Name: _____ Block: _____

IV. Acids & Bases (part 3)

IV.14-15 Calculations involving K_a and K_b (*Used for the WEAK A & B*)

You will be able to:

- Given the Ka, Kb, and initial concentration, calculate any of the following: [H_3O^+], [OH-], pH, pOH
- Calculate the value of Ka or Kb given the pH and initial concentration
- Calculate the initial concentration of an acid or base, given the appropriate Ka, Kb, pH, or pOH values

---

Remember: *WEAK acids/bases do not ionize completely.*

- The _____, the _____ is produced.

Therefore, a lower _____ means a _____ acid.

There are 3 TYPES of calculations involving Ka and Kb for weak acids and bases.
The following examples are interchangeable for **ACIDS** and **BASES**.

Calculations involving weak bases are similar to calculations involving weak acids, with 2 changes:

- _____
- _____

Q TYPE 1: Given [WA] and K_a, **find [H_3O^+] (or pH)**

Example 22: What is the pH of a 0.500 M solution of benzoic acid (C_6H_5COOH)?

Step 1: Look up the Ka on the B-L table	
Step 2: Write out ionization equilibrium with an ICE table.	
Step 3: Write Ka expression & substitute values.	
Step 4: State assumption. **Assumption can ONLY be made if percent dissociation is less than 5%.** * Show calc for percent dissociation.	

Step 5: Assumption reduces equation. Solve for x ($[H_3O^+]$).	
Step 6: Convert to pH. (Ka limits to 2 SD's.)	

Q TYPE 2: Given [WA]/[WB] and $[H_3O^+]$/[OH-] (or pH/pOH), **find K_a or K_b**

Example 23: At a certain temp, a 0.20 M solution of K_2SO_3 has a pH of 10.25. Calculate the Kb of SO_3^{2-} at this temp.

Step 1: Write out dissociation equation of salt. Identify the weak base.	
Step 2: Calculate [OH-] from pH (pH → pOH → $[OH^-]$)	
Step 3: Write hydrolysis equation and an ICE table. (It is called *hydrolysis* this time because SO_3^{2-} is an ion.)	
Step 4: Write the Kb expression and substitute the values from the [E]'s in our ICE table	
Step 5: Solve for Kb to correct SD's	

Q TYPE 3: Given $[H_3O^+]$ (or pH) and K_a, **find [WA]**

Example 24: Find the concentration of HCOOH needed to form a solution with pH = 2.69.

Step 1: Convert pH to $[H_3O^+]$ **This is the $[H_3O^+]$ at equilibrium.**	
Step 2: Write out ionization equilibrium with an ICE table. *Calc change in concentrations using molar ratios.*	

Step 3: Write Ka expression & substitute values. Find Ka for HCOOH on the acid table.	
Step 4: Solve for [WA] with correct SD's	

In written response questions, you will have to show your exact calculations! You may state assumption if you can prove that the base/acid is less than 5% ionized.

SHORTCUT FOR MULTIPLE CHOICE ONLY:

Example 25: The pH of 2.0 M acetic acid is...

Step 1: Use MC shortcut option to calc [H_3O^+] (see Ex. 16 in ABpt2)	
Step 2: Look up Ka value in table. Solve for [H_3O^+]	
Step 3: Calculate pH. Select best answer	

Do Hebden set 29: Ka calcs -p. 152 #77-80, 83
Kb calcs – p. 153 #85-87, 91

Date: _____

IV.13 Hydrolysis

You will be able to:

- Write a dissociation equation for a salt in water
- Write net ionic equations representing the hydrolysis of ions in solution
- Predict whether a salt solution would be acidic, basic, or neutral (compare Ka and Kb values, when necessary)
- Determine whether an amphiprotic ion will act as an acid or base in solution (compare Ka and Kb values)
- Calculate the pH of a salt solution from relevant data, assuming that the predominant hydrolysis reaction is the only reaction determining the pH

HYDROLYSIS is _____

_____

Only the reactions between ions and water will be considered in this section. Yes, there may be reactions between the ions, but we are not looking at that for the moment.

