Chemistry Unit 7

Primary reference: *CHEMISTRY*, Addison-Wesley

Topic	Essential Knowledge	Study Support
Scientific Investigation 1.7 SOL 1a,b,d,e,f,g	Always add acid slowly to water to prevent splattering. Review the equation to calculate percent error. Burets are used to add titrant during a titration.	
Atomic Structure and Periodic Relationships 2.7 SOL 2h	**Mixtures** may be homogeneous or heterogeneous. Homogeneous solutions do not scatter light, whereas heterogeneous liquids do scatter light (Tyndall Effect). Heterogeneous mixtures can be suspensions (particles separate upon standing) or colloids (particles stay dispersed upon standing).	**Ch 2:** Read pp 32-34 **Ch 17:** Read pp 490-491
Nomenclature, Formulas, and Reactions 3.7 SOL 3c, 3e	Empirical formulas show the lowest whole number ratio of element atoms present in a compound, whereas molecular formulas show the actual number of element atoms present in a molecule. Example: The molecular formula of benzene is C_6H_6, whereas the empirical formula is CH. In a **neutralization reaction**, an acid reacts with a base to form water and a salt. Example: HCl + NaOH $\rightarrow$ NaCl(salt) + H_2O. This is a type of double replacement reaction.	**Ch 7:** Read pp 194-195. Ch 21: Read pp 613-618
Molar Relationships 4.7 **SOL 4c, 4d**	A **solution** is a **homogeneous mixture** because the separate parts of the mixture cannot be seen. The **solvent** (usually water) is the part of the solution that is present in the largest amount. The **solute** is the substance that is dissolved. A **saturated** solution has all the dissolved solute that it can hold and can be identified by undissolved particles on the bottom after mixing. An **unsaturated** solution can still hold more solute. A **supersaturated** solution is a temporary condition where the solvent holds more solute than it can normally hold. The extra solute will eventually **precipitate** (crystallize). **Molarity** is a way of expressing concentration. $\boxed{\text{Molarity = moles solute/Liters solvent}}$ or $\boxed{M = mol/L}$ For dilutions use $\boxed{M_1V_1 = M_2V_2}$ where M = molarity and V = volume in mL or L **pH** is a number scale ranging from 0 to 14 that represents the acidity of a solution. The pH number shows the hydrogen (hydronium) ion concentration. The pOH number shows the hydroxide ion concentration. The higher the hydronium $[H_3O^+]$ concentration, the lower the pH number. $\boxed{\text{pH = log[H+] pOH = log[OH-] and pH + pOH = 14}}$ **Arrhenius Acids** are compounds that increase the concentration of hydrogen ions $[H^+]$ when they dissolve in water. Acid solutions have a pH below 7, taste sour and turn litmus paper red. **Arrhenius Bases** are compounds that increase the concentration of hydroxide $[OH^-]$ when they dissolve in water. Bases have a pH greater than 7, taste bitter, feel slippery and turn litmus paper blue. **Bronsted and Lowry** describe acids as proton donors and bases as proton acceptors. A proton (H^+) is also known as a hydrogen ion. In a sample of pure water a very small number of **water molecules dissociate**, producing equal concentrations of both hydrogen ions $[H^+]$ and hydroxide ions $[OH^-]$. The pH of pure water is 7. $H_2O \leftrightarrows H^+ + OH^-$ **or** $2H_2O \leftrightarrows H_3O^+ + OH^-$ Titration is a method used to determine the [OH-] or $[H_3O+]$ concentration. Indicators show color changes at certain pH levels.	**Ch 17:** Read pp 475-485 **Ch 18**: Read pp 500-505 **Ch 18**: Read pp 509-513 **Ch 20**: Read pp 577-600
Phases of Matter and Kinetic Molecular Theory 5.7 SOL 5g	Solutions can be a variety of solute/solvent combinations: gas/gas, gas/liquid, liquid/liquid solid/liquid or solid/solid. **"Like dissolves like"** Polar substances dissolve ionic or polar substances; non-polar substances dissolve non-polar substances. As the number of non-volatile solute particles increases, the freezing point decreases and the boiling point drops. In other words, the solution stays liquid over a wider temperature range than the pure solvent	**Ch 17:** Read pp 482-485, 490-493

Unit 7 Objectives

I) Empirical Formulas
II) Water's Properties
III) Liquid Solutions
  A) Dissolving Solutes
  B) Calculating Solution Concentrations
  C) Colligative Properties
IV) Non-liquid solutions
V) Acids and Bases
  A) Types of Acids and Bases
    1) Arrhenius
    2) Bronsted-Lowry
    3) Water—an Acid and a Base
  B) Acid Strength Calculations (pH and pOH)
  C) Neutralization Reactions and Titrations
  D) Comparing Acids and Bases
  E) Handling Acids and Bases Safely
  F) Dilutions with Acids and Bases
  G) Titration Calculations

