Forces of attraction between particles (including the noble gases and also different parts of large molecules) are important in determining many macroscopic properties of a substance, including how observable physical state changes with temperature.

BIG IDEA 3 - Changes in matter involve the rearrangement and/or the reorganization of atoms and /or the transfer of electrons.

- Enduring Understanding 3.C: Chemical and physical transformation may be observed in several ways and typically involve a change in energy.

BIG IDEA 5 - The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.

- Enduring Understanding 5.A: Two systems with different temperatures that are in thermal contact will exchange energy. The quantity of thermal energy transferred from one system to another is called heat.
- Enduring Understanding 5.B: Energy is neither created nor destroyed, but only transformed from one form to another.
- Enduring Understanding 5.C: Breaking bonds requires energy, and making bonds releases energy.
- Enduring Understanding 5.D: Electrostatic forces exist between molecules as well as between atoms or ions, and breaking the resultant intermolecular interactions requires energy.

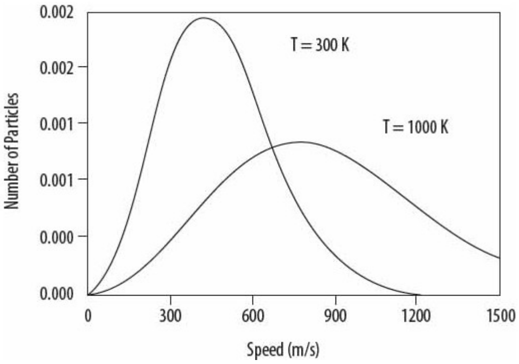
Thermochemistry: Hot, hot, hot!

Thermochemistry: deals with the _______________ changes that occur during chemical reactions.

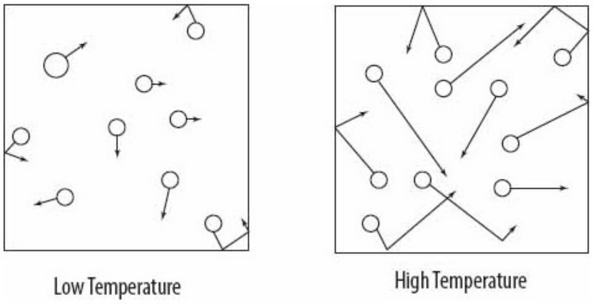
Temperature: measure of the average ______________ energy of the particles of a substance

- → Temperature is an ______________ property: amount of matter doesn't affect it!
- → The Kelvin temperature is directly _________________ to the __________ kinetic energy. For example, doubling the Kelvin temperature doubles the average kinetic energy.
- → As absolute zero is approached (0 K), the particles approach zero kinetic energy.

A Maxwell-Boltzmann distribution shows how the particles at a high temperature have greater kinetic energies than those at a low temperature.



Particle diagrams with _____________ can be used to illustrate differing kinetic energies at differing temperatures.

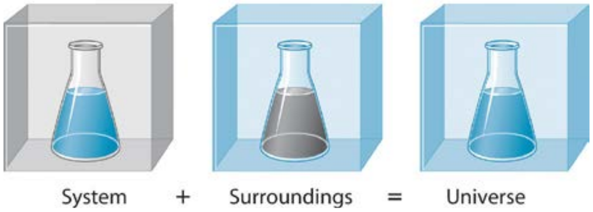


Thermal energy: internal energy of an object due to the _____________ energy of its particles

Heat (q): amount of thermal energy capable of being _________________ from one object to another

- → heat is an ______________ property (depends on how much of a substance you have), unlike temperature.
- → Heat always flows from a _________________ object to a _________________ object.

In thermochemistry, the universe is divided into two halves:



a. the ________________: the substance of interest

b. the ___________________________: whatever is outside the system

How to calculate heat transferred: mCAT!

$$q = mC\Delta T$$

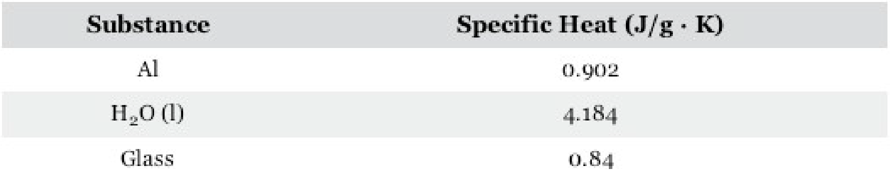
q = heat transferred
m = mass of substance
c = specific heat capacity
$\Delta T = T_{final} - T_{initial}$ = change in temperature

ΔT Object	Sign of ΔT	Sign of *q*	Direction of Heat Transfer
Increase	+	+	Heat transferred into object
Decrease	–	–	Heat transferred out of object

Specific Heat Capacity (C): the amount of heat (energy) required to raise temperature of 1 g of a substance by 1 K (1 °C)

- Units are $\frac{\text{J}}{\text{g °C}}$ or $\frac{\text{J}}{\text{g K}}$
- Metals have relatively _____ specific heats - relatively less energy is required to raise their temperatures.
- Water has a relatively ________ specific heat - requires much more energy to achieve a similar temperature change.

$$\text{Specific Heat Capacity } (C_p) = \frac{\text{quantity of heat supplied}}{\text{(mass of object)(temperature change)}}$$



Substance	Specific Heat (J/g · K)
Al	0.902
H_2O (l)	4.184
Glass	0.84

*Note: the mCAT equation can also be used to calculate heat using moles!

$$q = nC\Delta T$$

n = moles of substance
C = *molar* heat capacity (units $\frac{\text{J}}{\text{mol °C}}$)
Watch out for units!!! ☺

<u>Important</u>: It does NOT matter if you are given °C or K, as long as initial and final temperatures are in the same unit.

Why? Let's try an example.

1. Calculate ΔT for a sample that started at 2°C and was heated up to 27°C:
2. Calculate ΔT for a sample that started at _______ K and was heated up to _______ K:

When adding a FINITE (________________) amount of energy:

- Matter with a low specific heat will change temperature __________
- Matter with a high specific heat will change temperature __________

When adding an unlimited supply of heat (e.g. sitting in the sun)

- Matter with low specific heat will change temperature more _______________
- Matter with high specific heat will change temperature more _______________

Let's Practice!

1. The specific heat (in J/g °C) of solid aluminum is 0.89, of solid iron is 0.45, of liquid mercury is 0.14, and of carbon graphite is 0.71.
   a) When the same amount of heat is applied to one gram of these substances, which one will reach the highest temperature? Explain.

   b) If each substance is heated until they are all the same temperature, which substance required the most heat energy? Explain.

2. You are given a 45 cm^3 sample of copper metal at 300 K and a 30 cm^3 sample of copper metal at 300 K. Which sample contains the most heat? Explain.

3. How much energy, in joules, does it take to heat 15 mL of water from 273 K to 305 K? (The density of water is 1.0 g/mL and the specific heat capacity of water is 4.18 J/g °C).

4. If 0.596 kJ of heat are removed from 29.6 g of water at 22.9°C, what will be the final temperature of the water? (The specific heat capacity of water is 4.18 J/g °C).

5. If 1.82 kJ of heat is required to raise the temperature of a sample of mercury 52°C , and the molar heat capacity of mercury is 28.1 J/mol °C, what is the mass of the sample of mercury?

6. The temperature of a 95.4 g piece of copper increases from 25.0°C to 48.0°C when the copper absorbs 849 J of heat. What is the specific heat of copper?

Calorimetry

Calorimetry: experimental technique used to measure change in ____________ of a chemical reaction or phase change.

1. Put a chemical reaction or phase change in contact with a __________________ bath.
2. Measure the change in temperature of the water bath and then calculate the energy gained or lost by water.
3. The energy change in the water is **EQUAL** and **OPPOSITE** to the heat change by the system!
    a. The system can be an object, a phase change, or a chemical reaction.
        ➔ energy gained by calorimeter = energy lost by the system

$$q_{H_2O} = -q_{object} \text{ or } -q_{rxn}$$

$$+[\mathrm{mC\Delta T}]_{H_2O} = -[\mathrm{mC\Delta T}]_{object} \text{ or } -[\mathrm{mC\Delta T}]_{rxn}$$

$$\text{where } \Delta\mathrm{T} = \mathrm{T}_{final} - \mathrm{T}_{initial} = \mathrm{T}_f - \mathrm{T}_i$$

➔ Remember the specific heat of water: ____________________

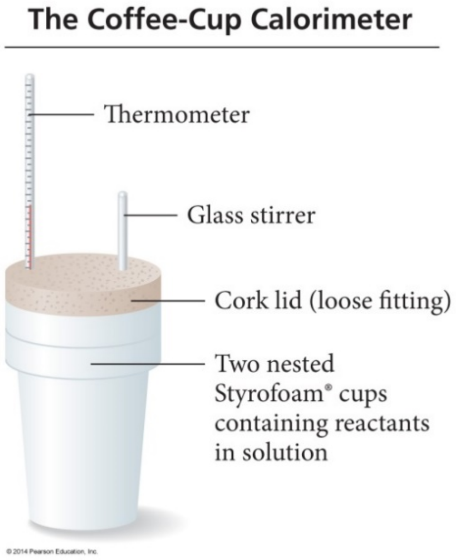
➔ When thermal equilibrium is reached, **BOTH** water bath and the object/reaction have the **SAME final temperature**!

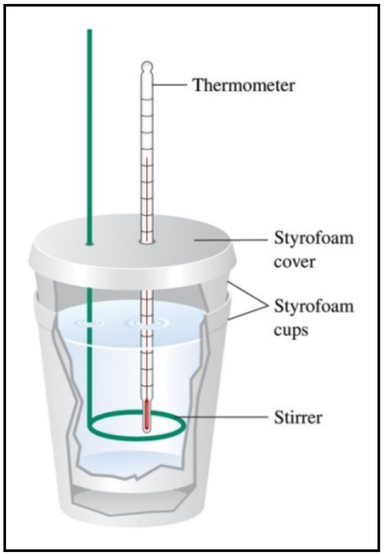
Warning: Experimental Error with Calorimetry!

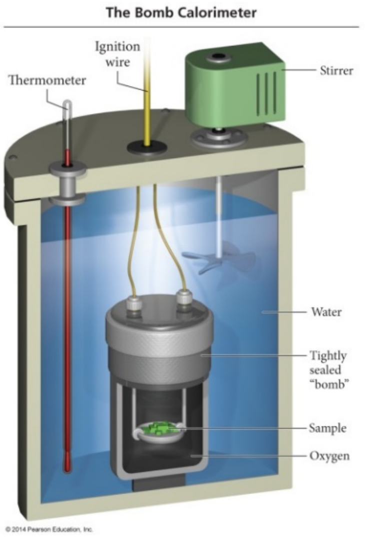
- We assume in the above equation that **ALL** energy lost by the system is gained only by the _____________, but that's not true!
- The calorimeter (container, usually a cup) can also _____________ heat (which means the calorimeter has a measurable heat capacity!), or heat can be lost to the air.
- Both of these errors would lead to a calculated heat (q) that was ____ than the actual heat exchange.

Types of Calorimeters:

1. _________________ **cup calorimeter**: coffee cups are commonly used as insulators in intro chemistry classes to measure temperature changes without a substantial loss of energy to the surroundings (they can be VERY effective!)
2. _____________ **calorimeters** are used at the professional level. Bomb calorimeters provide greater insulation and reduce heat loss to the surroundings (thus minimizing error).







Guided Practice

1. A 5.037 g piece of iron heated to 100.°C is placed in a coffee cup calorimeter that initially contains 27.3 g of water at 21.2°C. If the final temperature is 22.7°C, what is the specific heat capacity of the iron (J/g°C)? The specific heat capacity of water is 4.18 J/g K.

2. A 376 g sample of gold at 400. K is placed in a coffee cup calorimeter containing 50.0 mL of water at 300. K. Determine the final temperature of the water (assuming that no heat is lost to the surroundings). The specific heat capacity of gold is 0.128 J/g°C.

More Practice: Yay!

3. When 25.0 g of a metal at 90.°C is added to 50. g of water at 25.0°C, the temperature of the water rises to 29.8°C. What is the specific heat capacity of the metal? Assume no heat was lost to the surroundings. The specific heat capacity of water is 4.184 J/g°C.

4. A 120. g sample of titanium at 394 K is placed in a coffee cup calorimeter containing 65.0 mL of water at 23.0°C. If no heat is lost to the surroundings, what will be the final temperature of the titanium (in °C)? The specific heat capacity of titanium is 0.523 J/g°C, specific heat capacity of water is 4.18 J/g°C, and the density of water = 1.00 g/mL.

5. A student is provided with a sample of an unknown metal, a coffee cup calorimeter, a temperature probe, and unlimited water. They are asked to identify the unknown metal.

a. Answering the following questions about the experimental method the student should use in this experiment.

i. What measurements will the student need to make in order to identify their unknown metal?

ii. What calculations will the student need to perform? Explain how the student could use the measurements described above to perform these calculations.

iii. How can the student use the results of their calculations to identify their unknown metal? What other information will they need?

Multiple Choice Practice FTW!

6. How much heat is required to raise the temperature of 100. g of Fe_2O_3 from 5.0°C to 25.0°C? The specific heat of Fe_2O_3 is 0.634 J/g°C.

a. 1.27 kJ b. 0.0634 kJ c. 1.58 kJ d. 0.845 kJ

7. For an experiment, 50.0 g of H_2O was added to a coffee-cup calorimeter. The initial temperature of the H_2O was 22.0°C, and it absorbed 300.J of heat from an object that was carefully placed inside the calorimeter. Assuming no heat is transferred to the surroundings, which of the following was the approximate temperature of the H_2O after thermal equilibrium was reached? Assume that the specific heat capacity of H_2O is 4.2 J/(g·K).

a. 21.3°C b. 22.0°C c. 22.7°C d. 23.4°C

8. A 50 g sample of a metal is heated to 100°C and then placed in a calorimeter containing 100.0 g of water (c = 4.18 J/g°C) at 20°C. The final temperature of the water is 24°C. Which metal was used?

a. Lead (c = 0.14 J/g°C) c. Iron (c = 0.45 J/g°C)

b. Copper (c = 0.20 J/g°C) d. Aluminum (c = 0.89 J/g°C)

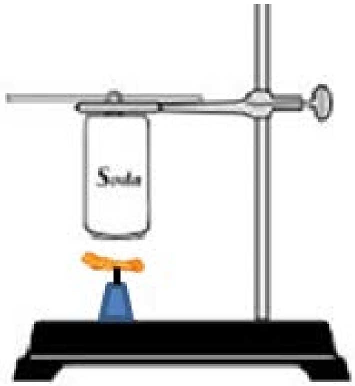
Unit 4 Part 2: Cheeto Lab (Energy of Foods)

Objective: To be able to calculate the number of Calories in one Cheeto by using the formula for specific heat.

Materials:

Thermometer, Ring stand w/ ring clamp, 100 mL graduated cylinder, Food holder (paper clip), Lighter, Stir rod, Electronic balance, Soda can, Cheetos, Water

Note: ALWAYS leave the plastic boat on a balance. It helps protect the balance from stains and marks. **DON'T set the can down anywhere - it is messy!**



Procedure:

1) Obtain one Cheeto and place on the exposed end of the paper clip. Find and record the initial mass of the food sample + food holder (paper clip). **Do NOT eat in the lab!**
2) Measure about 70 mL of water with the graduated cylinder. Record the precise volume and mass of the water in the data table. Pour the water into the soda can.
3) Assemble calorimetry apparatus as shown: the can should be suspended *about* 2.5 cm above the food sample. Include a piece of aluminum foil underneath the food sample for easier clean-up.
4) Place the thermometer into the soda can. Record the initial temperature and place the information in the data table.
5) Contact your teacher to help set the Cheeto on fire. Be sure the burning Cheeto is directly beneath the center of the can. Make sure the ENTIRE Cheeto is burning before removing the lighter.
6) Record the highest temperature of the water as the *final temperature* in the data table.
7) Determine and record the final mass of the burned Cheeto + food holder (paper clip) in the data table.
8) Place burned food sample into the trash once it has cooled - keep the food holder (paper clip) and aluminum foil.

**** ONLY throw away the Cheeto ****

Data:

	Flamin' Hot Cheeto
Initial mass of food and holder	
Final mass of food and holder	
Volume of water	
Mass of water	
Initial temperature, T_i	
Final temperature, T_f	

Calculations/Analysis:

1. Show the calculation for the change in temperature of the water, ΔT.

2. Calculate the heat absorbed by the water, q. For water, C is 4.18 J/g°C. **Convert your final answer to kJ.**

3. Show the calculation to determine the mass (in g) of your food sample that burned.

4. Show the calculation to calculate the energy content (in kJ/g) of your food sample. (Use answers #2 and #3).

5. Based on your answer to #4, calculate the number of kilojoules in a 28 g serving of Cheetos. **USE DIMENSIONAL ANALYSIS TO RECEIVE CREDIT.**

6. Based on your answer to #5, calculate the number of Calories in a 28 g serving of Cheetos. (There are 4.18 kJ in one Calorie.) **USE DIMENSIONAL ANALYSIS TO RECEIVE CREDIT.**

7. According to the package, how many Calories are in a 28 g serving of Cheetos? ________________
8. Calculate your percent error:

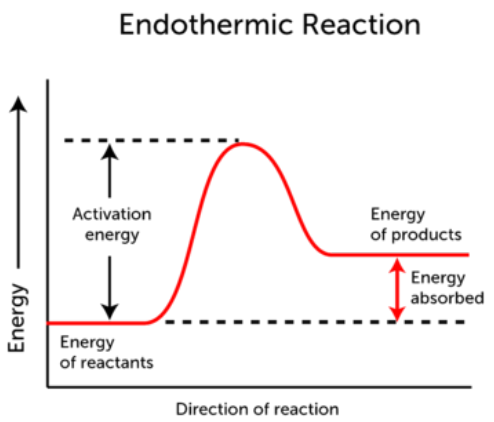
9. Identify one or more sources of error that might have led to your calculated percent error.

