

Solutions - Practice Test 1 - CHEM 110 Exam 3

1A

Noble gases are entirely uncharged, and do not have polar covalent bonds, ionic bonds, or dipole moments. The only intermolecular forces that exist in noble gases are London dispersion forces. Although they are very weak forces, they will cause noble gases to liquefy at very low temperatures.

2A

Remember guys, the ideal gas law can be rewritten to include molar mass as shown below:

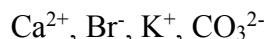
$$\text{MolarMass} = \frac{mRT}{PV} = 68.1 \text{ g/mol}$$

The little m represents mass (in g), T is temperature (268 in K), P is pressure (0.95 in atm), and V is volume (0.85 in L).

3A

This is an annoying way of testing your knowledge of solubility rules. If a reaction occurs it is because two ions have come together that are insoluble and so they drop out of solution as a precipitate (which just means a solid compound forming from a solution).

The best way to look at this is to imagine you have all four ions present from the two solutions:



By looking at solubility rules we can see that CaCO_3 is insoluble and so it will form a precipitate.

4D

This is a partial pressure question, and requires us first to convert the mass of Ne to moles:

$$10\text{g} \times (1 \text{ mol} / 20.18\text{g}) = 0.496 \text{ mole Ne}$$

Then we set up the partial pressure eqn:

$$P_{Ne} = X_{Ne} \times P_{Tot}$$
$$4 = \left(\frac{0.496}{0.496 + n_{Ar}} \right) \times 7$$

and we can solve for moles of Ar

$$n_{Ar} = 0.372 \text{ mol}$$

$$\text{Lastly we convert to mass} = (0.372 \text{ mol Ar}) \times (39.948 \text{ g / mol}) = 14.8 \text{ g}$$

5C

We set up the rate of effusion eqn:

$$\frac{r_{H_2}}{r_{Unk}} = \sqrt{\frac{M_{Unk}}{M_{H_2}}}$$

$$\frac{r_{H_2}}{\frac{1}{5} r_{H_2}} = \sqrt{\frac{M_{Unk}}{2.016}}$$

$$5 = \sqrt{\frac{M_{Unk}}{2.016}}$$

$$M_{Unk} = 50.4 \text{ g / mol}$$

6D

For this problem, we must add up all the number of each type of element present in the diagram, remembering to include all the necessary hydrogens. This gives us the molecular formula of

C₁₀H₁₂. All we do now is find the ratio of atoms to one another and we've got the empirical formula: C₅H₆.

7A

ppm is simply the mass (or volume, or moles) of the solute divided by the entire mass and then multiplied by a million:

$$\frac{0.0095 \text{ g}}{5 \text{ g}} = 1900 \text{ ppm} = 1.9 \times 10^3 \text{ ppm}$$

Remember to make sure that all values are in the same units before calculating.

8A

The question sets up STP, where we know that ONE mole will have volume 22.4 L, regardless of the type of gas.

So all we have to do is figure out which answer has one mole by multiplying the masses given by the molar mass of each molecule.

The best choice is (a) since the molar mass of CO₂ is approximately 44.01 g/mol.

9A

The primary factor affecting speed at any given temperature is mass (or molar mass). The larger the mass the slower the speed.

Remember, for this problem you have to assume that oxygen and nitrogen are in their standard diatomic molecules of O₂ and N₂ and so their molar masses will be double their atomic masses.

10E

Remember, the surest way to be certain if they are structural isomers is if they have the exact same number and type of atoms. Analysis of the options shows that i, ii and iv each has 6 carbons and 12 hydrogens.