<u>Learning Objectives (sol objectives)</u>

1. Identify empirical and Molecular Formulas(3c)
2. Convert molecular formulas to empirical formulas(3c)
3. Determine a compounds empirical formula from percent composition data(3c)
4. Determine a compounds molecular formula from percent composition data and molar mass data (3c)
5. Explain hydrogen bonding's impact on water's physical properties. (5e,d)
6. Distinguish between solutions, suspensions, colloids and emulsions (2h)
7. Recognize and predict that polar solvents dissolve polar solutes, and non-polar solvents dissolve non-polar solutes. "like dissolves like"(5c)
8. Explain which factors increase a solute's solubility <u>concentration</u>.(4c)
9. Interpret a solubility graph to predict how much solute will dissolve in solvent.(4c, 1g)
10. Use the molarity equation to calculate molarity, volume and moles for word problems. (4c)
11. Use the molarity dilution equation to calculate solution molarity or volume. (4e)
12. Explain how solutes affect solvent's boiling point and freezing point (Colligative properties)(5g)
13. Predict which solute will have the largest impact on a solutions boiling or melting point.(5g)
14. Define an acid and base using the Arrhenius and Bronsted-Lowry definitions(4d)
15. Identify acids and bases from their chemical formula(4d)
16. Explain the difference between a strong and weak acid(4d)
17. Explain that strong electrolytes dissociate completely and weak electrolytes dissociate partially.(4d)
18. Memorize that $[H^+]x[OH^-] = 1 \times 10^{-14}$ at 25°C in pure water(4d)
19. Calculate hydrogen and hydroxide ion concentrations in pure water(4d)
20. Convert between pH and pOH units to hydrogen and hydroxide ion concentrations in pure water.(4d)
21. Identify a solution as acidic or basic depending on its pH(4d)
22. Identify the products of an acid and base reaction(3e)
23. Identify a neutralization reaction(3e)
24. Identify the equipment typically used in a titration(1a)
25. Explain what an indicator is in an acid-base titration(4d, 1a)
26. Compare and contrast the properties of acids and bases(4d)
27. Explain how to safely handle acids and bases in the laboratory(1b)
28. Calculate the molarity of an unknown acid or base from titration data(4c)
29. Use a buret correctly.(1a)

Chapter 7 Empirical & Molecular Formula Notes

What is a molecular formula?	
What is an Empirical Formula?	

Write the empirical formulas for the following molecular formulas

Molecular Formula	Empirical Formula
N_2O_4	
C_6H_{12}	
$C_2H_2O_4$	
C_5H_{12}	

Empirical formulas are determined from lab data.

Example 1: A detective finds an unidentified powder next to a dead body. The forensic lab identifies the compound as containing 16.4% arsenic and 83.6% iodine. What is the empirical formula?

Approach:
1. Convert percent to grams of element
2. Find moles of each element
3. Make the moles into subscripts
4. Convert subscripts to whole numbers.

Example 2: A gas cylinder contains a gas that is 25.9% N and 74.1% O by mass. What is the empirical formula of the gas?

Example 3: A sample contains 8.57 grams of hydrogen and 59.43 grams of nitrogen. What is the compound's empirical formula?

Example 4: A compound contains 38.7% carbon, 9.7% hydrogen, and 51.6% oxygen. What is the compound's empirical formula?

Answers 3: NH_2, 4: CH_3O

Determining True Molecular Formulas Notes

Steps to solve a molecular formula problem:

1. Determine the empirical formula
2. Molecular mass/empirical mass = formula multiplier
3. Multiply all the empirical formulas subscripts by the multiplier

Example Problems

1. The empirical formula of a compound is NO_2. Its molecular mass is 92 g/mol. What is its molecular formula?

2. The empirical formula of a compound is CH. Its molecular mass is 65.09 g/mol. What is the compound's molecular formula?

3. A compound is found to be 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. Its molecular mass is 60.0 g/mol. What is its molecular formula?

Ch 17 & 18 solution notes (1st part with PPT)

1. What is a solution?	
2. What is a solute?	
3. What is a solvent?	
4. What is an aqueous solution?	
5. What is the difference between NaCl dissolving in water and dextrose dissolving in water?	$NaCl\ (s) \rightarrow$ $C_6H_{12}O_6(s) \rightarrow$
6. What does the expression "like dissolves like" mean?	
7. What is an electrolyte?	
8. What is the difference between a strong and weak electrolyte?	Strong Electrolytes ionize ____________________________ and weak electrolytes ionize ________________________________ non-electrolytes don't ionizie

Solubility Processes and Definitions

Solubility: Maximum amount of a substance that will dissolve in 100 mL of water at a specific temperature.

What happens as we begin adding NaCl to water?(pHet demo)

Unsaturated: Solution:____

____

Saturated: Solution ____

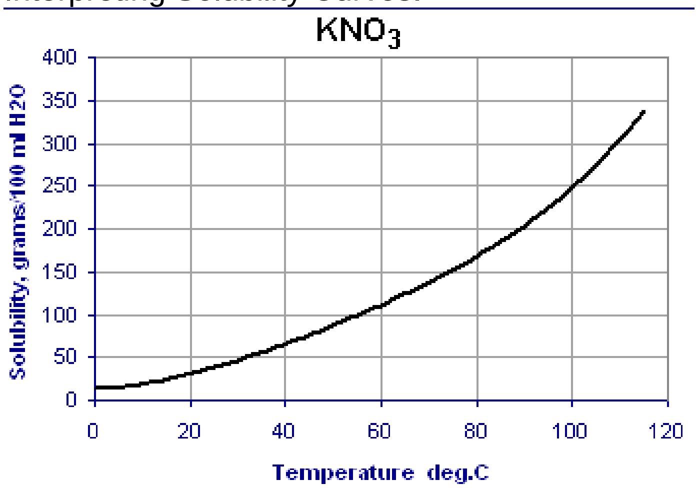
____

NaCl(s) ⇆ NaCl(aq)

