Aqueous Ions and Reactions

(ions, acids, and bases)

<u>Demo</u>
$NaCl(aq) + AgNO_3 (aq) \rightarrow AgCl (s)$
Two clear and colorless solutions turn to a cloudy white when mixed

<u>Demo</u>
Special Light bulb in water can test for complete circuit

			Light
Sugar	$C_{12}H_{22}O_{11}$	Nonelectrolyte	No
Salt	NaCl	Electrolyte	Yes
Vinegar	$C_2H_4O_2$	Weak Electrolyte	Dim
Hydrochloric	HCl	Strong Electrolyte	Bright

<u>Solutions General Considerations</u>

Solution → homogenous mixture of molecules, atoms or ions
Ex: Solid Brass (zinc in copper), gas (air), liquid (tea)

Solvent→ main component is usually liquid

Solute→ other components dissolved in solvent

Miscible→ mix uniformly
Ex: Alcohol and Water

Immiscible→ do not mix
Ex: Oil and Water

Solubility→ the maximum amount of solute that will dissolve in solvent
Soluble
Insoluble → less than 0.1g of solute in 1000g solvent

Dilute ---> concentrated solutions

Saturated Solution is the equilibrium between dissolved and pure
Ex. Sugar <---> Sugar in Solution

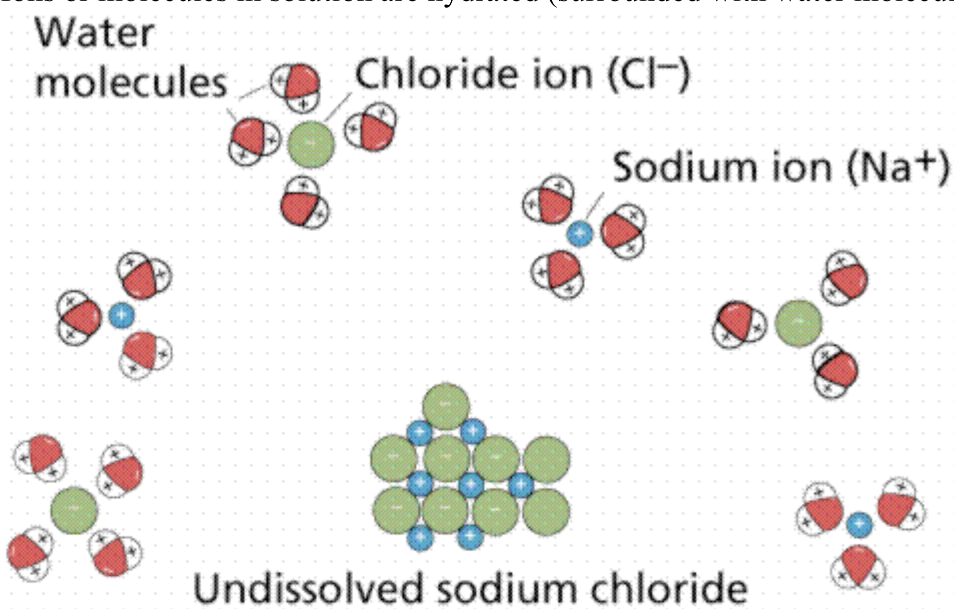
Unsaturated→ lower concentration than saturated
Supersaturated → more concentrated (not stable)

Solutes may be:
Electrolytes (ions)→ conduct electricity
Nonelectrolytes (molecules) → do not conduct
Demo demonstrates these properties

Ex:

NaOH	Na^+ OH^-
HCl	H^+ Cl^-
NaCl	Na^+ Cl^-
Glucose (sugar)	$C_6H_{12}O_6$
Sucrose	$C_{12}H_{22}O_{11}$

Ions or molecules in solution are hydrated (surrounded with water molecules)



(http://www.emc.maricopa.edu/faculty/farabee/BIOBK/BioBookCHEM2.html)

Strong Electrolytes dissociate completely (100%)
Weak Electrolytes partially dissociate

Ex. of Weak Electrolyte
CH_3COOH <----- --> $C_2H_3O_2^- + H^+$
(Acetic Acid)
Percent Dissociation = 1%
Equilibrium lies to the left

Ex. of Strong Electrolyte
$HCl \rightarrow H^+ + Cl^-$
Percent Dissociation = 100%

Shift to the right completely

Other Strong Electrolytes that need to be MEMORIZED
HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$

Ionic Reactions

Molecular: NaCl (aq) + $AgNO_3$ (aq) → AgCl (s) + $NaNO_3$ (aq)