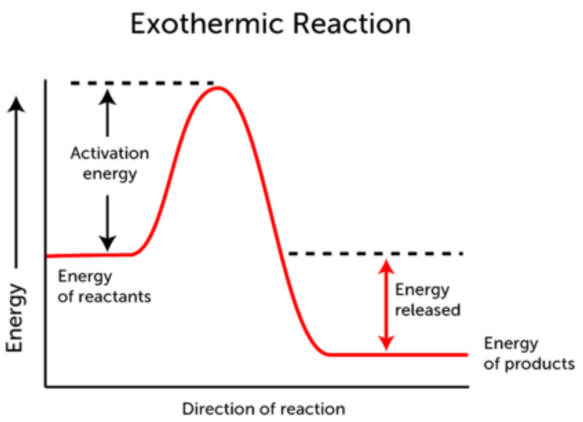
Enthalpy and Thermochemical Equations

Enthalpy Change (ΔH): amount of energy **absorbed** or **released** as heat by a system *per mole of reaction (mol_{rxn})* when the pressure is constant; measured in units of ___________ = $\frac{\text{kJ}}{\text{mol}_{\text{rxn}}}$

Note: Mastering Chem uses kJ (no per mol)

$$\Delta H_{rxn} = \frac{q}{mol_{rxn}}$$


Endothermic Reaction


Exothermic Reaction

- The magnitude of energy change is directly proportional to the ___________ of reactants and products involved in the change.
- __________________ energy is the minimum amount of energy the reactants need for the reaction to proceed.

Thermochemical equation: chemical equation that includes the ___________ change (the energy value)

Endothermic	**Exothermic**
Energy is __________________ by system	Energy is __________________ by system
Heat energy from surroundings is changed to potential energy of the system – temperature of the system doesn't necessarily increase!	Potential energy of the system is changed to **heat** energy of the surroundings – temperature of the system doesn't necessarily decrease!
$\frac{+q}{\text{mol}_{\text{rxn}}} = +\Delta H_{\text{rxn}}$	$\frac{-q}{\text{mol}_{\text{rxn}}} = -\Delta H_{\text{rxn}}$
bonds/attractions	Form
The energy required to break reactant bonds is _____________ than the energy released by forming product bonds.	The energy required to break reactant bonds is _____________ than the energy released by forming product bonds.
Energy appears in	Energy appears in
2 examples of the SAME information:	2 examples of the SAME information:
$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ **ΔH_{rxn} = +92 kJ/mol_{rxn}**	$N_2 + 3H_2 \rightarrow 2NH_3$ **ΔH_{rxn} = −324 kJ/mol_{rxn}**
positive = add(+) to **reactants**	negative = add(+) to **products**
$+ NH_3 + H_2O \rightarrow NH_4^+ + OH^-$	$N_2 + 3H_2 \rightarrow 2NH_3 +$

Phase changes and enthalpy:

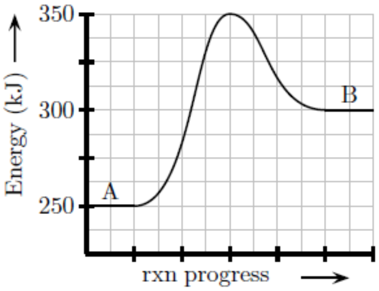
gas

liquid

solid

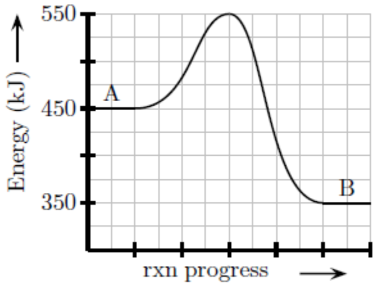
Let's Try! Complete the chart below.

Equation with Separate ΔH°_{rxn}	ΔH°_{rxn} within the Equation	Endo- or exothermic?
	$K + M \rightarrow N + 45\ kJ$	
$D \rightarrow E + F$ $\Delta H^\circ_{rxn} = 127\ kJ/mol_{rxn}$		
$A + B \rightarrow C + D$ $\Delta H^\circ_{rxn} = -35\ kJ/mol_{rxn}$		



1. What is the change in enthalpy (ΔH) for the reaction $A \rightarrow B$

2. Is this reaction endothermic or exothermic?



3. What is the change in enthalpy (ΔH) for the reaction $A \rightarrow B$

4. Is this reaction endothermic or exothermic?

5. For each example below, identify if the process is endothermic or exothermic, and explain why.

Process	Endo or Exothermic?	Why?
$H(g) + H(g) \rightarrow H_2(g)$		
$F(g) + e^- \rightarrow F^-(g)$		
$F^-(g) \rightarrow F(g) + e^-$		
$N_2(g) \rightarrow N_2(l)$		

Energy Stoichiometry! ☺

Enthalpy is commonly measured in kJ/mol_{rxn}, but what is a mole of reaction?

$1\ mol_{rxn}$ = 1 mole of reaction = stoichiometric # of reactants/ products

For the combustion of ethane: $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O + 3{,}120\ kJ$

When ____ mole of reaction has occurred,

- ____ mol of C_2H_6 reacted
- ____ mol of O_2 reacted
- 3120 kJ energy ___________________
- ____ mol of CO_2 were produced
- ____ mol of H_2O were produced

Luckily for us, the enthalpy of a reaction, when measured in kJ/mol_{rxn}, can act as a _________________________ factor between the amount of chemicals which react and the energy that is absorbed or released by the reaction!

Example 1: Give the following reaction, $2\ Fe + 3\ CO_2 \rightarrow 3\ CO + Fe_2O_3$ ($\Delta H = +24.7\ kJ/mol_{rxn}$) what energy change occurs when 6.0 moles of carbon dioxide react?

Example 2: Give the following reaction, $N_2 + 3\ H_2 \rightarrow 2\ NH_3$ ($\Delta H = -324\ kJ/mol_{rxn}$) what mass of hydrogen must have reacted if 525 kJ of heat were released?

Independent Practice

1. The heat of combustion of gaseous propane (C_3H_8) is $-2220\ kJ/mol_{rxn}$.
    a. Write the balanced thermochemical equation for the combustion of propane.

    b. What is the sum of the coefficients when the reaction is written and balanced? ______

    c. Is the reaction endothermic or exothermic? ________________

    d. What mass of propane must be burned to release 5,550 kJ of heat?

2. Solid sodium hydrogen carbonate, $NaHCO_3$, decomposes into solid sodium carbonate, Na_2CO_3, liquid water, and carbon dioxide gas. (ΔH_{rxn} = +85 kJ/mol_{rxn})

a. Write the balanced thermochemical equation for this reaction.

b. Is the reaction endothermic or exothermic? _______________

c. What is the energy change that occurs when 2.25 mol of $NaHCO_3(s)$ decomposes?

3. When liquid ethanol, $C_2H_6O(l)$, burns, it reacts with $O_2(g)$ to produce $CO_2(g)$ and $H_2O(l)$ and 1368 kJ of heat.

a. Write the balanced thermochemical equation for this reaction.

b. Is the reaction endothermic or exothermic? _______________

c. If the reaction takes place at 22°C and 0.92 atm, what volume of carbon dioxide gas will be produced from this reaction if 998.6 kJ of heat is released?

4. When barium hydroxide octahydrate, $Ba(OH)_2 \bullet 8H_2O$ is mixed in a beaker with ammonium thiocyanate, NH_4SCN, a reaction occurs. The beaker becomes very cold.

a. Have the surroundings gained or lost heat? How do you know?

b. Has the system gained or lost heat? How do you know?

c. Is the reaction endothermic or exothermic? ___________________

5. Two aqueous substances in a glass beaker chemically react, and the temperature of the water in the beaker rises.

a. Have the surroundings gained or lost energy? How do you know?

b. Has the system gained or lost energy? How do you know?

c. Is the reaction endothermic or exothermic? ____________________

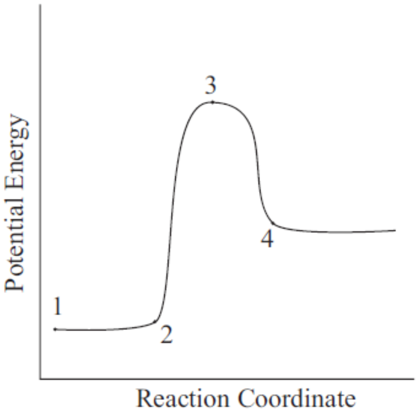
6. Which of the following is true for an exothermic reaction?

a. The strength of the bonds in the products exceeds the strength of the bonds in the reactants.

b. The activation energy is always greater than the activation energy for an endothermic reaction.

c. Energy is absorbed over the course of the reaction.

d. The products have less thermal energy than the reactants.



7. The distance between which two points is equal to the enthalpy change for this reaction?

a. Points 1 and 2
b. Points 1 and 3
c. Points 1 and 4
d. Points 2 and 3

8. The distance between which two points is equal to the activation energy for this reaction?

a. Points 1 and 2
b. Points 1 and 3
c. Points 1 and 4
d. Points 3 and 4

9. Consider the following reaction showing photosynthesis:

$$6\ CO_2(g) + 6\ H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g) \qquad \Delta H = +2800\ \text{kJ/mol}$$

Which of the following is true regarding the thermal energy in this system?

a. It is transferred from the surroundings to the reaction.

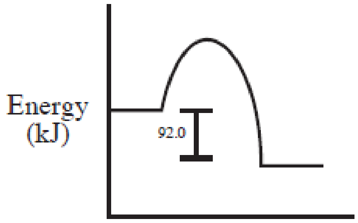
b. It is transferred from the reaction to the surroundings.

c. It is transferred from the reactants to the products.

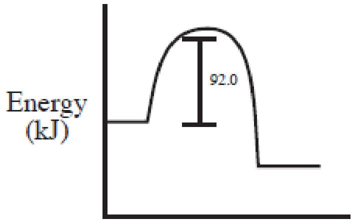
d. It is transferred from the products to the reactants.

$$C(s) + 2\,S(s) \rightarrow CS_2(l) \qquad \Delta H = +92.0\ \text{kJ/mol}$$

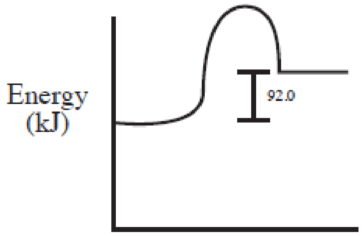
10. Which of the following energy level diagrams gives an accurate representation of the above reaction?



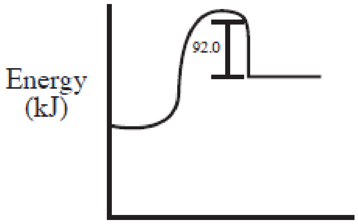
(A)



(B)



(C)



(D)

$$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + H_2O(l) \qquad \Delta H = -890\ \text{kJ/mol}$$

11. Regarding the reaction above, how much heat is absorbed or released when 2.0 mol of CH_4(g) reacts with 2.0 mol of O_2(g)?

a. 890 kJ of heat is released.

b. 890 kJ of heat is absorbed.

c. 1780 kJ of heat is released.

d. 1780 kJ of heat is absorbed.

13. Identify each of the following phrases/pictures as describing an endothermic or exothermic process:

_____ a. The reactants have more potential energy than the products.

_____ b. A liquid evaporates.

_____ c. ΔH is negative.

_____ d. An aqueous chemical reaction occurs, and the temperature of the water drops.

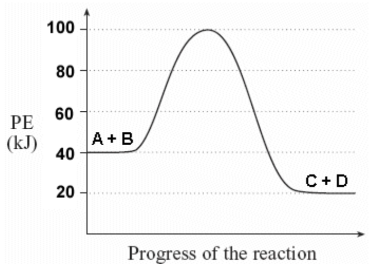
_____ f. The products have more potential energy than the reactants.

_____ g. A chemical reaction occurs, and the container becomes hot to the touch.

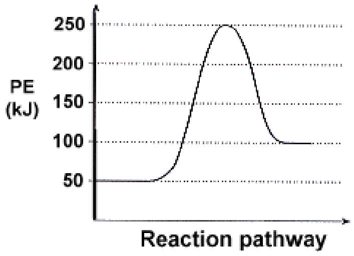
_____ h. ΔH is positive.

_____ i. A gas condenses.

_____ e.



_____ j.



Aqueous Calorimetry

Although many calorimetry labs involve a solid, there are multiple times when you will need to determine the energy change of an aqueous reaction. Two common examples:

1. Determine the enthalpy of neutralization, ΔH_{neut}, for an acid-base neutralization
2. Determine the enthalpy of solution, ΔH_{soln}, when dissolving a salt.

Normal calorimetry math still applies, but… you must __________ the masses of all chemicals reacting (including water if something is being dissolved), because the heat absorbed or released is shared by all chemicals present in the system.

$$+q_{rxn} = -[\mathrm{m}_{total}\mathrm{C}\Delta\mathrm{T}]_{calorimeter}$$

To calculate the enthalpy of reaction (or phase change), you will need to divide q_{rxn} by the moles of reaction.

$$\Delta H_{rxn} = \frac{q_{rxn}}{mol_{rxn}}$$

There are 4 important assumptions made during aqueous calorimetry labs*

1. It is assumed that the density of the aqueous solution is the __________ as for water, 1.00 g/mL.
(100 mL of solution is said to have a mass of 100 g).

2. It is assumed that the volumes of reactants are ______________________.
(100 mL of reactant A + 200 mL of reactant B = 300 mL of reaction mixture)

3. It is assumed that the specific heat capacity of an aqueous reaction mixture is the same as water, 4.18 J/g °C.

4. It is assumed that heat is _______ lost to, or absorbed by, the surroundings.

*If these assumptions are not correct, the problem should specifically tell you otherwise.

Experimental Error with Calorimetry! (The flaws of assumption #4)

- We assume that **ALL** energy lost by the system is gained *only* by the water, but that's not true! **The calorimeter can also absorb heat, or heat can be lost to the surrounding air.**
- **Both of these errors lead to a calculated heat (q) that's SMALLER than the actual heat exchange**, and thus the calculated ΔH would be **<u>smaller</u>** in **MAGNITUDE** (less negative or less positive) than the actual ΔH.
  - Decreased ΔT = decreased q = decreased ΔH

1. A student conducts an experiment to determine the enthalpy of neutralization for sodium hydroxide reacting with phosphoric acid in the reaction shown below. The student combines 25.0 mL each of equimolar solutions of acid and base, both having an initial temperature of 22.50°C, in a coffee cup calorimeter. Each solution has a density of 1.00 g/mL and a specific heat of 4.18 J/g °C. The student collects data until a maximum temperature of 26.40°C is reached.

$$3\ \mathrm{NaOH}(aq) + \mathrm{H_3PO_4}(aq) \rightarrow 3\ \mathrm{H_2O}(l) + \mathrm{Na_3PO_4}(aq)$$

a. Calculate q_{rxn} for this neutralization reaction.

b. Calculate the ΔH_{neut} if the initial molarity of both solutions is equal to 0.80 M. Report your answer in kJ/mol.

2. An experiment was conducted in which 5.19 g of Na_2CO_3 was dissolved into 75.0 g of distilled water. It was observed that the entire system increased temperature by 3.8°C. Calculate the enthalpy of dissolution for sodium carbonate. (Assume that the solution has a density of 1.00 g/mL and a specific heat of 4.18 J/g °C.)

3. A 100.0 mL sample of 0.300 M NaOH is mixed with a 100.0 mL sample of 0.300 M HCl in a coffee cup calorimeter. If both solutions were initially at 35.00°C and the temperature of the resulting solution was recorded as 37.00°C, determine the heat absorbed or released in kJ/mol_{rxn} for the neutralization reaction between aqueous NaOH and HCl. Assume that no heat is lost to the calorimeter or the surrounding, and that the density and the heat capacity of the resulting solution are the same as water.

Multiple Choice Practice FTW! Questions 3 – 5 refer to the following.

Inside a calorimeter, 100.0 mL of 1.0 M hydrocyanic acid (HCN), a weak acid, and 100.0 mL of 0.50 M sodium hydroxide, NaOH, are mixed, and chemical reaction below occurs. The temperature of the mixture rises from 21.5°C to 28.5°C. The specific heat of the mixture is approximately 4.2 J/g°C, and the density is identical to that of water.

$$HCN(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCN$$

4. What is the approximate amount of heat released during the reaction?

c. 1.5 kJ b. 2.9 kJ c. 5.9 kJ d. 11.8 kJ

5. If the experiment is repeated with 100.0 mL of 2.0 M HCN and 100.0 mL of 0.50 M NaOH, what would happen to the values for ΔT and ΔH_{rxn}?

	ΔT	**ΔH_{rxn}**
(A)	Increase	Increase
(B)	Stay the same	Stay the same
(C)	Decrease	Stay the same
(D)	Stay the same	Increase

6. If the experiment is repeated for a third time with 100.0 mL of 1.0 M HCN and 100.0 mL of 1.0 M NaOH, what would happen to the values for ΔT and ΔH_{rxn} (relative to the original experimental results)?

	ΔT	**ΔH_{rxn}**
(A)	Increase	Stay the same
(B)	Stay the same	Stay the same
(C)	Decrease	Stay the same
(D)	Stay the same	Increase

Hess's Law

Hess's Law: Sometimes it is impossible or impractical to measure the ΔH of a reaction by using a calorimeter. In these situations, ΔH can be calculated using Hess's Law!

