I. Chemistry and Biology Review

Review Activity: Divide into four groups. Each group will review one of the sets of concepts shown below. After reviewing the concepts among your group, you will then be asked to briefly explain your answers to the rest of the section.

Group 1 (Units): Define the metric units used to measure or describe mass, length, electrical charge, temperature, rates of chemical reactions, and energy. If a physical characteristic can be described by more than one unit, describe how the different units might relate.

mass= kg, rate= s^{-1}, length (m), energy (J; kcal/mol); elementary charge unit (the charge of an electron if negative, the charge of a proton if positive)

"kcal/mol" is used in the notes to define bond strength. It is a useful way of standardizing the energy of a bond per an amount (a mole) of whatever substance (like water) is involved in that bond

The SI unit for charge is the Coulomb, which is much too large for our purposes, which is why we as chemists use the elementary formal charge unit, where a proton = +1 and an electron = -1.

Group 2 (Orders of Magnitude): Determine the metric unit prefixes (which describe orders or magnitude) for 10^6, 10^3, 10^{-3}, 10^{-6}, 10^{-9}, 10^{-12}, and 10^{-15}. What is the term for a 10^{-10}th of a meter? Develop a question to ask the rest of the class regarding units.

10^6= "mega," 10^3= "kilo," 10^{-3}= "milli," 10^{-6}= "micro," 10^{-9}= "nano," 10^{-12} = "pico," and 10^{-15} = "femto"

A 10^{-10} m is called an "angstrom."

Group 3 (The Mole): Define the concepts of moles, molarity, and Avogadro's number, and how they relate to each other using an example molecule, such as NaOH.

Mole- The SI unit for an amount of a substance in grams which equals its atomic mass (so for NaOH, the molar mass would equal 40 grams). A very helpful concept that allows us to balance chemical equations, etc..

Molarity- number of moles of a solute in a liter of solvent. The unit of molarity is mols / liter.

Avogadro's number- 6.02 x 10^{23}- describes the number of molecules within a mole, For example, a mole of NaOH weighs 40g and has 6.02 x 10^{23} molecules of NaOH within it.

Group 4 (The Cell): Briefly define the parts of the cell, such as the nucleus, mitochondria, plasma membrane, cytoplasm, ribosomes, cytoskeleton. Develop a question for the rest of the class including a diagram of a cell.

Nucleus- where the eukaryotic hereditary info (DNA) is kept, enclosed by the nuclear envelope

Mitochondria- Organelle where food is oxidized to produce ATP, the primary energy currency of the cell.

Plasma membrane- Every cell is confined by a membrane. They behave kind of like a fluid, kind of like a solid, which is why they are sometimes called "plasma" membranes.

Ribosomes- are large, complex molecules in the cell responsible for translating the RNA sequence into the appropriate amino-acid sequence of a protein. Ribosomes themselves consist of very large RNA strands in conjunction with dozens of different proteins.

Cytoplasm- The cytoplasm is the gel-like substance within a cell. For prokaryotes, it is all of the space within the cell. In the case of eukaryotes, it does not include the nucleus, and is therefore the space between the nucleus and the cell membrane.

Cytoskeleton- The cytoskeleton is- aptly- the skeleton of the cell (duh). It consists of microtubules and microfilaments (both made out of protein) that provide a cell with shape, organization, and support as the cell moves and divides.

II. Valence Electrons and Bonding

- The number of valence electrons in an atom determines how many bonds it can form. It is very useful to know the number of valence electrons present in carbon, hydrogen, oxygen and nitrogen in their free and neutral states, as well as the number of lone pairs. For phosphorous and sulfur, it is sufficient to know the number of bonds that each can form while still remaining neutral.

Fill out the following table for biology's most common elements. Assume that each atom is **neutrally charged** and not bonded to any other atom(s).