Complete Ionic: Na^+ + Cl^- + Ag^+ + NO_3^- → AgCl (s) + Na^+ + NO_3^-

Net Ionic: Ag^+ (aq) + Cl^- (aq) → AgCl (s) (precipitate)
In a Net Ionic the spectator ions are left out

Think in terms of ions in solution and not just formulas as written

Na^+ (aq) means aqueous or in water
If aq not given Na^+ can still assume in water solution if discussing solutions
but for something like NaCl must specify NaCl(s) or NaCl(aq)

Metathesis → double displacement
Two compounds in solution and exchange of cations

Actual change:
NaCl + $AgNO_3$ → AgCl + $NaNO_3$
Net Ionic Reaction: Ag+ (aq) + Cl- (aq) → AgCl (s)

OR no real change
KCl + $NaNO_3$ → KNO_3 + NaCl
No Reaction just ions in solution
K^+ Cl^- Na+ NO_3^-

Reaction will occur if:
1. Precipitate formed
2. Weak electrolyte formed
3. Gas formed
These will have the effect if remove product the equilibrium shifts to the right and more product is formed (LeChatlier's Principle)

Solubility Rules

(Don't have to learn for my course but know how to use this information)
Salts always soluble: alkali metal, NH_4^+, NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$

Salts soluble with exception: Cl^-, Br^-, I^- except Ag^+, Pb^{2+}, Hg_2^{2+}
SO_4^{2-} except Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

Salts insoluble: metal oxides and hydroxides except Ca^{2+}, Sr^{2+}, Ba^{2+}, alkali (give base)
CO_3^{2-}, PO_4^{3-}, S^{2-}, SO_3^{2-} except NH_4^+, alkali metal

Solubility and Precipitation Reactions

Reaction Occur if:
1. Precipitate Formed

Soluble (dissolve > 10g/L)	NaCl	$NaCl \rightarrow Na^+ + Cl^-$ (lot)
Partially Soluble	$PbCl_2$	$PbCl_2 \longleftrightarrow Pb^{2+} + 2Cl^-$ (little)
Insoluble (dissolve < 1g/L)	AgCl	$AgCl \longleftrightarrow Ag^+ + Cl^-$ (very little)

Soluble will have no precipitate
Precipitate will form in partially soluble or insoluble if ions > 0.1M

In next course can assign numerical values but this is just qualitative

Not expected to memorize the solubility rules
But be able to use and know ion charges

Example:
Given: $Ca(OH)_2$ soluble and $Mg(OH)_2$ insoluble
Ocean Water contains 0.13% by weight Mg
How could you separate Mg from ocean water?

$Ca(OH)_2$ + sea water (Mg^{2+}) $\rightarrow Ca^{2+} + Mg(OH)_2$ (s)
$Ca(OH)_2 (aq) \rightarrow Ca^{2+} + 2OH^-$
$Mg^{2+} + 2OH^- \rightarrow Mg(OH)_2$ (s) Net Reaction

In Freeport Texas there is a large plant that to produce magnesium from ocean water

Naming Oxoacids
In problems look up if precipitate forms
On the exam it will tell you if insoluble or precipitate, otherwise assume soluble

$NaCl + NH_4NO_3 \rightarrow$ No Reaction
THINK IONS

2. Weak Electrolyte Formed

$HC_2H_3O_2 (aq) \longleftrightarrow H^+ (aq) + C_2H_3O_2^- (aq)$
99% acetic acid 1% acetate ion
If produce acetate it will quickly attract H^+

HCl (aq) + $NaC_2H_3O_2$ (aq)
$Na^+ + Cl^-$ = NO
$H^+ + C_2H_3O_2^-$ = YES
Net Ionic: $H^+ + C_2H_3O_2^- \rightarrow HC_2H_3O_2$
Spectator: Na^+ and Cl^-

If it forms water then it is a weak electrolyte
Fe_2O_3 (s) + $6H^+$ (aq) $\rightarrow$ $2Fe^{3+}$ (aq) + $3H_2O$
Metal oxide is soluble in acid and not in water (pure)

<u>3. Gas Formed</u>

KNOW (there are others):

CO_2	carbon dioxide
SO_2	sulfur dioxide
NH_3	ammonia
H_2S	hydrogen sulfide
H_2	hydrogen

Gas Directly
$2HCl$ (aq) + Na_2S (aq) $\rightarrow$
$2H^+ + 2Cl^-\ 2Na^+ + S_2^- \rightarrow H_2S$ (g) + $2NaCl$ (aq)

