"A" students work (without solutions manual) ~ 10 problems/night.

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Reactions in Aqueous Solutions

Water – the stuff of life and of the earth.

Reactions – carried out in water, either as a bucket or glass of water or as adsorbed water onto surfaces.

Types of common reactions

- precipitation
- acid-base
- oxidation-reduction reactions

All of above reactions involve **IONS** in solution and are affected by

1. Charge of ions
2. Size of ions
3. Number of ions/volume

$$mass \rightarrow particles \rightarrow \frac{particles}{volume}$$

$$g \rightarrow moles \rightarrow molarity$$

$$molarity(M) \equiv \frac{moles\ of\ \ solute}{liters\ of\ \ solution}$$

$$[A] = \frac{moles\ \ A}{liters\ of\ \ solution}$$

Example: Solve either stepwise or chainwise. **Stepwise**

If a bottle is labeled as concentrated hydrochloric acid (12.0 M) how many Moles are there in 50.0 mL?

Concentrated
[HCl] = 12.0M
Moles in 50.0 mL

$$[A] = \frac{moles\ \ A}{liters\ of\ \ solution}$$

$$[HCl] = \frac{12.0molesHCl}{L}$$

$$(50mL)\left[\frac{L}{10^3 mL}\right] = 0.0500L$$

$$\frac{12.0molesHCl}{L} = \frac{xmolesHCl}{0.0500L}$$

$$x = \left[\frac{12.0molesHCl}{L}\right]0.0500L$$

$$x = 0.600molesHCl$$

Sig Figs

Concentrated – **relatively** large (>?>6M)

Dilute – **relatively** small (<?<0.1 M)

Example: **Chainwise**

If a bottle is labeled as concentrated hydrochloric acid (12.0 M) how many Moles are there in 50.0 mL?

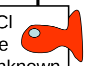
[HCl] = 12.0M= known
Moles in 50.0 mL= unknown

$$(50.0mL)\left[\frac{1L}{10^3 mL}\right]\frac{12.0molesHCl}{L} = x$$

$$x = 0.600molesHCl$$

Suppose you need to react 0.25 mol of HCl with a second substance. What volume of 12.0 M HCl would you use?

Need 0.25 mol HCl
Second substance
Vol 12.0 M HCl unknown



Stepwise solution

$$\left(\frac{12.0molesHCl}{1L}\right) = \frac{0.25molHCl}{xL}$$

$$xL = \frac{0.25molHCl}{\left(\frac{12.0molesHCl}{1L}\right)}$$

$$xL = \frac{0.25}{12.0}L$$

$$xL = 2.083x10^{-2}L$$

$$x\ mL = 2.083x10^{-2}L\left[\frac{10^3 mL}{L}\right]$$

$$x\ mL = 20.83mL = 21mL$$

Sig figs

Chainwise solution

$$0.25molHCl\left[\frac{1L}{12moleHCl}\right]\left[\frac{10^3 mL}{L}\right] = 20.8333mL$$

$$20.8333mL = 21mL$$

Most ionic compounds dissolve **completely** in water to the individual ions.

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

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

$$PbCl_{2(s)} \rightarrow PbCl_{2(aq)} \xrightarrow{completely} Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$

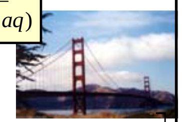
$$PbCl_{2(s)} \xrightarrow{completely} Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$

OJO!!!! – VERY IMPORTANT – Must conserve mass
- Equation must balance

Example. If you have 3.0 M $PbCl_2$ what concentration of Cl^- ions do you have?

$$PbCl_{2(s)} \xrightarrow{completely} Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$

Treat as conversion factors



$$\left[\frac{3.0molesPbCl_2}{1L}\right]\left[\frac{2molesCl^-}{1molePbCl_2}\right] = \left[\frac{6.0molesCl^-}{1L}\right]$$

$$\left[Cl^-\right] = 2\left[PbCl_2\right]$$

The concentration of Cl^- ion is double that of $PbCl_2$

$$\left[Cl^-\right] = 2(3.0) = 6.0M$$

Another example:
Write the chemical equation for the solution of ammonium phosphate

