CHEM 110 Final Exam (Old Material)
Practice Test Solutions

1C

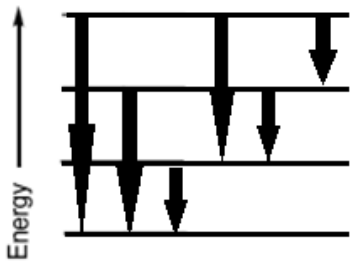
Start with the elemental electron configuration. Cu = [Ar] $4s^2$ $3d^9$. Next Remember that it is better to have half-filled subshells rather than randomly filled so a 4s electron will jump into the d. This exception to the rule occurs for all the group 6B and 1B elements. The new electron configuration is Cu=[Ar] $4s^1$ $3d^{10}$. Next step is to ionize an electron to make this a cation. The ionized electron must come from the highest energy level, which would be 4s. The final electron configuration is Cu^+=[Ar] $3d^{10}$. From here, you can see that there are no unpaired electrons, so this ion is diamagnetic. Since the ion has a positive charge, it is a cation, and this element is obviously a transition metal.

2A

Start by relating energy to wavelength. We know E=hv and $c = \lambda * \upsilon$. We can rearrange these to $E = \frac{hc}{\lambda}$. Solve this for λ and we get $\lambda = \frac{hc}{E}$. The energy in this problem has to do with electrons jumping levels so $E = R_H(\frac{1}{n_i^2} - \frac{1}{n_f^2})$. We can simplify all of this into one equation to answer this problem or we can break it into steps. Breaking it into steps is better to avoid calculator errors. For these problems, use lots of parenthesis to ensure you aren't imputing the data incorrectly. $E = \ 2.18 * 10^{-18} J\left(\frac{1}{6^2} - \frac{1}{2^2}\right) = \ -4.84 * 10^{-19} J.$ The negative sign infers that the energy is flowing out of the system, so we can get rid of it now. $\lambda = \frac{(6.63*10^{-34})(3*10^8)}{4.84*10^{-19}} = \ 410.5 * 10^{-9} m \ or \ 411 \ nm.$ This wavelength is in the Violet range.

3D

For these problems count the number of possible emission lines. Remember photons are emitted when electrons jump from a high level to a lower level. Below is a representation of every possible jump.



4C

To answer this problem refer to Columbs law, it's on the data sheet. $E = \frac{k_l Q_1 Q_2}{d}$. An electron has a (-1) charge and a proton has a (+1) charge. Since k_l is a constant, and we're not looking for exact values, it can be neglected.

I) would be $E = \frac{(-1)(+2)}{3} = -\frac{2}{3}$.

II) $E = \frac{(-1)(-2)}{1.5} = \frac{4}{3}$.

III) $E = \frac{(-1)(+1)}{3} = -\frac{1}{3}$.

The sign on energy values only define direction, so according to magnitudes, II>I>III.

5B

For this problem you must remember atomic trends and the names of the groups: Lanthanides (Not real, maybe "Lanthanoid" which is an f series.), Halogens (Group 7A), Actinides (Not real, maybe Actinoid series), Alkali Metals (Group 1A), Chalcogens (Group 6A). Remember that electronegativity values increase as you go to the right and up on the table. The atomic radii trend is the exact opposite. Radii increases as you progress towards the bottom left of the table. Therefore the highest electro negativities and smallest radii would be the furthest right group, which would be Halogens.

6B

The key to this question is balancing charges on each side and choosing the correct starting point. The question asked for the second ionization energy, so the reactant must be a first-ionized Mg, which is Mg^+. After this Mg+ is ionized, the products should be a double ionized Mg and an electron, so Mg^{2+} and e^-. ($Mg^+ \rightarrow Mg^{2+} + e^-$) Answer choice (c) is not correct because only 1 electron is being ionized at a time and the charges do not work.

7C

Set up a system of equations and solve them. Let's set the abundance of ^{81}Br to be (a) and the abundance of ^{77}Br to be equal to (b). Now we know:

$$Atomic\ Mass = \Sigma(mass * abundance).$$

Atomic mass of Br is 79.904 (from the periodic table). So [79.904=81a + 77b].
Also, since (a) and (b) are the only abundances, they must add to 100%. [a+b=1].

Here is the system:

1) [79.904=81a + 77b]
2) [1= a + b]

You are free to use any method you choose to solve the system. Here is one way:

Multiply equation 2 by (-77) then add it to equation 1.

(-77) *[1= a+b] → [-77=-77a -77b]

[-77=-77a-77b] + [79.904=81a + 77b] = [2.904=4a]

a=0.726= 72.6% This is the abundance of 81Br.

8B

First convert Watts to Joules.

$$5\ Watts * \frac{1\ \frac{J}{s}}{1\ W} * 60s = 300\ J$$

This is work being done by the system, so the sign for this will be negative. (-300J)
The problem also states that 500 J of heat is entering the system, the sign convention for this would be (+500J).
Finally:

$$\Delta E = q + w = (+500) + (-300) = 200\ J$$

9C

Electron affinity is the energy change when an electron is forced onto an atom: when anions are formed.
I like to compare it to the price the element would pay to get an electron. Fluorine has a very high electronegativity so it really wants an electron, so it is willing to pay more for it hence the electron affinity value of fluorine is very high. Some elements don't want electrons (low electronegativities) so it is going to cost you money (energy) to make that element an anion.