Hess's Law: Combining two or more reactions to achieve a goal reaction

1. When adding given reactions, they combine to produce the _______ reaction.
2. $\Delta H_{\text{new rxn}} = \Delta H_{\text{rxn 1}} + \Delta H_{\text{rxn 2}} + \Delta H_{\text{rxn 3}} + \ldots$

<u>Note</u>: Sometimes you have to manipulate your given reactions to produce your goal reaction. If so, you also need to manipulate ΔH_{rxn} using the following rules:

Rule 1: If you <u>reverse the reactions</u>, then <u>change the sign of ΔH</u>. For example,

$$A + B \rightarrow C \qquad \Delta H = 60\ \text{kJ}$$
$$C \rightarrow A + B \qquad \Delta H = -60\ \text{kJ}$$

Rule 2: If you <u>multiply the reaction by a coefficient</u>, then <u>multiply value of ΔH by same coefficient</u>. For example,

$$A + B \rightarrow C \qquad \Delta H = 60\ \text{kJ}$$
$$2A + 2B \rightarrow 2C \qquad \Delta H = 120\ \text{kJ}$$

Rule 3: Rule 1 and 2 can be combined! For example, if the first reaction is tripled and reversed,

$$A + B \rightarrow C \qquad \Delta H = 60\ \text{kJ}$$
$$3C \rightarrow 3A + 3B \qquad \Delta H = -180\ \text{kJ}$$

<u>Strategy</u>: Find things in your goal equation that appear in only one of the available reactions and make them match by flipping equations or multiplying/dividing coefficients. Then arrange equations to cancel out things that do ________ appear in the "goal." **Whatever you do to the equation, you must do to ________!**

<u>When combining reactions</u>:

1. Reagents on the <u>same side</u>: add together
2. Reagents on the <u>opposite side</u>: subtract (from the side with the greatest amount)

<u>Example</u>: Find the enthalpy change for the formation of $PbCl_4$ by the reaction of lead (II) chloride with chlorine.

$$PbCl_2(s) + Cl_2(g) \rightarrow PbCl_4(l) \qquad \Delta H = ?$$

Use the following thermochemical equations:

$$Pb(s) + 2\ Cl_2(g) \rightarrow PbCl_4(l) \qquad \Delta H = -329.2\ \text{kJ/mol}_{\text{rxn}}$$
$$Pb(s) + Cl_2(g) \rightarrow PbCl_2(s) \qquad \Delta H = -359.4\ \text{kJ/mol}_{\text{rxn}}$$

Guided Practice

1. Given the following information, find the heat of formation for methane: $C(s) + 2\,H_2(g) \rightarrow CH_4(g)$

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H^\circ_{comb} = -891\ \text{kJ mol}^{-1}$$
$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^\circ_f = -394\ \text{kJ mol}^{-1}$$
$$H_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow H_2O(l) \qquad \Delta H^\circ_f = -286\ \text{kJ mol}^{-1}$$

2. Find the enthalpy change for the formation of pentane, C_5H_{12}, by the reaction of carbon with hydrogen.

$$5\,C(s) + 6\,H_2(g) \rightarrow C_5H_{12}(g)$$

Use the following thermochemical equations:

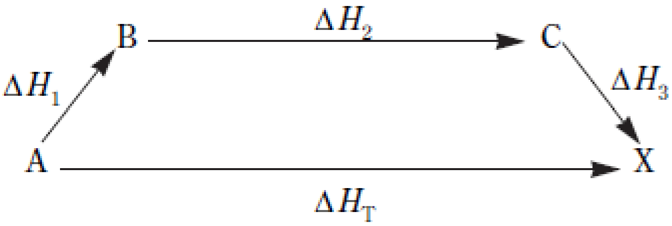
$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ \text{kJ/mol}_{rxn}$$
$$H_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow H_2O(l) \qquad \Delta H = -285.8\ \text{kJ/mol}_{rxn}$$
$$C_5H_{12}(g) + 8\,O_2(g) \rightarrow 5\,CO_2(g) + 6\,H_2O(l) \qquad \Delta H = -3535.6\ \text{kJ/mol}_{rxn}$$

3. Calculate the heat of formation for sulfur dioxide, $SO_2(g)$, from its elements sulfur and oxygen. Use the balanced chemical equation and the following information.

$$S(s) + \frac{3}{2}\,O_2(g) \rightarrow SO_3(g) \qquad \Delta H = -395.2\text{kJ/mol}_{rxn}$$
$$2\,SO_2(g) + O_2(g) \rightarrow 2\,SO_3(g) \qquad \Delta H = -198.2\text{kJ/mol}_{rxn}$$

$$\mathrm{A} \rightarrow \mathrm{X}$$

4. The enthalpy change for the reaction represented above is ΔH_T. This reaction can be broken down into a series of steps as shown in the diagram:



A relationship that must exist among the various enthalpy changes is:

A $\Delta H_T - \Delta H_1 - \Delta H_2 - \Delta H_3 = 0$

B $\Delta H_2 - (\Delta H_3 + \Delta H_1) = \Delta H_T$

C $\Delta H_3 - (\Delta H_1 + \Delta H_2) = \Delta H_T$

D $\Delta H_T + \Delta H_1 + \Delta H_2 + \Delta H_3 = 0$

$$\mathrm{C}(s) + 2\mathrm{H}_2(g) \rightarrow \mathrm{CH}_4(g) \qquad \Delta H^\circ = x$$

$$\mathrm{C}(s) + \mathrm{O}_2(g) \rightarrow \mathrm{CO}_2(g) \qquad \Delta H^\circ = y$$

$$\mathrm{H}_2(g) + \frac{1}{2}\mathrm{O}_2(g) \rightarrow \mathrm{H}_2\mathrm{O}(l) \qquad \Delta H^\circ = z$$

5. Based on the information given above, what is ΔH° for the following reaction?

$$\mathrm{CH}_4(g) + 2\,\mathrm{O}_2(g) \rightarrow \mathrm{CO}_2(g) + 2\,\mathrm{H}_2\mathrm{O}(l)$$

a. x + y + z

b. x + y – z

c. y + z – 2x

d. y + 2z – x

Enthalpy of Neutralization Lab

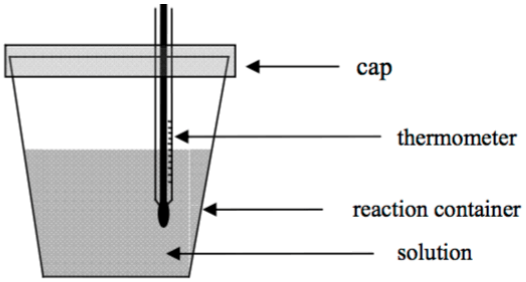
Goal:

In this lab you will be calculating ΔH_{neut} for the reaction of HCl and NaOH.

Chemicals	Equipment
hydrochloric acid, HCl (0.25 M, 0.50 M)	Coffee cup calorimeter
sodium hydroxide, NaOH (0.25 M, 0.50 M)	Temperature probe
	Two 100 mL Graduated cylinders

Procedure

1. Set up the data collection device as indicated by your instructor.
2. Using a graduated cylinder, place 30.0 mL of 0.25 M HCl solution into an insulated cup.
3. Allow your temperature to sit in the acid solution for a few minutes and then record the initial temperature of the 0.25 M HCl solution.
4. Measure 30.0 mL of 0.25 M NaOH into a second graduated cylinder.
5. When your group is ready to start data collection, quickly and carefully pour the NaOH solution into the insulated cup with the acid, cover with a lid, and **<u>stir with the temperature probe</u>** until the temperature no longer changes.



6. Record the highest temperature reached in your data table.
7. Dispose of solution down the drain, rinse the Styrofoam cup, and repeat steps 2 – 6 (for your second trial).
8. Now you're going to run two more trials, but with ONE change. Each table will be assigned a different change! Just like the first experiment, you will be running two trials.

Tables: __________ a. Change the molarity of **ONE** of the two solutions to 0.50 M

Tables: __________ b. Change the molarity of **BOTH** solutions to 0.50 M

Tables: __________ c. Double the volume of **ONE** of the two solutions

Tables: __________ d. Double the volume of **BOTH** solutions

9. When you have completed all four trials (or run out of time), clean up your station: dispose of solution down the drain (flush with water afterwards), rinse out Styrofoam cup, wipe off temperature probe and turn it off (by holding down the power button until the light stays red for a few seconds and then turns off), and return all supplies to the supply box.

Data Table: Enthalpy of Neutralization Lab

	$\mathbf{V_a}$ (mL)	$\mathbf{V_b}$ (mL)	$\mathbf{M_a}$	$\mathbf{M_b}$	$\mathbf{T_i}$ (°C)	$\mathbf{T_f}$ (°C)
Experiment 1, Trial 1			0.25 M	0.25 M		
Experiment 1, Trial 2			0.25 M	0.25 M		
Experiment 2, Trial 1						
Experiment 2, Trial 2						

Calculations Table #1: Data-crunching

	mass (g)	**ΔT** (°C)	$\mathbf{q_{rxn}}$ (J)
Experiment 1, Trial 1			
Experiment 1, Trial 2			
Experiment 2, Trial 1			
Experiment 2, Trial 2			

Calculations Table #2: Enthalpy of Neutralization

	$\mathbf{q_{rxn}}$ (kJ)	$\mathbf{mol_{rxn}}$	**ΔH** (kJ/$\mathrm{mol_{rxn}}$)
Experiment 1, Trial 1			
Experiment 1, Trial 2			
Experiment 2, Trial 1			
Experiment 2, Trial 2			

Calculations: Show the calculation for one trial to determine the enthalpy of neutralization, ΔH_{neut}, in kJ/mol. (Hint: you will need to calculate both q_{rxn} and mol_{rxn} to do so.) Complete the calculations table above for the remaining trials: you do not need to show work. ☺

Analysis Part I: Use your data to answer the following questions.

1. For each term in the equation $q = mC\Delta T$,
    a. identify what it measures or signifies;
    b. identify its unit of measure;
    c. identify if it is a measurement, a calculation, or an accepted value.

2. Based on the variable that your group changed, identify patterns in your data and calculations by answering the following questions:
    a. did the total mass increase, decrease, or remain unchanged? Justify your answer.

    b. did the moles of reaction, mol_{rxn}, increase, decrease, or remain unchanged? Justify your answer.

    c. did heat released, q_{rxn}, increase, decrease, or remain unchanged? Justify your answer.

    d. did change in temperature, ΔT, increase, decrease, or remain unchanged? Justify your answer.

    e. did enthalpy of neutralization, ΔH_{neut}, increase, decrease, or remain unchanged? Justify your answer.

Analysis Part II: Use the **class data set** to answer the following questions.

3. When the molarity of ONLY one of the solutions was changed,
   a. did the total mass increase, decrease, or remain unchanged? Justify your answer.
   b. did the moles of reaction, mol_{rxn}, increase, decrease, or remain unchanged? Justify your answer.
   c. did heat released, q_{rxn}, increase, decrease, or remain unchanged? Justify your answer.
   d. did change in temperature, ΔT, increase, decrease, or remain unchanged? Justify your answer.
   e. did enthalpy of neutralization, ΔH_{neut}, increase, decrease, or remain unchanged? Justify your answer.

4. When the molarity of BOTH of the solutions was changed,
   a. did the total mass increase, decrease, or remain unchanged? Justify your answer.
   b. did the moles of reaction, mol_{rxn}, increase, decrease, or remain unchanged? Justify your answer.
   c. did heat released, q_{rxn}, increase, decrease, or remain unchanged? Justify your answer.
   d. did change in temperature, ΔT, increase, decrease, or remain unchanged? Justify your answer.
   e. did enthalpy of neutralization, ΔH_{neut}, increase, decrease, or remain unchanged? Justify your answer.

5. When the volume of ONLY one of the solutions was doubled,
    a. did the total mass increase, decrease, or remain unchanged? Justify your answer.

    b. did the moles of reaction, mol_{rxn}, increase, decrease, or remain unchanged? Justify your answer.

    c. did heat released, q_{rxn}, increase, decrease, or remain unchanged? Justify your answer.

    d. did change in temperature, ΔT, increase, decrease, or remain unchanged? Justify your answer.

    e. did enthalpy of neutralization, ΔH_{neut}, increase, decrease, or remain unchanged? Justify your answer.

6. When the volume of BOTH of the solutions was doubled,
    a. did the total mass increase, decrease, or remain unchanged? Justify your answer.

    b. did the moles of reaction, mol_{rxn}, increase, decrease, or remain unchanged? Justify your answer.

    c. did heat released, q_{rxn}, increase, decrease, or remain unchanged? Justify your answer.

    d. did change in temperature, ΔT, increase, decrease, or remain unchanged? Justify your answer.

    e. did enthalpy of neutralization, ΔH_{neut}, increase, decrease, or remain unchanged? Justify your answer.

Enthalpy of Formation (ΔH_f°)

Calculating Enthalpy Changes of Reactions from Heat of Formation

Standard enthalpy (heat) of formation (______) : change in enthalpy that accompanies the formation of _______ mole of the compound in its standard state from its component ____________________ their standard states.

Note: The naught, ____, simply indicates standard conditions (**1 atm** and **25°C**).

Examples

→ $CH_4(g)$ ΔH_f° = -74.9 kJ/mol_{rxn}

→ $SO_3(g)$ ΔH_f° = -396 kJ/mol_{rxn}

Note: you will see ___________________ coefficients to ensure only ____ mole of compound is formed.

The ΔH_f° for _______________ (in their standard state) **is always _____ kJ/mol_{rxn}!**

Now you try! Write the formation reaction of liquid NH_3:

The ___________________ change for a chemical reaction, or amount of heat released or absorbed, can be determined by the following formula, which is known as Big Momma's Equation:

$$\Delta H_{rxn}^\circ = \Sigma\left[n\Delta H_f^\circ(products)\right] - \Sigma\left[n\Delta H_f^\circ(reactants)\right]$$

Example:

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

Enthalpy Changes of Different Types of Reactions

You will encounter a variety of _______________________ following the ΔH, however, they are simply indicating a ___________________ type of reaction or change of state.

Examples

ΔH_{comb}° = Enthalpy of Combustion (Heat Energy Released during Combustion Reactions)

ΔH_{neut}° = Enthalpy of Neutralization (Heat Energy Released during Acid-Base Neutralization Reactions)

ΔH_{soln}° = Enthalpy of Solution (Heat Energy Released/Absorbed Dissolving a Solute in Water)

ΔH_{vap}° = Enthalpy of Vaporization (Heat Energy Absorbed to Convert from Liquid to Gas Phase)

ΔH_{fus}° = Enthalpy of Fusion (Heat Energy Absorbed to Convert from Solid to Liquid Phase)

Let's Practice!

1. For each reaction below, fill in the box with a subscript that specifies the type of reaction:
   a. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$ ΔH ☐ = −55.8 kJ/mol
   b. $H_2O(s) \rightarrow H_2O(l)$ ΔH ☐ = +6.02 kJ/mol
   c. $CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$ ΔH ☐ = −80.52 kJ/mol

2. Circle each of the following which will have a standard heat of formation equal to zero.

   $H_2(g)$, $Hg(s)$, $CO_2(g)$, $H_2O(l)$, $Br_2(l)$, $N_2(l)$

3. Write a balanced molecular equation representing the enthalpy for standard heat of formation reaction of ethanol, $C_2H_5OH(l)$.

4. Use the information provided and the balanced equation to determine ΔH_f° of carbon tetrachloride, CCl_4.

   $CH_4(g) + 4\,Cl_2(g) \rightarrow CCl_4(g) + 4\,HCl(g)$ $\Delta H^\circ_{rxn} = -389\ \text{kJ/mol}_{rxn}$

Substance	**ΔH°_f**
$CH_4(g)$	$-75\ \text{kJ/mol}_{rxn}$
$HCl(g)$	$-92\ \text{kJ/mol}_{rxn}$

5. Use the data regarding the standard enthalpies of formation to calculate $\Delta H°_{comb}$ for the following reaction:

$$2\ C_3H_6(g) + 9\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

Substance	**$\Delta H°_f$(kJ/mol)**
$C_3H_6(g)$	20.9 kJ/mol
$CO_2(g)$	−393.5 kJ/mol
$H_2O(l)$	−286 kJ/mol

6. Describe in words what process you would follow to calculate the enthalpy of combustion for butane, C_4H_{10}. What information would you need? What would you need to do before you could complete the calculation?

$$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$$

7. Gaseous hydrogen and chlorine combine in the reaction above to form hydrogen chloride with an enthalpy change of −460 kJ/mol. What is the value of the heat of formation of HCl(g)?

   a. −9,20 kJ/mol
   b. −230 kJ/mol
   c. +230 kJ/mol
   d. +460 kJ/mol

8. If the standard enthalpies of formation of HBr(g) and $Br_2(g)$ are $-36\ kJ\ mol^{-1}$ and $+31\ kJ\ mol^{-1}$ (at 298 K) respectively, what is $\Delta H°_{rxn}$ for the following reaction?

$$H_2(g) + Br_2(g) \rightarrow 2\ HBr(g)$$

   a. $-103\ kJ\ mol^{-1}$
   b. $-67\ kJ\ mol^{-1}$
   c. $+67\ kJ\ mol^{-1}$
   d. $+103\ kJ\ mol^{-1}$

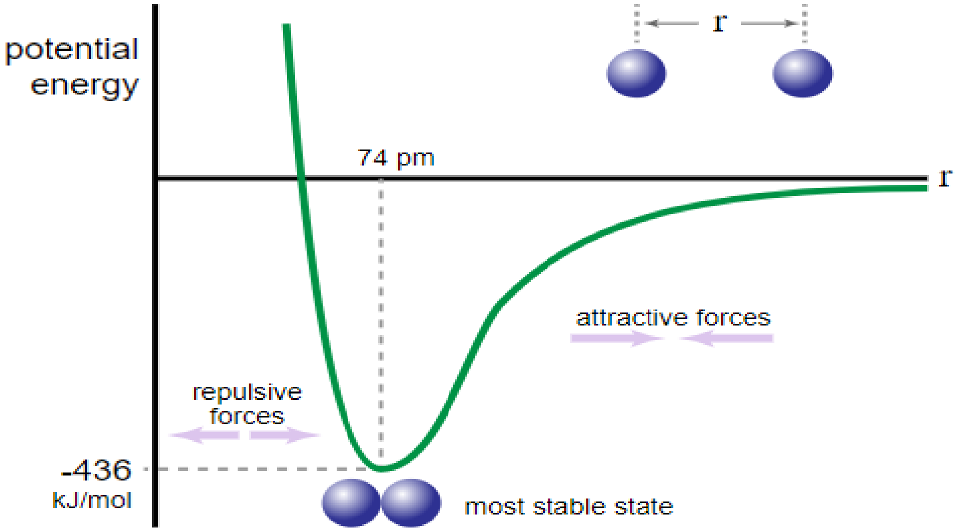
Bond Energy and Enthalpy

A chemical bond forms when a system of bonded atoms is lower in potential energy than that of independent atoms. Since bond-making is an electrostatic process, it is governed by Coulomb's Law.