Dynamic Equilibrium ____

____

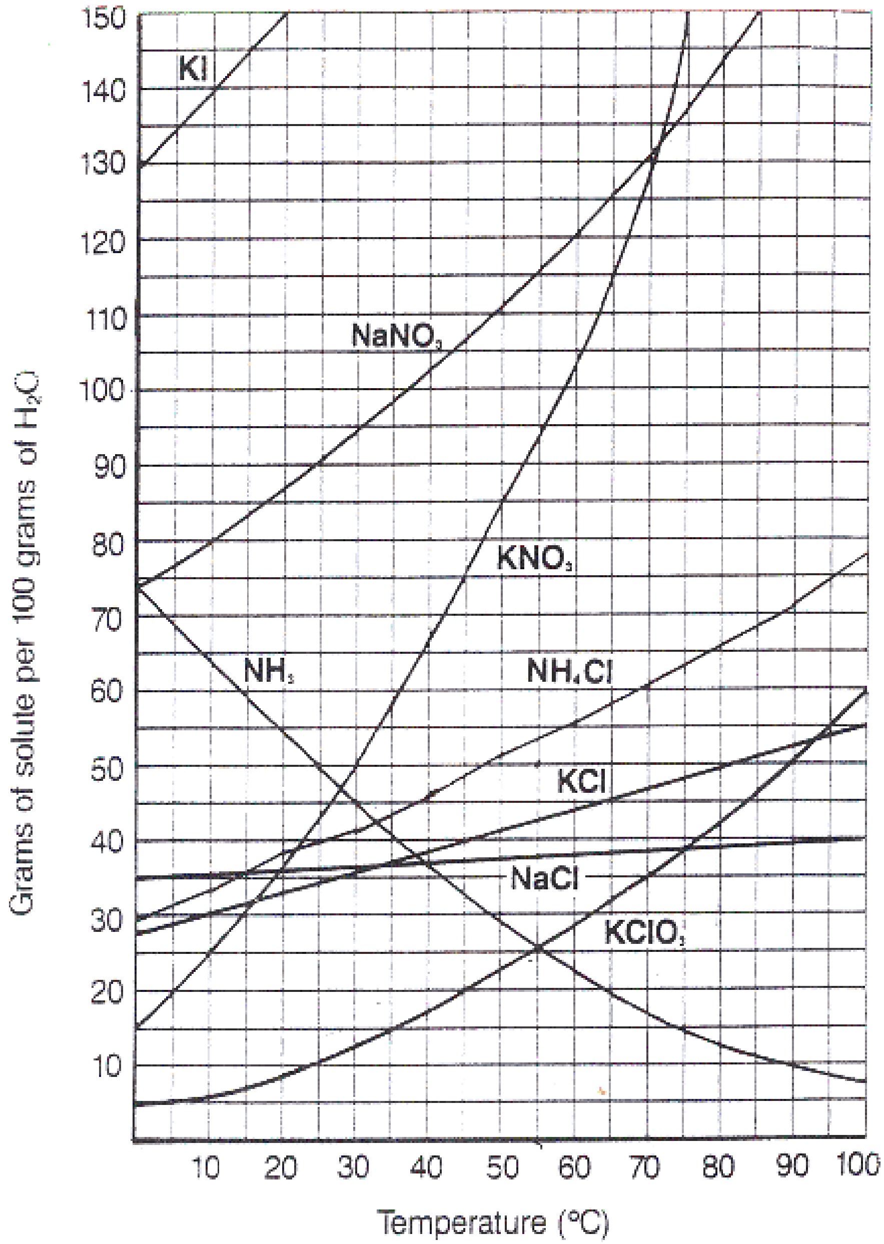
Interpreting Solubility Curves.



How many grams of KNO_3 can you dissolve in 100 grams of water at 100°C?

How many grams of KNO_3 can you dissolve in 200 grams of water at 20°C?

Use solubility curves to predict how much solute will dissolve in a solvent at any given temperature



Which compound is least soluble at 30°C?

Which compound is most soluble at 10˚C?

How many grams of KCl will dissolve in 100 mL of water at 80°C?

- In general, do compounds become more soluble as temperature increases?
- Which compound is most soluble at 20 degrees C?
- Which compound is least soluble at 60 degrees C?
- How many grams of $NaNO_3$ will dissolve in 100 mL of water at 60 degrees C?

Calculating Concentrations of Solutions

A) Molarity: moles of solute dissolved in 1 liter of solution.

