1. a) i. Cr or Mn

ii.

X: Cr	[Ar] $4s^1 3d^5$	*or*	Mn: [Ar] $4s^2 3d^5$
X^+: Cr^+	[Ar] $4s^0 3d^5$ or [Ar] $3d^5$	*or*	Mn^+: [Ar] $4s^1 3d^5$

b.   i. Give the number of core electrons for Mg^{2+}. Support your answer with the orbital box diagram for the ground-state electron configuration of Mg^{2+} (no noble gas abbreviation) and briefly (<15 words) explain your reasoning.

Mg^{2+} has 10 core electrons:

↑↓	↑↓	↑↓	↑↓	↑↓

     1s   2s   2p
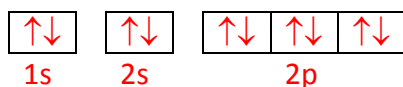
[3]

(or "2 core electrons" if you consider the n=2 electrons the new 'valence' shell)

Core electrons are 'inner shell' electrons (any below valence shell)

ii. Is Mg^{2+} paramagnetic or diamagnetic? Briefly explain.
Diamagnetic. All electrons are paired

c. Why do Li, Na, and K have similar chemical properties?
They all have one valence electron in a *s* orbital: ns^1

2. a. What is the general relationship between the size of an atom and its first ionization energy? Explain the factors that give rise to this relationship.

[5]   The general relationship is that as the size of an atom increases, its first ionization energy decreases. As valence electrons get further away from the nucleus they are held less tightly and are therefore easier to remove.

b.   What are the factors that determine the relative size of atoms? Based on your answer, arrange K, Ga, and Cl in order of increasing atomic radius.
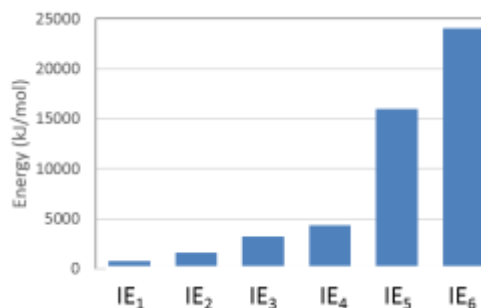
Smallest Radius		Largest Radius
Cl	Ga	K

[2]

c.   The figure below shows the successive ionization energies (IE) for an element in the third period of the periodic table.

Identify the element and justify your choice

Si – jump at 5th ionization energy indicates removal of core electron.

[3]

3. The table below contains the lattice energies (kJ/mol) of various sodium and magnesium compounds.

| Cation/Anion	Fluoride	Chloride	Bromide
Sodium	-940	-780	-740
Magnesium	-2920	-2500	-2400

Why do the magnesium compounds have lattice energies that are so much larger in magnitude than those of the sodium compounds? Provide <u>two</u> factors that cause these differences, specifying in what ways they impact the lattice energies of these sets of compounds.

According to Coulomb's law, lattice energy is evaluated as a product of charges and this term is doubled for the magnesium containing compounds. The second factor is the distance between ions. The magnesium ion is smaller than the sodium ion; both are isoelectronic but the magnesium ion has more protons which reduces the size.

4. The S – O bond length in the sulfur trioxide (SO_3) molecule is measured experimentally as 142pm. The S – O bond length in the sulfite anion (SO_3^{2-}) is longer and is measured experimentally as 151 pm.

(Note: All atoms follow the Octet Rule here. None are hypervalent.)

Explain the difference between the experimentally observed bond lengths using Lewis structures. Draw all relevant structures and include any non-zero formal charges.

