Metric and English Units and Unit Conversions

Activity 1

The purpose of this activity is for you to practice using units of energy, to review your metric system prefixes and conversions, and to practice some unit conversion using unit factor analysis (dimensional analysis) as a problem solving technique.

1. Complete the equations with the necessary variables (refer to this section's Guided Practice for help):

$\boxed{\_\_ = \tfrac{1}{2} mv^2}$ $\boxed{U_E = k\frac{qQ}{\_\_}}$

$\boxed{F = \_\_ a}$ $\boxed{U_G = mg\_\_}$

2. What are the units of force and the units of energy? Please also write these in base units.
3. Another equation for force is pressure multiplied by area. Using base units for force and area, derive the base units for pressure. These units are equivalent to Pascals.
4. Order these values from least to greatest amount of energy (hint: convert them all to one type of unit first): 150 eV, 16 Cal, 3 foot-lb, 1 J, 10 L-atm

Activity 2

The purpose of this activity is to practice using values and units in equations. Please calculate the values using the proper significant figures.

1. How many joules of gravitational potential energy are represented when 20kg worth of chemistry textbooks are sitting on the top of your bookshelf 3 meters above the ground? (Recall: acceleration due to gravity is 9.81 m/s^2)
2. You're really hungry as you stare at the stack of chemistry textbooks atop your bookshelf. How many calories (cal) are stored in the gravitational potential energy?

Activity 3

The purpose of this activity is to practice using dimensional analysis to convert between units. Please calculate the values using the proper significant figures.

1. Express 64 milliliters in liters.
2. Express 428 ^{m2} in k^{m2}.
3. Express 0.38 nanomoles in micromoles.

4. Express 1,296 mm in m.
5. Express 18 kilograms in milligrams.
6. Express 3.5 m^3 in cm^3.

Introduction to the Periodic Table

Activity 1

The purpose of this activity is to verify that you are familiar with the history and the basic language associated with the periodic table.

1. Mendeleev defined a law that states the properties of the elements are periodic functions of their atomic number. This is known as ____________________.
2. Elements on the modern periodic table are arranged in order of ____________________ atomic number and weight.

Activity 2

The purpose of this activity is to check your understanding of the layout of the periodic table.

1. In the periodic table, each horizontal row is referred to as a ____________________. On the other hand, each vertical column is called a ____________________ or a ____________________.
2. In a group or family elements have ____________________ properties.
3. The elements of Group IA in the periodic table, except hydrogen, belong to the ____________________ family, while the elements of Group IIA belong to the ____________________ family.
4. The following elements: F, Cl, Br, and I belong to the ____________________ family.
5. The elements in Group 18 (8A) are part of the ____________________ family.
6. The periodic table is also divided into blocks that include the metals and the non-metals. Metals exist on the ____________________ side of the periodic table. Nonmetals exist on the ____________________ side of the periodic table.
7. The elements with properties intermediate between metals and nonmetals are known as ____________________ or semi-metals.
8. Elements in the A groups are called ____________________ elements.
9. Elements in the B groups are known as the ____________________ elements.

Isotopes and Atomic Mass

Activity 1

The purpose of this activity is for you to practice using the symbols associated with the atoms of the elements and to understand how the atomic weights are calculated based on relative percentages of naturally occurring isotopes.

1. The number of protons in the nucleus is referred to as the ____________________.
2. The total number of particles in the nucleus is referred to as ____________________.
3. An atom of indium, In, with a mass number of 113 has __________ protons, __________ electrons and __________ neutrons.
4. The isotopic notation of carbon–12 is: _________ and that of carbon–13 is _________.

5. Naturally occurring copper consists of two isotopes: 69.1% ^{63}Cu with a mass of 62.9 amu and 30.9% ^{65}Cu, which has a mass of 64.9 amu. Calculate the atomic mass of copper.

6. Naturally occurring lead is composed of four isotopes: ^{204}Pb with a 1.40% abundance and a mass of 203.97 amu, ^{206}Pb with a 24.10% abundance and a mass of 205.97 amu, ^{207}Pb with a 22.10% abundance and a mass of 206.98 amu and ^{208}Pb with a 52.40% abundance and a mass of 207.98 amu. What is the atomic mass of lead?

Elements and Symbols

Activity 1

The purpose of this activity is to practice your knowledge of names and symbols of the first 92 elements listed on the periodic table of the elements Given the name of a particular element you should be able to write the elemental symbol and vice versa.

1. Practice naming the elements and their symbols here:

Elemental Name	Elemental Symbol
Copper	
	Ca
Cobalt	
	Cd
Chromium	
	Cl
Caesium (or Cesium)	
	Ce
Californium	

Activity 2

The purpose of this activity is to practice your comprehension of the number of electrons and protons in neutral atoms.

1. How is the atomic number of an element related to the symbol “Z”?

2. How many protons and electrons are in the following neutral atoms:
   Titanium
   Platinum
   Thorium

Early Experiments

Activity 1

1. Match the term on the left with the definition on the right.

_______ Law of Conservation of Mass	a. when two elements can combine to create different compounds, the element which combines variably has whole number ratios across the different compounds.
_______ Law of Definite Proportions	b. matter is neither created nor destroyed, only arranged in a way during a chemical reaction.
_______ Law of Multiple Proportions	c. all matter is composed of atoms, which are indivisible particles.
_______ One of Dalton's Principles	d. a particular compound will always contain the same ratio of elements.

Activity 2

1. Match the experiment on the left with the set-up (a–c) and the outcome (i–iii) below.

_________ Rutherford's Gold Foil Experiment
_________ Millikan's Oil Drop Experiment
_________ Thomson's Cathode Ray Experiment

a. a voltage sends a beam of particles between electrodes at either end of an evacuated tube.
b. oil droplets are sprayed out of a nozzle, become charged and then are directed into an electric field.
c. a stream of alpha particles directed to a piece of gold foil are detected around the perimeter of the testing area.
i. the charge of an electron is $1.6\text{x}10^{-19}$ J.
ii. atoms have a small, positively charged center called a nucleus surrounded by electrons.
iii. electrons are extremely small, negatively charged particles.

Introduction to Electromagnetic Radiation

Activity 1

1. What are the two components that make up electromagnetic radiation?

2. Please match the symbol on the left to the appropriate description on the left.
   _____ E a. wavelength, m
   _____ λ b. the speed of light in a vacuum, 3.00 x 10^8 m/s
   _____ ν c. energy, J
   _____ h d. Planck's constant, 6.626 x 10^{-34} Js
   _____ c e. frequency, Hz or s^{-1}
3. Please write at least two valid equations that use the symbols listen in question 2 above.

Activity 2

1. List the regions of electromagnetic radiation in decreasing order from highest energy to lowest energy.

2. What is the range of wavelengths for visible light? Range of frequencies?

3. A hypothetical wave has a frequency of 1.0 x 10^{15} Hz. What is its hypothetical, approximate wavelength? State in which region of light this wave exists.

4. A hypothetical wave has a wavelength of 10^{-5} m. What is its hypothetical, approximate frequency? State in which region of light this wave exists.

5. Electrons are __________ charged and so are attracted to the __________ side of a static electric field. In an oscillating electric field, electrons also __________.
6. Electrons are __________ charged and so are attracted to the ________ side of a static electric field. In an oscillating electric field, electrons also __________.
7. How would doubling the frequency (v, pronounced nu, not vee) of a given wave affect the following values?
   a. λ
   b. c
   c. E
   d. h
8. Without calculations, rank the frequency of the following waves from highest to lowest frequency: waves with wavelengths of 300 nm, 1 m, 1 km, and 300 Å.

Blackbody Radiation and Ultraviolet Catastrophe

Activity 1

1. Classical mechanics made a prediction about the power/intensity emitted by black body radiators at shorter and shorter wavelengths. What was this prediction and did it match the

experimental observations?

2. The scientist _______________________ suggested a solution to this problem dubbed the _______________________.
3. This scientist's suggestion indicated that the emission of a particular frequency requires an oscillator to have a _______________________ energy. Since physical objects have a limited number of oscillation energies, they have a limited number of emitted frequencies.

The Photoelectric Effect

Activity 1

1. Draw a picture illustrating the photoelectric effect. Describe the figure and explain how frequency and work function (Φ) relate to the kinetic energy of the emitted electron.

2. You shine 500 nm light on a metal and electrons come off. What will happen if you shine 400 nm light of the same intensity on the metal?
   a. fewer electrons will come off but with the same velocities
   b. nothing
   c. electrons will come off with higher velocities
   d. the two situations will be identical
3. Explain your answer choice in question 2 above. Devise experiments that would produce the other three outcomes that you did not choose.

Activity 4

1. Describe the relationship between frequency of incident light and the kinetic energy of ejected electrons.
2. A packet of quantized energy associated with electromagnetic radiation is called a _________.
3. Famous scientist ___________ proposed the idea of a photon after interpreting the results of the ___________ effect. He noted that increasing the intensity of light corresponded to ___________ electrons being ejected. However, if no electrons were being ejected in the first place, ___________ would happen even if the intensity were increased. This led him to develop the idea of ___________ of light rather than waves of light energy that interact individually with electrons.

Activity 5

1. A photon with enough energy, 5.1 electron volts (eV) of energy to be precise, will eject an electron from a piece of gold. What frequency and wavelength does light with this energy have? Note: 1eV = 1.60 x 10^{-19} joules.

2. Recalling that photon of 5.1 electron volts (eV) of energy will eject an electron from a piece of gold, what would happen if you were to shine a light of 6.5 eV on the gold surface? How is this the same or different from using light of 3.0 eV? What if the metal was Cesium (Φ = 2.1 eV) or Platinum (Φ = 6.35 eV) instead?

Wave-Particle Duality and Uncertainty

Activity 1

1. Young's double slit experiment and the observation of diffraction patterns suggest that light behaves as a ______________________.
2. The ultraviolet catastrophe outcomes and the photoelectric effect suggest that light behaves as a ______________________.
3. What does the phrase wave-particle duality mean in your own words?

Activity 2

1. What did Louis de Broglie propose about matter both large and small?
2. What equation describes his idea? Please indicate what each variable means (include units).
3. Please calculate the wavelength of a neutron traveling at the speed of light (in a vacuum).

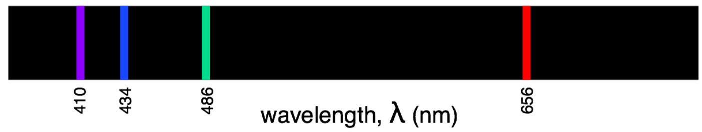
Activity 3

1. Briefly explain Heisenberg's uncertainty principle in your own words.
2. What equation describes this principle? Please indicate what each variable means (include units).
3. Assume that an atom is approximately 0.4nm in diameter and that this value is the uncertainty in position for an electron in an atom. What would the uncertainty in momentum be for this electron? What would the uncertainty in velocity be for this electron?

Line Emission Spectra and Electrons

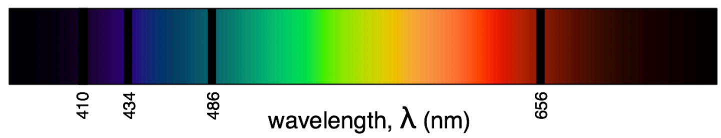
Activity 1

1. Identify the following spectra as either continuous, emission or absorption:



   a. ___________

   b. ___________



   c. ___________

2. Would you expect the line of spectra of different elements to be the same or different? Why?

Activity 2

1. What is Bohr's model of the atom?
2. What is Rydberg's equation?
3. How did the Bohr's model and Rydberg's equation each explain line spectra?

Activity 3

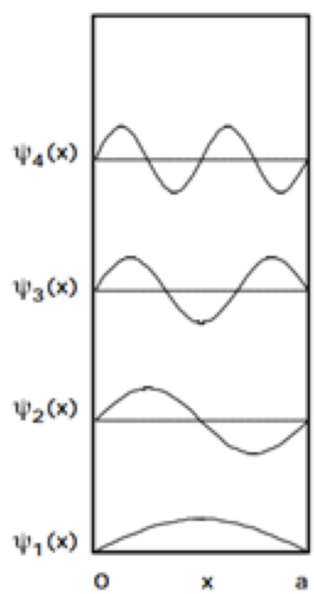
1. There are multiple transition series of the hydrogen atom. The Balmer series describes the transitions between the energy level n = 2 and energy levels 3 and above. These transitions occur in the visible range. Below are two other well-known transition series. From which energy levels do these series always originate and in which range of light do these transitions occur?
   a. Paschen –
   b. Lyman –

2. Use the Rydberg equation to calculate the energy in joules, wavelength in nm and frequency in Hz of the transition between n = 7 and n = 3 for the hydrogen atom.
3. Which absorption transition in the Balmer series corresponds to an energy difference of 3.029×10^{-19} J?
4. Rank the following five electron transitions in the hydrogen atom from highest energy to lowest energy (think of the Rydberg equation)
   2→1 4→2 4→3 5→2 5→1
5. Rank the following five electron transitions in the hydrogen atom from longest to shortest wavelength (think of the Rydberg equation)
   3→1 5→2 7→3 4→2 14→6

Quantum Mechanical Theory of the Hydrogen Atom

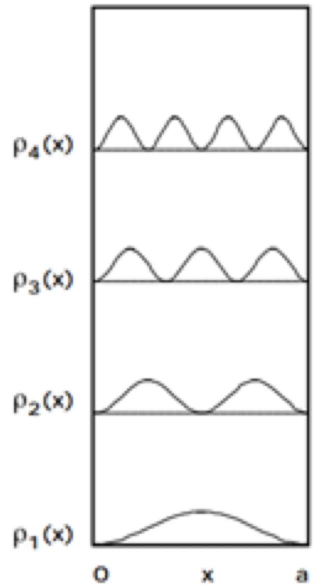
Activity 1

1. Erwin Schrödinger derived an equation that provides the infinite number of wave functions and associated energies for very small objects. One way we can begin to understand the importance of his solutions is to consider the _______________ in a _______________ model.
2. If we consider a particle in a one dimensional box, the boundary conditions for the walls is that they have _______________ potential energy. This means that the probability of finding the particle at the wall is _______________.
3. The equation for the energy of a particle in a one dimensional box is:
$$E_n = \frac{n^2h^2}{8mL^2}$$
Where E_n is the energy of the particle in joules, h is Planck's constant in joule-seconds, m is the mass of the particle in kg, and L is the length of the box in meters.
The principle quantum number n is a positive integer value (1, 2, 3, … etc.). Can the energy of the particle ever be zero inside the box? Why or why not?
Can the energy of the particle be any value inside a given box? Why or why not?
4. Using the equation for the energy of a particle in a one dimensional box, please describe the affect the changing the length of box would have on the energy of the particle.
How about changing the energy level (n)?
Which has a greater effect on the energy of the particle – halving the particle's mass or halving the length of the box?
5. By using the boundary conditions, we can plot the wave functions inside a one dimensional box with infinite walls. The particle in a box depiction below maps several wave functions and stacks these plots on top of each other. Also, ψ_3 indicates the wave function at n = 3.



http://chemwiki.ucdavis.edu/Physical_Chemistry/Quantum_Mechanics/05.5%3A_Particle_in_Boxes/3.5%3A_Quantum_Mechanics_of_Some_Simple_Systems

The probability function is simply the wave function squared (where p_3 indicates the probability function at n = 3):



http://chemwiki.ucdavis.edu/Physical_Chemistry/Quantum_Mechanics/05.5%3A_Particle_in_Boxes/3.5%3A_Quantum_Mechanics_of_Some_Simple_Systems

Is it equally likely to find the particle at any location in the box?

Areas of zero probability are called "nodes." How do the number of nodes change as n increases?

Activity 2

1. Bohr's model broke down because Newtonian physics could not account for the uncertainties related to very ________________ objects.
2. Quantum mechanics was the new physics used to model the extremely small. Match the term or person on the left with the best description on the right.

