Chemistry 12

Acid/Base II

**Chemistry 12 Acid/Base II**

**Helpful Equations**

1. Strong acid in water: 100% dissociation

 HNO3 + H2O ⇒ H3O+ + NO3-

 .10M .10M .10M pH = -log(.10) = 1.00

 H2SO4 in water:

 First proton is strong: H2SO4 + H2O ⇒ H3O+ + HSO4-

 .10M .10M .10M

 Second proton is weak: HSO4- + H2O ⇔ H3O+ + SO42-

 .10M <.10M <0.10M

2. Weak acid in water: not 100% dissociation

 H3PO4 + H2O ⇔ H3O+ + H2PO4-

 .10M 0.0279M 0.0279M

 pH = -log(0.0279) = 1.56

3. Strong base (hydroxide base) in water: 100% dissociation (water not in reaction; just a dissociation)

 Sr(OH)2 ⇒ Sr2+ + 2OH- pOH = -log(.20) = 0.70

 .10M .10M .20M pH = 13.30

4. Weak base in water: not 100% dissociation

 NH3 + H2O ⇔ NH4+ + OH-

 .10M 1.34 x 10-3M 1.34 x 10-3M

 pOH = -log(1.34 x 10-3) = 2.87 pH = 11.13

5. ANY reaction that involves a STRONG acid or base goes 100% to completion. So a weak acid with a strong base is 100% due to the strong base. A strong acid with a weak base is 100% due to the strong acid.

 Example: H3PO4 + 3NaOH ⇒ 3H2O + Na3PO4

 weak strong

The OH- ions take all three protons off of each H3PO4 molecule, such as in a titration. If H3PO4 was merely in water, only one proton would come off at less than 100% like #2 earlier.

6. Weak acid and weak base:

 NH4+ + SO42- ⇔ NH3 + HSO4- side with weak acid is favoured

I) Weak Acid Equilibrium and Ka

How do strong acids behave in water?

Write an equation for, and give the [H3O+] and pH for a 0.0010M solution of HCl.

How do weak acids behave in water?

Write an equation for a 0.010M HF solution.

It is not as simple to find the pH of a 0.010M solution of a weak acid as you must first have information on the extent of dissociation for the weak acid in question. Ka, the weak acid equilibrium constant, helps to determine this information and provides a means to solve **weak acid** problems. Ka is a Keq for weak acids, so it is simply an equilibrium constant, and therefore all Keq rules apply.

HF(aq) + H2O(l) ⇔ Ka =

Notice the Ka values \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ as you go down the table because the acids are getting progressively \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (creating \_\_\_\_\_\_\_\_\_\_ H3O+ in solution).

In Chemistry 12, we work on quanititative problems that involve weak acids and weak bases.

**Weak acid quanititative problems can be broken into three types.**

**Type 1 Problems: Finding pH of a weak acid solution**

Example: For the reaction:

 H2S(aq) + H2O(l) ⇔ H3O+(aq) + HS-(aq)

Calculate the [H3O+] and pH if the [H2S]i = 0.050M.

Example:

Calculate the pH of a 0.450M solution of H2PO4-.

\*Consider only the first proton dissociating from a polyprotic weak acid, as the dissociation of the second proton is negligible compared to the first.

Sometimes when salts dissolve in water, one of the ions can act as a weak acid in solution.

Example: What is the pH of a 0.100M NH4Cl solution?

**Type 2 Problems: Calculating the Initial Concentration of a Weak Acid**

Example: What [Fe(H2O)63+] would be required to produce a pH of 4.120 ?

**Assignment 1**:

1. Calculate the pH of a 0.50M solution of H3BO3.

2. Calculate the pH of a 0.235M solution of NaH2PO4.

Hebden p.128 #33 & p. 152 #79, 76, 78

**Type 3 Problems: Finding the Ka of an Unknown Weak Acid**

Example: A 0.20M solution of the weak acid, HA, has a pH of 1.32. Calculate the Ka of the weak acid and use this to identify it.

Example: A 2.00M diprotic acid has a pH of 0.50. Calculate the Ka value.

II) Weak Base Equilibrium and Kb

Write a reaction for and find the pH for 0.10M NaOH. Why is it considered a ‘strong’ base?

Write a reaction for 0.10M NH3 solution . Why is it considered a ‘weak’ base? How would you find the pH of a weak base solution?

Weak base problems such as the one previous can be solved using the Kb constant, a Keq for weak bases (similar to Ka for weak acids).

