

CHEM 112 Exam 3 Practice Test – Solutions

1A

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.

2C

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.

3B

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.

4E

A low concentration of hydronium indicates a basic solution. So we are looking for the most basic solution which will have the highest pH.

5B

The chemical formula for this strong base is $\text{Sr}(\text{OH})_2$. In solution, every molecule will release two hydroxide ions and so the $[\text{OH}^-] = 0.050 \text{ M} \times 2 = 0.10 \text{ M}$.

Lastly, to get pOH:

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.1) = 1$$

6A

Since we are dealing with a weak acid (strong acids don't have k_a values), we can skip straight to the equilibrium formula:

$$k_a = \frac{x^2}{M_i}$$

where x is equal to $[\text{H}_3\text{O}^+]$, and M_i is the molarity of the acid.

Solving for x we get:

$$x = \sqrt{k_a \times M_i} = 0.0027 = [\text{H}_3\text{O}^+]$$

so

$$\text{pH} = -\log(0.0027) = 2.6$$

7B

Just like the previous question we are dealing with a weak acid and so we can write:

$$k_a = \frac{x^2}{M_i}$$

Since we've already got $M_i = 0.4 \text{ M}$ all we need is x , which is equal to $[\text{H}_3\text{O}^+]$.

We use the pH to do this:

$$x = [\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.5}$$

Lastly we plug in our values to get:

$$k_a = \frac{(10^{-4.5})^2}{0.4} = 2.5 \times 10^{-9}$$

8E

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.

9C

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.

10C

To figure out how many moles were added we'll need to get the initial concentration of sodium acetate. Since sodium acetate is a weak base and was added to water we can skip the ICE table and go straight to the equilibrium constant equation for any weak base:

$$K_b = \frac{x^2}{M_i}$$

Remember M_i is equal to the initial molarity of sodium acetate, and since we're using a weak base we need to use K_b , and remember that x is equal to $[\text{OH}^-]$.

We can use the pH to get $[\text{OH}^-]$:

$$\text{pOH} = 14 - \text{pH} = 5.5$$

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-5.5} = 3.2 \times 10^{-6} \text{ M}$$

We can use K_a of acetic acid (the conjugate acid) to get K_b :

$$K_b = \frac{K_w}{K_a} = \frac{10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

Next we solve for M_i :

$$M_i = \frac{x^2}{K_b} = \frac{(3.2 \times 10^{-6})^2}{5.6 \times 10^{-10}} = 0.018 \text{ M}$$

Lastly, all we need do is multiply M_i by volume to get the number of moles of sodium acetate:

$$0.018 \text{ M} \times 0.05 \text{ L} = 9 \times 10^{-4} \text{ moles}$$

11D

This is a strong acid / strong base titration. To begin, we must compare the number of moles of acid with the number of moles of base and see which one is our limiting reactant:

$$\text{Moles HClO}_4 = 0.15 \text{ L} \times 0.1 \text{ M} = 0.015 \text{ mols}$$

$$\text{Moles OH}^- = 0.1 \text{ L} \times 0.5 \text{ M} = 0.05 \text{ mols}$$

Since there are fewer moles of acid they will all be consumed by the base and we will be left with: $0.05 - 0.015 = 0.035$ mols OH^- .

All we need do now is calculate $[\text{OH}^-]$ and then convert to pH:

$$[\text{OH}^-] = (0.035 \text{ mol}) / (0.25 \text{ L}) = 0.14 \text{ M}$$

(note that the volume is the combined volumes of both the acidic and basic solutions)

$$\text{pOH} = -\log(0.14) = 0.85$$

$$\text{pH} = 14 - \text{pOH} = 14 - 0.85 = 13.15$$

12A

The primary requirement of any buffer system is the presence of a weak acid and its conjugate base. Since (a) is the only option that has a weak acid (HF) and its conjugate base (F^-) that is the correct answer.

Remember that RbF in solution becomes Rb^+ and F^- since Rb is a group 1 metal.

13B

This is a standard weak acid / strong base titration.