SO_3 :
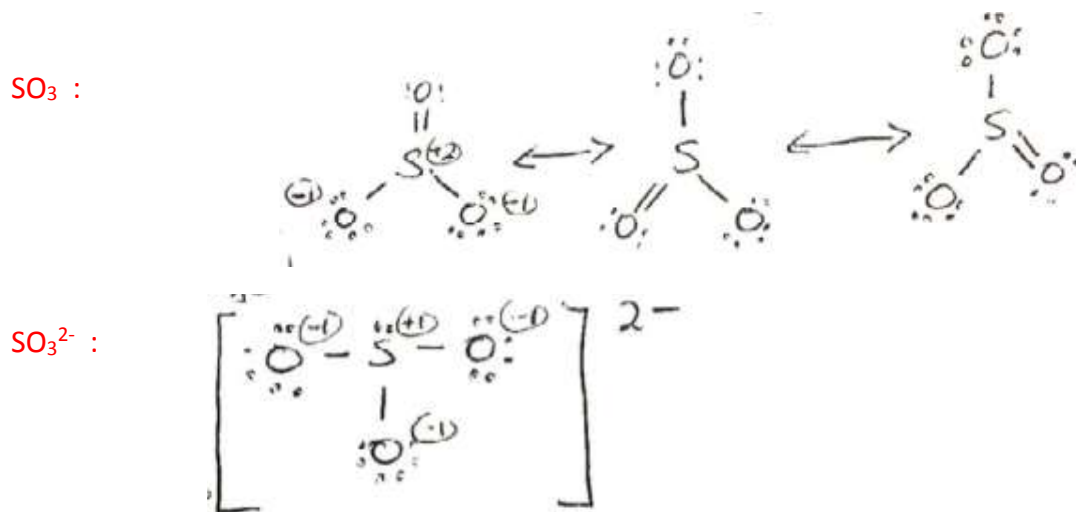
SO_3^{2-} :

S-O bond length depends on type of bond; double bonds are shorter than single bonds.

∴ SO_3 bond length is shorter since S-O bond has contributions of single bonds and double bond in resonance hybrid.

∴ SO_3^{2-} bond length is longer since SO_3^{2-} has only single bonds.

5. Rank the following species by expected F-N-F bond angle, from smallest to largest angle: NF_3, NF_4^+, NF_2^-

NF_2^- $\qquad\qquad$ NF_3 $\qquad\qquad$ NF_4^+
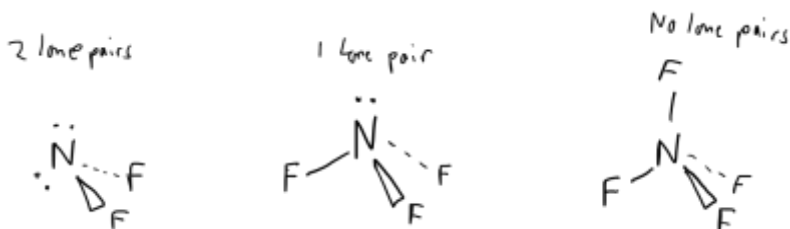
smallest bond angle $\qquad\qquad\qquad\qquad$ largest bond angle

Briefly justify your ranking (30 words or less). Support your explanation with 3-D structures.
[6]

Lone pairs exert greater repulsion than bonding pairs
the N-F bonds are pushed closer together where there are more lone pairs
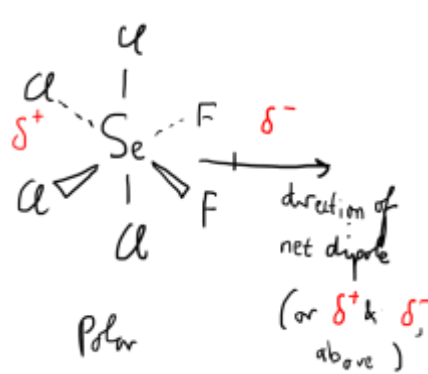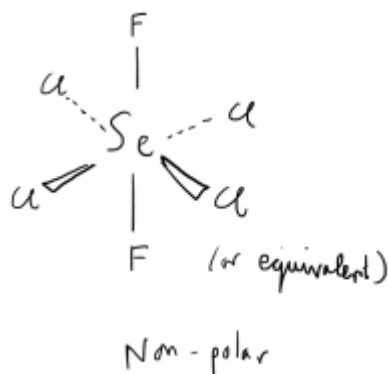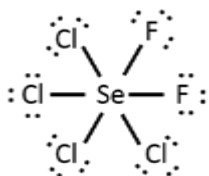
structures:

b. Complete this table by naming the molecular & electron group geometry of the species in a.

|	NF_3	NF_2^-	NF_4^+
Electron group geometry	tetrahedral	tetrahedral	tetrahedral
Molecular geometry	Trigonal pyramidal	Bent / V-shaped / angular	tetrahedral

[6]

c. The Lewis structure of SeF2Cl4 is shown below. There are two possible (3-D) molecular geometries for this compound, only one of which is polar. Draw 3-D representations of both the geometries. Indicate which one is polar and show the net dipole on your drawing.

[6]

6. Using the table provided below, determine the ΔH of reaction for the following reaction:

$$C_2H_2 (g) \; + \; 2F_2 (g) \; \rightarrow \; C_2H_2F_4 (g)$$
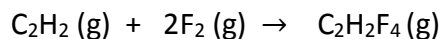
Your work should show all Lewis structures.