The larger the Kb, the stronger the base (the more H+ it will accept).

Write a Kb expression: NO2-(aq) + H2O(l) ⇔ HNO2(aq) + OH-(aq)

 Kb = but since Kb =

 [OH-] = [HNO2]…

**Relationship of Kw, Ka, and Kb for a Conjugate Acid-Base Pair**

Write the Write the

Ka for HF: Kb for F-:

multiply the Ka of

HF by the Kb of F- :

What results?

Conclusion?

The acid-base table only lists acid Ka values. Using what you learned above, how would you get the Kb for the corresponding conjugate base?

Example:

 Determine Kb for

 the weak base SO42-:

Example:

 Determine Kb for

 HCO3-:

**Practice Questions:** Determine the Kb values for the following:

a) HPO42- b) H2PO4-

**Assignment 2**: Hebden p. 152 #77, 80, 82 & p.130 #35bce, 36

**Weak base quantitative problems can be broken into three types.**

**Type 1 Problems: Finding the pH of a weak base solution**

Example: Calculate the [OH-] and pH for a 0.25M solution of the weak base HS-.

Sometimes when salts dissolve in water, one of the ions can act as a weak base in solution, such as the salt in the next example:

Example: Calculate the pH of a 0.100M solution of K2C2O4.

**Type 2 Problems: Calculating the Initial Concentration of a Weak Base**

Example: A solution of NO2- has a pH of 8.900. Calculate the [NO2-] that would have been required to make this solution.

Example: A solution of ammonia, NH3, has a pH of 10.50. Calculate the [NH3] used to make the solution.

**Type 3 Problems: Finding the Kb of an Unknown Weak Base**

Example: A 0.44M solution of the weak base B- has a pH of 11.12. Calculate the Kb for this base, and the Ka for the conjugate acid, HB at 25°C.

**Assignment 3**: Hebden p.153 #84-89

III) Writing Formula (Molecular), Complete Ionic, and Net Ionic Equations for Acid/Base Reactions

1. **Strong Acid/Strong Base** (Neutralization):

 F: HCl(aq) + NaOH(aq) 🡪 \_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_\_\_

 C: 🡪

 N: 🡪

Since 100% of strong acids and bases dissociate, they should be written as ions in the complete ionic and net ionic equations. If the resulting salt is low solubility and precipitates, it is included in the net ionic equation.

2. **Weak Acid/Strong Base**:

 F: HF(aq) + KOH(aq) 🡪 \_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_\_\_

 C: 🡪

 N: 🡪

Since less than 5% of weak acids and bases dissociate, don’t split them into ions for the complete ionic and net ionic equations, since the majority of weak acid molecules stay intact.

3. **Strong Acid/Weak Base**:

 F: HCl(aq) + NaCN(aq)­ 🡪 \_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_\_\_

 C: 🡪

 N: 🡪

Many weak bases **originate as salts** since weak bases often have a negative charge. The salt will dissociate 100% into ions, and the weak base component will then react 100% due to the strong acid present.

Here is an example when the weak base does not originate as a salt:

 F: HCl(aq) + NH3(aq) 🡪

 C: 🡪

 N: 🡪

Sometimes when an acid and base react, only a salt is produced as the base does not contain OH-, so no water can form.

4. **Weak Acid/Weak Base**

 F: NH3(aq) + HF(aq)

 C:

 N:

**Assignment 4** Write Formula (Molecular), Complete Ionic, and Net Ionic Equations for the following Acid/Base reactions:

1. HClO4 (aq) + KOH (aq) 🡪

2. HBr (aq) + NaCH3COO (aq) 🡪

3. HCOOH (aq) + LiOH (aq) 🡪

4. HI (aq) + NH3 (aq) 🡪

5. Sr(OH)2 (aq) + HNO3 (aq) 🡪

6. NaCN (aq) + NH4Cl (aq) 🡪

IV) Hydrolysis

When any acid (strong or weak) reacts with a strong hydroxide base, the reaction is 100% (due to the strong base):

Examples:

 HBr(aq) + Ca(OH)2(aq) ⇒

CH3COOH + KOH ⇒

In general, the products are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

These reactions are called \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ reactions.

When a strong acid reacts with a base that does not contain hydroxide, it is still a neutralization reaction, however the only product is a salt. The reaction is 100% because a strong acid is reacting.