11E

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:

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

And:

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

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

$$\begin{aligned}q_{\text{surr}} &= q_{\text{cooling}} + q_{\text{freezing}} \\q_{\text{cooling}} &= mC_p\Delta T = (64\text{g})\left(4.184\frac{\text{J}}{\text{gK}}\right)(0 - 15) = -4.02\text{ kJ} \\q_{\text{freezing}} &= m\Delta H_{\text{freezing}} = (64\text{g})\left(-334\frac{\text{J}}{\text{g}}\right) = -21.4\text{ kJ} \\q_{\text{surr}} &= -4.02\text{ kJ} + -21.4\text{ kJ} = -25.42\text{ kJ}\end{aligned}$$

Going back to the first equation:

$$\begin{aligned}q_{\text{sys}} &= -q_{\text{surr}} \\m\Delta H_{\text{vap}} &= -(-25.42\text{ kJ})\end{aligned}$$

Solve for m:

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

12E

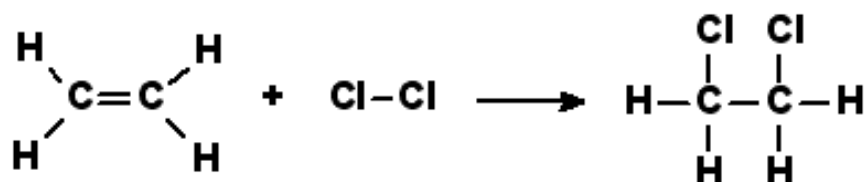
- I) False - The solubility of solids is determined by the molecules and temperature.
- II) False- Solubility of gases increase the temperature decreases.
- III) False- Solubility of gasses decrease as pressure decreases.

13D

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)

14D

Halogens can take part in addition reactions at room temperature by attacking the double or triple bonds of organic compounds. A pi bond becomes uncoupled, and two new sigma bonds form.



15C

Gasses behave most ideally at HIGH temperatures and LOW pressures.

Why?

Just because.

16D

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 H₂O:

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

So:

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

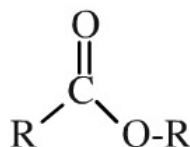
We need to find the mole fraction of I₂ so we will refer to the partial pressure equations:

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

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

17A

COOH is a carboxylic acid functional group and an ester is represented below.



18D

Structural isomers are compounds with the same molecular formula, but different atomic connectivity in their structures. The molecular formula of the molecule shown and the molecule in I and II is C₅H₁₁NO₂, but the molecular formula of the molecule in III is C₅H₉NO₂.

19D

The empirical formula of a molecule is the most reduced whole number of atoms of each element present in a compound. The molecular formula of the compound given is $C_{18}H_{24}O_2$, so the most reduced whole number form is found by dividing each subscript by two.

20A

Noble gases are entirely uncharged, and do not have polar covalent bonds, ionic bonds, or dipole moments. The only intermolecular forces that exist in noble gases are London dispersion forces. Although they are very weak forces, they will cause noble gases to liquefy at very low temperatures.

Solutions - Practice Test 2 - CHEM 110 Exam 3

1D

The factor that determines pressure in this situation is number of moles. There is no way to determine the exact number so assume each molecule or dot is one mole. So basically, the more dots or molecules, the greater the pressure. The answer would be D. i and ii have the same number of dots and molecules combined, so therefore, they have the greater pressure.

2C

Find the number of moles in the initial case, then use moles to find temperature in the final situation. Using the ideal gas law, solve for n:

$$(3\text{L})(1\text{ atm}) = (n\text{ mols})(R)(37+273\text{K}) = \\ n = 0.11793\text{ moles.}$$

Since the moles of the substance has to stay the same, use moles with the new pressure and volume to find the new temperature:

$$(.9\text{ L})(4\text{ atm}) = (.11793\text{ moles})(R)(T) \\ T = 372\text{ K (which is }99\text{ C)}$$

3D

Find the mole fraction of gas that is gas B. Divide the parts of gas B by the total number of parts. Here we would get 4 parts B / 13 parts total. This yields the ratio of pressures in the tank. Now by multiplying the total pressure by the ratio of gas B, we will get the partial pressure of B.
 $(4/13) \times (2.43) = 0.75\text{ atm.}$

4A

Set up a system of equations. Since we know that the atmosphere is only N and He gas, the partial pressures of the two should add up to the total pressure of the atmosphere. We can show this as:

$$P_{\text{total}}X_{\text{He}} + P_{\text{total}}X_{\text{N}} = P_{\text{total}}.$$

We also know that the $X_{\text{N}} = 5X_{\text{He}}$, so we can use substitution to solve and we end up with

$$P_{\text{total}}(X_{\text{He}}) + P_{\text{total}}(5X_{\text{He}}) = 900.$$

We want to solve for $P_{\text{total}}X_{\text{He}}$ because we want the partial pressure of He gas. So solving yields:

$$6P_{\text{total}}(X_{\text{He}}) = 900$$

$$P_{\text{total}}(X_{\text{He}}) = P_{\text{He}} = 150$$

5A

Remember that u_{rms} is given by the equation:

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314)(373)}{0.220}} = 206 \text{ m/s}$$

where we use molar mass, M , in kg/mol.

