Solutions - Practice Test - CHEM 110-B Exam 3

1C

This is almost a typical heat absorption question except for the roundabout way we are given ΔH.

Essentially, we have to assume that the amount of heat released will be proportional to the amount of mass. So, since the question deals with burning 2.5g (which is one-third of 7.5g) we can expect it will release one-third the amount of energy, which is 14 kJ, or 14,000 J.

Now we set up the standard calorimetry equation:

$$\Delta H = mc\Delta T = mc(T_f - T_i)$$

$$T_f = \frac{\Delta H}{mc} + T_i = 633\,°C$$

2A

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

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

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).

3D

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.

4D

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

10g x (1 mol / 20.18g) = 0.496 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

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

Lastly we convert to mass = (0.372 mol Ar) x (39.948 g / mol) = 14.8 g

5C

We set up the rate of effusion eqn:

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

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

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

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

6C

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:

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

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

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

7D

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) x (2.43) = 0.75 atm.

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 CO_2 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 O_2 and N_2 and so their molar masses will be double their atomic masses.

10A

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.

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:

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

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:

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

12A

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:

50g x (1 mol / 18g) = 2.78 mol

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

$$q_1 = mc\Delta T = (50)(2.089)(0 - -5) = 522.25 \text{ J}$$
$$q_2 = n\Delta H_{vap} = (2.78)(6.02) = 16.74 \text{ kJ} = 16{,}740 \text{ J}$$
$$q_3 = mc\Delta T = (50)(4.184)(3 - 0) = 627.6 \text{ J}$$
$$q_{Total} = q_1 + q_2 + q_3 = 17{,}890 \text{ J} = 18 \text{ kJ}$$

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)

14C

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 I_2 is heaviest, then bromine, then chlorine, and lastly methane. Thus the order is answer C.

15C

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

16D

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.

17D

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.

18B

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.

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

19A

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 $\Delta K = \Delta\ °C$

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.