Net Ionic Equation for hydrolysis: _____

SPECTATOR IONS: *Which ions DON'T react with water?*

Spectator **CATIONS** (look on periodic table)

Group 1 (Alkali Metal ions) eg. Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+
Group 2 (Alkaline Earth ions) eg. Be^+, Mg^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}

Spectator **ANIONS** (look on acid table)

- Conjugate bases of strong acids.
- ***Top 5 ions on the right*** side of table.
- ClO_4^- I^- Br^- Cl^- NO_3^-
  *(HSO_4^- is not a spectator – it is **amphiprotic** – will be dealt with later)*

Example 26: Write the net-ionic equation for the hydrolysis taking place in aqueous magnesium sulphate.

Step 1: Write the dissociation equation for the salt.	
Step 2: Determine any spectators.	
Step 3: Write acid-base reaction.	

WILL A SALT ACT AS AN ACID OR BASE IN WATER?

Process – if given **SALT** (DISSOCIATE → ELIMINATE → EVALUATE)

1. Write **dissociation** equation
2. Eliminate **spectators**
3. Remaining ions → **left** side of table – undergo **acid** hydrolysis is –produce H_3O^+
   → **right** side of table – undergo **base** hydrolysis – produce OH
   → **amphiprotic –** determine $\mathbf{K_a}$ and $\mathbf{K_b}$ to find ***dominant*** **hydrolysis.** (greater value = dominant hydrolysis)

If both ions in the salt are SPECTATORS, the solution will be _____.

- **Ex: KBr, NaCl, $Ca(NO_3)_2$, etc.**
- Dissociation equation: _____
- No ions will react with water, therefore _____.

Example 27: Is the salt $Fe(H_2O)_6I_3$ acidic, basic or neutral in aqueous solution?

Step 1: Write the dissociation equation for the salt.	
Step 2: Eliminate spectators.	

Step 3: Compare Ka and Kb values. The greater value will determine whether salt acts as acidic, basic, or neutral.	

Hydrolysis when there is an *AMPHIPROTIC* ion :

Example 28: Is the salt $LiHCO_3$ acidic, basic or neutral in aqueous solution?

Step 1: Write the dissociation equation for the salt.	
Step 2: Eliminate spectators.	
Step 3: For the AMPHIPROTIC ion, determine whether ion preferentially acts as acid or base. Compare Ka and Kb values.	

Hydrolysis When *BOTH* Cation and Anion hydrolyze:

Example 29: Determine whether the salt NH_4CN (ammonium cyanide) is acidic, basic or neutral.

Step 1: Write the dissociation equation for the salt	
Step 2: Determine which ion is the acid and which is the base. Find the Ka and Kb values for each ion.	
Step 3: Compare Ka and Kb values.	

SUMMARY:

IfThen the salt is:

Ka (cation) > Kb (anion) **Acidic**

Kb (anion) > Ka (cation) **Basic**

Ka (cation) = Kb (anion) **Neutral**

Do Hebden set 30: Read examples on pp145-147.

Do p. 148 #69acegi, 70acegi, 73

PUTTING IT ALL TOGETHER (*Test yourself!)*

Calculate the pH of a 0.24 M solution of the salt aluminum nitrate. Show all your steps. State any assumptions used.

Date: _____

IV.17 Indicators

You will be able to:

- Describe an indicator as a mixture of a weak acid and its conjugate base, each with distinguishing colours
- Describe the term transition point of an indicator, including the conditions that exist in an equilibrium system
- Describe the shift in equilibrium and resulting colour changes as an acid or base is added to an indicator
- Predict the approximate pH at the transition point using the Ka value of an indicator
- Predict the approximate Ka value for an indicator given the approximate pH range of the colour change

We have used indicators in many labs to identify acidic or basic solutions, or even to signify the equivalence point of a titration.
So, what is an indicator anyway?