6A

Using the equation for heat energy and temperature change we can calculate specific heat capacity:

$$q = mc\Delta T,$$

$$c = \frac{q}{m\Delta T} = \frac{2200}{(45)(188)} = 0.26$$

where we remember to convert the energy from kJ to J.

7A

This is really a 3 –step calculation, where we must first calculate the heat absorbed as the ice increases from -5 to 0, then the heat required to melt the ice to water, and finally the heat required to raise the temperature of water from 0 to 3.

Notice that we can use mass when calculating the temperature changes, but for the heat of fusion, we will need convert the mass of water to moles:

$$50\text{g} \times (1 \text{ mol} / 18\text{g}) = 2.78 \text{ mol}$$

Then we can find the 3 separate energies and add them to get the total heat absorbed:

$$q_1 = mc\Delta T = (50)(2.089)(0 - -5) = 522.25 \text{ J}$$

$$q_2 = n\Delta H_{\text{vap}} = (2.78)(6.02) = 16.74 \text{ kJ} = 16,740 \text{ J}$$

$$q_3 = mc\Delta T = (50)(4.184)(3 - 0) = 627.6 \text{ J}$$

$$q_{\text{Total}} = q_1 + q_2 + q_3 = 17,890 \text{ J} = 18 \text{ kJ}$$

8C

Since we want to list the options in increasing order, we start with the lowest vapor pressure, which corresponds to the strongest intermolecular force. In all the options, the molecules are nonpolar, meaning they are all reliant on London forces as their only intermolecular force. When we are tied like this, we look at molecular weight to decide the relative strengths: the heavier the molecule, the stronger its intermolecular force. So I₂ is heaviest, then bromine, then chlorine, and lastly methane. Thus the order is answer C.

9B

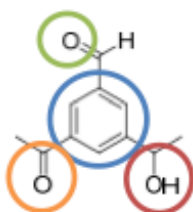
Based on the solubility rules provided on your data sheet, CaSO₄ is insoluble and will not dissolve in water.

10D

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.



11A

In this equation, “a” corrects for attractive forces between gas molecules, and “b” corrects for the frequency of collisions.

12D

If the substances were all at STP, they would all be the same at the same temperature (0C). Obviously, they are not, so they can not be at STP. I is false.

We assume the atmosphere to have a pressure of 1 atm and when the vapor pressure of a substance is equal to the atmospheric pressure, that substance has undergone a complete phase change (evaporation through boiling). So looking at the graph, it takes the most temperature to make line D reach 1atm, so D must have the highest boiling point.

Opposite to that, line A must have the lowest boiling point, given that weak intermolecular forces result in low boiling points, and so III is also correct.

13D

Which of the stages requires energy? Usually, it takes energy to break bonds and it releases energy when bonds are formed. I and II are separating bonds so this requires energy. These too are therefore endothermic because they require energy.

14D

Recall phase diagrams represent the three phases of any substance solid, liquid, and gas. The segment top left represents solid, top right represent liquid, and bottom right represents gas. This is due to pressures and temperatures on the axis. When looking at a substance on a line, both phases are present. Similarly, when on the point in between all the lines, (point A above), all the phases are present. Point E is on the line separating solid and gas, so both phases must be present in the representation. Answer A is just gas (free molecules). Answer B is liquid and gas (point C). Answer C is liquid and dissolved water. Answer D is solid and gas. Answer E is just liquid. The answer is D.

15B

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 r_1 and plug in the values of the molar masses.

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

16C

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.

17A

(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.

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

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

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

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

18B

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.

19B

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.

20A

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.

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

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

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