_____ Schrödinger	a. the solutions to the Schrödinger equation that when squared, gives the probability function.
_____ wave functions	b. the solutions to the Schrödinger equation other than the wave functions.
_____ quantum mechanics	c. the lowest energy solution to the Schrödinger equation.
_____ energies	d. the scientist who derived the Schrödinger equation.
_____ ground state	e. the new physics created describe physics at the atomic level.

Activity 3

1. How are the energy solutions to the Schrödinger equation related to Rydberg's solutions and Bohr's model?

2. The probability function is the wave function ________________________. By observing the probability functions we can define ________________________, or areas of electron density around the nucleus.
3. The idea of wave-particle duality tells us that electrons can behave in way similar to both ________________________ and ________________________.

Quantum Mechanical Solutions – Quantum Numbers for the H Atom

Activity 1

1. Please fill in the following table for the quantum numbers n, ℓ and m_ℓ.

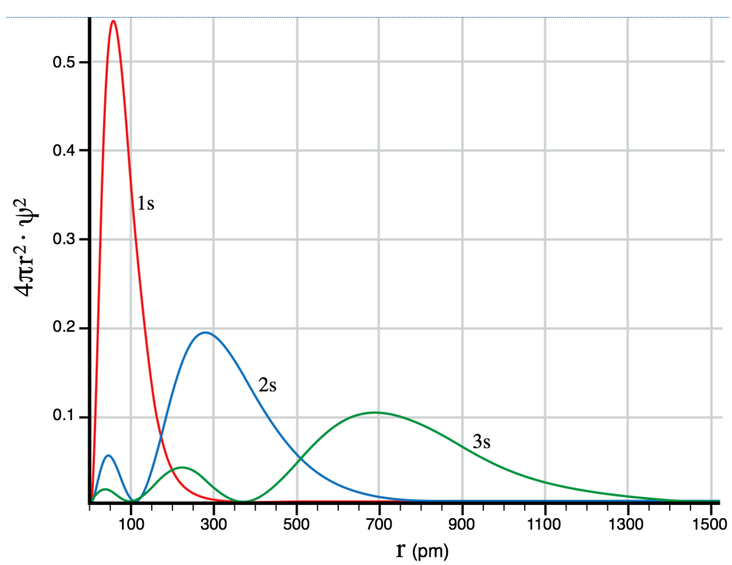
QuantumNumber	Meaning	Allowed Values
n		
ℓ		
m_ℓ		

Quantum Mechanical Solutions - Orbitals

Activity 1

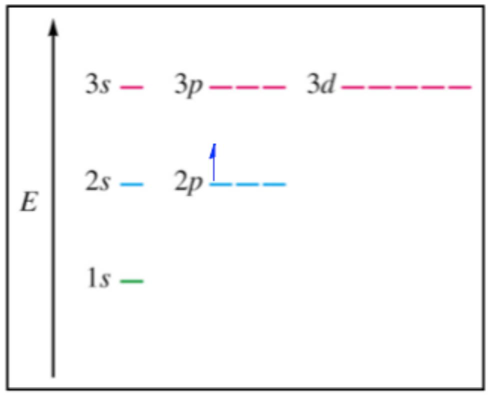
1. Make note of a word or two to describe the shapes of the s, p and d orbitals.

2. Below is an example of a radial distribution function. It is related to the probability function (wave function squared). By observing the probability functions we can define orbitals, or areas of electron density around the nucleus. Describe differences in the radial distributions of the 1s, 2s and 3s orbitals shown below.



Activity 2

1. The following orbital notation diagram is for the one electron of hydrogen. Is the electron in its ground state or is it in an excited state? How could you tell?



2. Write out the electron configuration for the state of hydrogen above and for the ground state electron of hydrogen.
3. Write a set of quantum numbers that could describe the electron shown on the diagram in question 1 above.

Activity 3

1. Which of the following sets of quantum numbers are allowed? If they are not allowed, how could you correct them?
    a. $n = 3, \ell = 2, m_\ell = -2$
    b. $n = 1, \ell = 1, m_\ell = 0$
    c. $n = 4, \ell = -2, m_\ell = 0$
    d. $n = 2, \ell = 1, m_\ell = -1$
2. Given the following quantum numbers (3, 2, -1) in what orbital will that electron be found?

3. Given the following quantum numbers (4, 3, 0) in what orbital will that electron be found?

4. Write the type of orbit described by each set of quantum numbers below (e.g. 3s, 2d, etc.).
    $n = 3, \ell = 0, m_\ell = 0$ Orbit:
    $n = 2, \ell = 1, m_\ell = 0$ Orbit:
    $n = 1, \ell = 0, m_\ell = 0$ Orbit:
    $n = 3, \ell = 2, m_\ell = -1$ Orbit:
    $n = 5, \ell = 3, m_\ell = -3$ Orbit:
    $n = 7, \ell = 2, m_\ell = 0$ Orbit:
    $n = 3, \ell = 2, m_\ell = -1$ Orbit:
    $n = 2, \ell = 1, m_\ell = 0$ Orbit:
    $n = 4, \ell = 1, m_\ell = -1$ Orbit:
5. Using quantum numbers, describe:
    a. The highest energy electron in a Boron atom.
    b. The lowest energy electron in a Xenon atom.
6. What is an element that can have a ground state electron described by: $n = 7, \ell = 1, m_\ell = 1$

7. For orbital diagrams A-D below, indicate what rule is being violated. If no rule is being violated, indicate what ground state element is described by the diagram.

	1s	2s	2p		
A.	↑↓	↑↓	↑↓	↑↓	
B.	↑↓	↑↓	↑↓	↑	↑
C.	↑↓	↑	↑↓	↑↓	↑
D.	↑↓		↑↓	↑↓	↑↓

Rules and Quantum Numbers

Activity 1

1. Match each idea on the left with the appropriate explanation or definition:

_____ Pauli Exclusion Principle	a. electrons must be placed in separate degenerate orbitals first before pairing.
_____ Hund's Rule	b. requires the use of m_s as the fourth quantum number.
_____ Aufbau principle	c. electron configurations in the ground state are built up from the lowest energy levels to higher energy levels.

2. The fourth quantum number, _______, can have the value of _______ or _______. Orbitals hold up to two electrons and because of the _______________ principle a fourth quantum number is needed.

3. Within multi-electron atoms, the fact that the 2s orbital is lower in energy than the 2p orbitals is called a "_______________________________________ within energy levels."

Activity 2

1. Identify the s, p, d and f blocks on the periodic table and write the order in which they are filled for a ground state atom (start with 1s and end with 7p)

Periodic Table of the Elements

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 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 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	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 65.41	31 Ga 69.72	32 Ge 72.64	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 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (281)	111 Rg (272)							

58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
90 Th 232.0	91 Pa 231.0	92 U 238.0	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)

2. The orbital notation of the electron configuration of cesium is $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^2$. Without looking at the periodic table, we know that cesium is in the _____ row and in the _____ block.

Ground State Electron Configurations

Activity 1

1. Write the ground state electron configuration for the following elements in three ways: the full electron configuration, the shorthand electron configuration and the orbital notation with the arrows.
   i. Beryllium
   ii. Nitrogen
   iii. Argon
2. Differentiate between the terms "subshell" and "orbital."

Activity 2

1. Write the ground state electron configuration for the following elements in three ways: the full ground state electron configuration, the shorthand electron configuration and the orbital notation with the arrows.
   i. Iodine
   ii. Thallium
2. Write a list of all the possible quantum numbers for an electron in a 4f subshell of bismuth (Bi).

Activity 3

1. Exceptions to ground state electron configurations occur in the ____ and ____ blocks when an atom is one electron shy of a ____ –filled or _________ filled subshell.
2. Write the ground state electron configuration for the following elements in three ways: the full ground state electron configuration, shorthand electron configuration and the orbital notation with the arrows.
   i. Copper
   ii. Molybdenum

Activity 4

1. Differentiate between the terms "core" and "valence."
2. The shell model of an atom is sometimes be referred to as the "planetary model." Explain why you think people use this nickname and explain what the sun and planets correspond to in the shell model.

Periodic Table and Reactivity

Activity 1

1. Consider lithium metal.
    a. Why don't we find lithium metal in its neutral atomic state in nature?
    b. In what electron configuration would you expect to find lithium metal?
    c. Comment of the similarities between your answer for part b and the electron configuration for helium.
2. Why do groups/families tend to share common chemical behaviors? In your explanation, discuss electron configurations and valence electrons.

Activity 2

1. There is a radioactive isotope of strontium that can cause bone and blood cancers. Why would strontium replace calcium in human bones (discuss valence electrons)?
2. Find a noble gas and a common ion that are both isoelectronic with $Sr^{2+}$and write down the electron configuration that these species all share.

Activity 3

1. Write down the charge the different groups tend to make when forming ions and explain why:
    a. Group IA –
    b. Group 2A –
    c. Group 5A –
    d. Group 6A –
    e. Group 7A –
    f. Why?
2. How is the effective nuclear charge of an atom different than its nuclear charge? Why is there a difference?
3. How does effective nuclear charge change as you move down a group of elements? As you move left to right across a period of elements?

Periodic Trends

Activity 1

1. As you move down a group what do we find about the ionization energy, electron affinity and atomic radii? How about as you move across a period from left to right?

2. Arrange the following groups of atoms in order of decreasing size:

    a. P, Sb, N, As –

    b. Br, V, N, Ba –

3. Which atom or ion in each of the following lists has the largest ionization energy:

    a. Mg, Se, Ba –

    b. Ca, Co, Se –

    c. O^{2-}, O, O^{2+} –

Activity 2

1. Explain how shielding affects ionization energy and atomic radii.
2. Here is a way to depict the process related to the definition of electron affinity using the element nitrogen:

   $$N\ (g) + 1e^- \rightarrow N^-\ (g)$$

   How could you depict the process related to the definition of ionization energy in a similar way using the element nitrogen?
3. Why do noble gases have exceptionally low electron affinities and exceptionally high ionization energies?
4. Write the electron configuration for silver. Try to give at least one example of an ion that would have an identical electron configuration.

Mole to Number Conversions

Activity 1

The purpose of this activity is to practice your understanding of the concept of the mole.

1. How are a mole of textbooks and a mole of water molecules similar? How are they different?
2. How is Avogadro's number related to the concept of a mole?
3. The atomic mass of carbon is approximately 12.0 amu. It represents the mass of ______ carbon atom(s).
4. 1 amu is equivalent to ______ grams.

Activity 2

The purpose of this activity is to practice your ability to convert between moles and number of atoms of elements within compounds.

1. You are given a beaker containing 27.8 moles of octane, C_8H_{18}. How many moles of carbon are present in the beaker?
2. In this same beaker, how many atoms of carbon are present?
3. In this same beaker, how many moles of hydrogen are present?
4. In this same beaker, how many atoms of hydrogen are present?

Mass to Mole conversions

Activity 1

The purpose of this activity is to practice your ability to convert between mass units and number units of elements in compounds.

1. A 109 gram sample of SF_6 contains how many moles of SF_6?
2. A 109 gram sample of SF_6 contains how many molecules of SF_6?
3. A 1.70 mole sample of Fe_2O_3 weighs how many grams?
4. This problem combines mass to mole and mole to number stoichiometry. Calculate the number of propane, C_3H_8 molecules, in 74.6 grams of propane.

Empirical and Molecular Formula

Activity 1

The purpose of this activity is for you to practice some basic compositional stoichiometry, which is based on the Law of Definite Proportions.

1. Water always has the chemical formula H_2O, regardless of its phase (solid, liquid, gas). Which law helps us explain this fact?
2. A different substance, hydrogen peroxide has the chemical formula H_2O_2. Both water and hydrogen peroxide have the same number of hydrogen atoms per molecule, but different numbers of oxygen atoms per molecule. Which law explains this fact?
3. Calculate the mass percent of each element in the compound disulfur dioxide, S_2O_2.
4. State the number of moles of each element in one mole of disulfur dioxide. What does the Law of Definite Proportions tell us about disulfur dioxide? What does the Law of Multiple Proportions imply about compounds made of sulfur and oxygen?

Activity 2

1. 32.0 grams of a compound was found to contain only carbon and hydrogen. There are 24.0 grams of carbon. What is the empirical formula of the compound?
2. Butyric acid is 54.5% carbon, 9.10% hydrogen and 36.4% oxygen. What is the empirical formula of butyric acid?
3. The molar mass of butyric acid is 88.0 g/mol. What is the molecular formula of butyric acid?

Ionic vs Covalent Compounds

Activity 1

The purpose of this activity is to practice recognizing, naming, and defining different types of compounds, namely ionic and covalent compounds. In general atoms of elements come together to form compounds. One can typically predict the type of compound based on the types of elements that have come together to form compounds. In this unit we focus on two types of compounds – ionic and covalent.

1. How are ionic and covalent compounds different? How are they similar?
2. Why do metals tend to lose electrons while nonmetals tend to gain electrons
3. Define cation, anion and isoelectronic
4. How are valence electrons involved in ionic bonding and covalent bonding differently?

Activity 2

A key skill in understanding both ionic and covalent bonding is the ability to predict the number of valence electrons in an atom.

1. Write out the electronic configuration of the following elements and state the number of valence electrons in each atom: Na, Mg, Al, Si, P, S, Cl, Ar.
2. Lewis dot symbols are a short hand method of representing the valence electrons of a particular atom of an element. Write the Lewis dot symbols for the following atoms: Na, Mg, Al, Si, P, S, Cl, Ar.
3. Write the Lewis dot symbols for the following elements: B, Al, Ga, In Tl. These elements are in group 13 (or 3A). How can the group number inform us about the number of valence electrons present in an atom?
4. Based on position on the periodic table state the number of valence electrons in Sb.

Ionic Compounds

Activity 1

The purpose of this activity is to practice your understanding of the formation of ions.

1. What does the phrase "achieving noble gas configuration" mean in the context of ionic bonding? How do metals and nonmetals achieve noble gas configuration?
2. Write the electron configuration for Sr atom and Sr cation.
3. Draw the Lewis dot structure of Sr atom and Sr cation.
5. Write the electron configuration for I atom and I ion.
6. Draw the Lewis dot structure of I atom and I ion.
7. Based on the ions that each of above elements tends to form, what is the predicted ionic compound between Sr and I.
8. Predict the formula for the ionic compound in formed between rubidium and sulfur.

Activity 2

The purpose of this activity is to practice your ability to form ionic compounds given a metal and a nonmetal and/or a polyatomic ion.

1. Predict the compound formed from magnesium and oxygen. Write the formula and correctly name the compound.

2. Calcium oxide always has the formula CaO. Manganese oxides however can take many different forms such as MnO, Mn_2O_3 and MnO_2. Please briefly explain why there are more manganese oxide compounds than calcium oxide compounds. In your answer include the terms "main group elements" and "transition metals."
3. Predict the compound formed between the ammonium ion and sulfur. Write the formula and correctly name the compound.
4. Predict the compound formed between sodium and permanganate ion. Write the formula and correctly name the compound.