Coulomb's Law tells us:

- The negative electrons of one atom and the positive nucleus of another atom _____________ each other.
- If the nuclei of two atoms get too close together, their like charges ___________ each other.

The optimum distance between two atoms is the ________ ______________, which represents the lowest energy state. The bond length is a balance between the attractive electrostatic forces between the nucleus of one atom and the electrons of another, and the repulsive forces between the positively charged nuclei and the negatively charged electrons.



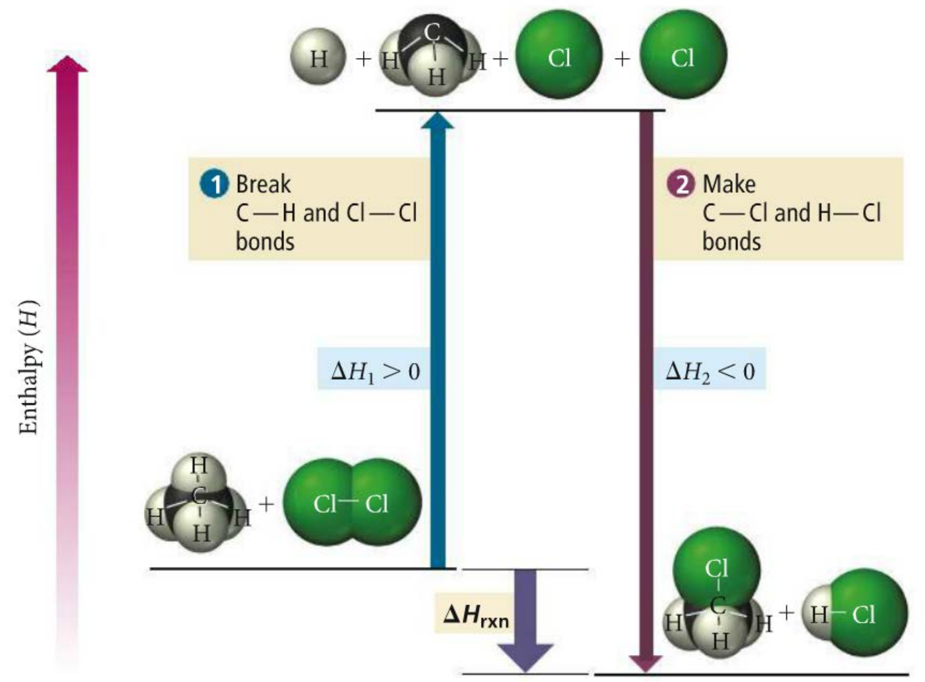
Breaking chemical bonds requires an input of energy to overcome the attractive forces. When new bonds are formed, energy is released. The difference between input and output determines whether a process is endothermic or exothermic.

Covalent Bond Energies and Enthalpy

- The enthalpy change for a reaction, ΔH_{rxn}, can also be understood in terms of bonds ___________________ (endothermic) and bond ___________________ (exothermic) during a chemical reaction.
- The total enthalpy change can be negative (exothermic) or positive (endothermic) depending on the relative magnitude of two (breaking and making) processes.

Endothermic: **+**ΔH^o (system gains energy)	**Exo**thermic: **−**ΔH^o (system loses energy)
➢ Breaking a chemical bond is always endothermic ➢ Hint: "**end**-ing a bond is **end**othermic"	➢ Forming a chemical bond is always exothermic

Estimating the Enthalpy Change of a Reaction from Bond Energies



The following expression can be written to express this relationship:

$$\Delta H^\circ_{rxn} = \Sigma H^\circ_{(bonds\ broken)} - \Sigma H^\circ_{(bonds\ formed)}$$

MEMORIZE! OR **MEMORIZE!**

$$\Delta H^\circ_{rxn} = \Sigma \mathrm{BE}_{(reactants)} - \Sigma \mathrm{BE}_{(products)}$$

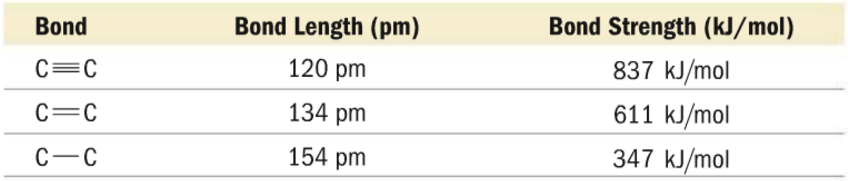
IMPORTANT: Draw the __________________ structures of the compounds involved in the reaction to determine the bonds being broken or formed.

But how will I know what bond energies to use?

You will be provided with the bond energies in the problem or be asked to look at a ______________.

Bond Energy Trends:

- Double and triple bonds are _______________ than single bonds, since double and triple bonds contain a greater number of ______________ than single bonds, so Coulombic attractions to the nuclei are stronger!
- Shorter bonds (with the atoms closer to each other) tend to be stronger than longer bonds with the atoms further apart, since the ____________ between the charges is smaller.



Bond	Bond Length (pm)	Bond Strength (kJ/mol)
C≡C	120 pm	837 kJ/mol
C=C	134 pm	611 kJ/mol
C—C	154 pm	347 kJ/mol

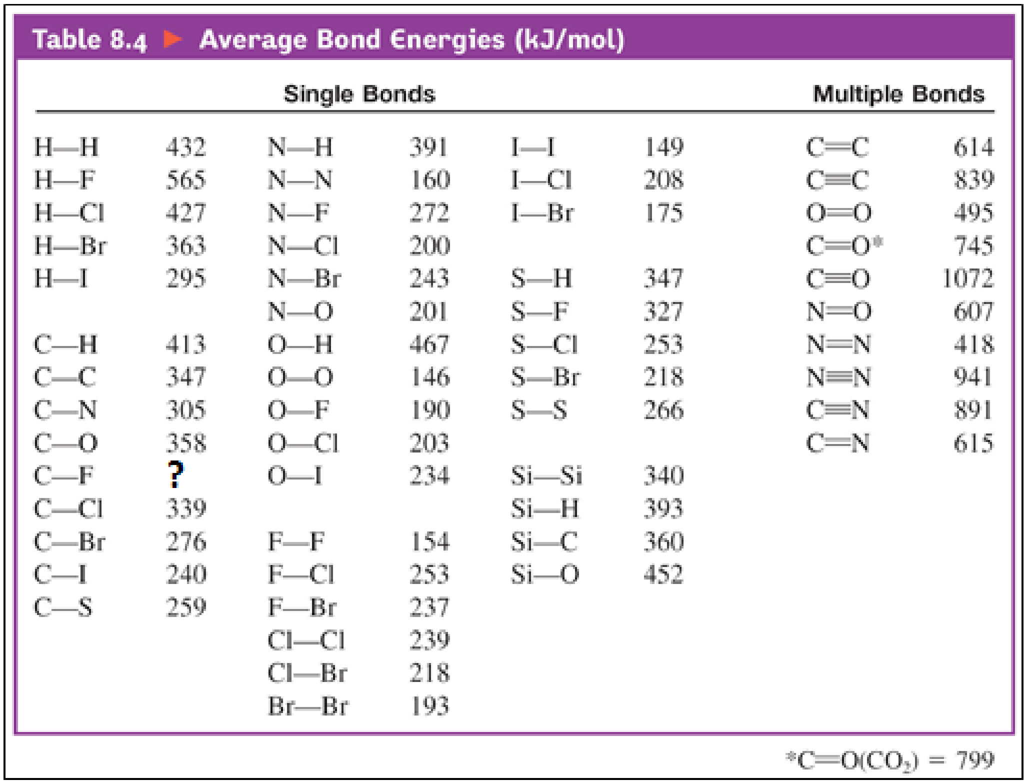


Table 8.4 ▶ Average Bond Energies (kJ/mol)

Single Bonds								Multiple Bonds	
H—H	432	N—H	391	I—I	149			C=C	614
H—F	565	N—N	160	I—Cl	208			C≡C	839
H—Cl	427	N—F	272	I—Br	175			O=O	495
H—Br	363	N—Cl	200					C=O*	745
H—I	295	N—Br	243	S—H	347			C≡O	1072
		N—O	201	S—F	327			N=O	607
C—H	413	O—H	467	S—Cl	253			N=N	418
C—C	347	O—O	146	S—Br	218			N≡N	941
C—N	305	O—F	190	S—S	266			C≡N	891
C—O	358	O—Cl	203					C=N	615
C—F	?	O—I	234	Si—Si	340				
C—Cl	339			Si—H	393				
C—Br	276	F—F	154	Si—C	360				
C—I	240	F—Cl	253	Si—O	452				
C—S	259	F—Br	237						
		Cl—Cl	239						
		Cl—Br	218						
		Br—Br	193						

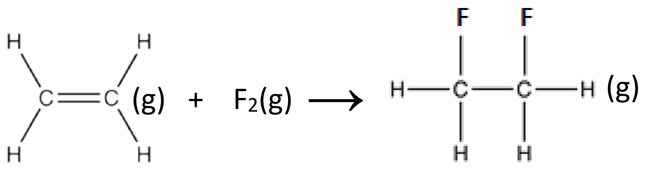
*C=O(CO_2) = 799

Let's Practice!

1. Using the bond energies from the reference chart above, estimate ΔH_{comb} for the following reaction. Does the reaction illustrate an endothermic or exothermic process?

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

2. **Estimate the carbon-fluorine bond energy**, given the remaining bond energies provided in the reference chart and the information provided in the equation below. Does the reaction illustrate an endothermic or exothermic process?



$\Delta H_{rxn} = -549$ kJ/mol

$$I_2(g) + 3\ Cl_2(g) \rightarrow 2\ ICl_3(g)$$

3. According to the data in the table below, what is the value of ΔH° for the reaction represented above? (Assume ICl_3 has three single I – Cl bonds).

Bond	Average Bond Energy (kJ/mol)
I—I	150
Cl—Cl	240
I—Cl	210

a. −870 kJ/mol b. −390 kJ/mol c. +180 kJ/mol d. +450 kJ/mol

Thermochemical Equations for Phase Changes

Thermochemical equations can also be written for _______________ changes!

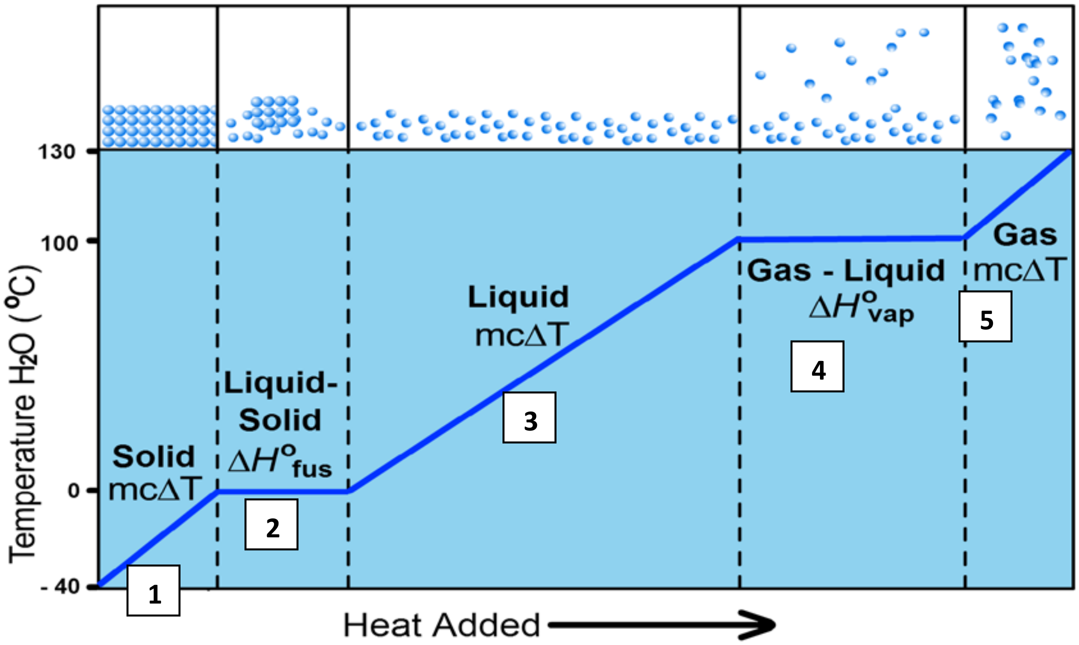
$$H_2O(s) \rightarrow H_2O(l) \qquad \Delta H = +6.02\ \text{kJ/mol}$$

$$H_2O(l) \rightarrow H_2O(s) \qquad \Delta H = -6.02\ \text{kJ/mol}$$

Physical changes (____________ changes, or changes of _________) are either endo- or exothermic:

- Endothermic physical changes
  1. Require an addition of heat/energy to overcome _______ (intermolecular forces) and change state (just like ending a bond, "end"ing an IMF is always "end"othermic!)
     - Solid → Liquid (Melting)
     - Liquid → Gas (Vaporization)
     - Solid → Gas (Sublimation)
- Exothermic physical changes
  1. Heat/energy is released as IMFs form
     - Gas → Liquid (Condensing)
     - Liquid → Solid (Freezing)
     - Gas → Solid (Deposition)

Heating Curve of Water



When heating a solid, liquid, or gas (1, 3, and 5):

- KE / Temp increases
- PE is constant
- Constant state

When changing phase in the forward direction (2 and 4):

- KE / Temp is constant
- PE increases
- Multiple states

Energy Changes

1. _____________ **(slanted) sections** (1, 3, and 5): Temperature (thus ____) changing
   - ➔ Use mCΔT
2. **Flat** (plateau) **sections** (2 and 4): 3 Ps: Plateau, Phase change, and Potential Energy Change
   - ➔ ___ change in temperature = KE constant
   - ➔ Use $q = n\,\Delta H$

How to Calculate Energy Needed to Change Temperature or Phase

1. **Temperature changes**: $q = m\, C\, \Delta T$

   This **starts with grams** and gives **heat in Joules**

   Note: is your specific heat a *molar* heat capacity? (C_m = molar heat capacity = J/**mol** K)

   ➔ Use $q = n\, C_m\, \Delta T$

2. **Phase changes**: TWO Options!
   - Melting $q = n\, \Delta H_{fus}$
   - Freezing $q = -n\, \Delta H_{fus}$

   n = moles of substance

   ΔH_{fus} = enthalpy of fusion

   - Vaporizing $q = n\, \Delta H_{vap}$
   - Condensing $q = -n\, \Delta H_{vap}$

   n = moles of substance

   ΔH_{vap} = enthalpy of vaporization

   These **start with moles** and give **heat in kilojoules**

3. Once you know the heat at each step of the process, add them together to find the heat of the process, but **make sure you are adding the same units of energy!** (usually kJ +kJ)

Hint: How to determine if it's a one-step or multi-step calculation.

1. Are you changing JUST state of matter or JUST temperature? **1-step calculation**
2. Are you changing state of matter AND temperature? **Multi-step calculation**

Examples: Determine each of the following changes.

1. Calculate the amount of heat gained (+) or lost (−) when melting 55.8 g of Ti at 1677°C. $\Delta H_{fus}(Ti)$ = +18.8 kJ/mol

2. Calculate the amount of heat gained (+) or lost (−) when converting 45.0 g of water at 20.0°C to steam at 115°C. (ΔH_{fus} = +6.02 kJ/mol, ΔH_{vap} = +40.7 kJ/mol, C_{solid} = 2.06 J/g °C, C_{liquid} = 4.18 J/g °C, and C_{gas} = 2.02 J/g °C)

3. Calculate the energy absorbed when heating 6.9 g of solid aluminum from 32°C to 320°C. The melting point of aluminum is 660°C, ΔH_{fus} = +10.8 kJ/mol, and the specific heat capacity of Al(s) is 0.903 J/mol °C.

4. Calculate the energy absorbed when melting 27.3 g of Al at 660°C. The melting point of aluminum is 660°C, ΔH_{fus} = 10.8 kJ/mol, and the specific heat capacity of Al(s) is 0.903 J/mol °C.

5. When water vaporizes at its normal boiling point, its ΔH_{vap} = +40.7 kJ/mol. A sample of water vapor was condensed at 25°C and 1.05 atm. Calculate the volume of water that was condensed, in liters, if 620. kJ of energy was released during the process. (ΔH_{vap} = +40.7 kJ/mol, C_{liquid} = 4.18 J/g °C, and C_{gas} = 2.02 J/g °C).

