Chapter 4 Problems:

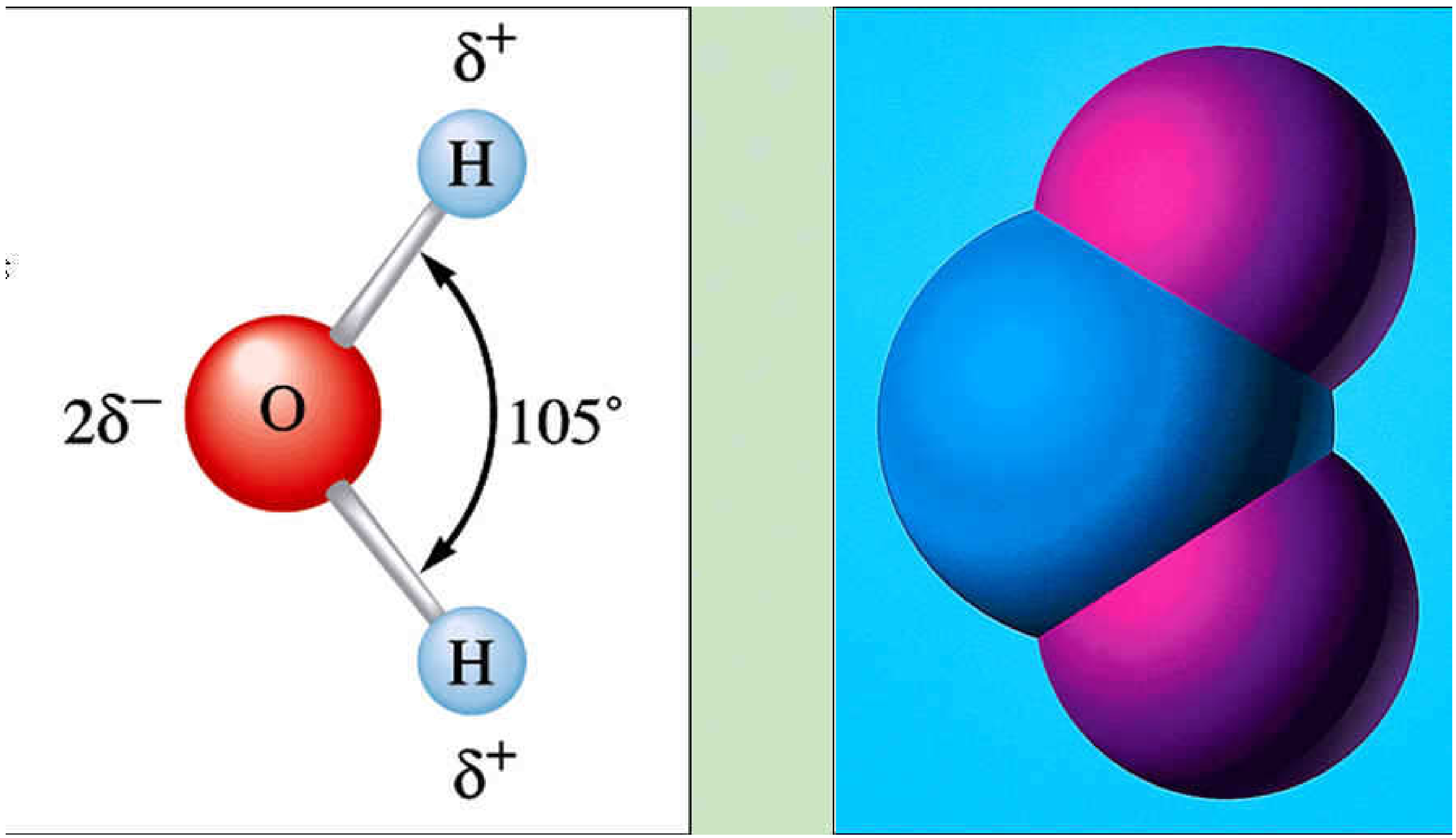
11 (a-d), 12, 18, 19, 23, 24a, 25, 36, 38, 41, 43, 47, 50, 51, 56, 57, 61, 63, 65b, 74, 76, 78, 81, 84, 87, 90, 92-95