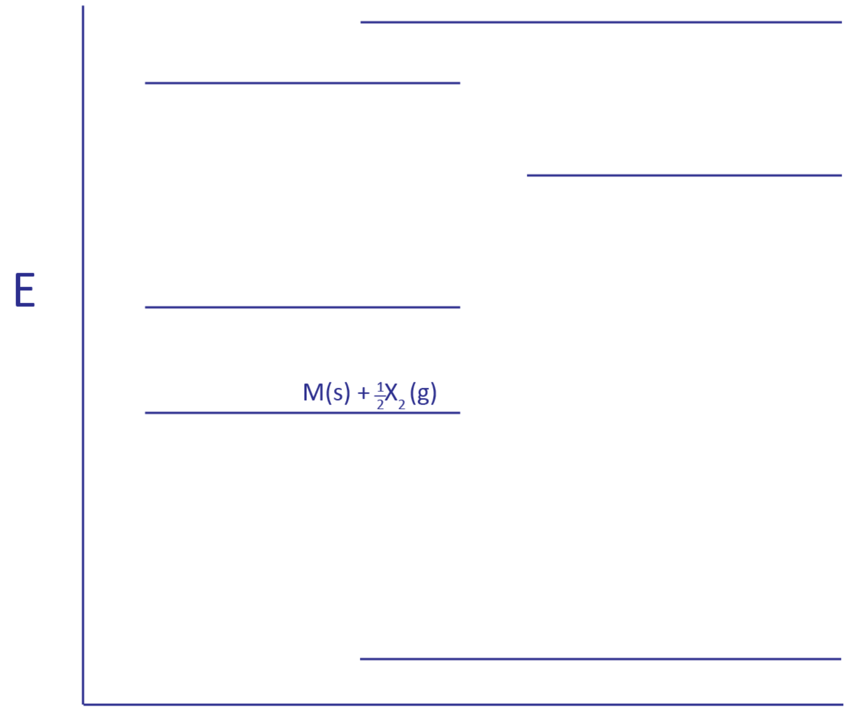
Lattice Energy

Lattice energy is defined as the amount of energy necessary to dissociate 1 mole of an ionic crystalline solid into its gas phase ions. Complete this activity to deepen the concept that different ionic compounds have different lattice energy.

Activity 1

In this activity, we will practice our understanding of ionic compounds and different lattice energies associated with ionic compounds.

1. Write the balanced chemical reaction for making a solid ionic compound with some alkali metal that is solid at room temperature and pressure (M) and some halogen that is a gas at room temperature (X_2).
2. Ionization energy values are positive because the process of ionization ________ energy.
3. Electron affinity values can either be negative or positive. The electron affinity for adding a single electron to neutral fluorine is large and negative. Conversely, the electron affinity for adding a single electron to neutral sodium is close to zero. Comment on why you think this is the case.
4. In this class we will consider ionization energy values and electron affinity values for compounds in the ________ phase.
5. Consider the alkai metal cation in the gas phase (M^+) and the halogen anion in the gas phase (X^-). What do you think most likely will happen if you brought those two ions together – will they have lower potential energy together or apart?
6. How does having a greater number of anions and cations affect the overall change in energy when going from separated gaseous ions to a solid crystal lattice?
7. Label the following energy diagram with the necessary steps to form a crystal lattice with an alkai metal (M) and a halogen (X). Please draw arrows between each subsequent step to show how the energy is changing. The first step has been filled in for you



Activity 2

1. Lattice energy is the amount of energy necessary to dissociate _________ of an ionic crystalline _________ into its _________ phase ions.
2. Order the following ionic compounds from lowest lattice energy to highest lattice energy: KF, RbF, LiF, CsF, NaF.
3. Rank the following salts from greatest lattice energy to least: $Fe_2(CO_3)_3$, FeI_3, $AlPO_4$, MgF_2, $MgCO_3$, AlN, NaI, MgO
4. How does ionic radius affect lattice energy? How does the magnitude of the charge affect lattice energy? Please include Coulomb's Law in your answer.
5. Aluminum hydroxide has a melting point of approximately 300˚C. Aluminum oxide has a melting point of approximately 2,072˚C. Why do you think these two ionic compounds have such different melting points?

Naming Ionic Compounds

Activity 1

Use this activity to practice your knowledge of recognizing formulas and naming ionic compounds. Follow up with creating your own flashcards to make sure you have memorized all the required polyatomic ions. This short activity is just a sample of the types of questions that could be asked on the quiz or the exam.

1. $PbSO_3$
2. $NiBr_3$
3. CaC_2O_4
4. KF

Activity 2

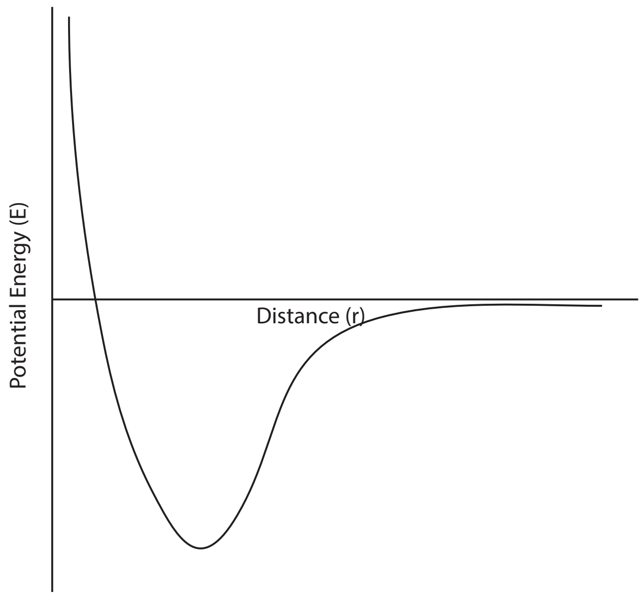
Use this activity to practice your knowledge of writing formulas for ionic compounds. Follow up with creating your own flashcards to make sure you have memorized all the required polyatomic ions. This short activity is just a sample of the types of questions that could be asked on the quiz or the exam.

1. Ammonium iodide
2. Vanadium (V) sulfide
3. Silver nitride
4. Barium phosphite

Covalent Compounds

ACTIVITY 1

1. The following potential energy diagram describes the potential energy for two nonmetal atoms. What do the axis labels mean? What does the diagram indicate about how these two atoms interact?



2. On the potential energy diagram above, draw a line from the x-axis to the graph that indicates the bond strength between the two atoms. Then, draw a line from y-axis to the graph that indicates the bond length between the two atoms. Label the area on the diagram that is dominated by attractive forces and the area that is dominated by repulsive forces.
3. When discussing the strength of covalent compounds we refer to the ________ ________ between atoms. When discussing the strength of ionic compounds we refer to the ________ ________ of the compounds.
4. Which carbon-carbon bond would you expect to be the strongest: a single, double or triple bond? The shortest?

Naming Covalent Compounds

ACTIVITY 1

Use this activity to practice your knowledge of recognizing formulas and naming covalent compounds. Follow up with creating your own flashcards to make sure you have memorized all the required prefixes and element names. This short activity is just a sample of the types of questions that could be asked on the quiz or the exam.

1. PCl_3
2. S_2O_5
3. N_2O_4
4. $SeCI_4$

ACTIVITY 2

Use this activity to practice your knowledge of writing formulas for covalent compounds. Follow up with creating your own flashcards to make sure you have memorized all the required prefixes and element names. This short activity is just a sample of the types of questions that could be asked on the quiz or the exam.

1. Dihydrogen monoxide (also known as water!)
2. Nitrogen monoxide
3. Iodine trichloride
4. Tetraphosphorus hexoxide

Line Drawings

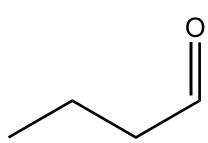
ACTIVITY 1

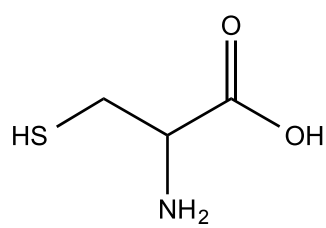
Chemists have different ways of representing compounds including systematic names, common names, chemical formulas, condensed structural formulas, and line drawings. The purpose of this activity is to practice how to interpret line drawings.

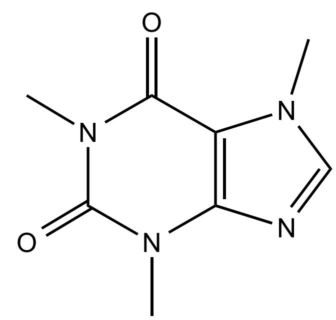
1. Carbon atoms are represented as corners in line drawings. Every carbon must have _____ bonds.
2. Hydrogen atoms are often not drawn in line drawings. How do you know if an implicit hydrogen is present?

ACTIVITY 2

Look at the following line drawings and write the associated chemical formula.







(This is caffeine)

The Polar Covalent Bond

The definition of the covalent bond is that electrons are shared between two atoms, whereas the definition of an ionic bond is that the electron(s) is donated from one atom to the other and the atoms are bound by a strong electrostatic attraction. In reality all chemical bonds exist between these two extremes.

ACTIVITY 1

1. Classify these bonds as pure covalent or polar covalent:
   - F–H
   - F–F
   - H–H
   - Br–I
   - O–O

2. How is an ionic interaction between two atoms different than a covalent interaction between to atoms?
3. Which element is the most electronegative? Of the polar covalent bonds in question #1, which has the greatest difference in electronegativity? Why does F–F not make the list of polar covalent bonds?
4. How are dipole moment and a difference in electronegativity connected? How do we conventionally draw a dipole moment for a bond?

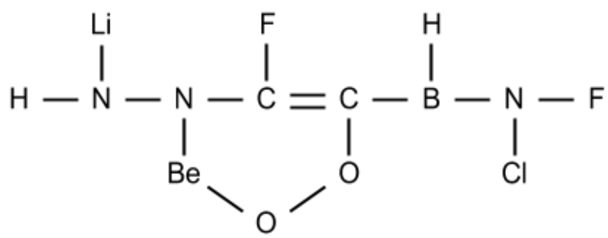
ACTIVITY 2

Use the concept of electronegativity and ΔEN to rank the polarity of the following bonds.

1. F–Si, F–S, F–Cl, F–P
2. Se–Se, Se–S, Se–O
3. C–N, C–F, C–C, C–O
4. N–F, N–N, N–Sb, N–P
5. What is the electronegativity of each of the following elements?

H	B	O	Li
F	C	Be	N

6. For the (completely fictitious) molecule below, calculate the ΔEN of each bond and determine whether it is ionic or covalent based on simple ΔEN.



Lewis Structures & Octet Rule

Chemists follow several guidelines when visually representing covalent compounds. The conventional method of drawing structures produces Lewis structures. The following activities will insure that you practice the skills necessary to draw and interpret Lewis structures.

Activity 1

1. In a conventional Lewis structure a covalent bond is represented as a _______. A lone pair is represented as a _______ of _______. A free radical is represented as a single _______.
2. An important skill in predicting the correct Lewis structure of a molecule is to be able to count up the total number of valence electrons on a molecule. State the total number of valence electrons on the following molecules: C_3H_8, SO_2, I_2, PH_3 and CH_3OH.
3. What is the octet rule and how does it help us draw Lewis structures?
4. State a rule that *might* help when beginning to draw Lewis structures (everyone learns differently so this rule may or may not be helpful to you).

Activity 2

The purpose of this activity is to use the rules of Lewis Dot Structures to predict the structures of some common molecules. It's important to remember that drawing Lewis structures is a process that sometimes requires trial and error. In later activities more skills will be added so that you can check the viability of your structure.

1. Draw the correct Lewis structures for the following molecules: OF_2, NH_2,I and CH_3Cl.
2. Draw the correct Lewis structures for the following molecules: N_2, N_2H_2 and N_2H_4.
3. How do structural formulae help us draw Lewis structures? Compare and contrast these structural formulae: CH_3OCH_3 and C_2H_5OH.

Formal Charge

Activity 1

When constructing Lewis structures there is often more than one possible way to put atoms together to form a molecule. Calculating the formal charge on each atom is an important method used to check the viability of your structure.

1. How is formal charge calculated?
2. Determine the formal charge on all the atoms in NH_3.
3. Determine the formal charge on all the atoms in CF_3CH_2OH.
4. Draw the following Lewis structures (Hint: the Lewis structures look a lot like how they're written.)
   a. H_3CCOOH (acetic acid)
   b. H_2NCONH_2 (urea)
   c. $H_3CH_2COCH_2CH_3$ (diethyl ether)

Activity 2

1. What are the formal charge characteristics for the most stable structure of a compound?
2. Draw two different structures for hydroxylamine, which has the chemical formula NOH_3. Then determine the formal charge on each atom in the structure and based on that calculation explain which is the better structure.
3. Suggest a viable Lewis structure for CH_3CN. Label the formal charge on each atom and use formal charge arguments to justify your Lewis structure.

ACTIVITY 3

1. What some examples of common polyatomic ions?
2. Draw the Lewis structure for the hydronium cation. Calculate the formal charge on each atom in the ion. Confirm that the sum of the FC's on all the atoms adds up to equal the charge on the polyatomic ion.
3. Draw the Lewis structure for hydroxide anion. Calculate the formal charge on each atom in the ion. Confirm that the sum of the formal charge of each atom in the ion equals the charge on the polyatomic ion.

Resonance Structures

Compounds that exhibit resonance have structures that are the averaged combination of multiple drawn structures. One way to recognize if a compound is capable of resonance is when you can draw equivalent Lewis structures with a double bond placed in different locations.

Activity 1

1. An incorrect way to describe resonance might be to say that a double bond is flipping back and forth between different atoms. Why is this an incorrect interpretation of resonance and what is a better description?
2. Benzene is a molecule that exhibits two resonance structures, however, the bond energies of all the carbon-carbon bonds within the benzene ring in are the same. Explain the relative strength of these bonds.

Activity 2

1. Draw the Lewis structure of the polyatomic anion, carbonate, CO_3^{2-}. Be sure to show all possible resonance structures. Comment on the relative bond strength of all the bonds in this ion. Discuss the formal charge on each atom in the averaged structure.
2. Draw the Lewis structure, including all resonance structures, for dinitrogen monoxide, N_2O.
3. Draw the Lewis structure of NO_2. What two features of this molecule are interesting?
4. Draw 3 resonant Lewis structures for OCN^-. Determine which makes the greatest contribution by assigning formal charge.

Lewis Structures & Exceptions to Octet Rule

Activity 1

Lewis structure rules don't always work. In some cases an atom can be stable with fewer than 8 electrons in the valence shell. This is called an incomplete octet.

1. Which elements have an incomplete octet?
2. Draw the correct Lewis structure for BeF_2.
3. Draw the correct Lewis structure for BH_2Cl.

Activity 2

Lewis structure rules don't always work. In some cases an atom can be stable with more than 8 electrons in the valence shell. This is called an expanded octet. If you use the S = N–A rule you can often predict that the octet rule is going to break down and will be expanded when the rule predicts fewer bonds than are necessary for the number of atoms in the molecule to be linked together.

1. Explain why nitrogen cannot have an expanded octet but phosphorus can.
2. Draw the Lewis structure for the XeO_4.

3. Draw the Lewis structure for ClF_3.
4. Draw the Lewis structure for SO_2Cl_2.
5. Draw the Lewis structure for SO_4^{2-}

Activity 3

Lewis structure rules don't always work. In some cases the sum of all the valence electrons in the molecule adds up to be an odd number. In this case you are going to have a single electron on the molecule. This is called an unpaired electron.

1. Which is more reactive – an unpaired electron or pair of electrons? What is another name for an unpaired electron?
2. The hydroxyl radical contains one hydrogen atom and one oxygen atom. It has a single unpaired electron on the oxygen. Draw the hydroxyl radical structure.
3. Draw the Lewis structure for the hydroxide anion (again!). Notice how it is different from the hydroxyl radical.

Electronic Geometry & VSEPR

The Valence Shell Electron Repulsion Model or VSEPR is an empirical model used to predict the three dimensional structure of molecules based on the notion that like charges repel. One can predict VSEPR geometries around central atom(s) in molecules based on the numbers of regions of high electron density around the central atom.

Activity 1

1. The electronic geometry around a central atom is determined by the number of _________ of _________ _________ _________ around the central atom.
2. A single region of high electron density can be a _________ bond, _________ bond, _________ bond or a(n) _________ pair of electrons.
3. As the number of regions of high electron density increases around a central atom the degree measurements of the bond angles around the central atom _________.
4. An unshared pair of electrons is considered to reside _________ to the central atom than a bonding pair of electrons.
5. What are the basic electronic geometries and their associated bond angles?