In summary, electron affinity values can be positive or negative depending on the electronegativities of the particular element.

10C

The Heisenberg Uncertainty Principle states that it is impossible to measure position and velocity at the same time. This principle also states that the Bohr model is bologna because it would be inappropriate to imagine electrons moving in orderly patterns.

11B

The answer is MgBr because it should be $MgBr_2$. It's subtle. Remember when naming ionic compounds, name the cation first then the anion with an –ide suffix.

12D

This type of molecule is called a hydrate molecule. For every molecule of $KAl(SO_4)_2$ there will be 12 water molecules attached. We could also say that for every 1 Al, we have 20 Oxygen molecules (8 from $(SO_4)_2$ and 12 from H_2O). So in 2 grams of alum we would have 20*(# of Aluminum atoms in 2 grams) or $20 * (2.6 * 10^{21}) = 5.2 * 10^{22}$ Oxygen atoms. Six grams of alum would just be 3 times greater than the current number.

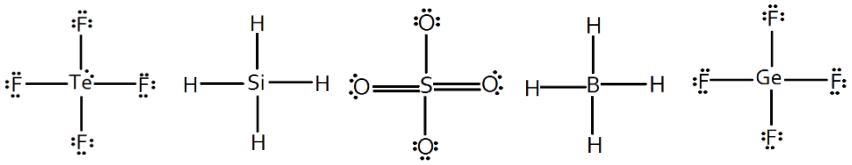
$$\#\ Oxygen\ atoms\ in\ 6\ grams = 3 * (5.2 * 10^{22}) = 1.56 * 10^{23}\ atoms\ of\ O$$

13C

Electron affinity is the energy change when an electron is forced onto an atom: when anions are formed. The reactants must be a neutral element and an electron and the products must be an anion. The only answer choice that satisfies this is (c).

14A&C

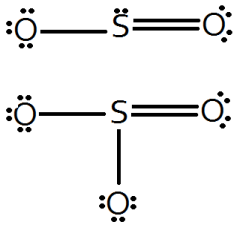
To solve this problem you must draw the Lewis Structures. Below are each of the Lewis structures.



SO_4^{2-} could be drawn without breaking the octet rule, but this structure has more acceptable formal charges.

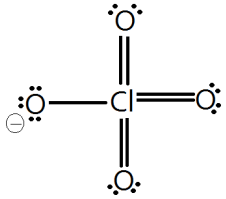
15A&B

Below are the Lewis structures for these molecules, the bonds lengths in each molecule will be the same because of resonance. The bond lengths in SO_2 are essentially 3/2= 1.5 and the bond lengths in SO_3^{2-} are 4/3=1.33. These values are called the bond order and are a result of resonance and delocalized bonding. Answers A and B are both correct. With answer choice A, make sure they mention the correct number of resonance forms. (There would be three resonance forms for SO_3^{2-} and 2 for SO_2). For B) each of the SO_3^{2-} bonds are 1.33 which is longer and weaker than the SO_2 (1.5) bonds.



16E

Below is the correct Lewis structure for perchlorate.



17D

To answer this problem, first measure the wavelength. In the diagram, two complete wavelengths are measured, so the wavelength would be half of this dimension. $\lambda = \frac{(6.1*10^{-7})}{2} = 335 * 10^{-9} m$. Next we have to convert wavelength to energy with this equation $E = \frac{\hbar c}{\lambda} = \frac{(6.63*10^{-34})(3*10^{8})}{355*10^{-9}} = 560.28 * 10^{-21} J$. We calculated the energy associated with 1 wavelength, but each of the answer choices have units of J/mole. It is necessary to convert to the energy associated with a mole of wavelengths to match units.

$$E = \frac{560.28 * 10^{-21} J}{1\ Wavelengt\hbar} * \frac{6.022\ * 10^{23}\text{Wavelengths}}{1\ \text{Mole}} = 3.92 * 10^{5}\ \frac{J}{mole}.$$

18E

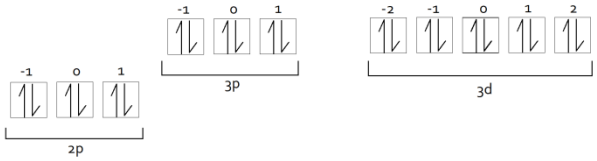
First start with the electron configuration of Cu. Keep in mind copper follows on of the d orbital exception rules.

$$Cu = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$$

Next, determine where m_l orbitals can be located. We know $m_l = -l\ to + l$ so l must be greater than 0. So we will have a m_l of +1 when $l = 1,2,3$. In this electron configuration we see $l = 1,2$ with the p and d orbitals. Here are the subshells with +1 orbitals.

$$2p^6\ 3p^6\ 3d^{10}$$

Below is an electron diagram of these electrons. Simply count the electrons in +1 orbitals.