Chapter 4 – Types of chemical reactions, and solution stoichiometry

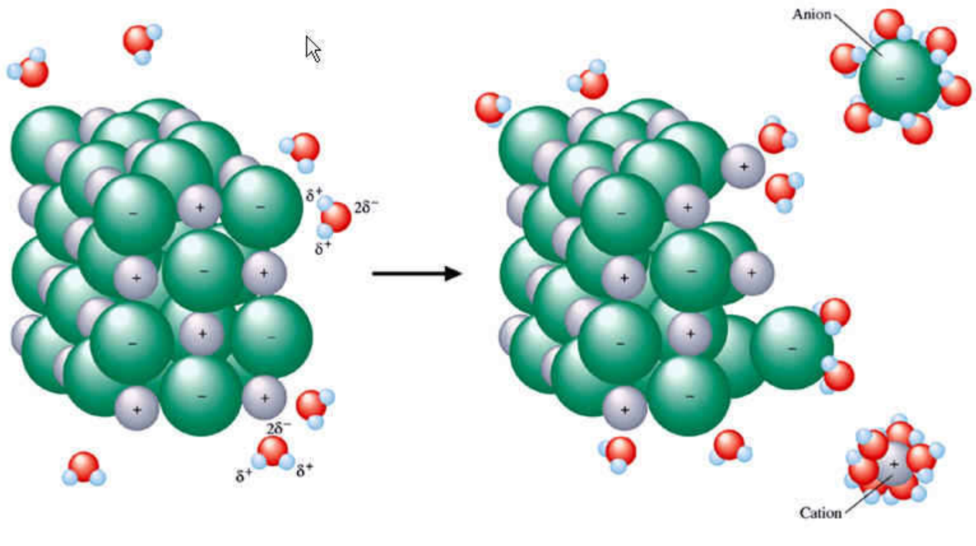
Properties of water:

bent or v-shaped

polar molecule



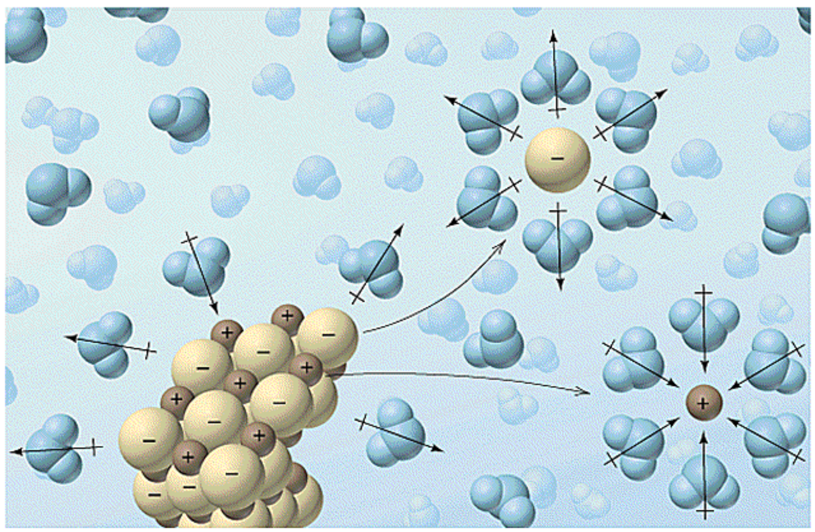
Ionic solids – many are soluble in water

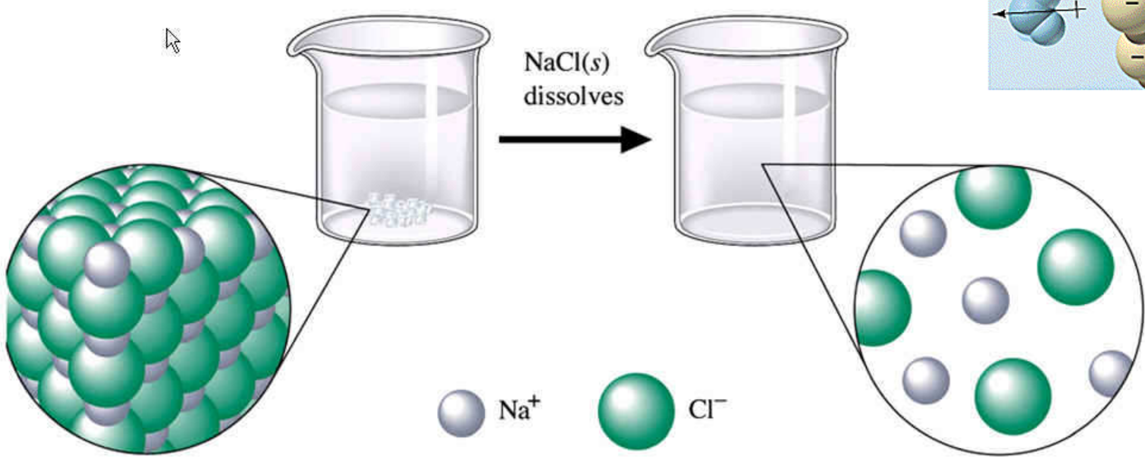


Dissociation equation of NaCl:

$$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Hydrated ion or hydration – when a cation or anion are completely surrounded by water molecules.





Solubility of ionic compounds in water depends on the strength of the bond between the cations and anions, and the bond between ions and water.

AgCl is not very soluble in water

NaCl is soluble in water

Which ionic solid has stronger bonds?

Non ionic compounds – soluble in water if they are polar.

Grease is non polar…and doesn't dissolve in water.

This table will need to be memorized for your a.p. test!