Activity 2

1. Draw a generic Lewis structure for a molecule that has linear electronic geometry around the central atom. Then, provide two examples of molecules that have linear electronic geometry.
2. Draw a generic Lewis structure for a molecule that has trigonal planar electronic geometry around the central atom. Then, provide two examples of molecules that have trigonal planar electronic geometry.
3. Draw a generic Lewis structure for a molecule that has tetrahedral electronic geometry around the central atom. Then, provide two examples of molecules that have tetrahedral electronic geometry.
4. Draw a generic Lewis structure for a molecule that has trigonal bipyramidal electronic geometry around the central atom. Then, provide two examples of molecules that have trigonal bipyramidal electronic geometry.
5. Draw a generic Lewis structure for a molecule that has octahedral electronic geometry around the central atom. Then, provide two examples of molecules that have octahedral electronic geometry.

Molecular Structure

The purpose of the following activities is to develop a solid understanding of the different molecular shapes that are associated with the basic electronic structures.

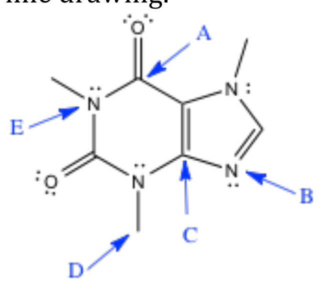
Activity 1

1. Why are there more molecular geometries than electronic geometries? When would a molecule have the same electronic and molecular geometry?
2. Lone pairs are not shared between two positive nuclei and they reside _______ to the central atom. Therefore, lone pairs take up more available bonding space around the central atom and cause bond angles to deviate from their electronic geometry predicted values. The angles between a lone pairs and a bonding regions will be ________ than predicted while the angles between the bonding regions will be ________ than predicted.

3. Which molecule do you expect to have a greater deviation in its bond angles from the electronic geometry predicted values: PCl_3 or SCl_2? Why?

Activity 2

1. State the correct molecular geometry and draw each molecule showing proper bond angles for the following molecules: BeH_2, SO_3, $SiBr_4$, PF_3, SF_4, and BrI_5. Include in your drawing dashes and wedges when appropriate to indicate when the bond might be in front of or behind the plane of the paper.
2. State the electronic and molecular geometries of the following molecules, and draw the molecule indicating with the drawing and labeling the correct bond angles: CH_2Cl_2, SO_3^{2-}, O_3 and ICl_4^-.
3. State the electronic and molecular geometry around all the atoms labeled on the following line drawing.



Molecular Polarity

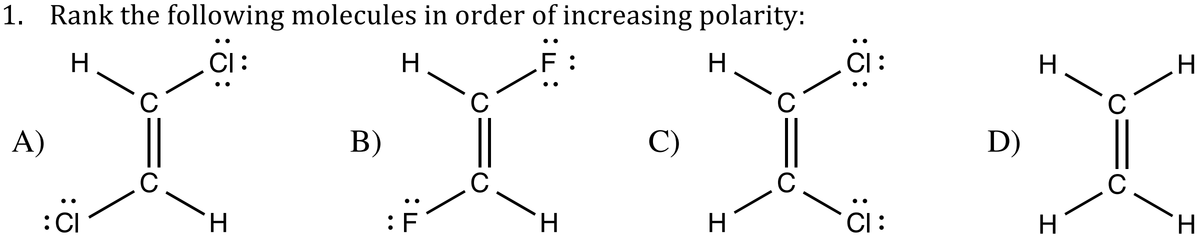
The polarity of a molecule depends on whether or not the molecule contains polar bonds, and if so, whether or not the molecule is symmetric. Being able to predict polarity is extremely important in understanding the physical and chemical properties of molecules.

Activity 1

1. A polar molecule has a permanent _____ _________ _________.
2. How can a molecule with polar bonds ever be nonpolar? Explain your answer and give an example of such a molecule.
3. Which is more important for understanding molecule polarity: the electronic geometry of a molecule or the molecular geometry of a molecule? Why?

Activity 2

1. Rank the following molecules in order of increasing polarity:



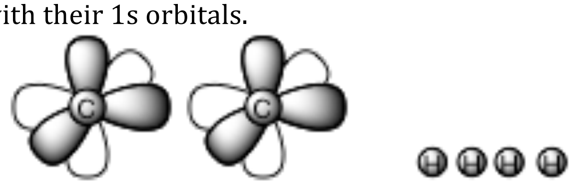
2. In the previous activities you have drawn the structures and stated the molecular geometry of the following molecules. Redraw all the following molecules and indicate if they are polar or nonpolar. Draw on the image the direction of the net dipole moment of the molecule – if it is polar: $SeCl_4$, BeF_2, BrI_5, SO_3, PF_3, $SiBr_4$, ICl_4^-, SO_3^{2-}, CH_2Cl_2 and O_3.

Valence Bond Theory

Activity 1

1. Below is a carbon atom depicted with its valence orbitals: the 2s orbital (sphere in the center) and the three 2p orbitals (three "dumbbells"). To the right of the carbon atom are four hydrogen atoms with their 1s orbitals.

   Please draw a methane molecule (CH_4) that shows how these orbitals mix and overlap. Indicate which orbitals are hybrid atomic orbitals and which orbitals are pure atomic orbitals.
2. Describe the bonds in methane. Which orbitals overlap and what name(s) does valence bond theory gives to the bonds?
3. Explain how carbon, with only 2 unpaired electrons, can form 4 bonds to fulfill its octet. Account for any energy changes (if you're putting in energy, where is it then released). What kind of hybrid orbitals will carbon then form?
4. Below are two carbon atoms depicted with their valence orbitals: the 2s orbital (sphere in the center) and the three 2p orbitals (three "dumbbells"). To the right of the carbon atoms are four hydrogen atoms with their 1s orbitals.



   Please draw an ethene molecule (C_2H_4) that shows how these orbitals mix and overlap. Indicate which orbitals are hybrid atomic orbitals and which orbitals are pure atomic orbitals.
5. Describe the bonds in ethene. Which orbitals overlap and what name(s) does valence bond theory gives to the bonds?
6. Name five common hybrid orbitals that you will encounter in this class.

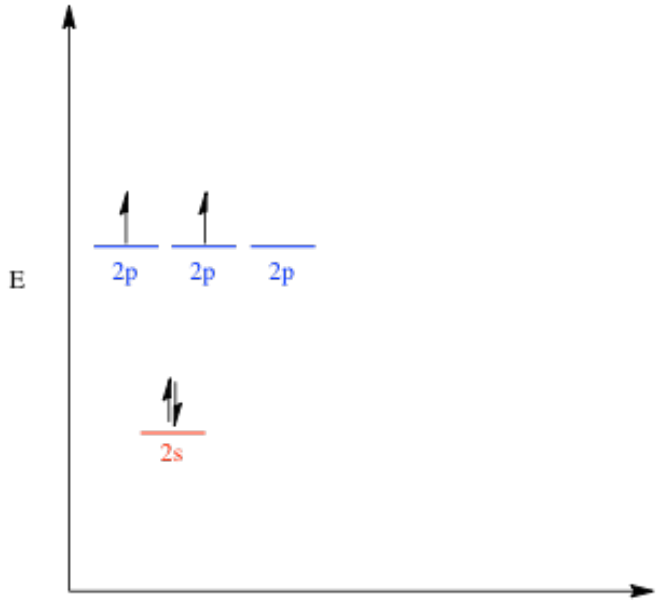
Activity 2

1. How does knowing the electronic geometry around an atom help you predict the hybridization that atom?
2. Any molecule with a linear electronic geometry will always have which type of hybridzation around the central atom?
3. An atom with sp^3d hybridization has which type of electronic geometry?

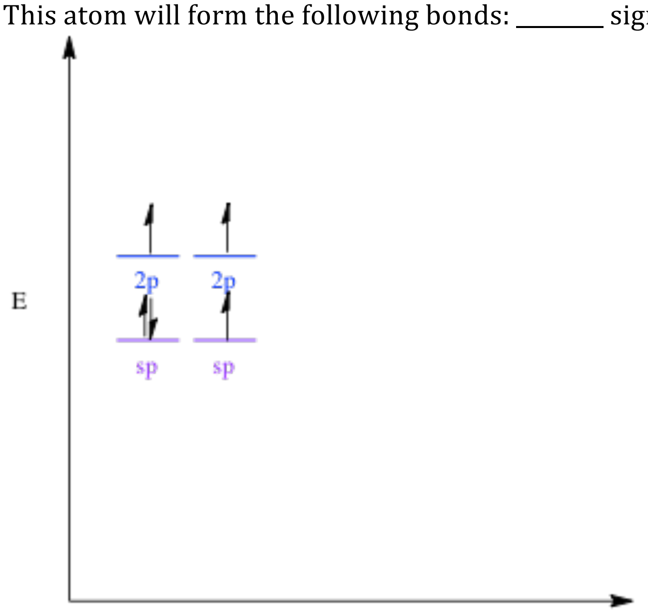
Activity 3

1. Consider the molecule ammonia, NH_3. According to the valence bond theory, what kinds of orbitals exist around the nitrogen atom? Discuss the relative energies of the orbitals around the nitrogen atom.
2. Shown below is an energy diagram for the valence atomic orbitals of some atom. In certain molecules, this atom forms three sigma bonds and one pi bond. What element is this atom?

Please draw the correct hybridized orbitals and any remaining pure atomic orbitals on the energy level diagram below showing their correct relative energies.
This atom is: ____________



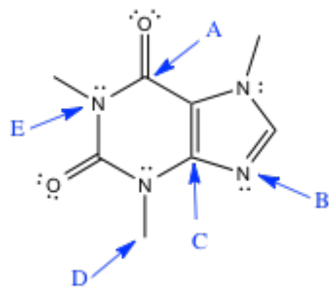
3. Shown below is an energy diagram for the hybridized and pure atomic orbitals of some atom. What element is this atom? What is its hybridization state? What kind of bonds will this atom form? Please draw the correct valence atomic orbitals on the energy level diagram below showing their correct relative energies.
This atom is: ____________
The hybridization state is: ____________
This atom will form the following bonds: ______ sigma bond(s) and ______ pi bond(s)



Activity 4

The purpose of this activity is to practice your understanding of valence bond theory.

1. Draw the Lewis structure of ethyne (actetylene), which has the molecular formula C_2H_2. Predict the hybridization of each C in the molecule. State the number of sigma and pi bonds in the molecule.
2. What is the hybridization of C in the molecule, CH_2O? State the number of sigma and pi bonds.
3. Predict the electronic geometry, molecular geometry, bond angles, hybridization, number of σ and π bonds around each atom with an arrow.



4. For formic acid, HCOOH:
   Draw the Lewis structure
   Use VSEPR to determine the geometry
   Identify bond angles
   Identify hybrid orbitals
   How many sigma and pi bonds? (Hint: There is an –OH group.)

5. Draw the Lewis dot structure for the given compounds. How many sigma bonds? How many pi bonds? Name the contributing orbitals.

   CH_3COCH_3 (Acetone) SF_4

Molecular Orbital Theory

Activity 1

The purpose of the following activity is to construct a further understanding of the quantum mechanical model of bonding called molecular orbital theory. While the valence bond theory is somewhat intuitive – the MO theory is more abstract and difficult to conceive mathematically. However, MO proves to be a more powerful and widely used modeling tool for predicting the lowest energy conformation of bonding and molecular shape.

1. Draw a Lewis structure for diatomic oxygen, O_2. How is this structure misleading when considering the paramagnetic properties of molecular oxygen?
2. The valence bond theory is considered a ____________ bonding model, while the molecular orbital theory is considered a ____________ bonding model. Valence bond theory allows us to look closely at individual ____________ while molecular orbital theory allows us to better consider the entire ____________.
3. Combining two atomic orbitals results in ______ molecular orbitals. Describe the energy of the molecular orbitals compared to the atomic orbitals.
4. There are two types of orbital interference that can occur when two atomic orbitals overlap. Which kind of interference gives rise to molecular anti-bonding orbitals? Which kind of interference gives rise to bonding molecular orbitals?

Activity 2

A result of the MO treatment of a molecule is a molecular orbital energy diagram. The MO energy diagram is analogous to the energy level diagram of an electron configuration of pure atomic orbitals. Once the MO diagram is generated from the model, all the electrons in the molecule are distributed into the MO diagram following the same rules as for atomic electron configurations. That is electrons fill from the lowest energy up, while following Hund's rule and the Pauli exclusion principle.

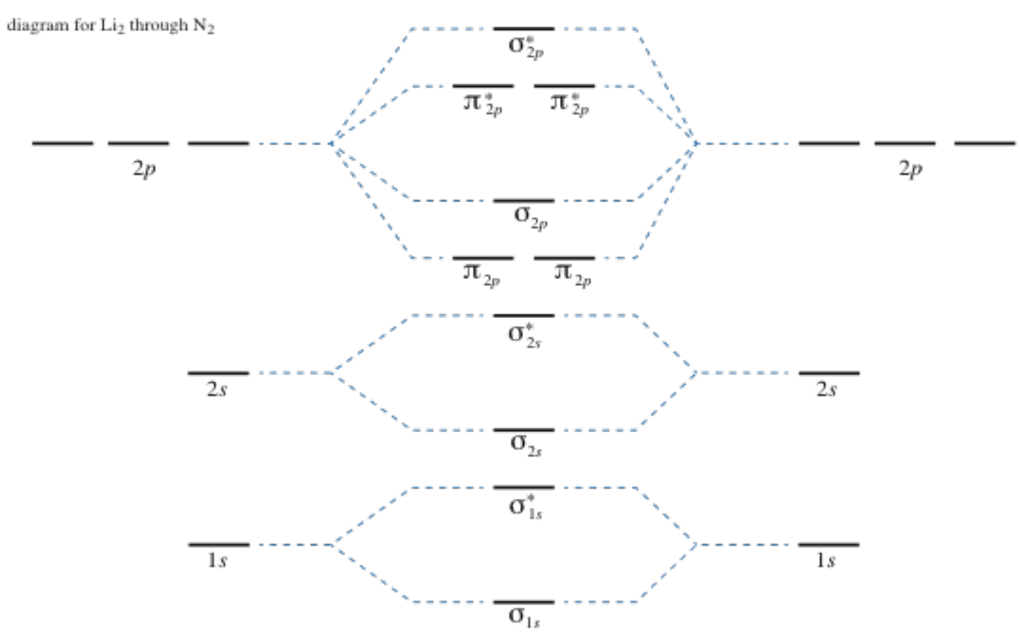
1. Fill in the following molecular orbital diagrams for some second row diatomics (B_2–F_2):

MO Diagram for B_2:

Bond Order of B_2:

Paramagnetic or diamagnetic:

HOMO and LUMO:

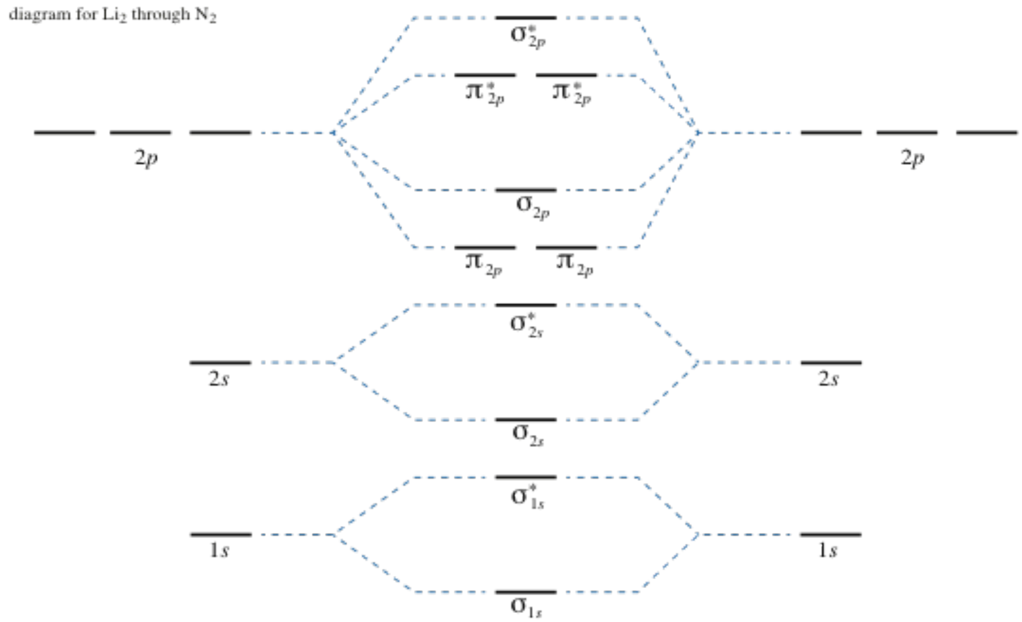


MO Diagram for C_2:

Bond Order of C_2:

Paramagnetic or diamagnetic:

HOMO and LUMO:

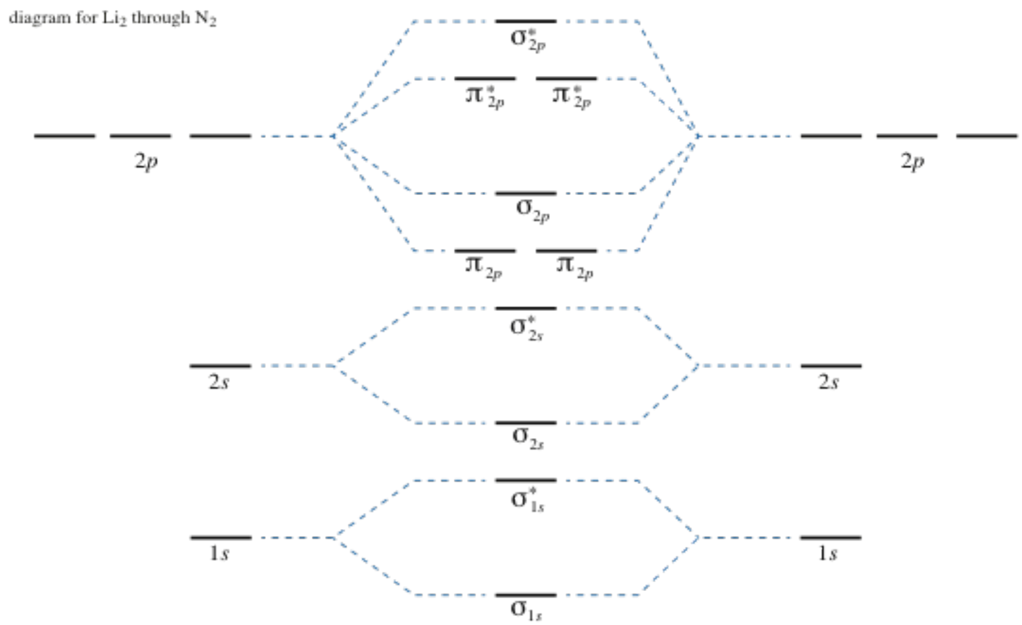


MO Diagram for N_2:

Bond Order of N_2:

Paramagnetic or diamagnetic:

HOMO and LUMO:

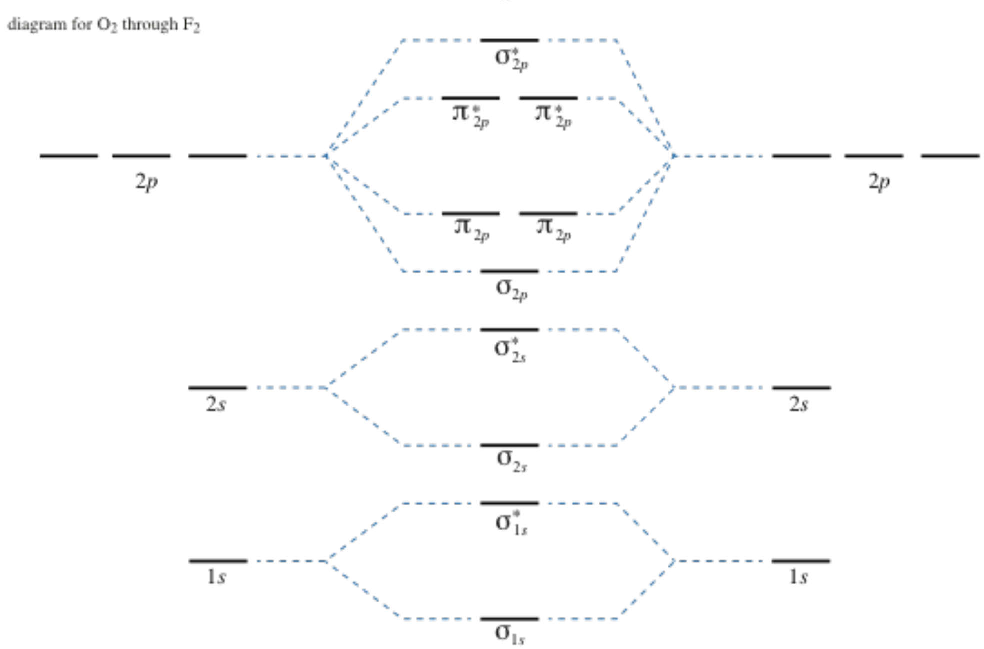


MO Diagram for O_2:

Bond Order of O_2:

Paramagnetic or diamagnetic:

HOMO and LUMO:

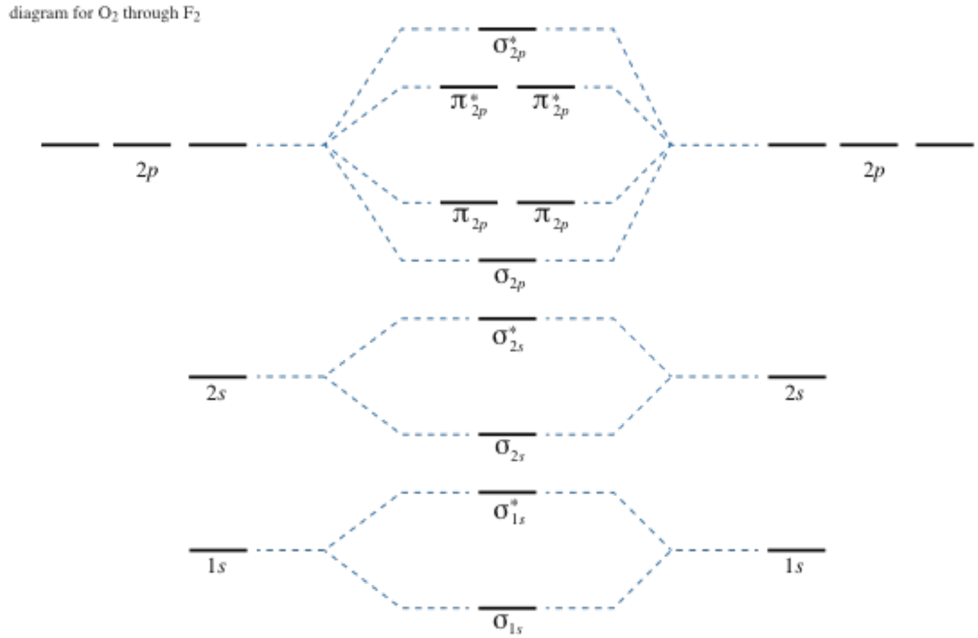


MO Diagram for F_2:

Bond Order of F_2:

Paramagnetic or diamagnetic:

HOMO and LUMO:



2. Sketch two molecular orbital diagrams to show why two hydrogen atoms when brought close enough together form a single bond but two helium atoms do not. Calculate the bond order of H_2 and He_2.

Activity 3

The purpose of this activity is to think broadly about bonding theories and to recognize that different theories have value in interpreting and predicting chemical properties and change.

1. Which bonding theory is most helpful when tracking the breaking and making of bonds across a chemical change? Why is this important in chemistry?
2. The MO theory is combined with the VB theory when thinking about ____________ in organic molecules. How do these two theories build on one another to explain this phenomenon?
3. How do the electrical conductive properties of a material relate to the relative localization of its electrons? Which bonding theory helps us explore electrical conductivity on a molecular level?

Balancing Chemical Reactions

Activity 1

The purpose of this activity is to check your understanding of the concept Law of Conservation of mass and apply that to balancing chemical equations.

1. How does the Law of Conservation of Mass help us balance chemical equations?
2. Balance the following equation:
   __$C_6H_{12}O_6$ + __O_2 → __CO_2 + __H_2O
3. Balance the following equation:
   __Co_2S_3 + __H_2 → __Co + __H_2S
4. Balance the following equation:
   __H_3PO_4 + __$Ca(OH)_2$ → __H_2O + __$Ca_3(PO_4)_2$

Activity 2

The purpose of this activity is to check if you remember how to name compounds!

1. Covalent compounds contain __________. Ionic compounds often contain a __________ and a __________ although sometimes they are formed using one or more of the __________ ions that a good student of chemistry should be able to recognize.
2. Name each of the compounds found in the chemical equations in question 3 in the previous activity.
3. Write down and balance this chemical reaction: magnesium nitride reacts with water to give magnesium hydroxide and nitrogen trihydride (ammonia).

Mole to Mole Conversions

Activity 1

The purpose of this activity is to check your understanding of the concept of mole-to-mole conversions across a chemical change.

1. How many moles of carbon dioxide are formed when 3.7 moles of C_3H_8 reacts with excess oxygen to form CO_2 and H_2O according to the following balanced chemical reaction:
   $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

Activity 2

The purpose of this activity is to check your ability to convert between moles of one reactant given and moles of other reactant needed.

1. How many moles of Fe are necessary to react completely with 0.91 moles of O_2 given the following balanced chemical reaction:
   $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
2. For the same balanced equation and assuming you have plenty of iron, how many moles of O_2 would be necessary to form 22 moles of Fe_2O_3?

Activity 3

The purpose of this activity is to check your ability to determine the limiting reactant and from there correctly predict the amount of product formed.

1. Given 2.1 moles of Al_2S_3 and 27 moles of H_2O how many moles of H_2S would be formed according to the following balanced chemical equation:
   $Al_2S_3 + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2S$

Mass to Mass conversions

ACTIVITY 1

The purpose of this activity is to check your ability to convert between mass of one reactant given and mass of product formed.

1. How many grams of water are formed when 63.0 g of C_3H_8 reacts with excess oxygen to form CO_2 and H_2O according to the following balanced chemical reaction:
   $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
2. Assuming you have plenty of iron, how many grams of O_2 would be necessary to form 3.7 g of Fe_2O_3 according to the following balanced chemical equation:
   $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

ACTIVITY 2

The purpose of this activity is to check your ability to convert between mass of reactant given and mass of other reactant needed to form product.

1. How many grams of Al_2S_3 are necessary to react completely with 29.2 g of H_2O given the following balanced chemical reaction:
   $Al_2S_3 + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2S$

ACTIVITY 3

The purpose of this activity is to check your ability to determine the limiting reactant and from there correctly predict the amount of product formed.

1. Given 28.7 g of Al_2S_3 and 83.9 g of H_2O how many grams of $Al(OH)_3$ would be formed according to the following balanced chemical reaction:
   $Al_2S_3 + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2$

Percent Yield

Activity 1

The purpose of this activity is for you check your understanding of percent yield.

1. The actual yield of a particular reaction was 67.50 grams whereas the theoretical yield was computed to be 89.88 g. What is the percent yield?
2. Imagine that you walk into the first day of chemistry lab and your TA informs you that you will be calculating the percent yield of a compound produced in a reaction. She explains that you should be able to provide a theoretical yield within five minutes and by the end of the lab session you should be able to provide the actual yield as well as the percent yield. As an excellent chemistry student you readily agree and start preparing the theoretical yield, however your lab partner thinks you should take a shortcut and just provide all three numbers now. Explain to your lab partner how you would provide the values for theoretical, actual and percent yield within your TA's time limits.

Activity 2

1. The percent yield of CO_2 for the reaction $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ is 68%. What mass of CO_2 is expected from the reaction of 6.5 g of propane (C_3H_8) with an excess of oxygen? What mass of CO_2 did the reaction produce in this experiment?
2. Your lab partner from the previous activity (problem #2) is upset that your reaction didn't have a 100% percent yield. Explain to your lab partner why a 100% yield would be surprising for most reactions.

Gas Laws

Activity 1

The purpose of this activity is to practice your understanding of Gas Laws.

1. Test yourself: how many of the gas law equations can you write down?
2. Torr, mmHg, atm, bar and Pa are all units of gas __________, which is the ratio of the combined __________ of the gas particle impacts and the surface __________ of the gas container.
3. Robert Boyle studied the relationship between __________ and __________ of a fixed amount of gas at a constant __________. Explain Boyle's law with words and an equation.
4. Jaques Charles studied the relationship between __________ and __________ of a fixed amount of gas at a constant __________. Explain Charles' law with words and an equation.
5. Lord Kelvin developed an absolute temperature scale called the __________ scale. It defines an absolute zero point at which substances have their minimum value of thermal __________ and at which pure substances exist as perfect __________.
6. The combined gas law combines which two gas laws? Explain the combined gas law with words and an equation.
7. Boyle's law, Charles' law and the combined gas law can be written as two state laws as well as one state laws. What is meant when we discuss the "state" of a gas sample?
8. Assuming a constant molar quantity of gas, how could you produce the following effects?
    a. Decrease pressure
    b. Decrease volume
    c. Increase pressure
    d. Increase volume

Activity 2

The purpose of this activity is to practice your mastery of the quantitative nature of gas laws.

1. A gas occupies 11.2 liters at 0.860 atm. What is the pressure when the volume is 15.0 liters? Assume that the temperature and amount of gas remain the same.
2. You're working on your chemistry homework with a friend. Your friend considers a problem describing a gas that is cooled from 120˚C to 40˚C. Your friend assumes that the volume must have decreased by a factor of three from 1.50L down to 0.50L. Is your friend correct? Why or why not?
3. A gas occupies 60.0 mL at 33˚C. What *change* in volume does this gas experience if it is cooled down to 5.00˚C? Assume that the pressure and amount of gas remain constant.
4. A 3250 mL gas sample at 24.5˚C has a pressure of 1825 mmHg. You change the temperature of the gas. The new volume is 4250 mL and the new pressure is 1.50 atm. What is the new temperature?

Activity 3

The purpose of this activity is to investigate Avogadro's law and start developing a small particle model of gases.

1. Amedeo Avogadro studied the relationship between __________ and __________ of a gas at a constant __________ and at a constant __________. Explain Avogadro's law with words and an equation.

2. How does a small particle simulator illustrate the microscopic behavior of gases? In your explanation, describe how are gases represented and the observance of the gas laws.
3. Imagine that you are an argon atom inside a sealed balloon filled with only argon gas. Describe your behavior and how your behavior affects the macroscopic measurements of pressure and volume. Describe what happens macroscopically and microscopically to the pressure and volume if the balloon begins to leak.

Ideal Gas Law

Activity 1

The purpose of this activity is to practice your understanding of the Ideal Gas Law

1. How is the ideal gas law different than the other gas laws previously discussed?
2. Explain the ideal gas law with words and an equation.
3. Write down different values for *R* that you could use when pressure is in atm and volume is in liters, when pressure is in torr and volume is in liters, when pressure is in Pa and volume is in m^3, and when you need to calculate joules.
4. What do we assume about ideal gases?
5. You are scuba diving in a large fish tank. While you are at the bottom of the tank, you release a balloon full of air and watch it as it rises to the surface. What do you notice about the volume of the balloon?

Activity 2

The purpose of this activity is to practice your understanding of the quantitative side of the Ideal Gas Law.

1. Give the different pressure and temperature conditions for both STP and SATP for gases.
2. What does the term “molar volume” mean? What is the value for molar volume of an ideal gas at STP?
3. Calculate the moles of gas present in a 910 mL sample at 38˚C and 650 Torr.
4. How many atoms of argon gas are in a 137 mL container if the pressure in the container is 8.80×10^{-5} mmHg and the temperature is 794 K?
5. What is the temperature of .75 moles of argon in a 18 L container with a pressure of 790 Torr?
6. Is it possible for 1 mole of air in an adult's lungs to be at STP? Explain and prove by use the ideal gas law.