First we must calculate the number of moles of weak acid and the number of moles of base:

$$\text{Moles HNO}_2 = 0.1 \text{ L} \times 0.5 \text{ M} = 0.05 \text{ mols}$$

$$\text{Moles OH}^- = 0.1 \text{ L} \times 0.05 \text{ M} = 0.005 \text{ mols}$$

Now we should look at the reaction that takes place and set up an ICE table:

I	0.05	0.005	0	-
C	-0.005	-0.005	+0.005	-
E	0.045	0	0.005	-

As the table suggests, since we're dealing with a strong base it reacts completely, hence at equilibrium there is no hydroxide left and all we've got in solution is some of the weak acid we started with, HNO₂, and some of its conjugate base, NO₂⁻.

The Henderson Hasselback Equation is perfect for situations like this since it's a direct link to pH when all you've got in solution is weak acid / conjugate base. So we plug in our values:

$$\text{pH} = -\log(4.5 \times 10^{-4}) + \log(0.005/0.045) = 2.4$$

(note that when using the Henderson Hasselback equation it's not necessary to convert our equilibrium mole values to concentrations)

14B

There are a couple of ways to solve this. One is time consuming, the other is not.

Under most circumstances you would need to calculate both the number of moles of weak acid, ammonium (NH_4^+), and that of the weak base, ammonia (NH_3), so that they could be used in the Henderson Hasselback equation. This is the time-consuming way.

Alternatively, you can hopefully recognize that both the acid and base solutions have the same volumes and concentrations and so their concentrations in the buffer are equal to one another:

$$[\text{NH}_4^+] = [\text{NH}_3]$$

This is referred to as the halfway point, and that means that $\text{pH} = \text{pK}_a$.

So all we really need to do is determine pK_a :

$$K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{1.77 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$\text{pH} = \text{pK}_a = -\log(5.6 \times 10^{-10}) = 9.25$$

15D

This is a weak acid / strong base titration.

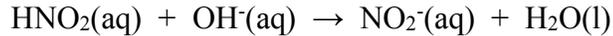
As with similar titrations we must first calculate the number of moles of acid base:

$$\text{Moles HNO}_2 = 0.1 \text{ L} \times 0.25 \text{ M} = 0.025 \text{ mols}$$

$$\text{Moles OH}^- = 0.05 \text{ L} \times 0.5 \text{ M} = 0.025 \text{ mols}$$

Immediately, you should recognize that the initial number of moles of weak acid are equal to the number of moles of hydroxide added, and so we are at the equivalence point.

This means that all of the weak acid and hydroxide will be used up and the only thing we'll have left in solution that can alter the pH is the conjugate base, NO_2^- . Furthermore, we know that we have the exact same number of moles of NO_2^- (0.025 mols) as we had of the hydroxide ions that were all used up during the titration. This is all useful information if you have any interest in saving time during your test. Still, for those who want to see a more detailed analysis, the ICE table below shows what's going on:



I	0.025	0.025	0	-
C	-0.025	-0.025	+0.025	-
E	0	0	0.025	-

As you can see, all we have left in solution is 0.025 moles of NO_2^- , and so to calculate pH we just have to treat this problem as we would any other weak base equilibrium problem by using the equation:

$$K_b = \frac{x^2}{M_i}$$

Remember that since we're dealing with a weak base we'll be needing K_b (not K_a) and also that $x = [\text{OH}^-]$ (not $[\text{H}_3\text{O}^+]$).

We get K_b using the same old equation:

$$K_b = \frac{K_w}{K_a} = \frac{10^{-14}}{4.5 \times 10^{-4}} = 2.2 \times 10^{-11}$$

Next we need the initial molarity, M_i , which is the concentration of our weak base:

$$[\text{NO}_2^-] = (0.025 \text{ mol}) / (0.15 \text{ L}) = 0.17 \text{ M}$$

(note that the volume is the combined volumes of both the acidic and basic solutions)

Now we can solve for x :

$$K_b = \frac{x^2}{M_i},$$

$$x = \sqrt{K_b \times M_i} = \sqrt{(2.2 \times 10^{-11}) \times (0.17)} = 1.9 \times 10^{-6} = [\text{OH}^-]$$

$$pOH = -\log[\text{OH}^-] = -\log(1.9 \times 10^{-6}) = 5.7$$

$$pH = 14 - pOH = 8.3$$

16B

CuCl would be least soluble in the solution that already contains the most Cu^+ and/or Cl^- .