TABLE 4.1 Simple Rules for the Solubility of Salts in Water

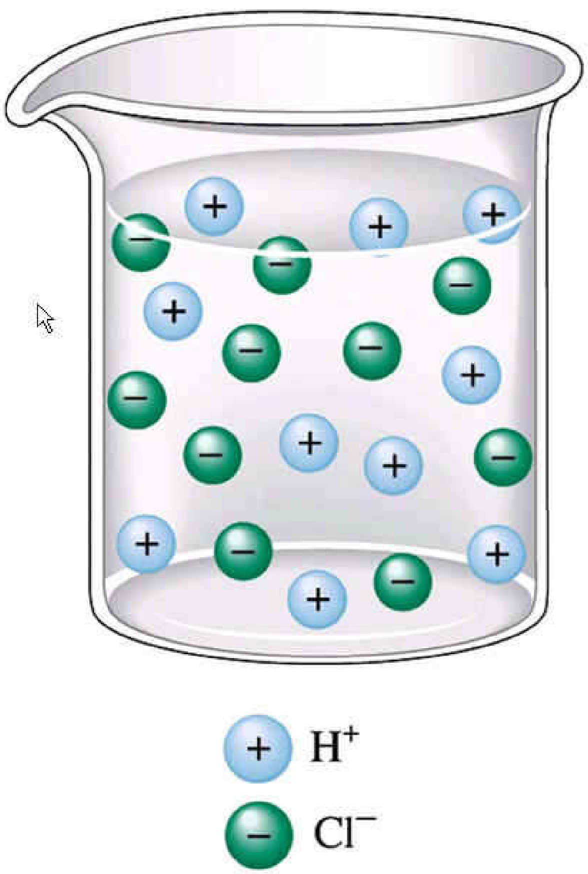
1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts containing the alkali metal ions (Li^+, Na^+, K^+, Cs^+, Rb^+) and the ammonium ion (NH_4^+) are soluble.
3. Most chloride, bromide, and iodide salts are soluble. Notable exceptions are salts containing the ions Ag^+, Pb^{2+}, and Hg_2^{2+}.
4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, Hg_2SO_4, and $CaSO_4$.
5. Most hydroxide salts are only slightly soluble. The important soluble hydroxides are NaOH and KOH. The compounds $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$ are marginally soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.

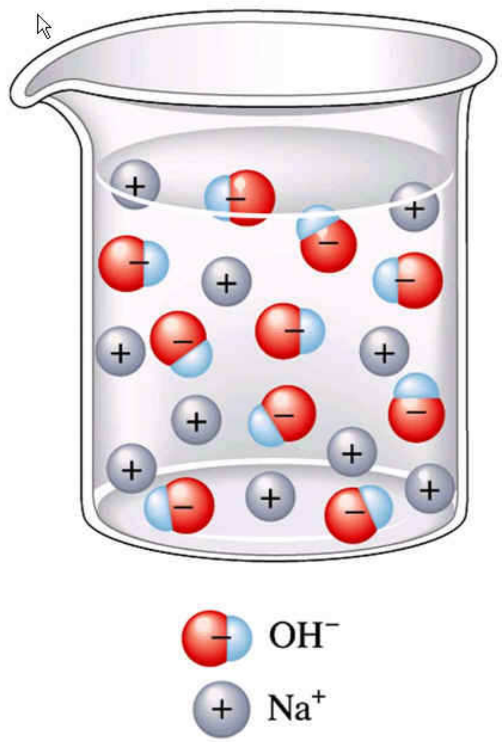
Aqueous solutions:

Electrolyte solution – a solution that is conductive to electricity because of the presence of dissociated ions in the solution.

Strong Electrolyte – contains a lot of dissolved ions

1) **soluble salts (ionic substances)**
2) **strong acids (HCl, $HClO_4$, HBr, HI, H_2SO_4, HNO_3**
3) **strong bases (alkali metal hydroxides)**