1) Symbol: M as in 4.0M NaCl stated as ________________________________

2) Formula and rearrangement:

B) Solving molarity problems
   1) Calculate moles of solute using molar mass
   2) Change all volumes to Liters
   3) Use the formula

55 mL = ________ L

328 mL = ________ L

0.42 L = _________ mL

C) Sample molarity problems
   1) 4.0 Liters of a dilute aqueous solution of sulfuric acid, H_2SO_4, contains 0.20 moles of sulfuric acid. What is the solution's molarity? (ans = 0.050 M)

   2) 1.2 Liters of dilute potassium hydroxide base, KOH, contains 0.25 moles of KOH. What is the solution's molarity.(Ans = 0.21 M)

   3) 15 mL of dilute sodium hydroxide base, NaOH, contains 1.3 grams of NaOH. What is the solution's molarity?(Ans = 2.2 M)

   4) 230 mLs of dilute hydrochloric acid, HCl, contains 1.68 grams of HCl. What is the solution's molarity?(Ans = 0.20 M)

$\boxed{M = \frac{\text{mol}}{\text{L}}}$ or $\boxed{\text{M x L = moles}}$

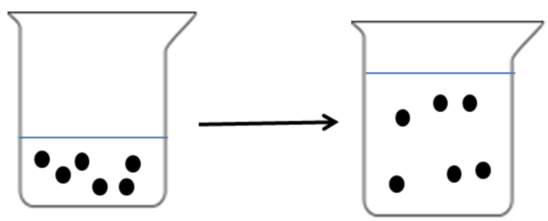
5) How many grams of NaCl will be needed to prepare 2.00 liters of 1.5M NaCl?(Ans = 180 g)

6) How many grams of $CaCl_2$ will be needed to prepare 500. mL of 1.3M $CaCl_2$?(Ans=72 g)

D) Making dilute solutions from concentrated solutions—calculations

1) Moles of solute before dilution = moles of solute after dilution

10 mL of concentrated solution



30 mL of dilute solution

mole solute 1 = mole solute 2

M_1L_1 = M_2L_2

2) Formula for diluting molar solutions: $\boxed{M_1V_1 = M_2V_2}$ volume units must match

E) Sample Molarity Dilution problems

1) How many Liters of 6.0M HCl will be required to prepare 52 Liters of 1.0 M HCl?(Ans=8.7L)

2) How many mLs of 1.5M NaOH will be required to prepare 250 mLs of 0.25 molar NaOH?
(Ans = 42 mL)

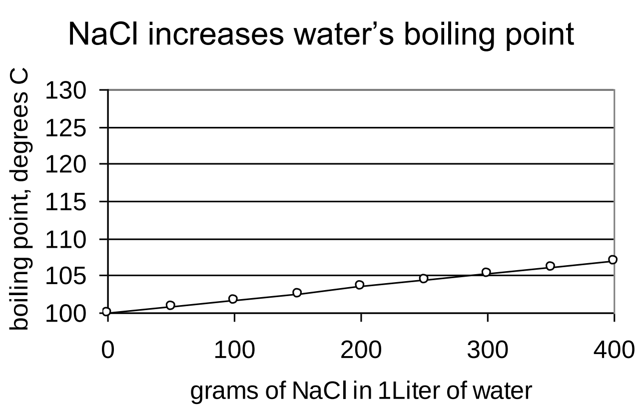
3) How many mLs of 3.0M glucose must be used to prepare 2 liters of 0.15 molar glucose? (Ans=100 mL)

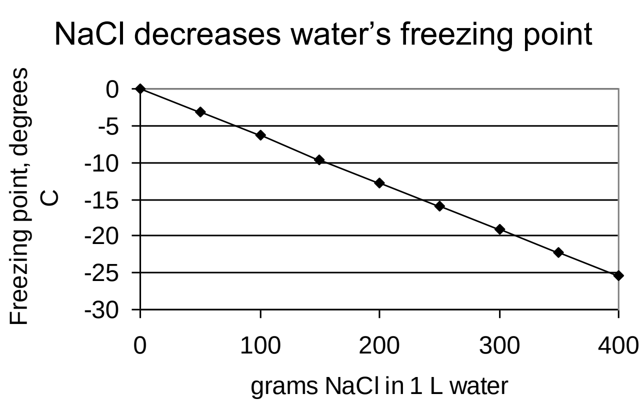
4) What will be the molarity of 15 mL of 1.5 molar nitric acid, HNO_3, diluted to 300 mL? (Ans = 0.075M)

5) What will be the molarity of 22 mL of 5.2 molar sodium hydroxide base diluted to 2.0 liters?(Ans = 0.0572 M)

III: Colligative Properties

Colligative properties depend on the number of ions or molecules that are dissolved in a solvent.





Solution freezing point decreases with increased particles(ions or molecules)
Examples:

Solution boiling point increases with increased particles (ions or molecules)
Examples:

A) Counting Particles for Colligative Properties. Particles are ions or molecules.

Dissociation: ____

____

Examples:

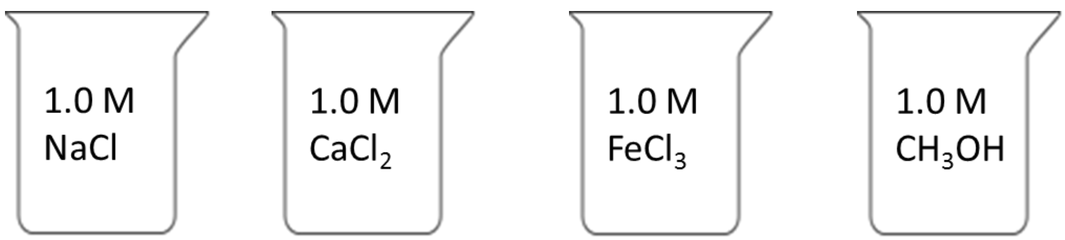
NaCl in water $\rightarrow$

$CaCl_2$ in water $\rightarrow$

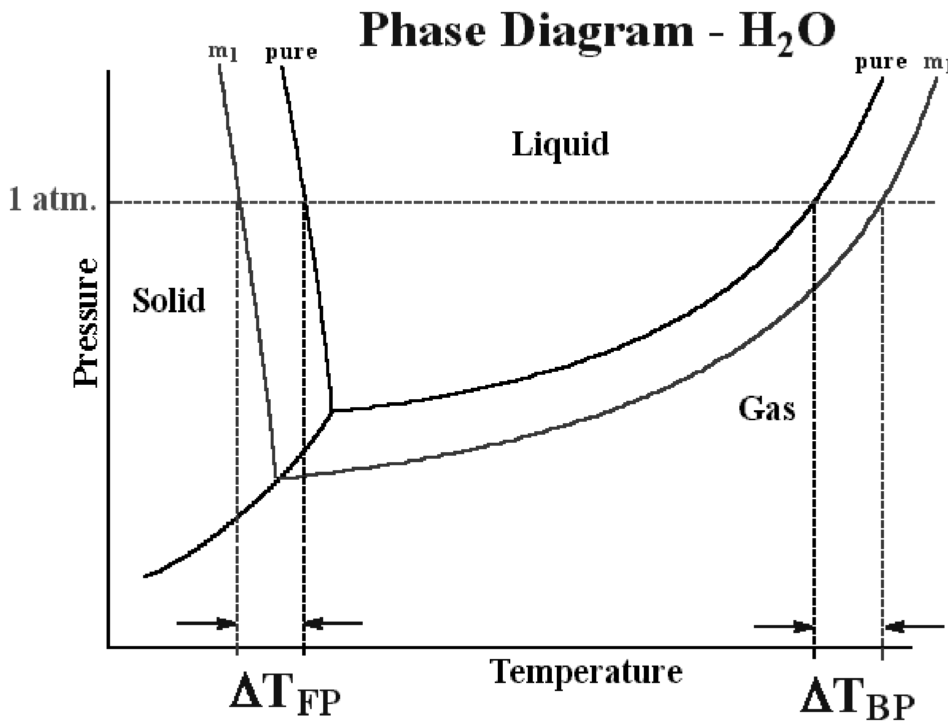
$FeCl_3$ in water $\rightarrow$

CCl_4 in water $\rightarrow$

If you have 0.1 M solutions of NaCl, $CaCl_2$, $FeCl_3$ and CCl_4, which one would have the highest boiling point?



Which one would have the lowest freezing point?



Other Homogeneous Solution Types and Examples

Solute/Solvent	Examples
Gas/gas	Air is a mix of N_2, O_2 and other gases
Gas/liquid	Soft Drinks have $CO_2(g)$ dissolved in water.
Liquid/liquid	Vinegar is acetic acid dissolved in water.
Solid/liquid	Salt water has NaCl dissolved in water
Solid/solid	Alloys are metals dissolved in other metals. Brass is copper and zinc Bronze is copper and tin

Solubility Curves

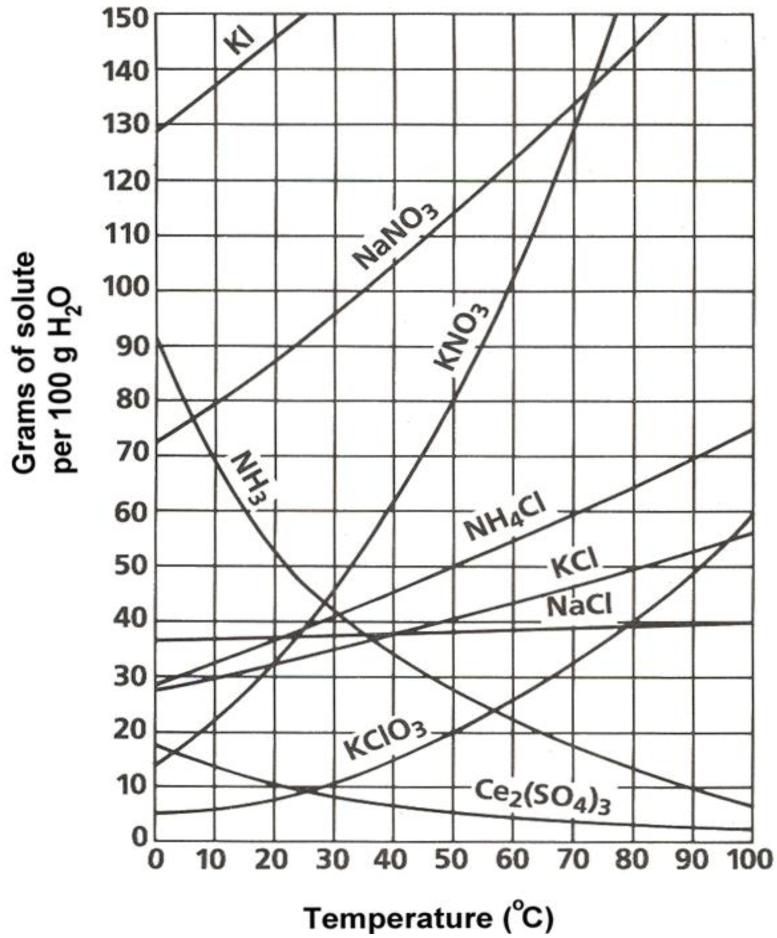
Easy:

1) Which substance is most soluble at 10°C?