Option (a) has 0.5 M Cu^+

Option (b) has 0.6 M Cl^- (since the ratio of chloride to calcium chloride is 2:1)

Options (c) and (d) contain no common ions and will not reduce solubility at all.

Option (e) has 0.4 M Cl^-

So clearly answer choice (b) has the highest concentration of the common ion and is thus the right answer.

17D

At the equivalence of a weak acid titration statement ii is correct. Statement i is true at the half-way point.

Also, at the equivalence point, the only thing left in solution that could affect pH is the conjugate base of the weak acid, and so since it's a base we can expect the solution to be basic too. Thus $\text{pH} > 7$.

Note that if this were a STRONG acid / strong base titration, then the pH would equal 7 at the equivalence point.

18B

For optimum buffer capacity we first need to find weak acid / conjugate base pairs. Answer choices (a) and (c) both use HCl (a strong acid) and cannot be candidates for the best buffer.

Since (b) and (d) both use the same legitimate weak acid / conjugate base pairs we must then look to see which pairs have the highest minimum concentrations. The minimum concentration for answer (d) is 0.15M, whereas for answer (b) it is 0.25M, so the most reliable of the two buffers is (b) as it has the highest minimum concentration.

19D

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.

20E

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.

21A

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:

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

22A

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.

23C

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:

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

24A

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 Ka equation of HCN, an initial concentration of H⁺ exists from HI (0.10 M). Therefore, you get an equation that looks like:

$$K_a = [x(0.10 + x)]/(0.70 - x)$$

Plug in the value for Ka 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 [CN⁻].

25E

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.

$$\text{pH} = \text{p}K_a + \log [\text{base}]/[\text{acid}]$$

26A

This question can be answered easily using the Henderson-Hasselbalch equation because it involves a weak acid/base conjugate pair. The Kb was given, so simply find the Ka 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:

$$\text{pH} = \text{p}K_a + \log [\text{base}]/[\text{acid}]$$

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...

27B

An atom in its elemental form always has an oxidation state of zero; therefore, Al(s) has an oxidation state of zero. Al(s) loses electrons (is oxidized) when it becomes Al(OH)₄⁻, because it then has an oxidation state of +3. A substance that is oxidized is a reducing agent.

28D

To balance a redox reaction, first separate it into the reduction and oxidation half reactions.



Balance each of these separately. First balance all atoms other than O and H (for example add Na⁺ to the right side of the oxidation to balance the Na atoms). Then, balance the O atoms, by adding the same number of H₂O molecules to the other side. Then, balance the H atoms (including those from the H₂O's just added) by adding H⁺ ions to the other side. Finally, balance the charge on each side by adding electrons to either the reactant side (reduction equation), or the product side (oxidation equation). The final step is to add both equations together, making sure that the number of electrons on the reactant side and product side are equal. For this to be true in this equation, the oxidation equation had to be multiplied by 2 and the reduction equation had to be multiplied by 5. Therefore, the final coefficient in front of Bi³⁺ is five.

29E

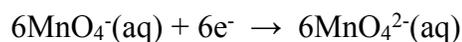
When balancing a redox reaction in basic solution you use the same approach as in acidic solution until the end when everything else has already been balanced. Then, you add OH⁻ ions to each side for every H⁺ present. In this case, there should be 14H⁺ on the reactant side, so we add 14 OH⁻ to both sides. After the H⁺ and OH⁻ on the reactant side neutralize to form H₂O, you will then have 14 H₂O on the reactant side and 7H₂O on the product side, which cancel out to leave 7 H₂O on the reactant side.