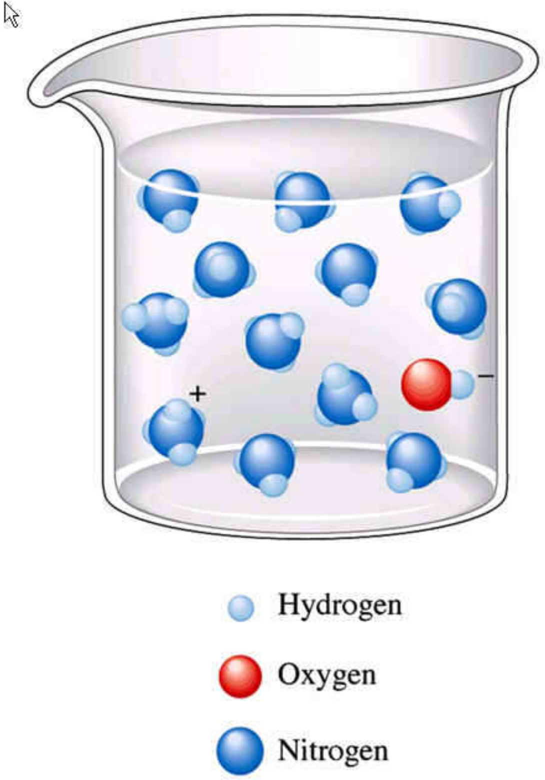




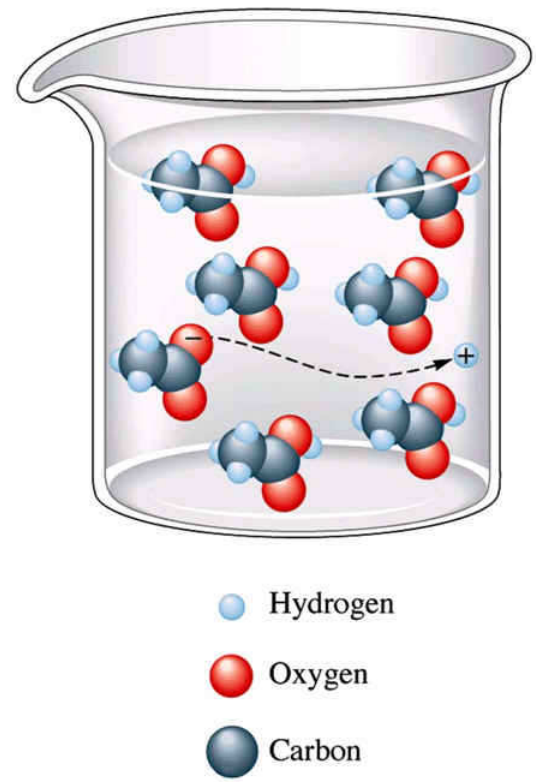
Weak Electrolyte – very few ions present

1) **weak acid – dissociate very little into ions.**
2) **weak base – dissociate very little into ions**

$$HC_2H_3O_2 + H_2O \leftrightarrow H_3O^+ + C_2H_3O_2^-$$



$$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$



Non-electrolytes – dissolve but don't produce any ions

- usually molecular compounds with only non-metal elements

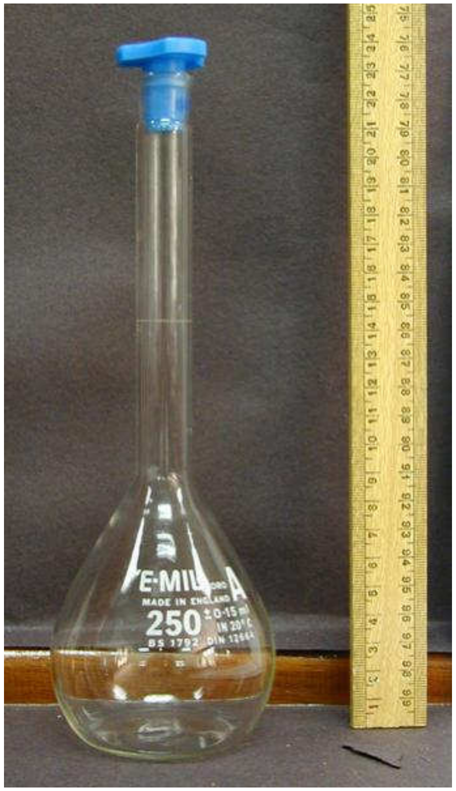
methanol (CH_3OH), glycerol ($C_3H_8O_3$)

Molarity – (M) - used to measure the concentration of solutions

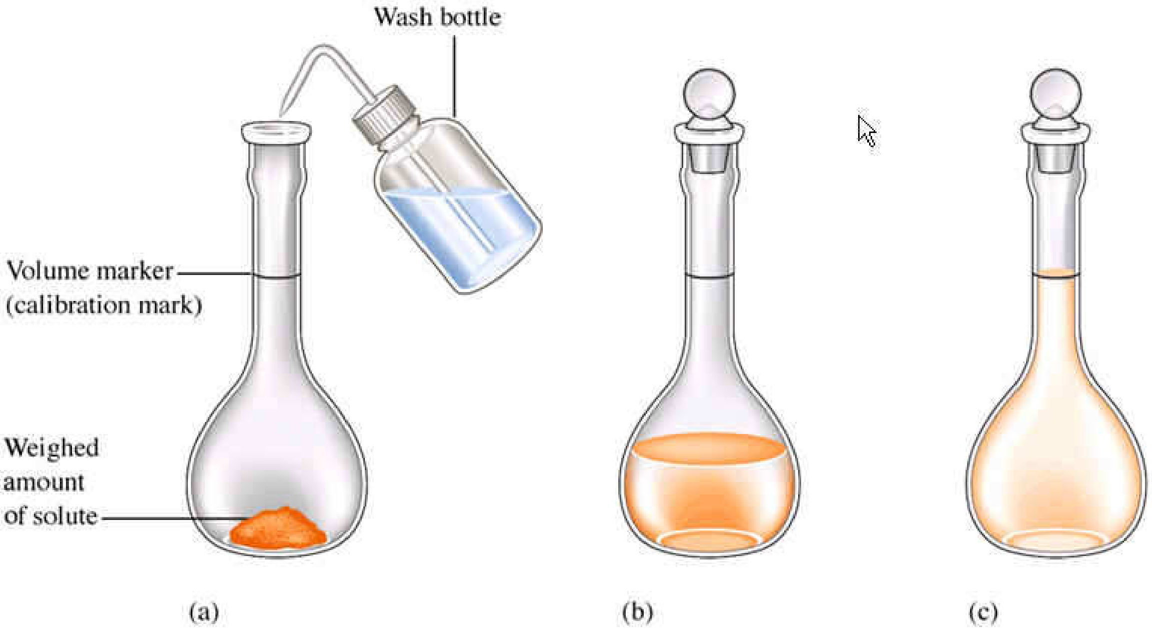
$$M = \frac{\text{moles solute}}{\text{liters solution}}$$