2) What is the solubility of KNO_3 in 100 g water that is 70°C?

3) If you have an unsaturated solution of cerium(III) sulfate and you want to add more solute, should you heat it or cool it? Explain.

4) How much potassium chloride can be added to 100 g of 90°C H_2O?



Medium:

5) What is the maximum mass of potassium nitrate that will dissolve in 200 g of water that is 323 K?

6) At approximately what temperature will 0.163 moles of potassium chlorate dissolve in 100 g of H_2O?

7) If you have 575 g of water (at 80°C), what is the maximum mass of sodium nitrate that will dissolve?

8) A student wants to completely dissolve 400 grams of potassium nitrate using warm tap water, which is 50°C. How much water should he use?

Hard:

9) An ammonium chloride solution with a volume of ~250 mL cooled from 70°C to 50°C. How many grams of the solute were present at 70°? How many grams of solute crystalized and precipitated once the temperature reached 50°C?

10) 30 g of ammonium chloride, 30 g of sodium nitrate, and 30 g of potassium chlorate are all mixed and dissolved in an unknown mass of water that has an initial temperature of 350 K. The temperature of the solution is then decreased from 350 K to 307 K. Which salt(s) will be the first salt(s) to visibly crystalize/precipitate? Justify your answer.

Molarity is the concentration of a solute in a solution.

Molarity is the <u>moles of solute</u> in the <u>volume of total solution</u> (measured in liters).

The molarity equation is:

…and the units are either

or

The molarity of a solution is most often expressed with a capital M. A solution with a molarity of **3 M** is spoken as "**three molar**"

THE MOLE IS THE GOAL, and if you don't have the moles of solute, you need to find them.

Water is the universal solvent. You can assume that whenever you see the phrase "solution of X," the X is the solute (with a certain molarity) and water is the solvent.

Square brackets mean "concentration of…" For example, [$C_6H_{12}O_6$] means "concentration of glucose. [K+] means "concentration of potassium ion."

You may be tasked with solving for any of the following:

- molarity (M) of the entire solution
- number of *moles* of solute (mol) or the *moles* of an ion in the solute after dissociation
- *mass* of the solute (g) after you calculate the moles
- volume of the solution (L)

Practice writing balanced ionic equations:

1) Sodium chloride completely dissolves in water to create a sodium chloride solution.

2) Calcium fluoride completely dissolves in water to create a calcium fluoride solution.

3) Lead(II) nitrate completely dissolves in water to create a lead(II) nitrate solution.

4) Copper(II) sulfate completely dissolves in water to create a copper(II) sulfate solution.

Molarity practice:

5) What is the molarity of ribose, $C_5H_{10}O_5(s)$ when 4 moles of it are dissolved to create 2 L or $C_5H_{10}O_5(aq)$ solution?

6) How many grams of ribose are in a 2.3 M solution with volume of 700 mL?

7) Write the balanced ionic equation for the complete dissociation of solid magnesium chloride into aqueous magnesium chloride ions:

    a. How many moles of Mg^{2+} ions are in the solution when a mole of $MgCl_2$ dissociates?

    b. How many moles of Cl^- are in solution?

8) Calculate $[Mg^{2+}]$ present in a 500 mL solution of completely dissociated $MgCl_{2.}$ The mass of magnesium chloride used was 62 g.

9) Calculate $[Cl^-]$ present in the same 500 mL solution of completely dissociated $MgCl_{2.}$ The mass of magnesium chloride used was 62 g.

10) In order to make a 1.5 L solution of potassium chloride with a molarity of 0.80 M, how many grams of the solute do you need?

11) 13.3 g of lithium bromide was dissolved in water, and a 0.2 M LiBr(*aq*) solution was created. Determine the volume of the solution in liters.

12) How many grams of glucose are needed to create 355 mL of a 12 M solution?

13) Scott measured 121.07 g of ammonium nitrate and created a solution with a volume of 5.38 × 103 mL. Find the concentration of the solute.

14) Find [Na+] when 9.00 g sodium fluoride completely dissociates in water, and a solution with a volume of 415 mL is created.

Dilutions:

The dilution equation is $\mathbf{M_1V_1 = M_2V_2}$. When you have a solution with a certain concentration and you add water to it, it will be diluted. (If you have very sweet koolaid and you want to lower the sugar concentration, you add water… which increases the volume).

M_1 is ____

V_1 is ____

M_2 is ____

V_2 is ____

Adding water to a concentrated solution will lower the molarity of the final solution (meaning it's *less* concentrated). The opposite can be done too: removing water from the solution will make the final solution *more* concentrated (because $V_2 < V_1$)

Chapter 20&21: Notes for Acids and Bases—the Basics

I) Classifying Acids and Bases

A) Arrhenius Definition of Acids and Bases

Acids____

____

Examples:

$HCl \xrightarrow{100\%}$

$HNO_3 \xrightarrow{100\%}$

1) Strong Acid: An acid that dissociates completely to H^+ and anions in water.