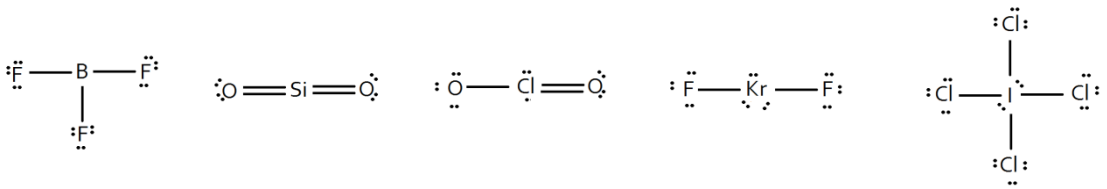


19E

Analyze the molecule to determine which bond is the most polar. Is it the C-Cl, C-I, C-Br, or C-F? Since C is the same in each of these bonds, we just need to compare electronegativity values. Whichever bond has the most electronegative element in it will be the most polar. Since F is the most electronegative element on the periodic table, the F-C bond is the most polar. This means in the overall molecule, F will attract the most electrons and have the highest partial negative charge.

20B

Draw the molecules to determine the correct answer. As your knowledge progresses, you will be able to recognize patterns to answer this question faster and easier. For instance, if more than one element is bonded to a central atom that belongs in Group 7, the octet rule cannot be satisfied. Noble gasses have a fulfilled octet when they aren't bonded, so bonding them will only break the octet rule. And always recognize BF_3 as an exception. Count the total number of electrons for each of these molecules and you'll find BF3 doesn't have enough to complete the octet, and ClO_2 will be an odd numbered molecule. Here are the Lewis structures regardless. There are other acceptable Lewis structures of ClO_2, but they all violate the octet rule.



21.

A. TeF_4
C. SO_4^{2-}

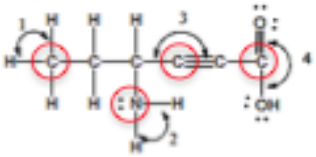
To solve this problem you must draw the Lewis Structures. Below are each of the Lewis structures.

SO_4^{2-} could be drawn without breaking the octet rule, but this structure has more acceptable formal charges.

22.

(e) 3 > 4 > 1 > 2

In order to determine the bond angles, start by identifying the central atom associated with each bond angle. The central atom is the vertex of the angle. For Bond Angles 1, 3, and 4, the central atom is a Carbon. The central atom in Bond Angle 2 is Nitrogen. These central atoms are marked below.



Now to determine the bond angles, you must determine the geometry of each separate bond.

Bond Angle 1:
The central atom C has 4 different atoms bonded to it (AX_4E_0) tetrahedral= 109.5°

Bond Angle 2:
The central atom N has 3 different atoms bonded to it (AX_3E_1) tetrahedral= <109.5°
Remember that lone electron pairs cause the bond angles to slightly decrease.

Bond Angle 3:
The central atom C has 2 different atoms bonded to it (AX_2E_0) linear = 180°

Bond Angle 4:
The central atom C has 3 different atoms bonded to it (AX_3E_0) tetrahedral= 120°

23.

(a) Ice is has a lower density than water.
(c) The boiling point of water is higher than that of H2S.
(e) Both A and C

Here we have a theory/ memorization problem.
Ice has a lower density than liquid water because when water begins to freeze, the molecules begin to orient themselves in a strict pattern because of the hydrogen bonding. When most compounds freeze, the molecules become very closely packed, but the hydrogen bonding in water forces water to actually expand when it freezes. For all extensive purposes, we will say that only water exhibits this trait. (Memorize)

Regarding the next two answer choices, we have to look at intermolecular forces. The trend for this goes as follows:

Stronger/more intermolecular forces → Higher Boiling/Melting Point.

Recall that a phase changes occur when the molecules gain energy and move apart from each other. So to understand this concept, imagine this process occurring, then ask: Is it easier or harder for the molecules to move away from each other when there are strong intermolecular forces. The answer is "it is more difficult" so that means more energy (heat) must be supplied for the phase change to occur.

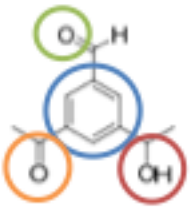
H_2O has hydrogen bonding, LDF, and Dipole-Dipole; while H_2S only has LDF and Dipole- Dipole.

24.

D. I, II, III, V,

Go through them individually and find them in the molecule. They are shown below circled. This is a memorization problem.

I) Aromatic- an unsaturated ring structure. BLUE
II) Alcohol- An OH bond in the chain with nothing else around it. RED
III) Aldehyde- Double Bond O at the END of a chain. GREEN
IV) Alkane- All single bonds- Obviously not.
V) Ketone- Double bond O in the middle of a chain. ORANGE
VI) Carboxylic Acid- Looks like an aldehyde and an Alcohol – not present.



25

(e) E = I, T-Shaped

First determine the shape, and then decide what the center molecule is.