 Example:

 HCl(aq) + NH3(aq) ⇒

The salts that are produced can be soluble or insoluble (use your solubility table). The insoluble salts will form a solid and precipitate out of solution. The soluble salts will stay in solution as cations and anions and may act as weak acids or bases. For example, the salt produced above was NH4Cl(aq).

Is it soluble (use your solubility table)?

Therefore, how will it actually exist in solution?

Is the cation or anion a weak acid or base?

So what will that ion do in solution?

The reaction you just wrote above is a **hydrolysis** reaction, and it can cause salt solutions to be acidic or basic (acidic in the example above due to H3O+ formation).

So, whenever a salt is dissolved in solution OR whenever a salt is formed due to an acid reacting with a base, the resulting salt solution may be acidic or basic if a **hydrolysis** reaction occurs. If not, the solution will be neutral.

The ions that make up the salts produced from the neutralization reactions may or may not undergo hydrolysis. Here are the guidelines:

1. Ions that will not undergo hydrolysis are…

2. Ions that will not undergo hydrolysis are…

3. Ions that will undergo hydrolysis are…

Write a dissociation equation for each salt, and then predict whether the resulting salt solutions will be acidic, basic, or neutral, and write any hydrolysis equations as support.

1. KCl(aq)

2. NH4NO3(aq)

3. Na2CO3(aq)

Sometimes, an ion is amphiprotic, so will it act more as an acid or a base?

4. NaHSO3(aq)

5. KHCO3(aq)

If an ion is amphiprotic, write an acidic hydrolysis with a Ka value, and a basic hydrolysis with a K­b value. Whichever K value is greater, that reaction will occur to a greater extent.

How do you predict if the solution is acidic or basic when both ions in the salt hydrolyze?

6.NH4NO2(aq)

7. Al(NO2)3(aq)

\*when Al3+, Cr3+, or Fe3+ exist in solution, they will gain six water molecules around them and then act as a weak acid (see table).

Summarize how you can predict whether a salt solution with two ions that hydrolyze will be acidic, basic, or neutral:

**Assignment 5**: Hydrolysis Exercises

1. Hydrolysis Mini-Lab – answer the questions from the lab for #1 of this assignment.

2. Write dissociation equations, and any hydrolysis equation(s) occurring when the following salts are added to water and predict whether the resulting solution will be acidic, basic, or neutral.

a) Na2HPO4 b) Cr(SO4)3

3. NH3 is titrated with HI. When the two react in the titration, what salt is formed? Does the salt undergo hydrolysis? If so, what is the hydrolysis equation and will the resulting pH be above or below 7?

4. In a titration, which of the following combinations would result in an equivalence point with pH greater than 7.0? \*Hint: find the resulting salt from each reaction and see if and how it undergoes hydrolysis

 A. HCl and NaOH

 B. HNO3 and NH3

 C. HBr and NaCH3COO

 D. CH3COOH and NaOH

5. Calculate the pH of a 0.20M KCN solution (Type 1 Kb problem after considering hydrolysis).

V) Indicators

Indicators are used to signal the equivalence point (when \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_) of an acid-base titration using a colour change. An indicator is a solution of a weak organic acid (an acid that contains \_\_\_\_\_\_\_\_\_\_\_), **HIn**, and its conjugate base, **In-**, at equilibrium. The acid form of the indicator, HIn, is a different colour than the conjugate base form, In-. The following is the general equilibrium for any acid-base indicator:

 HIn + H2O ⇔ In- + H3O+

Let’s look at how an indicator equilibrium works in solution using the indicator bromthymol blue:

 HIn + H2O ⇔ In- + H3O+

For bromthymol blue: yellow blue

If [HIn] > [In-], the system favours the \_\_\_\_\_\_\_side and the solution will be a \_\_\_\_\_\_\_\_\_\_\_\_\_ colour. If [In-] > [HIn], the system favours the \_\_\_\_\_\_\_\_\_\_\_ side and the solution will be a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ colour.

What happens to the equilibrium if bromthymol blue is put into an acidic solution, and what is the resulting solution colour?

 HIn + H2O ⇔ In- + H3O+

 yellow blue

What will happen if bromthymol blue is put into a basic solution, and what is the resulting solution colour?

 HIn + H2O ⇔ In- + H3O+

 yellow blue

During a titration, pH is constantly changing as base is being added to acid (or *visa versa*). If an indicator such as bromthymol blue is present, it will eventually undergo a colour change due to the continual change in [H3O+] and resulting shift of the indicator equilibrium.