Problem: How would you prepare 1.50 Liters of 0.192 M NaOH?

Add 11.5 g NaOH + enough water to make 1.50 Liters of solution.



Is this solution really 0.192 M NaOH?



Problem: Determine the concentration of each ion in a 1.5 M $Fe(ClO_4)_{3\ (aq)}$ solution.

Determine the concentration of each ion in a 1.5 M $H_2C_2O_{4(aq)}$ solution.

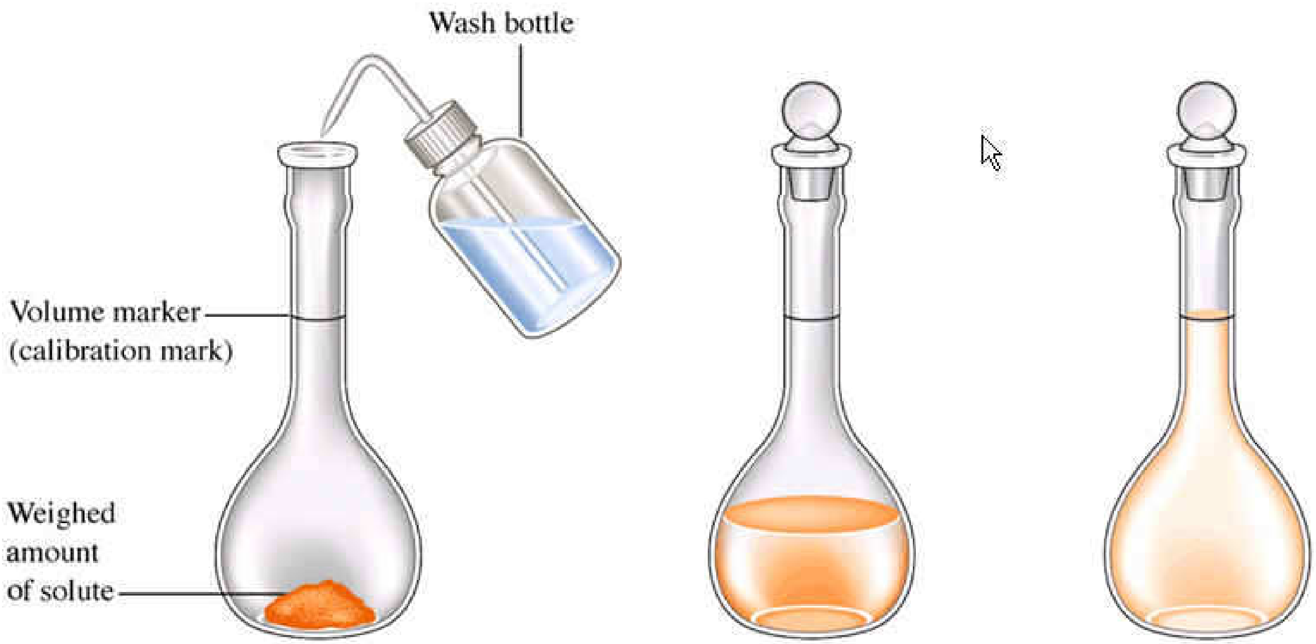
Blood serum is approximately .14 M NaCl. What volume of blood contains 1.0 mg NaCl?

Dilution

$$M_1 \ V_1 = M_2 \ V_2$$

Stock solution – concentrated solution, used to dilute and make a dilute solution.

Problem: 1.5 L of .10 M H_2SO_4 are needed. You have 18.3 M Stock H_2SO_4. How would you make this solution?