The shape is T-Shaped. You can see the center molecule has 5 electron domains (Trigonal Bi-pyramidal) and 2 lone election pairs meaning T- Shaped. To count electron domains, just count how many bonds the central atom is forming and count the number of lone electron pairs.

Or you can use the AXE formula to see that it is AX_3E_2 which corresponds to T shaped. Remember A stands for the center atom, X is the number of atoms bonded to the central atom, and E is the number of lone electron pairs on the central atom. (As an aside, Adding X and E will also tell you the number of electron domains.)

Now that we know the molecule is T-shaped, we can eliminate each answer choice except B and E leaving Sulfur and Iodine as the only possibilities for E. Again there are different methods to accomplish this, but the best way is to check the formal charge on each element if they were the central molecule.

If sulfur were the central atom, the formal charge would be:

(6 valence)- (3 bonds)-(4 lone electrons) = -1

If I were the central atom, the formal charge would be:

(7 valence) – (3 bonds)-(4 lone electrons) = 0

Iodine would be the best fit for the central atom.

26.

(a) C_3H_4O

Find the percent composition of each answer choice. This is the most assured way to complete this problem. You could also draw them to check which do not have a double bond.

$$a)\ mass\ \%\ C = \frac{3(12g)}{3(12g) + 4(1g) + 1(16g)} = 64.3\%$$

$$b)\ mass\ \%\ C = \frac{3(12g)}{3(12g) + 7(1g) + 1(16g)} = 61.0\%$$

$$c)\ mass\ \%\ C = \frac{3(12g)}{3(12g) + 2(1g) + 1(16g)} = 66.6\%$$

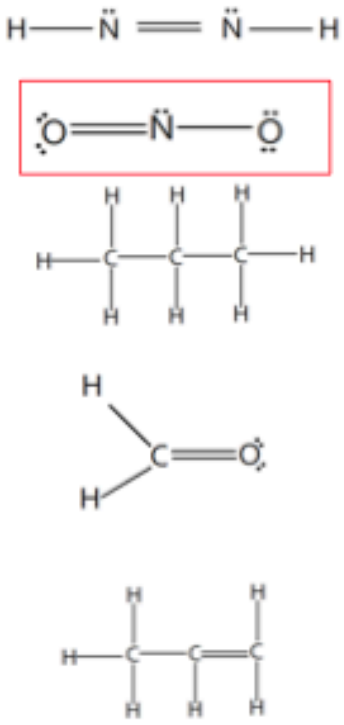
$$d)\ mass\ \%\ C = \frac{4(12g)}{4(12g) + 7(1g) + 1(16g)} = 67.6\%$$

27.

(b) NO_2

We only have delocalized π bonding when the structure has resonance. Recall that resonance occurs when electrons have the ability to move without changing anything else in the molecule.

Here are the Lewis structures:



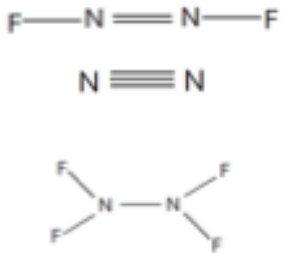
Only NO_2 can have delocalized π bonding. No other molecule can move the double bond without also moving a hydrogen, or breaking the octet.

28.

(d) N_2 N_2F_2 N_2F_4

The magnitudes of the lengths given do not physically mean anything to us in this problem. If you look at the answer choices, you see that the lengths are arranges in increasing order, so the purpose of this problem is to arrange the molecules by increasing bond length. In order to accomplish this, you have to draw each molecule.

Here are the molecules (lone electrons not shown):



Since N_2 has a triple bond, the N-N bond length is the smallest. The double bond in N_2F_2 is shorter that the single bond in N_2F_4. So $N_2 < N_2F_2 < N_2F_4$

29.

(d) DA

Recall that polarity is defined as the difference of electronegativites. To solve this problem, find the difference between of the electronegativites; the highest magnitude is the most polar.

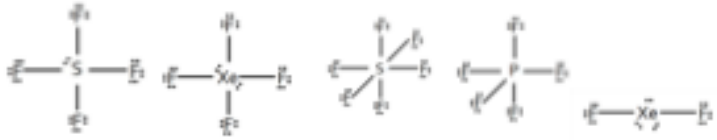
a) B – C = (2.01 – 2.04) = 0.03
b) A – B = (1.01 – 2.01) = -1
c) C – D = (2.04 – 3.05) = -1.01
d) D – A = (3.05 – 1.01) = 2.04
e) B – D = (2.01 – 3.05) -1.04

30.

(a) SF_4

Draw each molecule. Usually with this type of problem, only one of the answer choices will actually be polar.

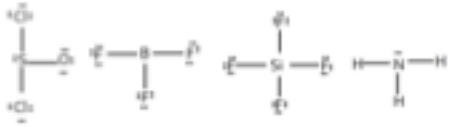
(a) SF_4 – See- Saw, Always Polar
(b) XeF_4 - Square Planer, non- Polar
(c) SF_6 – Octahedral, non-Polar
(d) PF_5 – Trigonal Pyramidal, non-Polar
(e) XeF_2- Linear, non-Polar.