DEFINE: An **INDICATOR** is _____
_____

- HIn is the _____
- In^- is the _____

Ex: An indicator HInd has a yellow acid form (**HIn**) and a red base form (In^-).

_____

ACCORDING TO LE CHATELIER'S PRINCIPLE:

- If **excess H_3O^+** is added, equilibrium shifts _____; Favours _____
  - **[HIn] > [In^-]**
  - Colour will be _____

Therefore, _____

- If **excess OH^-** is added, equilibrium shifts _____; Favours _____
  - **[HIn] < [In^-]**
  - Colour will be _____

Therefore, _____

- If acid/base solution has **[HIn] = [In^-]**
  - Colour will be _____

Therefore, _____

Using the A-B Indicator chart in Data Book:

ACID-BASE INDICATORS

Indicator	pH Range in Which Colour Change Occurs	Colour Change as pH Increases
Methyl violet	0.0 – 1.6	yellow to blue
Thymol blue	1.2 – 2.8	red to yellow
Orange IV	1.4 – 2.8	red to yellow
Methyl orange	3.2 – 4.4	red to yellow
Bromcresol green	3.8 – 5.4	yellow to blue

Example 30: When a drop of 0.1M HCl is added to the indicator bromcresol green, the colour is yellow. When a drop of 0.10M NaOH is added to the indicator, the colour is blue.

- What colour is the acid form of bromcresol green (HIn)? _____
- What colour is the base form of bromcresol green (In^-)? _____
- What would the colour be if [HIn] = [In^-] for bromcresol green? _____

DEFINE: The **TRANSITION POINT** is _____
_____

AT TRANSITION POINT (or END POINT):

[HIn] = [In^-] Colour is 50/50 mix of **acid/base colours.** $K_{a\ (indicator)}$ = **[ ]** *Reasoning:*

K_a = [ ][ ] = [ ]
[ ]***Reasoning:*****pKa = pH**
pKa = _____
So, _____
_____
Then, _____

Finding the transition point and K_a of an Indicator

Example 31: Find the K_a of Alizarin Yellow

Step 1: Look on the Indicator table. Find the midpoint of the pH range.	
Step 2: Remember at transition point, pKa = pH. Solve for K_a.	

Example 32: What is the colour of indigo carmine indicator in 0.01 M $Ca(OH)_2$?

Step 1: Look on the Indicator table. Find the midpoint of the pH range.	
Step 2: Determine [OH-] in solution, convert to [H_3O^+] to calculate pH	
Step 3: Locate pH of solution on pH range of indicator to determine colour	

Example 33: Indicator X ($K_a = 1.7 \times 10^{-4}$) is orange in acid and green in base form. What colour is it in 0.0001 M HCl?

Step 1: Calculate transition point pH from Ka	

Step 2: Determine $[H_3O^+]$ in solution, convert to to calculate pH	
Step 3: Locate pH of solution on pH range of indicator to determine colour	

Mixtures of Indicators

UNIVERSAL indicator is a mixture of _____, _____, and _____.

- It gives a spectrum of colours (ROY G BIV) that represent the range of pH values.

Do Hebden set 30 p. 162 #109, 112, 116, 120

"INDICATOR PRACTICE" Worksheet

Date: _____

IV. 16 & IV.18: Acid-Base Titrations

You will be able to:

- Design, perform, and analyze a titration experiment of: strong acid + strong base, strong base + weak acid, and strong acid + weak base.
- Explain the difference between the equivalence point of a SA-SB titration and the equivalence point of a WA/WB-SB/SA titration
- Interpret titration curves plotted from experimental data
- Select indicators whose transition point coincides with the equivalence point of the titration reaction
- Calculate the concentration of an acid or base using titration data or similar data (ex. grams or moles)
- Calculate the volume of an acid or base of known molarity needed to completely react with a given amount of acid or base
- Calculate the pH of a solution formed when a strong acid is mixed with a strong base

