CHEM 112 Exam 2 - Practice Test – Solutions

1C

The Second Law of Thermodynamics says that $\Delta S_{Univ} > 0$, which basically means that the total entropy of the universe is increasing, and so statement iii is correct. As for entropy being equal to energy of the universe, that is incorrect…and just something to throw you off.

2E

The strongest conjugate base will come from the weakest acid. So basically the question is asking us which of these is the weakest base. Straight away we can eliminate answers (a) and (b) since they are very strong acids. The trend for these weaker oxo-acids is that acidity decreases as you go down the group. Since iodine is lowest in the group, HIO_2 is the weakest acid.

3C

The only compounds that could form basic solutions are those which have anions that are the conjugate bases of weak acids.

NO_3^- and Cl^- are both the conjugate bases of really strong acids (HNO_3 and HCl) and so they are not basic at all.

CO_3^{2-} and F^-, however, are both the conjugate bases of weak acids (HCO_3^- and HF) so they will form basic solutions and thus we pick choices iii and iv.

4B

The equation we are looking to get K for can be obtained by modifying the first equation that is given. The equation needs to be flipped and multiplied by 4. Thus the new K value can be calculating by making the appropriate changes to the given K_c:

$$K = \frac{1}{(K_c)^4} = 2.3 \times 10^{-16}$$

5B

Look at the K value. We know that it is calculated by the

(concentrations of the products) / (concentrations of reactants). So if the answer is really large, either the concentrations of the products is really large, or the concentration of the reactants is really low. Either way, when K is large, there are more products than reactants.

6D

The corresponding ICE table would look as follows:

	A(g) +	2B(g) ⇌	2C(g)
I	2	1	0
C	-x	-2x	+2x
E	1.7	1-2x	2x

Since we know that 2-x = 1.7, we can solve for x = 0.3.

Thus at equilibrium the pressure of B is 1.4 atm and the pressure of C is 0.6 atm.

7A

For this problem we must combine the reactions given to end up with the target reaction for which we are solving for K.

Notice that the first reaction is the only one to have X_2 except that we need it on the reactants side and we need two moles, not one. So flip the first reaction and multiply it by 2 to get:

$$2\ X_2 \rightleftharpoons 4\ X, \qquad K = 1/9$$

Remember that when dealing with equilibrium constants if we flip the reaction we must take the reciprocal of the K value, and also that if we multiply the reaction by a factor of two, we must raise the K value to the power of 2.

The second reaction has Y_2 exactly where we want it and with the correct number of moles, so we leave the reaction as is:

$$4\ X + Y_2 \leftrightarrow 2\ X_2Y \qquad K = 2.0$$

Finally, if we add the two equations together we'll be left with the target equation. All we need do now is multiply together the two K values: 1/9 x 2.0 = 0.22.

8E

Remember that although K is always products over reactants, it only includes gaseous(g) and aqueous(aq) species.

9E

The main things to look for when entropy is increasing ($\Delta S > 0$) is first a change in phase (like going from solid to liquid or liquid to gas), and then to look for an increase in the number of moles.

Answer choice (c) does show an increase in moles as the reaction goes from reactants to products, but the phase change (going from gas to solid) is a dramatic decrease in entropy and is more relevant to the overall change in entropy than the increase in moles.

The only answer that works is (e) which shows a solid breaking up into two ions, thus increasing the number of moles.

10E

Although there are several factors that determine relative entropy, phase is the most important. Since answer choice (e) is the only option that's in the gas phase, it is the right answer.

However, if you were looking to put the answers in order from least to greatest entropy it would be: (b) < (a) < (c) < (d) < (e).

11C

Expanding the volume of a gas container allows for more room for the gas molecules to migrate and for entry of more gas molecules, which leads to an increase in the randomness of their motion.

12D

The equation relating K_p and K_c shows that if the number of moles of gas reactants equals the number of moles of gas products then $K_P = K_c$. Only answer choice D does this.

13A

By removing hydrogen, you decrease the pressure (or concentration if you like) of one of the products which is supposed to shift the equilibrium to the right. Thus we pick A.

14B

To calculate K you'll first need to get the concentrations of all the necessary reactants and products. In each case the concentration at equilibrium is (3 mol / 5 L) = 0.6 M.

Now we just set up the equation for K:

$$K = \frac{[C]}{[A]^2} = \frac{0.6}{(0.6)^2} = 1.67$$

Notice that B does not play a part in calculating K since it's in the liquid phase.

15A

We know:

$$K_p = \frac{P_B^3}{P_A^2}$$

To solve this formula, we must determine the equilibrium partial pressure values of both A and B. To accomplish this, we must form an ICE table:

	2A	3B
I (initial)	1	0
C (change)	-2x	+3x
E (equilibrium)	1-2x	3x

Build the ice table from the information provided in the problem statement. Notice that the reactant side has negative change values and the coefficient is determined by the balanced reaction. The equilibrium row is just the sum of initial and change cells.