Gas Density

Activity 1

The purpose of this activity is to demonstrate a thorough understanding of the difference between number density and mass density

1. The ratio of the mass of a substance and the volume that the mass occupies is what is considered to be the __________ of that substance.
2. __________ density is the ratio of mass and volume while __________ density is the ratio of moles or molecules and volume.
3. As some given pressure and temperature, how could two 500 mL closed containers filled only with ideal gas(es) have the same number density but different mass density? Provide a theoretical example.
4. Consider two balloons each filled with the same amount of helium gas. Both balloons are at SATP. Then, you take one balloon and carefully place it into an open container of liquid

nitrogen (77K). The other balloon remains unchanged. Draw both balloons in their initial and final states and include values for number and mass density in each drawing.

5. An adult's lungs can hold about 6L. What mass of air can an adult hold at a pressure of 102 kPa? Normal body temperature is 37 °C and air is about 20% oxygen and 80% nitrogen. (101,325 Pa = 1 atm)

ACTIVITY 2

The purpose of this activity is practice your understanding using gas data to compute molar mass of gas.

1. Use your knowledge of the ideal gas law to write two equations to solve for molar mass of an ideal gas – one where mass is a variable and the other where density is a variable.
2. Given that a 2.16 gram sample of gas occupies 0.43 L at a pressure of 1.2 atm at a temperature of 298 K. Calculate the molar mass of the gas.
3. A gas has a molar mass of 100.0 g/mol. At 25˚C, 1.40 moles of the gas exerts a pressure of 380 torr. What is the density of the gas under these conditions?
4. The unknown gas sample at 137°C and 0.989 bar, has a density of 0.698 g/L. Calculate the molar mass of this unknown gas.
5. A gas exerts a pressure of 1.12 atm in a 4 L container at °19C. You know the density of the gas is 1.5 g/L. What is the molecule?

Gas Mixtures

Activity 1

The purpose of this activity is for you check your understanding of gas mixtures.

1. Explain the relationship between the partial pressure of a gas in a mixture and the total pressure of the mixture with words and an equation.
2. Explain Dalton's Law of partial pressure with words and an equation.
3. A mole fraction is considered a measure of concentration. Explain why this is true.
4. True or False? It is impossible to actually measure the individual pressures of a particular gas within a mixture of gases. Explain your answer.

Activity 2

1. In the troposhere of Titan, Saturn's largest moon, the atmospheric pressure is about 1.5 bar. At this point in the atmosphere there is approximately 4.9% methane (the vast majority of the atmosphere is nitrogen) by number. Calculate the partial pressure of methane in the stratosphere of Titan.
2. You are the Titan expert of your lab group and you decide to recreate the atmosphere in the stratosphere of Titan in a container with a fixed volume of 7.00L. You dutifully remind your lab partners that the stratosphere of Titan is, by number, 98.4% nitrogen gas, 1.40% methane gas and the rest is hydrogen gas. Finally, you remind them to make sure all calculations are for a total pressure of 1.50 bar when the mixture is at a temperature –179˚C in order to best mimic the conditions on Titan. Luckily, you already created a freezer that maintains a perfect temperature of –179˚C into which you will place your model Titan atmosphere. Draw a small particle model of what is going on in the container when the mixture is at a final pressure of 1.5 bar and final temperature of –179˚C. Determine the pressure of each gas in the container and compute how many grams of each gas you will place into this container (after first evacuating it of course).

Writing Gas Law Reactions

Activity 1

1. Correctly write and balance the chemical reaction that occurs between iodine gas (diatomic) and hydrogen gas to produce HI gas.
2. Draw a small particle model of this reaction showing the reactant gases present as a mixture in a container on the left hand side of the equation and the product gas shown on the right hand side of the equation. In your drawing assume that you have exact stoichiometric amounts of the reactants and products needed to completely react away both reactants.
3. Assuming that your reactants and products are at the same temperature and pressure. What is the relationship between the volume of the reactants and the volume of the products?
4. Comment on the relationship between the moles of the reactants and products and the volumes of the reactants and products. What gas law is associated with this relationship?

Activity 2

1. Correctly write and balance the chemical reaction that occurs between ethanol in the gas phase and oxygen gas to produce carbon dioxide gas and water gas.
2. Draw a small particle model of this reaction showing the reactant gases present as a mixture in a container on the left hand side of the equation and the product gases shown on the right hand side of the equation. In your drawing assume that you have exact stoichiometric amounts of the reactants and products needed to completely react away both reactants.
3. Assuming that your reactants and products are at the same temperature and pressure. What is the relationship between the volume of the reactants and the volume of the products?
4. Comment on the relationship between the moles of the reactants and products and the volumes of the reactants and products. What gas law is associated with this relationship?

Gas Stoichiometry Problems

ACTIVITY 1

1. Some fire retardant substances contain solid magnesium hydroxide, which at high temperatures, absorbs heat energy and decomposes into solid magnesium oxide and water vapor. Write the balanced chemical equation (include the phases) and calculate the number of moles of water vapor produced with 3.2 moles of solid magnesium hydroxide.
2. Consider the chemical equation in the previous question. While testing the fire resistant properties of a certain material containing an unknown amount of magnesium hydroxide, you capture and measure the amount of water vapor produced by the complete combustion of the material. Assuming the water vapor comes only from the decomposition of magnesium hydroxide, how many grams of magnesium hydroxide must be present in the material if you capture a total volume of 0.44L with a pressure of 2.5 atm at the reaction temperature of 332˚C?
3. A small piece of solid iron is placed into a 5.00 L container of oxygen gas at room temperature and pressure (1.00 atm) and a small amount of water (assume volume of container is the voume of the gas). The container is sealed and after some time, it is observed that the iron has rusted. The container is still 5.00 L but the pressure is now only 0.800 atm. How many grams of oxygen gas were used to form rust according to this chemical equation: $2Fe(s)+3O_2(g) \rightarrow 2Fe_2O_3(s)$? How many grams did the original sample of iron weigh?
4. Balance the reaction below. If it goes to completion, what total volume will it occupy at STP? Will the volume have increased or decreased? By how much? (Assume ideality.) $C_3H_6O(l)+O_2(g) \rightarrow CO_2(g)+H_2O(g)$

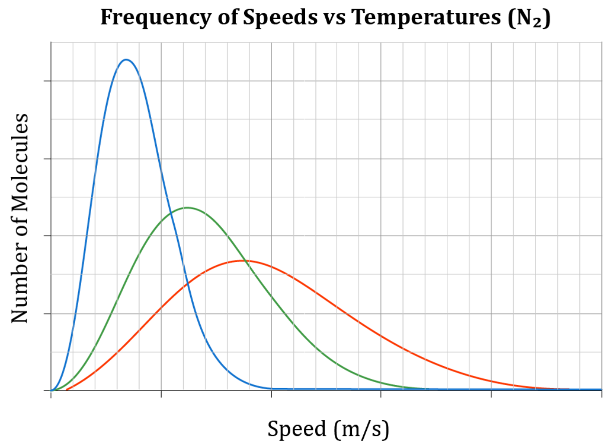
ACTIVITY 2

1. If 4.00 L of S_8 gas and 8.00 L of O_2 were allowed to react, how many liters of SO_2 gas could form. Assume that all gases are at the same temperature and pressure and that the limiting reactant is used up. What would be the final volume of the mixture?
2. Assume that you burn 3.0 L of methane in 12.0 L of oxygen at 370 °C. What volume of carbon dioxide gas can form? What volume of steam (water vapor) is formed? Assume all gases are at the same temperature and pressure and that all the limiting reactant is used up.

Gas Distribution, T & KE, v_{rms}

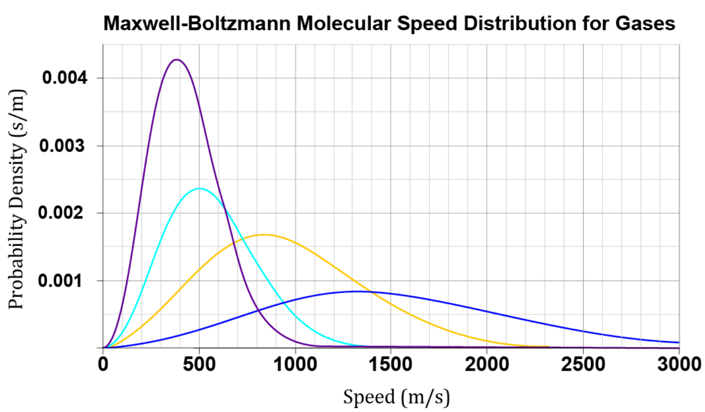
Activity 1

1. Consider the graph of the distribution of velocities of nitrogen gas at different temperatures.



Which curve is associated with the lowest temperature? Describe the shape of the curve.

2. Consider the graph of the distribution of velocities of different gases at the same temperature.



You know that the four gases represented on this graph are helium, neon, argon and krypton. Identify which curve is associated with each gas.

3. These distributions are named after __________.
4. How do the distributions above inform us about the motion of gases? Include in your explanation the relationship between speed and temperature and the relationship between speed and molecular weight.

Activity 2

1. Temperature and kinetic energy are __________ proportional.
2. Consider a textbook on a metal tabletop. Your place one hand on the textbook and the other on the metal tabletop. The metal tabletop feels cooler than the textbook, however and infrared thermometer revels that they are both room temperature. What can you say about the kinetic energy of the molecules that comprise the textbook versus the molecules that comprise the metal tabletop?
3. Write two equations solving for kinetic energy – one with Boltzmann's constant and the other with the gas constant. What is the difference between these two equations?
4. The gas constant *R* is the product of two fundamental constants: __________ constant and __________ number. Therefore, *R* is also a __________ constant.

ACTIVITY 3

1. Describe the relationship between kinetic energy with words and an equation.
2. Describe root mean square velocity with words and an equation.
3. What is the root mean square velocity of neon gas at 311˚C?
4. An unidentified gas has a velocity of 753 $m \cdot s^{-1}$ at STP. What is the identity of this gas? (Yes, this problem contains enough information to answer the question. Yes, it is hard.)

Gas Diffusion and Effusion

ACTIVITY 1

1. Draw small particle models of diffusion and effusion and explain each.
2. You calculated that a certain gas travels at 332 m/s at room temperature. You friend offers to travel 332 meters away from you with a highly sensitive detector so that you can allow the gas sample to diffuse and they can verify if your calculation of it covering 332 meters in one second is correct. Explain to your friend why this experimental set-up will probably not work.
3. Why are diffusion and effusion so much slower than the actual velocity of a gas, (e.g. a gas molecule moving at 1000 $km \cdot hr^{-1}$ diffuses at a tiny fraction of that rate)?
4. Calculate the ratio of the rate of diffusion of oxygen gas to nitrogen gas.
5. The rates of effusion of argon and unknown noble gas are carefully measured. It is found that argon gas effuses at a rate 1.81 times faster than the unknown noble gas. Determine the identity of the unknown gas.

The Kinetic Molecular Theory

ACTIVITY 1

1. The Kinetic Molecular Theory is based on 5 assumptions must be made. Select the five correct assumptions:
   - Particles exhibit attractive and repulsive forces.
   - Particles do not exhibit attractive and repulsive forces.
   - Particles have negligible volume themselves.
   - Particles have intrinsic volume
   - The average kinetic energy of a sample of gas is directly proportional to the absolute temperature of the gas.
   - The average kinetic energy of a sample of gas is inversely proportional to the absolute temperature of the gas.
   - Particles are not in constant motion.
   - Particles are in constant motion.
   - Particles lose energy during collisions.
   - Particles do not lose energy during collisions
2. From which kinetic molecular theory assumption can we support the ideal gas law notion that more moles of gas in a container relate to a greater the pressure (remember that pressure is due to the impacts of the gas particles)? Explain your answer.
3. Explain the common assumption the volume of the container is equal to the volume of the gas using the KMT assumptions.
4. Explain the ideal gas law notion that pressure and volume are inversely proportional using the KMT assumptions.

5. From which kinetic molecular theory assumption can we support the ideal gas law notion pressure and temperature are directly proportional? Explain your answer.

ACTIVITY 2

1. Fully describe – using macroscopic observations, small particle model ideas, and KMT – the behavior of a gas in a balloon that is warmed from 94K to 280K.
2. Fully describe – using macroscopic observations, small particle model ideas, and KMT – the behavior of a gas that exists in a closed cylinder with a moveable piston that starts at a volume of 1 L and is expanded to a volume of 2 L by pulling the piston at constant temperature.
3. Fully describe – using macroscopic observations, small particle model ideas, and KMT – the behavior of a gas in a 2 L closed container on a tabletop that starts STP and then more moles of gas are added. Assume the container is a constant volume container.

Deviations from Ideal Behavior

ACTIVITY 1

1. One of the assumptions of the Kinetic Molecular Theory is that gas particles do not feel attractive forces between themselves. How do we know that this assumption breaks down?
2. We tend to assume that all gases behave ideally. Under what sorts of conditions to gases begin to behave non-ideally?
3. At 1 atm pressure, different gases require different temperature conditions in order to condense into liquids (and/or solids). What does this indicate about the assumption that all gases behave equally ideally?

ACTIVITY 2

1. The hard sphere model accounts for the ___________ of the gas particles themselves. The term V represents the volume of the ___________. The term b is the ___________ ___________ of the gas particles. The b factor becomes more important as pressure ___________. Please write down the equation for the hard sphere model.
2. The van der Waals model accounts for the ___________ ___________ between gas particles with the a term. It also includes the b term. These terms become more important as temperature ___________ (for the a term) and pressure ___________ (for the a term). Both terms are multiplied by ___________ indicating that their effects are dependent on the ___________ of gas. Please write down the van der Waals equation.
3. Organize the following gases from lowest to highest expected b value: NO, Ne, Cl_2 and CO_2 Explain how you determined the order.
4. You place 0.2 moles of an unknown gas in a fixed volume container of 5.0L at some temperature and pressure. Then, you carefully increase the temperature while measuring pressure (in bar) and plot a pressure vs. absolute temperature graph. You measure the slope and find it to be 3.33054×10^{-3} bar/K and extrapolate back to find that the y-intercept is 6.672×10^{-3} bar. Solve for the b and a term and their correct units for the unknown gas using the van der Waals equation.
5. By looking at tables of van der Waals constants, you discover that the unknown gas in the problem above is ammonia. Would you expect methane to have a higher or lower value for a? Which would you expect to have a lower condensation temperature and why?
6. Rank the following gases from most to least ideal in terms of the van der Waals coefficient b: CO_2, SF_6, O_2, H_2, He, CH_4, Rn,

7. Rank the following gases from least to most ideal in terms of the van der Waals coefficient a: N_2, H_2, HCl, HF, NH_3

Electrostatic Forces

ACTIVITY 1

1. The electrostatic force depends on two things. What are they?
2. Describe the relationship between electrostatic force and the two variables you listed above.

ACTIVITY 2

1. True or false? Electrostatic forces cannot be observed on the macroscopic scale because they are so small. Give real world evidence to back up your claim.

Ion-Ion forces

ACTIVITY 1

1. Give an example of two ions that have the same charge but different charge densities.
2. _________ energy is the amount of energy required to completely ionize _________ mole of a solid _________ of an ionic substance.
3. What microscopic qualities would you predict for an ionic substance that has a relatively low melting point (compared to other ionic compounds).

ACTIVITY 2

1. What is the charge on the following ions in the following ionic compounds?
    a. KI
    b. $(NH_4)_2S$
    c. BaS
    d. K_2O
    e. $NaMnO_4$
    f. KNO_2
    g. $TiCl_4$
2. Rank the following in order of increasing lattice energy:
    a. $LiCl$, LiI, LiF, $LiBr$
    b. AlF_3, Al_2O_3
    c. KF, CsF, LiF, NaF

Dipole-Dipole Forces

ACTIVITY 1

1. What kind of molecules experience dipole-dipole forces in their condensed phases?
2. Explain the dipole-dipole force on a microscopic scale.