Bonds broken: 1 C=C   614 kJ/mol   Bonds formed: 1 C-C   347 kJ/mol

                                   4 C-H   4 × 413 kJ/mol   4 C-H   4 × 413 kJ/mol

                                   1 F-F   154 kJ/mol   2 C-F   2 × 453 kJ/mol

            Total: 2420 kJ/mol   Total: 2905 kJ/mol

$\Delta H°_{reaction} = \; \Sigma \, (\Delta H)_{bonds \; broken} \; - \Sigma \, (\Delta H)_{bonds \; formed}$

                   =   2420 kJ/mol -   2905 kJ/mol

                   =   - 485 kJ/mol

Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)
O-O	142	C≡C	839	C-Br	276
O=O	498	C≡N	891	C-I	240
H-O	464	C=O	745 (799 in CO_2)	F-F	154
H-H	436	C-H	413	Cl-Cl	243
C-C	347	C-F	453	Br-Br	218
C=C	614	C-Cl	339	I-I	149

7. Consider the following orbital box diagram for carbon.
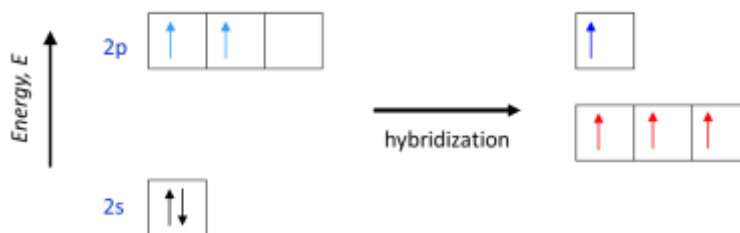
a. What is the hybridization of the carbon atom?

[1] sp^2

a. Justify your answer above in terms of the number of hybridized and unhybridized orbitals.

[3]   There are three equivalent hybrid orbitals (sp^2) and 1 unhybridized p orbital (to make a maximum of 1 pi bond). Three orbitals (two p, one s) were hybridized to make three new ones (three sp^2).

c. On the molecule shown below, indicate (e.g., with an arrow) *ONE* carbon atom which has the hybridization you chose in part a.

[2]
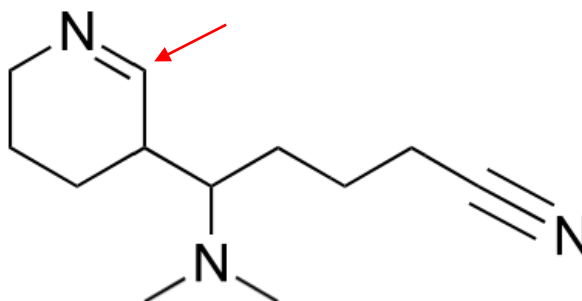
8. Observe the following orbital overlap diagram (hybridization and bonding scheme):

π: C(p)-N(p) (all one bond)

σ: N(sp²)-C(sp³)

σ: C(sp²) – H(s)

σ: C(sp²)-N(sp²)