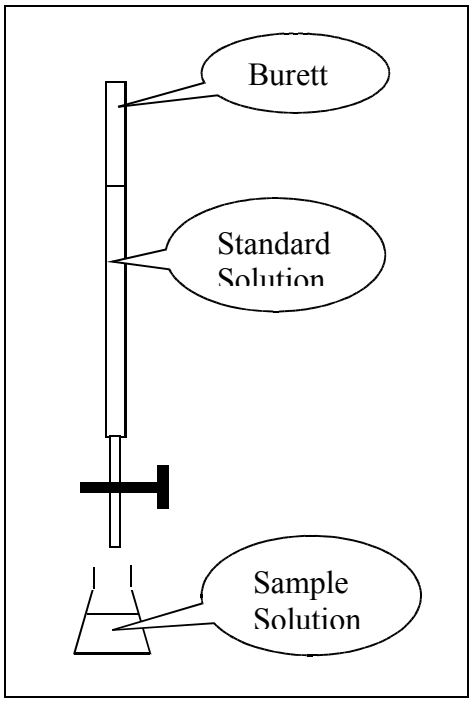
Review of titrations

TITRATION: _____
_____
_____

- **STANDARD** solution (**TITRANT**) = _____
- **SAMPLE** solution = _____ concentration.
- **EQUIVALENCE POINT** (or "stoichiometric point") is _____
  _____

- **INDICATOR** signifies the _____

The <u>best</u> indicator will have the pH at equivalence point withing their transition range.



<u>Important info in titration problems:</u>

- concentration of acid — concentration of base — base/acid mole ratio
- volume of acid — volume of base

I. Titration of STRONG ACID and STRONG BASE

Example 34: We have 150 mL of NaOH at an unknown concentration. 75 mL of 0.300 M HCl must be added to reach the equivalence point. What is the [NaOH]?

Step 1: Write out the reaction equation.	
Step 2: Calculate moles of standard solution used to reach equivalence point (mol = CV)	

Step 3: Use molar ratio to convert to moles of sample solution	
Step 4: Use volume of sample solution to find [NaOH] (C = mol/V)	

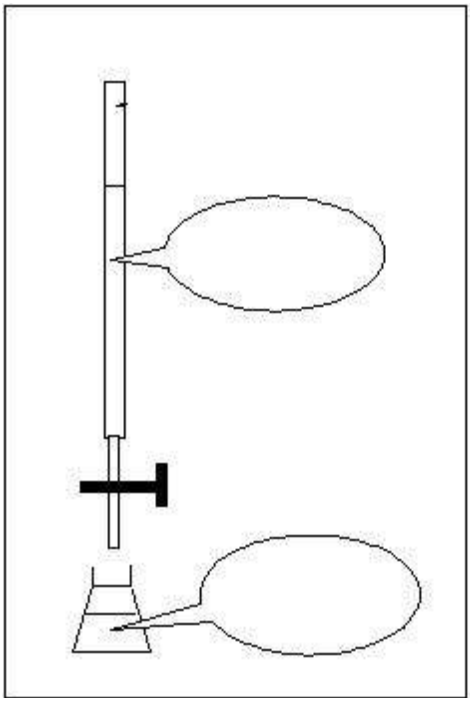
Example 35: 300 mL of unknown [H_2SO_4] is titrated with 600 mL of 0.400 M KOH. What is the [H_2SO_4]?

Step 1: Write out the reaction equation.	
Step 2: Calculate moles of standard solution used to reach equivalence point (mol = CV)	
Step 3: Use molar ratio to convert to moles of sample solution	
Step 4: Use volume of sample solution to find [H_2SO_4] (C = mol/V)	

TITRATION CURVES

Titration Story:

A base of known concentration (ex: 0.01 M NaOH) is slowly added to a measured volume of an acid of known concentration (25.0 mL of 0.01M HCl). Meanwhile, the pH of the mixture is monitored by a pH meter. The results can be plotted in a graph of *pH vs. volume of base* added.