2) Weak Acid: An acid that only partly dissociates to H^+ and anions in water.

Example: $HF \xrightarrow{1\%} H^+(aq) + F^-(aq)$

3) Strong acids are strong electrolytes, weak acids are weak electrolytes.

Bases____

____

Examples

$NaOH \xrightarrow{100\%}$

$Ca(OH)_2 \xrightarrow{100\%}$

B) Bronsted-Lowry Definition of Acids and Bases

Acid____

____

Base____

$$NH_3(aq) + H_2O \leftrightarrow NH_4^+(aq) + OH^-(aq)$$

Water—an acid and a base

$$H_2O + H_2O \xrightarrow{1x10^{-5}\%} H_3O^+ + OH^-$$

C) In pure water at 25°C:
H^+ concentration = 1 x 10^{-7}M and OH^- concentration = 1 x 10^{-7}M

1) Symbol for concentration in molarity: _____________

II) Calculating $[H^+]$ and $[OH^-]$ in aqueous solutions.

$\boxed{[H^+]x[OH^-] = 1 \text{ x } 10^{-14} \text{ ALWAYS, ALWAYS, ALWAYS at 25°, 1 atmosphere.}}$

We make solution of 0.01 M HCl in water. What is the concentration of OH^- ions?

We make a solution of 0.0001 M NaOH in water. What is the concentration of H^+ ions?

pH and pOH—A more convenient way to express H^+ and OH^- concentration.

pH = -log$[H^+]$
pOH = -log$[OH^-]$
pH + pOH = 14 (ALWAYS in water at 25°C)

Base 10 Logarithms--Logarithms are exponents

The log of 100 = 2

The log of 1000 = 3

The log of 10,000 = ____

The log of 0.1 = -1

The log of 0.01 = -2

The log of 0.001 = _____

Using your calculator to find logarithms

Log(120) = 2.08 or $10^{2.08}$ = 120

Log(0.043) = -1.37

pH is a way of expressing [H+]

$\boxed{pH = -\log[H+] \qquad pOH = -\log[OH-] \quad \text{and} \quad pH + pOH = 14}$

1. Find the pH of 0.0034 M HCl (Ans = 2.47)

2. Find the pH of 0.00023 M HNO_3(Ans = 3.64)

3. Find the pOH of 0.0012 M NaOH (Ans = 2.92)

4. Find the pH of 3 x 10^{-5} M NaOH--careful (Ans = 9.48)

Now, let's go the other way.

Example: A solution has a pH of 2.3. What is the [H+]?

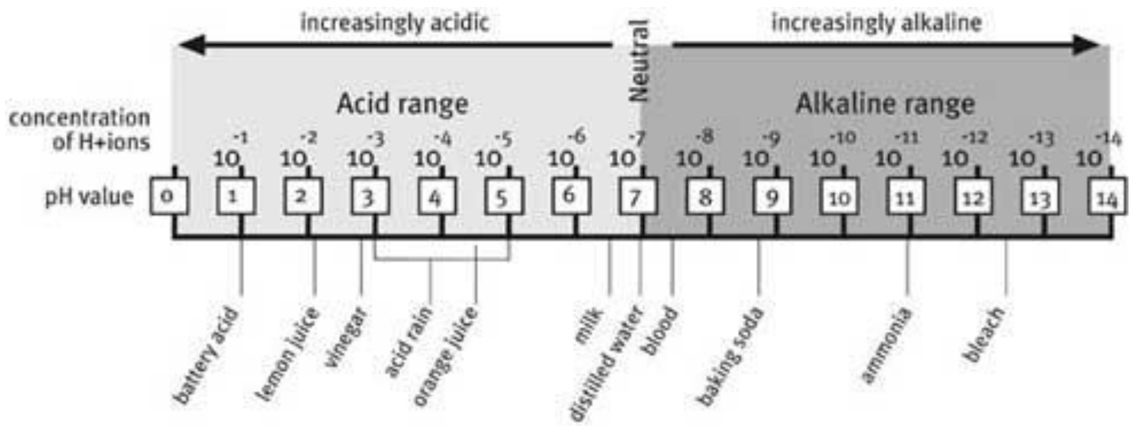
$10^{-2.3}$ = 0.0050 M

1. A solution has a pH of 4.32. What is the [H+]?(Ans = 4.79 x 10^{-5} M)

2. A solution has a pH of 5.6. What is the [H+]? (Ans = 2.51 x 10^{-6} M)

3. A solution has a pOH of 6.8. What is the [OH-]? (Ans = 1.58 x 10^{-7} M)

pH and Identifying acidic and basic solutions



1) ACIDIC: pH less than 7

2) BASIC: pH greater than 7

3) NEUTRAL: pH = 7

III) Neutralization Reactions

Neutralization____

____

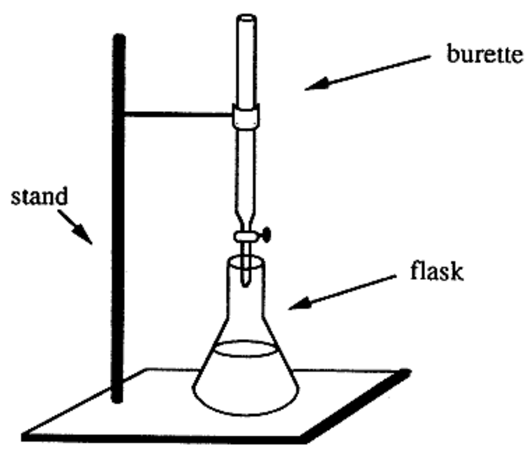
Example: $HCl + NaOH \rightarrow$ reaction type?:

Example: $HNO_3 + KOH \rightarrow$

Complete neutralization when pH = 7 and $[H^+] = [OH^-]$

Titrations

A method used to determine the concentration of an acid or base using a buret filled with a standard solution, a sample of unknown concentration, and an indicator (frequently phenolphthalein).