We memorized these ions NH_4^{+} PO_4^{3-}

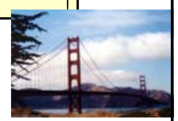
Molecular charge neutrality requires $(NH_4)_3PO_{4(s)}$

$$(NH_4)_3PO_{4(s)} \xrightarrow{completely} 3NH_{4(aq)}^{+} + PO_{4(aq)}^{3-}$$

stoichiometry 1: 3: 1

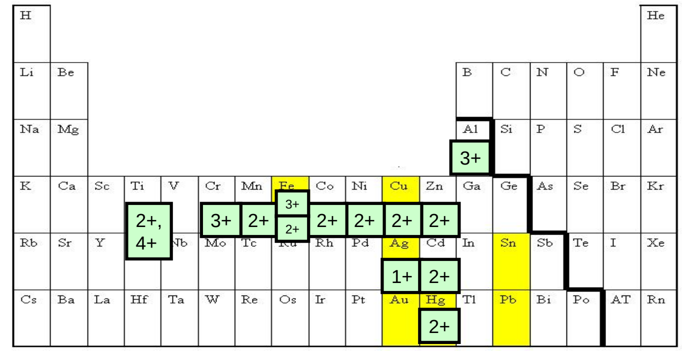
$$\left[\frac{molesPO_{4(s)}}{L}\right]\left[\frac{3molesNH_{4(aq)}^{+}}{1molePO_{4(s)}}\right] = \left[\frac{molesNH_{4(aq)}^{+}}{L}\right]$$

$$\left[PO_4^{3-}\right]3 = \left[NH_4^{+}\right]$$



What is the concentration of ammonium in solution compared to phosphate?

These are the most common charges of Transition metal ions and aluminum



Example Give concentration in molarity of each ion in 0.080 M zinc phosphate

$$Zn_3(PO_4)_{2(s)} \rightarrow 3Zn_{(aq)}^{2+} + 2PO_{4(aq)}^{3-}$$

Stoichiometry 1: 3: 21

$$\left[\frac{0.080molesZn_3(PO_4)_2}{L}\right]\left[\frac{3molesZn^{2+}}{1moleZn_3(PO_4)_2}\right] = 0.24molesZn^{2+}$$

$$\left[\frac{0.080molesZn_3(PO_4)_2}{L}\right]\left[\frac{2molesPO_4^{2-}}{1moleZn_3(PO_4)_2}\right] = \frac{0.16molesPO_4^{2-}}{L} = 0.16MPO_4^{3-}$$

Water – the stuff of life and of the earth.

Reactions – carried out in water, either as a bucket or glass of water or as adsorbed water onto surfaces.

Types of common reactions

- **precipitation**
- acid-base
- oxidation-reduction reactions

Which Rules Apply?

FITCH Rules

General
G1: Suzuki is Success
G2. Slow me down
G3. Scientific Knowledge is Referential
G4. Watch out for Red Herrings
G5. Chemists are Lazy

Chemistry
C1. It's all about charge
C2. Everybody wants to "be like Mike"
C3. Size Matters

$$E_{el} = k\left(\frac{q_1 q_2}{r_1 + r_2}\right)$$

C4. Still Waters Run Deep
C5. Alpha Dogs eat first

Who precipitates and who stays soluble?
Depends upon who reacts with whom?

Low Charge Density
Intermediate Charge Density
High Charge Density

	No NO_3^-	Clean Cl^-	Socks SO_4^{2-}	Oh OH^-	Card me CO_3^{2-}	PleeeeaSe!! PO_4^{3-}	S^{2-}
Group 1 cations (1+)	Weak Electrostatic Interaction = Soluble						
Group 2 cations (2+)			$BaSO_4$	$Mg(OH)_2$			
Transition metal cations (usually 2+)		AgCl			Strong Electrostatic Interaction results in precipitation		

Extremes
1. Low -Low charge density ion interactions - weak electrostatic energy **Stay soluble**
2. High -High charge density ion interactions - strong electrostatic energy **Precipitate**
3. High- Intermediate charge density ion interactions – generally strong electrostatic energy – **precipitate**

In Between
1. Low – High charge density ion interactions – generally weak electrostatic energy: **Soluble with exceptions**
2. Low – Intermediate charge density ion interactions – generally weak electrostatic energy: **Soluble with exceptions**
3. Intermediate – Intermediate – charge density ion interactions – generally weak electrostatic energy: **soluble with exceptions**

Example
What happens when a solution of $CuSO_4$ is mixed with a solution of $NaNO_3$?
What happens when a solution of Na_2CO_3 is mixed with a solution of $CaCl_2$?

First assume everbody goes into solution or already in solution
Identify who is **charge dense** and who is **not charge dense**
High/high always precipitates
Low/low never precipitates
Low/High generally does not precipitate

No Clean Socks

$$CuSO_{4(aq)} + NaNO_{3(aq)} \rightarrow Cu^{2+}_{aq} + SO^{2-}_{4aq} + Na^{+}_{aq} + NO^{-}_{3aq}$$

Transition Metal ion — **Group 1** — **No Clean Socks**

no reaction

Example.
What happens when a solution of $CuSO_4$ is mixed with a solution of $NaNO_3$?
What happens when a solution of Na_2CO_3 is mixed with a solution of $CaCl_2$?