31.

(b) 1 and 4 only

Draw them. Determine which is polar.



1. $SOCl_2$ Pyramidal- Always Polar
2. BF_3 Trigonal Planar, non-polar
3. SiF_4 tetrahedral, non-polar
4. NH_3 pyramidal, Always polar

32.

(a) The π electrons in SO_3^{-2} are delocalized. - TRUE
This is true. Structures with resonance have delocalized bonding.

(b) All C–C bond lengths in benzene are the same. - TRUE
For structures with resonance, we look at the bond order which is similar to an average bond length.

(c) More energy is required to break a carbon–nitrogen triple bond than energy needed to break a carbon–nitrogen double bond. – TRUE
Triple bonds require more energy to be broken.

(d) Hybridization allows for all F–C–F bond angles in CF_4 to be the same. – TRUE
It's the reason for hybridization.

(e) The bonding in benzene, C_6H_6, includes 12 sigma (σ) bonds and 3 localized pi (π) bonds. – FALSE
The pi bonds are delocalized.

33.

(e) None of the above.

Count the number of electron domains on each carbon, then assign the correct hybridization. You can use AXE to help you. (A= central atom, X= number of elements bonded to central atom, E= number of lone electron pairs on central atom; Number of electron domains is X + E) Also, recall the sum of the hybridization (s + p + d) must be equal to the number of electron domains.

C1 → AX_3E_0 → 3 electron domains → sp^2
C2 → AX_3E_0 → 3 electron domains → sp^2
C3 → AX_2E_0 → 2 electron domains → sp

34.

(a) TeF_4

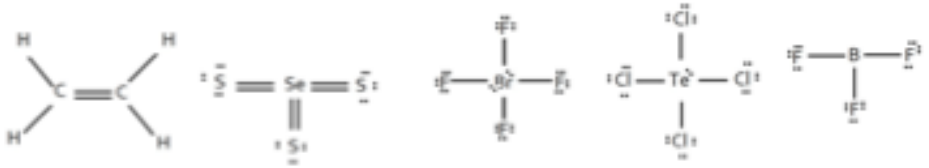
Believe it or not, you have to draw the Lewis structures for each of the answer choices! Whichever does not have 4 electron domains is the answer.

Here are the Lewis structures:

35.

(d) $TeCl_4$

There is a really quick and easy way of answering this problem. Just kidding. Draw each Lewis structure and determine which is not planar. Any electron domains over 3 can be 3D.



(a) C_2H_4 – Organic and planar. (no electron domains over 3)
(b) SeS_3 – Trigonal Planar
(c) BrF_4^- - Square Planar
(d) $TeCl_4$ – See Saw – 3D
(e) BF_3 – Trigonal Planar

36.

(b) Each carbon is sp^2 hybridized.

Look at each carbon and determine the number of electron domains for each. Every carbon will have 3 things bonded to it and no electron pairs. Therefore, each carbon is sp^2 hybridized. Do not forget to fill in the missing hydrogens!

37.

(d) I

Fill in the missing hydrogens then write the formula. It will be C_6H_5OX. Now find the mass percent for C and solve for the unknown.

$$mass\ \%\ Carbon = 0.327 = \frac{6(12g)}{\left(6(12g) + 5(1g) + 1(16g) + 1\ (mass\ x)\right)}$$

Solve for mass x

$$0.327 = \frac{72}{93 + mass\ x}$$

$$(93 + mass\ x) = \frac{72}{0.327}$$

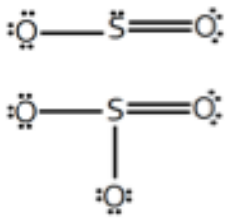
$$(mass\ x) = 220.2 - 93 = 127.2$$

So X must be Iodine.

38.

(a) All S–O bonds in SO_3^{2-} are the same because of resonance.
(b) Each of the S–O bonds in SO_3^{2-} are weaker than S–O bonds in SO_2.

Below are the Lewis structures for these molecules, the bonds lengths in each molecule will be the same because of resonance. The bond lengths in SO_2 are essentially 3/2= 1.5 and the bond lengths in SO_3^{2-} are 4/3=1.33. These values are called the bond order and are a result of resonance and delocalized bonding. Answers A and B are both correct. With answer choice A, make sure they mention the correct number of resonance forms. (There would be three resonance forms for SO_3^{2-} and 2 for SO_2). For B) each of the SO_3^{2-} bonds are 1.33 which is longer and weaker than the SO_2 (1.5) bonds.



39.

