**CHE2060 Lecture 1 Examples**

***Not graded***

**1.1 Very brief history of the development of chemistry**

**1.2 What is organic chemistry?**

1. When a solution of sodium chloride in water is mixed with silver nitrate, a white precipitate forms immediately. When tetrachloromethane is mixed with silver nitrate, no precipitate is produced. Explain these events in terms of the types of bonds present in the two compounds.

2. Here’s a thought question: Why is carbon the basis of the molecules of life? Why not another atom? What makes carbon especially well suited to be the basis of a wide variety of molecules?

**1.3 Atomic models: nuclear to quantum**

3. Think about the nuclear model of the atom, and where each subatomic particle (proton, neutron and electron) is located. Now think about electrostatic force as it applies in everyday life and the science of physics. How does the nuclear structure of the atom appear to contradict everyday electrostatic force.

**1.4 All about orbitals**

4. While a number of the atoms involved in organic chemistry include d orbitals, none have variable charges like the transition metals do. Why?

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6. How many electrons can be held in the:

3d orbital

2p orbital

1s orbital

**1.5 How orbitals fill: electron configuration**

7. Using the atom phosphorous and box-arrow diagrams, create electron configuration diagrams that specifically violate each of these rules.

 a. Pauli’s exclusion principle

 b. the Aufbau principle

 c. Hund’s rule

8. Write the electron configuration of:

a. atomic chlorine

b. ionic chloride

9. All electrons in each atom occupy an assigned (or “ground-state”) energy level.

a. Can that energy level be increased?

b. If so, how?

c. Is the change permanent or reversible?

**1.6 Basic bonding: valence electrons & molecular orbitals**

10. Match the orbital overlap shown below with the appropriate type of interaction. Which is:

* bonding
* anti-bonding
* non-bonding



11. How would Figure 1.15 of Chapter 1 change if:

the bond were weaker than that of H2

the bond were stronger than that of H2

12. If you were able to measure the energy level caused by moving two atoms towards one
 another, how would you determine the optimal bond length between those two atoms?

13. How would Figure 1.15 of Chapter 1 change if:

the bond were weaker than that of H2

the bond were stronger than that of H2

**1.7 Lewis dot structures of molecules**

14. Is it possible to have a double bond between the carbon and the oxygen of the methoxide ion (CH3O-1)? The charge is on the oxygen atom. Draw a Lewis dot structure and explain.

15. Complete this Lewis dot structure by adding bonds and electron pairs where needed.



16. Draw Lewis dot structures for these ions. Calculate formal charge to determine which atom carries the charge, and what the charge is.

1. C2H5 anion
2. CH3O cation
3. CH6N cation
4. CH5O cation
5. C3H3 anion (Note that all H are on the same C.)

**1.8 Electronegativity & bond polarity**

17. Consider the N—Br bond. Given the fact that nitrogen and bromine have nearly the same electronegativity (3.04 and 2.96), answer the following questions. Is the N—Br bond covalent, polar covalent, or ionic? Explain your answer. Is the N—Br bond a stable bond?

18. Account for the differences in bond length in the table shown below.

|  |  |
| --- | --- |
| **Molecule** | **C – Cl bond length (pm)** |
| H3C – Cl | 178 |
| ClH2C – Cl | 177 |
| Cl2HC – Cl  | 176 |
| FH2C – Cl  | 176 |
| Cl3C – Cl  | 175 |
| F3C – Cl  | 172 |

19. Which of these are purely covalent and which include ionic bonds?



**1.9 Resonance: a critical concept**

20. A commonly accepted structure of benzene is shown below. Experiments show that all of benzene’s bonds are equal in length.

1. So why is the structure shown below not quite accurate?
2. Draw a more accurate representation of benzene.



21. Acetic acid has the formula,CH3COOH.

a. Draw a Lewis dot structure for acetic acid.

b. Does the molecule have resonance?

c. If so, which resonance structure is the major contributor and why?

**1.10 Orbital hybridization: key to carbon’s “flexibility”: sp3, sp2 & sp**

22. Label each statement TRUE or FALSE:

a. The number of hybridized orbitals may differ from the number of orbitals that are blended to make them.

b. Free electron pairs can occupy hybridized orbtials.

c. Bonding hybridized orbitals exert more repulsive power then hybridized orbitals that hold lone (unbonded) electron pairs.

23. Rank these bonds in terms of reactivity, from most reactive to least reactive. Explain your answer.

* Double
* Single
* Triple

24. Considering the repulsion that exists between electrons in different bonds, give a reason why a planar geometry for methane would be less stable than tetrahedral geometry.

**1.11 Free electron pairs & radicals**

25. How does a molecules instability (or high energy state) influence its chemical reactivity and why?

**1.12 VSEPR: classifying molecular geometry & orbital hybridization**

25. Use VSEPR to predict the geometry & orbital hybridization of the following molecules:

PH3

SiH4

CCl4

:CH3-1

BH3

NH4+1

BH4-1

BeF2

CH3+1

26.Indicate which of the following molecules have a dipole moment. Using the geometry of orbital hybridization, draw a three-dimensional representation of the molecule and show the direction of the dipole for the molecule using an arrow.

* CCl4
* CFBr3
* CH3NH2
* CH3Cl
* CH2=CHBr
* CH3OCH3
* CH3CH=CH2
* CH3CHClCH3