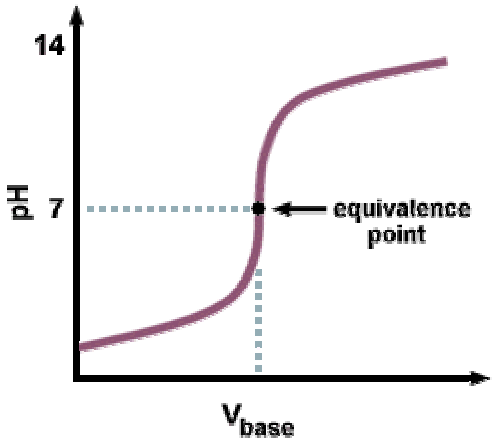


The curve on the graph that results from this is called a **TITRATION CURVE**.

We can calculate the pH of the mixture in the beaker throughout the titration. *There are 4 stages:*

1. Acid before any base is added
2. Based added but acid in excess
3. Equivalence point
4. Base in excess

STRONG ACID-STRONG BASE TITRATION CURVE



Remember: **SA** + **SB** → SALT + H_2O

Therefore: The *SALT* formed from a **SA-SB** titration is ALWAYS _____

Since there is no SA, no SB and just H_2O and a NEUTRAL salt, the pH of the solution formed will be **7.00**

Conclusion: _____

A good **INDICATOR** for this titration would be _____.

Do Hebden set 31 p. 158 #94-97, p. 167 #24

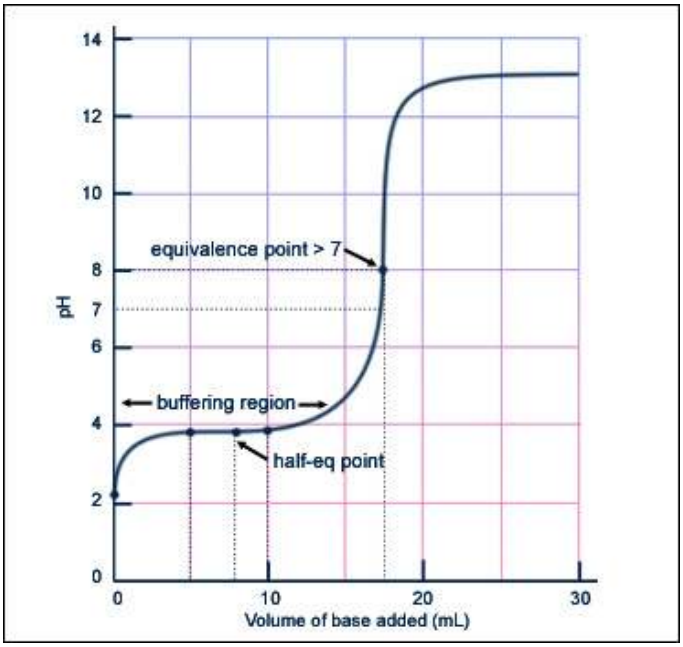
Read p. 155-157 examples "Partial Neutralization", "% Purity" and "Molar Mass", then try p. 158 #98-107

II. Titration of STRONG BASE and WEAK ACID

Example 36: 150 mL of unknown [CH_3COOH] is titrated with 220 mL of 0.250 M NaOH to reach the equivalence point. What is the [CH_3COOH]? **Calc the same as SB-SA**

Step 1: Write out the reaction equation.	
Step 2: Calculate moles of standard solution used to reach equivalence point (mol = CV)	
Step 3: Use molar ratio to convert to moles of sample solution	
Step 4: Use volume of sample solution to find [CH_3CHOOH] (C = mol/V)	

STRONG BASE-WEAK ACID Titration Curve



Conclusion: _____

_____

A good **INDICATOR** would be _____

_____

Calculating the K_a and concentration of the weak acid: *(summary of p. 169-170)*

$pH_{1/2}$ = _____ = pKa

*So, the **K_a of the weak acid** can be calculated from the $pH_{1/2}$!*

K_a = antilog (-pKa)