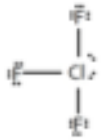
(b) Trigonal bipyramidal

The molecule only contains Cl and F, so the $mass\ \%\ Cl + mass\ \%\ F = 100$. since we know mass percent Cl, we can solve for mass % F. (61.65 %)
Now we can assume we have 100 grams of the substance meaning we have 38.35 grams of Cl and 61.65 grams of F. With these masses, we can determine the number of moles, and then find the mole to mole ration known as the empirical formula.

$$38.35\ g\ Cl * \frac{1\ mole\ Cl}{35.45\ g\ Cl} = 1.08\ mole\ Cl$$

$$61.65\ g\ F * \frac{1\ mole\ F}{18.99\ g\ F} = 3.25\ mole\ F$$

Divide through by the smallest number of moles to get the ratio → ClF_3
Now draw this molecule:



The central atom Cl has 5 electron domains so this molecule is Trigonal bi-pyramidal. (AX_3E_2 → X + E = 5 electron domains.)

40.

A. I > II > III

When dealing with boiling points on this exam, we must look at intermolecular forces. Recall that more intermolecular forces mean a higher boiling point. Hydrogen bonding is stronger that dipole-dipole which is stronger than LDF. EVERYTHING has LDF and it increases in strength as mass increases.

For this problem, none of the choices have hydrogen bonding or dipole-dipole interactions. Therefore we must rely solely on LDF forces. The boiling point will increase as the mass increases. The masses are as follows:

$I)\ 8(12) + 18(1) = 114\ g$
$II)\ 6(12) + 14(1) = 86\ g$
$III)\ 4(12) + 10(1) = 58\ g$

I will have the highest boiling point and III will have the lowest based on LDF and mass.

Problem 41:

(a) 1.02 m $CaCl_2$

When facing a problem that refers to solutions and a change in boiling point, you should immediately refer to:

$$\Delta T_b = K_b m\, i$$

DO NOT forget the Van Hoff Factor i. Remember that

$$i = number\ ions\ solute\ dissolves\ into$$

We already know

$$\Delta T_b = (101.56 - 100) = 1.56\ and\ K_b = 0.51$$

With this, we can solve for

$$mi = \frac{\Delta T_b}{K_b} = \frac{1.56}{0.51} = 3.06.$$

Now we must find the answer choice that matches $mi = 3.06$.

For Answer choice (a)

$$m = 1.02, and\ i = 3\ so\ mi = 3.06$$

Problem 42:

(e) 87.9 g CCl_2F_2

To answer this problem, we must refer to conservation of energy. In a closed system, no energy is created or destroyed; or energy lost by the system must be gained by the surroundings. All of the energy absorbed by the evaporating CCL_2F_2 must have been released by the H_2O. So we will begin by defining the system as the CCL_2F_2 and the surroundings as H_2O. Now we know:

$$q_{sys} = -q_{surr}$$

And:

$$q_{sys} = m\Delta H_{vap}$$

For the surroundings, the water must first cool, and then completely freeze.

$$q_{surr} = q_{cooling} + q_{freezing}$$

$$q_{cooling} = mC_p\Delta T = (64g)\left(4.184\frac{J}{gK}\right)(0-15) = -4.02\ kJ$$

$$q_{freezing} = m\Delta H_{freezing} = (64g)\left(-334\frac{J}{g}\right) = -21.4\ kJ$$

$$q_{surr} = -4.02\ kJ + -21.4\ kJ = -25.42\ kJ$$

Going back to the first equation:

$$q_{sys} = -q_{surr}$$

$$m\Delta H_{vap} = -(-25.42\ kJ)$$

Solve for m:

$$m = \frac{25.42 * 10^3\ J}{289\frac{J}{g}} = 87.9\ g$$

Problem 43:

(d) 1.30 atm

Whenever you see the words pure and pressure in the same sentence (or anything along those lines), refer to Raoult's Law:

$$P_a = X_a P_a°$$

This law accounts for the pressure change associated with the addition of more intermolecular forces.

$$P_a\ (Vapor\ pressure\ of\ solution) <\ P_a°\ (Vapor\ pressure\ of\ pure\ specimen)$$

We need to find:

$$X_a = \frac{moles\ of\ solvent}{moles\ of\ solvent + moles\ of\ solute}$$

$$moles_{solvent} = 200g\ C_3H_8O * \frac{1\ mole\ C_3H_8O}{3(12) + 8(1) + 1(16)\ g\ C_3H_8O} = 3.33\ moles_{solvent}$$

$$moles_{solute} = 85.2\ g\ C_3H_8O_3 * \frac{1\ mole\ C_3H_8O_3}{3(12) + 8(1) + 3(16)\ g\ C_3H_8O_3} = 0.926\ moles_{solute}$$

$$X_a = \frac{3.33}{3.33 + 0.926} = 0.782$$

To solve:

$$P_a° = \frac{P_a}{X_a} = \frac{1.02\ atm}{0.782} = 1.30\ atm$$

Problem 44:

(e) None of the above

I) False - The solubility is of solids is determined by the molecules and temperature.
II) False- Solubility of gases increase the temperature decreases.
III) False- Solubility of gasses decrease as pressure decreases.