ACTIVITY 2

1. Which of the following molecules would have dipole-dipole forces in the condensed phase?

NH_3 BH_3 H_2O CO_2 PF_3 CF_4 SeF_4

ACTIVITY 3

1. _________ bonding occurs between molecules where hydrogen is covalently to _________, _________ or _________. These bonds have high _________ and pull most of the electron density from the hydrogen atom. The hydrogen atom can then get _________ to the negative areas of neighboring molecules.

2. Which kind of intermolecular force is generally stronger: dipole-dipole forces or hydrogen bonds? Why?
3. Is a hydrogen bond between two molecules a type of covalent bond? How do we often represent hydrogen bonds in a drawing?
4. Circle the compounds below that have H-bonding in the condensed phase. For those that do have H-bonding, draw a molecular view of the condensed phase and indicate the H-bonding with dashed line.
   NF_3 NH_3 C_2H_4O $C_2H_4NH_2COOH$ (keep C in a chain) CH_2O C_3H_5OOH

Dispersion Forces

Activity 1

1. van der Waals forces are also called _________ forces, _________ _________ forces and _________ _________ -_________ forces.
2. True or false? Dispersion forces exist only in non-polar condensed phase matter. Explain your answer.
3. How do size and shape influence dispersion forces? Is it possible to have a non-polar molecule with a higher boiling point than a molecule with hydrogen bonding capabilities?

Activity 2

1. Circle the compounds in which the only IMF is the dispersion force:
   PCl_5 SCl_4 H_2O O_2 CH_2I_2 C_6H_6 (benzene)
2. Consider the following boiling point data.

Substance	Atomic/Molecular Weight (g/mol)	Boiling Point (˚C)
He	4	–269˚C
Ne	20	–246˚C
Ar	40	–186˚C
Kr	83.8	–152˚C
Xe	131.3	–107˚C
Rn	222	–62˚C
F_2	38	–188˚C
HF	20	+19.5˚C
Cl_2	70.9	–34.1˚C
HCl	36.46	–85.05˚C
Br_2	159.8	+59.4˚C
HBr	80.9	–66.8˚C
I_2	253.8	+185˚C
HI	129	–35.36˚C

3. The boiling points of the noble gases increases as you go down the group. Please explain.
4. Ne and HF both have a molecular weight of 20 g/mol but their boiling points are different. Please explain.
5. The boiling point of F_2 is lower than the boiling point of HF. Is this the typical pattern for the other halogens? Please explain.
6. The boiling point of I_2 is greater than the boiling of HF. Please explain.

Properties of Liquids

Activity 1

1. Vapor pressure is dependent on the ______________ and ______________ of the sample.
2. Define dynamic equilibrium. Explain this term in the context of vapor pressure.
3. Explain how intermolecular forces affect vapor pressure.

Activity 2

1. Compare and contrast the surface molecules of a liquid to the "bulk" molecules below the surface. How do their differences explain the liquid property of surface tension?
2. Explain how intermolecular forces affect surface tension.
3. Which liquid would expect to form a better droplet when a small amount is dropped out of a pipette? Explain your answer.
   a. C_2H_3OOH (acetic acid) or $HOCH_2CHOHCH_2OH$ (glycerol)
   b. $CH_3CH_2OHCH_3$ (isopropanol aka rubbing alcohol) or H_2O (water)

Activity 3

1. Viscosity is the ______________ to ______________.
2. At a given temperature, viscosity depends not only on the ______________ but also the ______________ of the molecule.
3. How does temperature affect the viscosity of a liquid? Explain both microscopically and macroscopically pouring out a jar of honey at 5˚C and at 30˚C.

Activity 4

1. The normal boiling point of a liquid is the temperature at which the ______________ pressure is equal to the atmospheric pressure of ______________ atm.
2. When a liquid boils, bubbles of gas form (choose one: below the surface/on the surface) of the liquid and the vapor pressure is (less than or equal to/greater than or equal to) the atmospheric pressure.
3. Why does it take longer to cook your pasta at high altitudes?
4. Explain the following trends in normal boiling point data:

Substance	Normal Boiling Point (˚C)
Methane (CH_4)	–161.5˚C
Ethane (C_2H_6)	–88.5˚C
Propane (C_3H_8)	–0.5˚C
Methanol (CH_3OH)	64.7˚C
Ethanol (C_2H_5OH)	78.5˚C
1-Propanol (C_3H_7OH)	97˚C
Water (H_2O)	100˚C

Activity 5

1. Predict the order of increasing capillary action for the following: H_2S; H_2O; CH_4; H_2; KBr
2. Put the following compounds in order from lowest boiling point to highest boiling point and justify your answer. CH_4; C_4H_{10}; C_2H_6; C_3H_8; C_5H_{12}

Properties of Solids

Activity 1

1. Fill in the table below:

Type of Solid	How Its Held Together	Example
Molecular		
Covalent/Network		
Metallic		
Ionic		

2. A metal ________ is a mixture of two metals such as bronze (tin and copper).

Activity 2

1. You are given several unknown solid compounds in the lab. You are directed to determine their types: Ionic, Covalent/Network, Molecular, or Metallic. You perform several tests and make the following data table:

Compound	Appearance	Softness/ Rigidity	Melting Point	Solubility in Water	Conductivity	Malleability/ Ductility
A	White and crystalline	Brittle and hard	338˚C	Dissolves easily	Conducts when in solution	Not malleable or ductile
B	Colorless and waxy	Soft	37˚C	Does not dissolve	Does not conduct	Very soft, but not malleable or ductile like a metal
C	Violet crystals	Brittle and hard	113.7˚C	Does not dissolve	Does not conduct	Not malleable or ductile
D	Silvery and shiny	Soft	962˚C	Does not dissolve	Conducts heat and electricity	Both malleable and ductile
E	Colorless and crystalline	Brittle and hard	1,713˚C	Does not dissolve	Does not conduct	Not malleable or ductile
F	White and crystalline	Brittle and hard	962˚C	Dissolves easily	Conducts when in solution	Not malleable or ductile

Identify which types of solid compounds A-F might be:

2. Your lab professor informs you that the six solids were I_2, Ag, SiO_2, $BaCl_2$, $C_{20}H_{42}$ and NH_4Cl. He also informs you that I_2 crystals are violet and that NH_4Cl has a melting point of 338˚C. Match each solid to its correct letter.

Modeling Energy Flow

ACTIVITY 1

1. Write two equations for kinetic energy – one related to the velocity of an object and the other related to the temperature of a sample.
2. What is an example of macroscopic potential energy? What is an example of microscopic potential energy? Write a relationship for both.
3. Compare and contrast kinetic and potential energy.

ACTIVITY 2

1. A _________ _________ is a measurable property that defines the state of a system.
2. Is temperature, *T*, a state function? Give an example of a state function.
3. Give an example of a gas law that describes a gas sample in a final and an initial state.
4. Which type of system does not interact with the surroundings either by matter or energy?
5. What type of system is a chemical reaction occurring in a non-insulated sealed container?

Work

Activity 1

1. Work is the energy of a _________ acting over a _________.
2. The combustion of gases is often associated with _________ work whereas electrical work is _________ work.
3. In chemical thermodynamics, work done _________ the system is given a negative sign and work done _________ the system is given a positive sign.
4. _________ is an organized transfer for energy.

Activity 2

1. Suppose a gas compresses by 185 mL against a pressure of 0.400 atm. How much work is done on the system due to this compression? Report your answer in units of Joules.
2. Calculate the work for the following systems at different temperatures:

   $2CH_3OH\ (g) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 4H_2O\ (g)$ (at 200˚C)

   $2CH_3OH\ (l) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 4H_2O\ (l)$ (at room temperature)

   In each case report the work in units of Joules and state if work was done on or by the system. In which system was the magnitude of the work greater? Explain.

Heat

ACTIVITY 1

1. Heat is the _________ of energy from a _________ temperature area to a _________ temperature area.
2. Describe the relationship between kinetic energy, thermal energy and heat.
3. Draw a microscopic model representing the differences between heat and work.

ACTIVITY 2

1. An _________ property, like volume, is _________ on the amount of substance in the sample. An _________ property, like density, is _________ of the amount of substance in the sample.
2. Heat capacity is an _________ property and has units of _________ per _________.
3. Give an example of heat conduction. How does the heat capacity of a material affect the process of heat conduction?

4. Specific heat capacity is an _________ property and has units of _________ per temperature per _________ unit. Molar heat capacity is an _________ property and has units of _________ per temperature per _________.
5. Which of the following materials would you expect to experience the greatest temperature change if 1 gram of each metal were exposed to the same amount of thermal energy? Explain your answer.

   Aluminum 0.89 J/g °C   Lead 0.129 J/g °C°   Steel 0.466 J/g °C

ACTIVITY 3

1. Calculate the amount of heat needed to raise the T of 1 gram of aluminum from 5°C to 65°C.
2. Calculate the amount of heat needed to raise the T of 1 gram of steel from 5°C to 65°C.
3. Why would building engineers consider the different heat capacities of their building materials? Do you think you might consider switching certain building materials if you live in Arizona versus Maine?

Conservation of Energy

ACTIVITY 1

1. The First Law of Thermodynamics states that energy can neither be _________ nor destroyed – it can only change _________. The total change in energy of the universe for any process is _________.
2. The two forms of energy are _________ and _________. A system's energy can be changed when energy flows in the form of _________ and/or _________.
3. Internal energy a _________ function.
4. Write an equation that relates, internal energy, heat and work.
5. By convention the sign associated with energy flowing _________ of a system is negative and the sign associated with energy flowing _________ a system is positive.

ACTIVITY 2

1. On a microscopic scale, we consider the potential energy of the _________ as well as the nucleus. Potential energy is the energy of _________.
2. What are the three types of molecular motion that contribute to the kinetic energy of a sample?
3. What are two equations for the kinetic energy of the motion of a sample of atoms – one in a single dimension and the other in three dimensions?
4. The rotational energy for a _________ molecule is kT because there are _________ axes of rotation.
5. What is the formula for the change in translational KE for one mole of a monatomic, ideal gas?
6. What is the measured change in internal energy for 2.5 moles of neon gas as it is cooled from 350 K to 100 K?

ACTIVITY 3

1. Assume that a chemical reaction occurring in a non-insulated container closed with a moveable piston that produces gas, moves the piston upwards and heats up the container. Define the system, surroundings and the sign of q. Does the system do work?
2. A system lost 11.9 J of heat, yet it was found that its internal energy increased by 85.5 J. Calculate w. Was work done on the system or did the system do work?
3. A system is heated by using 65.3 kJ of heat and does 34.3 kJ of work. What is the change in internal energy? Based on the words in this question what are the signs of q and w?

Enthalpy Defined

Activity 1

1. The change in enthalpy is equal to the amount of __________ transferred at constant __________.
2. __________, q, is not a state function, but __________, H, is a state function.
3. When enthalpy is derived from the First Law of Thermodynamics, we find that enthalpy, H, is the sum of __________ and __________.
    4. Describe heat flow during a phase change where thermal energy is added to a substance.
    5. When transferring heat to a substance you observe that temperature of the substance is increasing. Describe this scenario on a microscopic scale.

Activity 2

1. An __________ process __________ heat energy and increases the enthalpy of the system. An __________ process __________ heat energy and decreases the enthalpy of the system.
2. Which types of phase changes are endothermic? Which types of phase changes are exothermic? Give an example of a common *chemical* process that is exothermic.

Enthalpy of Physical Change

Activity 1

1. You see your breath on a cold day. What type of physical change is this and is it endothermic or exothermic?
2. You place some ice into a glass on a warm day. As the ice begins to melt, you make regular temperature measurements of the ice water mix. What do you expect the temperature to be each time you take a temperature measurement of the ice water mix? What do you expect to happen to the temperature of the liquid once all the ice has melted?
3. Explain why, on a microscopic scale, you feel cooler when your sweat vaporizes.
4. If you were given the value of $\Delta H_{vaporization}$ for a particular substance, how would you know the value for $\Delta H_{condensation}$?

Activity 2

1. Draw a heating curve with all the steps required to change a solid below its melting point to a gas above its boiling point. Write the equation necessary to quantify the amount of heat associated with each numbered section of the curve.
2. The molar heat of vaporization of Hg is 5.92 kJ/mol at its boiling point, 357 °C. How much heat is released by 120 g of gaseous Hg at 357 °C if it completely vaporizes?
3. Calculate the amount of heat needed to heat 55 g of ice at –30 °C and to completely melt the ice at 0 °C.
4. How much heat is absorbed when 5mol of liquid benzene vaporizes completely at its boiling point (ΔH_{vap} =30.5 kJ/mol)? Does the temperature of the liquid change during this process?

Enthalpy of Chemical Change

Activity 1

1. Draw reaction path diagrams for endothermic and exothermic reactions. Label the change enthalpy, the enthalpy of the reactants and the enthalpy of the products.
2. How is a phase change different from a chemical change?
3. If you treat "heat" as a reactant or a product, how can you indicate whether a reaction is endothermic or exothermic in a chemical equation? For example, is the following reaction endothermic or exothermic and how can you tell: $6CO_2 + 6H_2O + \text{heat} \rightarrow C_6H_{12}O_6 + 6O_2$?

Coffee Cup Calorimetry

Activity 1

1. What is the purpose of a calorimeter?
2. Coffee cup calorimetry is more formally known as ___________ calorimetry. Conveniently, a styrofoam coffee cup generally works for this type of calorimetry because styrofoam isolates the heat flow making it a good ___________ wall.
3. Because a ___________ calorimeter isolates a chemical or physical change in constant pressure, the measured ___________ flow tells us the change in enthalpy of the ___________.

Activity 2

1. When 0.113 g of liquid benzene, C_6H_6, burns in excess oxygen in a calibrated constant pressure calorimeter the temperature of the calorimeter rises by 8.60˚C and carbon dioxide gas and liquid water are formed.
   In this example, what is the system? What are the surroundings?
2. Write the balanced chemical equation for the combustion of benzene and include the word "heat" on the appropriate side of the equation. Also draw an energy level diagram showing how the potential energy of the system changes as a result of the combustion.
3. Given that there are only two types of energy, KE and PE, what must have happened to the PE of the system – an increase, decrease or no change?
4. What must have happened to the energy of the surroundings?
5. Given that the heat capacity of the calorimeter is 551 J/˚C, how might one calculate the amount of energy that flowed from the system to the surroundings?
6. Calculate the amount of heat that flowed from the system to the surroundings.
7. Heat flow at constant pressure is defined as change in enthalpy or ΔH. Calculate the change in enthalpy of the combustion per mole of liquid benzene using the data given.

Bomb Calorimetry

Activity 1

1. A ___________ calorimeter isolates a chemical or physical change in constant volume. The measured heat flow tells us the change in ___________ ___________ of the system/reaction.
2. When the consumption or production of gases is a key part of a chemical or physical process it is preferable to use a ___________ calorimeter. Why?

Activity 2

1. 1.14 g of liquid octane (C_8H_{18}) is combusted in a bomb calorimeter surrounded by 1 L of water. The combustion produces carbon dioxide gas and liquid water. The initial temperature of the water and calorimeter hardware is 25˚C. The final temperature of the water and calorimeter hardware is 33˚C. The heat capacity of the calorimeter (not including the water!) is 456 J/˚C. Determine the heat flow into the surroundings (water + calorimeter). Recall that the density of water is 1 g/mL.
2. From the data, determine the change in internal energy of octane per mole combusted.
3. Based on the change in gas moles for the combustion of octane, starting and ending at room temperature, determine the molar enthalpy of combustion of octane.