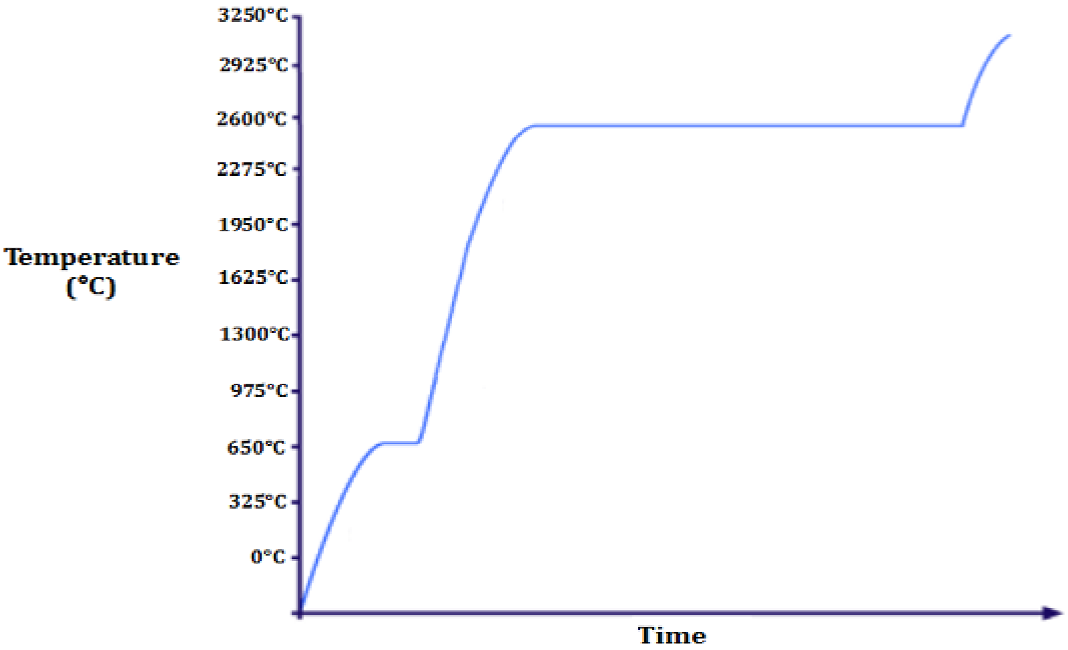
Part II: Determine the amount of heat gained (+) or lost (-) during each of the following changes. (For water, ΔH_{fus} = +6.02 kJ/mol, ΔH_{vap} = +40.7 kJ/mol, C_{solid} = 2.06 J/g °C, C_{liquid} = 4.18 J/g °C, and C_{gas} = 2.02 J/g °C).

6. 220.0 g of ice at −35.0°C is converted to liquid water at 50.0°C.

7. 5.00 g of steam at 155°C is converted to liquid water at 100.0°C

Multiple Choice Practice FTW!

8. Aluminum metal can be recycled from scrap metal by melting the metal to evaporate impurities.



Calculate the amount of heat needed to purify 1.00 mol of Al originally at 298 K by melting it. The melting point of Al is 650°C. The molar heat capacity of Al is 24.0 J/mol K, and the heat of fusion of Al is 10.7 kJ/mol.

a. 15,300 kJ b. 30,600 kJ c. 10.7 kJ d. 25.9 kJ

9. Benzene, C_6H_6, has a melting point of 5.5°C and a boiling point of 80.1°C. Given the other constants below for benzene, which of the following processes would require the least amount of energy?

$$c_{solid} = 118.4\ \mathrm{J \cdot mol^{-1} \cdot K^{-1}}$$

$$c_{liquid} = 134.8\ \mathrm{J \cdot mol^{-1} \cdot K^{-1}}$$

$$c_{vapor} = 82.44\ \mathrm{J \cdot mol^{-1} \cdot K^{-1}}$$

$$\Delta H_{vap} = 30.77\ \mathrm{kJ \cdot mol^{-1}}$$

$$\Delta H_{fus} = 9.9\ \mathrm{kJ \cdot mol^{-1}}$$

a. Vaporizing 2.0 mol of benzene at 80.1°C.

b. Warming 2.0 mol of benzene from 91°C to 101°C.

c. Warming 2.0 mol of benzene from −10°C to 0°C.

d. Melting 2.0 mol of benzene at 5.5°C.

Study Guide or: How I Learned to Stop Worrying and Love Thermochem 💙

First: determine what you're being asked to find: **ΔH_{rxn} or something else?**

1. **Asked to find ΔH_{rxn}?** 3 Options!

Given?	Asked to find?	Use:
1. A lot of bond energies 2. A reaction **without** ΔH_{rxn}	ΔH_{rxn}	1. Draw the Lewis structures. 2. Use the following: $\Delta H_{rxn} = \Sigma(BE_{reactants}) - \Sigma(BE_{products})$
1. A lot of heats of formation (ΔH_f) 2. A reaction **without** ΔH_{rxn}	ΔH_{rxn}	$\Delta H_{rxn} = \Sigma\Delta H_f\,(products) - \Sigma\Delta H_f\,(reactants)$
1. Multiple reactions **with** ΔH 2. A goal reaction **without** ΔH_{rxn}	ΔH_{rxn}	**Hess's Law!** Rearrange the equations to make the goal equation, then combine your new ΔH's (remember, what you do to an equation you must do to ΔH!)
1. info to calculate q (using $mC\Delta T$ or $nC\Delta T$) 2. Moles/grams of a chemical	ΔH_{rxn}	$\Delta H_{rxn} = \frac{q}{mol_{rxn}}$

2. **Asked to find something else?** 3 Options!

Given?	Asked to find?	Use:
1. A reaction **with** ΔH_{rxn} 2. **Either**: a. g or mol of a substance b. energy change (J or kJ)	1. **Either**: a. g or mol of a substance b. energy change (J or kJ)	**Stoich!** Don't forget to convert between moles of your substance and $moles_{rxn}$
1. A phase change (vaporizing, condensing, freezing or melting) 2. ΔH_{vap} or ΔH_{fus}	Energy change (heat absorbed or released)	$q = n\Delta H$
1. A temperature change 2. Mass or moles of a substance 3. Heat capacity ($\frac{J}{g\,°C}$ or $\frac{J}{mol\,°C}$)	Energy change (heat absorbed or released)	$q = mC\Delta T$

Unit 4 Multiple Choice Practice

1. Which of the following is true for an endothermic reaction?
    - a. The activation energy is always greater than the activation energy for an exothermic reaction.
    - b. Energy is released over the course of the reaction.
    - c. The strength of the bonds in the reactants exceeds the strength of the bonds in the products.
    - d. The products have more thermal energy than the reactants.

2. If the standard enthalpy of formation for HBr(g) is −35 kJ mol^{-1}, what is $\Delta H°_f$ for $Br_2(g)$?

$$\mathrm{H_2}(g) + \mathrm{Br_2}(g) \rightarrow 2\ \mathrm{HBr}(g) \qquad \Delta H_{rxn} = -110\ \mathrm{kJ\ mol^{-1}}$$

a. 0 kJ mol^{-1} b. −40. kJ mol^{-1} c. +40. kJ mol^{-1} d. +75 kJ mol^{-1}

$$\mathrm{MgO}(s) \rightarrow \mathrm{Mg^{2+}}(aq) + \mathrm{O^{2-}}(aq)$$

3. Is the process above endothermic or exothermic? Why?
    - a. Exothermic, because one solid particle becomes two aqueous particles.
    - b. Endothermic, because one solid particle becomes two aqueous particles.
    - c. Exothermic, because overcoming the Coulombic forces within MgO(s) releases energy.
    - d. Endothermic, because overcoming the Coulombic forces within MgO(s) requires the input of energy.

4. Consider the following reaction showing photosynthesis:

$$6\ \mathrm{CO_2}(g) + 6\ \mathrm{H_2O}(l) \rightarrow \mathrm{C_6H_{12}O_6}(s) + 6\ \mathrm{O_2}(g) \qquad \Delta H = +2800\ \mathrm{kJ/mol}$$

How much thermal energy is absorbed or released when 2.0 moles of glucose, $C_6H_{12}O_6(s)$, is metabolized in the presence of excess oxygen gas, releasing CO_2 and H_2O?

a. 2800 kJ absorbed b. 5600 kJ absorbed c. 2800 kJ released d. 5600 kJ released

Use the following information to answer the next three questions.

$$3\ \mathrm{H_2}(g) + \mathrm{N_2}(g) \rightarrow 2\ \mathrm{NH_3}(g) \quad \Delta H = -81.0\ \mathrm{kJ/mol}$$

5. Gaseous hydrogen and nitrogen combine in the reaction above to form ammonia, NH_3. What is the value of the heat of formation of NH_3(g)?

   a. −40.5 kJ/mol b. −81.0 kJ/mol c. +40.5 kJ/mol d. +81.0 kJ/mol

6. If 28.0 g of N_2(g) is combined with 28.0 g of H_2(g) in a sealed flask and the reaction is allowed to go to completion, what chemical species will still be present inside of the flask?

   a. NH_3(g) only c. NH_3(g) and H_2(g)
   b. NH_3(g) and N_2(g) d. NH_3(g), N_2(g) and H_2(g)

7. If 28.0 g of N_2(g) is combined with 28.0 g of H_2(g) in a sealed flask and the reaction is allowed to go to completion, how much heat is absorbed or released?

   a. 81.0 kJ absorbed b. 375 kJ absorbed c. 81.0 kJ released d. 375 kJ released

$$\mathrm{H_2O}(l) \rightarrow \mathrm{H_2O}(s)$$

8. Is the process above endothermic or exothermic? Why?

   a. Exothermic, because forming new intermolecular attractions between water molecules releases energy.
   b. Endothermic, because energy is needed to form new intermolecular attractions between water molecules.
   c. Exothermic, because forming new bonds between water molecules releases energy.
   d. Endothermic, because energy is needed to form new bonds between water molecules.

Bond	Average Bond Dissociation Energy (kJ/mol)
C − H	415
O = O	495
C = O	798
O − H	463

$$\mathrm{CH_4}(g) + 2\ \mathrm{O_2}(g) \rightarrow \mathrm{CO_2}(g) + 2\ \mathrm{H_2O}(g)$$

9. Use the information provided above to calculate the molar enthalpy of combustion, ΔH°_{comb}, for methane gas.

a. −798 kJ mol^{-1} b. −1226 kJ mol^{-1} c. +798 kJ mol^{-1} d. +1226 kJ mol^{-1}

$$2\ \mathrm{H_2O}(l) \rightarrow 2\ \mathrm{H_2}(g) + \mathrm{O_2}(g)$$

10. Based on the information in the table below, calculate ΔH°_{rxn} for the reaction shown above.

Bond	Average Bond Dissociation Energy (kJ/mol)
H − H	430
O = O	500
O − H	460

a. −900 kJ mol^{-1} b. −480 kJ mol^{-1} c. +480 kJ mol^{-1} d. +900 kJ mol^{-1}

$$\mathrm{F}(g) + e^- \rightarrow \mathrm{F}^-(g)$$

11. Is the process above endothermic or exothermic? Why?

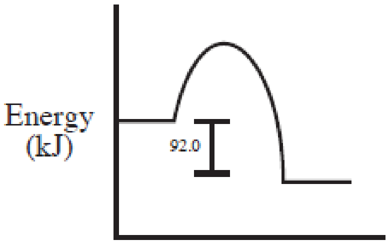
a. Exothermic, because of the attraction between F(g) and an electron.
b. Endothermic, because of the attraction between F(g) and an electron.
c. Exothermic, because the process decreases the total number of particles from two to one.
d. Endothermic, because the process decreases the total number of particles from two to one.

12. When ammonium nitrate (NH_4NO_3) dissolves in water, the temperature of the water decreases dramatically. During this reaction, energy transfers from:

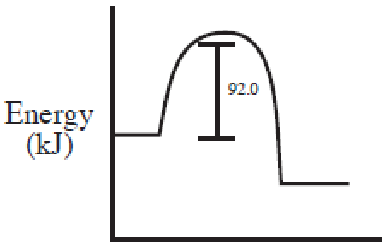
   a. the reactants to the products.
   b. the reactants to the surroundings.
   c. the surroundings to the system.
   d. the products to the surroundings.

$$\mathrm{CS_2}(l) \rightarrow \mathrm{C}(s) + 2\ \mathrm{S}(s) \qquad \Delta H = -92.0\ \mathrm{kJ/mol}$$

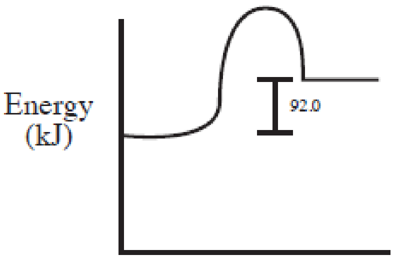
13. Which of the following energy level diagrams gives an accurate representation of the above reaction?



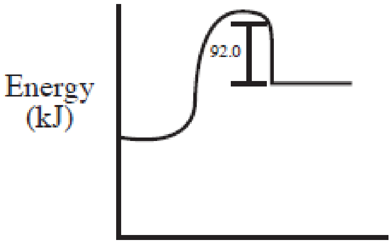
(A)



(B)



(C)



(D)

$$\mathrm{CS_2}(l) \rightarrow \mathrm{C}(s) + 2\ \mathrm{S}(s) \qquad \Delta H = -92.0\ \mathrm{kJ/mol}$$

14. Regarding the reaction above, how much heat is absorbed or released when 1.0 mol of solid carbon reacts with 1.0 mol of solid sulfur to produce CS_2(l)?

   a. 46 kJ of heat is released.
   b. 46 kJ of heat is absorbed.
   c. 92 kJ of heat is released.
   d. 92 kJ of heat is absorbed.

15.

16. The enthalpy change for which of the following reactions would be equal to the enthalpy of formation for ethanol (CH_3CH_2OH)?

    a. $CH_3 + CH_2 + OH \rightarrow CH_3CH_2OH$

    b. $4\ C + 6\ H_2 + O_2 \rightarrow 2\ CH_3CH_2OH$

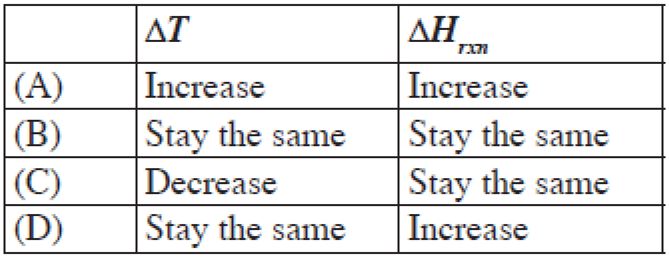
    c. $2\ C + 6\ H + O \rightarrow CH_3CH_2OH$

    d. $2\ C + 3\ H_2 + \frac{1}{2}\ O_2 \rightarrow CH_3CH_2OH$

17. Inside a calorimeter, 100. mL of 1.0 M formic acid (HCOOH), a weak acid, and 100. mL of 2.0 M sodium hydroxide, NaOH, are mixed. The specific heat of the mixture is approximately 4.2 J/g°C, and the density is identical to that of water. If the temperature of the combined solution rises 8.0°C during the course of the reaction, calculate the approximate amount of heat released.

    a. 0.84 kJ b. 3.4 kJ c. 6.7 kJ d. 11.8 kJ

18. Consider the previous question. If the experiment is repeated with 150. mL of 1.0 M HCOOH and 50.0 mL of 2.0 M NaOH, what would happen to the values for ΔT and ΔH_{rxn}?



	ΔT	ΔH_{rxn}
(A)	Increase	Increase
(B)	Stay the same	Stay the same
(C)	Decrease	Stay the same
(D)	Stay the same	Increase

Unit 4: AP Free Response Practice #1 [2018 #1, shortened, 7 points]

$$\mathrm{Na_2S_2O_3}(aq) + 4\ \mathrm{NaOCl}(aq) + 2\ \mathrm{NaOH}(aq) \rightarrow 2\ \mathrm{Na_2SO_4}(aq) + 4\ \mathrm{NaCl}(aq) + \mathrm{H_2O}(l)$$

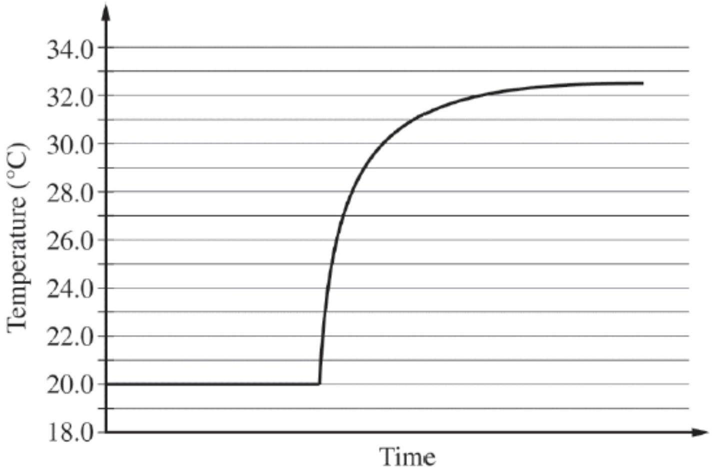
1. A student performs an experiment to determine the value of the ethalpy change, ΔH°_{rxn}, for the oxidation-reduction reaction represented by the balanced equation above.

In the experiment, the student uses the solutions shown in the table below.

Solution	Concentration (*M*)	Volume (mL)
$Na_2S_2O_3(aq)$	0.500	5.00
$NaOCl(aq)$	0.500	5.00
$NaOH(aq)$	0.500	5.00

a. Using the balanced equation for the oxidation-reduction reaction and the information in the table above, determine which reactant is the limiting reactant. Justify your answer. [*1 point*]

The solutions, all originally at 20.0°C, are combined in an insulated calorimeter. The temperature of the reaction is monitored, as shown in the graph below.



b. According to the graph, what is the temperature change of the reaction mixture? [*1 point*]

c. The mass of the reaction mixture inside the calorimeter is 15.21 g.

   i. Calculate the magnitude of the heat energy, in joules, that is released during the reaction. Assume that the specific heat of the reaction mixure is 3.94 J/(g °C) and that the heat absorbed by the calorimeter is negligible. [*1 point*]

   ii. Using the balanced equation for the oxidation-reduction reaction and your answer to part (a), calculate the value of the enthalpy change of the reaction, ΔH°_{rxn}, in kJ/mol_{rxn}. Include the appropriate algebraic sign with your answer. [*2 points*]

The student repeats the experiment, but this time doubling the volume of each of the reactants, as shown in the table below.

