

CHEM 112 Final Exam (Old Material)

Practice Test – Solutions

1B

Since we're looking for the overall change in concentration over time, we can use the balanced equation to write the following:

$$-\frac{\Delta[N_2]}{\Delta T} = -\frac{1}{3} \frac{\Delta[F_2]}{\Delta T} = \frac{1}{2} \frac{\Delta[NF_3]}{\Delta T}$$

We can then solve for $\Delta[NF_3]/\Delta T$ to get:

$$\frac{\Delta[NF_3]}{\Delta T} = -\frac{2\Delta[N_2]}{\Delta T} = -\frac{2}{3} \frac{\Delta[F_2]}{\Delta T}$$

Either answer would work, but only one option is given, B.

2D

Since the reaction has only one step, its coefficients can be used as the exponents in the reaction's rate law. Therefore, the rate law is:

$$\text{Rate} = k[A][B]^2.$$

The units of rate are M/s and M for concentration, which leads to $M^{-2}s^{-1}$ for the units of k.

3D

By looking at experiments 1 and 2, we can see that tripling the concentration of $[ClO_2]$ causes the rate to increase by a factor of 9. Thus, the order is 2 with respect to $[ClO_2]$.

4A

To get the rate constant we first need to get the order with respect to hydroxide. By looking at experiments 2 and 3, we can see that tripling the hydroxide concentration causes the rate to increase by a factor of 3. Thus the order is 1 with respect to [OH]. At this point, you can take any experiment's data and plug in all the values that are now known. The only unknown variable is k , and so we can solve for it:

#1:

$$0.0248 = k[0.060]^2[0.030]^1$$

$$k = 229.6$$

5A

A straightforward question. We can relate half-life to rate constant using the equation:

$$t_{1/2} = \frac{0.693}{k}$$

$$k = 0.00385/s$$

6D

For a compound to be an intermediate, it must be produced during one of the first steps of the mechanism, and then completely used up in a later step. Thus intermediates never appear in the general reaction. As a rule, intermediates are first seen as products, and later as reactants. Only D fits this description.

7E

This is a conceptual question dealing with the application of the collision model theory.

8C

For a first order reaction, the rate law would look as follows:

$$M/s = k[M]$$

Thus if we solve for k, we get:

$$k = 1/s$$

These are the correct units for every first order reaction.

9B

The activation of the reverse reaction is the sum of the E_a of the forward reaction and the total change in energy with the opposite sign.

10C

Considering that after one hour the radioisotope would have experienced a little under 3 half-lives, we can expect the percentage remaining to be close to 12.5%. Remember that after one half-life 50% remains, after two half-lives 25% remains, and after three half-lives 12.5% remains. Since one hour doesn't quite make 3 half-lives, we expect there to be a little more left than 12.5%.

11B.

At the normal boiling point, the vapor pressure of propanol (l) equals the atmospheric pressure. Therefore, this is an equilibrium state, so ΔG is equal to zero. Use the formula $\Delta G = \Delta H - T\Delta S$, plug in zero for G and the S value given to solve for enthalpy. Be sure to convert temperature to Kelvin and Joules into kJ.

12D

Since we are dealing with a weak acid alone in solution, we can skip straight to the equation: $K_a = (x^2) / (M_i - x)$. To get pH we'll need to solve for x , the hydronium concentration, and so we are allowed to assume the x in the denominator is negligible and to not include it. Once we get the hydronium concentration we can simply convert to pH by taking its negative log.

13E

The one requirement to be a Lewis base is to have a lone pair of electrons. Since option E has absolutely no lone pairs of electrons when you draw its Lewis structure, it is least likely to act as a base.

14D

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.

15A

This is a strong acid / strong base titration. We first need to determine the number of moles of hydronium and of hydroxide:

$$\text{H}_3\text{O}^+: 0.05\text{M} \times 0.13\text{L} = 0.0065 \text{ moles}$$

$$\text{OH}^-: (0.05\text{M} \times 0.055\text{L}) \times 2 \text{ mol OH}^- / 1 \text{ mol Ba(OH)}_2 = 0.0055 \text{ moles}$$

If we set up an ICE table for this reaction we'll see that all the hydroxide is used up and we are left with 0.0010 moles of H_3O^+ .