Gas Indirectly
Vinegar and baking soda (mix in medicine bottle, shake and throw in air)
H^+, $C_2H_3O_2^-$, $NaHCO_3$ (sodium bicarbonate)
$H^+\ Na^+\ HCO_3^- \rightarrow H_2CO_3$ (aq) $Na^+\ C_2H_3O_2^-$ <----> CO_2 (g) + H_2O
The carbonic acid decomposes
? if acetic acid was only 1% would there be enough?
This would drive the reaction to the right and pull H^+ off vinegar

Alka Seltzer, citric acid and sodium bicarbonate fizz in water

Cases in medical literature where eating baking soda the gas blows the stomach apart

Naming Acids
<u>Naming Binary Acids</u> hydro___ic acid (nonmetal root)

Binary Acids Table	**Compound Name**	**Aqueous Solution**
HCl	hydrogen chloride	hydrochloric acid
H_2S	hydrogen sulfide	hydrosulfuric acid
HF	hydrogen fluoride	hydrofluoric acid

use _____ic acid OR ____ous acid
If root name is ate-->ic
ite-->ous

Oxoacids Table	**Compound Name**	**Aqueous Solution**

H_2SO_4	hydrogen sulfate	sulfuric acid
H_2SO_3	hydrogen sulfite	sulfurous acid
HNO3	hydrogen nitrate	nitric acid
HNO_2	hydrogen nitrite	nitrous acid

If prefix then continue to use
Example:
Perchloric acid $HClO_4$
Hypochlorous acid HClO

Naming Acid Salts
Use bi in place of hydrogen

H_2SO_4 hydrogen sulfate = bisulfate
(Ion: SO_4^{2-} sulfate)

$NaHCO_3$ sodium bicarbonate (baking soda)
(Ions: Na^+ HCO_3^-)

Several Hydrogens then indicate
H_3PO_4 -----> phosphoric acid
$H_2PO_4^-$ -----> dihydrogen phosphate
HPO_4^{2-} -----> hydrogen phosphate
PO_4^{3-} -----> phosphate

Acid and Bases in Aqueous Solutions
Application observe in lab: acid turns litmus red
Base turns litmus blue

Acid-Base Concepts	**Acid**	**Base**
Arrhenius	Substance produce H^+ (H_3O^+ Hydronium) in water	Produce OH^- (Hydroxide) in water
Bronsted-Lowry	Donate Proton	Accept Proton
Lewis-	Accept electron pair	Donate electron pair
Solvent System	Give solvent cation NH_4^+	Give solvent anion NH_2^- in liquid ammonia

Common Strong Acids
HCl
HBr
HI
HNO_3
H_2SO_4
$HClO_4$

Common Strong Bases

$NaOH$
KOH
$Ca(OH)_2$
$Ba(OH)_2$

Arrhenius Acids and Bases
Acid increase concentration of hydronium ion ($H3O^+$)
Base increase concentration of hydroxide (OH^-)

Acid: $HCl\ (aq) + H2O \rightarrow H_3O^+\ (aq) + Cl^-$
$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$

Base: $Ca(OH)_2\ (s) \rightarrow Ca^{2+}\ (aq) + 2OH^-\ (aq)$

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

Often:
Acid + Base → Water + Salt
$HCl + NaOH \rightarrow H_2O + NaCl$ (aq) Note: it stays as ions Na^+ and Cl^-

Bronsted-Lowry Acids and Bases
Acid is a substance that can donate a proton
Base is a substance that can accept a proton

Reaction involves the transfer of protons from acid to base

Acid 1 + Base 2 <-----> Acid 2 + Base 1

$HC_2H_3O_2 + H_2O \leftrightarrow H_3O^+ + C_2H_3O_2^-$
Conjugate Acid-Base pair $HC_2H_3O_2$ and $C_2H_3O_2^-$, H_2O and H_3O^+
Equilibrium lies more to left so H_3O^+ is a stronger acid than acetic acid

Water can act as an acid or base

Acid 1		Base 2		Acid 2	Base 1
H_2O	+	NH_3	<--- -->	NH_4^+	OH^-

NH_4 is a stronger acid than H_2O
OH^- is a stronger base than NH_3

When an acid gives up a proton it forms a base
When base accepts proton it forms an acid
$HA + H_2O \leftrightarrow H_3O^+ + A^-$
acid base acid base

HA and A^- are conjugate acid-base pair

H_3O^+ and H_2O are conjugate acid-base pair

Monoprotic: means donate one proton (ex. HCl and $HC_2H_3O_2$)