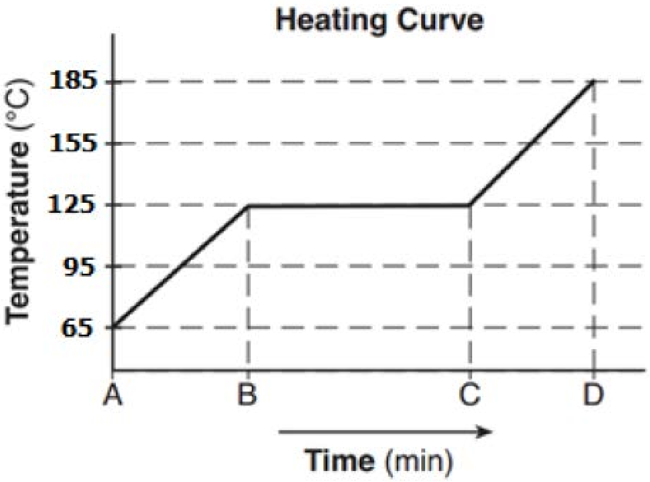
Solution	Concentration (*M*)	Volume (mL)
$Na_2S_2O_3(aq)$	0.500	10.0
$NaOCl(aq)$	0.500	10.0
$NaOH(aq)$	0.500	10.0

d. The magnitude of the enthalpy change of the reaction, ΔH°_{rxn}, in kJ/mol_{rxn}, calculated from the results of the second experiment is the same as the result calculated in part (c)(ii). Explain this result. [*1 point*]

e. Write the balanced net-ionic equation for the given reaction. [*1 point*]

Unit 4: AP Free Response Practice #2 [10 points]

1. Octane, C_8H_{18}, is a hydrocarbon and an alkane that is a common component in gasoline. A 1.00 mole sample of octane is a liquid at 65°C. The sample is heated uniformly to 185°C. The heating curve for the sample at standard pressure is shown below.



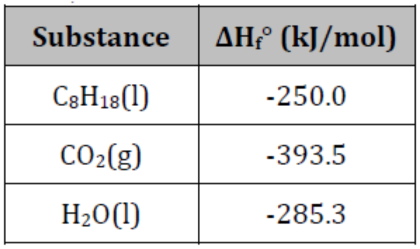
   a. Determine the boiling point of octane at standard pressure. (1 point)

   b. Calculate the amount of heat needed to completely vaporize 1.00 mol of the sample of octane originally at 65°C. The molar heat capacity of the substance in the liquid phase is 255 J/(mol °C), and the heat of vaporization of the substance is 41.0 kJ/mol. (2 points)

A second 1.00 mol sample of liquid octane is combusted with excess oxygen gas in a bomb calorimeter.

   c. Write a balanced equation for the complete combustion of liquid octane, which yields $CO_2(g)$ and $H_2O(l)$. (1 point)

d. Using the heat of formation data in the table below, calculate the ΔH° for the reaction in part (c). (2 points)



Substance	**ΔH_f° (kJ/mol)**
$C_8H_{18}(l)$	-250.0
$CO_2(g)$	-393.5
$H_2O(l)$	-285.3

e. Is the amount of heat required to completely vaporize 1.00 mol of liquid octane originally at 65°C greater than, less than, or equal to the amount of heat released in the combustion of 1.00 mole of liquid octane? Justify your answer. (2 points)

f. All the heat evolved in the combustion of 1.00 mole of liquid octane is transferred to a sample of liquid water. The temperature of the water increases by 72°C. What was the mass of the sample of water in kilograms? The specific heat capacity of water is 4.18 J/g°C. (2 points)

Unit 4: AP Free Response Practice #3 [2013 FR #3, 10 points]

$$\mathrm{MgO}(s) + 2\,\mathrm{H}^+(aq) \rightarrow \mathrm{Mg}^{2+}(aq) + \mathrm{H_2O}(l)$$

3. A student was assigned to the task of determining the enthalpy change for the reaction between solid MgO and aqeous HCl represented by the net ionic equation above. Ther student uses a polystyrene cup calorimeter and performs four trials. Data for each trial are shown in the table below.

Trial	Volume of 1.0 *M* HCl (mL)	Mass of MgO(*s*) Added (g)	Initial Temperature of Solution (°C)	Final Temperature of Solution (°C)
1	100.0	0.25	25.5	26.5
2	100.0	0.50	25.0	29.1
3	100.0	0.25	26.0	28.1
4	100.0	0.50	24.1	28.1

a. Which is the limiting reactant in all four trials, HCl or MgO? Justify your answer. [*1 point*]

b. The data in one of the trials is inconsistent with the data in the other three trials. Identify the trial with the inconsistent data and draw a line through the data from that trial in the table above. Explain how you identified the inconsistent data. [*1 point*]

For parts (c) and (d), use the data from one of the other three trials (i.e., not from the trial you identified in part (b) above.) Assume the calorimeter has a negligible heat capacity and that the specific heat of the contents of the calorimeter is 4.18 J/(g °C). Assume that the density of the HCl(aq) is 1.0 g/mL.

c. Calculate the magnitude of q, the thermal energy change, when the MgO was added to the 1.0 M HCl(aq). Include units with your answer. [*2 points*]

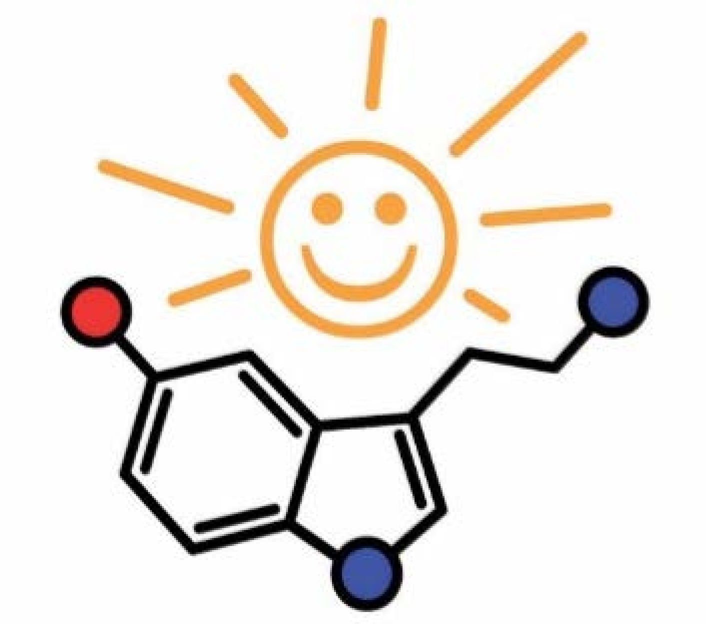
d. Determine the student's experimental value of ΔH° for the reaction between MgO and HCl in units of kJ/mol$_{rxn}$. [*2 points*]

e. Enthalpies of formation for substances involved in the reaction are shown in table below. Using the information in the table, determine the accepted value of ΔH° for the reaction between MgO(s) and HCl(aq). [*2 points*]

Substance	ΔH_f° (kJ/mol)
MgO(*s*)	−602
$H_2O(l)$	−286
$H^+(aq)$	0
$Mg^{2+}(aq)$	−467

f. The accepted value and the experimental value do not agree. If the calorimeter leaked heat energy to the environment, would it help account for the discrepancy between the values? Explain. [*1 point*]

AP Chemistry FTW!



Unit 5:

Thermodynamics

AP Chemistry Unit 5 Objectives

BIG IDEA 5 - The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.

- Enduring Understanding 5.E: Chemical or physical processes are driven by a decrease in enthalpy or an increase in entropy, or both.

BIG IDEA 6 - Any bond or intermolecular attraction that can be formed can be broken. These processes are in a dynamic competition, sensitive to initial conditions and external perturbations.

- Enduring Understanding 6.A: Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal.
- Enduring Understanding 6.D: The equilibrium constant is related to temperature and the difference in Gibbs free energy between reactants and products.

Entropy: Let the chaos begin!

Entropy (S): the extent of randomness or ________________ in a chemical or physical system. (ΔS = change in entropy)

- The second law of thermodynamics states that entropy of the universe will _________________ over time.
    - o + ΔS implies ____________________ entropy and − ΔS implies ____________________ entropy
    - o Nature tends towards_______________________________!
- The more ________________ available (the more space, or more places something can move) the greater the entropy.

The entropy change for a system (Δ) is calculated from the absolute entropies of the products and reactants.

$$\Delta S^{\circ}_{system} = \sum[S^{\circ}(\text{products})] - \sum[S^{\circ}(\text{reactants})]$$

Units of Entropy: $\frac{\mathrm{J}}{\mathrm{mol \cdot K}} = \mathrm{J \cdot mol^{-1} \cdot K^{-1}}$

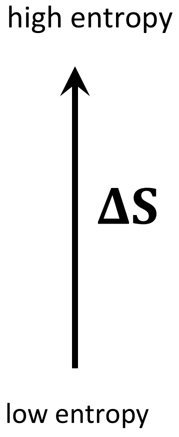
Entropy of an element in its most stable form is NOT zero!

- The third law of thermodynamics states that the entropy of a perfect, pure crystal at ________ K is given a value of zero.
- Thus, all absolute entropies for substances in the ________ world (above 0 K, not a pure substance, not perfect crystals) are _____ → even elements!

Two biggest factors for evaluating ΔS_{rxn}

1) Change in ____________ of matter (gas ____ aqueous ____ liquid ____ solid)
2) Change in_____ of particles (____ particles = ____ places for particles to be = ____ entropy)

Entropy Relationships

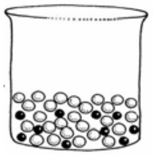


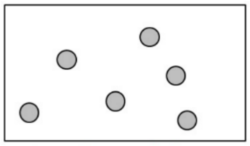
gas

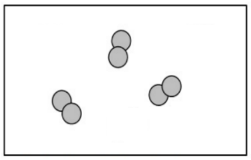
aqueous

liquid

solid







In order of lowest to highest entropy:

Solids < Liquids < Aqueous solutions << Gases < More moles of Gas

Rules for comparing entropy of individual molecules or materials:

→ Entropy ____ with <u>higher</u> temperature: more motion, more possible arrangements

→ Entropy ____ if substance <u>dissolves</u> in a solvent: more possible arrangements

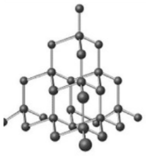
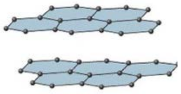
→ Entropy of a ________:

  - ■ ____with ____ pressure
  - ■ ____with ____ volume

→ Entropy is ____ for __________________ bonded compounds than for very strong covalent bonds: atoms have more wiggle room, more positional entropy

→ Entropy ____ as the _________________ (# of atoms, # of heavier atoms, # of e⁻, etc) of a molecule ____

Examples:

Less Entropy	**More Entropy**	**Why?**
diamond 	graphite 	Graphite has fewer bonds, more possible arrangements
butane gas (2 atm)	butane gas (1 atm)	Decreasing gas pressure increases volume, allowing more possible positions
$F_2(g)$	$Cl_2(g)$	$Cl_2(g)$ has more electrons/molar mass than $F_2(g)$, thus more possible arrangements of particles

Practice:

1) For the following reactions, is the entropy of the reaction *increasing* or *decreasing*?

a. $2\ NO_2(g) \rightarrow N_2O_4(g)$ Entropy is ________________________ Why?

b. $H_2(g) + Br_2(g) \rightarrow 2\ HBr(g)$ Entropy is ________________________ Why?

c. $Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$ Entropy is ________________________ Why?

2) Place the following in order of increasing entropy:

a) methane (CH_4), propane (C_3H_8), ethane (C_2H_6)

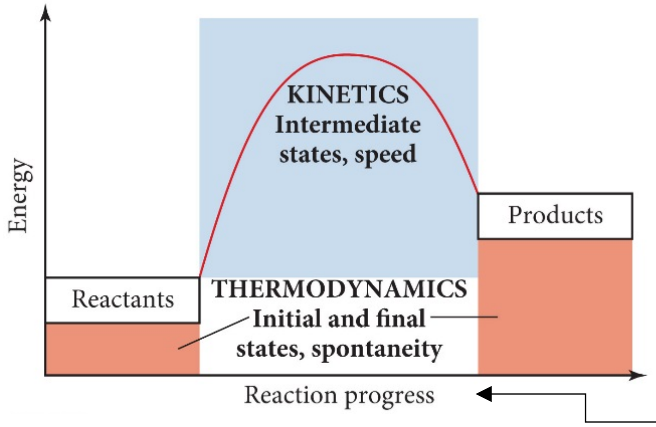
b) $NaCl(s)$, $LiCl(s)$, $RbCl(s)$, $KCl(s)$

c) O_2 (1 atm), O_2 (3 atm), O_2 (0.25 atm)

Gibbs Free Energy: How to Determine Thermodynamic Favorability

A clarifying note:

Thermodynamics (will a reaction happen spontaneously?) **vs Kinetics** (how fast will this reaction happen?)



Note x axis!

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and now back to… **Thermodynamically Favorable Processes**

Thermodynamically favored (aka spontaneous) processes take place **without *net* input** of energy from external sources.

Neither entropy (ΔS) nor enthalpy (ΔH) alone can tell us if a reaction is spontaneous:

→ must use a combination of the two: ________ **Free Energy**!

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

T = temperature in Kelvin

ΔH = enthalpy change (units of **kJ/mol** or **kJ mol^{-1}**)

ΔS = entropy change (units of **J/(mol K)** or **J mol^{-1} K^{-1}**)

Notice difference in units! You must make both units the same (J or kJ) before you calculate ΔG.

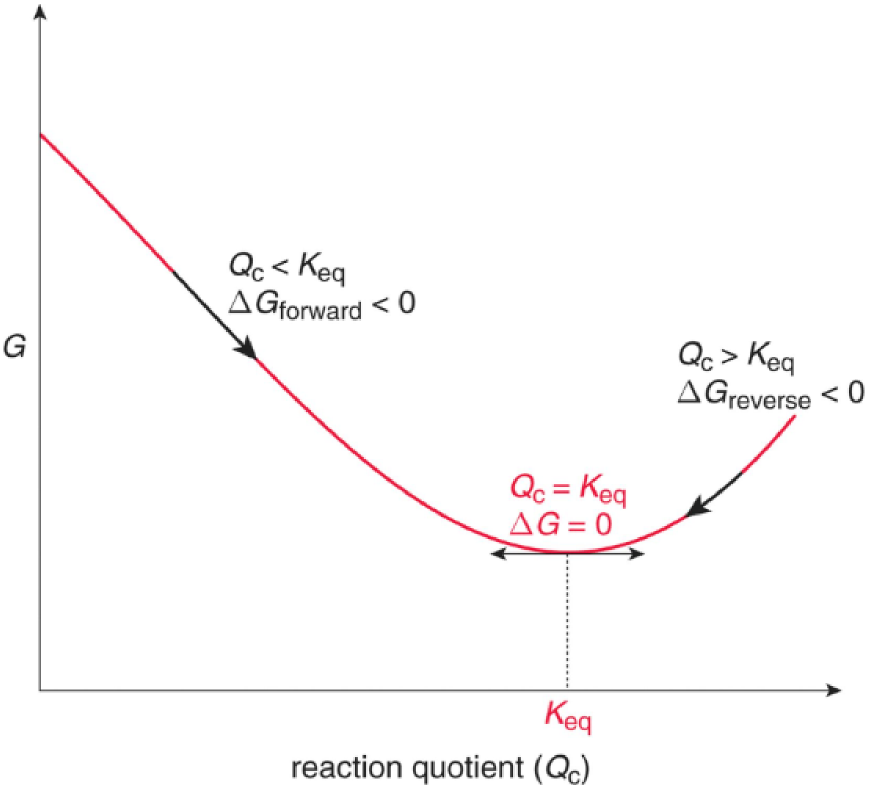
Δ**G** (Gibb's free energy) is the energy associated with a chemical reaction that can be used to do work.

	AP Units (Use on quizzes and tests)	**Mastering Chem Units**
ΔG	$\frac{\text{kJ}}{\text{mol}_{rxn}}$ *or* $\frac{\text{J}}{\text{mol}_{rxn}}$	kJ or J
ΔH	$\frac{\text{kJ}}{\text{mol}_{rxn}}$ *or* $\frac{\text{J}}{\text{mol}_{rxn}}$	kJ or J
ΔS	$\frac{\text{J}}{\text{mol} \cdot \text{K}}$	$\frac{\text{J}}{\text{K}}$

What does ΔG tell us?

The sign of ΔG tells if any given process is thermodynamically favored:

- $-\Delta \mathbf{G}$: Thermodynamically favored (spontaneous)
- $+\Delta \mathbf{G}$: Not thermodynamically favored (non-spontaneous)
- $\Delta \mathbf{G} = \mathbf{0}$: At equilibrium



Equations to calculate Gibb's Free Energy Change:

Given on formula chart:

$\Delta G^o = -nFE^o$ n = e^- transferred, F = Faraday's constant, E^o = standard potential

$\Delta G^o = -RT \ln K$ R = 8.314 J/(mol K), T = temp (K), ln(k) = natural log of K

$$\Delta G^o_{\text{reaction}} = \sum [G^o_f(\text{products})] - \sum [G^o_f(\text{reactants})]$$

Not given on formula chart:

$K = e^{-\Delta G^o / RT}$ e = 2.718, R = 8.314 J/(mol K), T = temp (K), ΔG must be in J/mol

Note: Elements DO have a value of zero for ΔG^o, because forming an element from itself results in no change.