Oh Card me PleaSe — Rx

$$Na_2CO_{3(aq)} + CaCl_{2(aq)} \rightarrow 2Na^{+}_{aq} + CO^{2-}_{3aq} + Ca^{2+}_{aq} + 2Cl^{-}_{aq}$$

Group 1 — Group 2 — Group 17

$$2Na^{+}_{aq} + 2Cl^{-}_{aq} + CO^{2-}{}_{3(aq)} + Ca^{2+}_{aq} \rightarrow CaCO^{2-}_{3s} + 2Na^{+}_{aq} + 2Cl^{-}_{aq}$$

FITCH Rules

General

G1: Suzuki is Success

G2. Slow me down

G3. Scientific Knowledge is Referential

G4. Watch out for Red Herrings

G5. Chemists are Lazy

Chemistry

C1. It's all about charge

C2. Everybody wants to "be like Mike"

C3. Size Matters

C4. Still Waters Run Deep

C5. Alpha Dogs eat first

$$E_{el} = k\left(\frac{q_1 q_2}{r_1 + r_2}\right)$$

Example.

What happens when a solution of $CuSO_4$ is mixed with a solution of $NaNO_3$?

What happens when a solution of Na_2CO_3 is mixed with a solution of $CaCl_2$?

Oh Card me PleaSe

$$Na_2CO_{3(aq)} + CaCl_{2(aq)} \rightarrow 2Na^+_{aq} + CO^{2-}_{3aq} + Ca^{2+}_{aq} + 2Cl^-_{aq}$$

Group 1 | Group 2 | Group 17

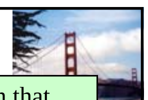
$$2Na^+_{aq} + 2Cl^-_{aq} + CO^{2-}{}_{3(aq)} + Ca^{2+}_{aq} \rightarrow CaCO^{2-}_{3s} + 2Na^+_{aq} + 2Cl^-_{aq}$$

Do not write redundant (on both sides) ions

Write NET ionic equation

$$CO^{2-}{}_{3(aq)} + Ca^{2+}_{aq} \rightarrow CaCO^{2-}_{3s}$$

Stoichiometry is important in these reactions



Example : Consider the net ionic equation for the reaction that occurs when solutions of $Cu(NO_3)_2$ and NaOH are mixed. What volume of 0.106 M $Cu(NO_3)_2$ solution is required to form 6.52 g of solid $Cu(OH)_2$?

1. Check for charge dense ions that can precipitate
2. Write a net ionic reaction which excludes spectators (low charge dense ions)
3. Count total moles of all precipitating reactants

$Cu^{2+}_{(aq)}$ Charge Dense Transition metal

$$Cu^{2+}_{(aq)} + OH^-_{(aq)} + \rightarrow CuOH_{x(s)}$$

$OH^-_{(aq)}$ Oh, Card me Please – charge dense

$$Cu^{2+}_{(aq)} + 2OH^-_{(aq)} + \rightarrow Cu(OH)_{2(s)}$$

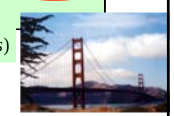
$NO^-_{3(aq)}$ No Clean Socks – not charge dense

$Na^+_{(aq)}$ Group – not charge dense

Stoichiometry is important in these reactions

Example : Consider the net ionic equation for the reaction that occurs when solutions of $Cu(NO_3)_2$ and NaOH are mixed. What volume of 0.106 M $Cu(NO_3)_2$ solution is required to form 6.52 g of solid $Cu(OH)_2$?

$$Cu^{2+}_{(aq)} + 2OH^-_{(aq)} + \rightarrow Cu(OH)_{2(s)}$$



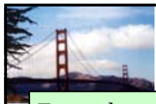
1: 2: 1

$$\left[6.52\,gCu(OH)_2\right]\left[\frac{1moleCu(OH)_2}{97.57\,gCu(OH)_2}\right] = 6.6823x10^{-2}\,moleCu(OH)_2$$

$$6.6823x10^{-2}\,moleCu(OH)_2\left[\frac{moleCu}{moleCu(OH)_2}\right]\left[\frac{moleCu(NO_3)_2}{moleCu}\right] = 6.6823x10^{-2}\,moleCu(NO_3)_2$$

$$6.6823x10^{-2}\,moleCu(NO_3)_2\left[\frac{1LCu(NO_3)_2}{0.106moleCu(NO_3)_2}\right] = 6.304133849x10^{-1}\,L$$

$$? = 0.630L$$



Stoichiometry is important in these reactions

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$$Cu^{2+}_{(aq)} + 2OH^{-}_{(aq)} + \rightarrow Cu(OH)_{2(s)}$$