Polyprotic Acids: can donate more than one proton
Examples: H_2SO_4 hydrogen sulfate is diprotic (sulfuric acid)
H_3PO_4 hydrogen phosphate is triprotic (phosphoric acid)

$H_2SO_4 + H_2O \rightarrow H_3O_+ + HSO_4^-$
$HSO_4^- + H_2O$ <----> $H_3O^+ + SO_4^{2-}$
Acid 1 Base 2 Acid 2 Base1

Amphiprotic: ion or molecule that can accept or donate a proton
In the example above HSO_4^- is amphiprotic

Types of Reactions
Chemical reaction – atoms are rearranged
(may or may not be obvious by simple observation)

Traditional Scheme:

Combination $2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$

Decomposition $2KClO_3\ (s) \rightarrow$ (heat and MnO_2 catalyst) $2KCl\ (s) + 3O_2(g)$

Displacement $Zn\ (s) + CuCl_2\ (aq) \rightarrow Cu + ZnCl_2\ (aq)$
$Zn\ (s) + 2HCl\ (aq) \rightarrow H_2\ (aq) + ZnCl_2\ (aq)$

Metathesis $AgNO_3\ (aq) + NaCl\ (aq) \rightarrow NaNO_3\ (aq) + AgCl\ (s)$
(Double Displacement)

Combustion $4\ Fe\ (s) + 3O_2\ (g) \rightarrow 2FeO_3\ (s)$
$C\ (s) + O_2\ (g) \rightarrow CO_2(g)$

Other Schemes: Redox and Other (such as acid-base)

Reaction in Solution

	Acid +	Base →	Salt +	Water
Conc.	0.15M	0.75M		
	H_2SO_4 (aq) + 2NaOH (aq) → Na_2SO_4 (aq) + $2H_2O$			
Sulfuric Acid		Sodium Hydroxide	Sodium Sulfate	Water

How many ml of NaOH to react completely with 50 ml of H_2SO_4?
?mL NaOH = (0.050L H_2SO_4)(0.15 mol H_2SO_4 / L)(2 mol NaOH/ 1 mol H2SO4)(1L/0.75 mol NaOH) = 0.020 L = 20 ml

H_2SO_4 = 0.0075 mol 50 ml 0.15M → 7.5 x 10^{-3}
NaOH = 0.0150 mol 25 ml 0.75M → 1.5 x 10^{-3}
Note: that this gives the needed 1:2 ratio
Need twice as much NaOH

Summary: vol A –(conc)--> mol A –(Balanced Eq.)--> mol B –($conc^{-1}$)--> vol B

Example:
How many grams $BaSO_4$ formed?
Ba^{2+} (aq) + SO_4^{2-} (aq) → $BaSO_4$ (s)
First you have to find how many moles formed?

20.0 ml of 0.150M $Al_2(SO_4)_3$
(3(SO_4)/ 1 $Al_2(SO_4)_3$)(20 x 10^{-3} L)(0.150 mol $Al_2(SO_4)_3$/ L) = 9.00 x 10^{-3} mol SO_4^{2-}

30.0 ml of 0.200M $BaCl_2$
(1 Ba/ 1 $BaCl_2$)(30 x 10^{-3} L)(0.200 mol $BaCl_2$/L) = 6.00 x 10^{-3} mol Ba^{2+}

Ba^{2+} and SO_4^{2-} combine 1:1 so excess of SO_4^{2-} which mean Ba^{2+} limiting reactant

6.00 x 10^{-3} mol $BaSO_4$ (233.4g $BaSO_4$/ 1mol $BaSO_4$) = 1.40g

List ions and possible combination (any insoluble? See rules)

Cations Anions
Al^{3+} and SO_4^{2-} is soluble

Ba^{2+} and Cl^- is soluble

Al^{3+} and Cl^- is soluble

Ba^{2+} and SO_4^{2-} is insoluble

Stoichiometry of Ionic Reactions
Various ways of expressing concentration
For example Weight percent of sulfuric acid 96% (w/w) H_2SO_4 or 96% H_2SO_4
Means 96% H_2SO_4 and 4g of H_2O in 100g of total acid solution

But for most uses molarity M (mol solute/ liter solution) will be used

Know how many ions in solution
0.2M $CaCl_2$ (aq)   0.2M Ca^{2+}(aq) and 0.4M Cl^-(aq)
since $CaCl_2 \rightarrow Ca^{2+} + 2Cl^-$

Stoichiometry is just like earlier calculations

Remember:
- (g) gas
- (s) solid
- (l) liquid
- (aq) aqueous in solution

Oxidation Numbers and Redox Reactions (covered in "Redox" material)