(a) (b) (c)

Problem: 15.00 mL KNO_3 was diluted to 125.0 mL, then 25.00 mL of this diluted solution was diluted further to 1000. mL. The final concentration was 0.00383 M. What was the concentration of the original solution?

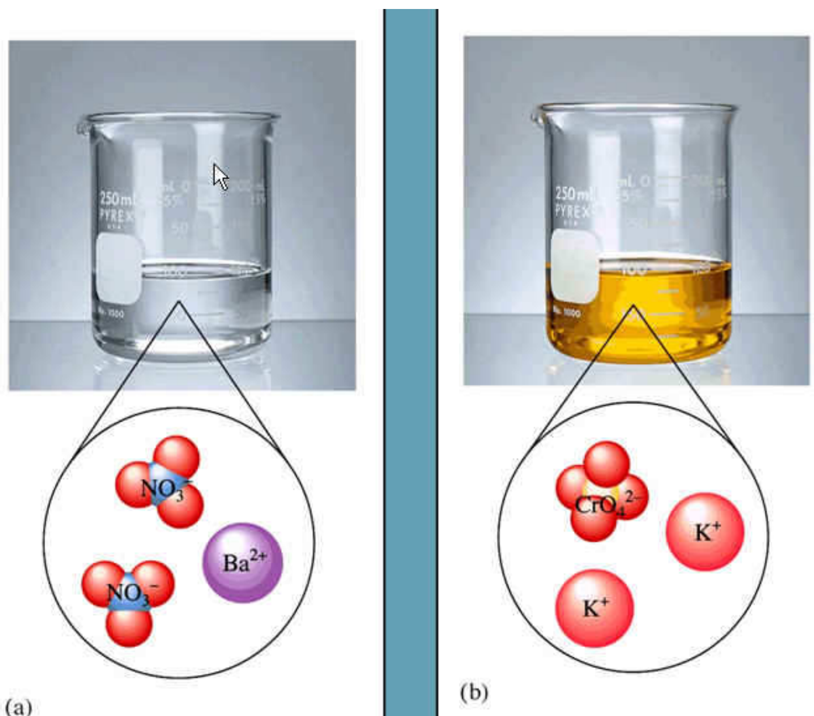
Types of Reactions:

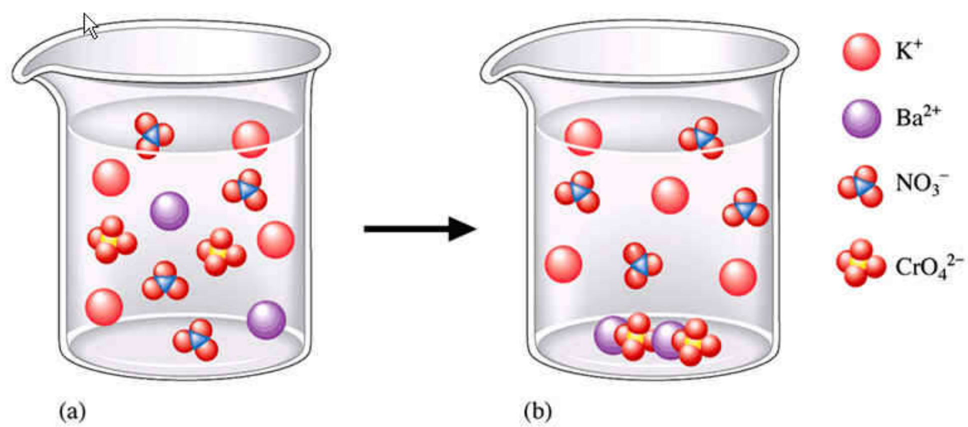
1) Precipitate or double replacement:

$$AB_{(aq)} + XY_{(aq)} \rightarrow AY + XB$$ **where AY and/or XB are insoluble solids**

What will happen if $Ba(NO_3)_{2(s)}$ is added to $K_2CrO_{4(s)}$??

What will happen if $Ba(NO_3)_{2(aq)}$ is added to $K_2CrO_{4(aq)}$??





The reaction between $Ba(NO_3)_{2(aq)}$ plus $K_2CrO_{4(aq)}$ can be represented by 1 of 3 equations:

1. **Molecular – shows the reactants and products as molecular aqueous substances.**
2. **Complete Ionic – shows the soluble reactants and products as dissociated ions.**
3. **Net Ionic – eliminates spectator ions from the complete ionic equation.**

Stoichiometry of precipitate reactions – (also called Gravimetric Analysis)

Problem: 1.25 L of 0.0500 M $Pb(NO_3)_2$ are added to 2.00 L of 0.0250 M Na_2SO_4

How many grams of $PbSO_{4(s)}$ will form?

Problem: A mixture contains $BaCl_2 \cdot 2H_2O$ and $Na_3PO_4 \cdot 12H_2O$

A student weighs 1.00 g of the mixture, and adds distilled water. What is the ppt.?

The ppt. is filtered, washed, dried, and then weighed at 0.48 grams.