Thus the pH can now be determined by calculating the final concentration of hydronium and then converting to pH:

$$[\text{H}_3\text{O}^+] = (0.0010 \text{ moles} / 0.185 \text{ L}) = 0.0054 \text{ M, and pH} = -\log[\text{H}_3\text{O}^+] = 2.27$$

16C

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.

17C

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.

18A

Understand that alone in solution, HClO 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 HClO, the Na^+ is considered a spectator ion and will not appear in the net reaction. Thus option A is the appropriate answer choice.

19C

By recognizing that this solution is essentially a weak acid, HF, and its conjugate base, which is the F^- ion, we can always calculate the pH using the Henderson Hasselbach equation:

$$\text{pH} = -\log(6.8 \times 10^{-4}) + \log\left(\frac{[0.4]}{[0.5]}\right) = 3.07$$

20A

A common ion exists between HCN, a weak acid, and HI, a strong acid that will completely dissociate. Therefore, to calculate the concentration of CN^- from the K_a equation of HCN, an initial concentration of H^+ exists from HI (0.10 M). Therefore, you get an equation that looks like:

$$K_a = \frac{[x(0.10 + x)]}{(0.70 - x)}$$

Plug in the value for K_a given and disregard the $+x$ and $-x$ in the equation because the value of x is so small that part can be ignored to make for an easier calculation. The value for x equals $[\text{CN}^-]$.

21E

Simply use the Henderson-Hasselbalch equation because it is a buffer system of a weak acid and its conjugate base. The concentration of the acid is 0.15 M and 0.23 M for the base.

$$\text{pH} = \text{pK}_a + \log [\text{base}]/[\text{acid}]$$

22A

This question can be answered easily using the Henderson-Hasselbalch equation because it involves a weak acid/base conjugate pair. The K_b was given, so simply find the K_a first by taking $(1.0 \times 10^{-14}/K_b)$ to get $K_a = 2.5 \times 10^{-9}$. Now plug in the known values into:

$$\text{pH} = \text{pK}_a + \log [\text{base}]/[\text{acid}]$$

Solve for [base]. THEN, multiply that number by 2.0 because the question asks for number of MOLES in a 2.0L solution, not the concentration. Tricky, tricky, tricky...

23A

No matter what temperature the reaction takes place, the product of $[\text{OH}^-] \times [\text{H}^+]$ will always equal the value of K_w . Therefore, if you take the square root of the given K_w , you can get the concentration of H^+ . Take the $-\log$ to get the pH.

24C

Most metal oxides are basic; they form metal hydroxides in water. The other statements all follow from the periodic trends outlined in the packet and in notes.

25B

The main objective will be to determine the pH (and then the pOH) which will require the number of moles of acetic acid and NaOH present. That calculation is as follows:

$$(0.05 \text{ L})(0.2 \text{ M}) = 0.01 \text{ moles acetic acid}$$

$$(0.035 \text{ L})(0.1 \text{ M}) = 0.0035 \text{ moles NaOH}$$

Since there are less moles of base, it is the limiting reagent in the reaction with the acid and will run out first. Therefore, $0.01 - 0.0035$ gives 0.0065 moles of acetic acid formed. The 0.0035 moles of OH^- from NaOH led to the production of 0.0035 moles of acetic acid's conjugate base (the acetate ion). The total volume in this solution is 0.085 L, which the number of moles of acetic acid and its conjugate base must each be divided into to get their concentration values. Now, plug those values into the Henderson-Hasselbalch equation to determine pH, subtract that value from 14, and you get your pOH.

26C

The buffer zone exists between $\text{pH} = 4$ and $\text{pH} = 6$, and this curve shows a weak acid being titrated with a strong base.

27C

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.

28D

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

29B

An endothermic reaction that decreases in entropy will never be spontaneous.

30C

Remember we can relate equilibrium reactions with free energy using:

$$\Delta G = \Delta G^0 + RT \ln(Q)$$

By increasing the pressure of gaseous products we cause Q to increase, which in turn causes the entire right side of the equation to increase, and so ΔG increases as well.