σ: C(sp³)-H(s)
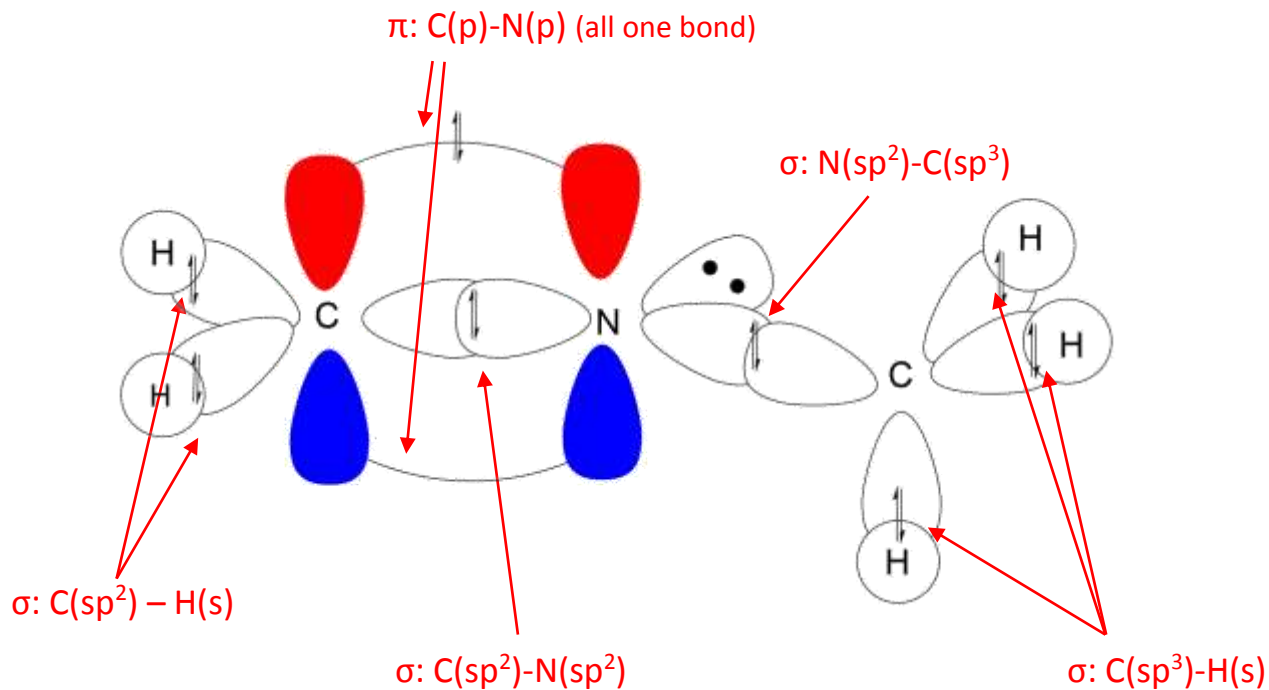
[8]

On the diagram, identify all unique bond types in this diagram and indicate what orbitals are overlapping to create that particular bond. You may use the formal notation shown below to denote the bonds. For example, for the molecule H_2, the covalent bond is a sigma bond created by the overlap of the s orbital from each H atom.
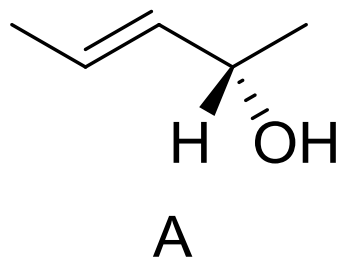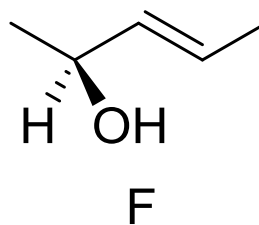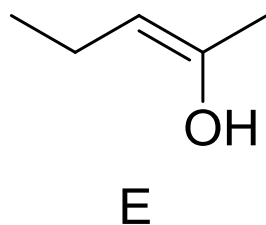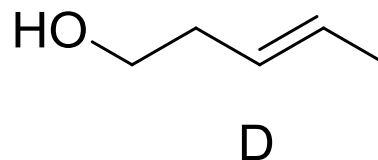
Example (notation used in *Tro*):   σ: H(s)-H(s)
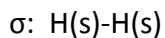
for the bond   H-H
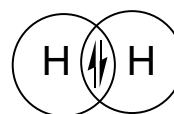
9. Consider Structure A, below:

H   OH

A

Compare it to the various other structures, B to F, below:

H   OH

B

HO   H

C

HO

D

OH

E

H   OH

F
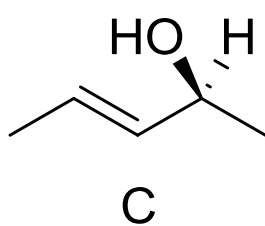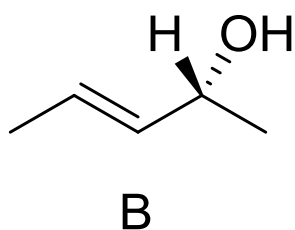
Answer the following questions by indicating option "B", "C", etc. or, if none of the above structures matches, write "none".

a. List all structures (among B to F) that are structural isomers of Structure A: D and E

[3]

b. List all enantiomers of Structure A: B

[3]

c. List all *cis-trans* diastereomers of Structure A: none

[2]

d. Of the six structures shown in this question, how many are chiral?   Four (A, B, C and F)

[2]