	Carbon	Hydrogen	Oxygen	Nitrogen	Sulfur*	Phosphorus*
Lewis Dot Structure of Free Atom	$\cdot\overset{\cdot}{\underset{\cdot}{C}}\cdot$	$H\cdot$	$\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}\cdot$	$\cdot\overset{\cdot\cdot}{\underset{\cdot}{N}}\cdot$	$\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{S}}\cdot$	$\cdot\overset{\cdot\cdot}{\underset{\cdot}{P}}\cdot$
# of Valence Electrons	4	1	6	5	6	5
# of Lone Pairs on Neutral Atom When Octet is Satisfied	0	0	2	1		
# of Bonds Neutral Atom Forms	4	1	2	3	2, 4 or 6	3 or 5

* Phosphorous and sulfur can undergo octet expansion

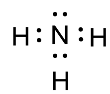
III. Formal Charge

- Atoms acquire formal charges when they have either more or less valence electrons than are needed to be neutral.
- A quick way to assess formal charges on C, H, O and N is to compare the number of valence electrons present in the neutral atom to the number of valence electrons that it is assigned when

bound to other atoms. An atom is assigned all of its lone pairs and half of the electrons in a covalent bond.

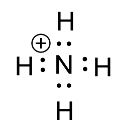
Example: Draw Lewis structures for the following molecules. Using your knowledge of how many valence electrons each atom has in its neutral state, assign appropriate formal charges to any relevant atoms within each molecule.

NH_3



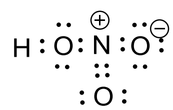
Ammonia

NH_4^+



Ammonium

HNO_3

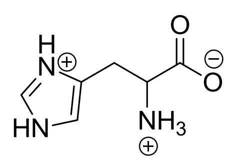


Nitric Acid

For HNO_3, the atoms could be connected in other ways and not violate the octet rule for individual atoms. All valid structures would be considered acceptable (even though nitric acid only has one structure).

Formal Charge= (# of valence electrons) – (# of lone pairs electrons) – (1/2 # of bond pair electrons)

Section Problem #1: Label any formal charges present in the following molecule.



Histidine

One-minute essay:
Take a minute to think about your answer to the two questions below, and then record your responses (along with your name) on the index card given to you at the beginning of section:
1. What is the difference between a full charge and a partial charge?

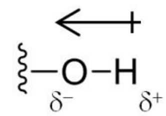
A full charge has an absolute value of 1, and is either positive in the case of a proton or negative in the case of an electron. A partial charge is a charge with an absolute value of less than 1.

2. How does this difference relate to whether a bond is ionic or permanent-dipole:permanent:dipole?

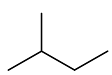
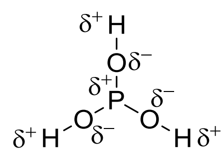
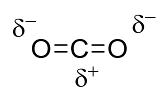
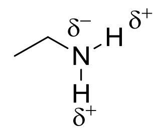
Ionic bonds are chemical bonds that form between two oppositely charged ions ("Ions" are atoms or molecules which have a formal positive or negative charge because they unequal numbers of protons and electrons).
Permanent-dipole:permanent-diple bonds are chemical bonds that form between two molecules that only have partial (rather than a full) charges. Partial charges are generated when there is a big difference in the electronegativity between two covalently bonded atoms.

IV. Electronegativity, Polar Bonds and Partial Charges

- Define Electronegativity: ***The tendency of an atom to attract electrons.***
- When there is a significant electronegativity difference (ΔEN) between two atoms in a bond, electrons are shared unevenly and there is a separation of charge in the bond. This results in a **polar covalent bond**. Polar bonds can lead to polar molecules.
- Commonly observed polar covalent bonds in the molecules of life are N–H, C–O, P–O and O–H bonds. Relatively non-polar bonds generally have ΔEN values of 0.4 or less. Electronegativity values are located in the lecture notes.
- Dipoles are often represented graphically as an arrow that points towards the partially negative atom, as shown to the right.

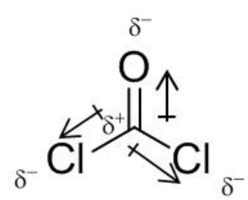


Section Problem #2: In the molecules shown below, determine which bonds are polar and label partial charges on all relevant atoms. Also, designate whether or not each molecule has a net dipole moment.

Net Dipole: **No** **Yes** **No** **Yes**

On the molecule below, label partial charges and draw an arrow to represent the dipole of each polar bond.

