Pre-Course Worksheet

Summer HSSP, 2018

These are some practice problems that test your understanding of some fundamental general chemistry topics that we will be using. Don't worry if you don't know everything. We will have a brief review in class, and we won't be using everything right away, but try to review any topics on your own that you are struggling with.

Atomic Orbitals

1. What is the shape of an s-orbital? A p-orbital? The d-orbitals?
   S-orbitals: spheres, p-orbitals: dumbbells, d-orbitals: dumbbell-ish (look it up online)
2. How many valence electrons does each of the following elements have in its neutral state?
   a. Na 1
   b. Cl 7
   c. O 6
   d. N 5
   e. C 4
   f. H 1
   g. Mg 2
3. How many electrons can reside in a single atomic orbital? Are there any requirements on the/those electron(s)? 2 electrons/orbital, they must have opposite spin.
4. What charge does each of the following elements typically form when ionized?
   a. Cl -1
   b. O -2
   c. Li +1
   d. Ca +2

Periodic Trends

1. Arrange the following elements in order of increasing size (atomic radius):
   a. Na, K, Li Li < Na < K (increase from top to bottom)
   b. O, N, C O < N < C (decrease from left to right)
   c. S, Br, Cl Cl < S < Br
2. Arrange the following ions and elements in order of increasing size (atomic radius):
   a. Na^+, Ne, F^-, Mg^{2+} $Mg^{2+} < Na^+ < Ne < F^-$ (when all species have the same number of electrons, the species with higher atomic number will exert a stronger pulling force on the electrons because of more protons, so they tend to be smaller -- see isoelectronic species)
3. Arrange the following elements in order of increasing electronegativity:
   a. O, N, C C < N < O (increase left to right)
   b. K, Li, Na K < Na < Li (increase from bottom to top)
   c. Cl, Br, I I < Br < Cl (increase from bottom to top)
4. Explain why the second ionization energy of sodium is vastly greater than its first ionization energy.
   Removing the second electron involves removing a core electron which requires much more energy than simply removing the outer valence electron.

Bonding

1. What is the difference between an ionic and a covalent bond?

Ionic bonds are between two atoms of very different electronegativities whereas the two elements in a covalent bond only have slightly different, if any, difference in electronegativity.

2. What is a sigma bond? A pi bond? Which type allows for rotation around the bonding axis?

Sigma bonds are formed when orbitals overlap head-on. Pi bonds are formed when orbitals overlap in parallel. Sigma bonds allow for rotation.

3. Is the bonding between each of the following pairs of elements normally ionic or covalent?
    a. N, O covalent
    b. Na, Br ionic
    c. K, O ionic
    d. C, H covalent

VSEPR Model and Hybridization

1. What is the shape of each of the following molecules?
    a. CH_4 tetrahedral
    b. NH_3 trigonal pyramidal
    c. H_2O bent or angular
    d. PCl_5 trigonal bipyramidal
    e. BH_3 trigonal planar
2. How many lone pairs are in each of the following molecules in 1? 0, 1, 2, 0, 0
3. What is the hybridization of the central atom in each of the molecules in 1? Sp3, sp3, sp3, dsp3, sp2

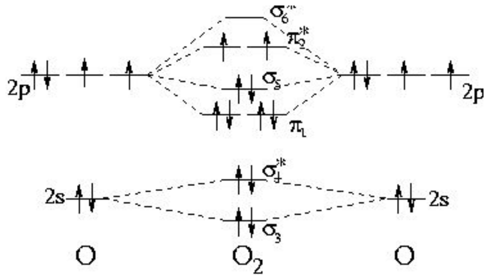
Molecular Orbitals

1. Do electrons in bonding orbitals raise or lower the energy of the bond? In other words, do they make the bond more stable or less stable?

They lower the energy. Electrons occupying bonding orbitals contribute to the bonding and stabilize the molecule. Non-bonding electrons do not stabilize and anti-bonding orbitals destabilize the bond.

2. Explain why oxygen gas (O_2) is diamagnetic.

If you draw the molecular orbital diagram, you will notice that the last two electrons occupy two different pi-2p* orbitals and are thus unpaired. Thus O_2 is diamagnetic. See the diagram below.

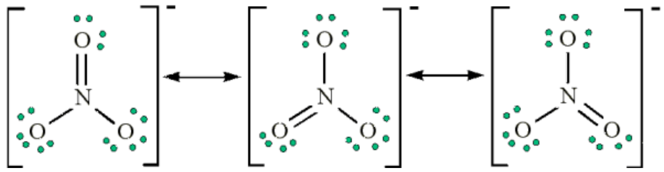


Lewis Structures and Resonance

1. Draw the Lewis structures for each of the following molecules/ions.
    a. CH_4
    b. NH_3
    c. SO_4^{2-} (by the way, do you know the name for this ion? Sulfate)
    d. NO
    e. C_2H_4

    It's kind of a pain to draw all these Lewis structures, so just look these up online.