To calculate the concentration of the weak acid:

$$[WA]_{EQ} = \frac{[H_3O^+]^2}{K_a}$$ <-- *from initial pH*

$$[WA]_{INT} = [WA]_{EQ} + X$$

III. Titration of a STRONG ACID and WEAK BASE

***NOTE*:** Graph gives pH values, but need pOH for all calculations! So, **FIRST <u>convert pH to pOH</u>**.

Conclusion: _____

_____

A good **INDICATOR** would be _____

_____

Summary of pH at Equivalence (Stoichiometric) Points for the three types of Titrations:

Reactants	*Salt Formed is…*	*pH at Equivalence Point*
Strong Acid – Strong Base	Neutral (conjugate base of SA)	= 7.00
Weak Acid – Strong Base	Basic (conjugate base of WA)	> 7.00
Strong Acid – Weak Base	Acidic (conjugate acid of WB)	< 7

Calculating the K_a or K_b and concentrations from a titration:

Example 37: The titration of a solution of benzoic acid (C_6H_5COOH) requires 28.4 mL of 0.125M NaOH. The initial pH of the benzoic acid solution is 2.628 and the pH is 4.191 after 14.2mL of the NaOH is added.

a) What is the Ka for benzoic acid?

Step 1: Determine what type of titration you are doing and outline the information you have.	
Step 2: Calculate Ka from the pKa = $pH_{1/2}$ relationship.	

b) What is the original concentration of the benzoic acid solution?

Step 3: Calculated the initial $[H_3O^+]$ from $pH_{initial}$	
Step 4: Set up Ka expression. Solve for $[WA]_{EQ}$ **Note: This is the conc at equivalence point, NOT initially!*	
Step 3: Calculate the ORIGINAL concention by adding equilibrium conc + dissociated ion	

Read examples on p. 171 and 175

Hebden set 32: p. 176 #125-127

\

Date: _____

IV.19-20 Buffers

You will be able to:

- Describe the tendency of buffer solutions to resist changes in pH
- Describe the composition of an acidic buffer and a basic buffer
- Describe qualitatively how the buffer equilibrium shifts as small quantities of acid or base are added to the buffer; the stress being the change in the concentration of the stronger acid (H_3O^+) or base (OH-)
- Describe in detail a common biological buffer system
- Outline a procedure to prepare a buffer solution
- Identify the limitations in buffer systems

DEFINE: A **BUFFER** is a solution containing _____ _____.

Purpose of a buffer:

- It _____ changes in pH when _____.
- Or we could say it _____ when acid or base is added.

Ex: $WA + H_2O \rightleftharpoons H_3O^+ + WCB^-$ *(WCB is "weak conjugate base")*

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

K_a = _____ = $[H_3O^+]$ = _____

Therefore, **pH = pKa**

When _____ of a weak acid and its conjugate base are added to water, the _____.

NOTE: A BUFFER requires **substantial** amounts of _____.

There are two kinds of **Buffer Solutions**:

- ***ACIDIC BUFFERS***: _____

Acidic Buffers are useful as buffers in the acidic range (solutions in which pH is 7 or lower)

Ex: "Mix 1.0 mol of CH_3COOH and 1.0 mol of $NaCH_3COO$ and dilute to 1.0 L solution."

K_a = _____

Buffer pH = pK_a = _____ = _____

- ***BASIC BUFFERS***: _____

Basic Buffers are useful as buffers in the basic range (solutions in which pH is 7 or higher)

Ex: "Mix 1.0 mol of NH_3 and 1.0 mol of NH_4NO_3 and dilute to 1.0 L solution."

K_a = _____

Buffer pH = pK_a = _____ = _____

Preparing a Buffer Solution:

Three concepts to consider:

- _____
- _____
- _____

Ex: Say you had some 1.0 M acetic acid (**CH_3COOH**) solution:

Original equilibrium: $CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

Now, lets add some sodium acetate (**$NaCH_3COO$**) to the equilibrium so that [CH_3COO^-] is 1.0 M.