The supernatant (liquid that passes through the filter paper is then divided into 2 beakers. To 1 beaker, $BaCl_{2(aq)}$ is added, the other beaker has $Na_3PO_{4(aq)}$ added.

The solution in the beaker that has $BaCl_{2(aq)}$ added forms a ppt., the other remains clear.

Determine the percent by mass of each salt in the mixture.

2) Acid/Base Reactions (neutralization) –

Problem: Write the molecular, complete and net ionic equations for the following:

a) **strong acid + strong base:** $HCl_{(aq)} + NaOH_{(aq)}$

b) **strong acid + weak base:** $HCl_{(aq)} + NH_{3(aq)}$

c) **weak acid + strong base:** $HC_2H_3O_{2\ (aq)} + NaOH_{\ (aq)}$

d) **weak acid + weak base:** $HC_2H_3O_{2\ (aq)} + NH_{3(aq)}$

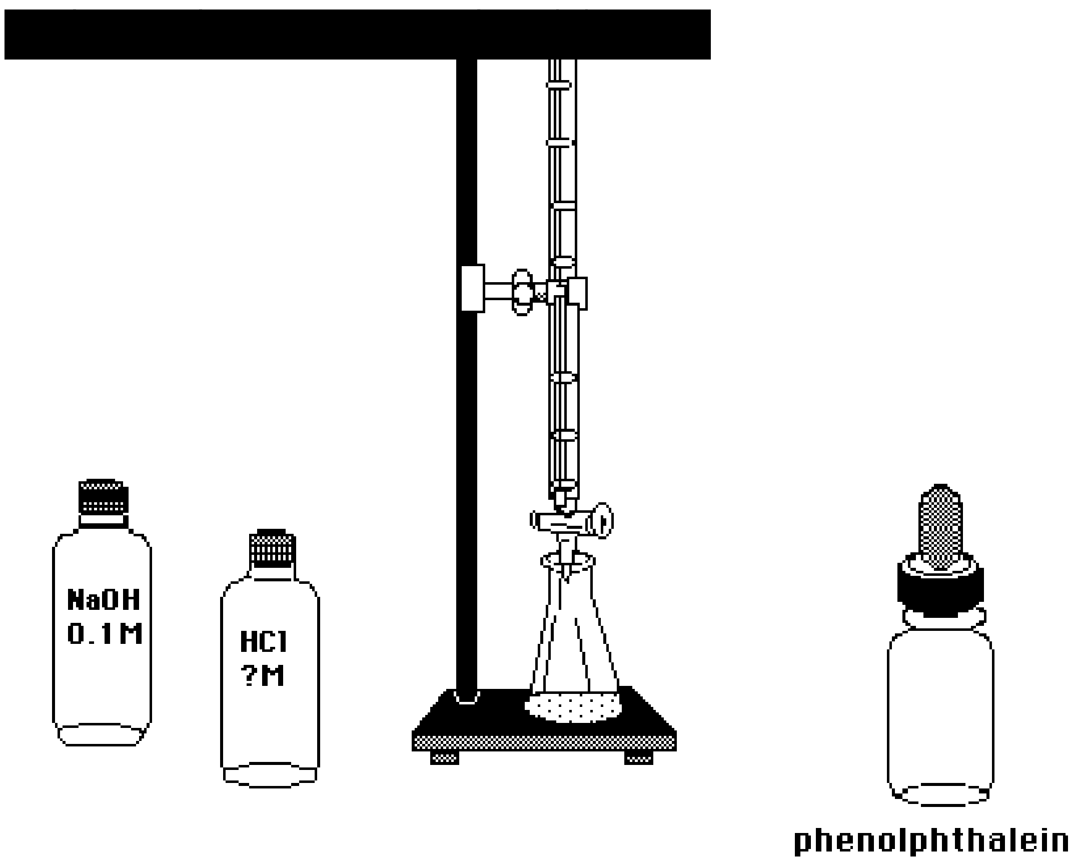
Acid – proton or H^+ donor
Base – proton or H^+ acceptor

Using these definitions, identify the acid and base for the last set of equations.

Problem: 28.0 mL of 0.250 M H_2SO_4 is reacted with 53.0 mL of 0.320 M KOH. Determine the grams of water that forms, and the grams of excess reactant left over.

<u>Titration</u> – lab technique used to determine the concentration of a substance.

<u>Add a titrant</u> (usually from a <u>buret</u>) to the <u>analyte</u> (solution you are testing), measure the volumes of each solution used at the <u>equivalence point</u> (where all reactants are consumed). The equivalence point is signaled by a substance, called an <u>indicator</u>, that changes color when a reactant is used up completely. The endpoint is when the indicator changes color.