Elements ARE zero for: ____________________

Elements are NOT zero for: ____________

$$\Delta G^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}$$

What does it mean when a reaction is "driven" by enthalpy, entropy, or both?

Nature favors:

1. ________ (increasing entropy, greater disorder)
2. ________ (exothermic, release of energy)

For a reaction to be thermodynamically favorable, ΔG must be negative, but that doesn't mean you need increasing entropy AND an exothermic process for a reaction to proceed spontaneously!

- **Driven by enthalpy:** a very exothermic reaction (_______) overcomes a decrease in entropy

Example: The formation of NaCl(s) from its elements is spontaneous and releases lots of heat.

$$2\ \mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{NaCl}(s)$$

This process clearly involves a decrease in entropy (−ΔS) since a gas molecule becomes a solid. However, it is a highly exothermic reaction (−ΔH). Because the large −ΔH is more negative than the positive −TΔS term, overall ΔG is negative, and the reaction is thermodynamically favored.

- **Driven by entropy:** an endothermic reaction occurs spontaneously because of a highly __________ ΔS

Example: $NaNO_3$(s) dissolves in water with a noticeable decrease in temperature.

$$\mathrm{NaNO_3}(s) \rightarrow \mathrm{Na^+}(aq) + \mathrm{NO_3^-}(aq)$$

This is an endothermic reaction, so ΔH is positive. However, $NaNO_3$ does dissolve, so ΔG must be negative. This means that ΔS must be very positive, so that the negative −TΔS term is sufficiently negative to overcome the positive ΔH, and the reaction is thermodynamically favored.

You can use the signs of ΔS and ΔH (together with the reaction ΔG = ΔH − TΔS) to predict qualitatively the spontaneity of a given reaction. You can show this graphically, which leads to the best graph of all time (but SHHHH! It's a secret. 😊)

You can also consider this in a table:

ΔH	ΔS	-TΔS	ΔG	**The reaction is thermodynamically:**
–	+	–	–	favored at all temperatures
+	–	+	+	not favored at all temperatures
–	–	+	+/–	favored only at low temperatures
+	+	–	+/–	favored only at high temperatures

Is it possible to have a reaction with –ΔG that doesn't happen?

Yes! Well, kind of. It is possible for a technically spontaneous reaction (i.e. **–**ΔG) to occur so _________________ that no measurable products form. This happens if a reaction has a VERY high ______________________ energy – we'll talk more about this during our next unit! When this happens, the reaction is said to be under __________________ control.

How can you "force" a reaction to run if it has +ΔG?

There are actually multiple ways! We already know about one of them.

1. Apply an external ____________ of energy, such as electricity in electrolysis reactions or light in photosynthesis.
2. ______________ thermodynamically unfavored reactions with thermodynamically favored ones, via their common intermediates.

$$\begin{array}{rr} \text{glucose} + \text{fructose} \leftrightharpoons \text{sucrose} & \Delta G = +27\ \text{kJ/mol} \\ \text{ATP} + \text{H}_2\text{O} \leftrightharpoons \text{ADP} + \text{P}_i & \Delta G = -30\ \text{kJ/mol} \\ \hline \text{glucose} + \text{fructose} + \text{ATP} \leftrightharpoons \text{sucrose} + \text{ADP} + \text{P}_i & \Delta G = -3\text{kJ/mol} \end{array}$$

Let's Practice!

1. Given the values for ΔH, ΔS, and T, determine whether each of the following sets of data represent thermodynamically favorable or non-thermodynamically favorable processes.

	ΔH (kJ/mol)	ΔS (J/mol K)	T (K)
a.	40	300	130
b.	40	300	150
c.	40	−300	150
d.	−40	−300	130
e.	−40	300	150

2. Using data for ΔH° and S°, calculate ΔG° for the following reaction at 25°C and 1 atm.

$$Cr_2O_3(s) + 2\ Al(s) \longrightarrow Al_2O_3(s) + 2\ Cr(s)$$

Substance	ΔH°_f(kJ/mol)
$Cr_2O_3(s)$	-1128 kJ/mol
$Al_2O_3(s)$	-1676 kJ/mol

Substance	S°(J/mol K)
Cr (s)	24 J/mol K
Al (s)	28 J/mol K
$Al_2O_3(s)$	51 J/mol K
$Cr_2O_3(s)$	81 J/mol K

3. Given the values of ΔG° you calculated in #2, calculate the equilibrium constant, *K*, for the following reaction (at 25°C):

$$Cr_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3(s) + 2Cr(s)$$

Is this reaction thermodynamically favorable? Use your calculated Gibb's free energy and equilibrium constant to justify your answer.

Multiple Choice Practice FTW!

$$2\ \mathrm{Al}(s) + 3\ \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{AlCl_3}(s)$$

4. The reaction above is not thermodynamically favored under standard conditions, but it becomes thermodynamically favored as the temperature decreases toward absolute zero. Which of the following is true?

   a. ΔS and ΔH are both negative.
   b. ΔS and ΔH are both positive.
   c. ΔS is negative, and ΔH is positive.
   d. ΔS is positive, and ΔH is negative.

$$\mathrm{H_2O}(l) \rightarrow \mathrm{H_2O}(s)$$

5. Which of the following is true for the above reaction?

   a. The value of ΔS is positive.
   b. The value of ΔG is positive at 298 K.
   c. The value of ΔH is positive.
   d. The reaction is favored at 1.0 atm and 298 K.

6. A chemical reaction has an equilibrium constant, K, equal to 1.0×10^{-6}. If, at a given point in the reaction, the value for the reaction quotient *Q* is determined to be 2.5×10^{-8}, what is true about Gibb's free energy at that moment?

   a. ΔG = 0
   b. ΔG > 0
   c. ΔG < 0
   d. The value of ΔG cannot be determined.

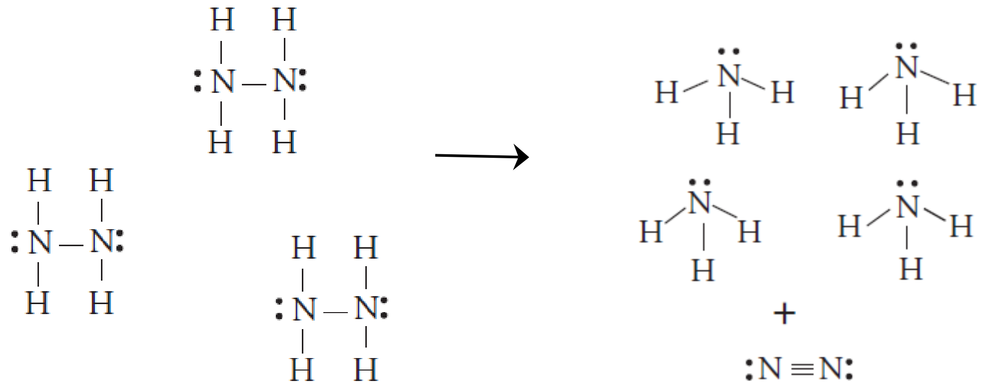
7. In which of the following reactions is entropy increasing?

   a. $2\ \mathrm{SO_2}(g) + \mathrm{O_2}(g) \rightarrow 2\ \mathrm{SO_3}(g)$
   b. $\mathrm{CO}(g) + \mathrm{H_2O}(g) \rightarrow \mathrm{H_2}(g) + \mathrm{CO_2}(g)$
   c. $\mathrm{H_2}(g) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{HCl}(g)$
   d. $2\ \mathrm{NO_2}(g) \rightarrow 2\ \mathrm{NO}(g) + \mathrm{O_2}(g)$

8. When calcium chloride ($\mathrm{CaCl_2}$) dissolves in water, the temperature of the water increases dramatically.

   a. Entropy
   b. Enthalpy
   c. Both enthalpy and entropy
   d. Neither enthalpy and entropy

9. The reaction shown in the diagram below is accompanied by a large increase in temperature. If all molecules shown are in their gaseous state, which statement accurately describes the reaction?



a. It is an exothermic reaction in which entropy increases.

b. It is an exothermic reaction in which entropy decreases.

c. It is an endothermic reaction in which entropy increases.

d. It is an endothermic reaction in which entropy decreases.

10. When solid ammonium chloride, $NH_4Cl(s)$, is added to water at 25°C, it dissolves and the temperature of the solution decreases. Which of the following is true for the values of ΔH and ΔS for the dissolving process?

	ΔH	ΔS
a.	Positive	Negative
b.	Positive	Positive
c.	Negative	Positive
d.	Negative	Negative

Enthalpy, Entropy and Gibb's Free Energy Activity

1. Baking Soda and Vinegar

Procedure: Put about 2 tablespoons of vinegar in one cup. Place about 2 tablespoons of water plus about 1 teaspoon of baking soda in a second cup. Stir the second cup. Pour the contents of the second cup into the first cup.

$$\mathrm{HC_2H_3O_2}(aq) + \mathrm{NaHCO_3}(s) \rightarrow \mathrm{NaC_2H_3O_2}(aq) + \mathrm{CO_2}(g) + \mathrm{H_2O}(l)$$

a) What do you observe through sight, sound, and touch?

b) Identify the sign of the entropy change for this reaction, ΔS. Justify your answer.

c) Based on your observations, identify the sign of the enthalpy change for this reaction, ΔH. Justify.

d) Fill out the potential energy profile for this reaction.

e) Based on your observations, identify the sign of the free energy change for this reaction. Justify your answer.

f) Calculate Gibbs free energy, given ΔH = 12.0 kJ/mol$_{rxn}$, ΔS = 189.1 J/(mol K), and a room temperature of 21°C.

g) What is the driving force of this reaction? (What is making the free energy value negative?)

2. Steel Wool and 9-V Battery

Procedure: Obtain a pinch of steel wool and increase its surface area. Touch the two leads of the battery to the steel wool and remove.

$$4\,\mathrm{Fe}(s) + 3\,\mathrm{O_2}(g) \rightarrow 2\,\mathrm{Fe_2O_3}(s)$$

a) The mass of the steel wool increases during this reaction. Explain why this happens.

b) Based on your observations, identify the sign of the enthalpy change for this reaction. Justify.

c) Fill out the potential energy profile for this reaction.

d) What is the role of the 9-V battery?

e) Identify the sign of the entropy change for this reaction. Justify your answer.

f) Based on your observations, explain the free energy change for this reaction.

g) Calculate Gibbs free energy, given $\Delta H = -1{,}644.4\ \mathrm{kJ/mol_{rxn}}$, $\Delta S = -543.4\ \mathrm{J/(mol\ K)}$, and room temp of 21°C.

h) What is the driving force for this reaction? Justify your answer.

Thermodynamics Study Guide

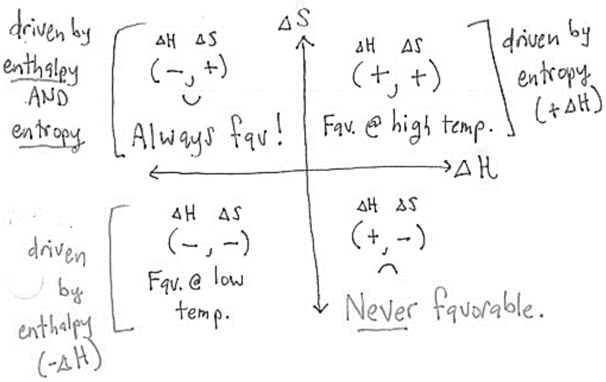
$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

Elements ARE zero for: $\Delta H^\circ, \Delta G^\circ$
Elements are NOT zero for: ΔS°

Nature favors:

1. ________ (increasing entropy, greater disorder)
2. ________ (exothermic, release of energy)

- **Driven by enthalpy:** a very exothermic reaction (_______) overcomes a decrease in entropy
- **Driven by entropy:** an endothermic reaction occurs spontaneously because of a highly __________ ΔS
- You can use the signs of ΔS and ΔH (together with the reaction ΔG = ΔH − TΔS) to qualitatively predict the spontaneity of a given reaction. You can show this graphically, which leads to the best graph of all time (but SHHHH! It's a secret. 😊)



E^o_{cell}	ΔG^o	K	**K & Q Relationship**	**Reaction Direction**	**Spontaneity in the Forward Direction** (as written)
+	−	K > 1	K > Q	Forward	thermodynamically favorable
−	+	K < 1	K < Q	Reverse	NOT thermodynamically favorable
= 0	= 0	K = 1	K = Q	No reaction	n/a

At standard conditions (1 M, 1 atm, 298 K): $$\Delta G^o = -nFE^o_{cell}$$

n = number of moles of electrons transferred in a **BALANCED** redox reaction
F = faraday's constant = 96,485 C/ mol e^- (charge on one mole of electrons)

At standard conditions (1 M, 1 atm, 298 K): $$\Delta G^o = -RT\ln K$$

R = = 8.314 J mol^{-1} K^{-1} T = temperature (in Kelvin) K = equilibrium constant

For both reactions, the units for $\Delta G^o = \frac{\text{joules}}{\text{moles}_{\text{reaction}}} = \frac{\text{J}}{\text{mol}_{\text{rxn}}}$

END-ing a bond (or IMF/attraction) is always END-othermic! (+ΔH)

PROBLEM SOLVING IN THERMODYNAMICS

Determining
ΔH = Enthalpy

Calorimetry:

$$q_{gained} = -q_{lost}$$

Assess everything losing and gaining heat in the system. Determine energy change (J) associated with each.

Types of Calorimetry:

1. Things in water:
    a. Solid, non-reacting metal
    $$[mC\Delta T]_{water} = -[mC\Delta T]_{metal}$$
    b. Aqueous/reacting things
    $$q_{rxn} = -m_{total}C\Delta T$$

Usually C = 4.18 J/g°C = C_{water}, but not always! Be careful.

2. Things separate from each other (often this means using a rxn to heat something up):
$$q_{gained} = -q_{lost}$$
$$[mC\Delta T]_{gained} = -[mC\Delta T]_{lost}$$

Hess's Law:

Using the ΔH's from equations that were added to get the "goal" equation.

Whatever you do to the equation, you must do to ΔH!

Ex. *Flipping an equation ΔH * (–1)*
*Multiply by two = ΔH * 2*
Divide by two = ΔH / 2

Standard Heats of Formation:

Amount of energy involved in the formation of **1 mole** of a substance.

$\Delta H_{rxn} = \sum \Delta H_f$ products - $\sum \Delta H_f$ reactants

*Don't forget to multiply by coefficients!

The Question to answer: **Is the reaction thermodynamically favorable?**

Bond Energy:

The energy it takes to break a bond. Higher BE implies stronger bonds. Higher bond orders have higher BE's.

$\Delta H_{rxn} = \sum \Delta H_{BE}$ reactants - $\sum \Delta H_{BE}$ products
$\Delta H_{rxn} = \sum \Delta H_{BE}$ broken - $\sum \Delta H_{BE}$ formed

**Be sure to draw structures to see exactly how many of each bond type are broken and formed!*

Determining
ΔS = Entropy

Qualitative:

$S_{solid} < S_{liquid} << S_{gas}$

Compare products to reactants to judge if entropy is increasing/decreasing. Entropy is greater on the side with more moles of gas… etc. Entropy is higher for molecules of higher structural complexity.

Quantitative:

$\Delta S_{rxn} = \sum S^\circ$ products - $\sum S^\circ$ reactants

Determining
ΔG = Free Energy

Using ΔH and ΔS:

When gas pressures all equal 1 atm.

$$\Delta G = \Delta H - T\Delta S$$

**Units of enthalpy are often kJ/mol while entropy is in J/mol K. Make them match!

Related to Q and K (Equil.):

$\Delta G^\circ = -RT\ln(K)$

$R = 8.314 \times 10^{-3}$ kJ/mol*K

Write in alpha order to predict shift.
K > Q, shift right, ΔG = – (favorable)
K = Q, at equilibrium, ΔG = 0
K < Q, shift left, ΔG = + (not favorable)

Le Châtelier's Principle:

Endo (heat as reactant): ↑ T, ↑ K
Exo (heat as product): ↑ T, ↓ K

Related to E°_{cell} (Electro):

$\Delta G^\circ = -nF(E^\circ_{cell})$

F = 96,485 C/mol e–
n = number of electrons transferred

Free Energy of Formation:

$\Delta G_{rxn} = \Sigma \Delta G_f$ products - $\Sigma \Delta G_f$ reactants

If ΔG is negative, the reaction IS thermodynamically favorable.
If ΔG is positive, the reaction is NOT thermodynamically favorable.
**Be able to predict spontaneity based on enthalpy, entropy, temperature, voltage, and equilibrium quotients.*

Unit 5 Multiple Choice Practice

$$3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$$

1. Gaseous hydrogen and nitrogen combine in the reaction above to form ammonia, NH_3. Using the table of entropies provided below, calculate the standard entropy change, $\Delta S°_{rxn}$, for the reaction shown at 25°C.

Substance	S° at 25°C (J/mol K)
$H_2(g)$	131
$N_2(g)$	192
$NH_3(g)$	193

a. −199 J/mol K
b. −386 J/mol K
c. +199 J/mol K
d. +386 J/mol K

$$MgO(s) \rightarrow Mg^{2+}(aq) + O^{2-}(aq)$$

2. Is the process above endothermic or exothermic? Why?
   a. Exothermic, because one solid particle becomes two aqueous particles, increasing the total possible positions.
   b. Endothermic, because one solid particle becomes two aqueous particles, increasing the total possible positions.
   c. Exothermic, because overcoming the Coulombic forces within MgO(s) releases energy.
   d. Endothermic, because overcoming the Coulombic forces within MgO(s) requires the input of energy.