Problem 45:

(d) $KClO_3$

The solution with the lowest concentration of Cl^- ions will have the lowest solubility at the given temperature. Find the temperature on the x-axis and find the line that is closest to 0 on the y- axis. (The lowest line has the lowest concentration)

Problem 46:

(b) 0.27 atm

We need to solve for partial pressure given a concentration. We know an equation that relates these two:

$$C_g = k_H P_g \quad meaning\ Concentration_{gas} = (Constant) * \left(Partial\ Pressure_{gas}\right)$$

$$P_g = \frac{C_g}{k_H} = \frac{2.3 * 10^{-3}\ M}{8.3 * 10^{-3}\ \frac{M}{atm}} = 0.27\ atm$$

Problem 47:

(c) O_2 at 10 atm and 1000 K

Gasses behave most ideally at HIGH temperatures and LOW pressures.
Why?
Just because.

Problem 48:

(d) 0.58

This is a typical problem to be found on the test. A reaction takes place under water, and the bubbles are collected in an upside down cup. Initially, there is no gas in the upside down cup. As the bubbles are collected, the volume and pressure of the gas within the cup can be measured. With these measurements, we can perform our calculations; however, there is one thing to account for beforehand. The total pressure in the upside down cup comes from the gas released in the reaction AND the vapor pressure of the H_2O:

$$P_{total} = P_{I_2} + P_{H_2O}$$

So:

$$P_{I_2} = P_{total} - P_{H_2O} = (0.267 - 0.11\ atm) = 0.157\ atm$$

We need to find the mole fraction of I_2 so we will refer to the partial pressure equations:

$$P_{I_2} = X_{I_2} * P_T$$

$$X_{I_2} = \frac{P_{I_2}}{P_T} = \frac{0.157\ atm}{0.267\ atm} = 0.58$$

Problem 49:

(b) 3

With rates of effusion use:

$$\frac{r_1}{r_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

We will set r_1 to be rate of H_2O. So solve for r1 and plug in the values of the molar masses.

$$r_1 = \frac{\sqrt{M_2}}{\sqrt{M_1}} * r_2 = \frac{\sqrt{18}}{\sqrt{2}} * r_2 = 3r_2 = r_1$$

Problem 50:

(c) Liquid C

The temperatures next to the liquids are the boiling points of each liquid. Boiling point and vapor pressure are inversely proportional. As boiling point increases, vapor pressure at a given temperature will decrease. So the highest vapor pressure at a constant temperature will belong to the liquid with the lowest boiling point.

Problem 51:

(f) 0.716 g

(K_f for water is 1.86°C/m)

When facing a problem that refers to solutions and a change in freezing point, you should immediately refer to:

$$\Delta T_f = K_f m\, i$$

DO NOT forget the Van Hoff Factor i. Remember that

$$i = number\ ions\ solute\ dissolves\ into$$

We already know

$$\Delta T_f = \big(0 - (-2)\big) = 2$$
$$K_f = 1.86$$
$$i = 3$$

So solve for m:

$$m = \frac{\Delta T_f}{K_f i} = \frac{2}{1.86 * 3} = 0.358 \frac{moles\ solute}{kg\ solvent}$$

We can now solve for the moles of solute knowing that we have 1 mole of solvent (.018kg) and the molar mass of $CaCl_2$.

$$0.358 \frac{moles\ solute}{kg\ solvent} * (0.018\ kg\ solvent) * \frac{1(40.1)+2(35.45) g\ CaCl_2}{1\ mole\ CaCl_2 (solute)} = 0.716\ g\ CaCl_2$$

Problem 52:

(e) 0.174 g

We know:

$$M_{solution} = \frac{moles_{solute}}{Liters_{solution}}$$

Find the moles of the solute:

$$moles_{solute} = MV = (1.25 * .01) = 0.0125\ moles\ Li_2CO_3$$

Now convert to mass of Li^+

$$0.0125\ moles\ Li_2CO_3 * \frac{2\ moles\ Li^+}{1\ mole\ Li_2CO_3} * \frac{6.941\ grams\ Li^+}{1\ mole\ Li^+} = 0.174\ g\ Li^+$$

Problem 53:

(b) Ar

When dealing with speed, refer to this equation:

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

Since R and T are constants, we see that speed is inversely proportional to the molar mass. Notice that the graph has speed on the x-axis. Lines with peaks further right will be have a faster speed.

Line A molecules are traveling the slowest, so they must have the highest mass. Line C is second fastest; therefore, it must be the second lightest.

Problem 54:

(d) Vapor Pressure

Intermolecular forces are the forces holding each molecule to one another. If these forces are very strong, it is difficult for molecules to break free. As intermolecular forces increase, the substance will want to become more of a solid. (This is an exaggeration, but it helps).

High intermolecular forces will raise the boiling point because it will take more energy to free the molecules from the liquid phase into to gas phase.

High intermolecular forces will raise the freezing (melting) point because it will take more energy to free the molecules from the solid phase into to liquid phase.

Viscosity is the resistance to flow. Solids do not flow. More intermolecular forces, means more solid like.

Vapor pressure will DECREASE as vapor pressure increases.

Cohesive forces are intermolecular forces.

Problem 55:

(b) H_2O

This diagram cannot represent H_2O because the line separating solid and liquid has a positive slope. The diagram for H_2O has a negative slope (points up and to the left) because the liquid phase is denser than the solid phase.