Chainwise solution

$$\left[6.52\,gCu(OH)_2\right]\left[\frac{1moleCu(OH)_2}{97.57\,gCu(OH)_2}\right]\left[\frac{moleCu}{moleCu(OH)_2}\right]\left[\frac{moleCu(NO_3)_2}{moleCu}\right]\left[\frac{1LCu(NO_3)_2}{0.106moleCu(NO_3)_2}\right] = ?$$

$$? = 6.304133849x10^{1}\,L$$

$$? = 0.630L$$

Stoichiometry is important in these reactions

Example: When aqueous solutions of sodium hydroxide and iron(III) nitrate are mixed, a red gelatinous precipitate forms. Calculate the mass of precipitate formed when 50.00 mL of 0.200 M NaOH and 30.00 mL of 0.125 M $Fe(NO_3)_3$ are mixed

1. Check for charge dense ions that can precipitate
2. Write a net ionic reaction which excludes spectators (low charge dense ions)
3. Count total moles of all precipitating reactants

Fe(III)
OH-

$$Fe^{3+}_{(aq)} + OH^{-}_{(aq)} + \rightarrow Fe(OH)_{x(s)}$$

$$Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)} + \rightarrow Fe(OH)_{3(s)}$$

Charge balance both sides of equation

$$\left[50mLNaOH\right]\left[\frac{1L}{10^3mL}\right]\left[\frac{0.200moleNaOH}{LNaOH}\right] = 0.0100moleNaOH$$

$$\left[30.00mLFe(NO_3)_3\right]\left[\frac{1L}{10^3mL}\right]\left[\frac{0.125moleFe(NO_3)_3}{LFe(NO_3)_3}\right] = 0.00375moleFe(NO_3)_3$$

Stoichiometry is important in these reactions

Example: When aqueous solutions of sodium hydroxide and iron(III) nitrate are mixed, a red gelatinous precipitate forms. Calculate the mass of precipitate formed when 50.00 mL of 0.200 M NaOH and 30.00 mL of 0.125 M $Fe(NO_3)_3$ are mixed

	Moles present	could make moles product	
Fe^{3+}	0.00375	0.00375	Smallest # is answer
OH^-	0.0100	0.0100/3 = 0.00333	

$$\left[0.00333moleFe^{3+}\right]\left[\frac{1moleFe(OH)_3}{1moleFe^{3+}}\right]\left[\frac{106.87\,gFe(OH)_3}{1moleFe(OH)_3}\right] = 0.356gFe(OH)_3$$

1: 3: 1

$$Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)} + \rightarrow Fe(OH)_{3(s)}$$

Water – the stuff of life and of the earth.

Reactions – carried out in water, either as a bucket or glass of water or as adsorbed water onto surfaces.

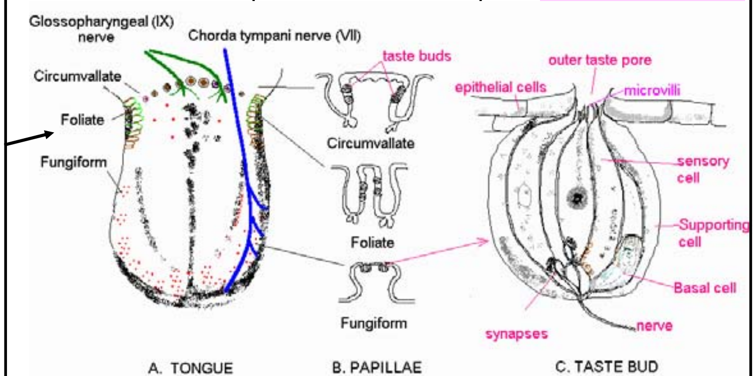
Types of common reactions

- precipitation
- acid-base
- oxidation-reduction reactions

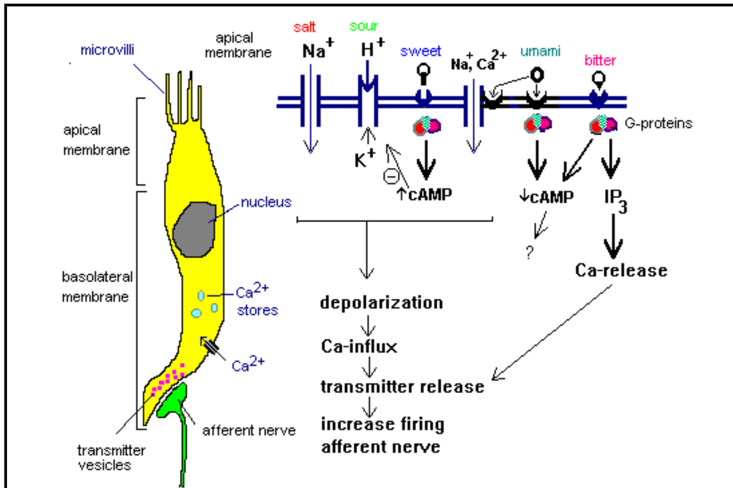
Predicting **who gets the proton** (acid-base) depends on most charge dense of competing reactions