If there is acid in a flask with some bromthymol blue, what colour will it be?

If base is continually added from the buret, what shift results in the equilibrium?

How does this affect [HIn] and [In-]?

What result will this have on the colour of the solution?

 HIn + H2O ⇔ In- + H3O+

 yellow blue

The point at which the colour is an equal mixture of the [HIn] colour and the [In-] colour is called the **transition point** for the indicator. Another name is the **endpoint**, as this is when a titration would come to an end as the endpoint signals that the equivalence point has been reached.

The endpoint occurs at different pHs for different indicators, as each indicator has its own unique equilibrium. The acid-base indicator table in the data booklet shows different indicators and the pH range of their colour changes. Most indicators change colour over a range of about 2 pH units. For example, bromthymol blue is yellow at pH 6.0 and below and blue at pH 7.6 and above. From 6.0 to 6.8, it’s yellow-green, at 6.8 it’s perfect green, and from 6.8-7.6 it’s blue-green.

It is very important to be able to distinguish between the two terms **equivalence point** and **endpoint**. The equivalence point is the point in the titration where moles of H3O+ = moles of OH-. The endpoint is the point in the titration where the colour of the indicator changes. If the indicator is chosen correctly, it will change the colour of the solution at or very near the equivalence point.

**Practice Questions:**

1. Which of the following indicators is red at pH 13?

 A. Orange IV

 B. Alizarin Yellow

 C. Indigo Carmine

 D. Thymol Blue

2. What colour is a 1 x 10-3 M NaOH solution containing the indicator Neutral Red?

Recall that the general equilibrium equation for an indicator is as follows:

 HIn(aq) + H2O(l) ⇔ In-(aq) + H3O+(aq)

Write the Ka equation for the above:

 Ka =

At the endpoint, what is true about [HIn] and [In-]? Therefore, what will the Ka reduce to?

 Ka =

So, **at the endpoint** (point of colour change), the [H3O+] equals the value of the Ka for the indicator.

It is easy to find the Ka of each indicator (remember, indicators are weak organic acids) using the indicator data table and some simple calculations.

i) Find the pH of the endpoint of the indicator using the table.

ii) Use the endpoint pH to find the [H3O+] at this point (2nd log(-pH))

iii) At the endpoint, the [H3O+] is equal to the Ka

Example:Find the Ka of Orange IV.

**Assignment 6**: Indicator Exercises

1. Which of the following chemical indicators has a Ka = 2.5 x 10-5 ?

 A. methyl orange B. phenolphthalein

 C. thymolphthalein D. bromcresol green

2. Find the Ka of Alizarin Yellow.

3. A weak acid is titrated with a strong base using the indicator phenolphthalein to detect the endpoint. What is the approximate pH at the transition point?

A. 7.0 B. 8.0 C. 9.0 D. 10.0

4. Read Hebden p.161 (bottom) & 162 on Universal Indicators

Do Hebden p.162 #108-112 and p.163 #116-119

5. The indicator thymol blue has two colour changes (see the acid-base indicator table). How is this possible?

VI) Buffers

What is a buffer?

What composes a buffer?

How would you make an HSO3‑/SO32- 2.0M buffer solution?

Would you have a buffer if you simply add 2.0M NaHSO3 to water? Why or why not?

Let’s suppose you’ve made a HSO3‑/SO32- 2.0M buffer:

 HSO3- + H2O ⇔ SO32- + H3O+

 2.0M 2.0M ~0.05M (very small compared

 To [HSO3-] & [SO32-])

**How does a buffer work?**

What determines pH in any solution?

Therefore, if [H3O+] and [OH-] can somehow be kept constant in the buffer solution, the pH will not drastically change.

What would happen if a small amount of HCl were added to the buffer solution?

 HSO3- + H2O ⇔ SO32- + H3O+

 2.0M 2.0M ~0.05M (v. small)

You could keep adding HCl (which immediately becomes H3O+) and as long as there is sufficient SO32- available in the solution, the H3O+ that forms due to HCl addition will react with SO32- to form H2O and HSO3- (a shift left in the equilibrium). Since most of the H3O+ reacts and is no longer present, the pH will not drastically change.

If HCl continues to be added, eventually SO32- will be depleted enough so that H3O+ will no longer have anything to react with. What happens in this situation?