10. For the following three molecules (i-iii):
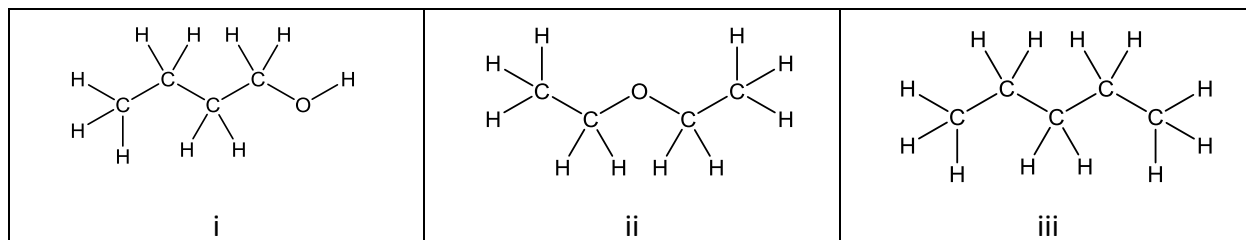
| i | ii | iii |

a. Put these molecules in order from lowest to highest boiling point. Justify your ranking by including all relevant intermolecular forces.

iii ‹ ii ‹ i
i : is a polar molecule (dipole-dipole forces) and also has H-Bonds that will give it a high boiling point
ii: is a polar molecule with dipole-dipole forces
iii : is a non-polar molecule with only dispersion forces making its boiling point lowest.

[6]

b. One of these compounds is more soluble in water than the two others. Identify that compound and explain this observation with the help of the intermolecular forces between solute and solvent (water). Limit your answer to <20 words.

Compound i is the most soluble in water. Its solution with water will have H-bonds (as well as dipole-dipole forces) making the most soluble compound of the three.

[4]

11. The strength of an acid depends upon the stability of its conjugate base.

a. Write the equation for the hydrolysis (dissociation reaction of acid in water) for the generic acid HA and water.

[1] $HA\ (aq) + H_2O\ (l) \rightleftarrows H_3O^+\ (aq) + A^-\ (aq)$

b. Explain how the numerical value of the K_a of an acid reflects the stability of that acid's conjugate base.

[3]   Greater stability of A- increases the ratio between A- and HA in Ka=[H3O+][A-]/[HA] which leads to a larger Ka value. Small Ka, low A-, less stability of A-, weaker acid.

(Question 11 continues on the next page.)

c. This stability of the conjugate base in turn depends upon factors which include size, resonance, and inductive effects. Complete the table below to show which factor is most important in determining which of the two species is the stronger acid. Add a check mark under the most important factor.

[6]

	Size	Resonance	Inductive Effect
Sulfuric acid (H_2SO_4) vs Sulfurous acid (H_2SO_3)		✓	
Hydrofluoric acid vs Hydrochloric acid	✓		
Acetic acid (CH_3COOH) vs chloroacetic acid ($CH_2ClCOOH$)			✓

d. HCO_3^- can act as a weak acid. Draw the complete Lewis structure (including any resonance and any non-zero formal charges) of the conjugate base of HCO_3^- and specifically explain the factors that lead to its stability.

Conjugate base of HCO_3^- is CO_3^{2-} :

[6]
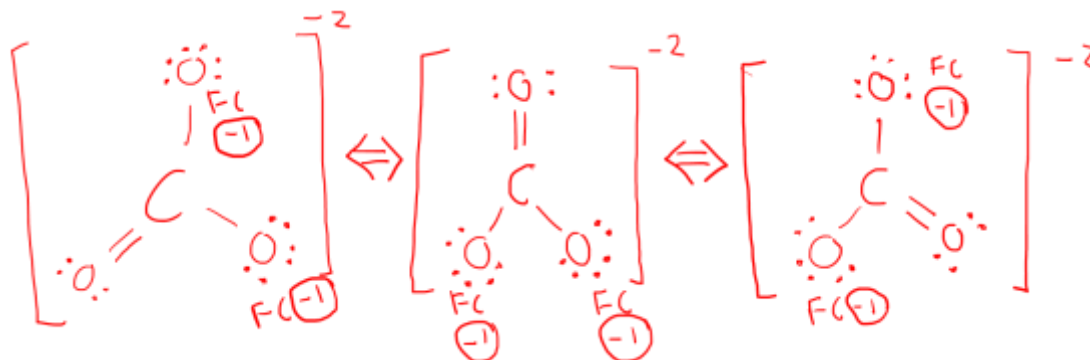
Formal (negative) charge is spread over all three oxygen atoms, reducing charge per atom and so stabilizing the base.