Acid-Base Reactions

Acid is a sour taste



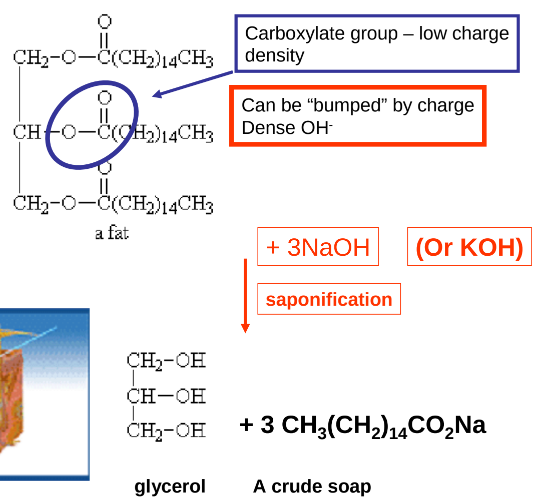
Foliate papillae are situated on the edge of the tongue slightly anterior of the circumvallate line. They are predominantly sensitive to sour tastes. Innervated by the glossopharyngeal (IXth cranial) nerve. On average 5.4 foliate papillae per side of the tongue, 117 taste buds per foliate papillae, total = 1280 foliate taste buds per tongue.

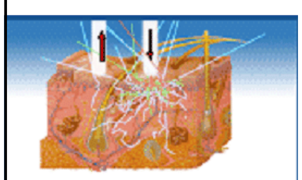


Fatty acids and oils fill your skin.

Bases dissolve

These compounds making soaps (feel slippery)





In 1849 a Harvard chemistry professor, John Webster argued with a colleague, Dr. Parkman, about money he owed the colleague for some dinosaur bones. He killed his fellow Professor and attempted to hide the crime by dissolving the body in a vat Of "lime" ($Ca(OH)_2$) and then burning the remaining bones.

Chemical Definition of Acid/Base

Svante Arrhenius

An acid is a species that produces proton ions in solution

Solution

Acid: HX

Strong acid

Weak acid

X^- is an anion who is partnered with the cation, H^+

$$HX_{aq} \rightarrow H^+_{aq} + X^-_{aq}$$

In a **STRONG** acid the anion can not provide enough electrostatic attraction to hold onto the H^+ (proton)

In a **WEAK** acid the anion is very charge dense and holds onto the H^+

Who gives up and who holds onto a proton?

Low Charge Density
Intermediate Charge Density
High Charge Density

No	Clean	Socks	Oh	Card me	PleeeeaSe!!
NO_3^-	Cl^-	SO_4^{2-}	OH-	CO_3^{2-} PO_4^{3-}	S^{2-}

H+

Low Charge Density Anions Give up protons | H_2O | **High Charge Density Anions hold onto protons**

STRONG acids

Weak to Intermediate

WEAK acids

$$HX_{aq} \rightarrow H^+_{aq} + X^-_{aq}$$

Completely falls apart

$$HX_{aq} \rightleftarrows H^+_{aq} + X^-_{aq}$$

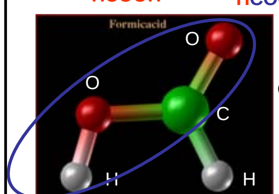
X- tries to hold onto some protons
Double arrow

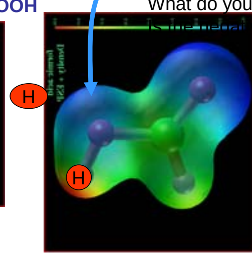
Another very common type of Weak Acid - ….COOH

Condensed Structural Formula gives a clue as to the Important Functional Group

HCOOH

HCOOH





Blue represents negative charge causes potential field

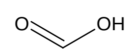
What do you see?
Is the negative charge fully localized?