2. Show all the resonance structures for NO_3^-. Also draw the resonance hybrid structure. What is the bond order of each N-O bond?



Bond order = 4/3 (each N-O bond has a total of 4 bonds across all 3 structures and then we divide by the number of structures -- the resonance structure implies that the true electron distribution is close to a 4/3 bond between the N-O rather than either a single or a double bond)

Intermolecular Forces

1. Identify the dominant intermolecular force within pure samples of each of the following compounds. Choose from: Van der Waals forces (London dispersion forces), dipole-dipole interactions, and hydrogen bonding.
    a. O_2 Van der Waals (non-polar and no hydrogen)
    b. H_2O Hydrogen bonding (polar and contains an H bonded to a small, electronegative atom)
    c. HCl Dipole-dipole (polar but H is not bonded to a small, electronegative atom -- Cl is rather large)
2. Do non-polar solvents tend to dissolve polar or non-polar solutes?

Follow the like-dissolves-like rule; non-polar solvents dissolve non-polar solutes.

Acids and Bases

1. What is the pH of a 0.1M solution of HCl? -log(0.1 M) = 1
2. Rank the following pKa's from strongest acid to weakest acid: 2, 10, 5 2, 5, 10
3. For a particular species, what is pKa + pKb? pKa + pKb = pKw = 14
4. List some strong and weak acids. Strong: H_2SO_4 (first deprotonation), HCl, HBr, HI Weak: HF, H_3PO_4
5. What is the conjugate base of H_2SO_4? What is the name of this acid? HSO_4^-, sulfuric acid
6. Is NH_3 typically an acid or a base? base
7. Write the chemical equation corresponding to the Ka of a hypothetical acid HA. $HA + H_2O \rightarrow H_3O^+ + A^-$

Equilibrium and Kinetics

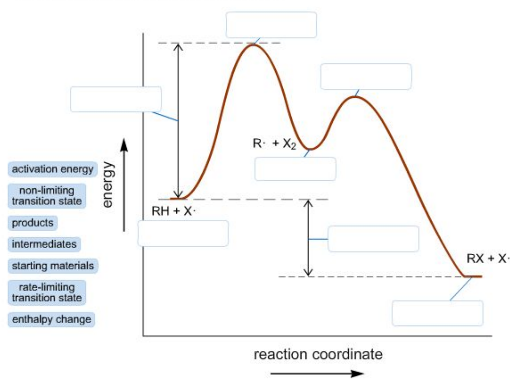
1. Consider the following equilibrium:

$$N_2\,(g) + H_2\,(g) \rightleftharpoons NH_3\,(g)$$

    a. Suppose more hydrogen gas was added to the system. In which direction would the equilibrium shift? Towards the product (Le Chatelier's principle)
    b. This reaction does not produce a good amount of product at room temperature. We heat up the system to make the reaction proceed forwards to produce more NH_3. Is the formation of NH_3 endothermic or exothermic? Endothermic (Le Chatelier's principle)
    c. This reaction, known as the Haber Process, also occurs too slowly to be useful. However, we add a small amount of iron to the reaction chamber to speed up the process. After the reaction, the iron is left unchanged; it is not consumed by the reaction. What is the term for the purpose the iron is serving in the reaction? Catalyst
2. Consider the following reaction energy diagram.
    a. Label the reactants, products, activation energy, transition states, and intermediate.
    b. If you completed this pre-test up to this point, remember that my favorite element is iodine.
    c. Is the reaction endothermic or exothermic? Exothermic; the products are of lower energy than the reactants.

d. Which step is the rate-limiting step in this reaction?

First step is rate limiting due to the higher activation energy.



Redox Chemistry

1. Identify the oxidation states of each atom in the following compounds.
    a. PO_4^{3-} P: +5, O: -2
    b. Fe_2O_3 Fe: +3, O: -2
    c. $Cr_2O_7^{2-}$ Cr: +6, O: -2

Nomenclature

1. Write the molecular formulas for each of the following compounds
    a. Ammonia NH_3
    b. Methane CH_4
    c. Nitrogen monoxide NO
    d. Carbon tetrachloride CCl_4
2. Write the systematic names for each of the following compounds.
    a. N_2O Dinitrogen monoxide
    b. HCl (g) Hydrogen chloride (note that HCl doesn't become hydrochloric acid until it dissolves in water)
    c. $CaCl_2$ Calcium chloride