Problem: 1.50 grams of KHP, potassium hydrogen phthalate, (molecular weight 204.0 g), a weak, monoprotic acid, is made into 1.00 L of solution. 25.0 mL of this solution is titrated to the endpoint with 18.6 mL NaOH. Determine the molarity of the NaOH

$M_aV_a = M_bV_b$	where M_a =	Molarity of acid
	V_a =	Volume of acid
	M_b =	Molarity of base
	V_b =	Volume of base

$M_aV_a = M_bV_b$ doesn't work for the following acid/base reaction:

$$3Ca(OH)_2 + 2H_3PO_4 \quad 6H_2O + Ca_3(PO_4)_2$$

$$\frac{MaVa}{2} = \frac{MbVb}{3}$$

Divide the acid by the coefficient in front of the acid, and the base by the coefficient in front of the base.

A 0.3518 g sample of waste water containing CCl_4 and benzoic acid was titrated to the endpoint with 10.59 mL of 0.1546 M NaOH. Calculate the percent by mass of benzoic acid ($HC_7H_5O_2$) in the sample.

3) Oxidation / Reduction Reactions – these reactions involve the transfer of electrons from 1 substance to another. (Redox)

Rules for determining oxidation numbers (real or apparent charge of an element):

1) The oxidation number of an atom in a pure element = 0

2) The oxidation number of a monoatomic ion is its charge

3) In a compound,
Fluorine = -1
Alkali Metals = +1
Alkaline Earth metals = +2

4) O in a compound is usually -2, except as a peroxide, O_2^{-2}, which is -1, and in OF_2, which is +2

5) H is +1 when combined with other non-metals

6) The sum of oxidation numbers = the charge of the molecule or ion.

Oxidation – the loss of electrons

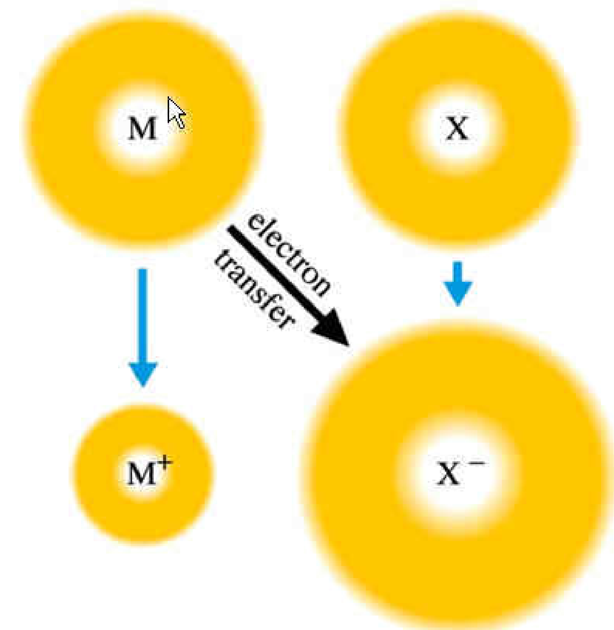
Reduction – the gain of electrons

LeO the Lion goes GeR

OiL RiG

Oxidizing Agent – The reactant that gets reduced, and causes something else to be oxidized.

Reducing Agent – The reactant that gets oxidized, and causes something else to be reduced.



Oxidized	Reduced
loses electrons	*gains* electrons
oxidation state *increases*	oxidation state *decreases*
reducing agent	*oxidizing* agent

Problem: Write an equation for the combustion of methane gas.

Identify the oxidizing agent, and the reducing agent. Which reactant is oxidized, and which is reduced?

Balancing Redox equations by the ½ reaction method in an acidic solution.

$$MnO_4^-{}_{(aq)} + Fe^{+2}{}_{(aq)} \rightarrow Fe^{+3}{}_{(aq)} + Mn^{+2}{}_{(aq)}$$

1) Write oxidation numbers for every element, then write 2 half reactions based on what changes oxidation numbers.
2) Balance all elements except H & O
3) Balance O by adding H_2O
4) Balance H by adding H^+
5) Balance the electrons, so that they cancel
6) Add the 2 half reactions, and cancel species that are common to both sides of the equation.

Balance the same reaction in a basic solution

Add OH^- to neutralize all the H^+, and add the same # of OH^- to the other side of the equation also. (H^+ + OH^- makes water)

$$P_4 + IO_3^- \rightarrow H_2PO_4^- + I^-$$

$$H_2O_2 \rightarrow H_2O + O_2$$

The 2 most useful substances in redox titrations are:

$Cr_2O_7^{-2}{}_{(aq)}$ → $Cr^{+3}{}_{(aq)}$

orange/yellow → green

$MnO_4^{-}{}_{(aq)}$ → $Mn^{+2}{}_{(aq)}$

purple → colorless

They are both reduced, and change colors when all the reactant is consumed.

(act as their own indicator)

Redox titration problem:

A 1.00 gram sample of a plant material, containing oxalic acid, $H_2C_2O_4$ is titrated with 24.0 mL of 0.0100 M $KMnO_4$. Determine the percent by mass of oxalic acid in the plant.