Dissociate temporarily

Proton actually resides
Somewhere between
Two highly negatively
Charged oxygen

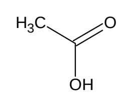
Some Weak Acids with COO- functional group **(Carboxylate)**

Formic Acid
(Ants)



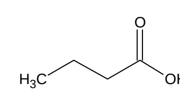
HCOOH

Acetic Acid
Vinegar
HOAc (OAc = acetate)

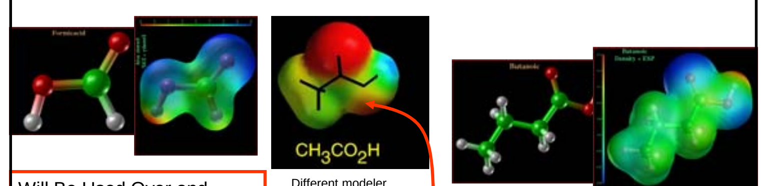


CH_3COOH

Butanoic Acid



$CH_3CH_2CH_2COOH$



Will Be Used Over and Over In Example Problems

Different modeler
Gave opposite color scheme (Red = large neg charge)

http://academic.reed.edu/chemistry/alan/ED/JCE/figlist.html

Chemical Definition of Acid/Base

Svante Arrhenius

An acid is a species that produces proton ions in solution
A base is a species that produces hydroxide ions in solution

$$MOH \rightarrow OH^-_{aq} + M^+_{aq}$$

- Solution
  - Acid: HX
    - Strong acid
    - Weak acid
  - Base:
    - Strong base
    - Weak base

Where M⁺ is some cation

Who gives up and who holds onto a hydroxide?

Low Charge Density
Intermediate Charge Density
High Charge Density

	No	Clean	Socks	Oh	Card me		PleeeeaSe!!
	NO_3^-	Cl^-	SO_4^{2-}	OH-	CO_3^{2-}	PO_4^{3-}	S^{2-}
H+	STRONG acids				WEAK acids		
Group 1 cations (1+)				Strong Bases			
Group 2 cations (2+)							

Weak bases are produced by an alternative manner

$$R + H_2O \rightarrow RH^+_{aq} + OH^-_{aq}$$

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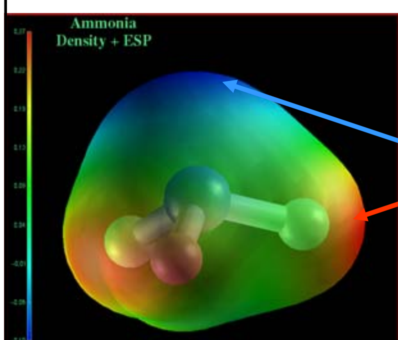
http://core.ecu.edu/phys/flurchickk/AtomicMolecularSystems/molecularStructures/molecularStructures.html



net neutral Molecule ammonia, NH_3

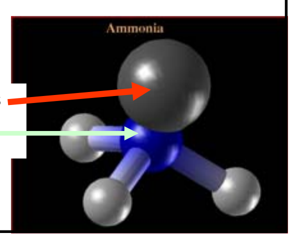
Is the charge evenly distributed?

Is there a region of high charge Density?

In this model blue is negative charge causing a potential field and red is Positive charge opposite potential field

Nitrogen typically contains a set of unbonded electrons

which give Large charge density near the N atom



$$R + H_2O \rightarrow RH^+_{aq} + OH^-_{aq}$$

Write reaction to show the unbonded electrons on R (where R is N in ammonia)

$$H_2O + :R \rightarrow OH^-_{aq} + RH^+_{aq}$$

Write the reaction to show that electrons on R are attracted to proton on water

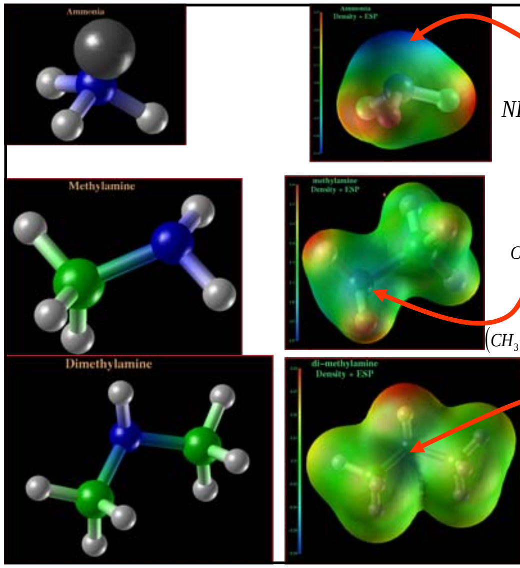
$$HO\cdot\cdot H + :R \rightarrow OH^-_{aq} + RH^+_{aq}$$