3. When ammonium nitrate (NH_4NO_3) dissolves in water, the temperature of the water decreases dramatically. Which is the primary driving factor behind this reaction?

a. Entropy
c. Both enthalpy and entropy
b. Enthalpy
d. Neither enthalpy and entropy

$$2\ AlCl_3(s) \rightarrow 2\ Al(s) + 3\ Cl_2(g)$$

4. The reaction above is spontaneous under standard conditions, but it becomes thermodynamically unfavorable as temperature decreases. Which of the following is true?

a. ΔS and ΔH are both negative.
c. ΔS is negative, and ΔH is positive.
b. ΔS and ΔH are both positive.
d. ΔS is positive, and ΔH is negative.

$$\mathrm{CH_3OH}(l) \rightarrow \mathrm{CH_3OH}(g)$$

5. For the boiling of methanol, CH_3OH, $\Delta H^\circ = +37.6\ \mathrm{kJ\ mol^{-1}}$ and $\Delta S^\circ = +111\ \mathrm{J\ mol^{-1}\ K^{-1}}$. Calculate the standard free energy change of this reaction, ΔG°_{rxn}.

   a. −4.5 kJ/mol
   b. −2.8 kJ/mol
   c. +2.8 kJ/mol
   d. +4.5 kJ/mol

$I_2(g) + 3\ Cl_2(g) \rightarrow 2\ ICl_3(g)$ $\Delta H = -390$ kJ/mol

6. Which of the following statements accurately describes the above reaction?

   a. The entropy of the products exceeds that of the reactants.
   b. $I_2(g)$ will always be the limiting reagent.
   c. This reaction is never thermodynamically favored.
   d. The temperature of the surroundings will increase as this reaction progresses.

$$\mathrm{H_2O}(l) \rightarrow \mathrm{H_2O}(s)$$

7. Is the process above endothermic or exothermic? Why?

   a. Exothermic, because forming new intermolecular attractions between water molecules releases energy.
   b. Endothermic, because energy is needed to form new intermolecular attractions between water molecules.
   c. Exothermic, because forming new bonds between water molecules releases energy.
   d. Endothermic, because energy is needed to form new bonds between water molecules.

$$\mathrm{H_2O}(l) \rightarrow \mathrm{H_2O}(g)$$

8. Which of the following is true for the above reaction?

   a. The value of ΔS is negative.
   c. The value of ΔH is positive.
   b. The value of ΔG is negative at 298 K.
   d. The reaction is favored at 1.0 atm and 298 K.

Use the following information to answer the next two questions.

Reaction 1: $N_2H_4(l) + H_2(g) \rightarrow 2NH_3(g)$ $\Delta H = ?$

Reaction 2: $N_2H_4(l) + CH_4O(l) \rightarrow CH_2O(g) + N_2(g) + 3H_2(g)$ $\Delta H = -37\ \text{kJ/mol}_{\text{rxn}}$

Reaction 3: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ $\Delta H = -46\ \text{kJ/mol}_{\text{rxn}}$

Reaction 4: $CH_4O(l) \rightarrow CH_2O(g) + H_2(g)$ $\Delta H = -65\ \text{kJ./mol}_{\text{rxn}}$

9. If reaction 4 (shown above) were repeated at a higher temperature, how would the reaction's value for ΔG be affected?

   a. It would become more negative because entropy is a driving force behind this reaction.
   b. It would become more positive because enthalpy is a driving force behind this reaction.
   c. It will stay the same because reaction 4 is never thermodynamically favorable.
   d. It will stay the same because reaction 4 is thermodynamically favorable at all temperatures.

10. Under what conditions would reaction 2 (shown above) be thermodynamically favored?

    a. It is always favored.
    b. It is never favored.
    c. It is only favored at low temperatures.
    d. It is only favored at high temperatures.

Unit 4: AP Free Response Practice #1 [2017 #2, shortened, 5 points]

1. Answer the following questions about the isomers fulminic acid and isocyanic acid.

   Fulminic acid can convert to isocyanic acid according to the equation below.

$$\mathrm{HCNO}(g) \rightleftarrows \mathrm{HNCO}(g)$$

fulminic acid *isocyanic acid*

Fulminic Acid	Isocyanic Acid
$\mathrm{H{-}C{\equiv}N{-}\underset{\cdot\cdot}{\ddot{O}}}{:}$	$\mathrm{H{-}\ddot{N}{=}C{=}\ddot{O}}{:}$

   a. Using the Lewis electron-dot diagrams of fulminic acid and isocyanic acid shown in the boxes above and the table of average bond enthalpies below, determine the value of ΔH° for the reaction of HCNO(g) to form HNCO(g). [*2 points*]

Bond	Enthalpy (kJ/mol)	Bond	Enthalpy (kJ/mol)	Bond	Enthalpy (kJ/mol)
N−O	201	C=N	615	H−C	413
C=O	745	C≡N	891	H−N	391

   b. A student claims that ΔS° for the reaction is close to zero. Explain why the student's claim is accurate. [*1 point*]

   c. Which species, fulminic acid (HCNO) or isocyanic acid (HNCO), is present in high concentration at equilibrium at 298 K? Justify your answer in terms of the thermodynamic favorability and the equilibrium constant. [*2 points*]

Unit 4: AP Free Response Practice #2 [LTF Free Response #2, 10 points]

$$N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g) \qquad \Delta G^\circ_{rxn} = -34.1\ \text{kJ mol}^{-1}$$

2. The following questions relate the the synthesis reaction represented by the chemical equation above.

   a. Is the reaction thermodynamically favorable or unfavorable under standard conditions at 298 K? Justify your answer. [*1 point*]

   b. In terms of the equilibrium constant, K, for the above reaction at 25°C

      i. Predict whether K will be greater than, less than, or equal to one. Justify your choice. [*1 point*]

      ii. Calculate its value. [*2 points*]

   c. Given the following data, determine the ΔH° for the above reaction. [*2 points*]

Substance	ΔH°_f (kJ mol^{-1})
$NH_3(g)$	−46.1

d. In terms of the standard entropy change, ΔS^{o}

i. Predict the sign of ΔS^{o} for the above reaction. Justify your answer. [*1 point*]

ii. Calculate the value of ΔS^{o} for the synthesis reaction given at 25°C. [*1 point*]

e. Using the data in the table below and the enthalpy of reaction, ΔH^{o}_{rxn} determined in part (c), calculate the approximate bond energy of the nitrogen-hydrogen bond in ammonia. [*2 points*]

Bonds	Approximate Bond Energy (kJ mol^{-1})
$\text{N}-\text{H}$	???
$\text{H}-\text{H}$	430
$\text{N}\equiv\text{N}$	960

STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			$E°$(V)
$F_2(g) + 2e^-$	→	$2F^-$	2.87
$Co^{3+} + e^-$	→	Co^{2+}	1.82
$Au^{3+} + 3e^-$	→	$Au(s)$	1.50
$Cl_2(g) + 2e^-$	→	$2Cl^-$	1.36
$O_2(g) + 4H^+ + 4e^-$	→	$2H_2O(l)$	1.23
$Br_2(l) + 2e^-$	→	$2Br^-$	1.07
$2Hg^{2+} + 2e^-$	→	$Hg_2{}^{2+}$	0.92
$Hg^{2+} + 2e^-$	→	$Hg(l)$	0.85
$Ag^+ + e^-$	→	$Ag(s)$	0.80
$Hg_2{}^{2+} + 2e^-$	→	$2Hg(l)$	0.79
$Fe^{3+} + e^-$	→	Fe^{2+}	0.77
$I_2(s) + 2e^-$	→	$2I^-$	0.53
$Cu^+ + e^-$	→	$Cu(s)$	0.52
$Cu^{2+} + 2e^-$	→	$Cu(s)$	0.34
$Cu^{2+} + e^-$	→	Cu^+	0.15
$Sn^{4+} + 2e^-$	→	Sn^{2+}	0.15
$S(s) + 2H^+ + 2e^-$	→	$H_2S(g)$	0.14
$2H^+ + 2e^-$	→	$H_2(g)$	0.00
$Pb^{2+} + 2e^-$	→	$Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	→	$Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	→	$Ni(s)$	–0.25
$Co^{2+} + 2e^-$	→	$Co(s)$	–0.28
$Cd^{2+} + 2e^-$	→	$Cd(s)$	–0.40
$Cr^{3+} + e^-$	→	Cr^{2+}	–0.41
$Fe^{2+} + 2e^-$	→	$Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	→	$Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	→	$Zn(s)$	–0.76
$2H_2O(l) + 2e^-$	→	$H_2(g) + 2OH^-$	–0.83
$Mn^{2+} + 2e^-$	→	$Mn(s)$	–1.18
$Al^{3+} + 3e^-$	→	$Al(s)$	–1.66
$Be^{2+} + 2e^-$	→	$Be(s)$	–1.70
$Mg^{2+} + 2e^-$	→	$Mg(s)$	–2.37
$Na^+ + e^-$	→	$Na(s)$	–2.71
$Ca^{2+} + 2e^-$	→	$Ca(s)$	–2.87
$Sr^{2+} + 2e^-$	→	$Sr(s)$	–2.89
$Ba^{2+} + 2e^-$	→	$Ba(s)$	–2.90
$Rb^+ + e^-$	→	$Rb(s)$	–2.92
$K^+ + e^-$	→	$K(s)$	–2.92
$Cs^+ + e^-$	→	$Cs(s)$	–2.92
$Li^+ + e^-$	→	$Li(s)$	–3.05

Polyatomic Ions

1st six-weeks

Nick the **C**amel ***ate*** an **I**cky **Cl**am for **S**upper in **P**hoenix with his **Br**os

NO_3^-	nitrate	ClO_3^-	chlorate	PO_3^{3-}	phosphite
NO_2^-	nitrite	ClO_2^-	chlorite	BrO_3^-	bromate
CO_3^{2-}	carbonate	SO_4^{2-}	sulfate	BrO_2^-	bromite
IO_3^-	iodate	SO_3^{2-}	sulfite		
IO_2^-	iodite	PO_4^{3-}	phosphate		

Consonants = # of Oxygen

Vowels = Charge

2nd six-weeks

ClO_4^-	perchlorate	IO^-	hypoiodite
ClO^-	hypochlorite	BrO_4^-	perbromate
IO_4^-	periodate	BrO^-	hypobromite

	Difference in Oxygen from ATE
Per_____ate	+1
Ate	0
Ite	−1
Hypo ____ite	−2

3rd six-weeks

$H_2PO_4^-$	dihydrogen phosphate	HCO_3^-	hydrogen carbonate *or* bicarbonate
HPO_4^{2-}	hydrogen phosphate	HSO_4^-	hydrogen sulfate

4th six-weeks

NH_4^+	ammonium	OH^-	hydroxide
$C_2H_3O_2^-$ *or* CH_3COO^-	acetate	H_3O^+	hydronium

5th six-weeks

MnO_4^-	permanganate	CrO_4^{2-}	chromate
CN^-	cyanide	$Cr_2O_7^{2-}$	dichromate

6th six-weeks

O_2^{2-}	peroxide	$C_4H_4O_6^{2-}$	tartrate
$S_2O_3^{2-}$	thiosulfate	$C_2O_4^{2-}$	oxalate

Acid Nomenclature	
Binary	Hydro_______ic
ate	ic
ite	ous

Polyatomic Ions

Br–Based Ions

BrO^-	hypobromite
BrO_2^-	bromite
BrO_3^-	bromate
BrO_4^-	perbromate

Cr–Based Ions

CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate

I–Based Ions

IO^-	hypoiodite
IO_2^-	iodite
IO_3^-	iodate
IO_4^-	periodate

P–Based Ions

PO_3^{3-}	phosphite
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate

Other Ions

CN^-	cyanide
O_2^{2-}	peroxide
MnO_4^-	permanganate

C–Based Ions

CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate *or* bicarbonate
$C_2H_3O_2^-$ *or* CH_3COO^-	acetate
$C_4H_4O_6^{2-}$	tartrate
$C_2O_4^{2-}$	oxalate

Cl–Based Ions

ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

N–Based Ions

NO_2^-	nitrite
NO_3^-	nitrate
NH_4^+	ammonium

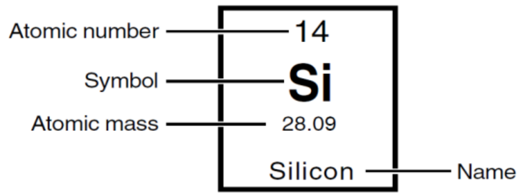
S–Based Ions

SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate
$S_2O_3^{2-}$	thiosulfate

Acid & Base Ions

H_3O^+	hydronium
OH^-	hydroxide

PERIODIC TABLE OF THE ELEMENTS



	1	2											3	4	5	6	7	8
1	1 **H** 1.008 Hydrogen																	2 **He** 4.00 Helium
2	3 **Li** 6.94 Lithium	4 **Be** 9.01 Beryllium											5 **B** 10.81 Boron	6 **C** 12.01 Carbon	7 **N** 14.01 Nitrogen	8 **O** 16.00 Oxygen	9 **F** 19.00 Fluorine	10 **Ne** 20.18 Neon
3	11 **Na** 22.99 Sodium	12 **Mg** 24.30 Magnesium											13 **Al** 26.98 Aluminum	14 **Si** 28.09 Silicon	15 **P** 30.97 Phosphorus	16 **S** 32.06 Sulfur	17 **Cl** 35.45 Chlorine	18 **Ar** 39.95 Argon
4	19 **K** 39.10 Potassium	20 **Ca** 40.08 Calcium	21 **Sc** 44.96 Scandium	22 **Ti** 47.90 Titanium	23 **V** 50.94 Vanadium	24 **Cr** 52.00 Chromium	25 **Mn** 54.94 Manganese	26 **Fe** 55.85 Iron	27 **Co** 58.93 Cobalt	28 **Ni** 58.69 Nickel	29 **Cu** 63.55 Copper	30 **Zn** 65.39 Zinc	31 **Ga** 69.72 Gallium	32 **Ge** 72.59 Germanium	33 **As** 74.92 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.90 Bromine	36 **Kr** 83.80 Krypton
5	37 **Rb** 85.47 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.91 Yttrium	40 **Zr** 91.22 Zirconium	41 **Nb** 92.91 Niobium	42 **Mo** 95.94 Molybdenum	43 **Tc** (98) Technetium	44 **Ru** 101.10 Ruthenium	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.75 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.91 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	57 ***La** 138.91 Lanthanum	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.85 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.2 Osmium	77 **Ir** 192.2 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** (209) Polonium	85 **At** (210) Astatine	86 **Rn** (222) Radon
7	87 **Fr** (223) Francium	88 **Ra** 226.02 Radium	89 **†Ac** 227.03 Actinium	104 **Rf** (261) Rutherfordium	105 **Db** (262) Dubnium	106 **Sg** (266) Seaborgium	107 **Bh** (264) Bohrium	108 **Hs** (277) Hassium	109 **Mt** (268) Meitnerium	110 **Ds** (271) Darmstadtium	111 **Rg** (272) Roentgenium							

***Lanthanide Series**	58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** (145) Promethium	62 **Sm** 150.4 Samarium	63 **Eu** 151.97 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.93 Terbium	66 **Dy** 162.5 Dysprosium	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.04 Ytterbium	71 **Lu** 174.97 Lutetium
†Actinide Series	90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactnium	92 **U** 238.03 Uranium	93 **Np** (237) Neptumium	94 **Pu** (244) Plutonium	95 **Am** (243) Americium	96 **Cm** (247) Curium	97 **Bk** (247) Berkelium	98 **Cf** (251) Californium	99 **Es** (252) Einsteinium	100 **Fm** (257) Fermium	101 **Md** (258) Mendelevium	102 **No** (259) Nobelium	103 **Lr** (262) Lawrencium

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL = liter(s), milliliter(s)
g = gram(s)
nm = nanometer(s)
atm = atmosphere(s)
mm Hg = millimeters of mercury
J, kJ = joule(s), kilojoule(s)
V = volt(s)
mol = mole(s)

ATOMIC STRUCTURE

$$E = h\nu$$
$$c = \lambda\nu$$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number = 6.022×10^{23} mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}\text{, where } a\,\mathrm{A} + b\,\mathrm{B} \rightleftarrows c\,\mathrm{C} + d\,\mathrm{D}$$

$$K_p = \frac{(P_\mathrm{C})^c(P_\mathrm{D})^d}{(P_\mathrm{A})^a(P_\mathrm{B})^b}$$

$$K_a = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$K_b = \frac{[\mathrm{OH}^-][\mathrm{HB}^+]}{[\mathrm{B}]}$$

$$K_w = [\mathrm{H}^+][\mathrm{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\mathrm{C}$$
$$= K_a \times K_b$$

$$\mathrm{pH} = -\log[\mathrm{H}^+]\text{, } \mathrm{pOH} = -\log[\mathrm{OH}^-]$$

$$14 = \mathrm{pH} + \mathrm{pOH}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$\mathrm{p}K_a = -\log K_a\text{, } \mathrm{p}K_b = -\log K_b$$

<u>Equilibrium Constants</u>

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$\ln[\mathrm{A}]_t - \ln[\mathrm{A}]_0 = -kt$$

$$\frac{1}{[\mathrm{A}]_t} - \frac{1}{[\mathrm{A}]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_{\text{A}}, \text{ where } X_{\text{A}} = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_{\text{A}} + P_{\text{B}} + P_{\text{C}} + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, $R = 8.314 \text{ J mol}^{-1} \text{K}^{-1}$
$= 0.08206 \text{ L atm mol}^{-1} \text{K}^{-1}$
$= 62.36 \text{ L torr mol}^{-1} \text{K}^{-1}$
1 atm = 760 mm Hg = 760 torr
STP = 273.15 K and 1.0 atm
Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

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

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -RT \ln K \\ &= -nFE^\circ \end{aligned}$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.10	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

***Lanthanide Series**	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.5	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)