If OH- is added to our buffer solution, a small amount of it will react with the small amount of H3O+ present, but the majority will react with the weak acid HSO3- in the following reaction:

 HSO3- + OH- ⇒

This is just like the original buffer equilibrium shifting to the right since HSO3- is turning into SO32- :

 HSO3- + H2O ⇔ SO32- + H3O+

 2.0M 2.0M v.small

The critical idea here is that almost all of the OH‑ being added reacts with HSO3- to make water and SO32-, thereby ‘getting rid’ of OH-, so the pH of the solution will not rise drastically. It will, however, rise slightly. Show why graphically:

Not quite all of the OH- added will react, thus the pH will rise slightly. If too much OH- is added, all of the HSO3- will eventually be used up and the buffer will collapse.

The key to a functional buffer is the large, equal concentrations of a weak acid and its conjugate base. The weak conjugate base (in our example SO32-) is present in large amounts to react with and deplete H3O+ added to the solution. The weak acid (in our example HSO3-) is present in large amounts to react with and deplete any OH- added to the buffer. Thus, the pH of the solution cannot drastically change as it is solely dependent on [H3O+] and [OH-].

Summary of how a buffer operates:



**Diluting a Buffer**

Does diluting a buffer with water affect its performance?

Take the example used previously:

 HSO3- + H2O ⇔ SO32- + H3O+

 2.0M 2.0M v. small

Adding water to the above equilibrium system will cause each concentration to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. Will it still work as a buffer?

**Acidic and Basic Buffers**

Each weak conjugate acid/base buffer system has a unique pH that it buffers. The pH can be found by using the Ka of the weak acid. For example:

 HSO3- + H2O ⇔ SO32- + H3O+

 2.0M 2.0M very small

Ka = = =

For the HSO3-:

 Ka = [H3O+] = 1.0 x 10-7

Therefore, the pH of HSO3-/SO32- buffer system is - log (1.0 x 10-7)

 = \_\_\_\_\_\_\_\_\_\_\_\_

Thus, the HSO3-/SO32- buffer system is called a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ buffer.

Weak acids that have Ka values greater than 1.0 x 10-7 have buffering pHs less than 7, thus they are called **acidic buffers**. Weak acids that have Ka values less than 1.0 x 10-7 will have buffering pHs greater than 7, and are deemed **basic buffers**.

Example: Find the buffering pH of an acetic acid / acetate buffer.

Example: Find the buffering pH of an HCN / CN- buffer.

**Assignment 7**: Buffer Exercises

Read Hebden p.177-181 (stop after first paragraph on 181)

1. What composes a buffer and how would you make one?

2. Do Hebden p. 181 -182 #131-140

3. Read Hebden p. 182-183: *Buffers in Biological Systems*

 What is the buffering pH in your blood? What two buffering systems contribute to this?

VII) Acid/Base Titration Curves

Titrations are commonly carried out to find the concentration of an acidic or basic solution.

A standard titration curve has an *x* axis that is the *Volume of Base (or Acid)* *Added* from the buret and the *y* axis is the *pH* in the flask.

**Strong Acid/Strong Base Titration Curve**

(strong acid in the flask, strong base in the buret)

Volume of base added to reach equivalence point

Starting pH in flask

Notice the general shape of the titration curve. The pH rises very slowly at the start of the titration, drastically in the middle region, and then very slowly again at the end. Why is this so?

At the start, in order to make the pH change 1 unit (from say 2 to 3) you have to add a large amount of OH- from the buret. pH 2 is an [H3O+] = \_\_\_\_\_\_\_\_\_\_ and pH 3 is an [H3O+] = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. Thus, you must add

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ - \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_OH‑, quite a large amount.

In the middle region, to change the pH from 6 to 7, you must add

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ - \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ OH-, one ten-thousandth of the OH‑ needed to change the pH from 2 to 3!

This is why, from about pH 4 to pH 10, you add very little OH- and the pH changes so quickly. After pH 10, the same effect takes place as early on.

An analogy using $: If you need to pay a $10 000 loan down to $1 000, it costs $9 000. This may take awhile to pay off! But suppose you had to pay a $1 000 loan down to $100, a $100 loan down to $10, a $10 loan down to $1, a $1 loan down to $0.10, and a $0.10 loan down to $0.01 (simulating the middle of the pH curve). This, in total, costs $999.99, a fraction of the first loan!

Here is an example of a strong acid/strong base titration reaction:

 HCl + NaOH ⇒

What products result?

Water is, of course, neutral. Is the resulting salt neutral? Why or why not?