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

Compounds in which H on NH_3 has been replaced by –C groups react Similarly

- NH_3
- CH_3NH_2
- $C_2H_5NH_2$
- $(C_2H_5)_2NH$
- $C_6H_5NH_2$

} R



NH_3

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

CH_3NH_2

$CH_3NH_2 + H_2O \rightleftarrows CH_3NH_3^+ + OH^-$

$(CH_3)NH + H_2O \rightleftarrows (CH_3)_2NH_2^+ + OH^-$

$(CH_3)_2NH$

Equations for ACID/BASE Reactions

In order to KNOW How **MUCH** acid or base is present we often react acids and bases together

Three types of reactions:	Three types of reactions:
Strong Acid + Strong Base	SA/SB
Strong Acid + Weak Base	SA/WB
Weak Acid + Strong Base	WA/SB

Chemists Are Lazy Notation

Strong Acid Strong Base Reaction

SA/SB

SA $HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO^-_{3(aq)}$

SB $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

Do not react
(Low Charge Density)
Spectators
Reactants are H+ and OH-

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O$ Net ionic equation SA/SB

Why is there a single arrow?

We are indicating that the protons and hydroxides we added as HNO_3 and NaOH Are not going back to NO_3^- and Na^+. Proton and hydroxide **must** find each other and react

Weak Acid Strong Base Reaction

WA/SB

WA $HX_{aq} \rightleftarrows H^+_{aq} + X^-_{aq}$

Remember: double arrow means X- is charge dense and tries to hold H+

SB $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

Spectator

Add These 2 rx $H^+_{aq} + OH^-_{aq} \rightarrow H_2O$

$HX_{aq} + OH^-_{aq} \rightarrow H_2O + X^-_{aq}$ **Net Ionic equation SB/WA**

$HF_{aq} + OH^-_{aq} \rightarrow H_2O + F^-_{aq}$ example

Strong Acid Weak Base Reaction

WA/SB

Remember: double arrow means an incomplete reaction

WB $R_{aq} + H_2O \rightleftarrows OH^-_{aq} + RH^+_{aq}$

SA $HX \rightarrow H^+_{aq} + X^-_{aq}$ Spectator

add $H^+_{aq} + OH^-_{aq} \rightarrow H_2O$

$$R_{aq} + H^+_{aq} \rightarrow RH^+_{aq}$$ **Net Ionic equation SB/WA**

$$NH_{3aq} + H^+_{aq} \rightarrow NH^+_{4aq}$$ example

Table

Reactants		Reacting Species	Net Ionic Eq
Strong Acid Strong Base	SA/SB	H^+, OH^-	$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O$
Weak Acid Strong Base	WA/SB	HX, OH	$HX_{(aq)} + OH^-_{(aq)} \rightarrow X^-_{(aq)} + H_2O$
Strong Acid Weak Base	SA/WB	H^+, R	$R_{(aq)} + H^+_{(aq)} \rightarrow HR^+_{(aq)}$

Example: Write a net ionic equation for each of the following reactions in dilute water solution

Nitrous acid with sodium hydroxide

Ous = least number of oxygen

HNO_3 is nitric acid so nitrous acid would be HNO_2

NO_2^- vs NO_3^-

↑ Low charge density anion

Higher charge density anion

Probability: wants to hold onto H^+

Probability: **weak acid**

NaOH; Na^+ is group 1

Low charge density cation

Na+ can not hold onto OH-

Strong base

Reaction is WA/SB

$$HX_{(aq)} + OH^-_{(aq)} \rightarrow X^-_{(aq)} + H_2O$$

$$HNO_{2(aq)} + OH^-_{(aq)} \rightarrow NO^-_{2(aq)} + H_2O$$

Example: Write a net ionic equation for each of the following reactions in dilute water solution

Ethylamine ($CH_3CH_2NH_2$) with perchloric acid $HClO_4$

$CH_3CH_2NH_2$ is an amine

N on amine has local negative charge

N reacts with water to steal a proton and leave a hydroxide

Weak Base

No Clean Socks

Cl^- and ClO_4^-

perchlorate ion is 5 atom with single charge, electrons spread out over oxygen implies it is low charge dense, can't hold proton.