Problem 56:

(a) 86.56 J/(mol·K)

In this problem, we are looking at energy changes during the heating of a single phase. This equation must be used:

$$q = mC_p\Delta T$$

We want to find the molar heat capacity so we can solve for C_p and ensure m is in moles.

$$100\ C_9H_{20} * \frac{1\ mole\ C_9H_{20}}{9(12) + 20(1) g\ C_9H_{20}} = 0.781\ mole\ C_9H_{20}$$

$$C_p = \frac{q}{m\Delta T} = \frac{234\ J}{0.781\ mole * 3.46K} = 86.59 \frac{J}{mol * K}$$

Keep in mind that $\Delta K = \Delta\ ^\circ C$

Problem 57:

(a) 3.94 x 10^{22}

We are given V, P, and T. Sounds like a classic PV=nRT equation with n as an unknown. This works because there is a simple stepping stone to convert from moles to molecules.

$$n = \frac{PV}{RT} = \frac{1.02\ atm * 1.62\ L}{0.08206 \frac{L*atm}{mole*K} * (\ 308.15\ K)} = 0.06535\ moles\ air$$

$$0.06535\ moles * \frac{6.022 * 10^{23}\ molecules}{1\ mole} = 3.94 * 10^{22}\ molecules$$

Problem 58:

(d) 0.43

Mole Fraction:

$$X_a = \frac{moles\ a}{Total\ moles} = \frac{(moles\ methylamine)}{moles\ methylamine + moles\ ammonia}$$

$$93\ g\ CH_5N * \frac{1\ mole\ CH_5N}{1(12) + 5(1) + 1(14)\ g\ CH_5N} = 3\ moles\ CH_5N\ (methylamine)$$

$$68\ g\ NH_3 * \frac{1\ mole\ NH_3}{1(14) + 3(1) g\ NH_3} = 4\ moles\ NH_3\ (ammonia)$$

$$X_{CH_5N} = \frac{3}{3+4} = 0.43$$

Problem 59:

(a) 117 nm

For light and energy we know:

$$E = hv$$

For light and wavelength, we know:

$$c = \lambda * v$$

We can combine these and solve for lambda:

$$E = \frac{hc}{\lambda} \qquad \lambda = \frac{hc}{E}$$

Energy must be in J/bond

$$1023\frac{kJ}{mol} * \frac{1000\ J}{1\ kJ} * \frac{1\ mole}{6.022 * 10^{23}\ bonds} = 1.698 * 10^{-18}\frac{J}{bond}$$

Now simply solve for wavelength:

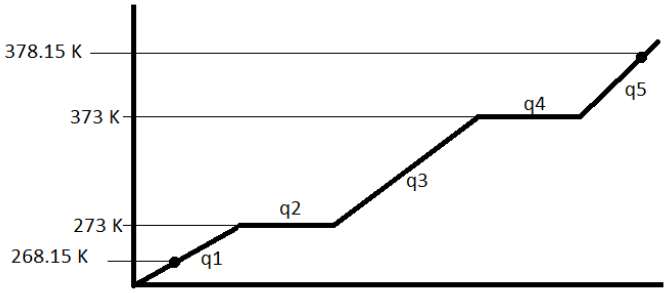
$$\lambda = \frac{(6.63 * 10^{-34})(3 * 10^{8})}{1.698 * 10^{-18}} = 117\ nm$$

Problem 60:

(e) 152 kJ

This problem includes 5 energy transitions:

- Heating of a solid: $q_1 = mC_p\Delta T$
- Melting phase change: $q_2 = m\Delta H_{fus}$
- Heating of a liquid: $q_3 = mC_p\Delta T$
- Vaporization phase change: $q_4 = m\Delta H_{Vap}$
- Heating of a gas: $q_5 = mC_p\Delta T$



Be sure to use the correct heat capacity for each different phase, and be sure that check units. Total heat will be $q_1 + q_2 + q_3 + q_4 + q_5 = q_{total}$

$$solid: q_1 = (50g)\left(2.03\frac{J}{g*K}\right)(273.15 - 268.15) = 507.5\ J\ or\ 0.508\ kJ$$

$$melting: q_2 = \left(\frac{50g}{18\left(\frac{g}{mol}\right)}\right)\left(6.01\frac{kJ}{mol}\right) = 16.69\ kJ$$

$$liquid: q_3 = (50g)\left(4.18\frac{J}{g*K}\right)(373.15 - 273.15) = 20900\ J\ or\ 20.9\ kJ$$

$$vaporize: q_4 = \left(\frac{50g}{18\left(\frac{g}{mol}\right)}\right)\left(40.67\frac{kJ}{mol}\right) = 112.97\ kJ$$

$$gas: q_5 = (50g)\left(1.84\frac{J}{g*K}\right)(378.15 - 373.15) = 460\ J\ or\ 0.460\ kJ$$

$$q_{total} = 0.508 + 16.69 + 20.9 + 112.97 + 0.460 = 152\ kJ$$