- When we do this the _____. However, by *LeChatelier's Principle*, the equilibrium will shift to the _____, causing _____ and _____.

Shifted equilibrium: $CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

Since the acid and the base are both *WEAK*, they don't neutralize each other like a mixture of a SA and SB would. They co-exist in this equilibrium unless disturbed! A **BUFFER SOLUTION** is prepared!

Example 38: How would you prepare a solution in which the pH is buffered close to 7.2?

Step 1: pH = pKa Calculate Ka of buffer	
Step 2: Find acid that has a similar Ka value (from BL table)	
Step 3: Prepare buffer by mixing equal amounts of WA and soluble salt of its conjugate base	

Explaining Buffer Equilibrium Shifts

Work through the example by filling in the blanks…..

Ex: A buffer solution is prepared using 1M NH_3 and 1M NH_4Cl *(Basic Buffer)*

a) Write the **equilibrium equation** describing this buffer.

_____

b) When a small amount of HCl (SA) is added, the [OH^-] quickly _____creases (the pH goes _____)

c) As a result, the equilibrium shifts to the _____, and the [OH^-] gradually _____creases. (the pH goes back _____)

d) So, as a result of adding HCl, there was a small **net** _____crease in the [OH^-] (a small **net** ___ crease

in pH)

Limitations of buffers:

- If there is _____ of conjugate base present, a maximum of _____ of H_3O^+ can be neutralized.
- If there is _____ of conjugate acid present, a maximum of _____ of OH^- can be neutralized.

Biological buffers:

For Hemoglobin to work properly, the pH of the blood needs to stay very close to 7.35

Equilibrium: $HHb + O_2 \leftrightarrows H_3O^+ + HbO_2^-$
hemoglobin *oxyhemoglobin*

- If pH < 7.20 (**"ACIDOSIS"**), _____.
- If pH > 7.20 (**"ALKALOSIS"**), _____.

TWO BUFFER SYSTEMS:

a) CO_2/HCO_3^-

_____

_____

"Hyperventilating" will lower $[CO_2]$ in the blood, and _____

b) $H_2PO_4^-/HPO_4^{2-}$

_____

Do Hebden set 33: p. 181 #132-133, 136-138, 140; p. 183

Date: _____

IV.21 Applied Acid-Base Chemistry

You will be able to:

- Write equations representing the formation of acidic solutions or basic solutions from non-metal and metal oxides
- Describe the pH conditions required for rain to be called acid rain (pH 5.0 and lower)
- Relate the pH of normal rain water to the presence of dissolved CO_2 (approximately pH 5.6)
- Describe sources of NO_x (automobile engines) and SO_x (fuels containing sulfur and smelters of sulfide ores)
- Discuss general environmental problems associated with acid rain

---

I. OXIDES

IN GENERAL:

- **METAL OXIDES** (IONIC) form _____.

Dissociation in water:

Ex: $Na_2O(s)$ --> _____ or $CaO(s)$ --> _____

Hydrolysis of O^{2-}: _____

Balanced equation of Na_2O and water: _____

- **NONMETAL OXIDES** (COVALENT) form _____.
  (*Exception:* Highly charges small ions such as $Al^{3+}_{(aq)}$, $Cr^{3+}_{(aq)}$, and $Fe^{3+}_{(aq)}$ form hydrated ions)

Ex: _____
_____
_____

II. ACID RAIN

Normal rain has a pH of _____.

- Naturally acidic due to _____ from non-human activity

_____

ACID RAIN has a pH of _____.

- It is a combination of _____, _____, _____, and _____.

Formation of acid rain: _____

_____

_____

Sources of NO_x (NO and NO_2): *(natural)* _____
(man made) _____

Sources of SO_x (SO_2 and SO_3): *(natural)* _____
(man made) _____

What are the environmental problems associated with ACID RAIN?

- _____
- _____
- _____
- _____