Thermochemical Equations

ACTIVITY 1

1. Here is a thermochemical equation for the combustion of cyclohexane at room temperature:
   C_6H_{12} (l) + $9O_2$ (g) → $6CO_2$ (g) + $6H_2O$ (l) $\Delta H = -3920$ kJ mol^{-1}rxn^{-1}
   What would happen to the value of the enthalpy if you reversed the chemical process (switched reactants and products)? What would happen to the value of the enthalpy if you reduced each of the stoichiometruc coefficients by a factor of 3?
2. Do you think the value of the change in enthalpy for the thermochemical equation is temperature dependent? Why or why not?

ACTIVITY 2

1. Here is the same thermochemical equation:
   C_6H_{12} (l) + $9O_2$ (g) → $6CO_2$ (g) + $6H_2O$ (l) $\Delta H = -3920$ kJ mol^{-1}rxn^{-1}
   How could you write this thermochemical equation differently but still provide the value for the change in enthalpy of the reaction?
2. For the two formats of thermochemical equations, how can you determine if the reaction is endothermic or exothermic?

ACTIVITY 3

1. How much heat is released when 8.40 g of C_6H_{12} is combusted according to the equation below?
   C_6H_{12} (l) + $9O_2$ (g) → $6CO_2$ (g) + $6H_2O$ (l) $\Delta H = -3920$ kJ mol^{-1}rxn^{-1}

Hess's Law

ACTIVITY 1

1. State Hess's Law and discuss its implication when calculating the change in enthalpy of a reaction.
2. Show how you would manipulate the listed thermochemical equations to get the desired reaction for the synthesis of anhydrous aluminum chloride, which is written:
   2Al (s) + $3Cl_2$ (g) → $2AlCl_3$ (s)

   ΔH_1 2Al (s) + 6HCl (aq) → $2AlCl_3$ (aq) + $3H_2$ (g) –149 kJ mol^{-1}rxn^{-1}
   ΔH_2 HCl (g) → HCl (l) –74 kJ mol^{-1}rxn^{-1}
   ΔH_3 H_2 (g) + Cl_2 (g) → 2HCl (g) –185 kJ mol^{-1}rxn^{-1}
   ΔH_4 $AlCl_3$ (s) → $AlCl_3$ (aq) –323 kJ mol^{-1}rxn^{-1}
   Note: Assume HCl (l) is equivalent to HCl (aq) for this problem.
3. Calculate the reaction enthalpy for the synthesis of anhydrous aluminum chloride using the thermochemical data provided with the thermochemical equations.

ACTIVITY 2

1. A formation reaction is a chemical reaction that produces __________ mole of a substance from its __________.
2. How is a value calculated for standard conditions noted in its abbreviation?
3. Was the reaction given for the synthesis of anhydrous aluminum chloride in the previous activity (question#2) a formation reaction? Why or why not? If not, please correct the equation so that it is a formation reaction.

ACTIVITY 3

1. Write the formation reaction for hydrogen sulfide gas (assume standard conditions) from solid sulfur (monatomic) and hydrogen gas.

2. Look up the value for the enthalpy of formation for hydrogen sulfide from a thermodynamic data table.
3. Use the enthalpies of formation from a thermodynamic data table to calculate the enthalpy for the combustion of hydrogen sulfide gas in oxygen when the products are sulfur dioxide gas and gaseous water.

ACTIVITY 4

1. Bond energy is the energy required to break the bonds in __________ mole of a substance in the __________ phase.
2. __________ bonds is an endothermic process while __________ bonds is an exothermic process.
3. Use a bond energy table to ***set-up*** the change in enthalpy for the combustion of hydrogen sulfide as you wrote it in the previous activity (question #3). Don't actually solve! Consider both resonance structures of SO_2. What does that mean for the bond energy of the sulfur–oxygen bonds?
4. What are the three methods discussed in this supplemental worksheet that allow you to calculate the change in enthalpy of a reaction? Which two give the most similar answers and why does the third method sometimes provide a different value than the other two?

Spontaneity

ACTIVITY 1

1. A ___________change is defined as one that happens in only one direction in ___________, which means on its own with no outside influences.
2. Give two examples of processes that are spontaneous at room temperature and pressure – provide one physical change and one chemical change.
3. Can spontaneous reactions ever be reversed? Does temperature ever play a role in the spontaneity of a reaction?

ACTIVITY 2

1. The second law of thermodynamics states that all spontaneous changes are accompanied by an ___________ in ___________entropy.
2. The change in entropy of the universe takes into account the change in entropy of the ___________ and the change in entropy of the ___________.
3. Entropy is a ___________ function that is a measure of the ___________of energy.
4. In thermodynamics, the word "___________" indicates a specifically defined path for some process.
5. Give an equation for the change in entropy using heat and absolute temperature. Explain the equation in words.

ACTIVITY 3

1. When dry ice sublimes, the sign of the change in entropy for the system must be _____, the sign of the change in entropy of the surroundings must be _____. The overall change in entropy of the universe must be _____. Given this situation what must be the relationship of the magnitude of the values for the system and the surroundings?
2. Consider the dissolution of a small amount of table salt into water at room temperature. Define the system and the surroundings. Is the $\Delta S_{sys} > $ or < 0 or is there no way to know? Is the $\Delta S_{sur} >$ or < 0 or is there no way to know? Please justify your answer.
3. Consider a chemical reaction in which a solid metal is placed into a beaker of water. Upon mixing, a gas is released as well as light and heat. Define the system and the surroundings. Is the $\Delta S_{sys} >$ or < 0 or is there no way to know? Is the $\Delta S_{sur} >$ or < 0 or is there no way to know? Please justify your answer.

Change in Entropy

ACTIVITY 1

1. What are the five things that will lead to a change in entropy of the system.
2. Describe how each of the five things you listed above could change to result in a decrease in entropy for a system.
3. How could a sample increase in positional disorder without increasing in thermal disorder?

ACTIVITY 2

1. Write three equations for the change in entropy of a system when it experiences a change in temperature, a change in volume and or a change in phase. Define each of the variables and give the units for each.
2. Calculate the change in entropy for 4.8 moles of H_2O gas when it increases in temperature from 400 K to 500 K. The heat capacity for H_2O gas is 36.57 J/molK.
3. Calculate the change in entropy of fusion for one mole of mercury. The melting point of mercury is –37.9˚C.
4. Calculate the change in entropy of fusion for 63 grams of mercury.

ACTIVITY 3

1. Write the equation for the change in entropy of the surroundings. Define each of the variables and give the units for each.
2. If you know the heat flow in or out of the system at constant pressure, could you calculate the change in entropy of the surroundings (given the temperature of the surroundings)?
3. Calculate the ΔS_{sur} for the vaporization of 216 g of benzene (C_6H_6) if the surrounding temperature is 25.0 °C.

ACTIVITY 4

1. Imagine you start with a beaker that has **45 g** of liquid nitrogen. The initial temperature of the liquid nitrogen is 77 K. If you leave this beaker in a room that has a constant temperature of 25 °C, the liquid nitrogen will *spontaneously* boil and you will end up with all gaseous nitrogen at 25 °C. Here we will calculate the entropy change for the process of the liquid nitrogen boiling and increasing in temperature to 25 °C.
   a. First identify what the system and surroundings are. (what are they composed of, what is their temperature, etc…) Often it is helpful to make a sketch or diagram.
   b. Now identify what the initial and final states of the system are. Again a diagram, equation, or sketch can be helpful. You should break this down into two steps: liquid nitrogen boiling, then the nitrogen gas warming up.
   c. Now look at the energy change for this process. Does energy flow into or out of the system? If so, as heat or work?
   d. Given that $\Delta H°_{vap}$=5.92 $kJmol^{-1}$ and that the heat capacity for nitrogen gas is 1.04 $Jg^{-1}°K^{-1}$. How much heat flows into the system for this process (in kJ)?
   e. What is the change in entropy for the boiling of the liquid nitrogen? What is the change in entropy for the warming of the nitrogen gas? What is the change in entropy for the system?
2. Now let's consider the surroundings for this particular system.
   a. Did heat flow into or out of the surroundings during this change?
   b. What is the entropy change for the surroundings?
3. Now let's consider the entropy change for the universe for this process.
   a. What is the entropy change for the universe?
   b. Did you expect this process to have a + or – ΔS_{univ}? Why?
   c. Is there a condition in which you would expect this process to not be spontaneous? If so what effect would such a condition have on the ΔS_{surr}?

Absolute Entropy

ACTIVITY 1

1. At absolute _________ Kelvin we can consider a perfectly ordered system. In this perfectly ordered system, there is no _________ disorder, temperature or _________ energy meaning no thermal disorder. In this perfect state we can define an absolute scale for _________. This theoretical system gives us an understanding of the _________ law of thermodynamics.
2. Which compound would you expect to have the lowest absolute entropy of the given pairs? Explain your choice.
   a. Br_2 (l) or Br_2 (g)?
   b. C (s) graphite or C (s) diamond?
   c. Fe_2O_3 (s) or Fe_3O_4 (s)?
   d. NH_3 (g) or NH_3 (aq)?

ACTIVITY 2

1. A system with high positional entropy has _________ microstates, which are all the possible _________ of energy in a system
2. Consider a small amount of potassium bromide (white crystalline solid) and a beaker of water. Draw these images.
3. Now dissolve the KBr into the beaker of water. Describe what happened to your salt.
4. Are there more or less microstates available when the salt is dissolved? Please explain.
5. Would you consider the above process to be spontaneous? If so, what must be true about the ΔS_{univ}? What about the ΔS_{sys}? What about the ΔS_{surr}?

ACTIVITY 3

1. What is the equation to calculate absolute entropy of a particular sample (not per mole)? Define each of the variables and give units for each.
2. What is the equation to calculate absolute entropy of a mole of a substance? Define each of the variables and give units for each.
3. Calculate the number of microstates **and** the number of molecules of a sample of a solid in which the molecules can take any one of three orientations with the same energy and the absolute entropy is 6.58051×10^{-22} J/K.

Change in Entropy from Absolute Entropy

ACTIVITY 1

1. Use the equation to determine the ΔS_{rxn} for the combustion of 1 mole of diamond (C (s, diamond)) at room temperature. You will need to look up the $S°$ values from a thermodynamic table.

ACTIVITY 2

1. Consider the system in Activity 1, above. Given that $\Delta H_f°(CO_2(g))$l = –393.51 kJ/mol, determine the change in entropy for the surroundings for this change under standard conditions.
2. What is the ΔS_{univ} for this particular change, the combustion of 1 mole of diamond at room temperature?
3. Is this change spontaneous or nonspontaneous? Why or why not?

Gibb's Free Energy

Activity 1

1. Both __________ __________ energy, G, and the change in entropy of the __________, $\Delta S_{universe}$, can be used to predict the spontaneity of physical or chemical change.
2. The change in the Gibb's free energy, ΔG, is actually derived from the change in entropy of the universe. What is the equation relating the change in Gibb's free energy for a process and the change in enthalpy and entropy for the same process? Define the variables and provide units for each.
3. What is the sign of ΔG for a spontaneous process? For a non-spontaneous process?
4. What are the four possible combinations of ΔH and ΔS and what kind of change do they lead to – always spontaneous, always non-spontaneous or temperature dependent?

Activity 2

1. How does the change in Gibb's free energy of formation under standard conditions, ΔG_f°, inform us about the stability of a compound?
2. Which compound would you expect to have the lowest ΔG_f° (under standard conditions) of the given pairs? Explain your choice.
   a. Br_2 (l) or Br_2 (g)?
   b. O (g) or O_2 (g)?
   c. Fe (s) or Fe_2O_3 (s)?
   d. O_2 (g) or Al_2O_3 (s)?
3. Explain the concept of thermodynamic stability. Explain how the temperature and pressure conditions of a given environment might change the stability of a compound.
4. Corrosion is a huge obstacle in industry. One of the most common and pervasive corrosion reactions is the formation of iron rust (iron oxides). Provide a thermodynamic explanation as to why the formation of iron oxides continues to present industrial challenges.

Activity 3

1. Calculate the ΔG_{rxn} of the following reaction using the given standard free energy of formation data:
   $6CO_2\ (g) + 6H_2O\ (l) \rightarrow C_6H_{12}O_6\ (s) + 6O_2\ (6)$
   $\Delta G_f^\circ\ _{CO_2\ (g)} = -394.36$ kJ/mol
   $\Delta G_f^\circ\ _{H_2O\ (l)} = -237$ kJ/mol
   $\Delta G_f^\circ\ _{C_6H_{12}O_6\ (s)} = -910$ kJ/mol

2. Here is a chemical reaction occurring under standard conditions:
   C (s, graphite) + ½ O_2 (g) + $2H_2$ (g)$\rightarrow$ CH_3OH (l)
   The associated thermodynamic data is:
   $S^\circ\ _{C\ (s,\ graphite)} = 5.740$ J/K mol $S^\circ\ _{O_2\ (g)} = 205$ J/K mol $S^\circ\ _{H_2\ (g)} = 131$ J/K mol
   $S^\circ\ _{CH_3OH\ (l)} = 127$ J/K mol $\Delta H_f^\circ\ _{CH_3OH\ (l)} = -239$ kJ/mol
   Calculate the ΔG_{rxn} under standard conditions from this data.

3. Is the above equation a formation equation? Explain. Look up $\Delta G_f^\circ\ _{CH_3OH\ (l)}$, and compare that tabulated value to the value you calculated here.

Predicting Spontaneity

Activity 1

1. Dry ice (solid CO_2) spontaneously sublimes in room temperature. What can you conclude about the *ΔH* of this system: CO_2 (s) → CO_2 (g)?

2. What can we conclude about the magnitude of the Δ*H* term in the dry ice example?
3. What can we conclude about changing the temperature and its effect on the spontaneity of this reaction? That is do you predict that decreasing the temperature will help the dry ice not sublime?
4. The equation for the formation of 2 moles of NH_3 under standard conditions is:

   N_2 (g) + $3H_2$ (g) → $2NH_3$ (g)

   The associated thermodynamic data is:

   $S°_{N_2 (g)}$ = 192 J/K mol $S°_{H_2 (g)}$ = 131 J/K mol $S°_{NH_3 (g)}$ = 193 J/K mol

   $\Delta H_f^\circ{}_{NH_3 (g)}$ = −46 kJ/mol

   a. Given this information calculate the ΔS_{univ}. Based on the ΔS_{univ} for this example would you predict that this reaction spontaneous? Why or Why not?
   b. Using the given thermodynamic data calculate ΔG_{rxn} at 298K.
   c. Based on your calculated value of ΔG_{rxn} is the formation of 2 moles of NH_3 under these conditions spontaneous?
   d. Is there a temperature at which point this reaction would become nonspontaneous? If so what is that temperature?

Activity 2

1. Choose the correct word to complete the statement and then give an example of that kind of reaction:
   a. An endothermic reaction for which the entropy of the system is increasing the reaction will be (choose one: always/never/sometimes) spontaneous. Give an example of such a reaction.
   b. An endothermic reaction for which the entropy of the system is decreasing the reaction will be (choose one: always/never/sometimes) spontaneous.
   c. An exothermic reaction for which the entropy of the system is increasing the reaction will be (choose one: always/never/sometimes) spontaneous. Give an example of such a reaction.
   d. An exothermic reaction for which the entropy of the system is decreasing the reaction will be (choose one: always/never/sometimes) spontaneous. Give an example of such a reaction.
2. What is the value of ΔG_{rxn} when a system is at equilibrium?
3. What is true about the forward process and the backward process of a system at equilibrium?
4. The condensation of CH_3OH gas to CH_3OH liquid is spontaneous when *ΔG* is negative. Using the data below, predict at what temperature is this process first spontaneous at 1 bar.

   $S°_{CH_3OH (g)}$ = 240 J/K mol $S°_{CH_3OH (l)}$ = 127 J/K mol

   $\Delta H_f^\circ{}_{CH_3OH (g)}$ = −201 kJ/mol $\Delta H_f^\circ{}_{CH_3OH (l)}$ = −239 kJ/mol