Therefore, what is the pH at the **equivalence point** of a strong acid/strong base titration?

Strong acid/strong base titrations result in a salt that does not hydrolyze, therefore the **equivalence point** is always 7.

An ideal indicator for a titration is one in which the colour change encompasses the equivalence point. List the ideal indicators for a strong acid/strong base titration.

However, in our lab, we used phenolphthalein to indicate the equivalence point of a strong acid/strong base titration, even though the endpoint of phenolphthalein is 9.1. Why is this okay?

**Weak Acid/Strong Base Titration** (strong base is in the buret)

Though two parts of this curve are different than the strong acid/strong base curve, the vertical rise is still present.

What is different compared to a strong/strong titration?

1.

2.

Why, for a weak acid/strong base titration, is the equivalence point between pH 8 & 9?

Which indicators would be **ideal** for a weak acid/strong base titration?

**Weak Base/Strong Acid Titration** (strong acid is in the buret)

Why does this curve start at a high pH and end at a low pH?

Characteristics include an initial dip in pH and an equivalence point pH of

5-6. This is because the salt produced will hydrolyze acidically:

 ex. HCl + NH3 ⇒ NH4Cl

 NH4Cl ⇒ NH4+ + Cl-

 NH4+ + H2O ⇔ NH3 + H3O+

 The extra H3O+ produced due to the

 hydrolysis of the salt lowers the equivalence

 point to pH 5-6.

List the ideal indicators for a weak base/strong acid titration:

**Assignment 8**: Titration Curve Exercises

1. Hebden p.176 #125

2. A student titrated a 25.00mL sample of 0.20M HX acid with 0.20M NaOH. The following data was collected.

|  |  |
| --- | --- |
| Volume of NaOH added (mL) | pH |
| 0.00 | 2.72 |
| 10.00 | 4.57 |
| 24.90 | 7.14 |
| 24.99 | 8.14 |
| 25.00 | 8.88 |
| 25.01 | 9.60 |
| 26.00 | 11.59 |
| 35.00 | 12.52 |

a) What volume of NaOH must be added to reach the endpoint?

b) Is HX weak or strong? How do you know?

**c)** Select an indicator that would be ideal for this titration and give the colour at the equivalence point.

3) Draw a typical curve for a weak base/strong acid titration (strong acid in the buret). Show all characteristics and estimate the endpoint pH.

IX) Acidic and Basic Anhydrides

How are acids and bases produced in nature?

Both are produced from **oxides**. What are oxides?

Acids are produced from **non-metal oxides** and bases are produced from **metal oxides**.

When **non-metal oxides** react with water in a **synthesis** reaction, an acid is formed. Thus, **non-metal oxides** are called **acidic anhydrides** (\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_). These reactions often occur in the atmosphere as most non-metal oxides are gases.

Examples:

 SO2(g) + H2O ⇒

 SO3(g) + H2O ⇒

 CO2(g) + H2O ⇒

When Group 1 and 2 **metal oxides** react with water in a synthesis reaction, a base is formed. **Metal oxides** are called **basic anhydrides**. Na2O(s) + H2O ⇒

 MgO(s) + H2O ⇒

 CaO(s) + H2O ⇒

X) Acid Rain

Fuels that contain sulfur are combusted in an industrial setting to form sulfur dioxide gas ( ). Some of that sulfur dioxide then reacts with oxygen in the air to produce sulfur trioxide gas ( ). What do we call these compounds?

What will they do when they go up into the atmosphere?

Thus, what results?

When fuel combusts in a car engine, it is so hot that N2 from the air reacts with O2 to form nitrogen monoxide gas ( ). Though the catalytic converter reverses this reaction, some nitrogen monoxide escapes through the exhaust and reacts with O2 in the air to make nitrogen dioxide (\_\_\_\_\_\_\_).

What happens next?

Acid rain is defined as rainwater that has a pH less than **5.6**.

It is important to note that even "normal" rainwater is acidic (between pH 5.6 and 7) due to atmospheric CO2 (an acidic anhydride) dissolving in water to produce carbonic acid.

 CO2(g) + H2O(l) ⇒

**Assignment 9**:

1) There are many environmental problems associated with acid rain. Read Hebden pages 187 & 188.

2) Do Hebden p. 185 #144, 145 & p.188 #147

XI) Acid/Base Testing

Suppose you had a 1.0M solution of a strong acid and weak acid but did not know which was which. Describe any testing you could do to identify each solution.