30D

First we need to separate the reaction into two half reactions, and then apply material and charge balances to each:

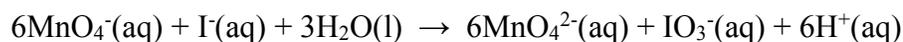


Once this is done we must multiply each reaction by the smallest whole number to get the electrons transferred to be equal. This means that we just multiply the first reaction by 6 and leave the other one alone to get:



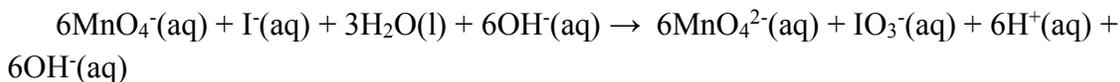
Remember that when balancing we add an H_2O for every oxygen needed and an H^+ for every hydrogen needed.

Now we combine the reactions again by adding together everything on the left side and everything on the right side (make sure to cancel out anything that exists on both sides):

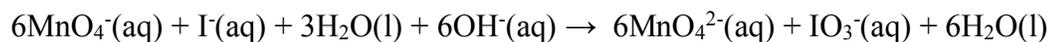


Note that this would represent the balanced reaction in acidic solution.

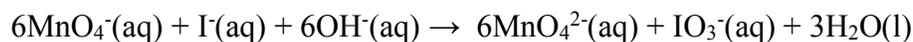
The next step is to take into consideration that we're in basic solution and so we must add an OH^- ion to each side of the reaction for every H^+ that's present to get:



As expected the $6\text{H}^+(\text{aq})$ and $6\text{OH}^-(\text{aq})$ will neutralize to form $6\text{H}_2\text{O}$ and the reaction becomes:



Lastly, we cancel out any extra H_2O that might be on both sides to get:



This is our balanced reaction in basic solution. The sum of the coefficients is 23.

31A

The oxidizing agent is the species that gets reduced. This can only be a reactant so (d) and (e) aren't even options to begin with. The quickest way to answer this problem is to assign oxidation numbers to each element on both the product's side and the reactant's side and see which elements show a change in number.

If this is done correctly you will notice that the Cr in $\text{Cr}_2\text{O}_7^{2-}$ has an oxidation number of +6 on the reactant's side and +3 on the product's side. This means that it since its charge is decreasing it must be gaining electrons and so it is being reduced. Thus, $\text{Cr}_2\text{O}_7^{2-}$ must also be the oxidizing agent.

As a side note, I'd like to point out that the oxidation number for S in $\text{S}_2\text{O}_3^{2-}$ is +2 but on the product's side it is +2.5. This is very unusual, and yet we still see an increase in oxidation number, which means that $\text{S}_2\text{O}_3^{2-}$ is still being oxidized and is thus the reducing agent.

32A

The oxidation state of Ag^+ is 1+ and it is reduced to $\text{Ag}(s)$, which has an oxidation state of zero. When the oxidation state of an ion/molecule is lowered, that ion/molecule is REDUCED.

33D

The question is really asking us to find out the molar solubility of $\text{Mn}(\text{OH})_2$ in a solution that already has hydroxide ions floating around. This is the common ion effect, and means that the initial concentration of hydroxide has to be taken into account when calculating solubility at equilibrium.

Since $\text{pH} = 11.5$, we know $\text{pOH} = 14 - 11.5 = 2.5$, and so $[\text{OH}^-] = 10^{-2.5} = 0.0032 \text{ M}$

To help calculate solubility we'll use the following ICE table:

	$\text{Mn}(\text{OH})_2(\text{s})$	\leftrightarrow	$\text{Mn}^{2+}(\text{aq})$	$+$	$2\text{OH}^-(\text{aq})$
I	-		0		0.0032
C	-		+x		+2x
E	-		x		0.0032 + 2x

We then set up the equation for the solubility product constant, K_{sp} :

$$K_{\text{sp}} = [\text{Mn}^{2+}][\text{OH}^-]^2 = [x][0.0032 + 2x]^2, \text{ where } x \text{ is the solubility}$$

Since K_{sp} is so small we can assume that x is negligible and the equation becomes:

$$K_{\text{sp}} = [x][0.0032]^2$$

$$x = (1.6 \times 10^{-13}) / (0.0032)^2 = 1.6 \times 10^{-8} \text{ M}$$

By examining the ICE table we can see that $[\text{Mn}^{2+}] = x$, and so $[\text{Mn}^{2+}] = 1.6 \times 10^{-8} \text{ M}$