To solve for x, again look at the problem statement. It tells us that at equilibrium, there is 0.45 atm of A. So, 1-2x= 0.45 → solve for x= 0.275. Now we also know at equilibrium, the partial pressure of B is 3x → 3(0.275) = 0.825 atm.

Plug the equilibrium values into the K_p equation.

$$K_p = \frac{(0.825\ atm)^3}{(0.45\ atm)^2} = 2.77$$

16A

Understand that alone in solution, HOCl does not break up, and so it will be present in the net reaction. On the other hand, NaOH does break up even before it reacts with anything else, and so we really have Na^+ and OH^- floating around separately. Since only the OH^- is needed for the reaction with HOCl, the Na^+ is considered a spectator ion and will not appear in the net reaction. Thus option A is the appropriate answer choice.

17D

A change in volume corresponds to a change in total pressure. Recall that when pressure is increased, the reaction will proceed towards the side with a fewer number of moles of gas. If the total pressure is decreases, the reaction will proceed towards the side with more moles of gas.

To solve this problem, balance each reaction and determine which has the same number of moles of gas on both sides. Be sure to only look at those in the gas phase (g).

(a) $2H_2$ (g) + O_2 (g)→ $2H_2O$ (g) 3 to 2

(b) 4HCl (g) + O_2 (g) → $2H_2O$ (g) + 4Cl (s) 5 to 2

(c) 2CO (g) + O_2 (g) → $2CO_2$ (g) 3 to 2

(d) $3H_2$ (g) + At_2 (s) + O_2 (g) → 2HAt (g) + $2H_2O$ (g) 4 to 4

(e) $2C_8H_{18}$ (g) + $25O_2$ (g) → $16CO_2$ (g) + $18H_20$ (g) 27 to 34

18C

For greater entropy, we first look to gas phase options. This leaves A and C. Since they both have the same number of atoms, we can use molecular weight as a quick guide to entropy. The greater weight corresponds to the greater entropy.

19D

The only candidates even worth considering are options D and E since the other choices are all strong acids. Since the trend for oxoacids involving halogens is that the strength increases as you go up the group, we find that HOBr will be the weakest option.

20C

Since all options are gas phase, we must first look at the number of atoms. C and D have the fewest atoms, and C has the lowest molecular weight, and thus is the best choice.

21A (NOT ON TEST)

No matter what temperature the reaction takes place, the product of [OH-] x [H+] will always equal the value of Kw. Therefore, if you take the square root of the given Kw, you can get the concentration of H+. Take the –log to get the pH.

22D

Decrease Pressure and Decrease Temperature

A change in volume corresponds to a change in total pressure. Recall that when pressure is increased, the reaction will proceed towards the side with a fewer number of moles of gas. If the total pressure is decreases, the reaction will proceed towards the side with more moles of gas.

For endothermic reactions ($\Delta H°_{rxn} > 0$) increasing the temperature will make the reaction shift to the right. The reaction needs energy to proceed, so providing more energy will allow it to proceed further.

For exothermic reactions ($\Delta H°_{rxn} < 0$), increasing the temperature shifts the equilibrium to the left. This is due to heat transfer principles that state heat must flow from high to low. Heat released by exothermic reactions must travel to low surrounding temperatures, so raising the temperature will slow the reaction.

This specific reaction is exothermic with 6 moles of gas as reactants and 7 moles of gas as products. To maximize a product we should decrease pressure and decrease temperature.

23D

The Second Law of Thermodynamics states that the entropy of the universe is always INCREASING.

24B

An endothermic reaction that decreases in entropy will never be spontaneous. This is because when we calculate free energy using $\Delta G = \Delta H - T\Delta S$, if we plug in positive values for ΔH and negative values for ΔS (with Temperature always being a positive value), then it's impossible to get anything but a positive number for ΔG, hence it can never be spontaneous.

25A

Reaction will proceed to the right.

We are given the concentration values at a random point in this reaction. We should calculate Q and compare it to K. Don't forget to balance the reaction.

$2NaCl\,(aq) + Pb(NO_3)_2\,(aq) \rightarrow PbCl_2\,(s) + 2NaNO_3\,(aq) \quad K_c = 1.0$

$$Q_C = \frac{[NaNO_3]^2}{[NaCl]^2[Pb(NO_3)_2]} = \frac{[1]^2}{[1.5]^2[0.85]} = 0.522 < K_c$$

Since Q<K the reaction must proceed to the right in order to achieve equilibrium.