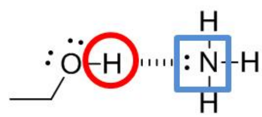


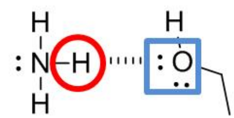
V. Hydrogen Bonding

- Hydrogen bonds are a special type of dipole-dipole interaction in which a hydrogen atom that is bonded to an electronegative atom (O, N, F, or Cl) is electrostatically attracted to the lone pair of another electronegative O, N, F or Cl atom.
- Properties of Hydrogen Bonds:
  1. The donor must be a hydrogen atom with a partial **positive** charge.
  2. The acceptor must have at least one **lone pair** and a partial **negative** charge.
  3. Hydrogen bonds are (**strongest**/weakest) when they are linear.

iv. The bond strength (increases/**decreases**) as bond distance increases.

Section Problem #3: Ammonia (NH_3) can hydrogen bond with ethanol (CH_3CH_2OH). Draw two ways in which this interaction can occur in the space below. For every hydrogen bond donor, circle the hydrogen involved in the bond and for hydrogen bond acceptors, draw a square around the atom that is donating the lone pair of electrons involved in the bond. Be sure to use optimal bond angles in your drawing.

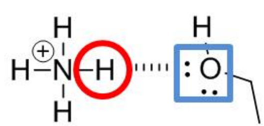




Donors are circled in red, and acceptors are boxed in blue. It is equally feasible to designate the entire O-H group or the N-H group as the donors. However, only the electronegative atom (ether O or N) should be labeled as the acceptor. Notice that in this problem, ethanol and ammonia can provide both H-bond acceptors and H-bond donors so the interaction can be drawn in two different ways.

If ammonium (NH_4^+) were used instead of ammonia, would the situation change? If so, how?

This is now an ion-dipole interaction. There can still exist an electrostatic attraction between a lone pair of electrons on the oxygen atom in ethanol and a properly oriented hydrogen atom on NH_4^+, similar to what occurs in NH_3. However, ammonium ion no longer has a lone pair on the nitrogen atom and thus cannot act as a hydrogen bond acceptor.



VI. Intermolecular Interactions

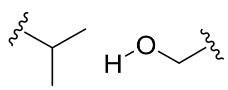
Review the different types of intermolecular interactions and highlight the relative strengths of each. Keep in mind that the strength of ionic interactions is attenuated in water.

- Molecules don't exist in isolation; they are involved in a range of intermolecular interactions with themselves & other molecules. These interactions all arise from attractive electrostatic interactions between various combinations of full charges, partial charges, and instantaneous charges.

Section Problem #4: Shown below are a series of intermolecular interactions.

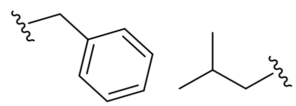
A. Identify the strongest type of interaction that is possible for each pair. Assume that all molecules lie in the same plane.

A.



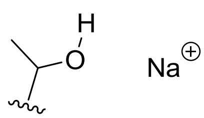
Dipole-Induced Dipole

B.



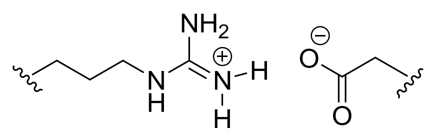
Induced Dipole-Induced Dipole

C.



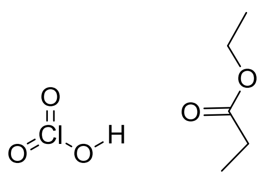
Ion-Dipole

D.



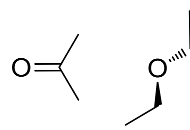
Ionic

E.



Hydrogen Bond (dipole-dipole)

F.



Dipole-Dipole

B. Which pairs could hydrogen bond?

Pairs D and E could form hydrogen bonds if the donors and acceptors were favorably oriented with respect to each other.

C. The chart shown below represents a continuum of increasing interaction strength. Fill in the blanks on the chart with the letters from the examples above such that each interaction is ordered from weakest to strongest.