Acid-Base Titration with a pH Meter.

Objective: Learn to interpret a pH curve to determine the concentration of an unknown acid solution using a base of known concentration.

Site: http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s5_5.swf

Look at the screen and record the following information:

Concentration of NaOH solution in buret	M
Initial pH of unknown acidic solution	
mL of unknown acidic solution	mL
Initial buret volume	mL

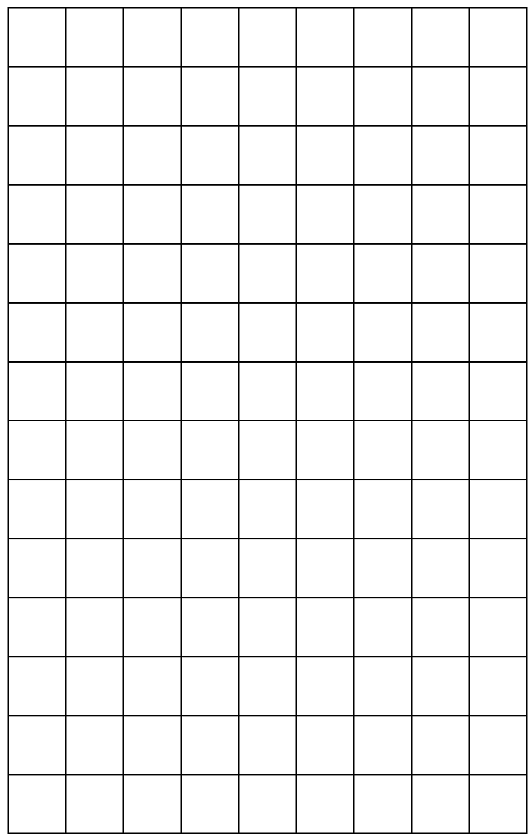
You will be recording the mLs of base and the corresponding unknown pH as you add base to the solution. Then you will graph mL vs pH to determine the endpoint.

mL	pH

Mark the equivalence point on the titration curve.

pH Titration Curve

pH of unknow solution



mL base added

Using the data, calculate the molarity of the unknown acid solution. Show your work.

Acid – Base Titration Calculations

Example 1: How many moles of potassium hydroxide would be required to completely neutralize 1.2 moles of hydrofluoric acid?

$$HF + KOH \rightarrow H_2O + KF$$

Example 2: How many moles of sodium hydroxide would be required to completely neutralize 0.60 moles of phosphoric acid?

$$H_3PO_4 + 3NaOH \rightarrow 3H_2O + Na_3PO_4$$

In titration, we add an aqueous solution in the buret to an aqueous sample. If the mole ratio in the balanced equation is 1:1, we can use the equation below to calculate molarity:

$$\mathbf{M_{acid} \times V_{acid} = M_{base} \times V_{base}}$$

Example 3: If 27 mL of 0.2 M KOH is needed to titrate a 15 mL sample of hydrofluoric acid,HF to its phenolphthalein endpoint, what is the molarity of the HF?(Ans = 0.36 M)

$$HF + KOH \rightarrow H_2O + KF$$

Example 4: If 30. mL of 0.1 M KOH is needed to titrate a 40. mL sample of nitric acid, HNO_3 to its phenolphthalein endpoint, what is the concentration of the HNO_3?(Ans = 0.075 M)

$$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(s) + H_2O(l)$$

Example 5: How many mLs of 0.50 molar HCl is needed to neutralize 620 mLs of 0.20 molar NaOH?(Ans = 250 mL)

$$HCl + NaOH \rightarrow H_2O + NaCl$$

D) Comparing Acids and Bases

	Acids	Bases
Ions produced in water		
Typical Compound formula		
Taste/feel		
pH range		
Litmus paper turns:		
Common Household Products		
Affect on human tissue and how to treat		

E) Safety in handling acids and bases

When you mix acid and water ______________________________

Why?______________________________

If you get acid or base on skin ______________________________

Why:______________________________

Using Sodium hydrogen carbonate (Baking Soda) to neutralize acid spills

$HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$

F) Putting acid/base chemistry together with molarity and dilutions:

What is the molarity of 18 grams of NaOH dissolved in 400 mL of solution?(Ans=1.1 M NaOH)

How many grams of H_2SO_4 would be required to prepare 320. mLs of 0.20M H_2SO_4?(Ans =6.3 g)

You need to prepare 2.0 liters of 0.50M HCl using a stock solution 6.0 M HCl. How many mLs of the stock solution do you need?(Ans=170 mL)

How many liters of 0.05M HCl can you prepare using 21 mL of 2M HCl (Ans=0.8 L)