Strong acid

Reaction is SA/WB

$$R_{(aq)} + H^+_{(aq)} \rightarrow HR^+_{(aq)}$$

$$CH_3CH_2NH_2 + H^+_{(aq)} \rightarrow CH_3CH_2NH^+_{3(aq)}$$

Example : Write a net ionic equation for each of the following reactions in dilute water solution

Hydrobromic acid (HBr) with potassium hydroxide

Br is group 17

Group 17 are singly charged

Br is one charge/1 atom

HBr = strong acid

K is group 1

Group 1 are singly charged

K is one charge/1 atom

KOH= strong base

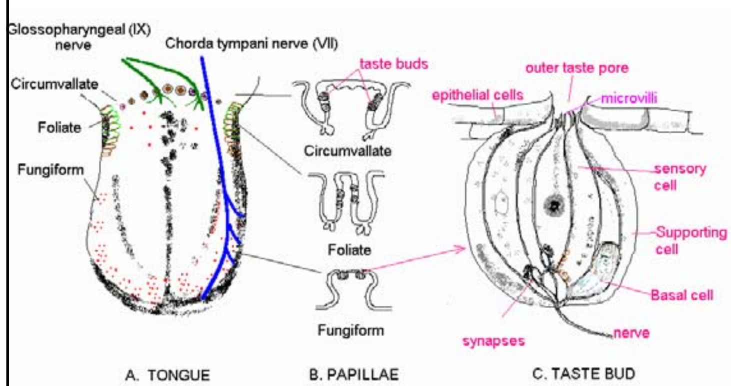
Reaction is SA/SB

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O$$

Knowing HOW MUCH stuff is present

How acidic is vinegar?

Is tasting the way to go?



At Equivalence Point

Types of problems:

Solve for volume
Solve for Molarity
Solve for Mass
Solve for moles

$[A] = \frac{n_A}{V_A}$ $\quad m_A = n_A MM_A$

$[A] = \frac{n_A}{V_A}$

molarity ⟷ Moles of acid ⟷ mass ($m_A = n_A MM_A$)

↕

Moles protons

↕

Moles OH

↕

Volume ⟷ Moles of Base ⟷ moles

$V_B = \frac{n_B}{[B]}$

$m_B\left(\frac{1}{MM_B}\right) = n_B$

$n_{H,eq\ pt} = n_{OH,eq\ pt}$

Example : In a titration it is found that 25.0 mL of 0.500 M NaOH is required to react with a 15.0 mL sample of vinegar. What is the molarity of acetic acid in the sample?

25.0 mL 0.500 M NaOH
Titrate
Vinegar = Acetic Acid
15.0 mL Acetic Acid
Molarity?

$n_{H,eq\ pt} = n_{OH,eq\ pt}$

Na+ - weak charge density = Strong base

Acetic Acid CH_3COOH = weak acid

$$HX_{(aq)} + OH^-_{(aq)} \rightarrow X^-_{(aq)} + H_2O \qquad \frac{0.8333 mol}{L} = 0.833 M$$

$$H[CH_3COO]_{aq} + OH^-_{(aq)} \rightarrow CH_3COO^-_{(aq)} + H_2O$$

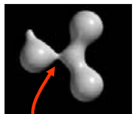
1: 1: 1: 1

$$\frac{(25 mL NaOH)\left(\frac{1 L}{10^3 mL}\right)\left(\frac{0.500 mole NaOH}{1 L NaOH}\right)\left(\frac{H[CH_3COO] equivalence point mole}{NaOH equivalence point mole}\right)}{0.015 L H[CH_3COO][} = \frac{1.52x10^{-3} mol H[CH_3COO]}{0.015 L}$$

Example : The principal ingredient of certain commercial antacids is calcium carbonate, $CaCO_3$. A student titrates an antacid tablet weighting 0.542 g with hydrochloric acid; the reaction is

$$CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O$$

If 38.5 mL of 0.200 M HCl is required for complete reaction, what is the percentage of $CaCO_3$ in the antacid table?



$$CO^{2-}_{3aq} + H^+_{(aq)} \rightarrow HCO^{-1}_{3aq}$$

$$H_2O \xleftarrow{H^+} OH^-_{aq} + CO_{2g}$$

Before we start
1. Problem is slightly unusual because of gas production

Notice the "pinch point"
Where there is little electron density = Weak bond
What do you predict happens?

Aqueous reactions with CO_3^{2-}
In presence of protons result in gas

http://academic.reed.edu/chemistry/alan/ED/JCE/figlist.html

Example : The principal ingredient of certain commercial antacids is calcium carbonate, $CaCO_3$. A student titrates an antacid tablet weighting 0.542 g with hydrochloric acid; the reaction is

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If 38.5 mL of 0.200 M HCl is required for complete reaction, what is the percentage of $CaCO_3$ in the antacid table?

Want: $\left[\frac{gCaCO_{3(s)}}{gtablet}\right]100$ $\left[\frac{0.3853gCaCO_{3(s)}}{0.542gtablet}\right]100 = 71.097\%$

71.1%

Need g $CaCO_3$ → Need Moles $CaCO_3$ → Moles HCl equivalence pt

$$[38.5mLHCl]\left[\frac{1L}{10^3mL}\right]\left[\frac{0.200molesHCl}{1LHCl}\right]\left[\frac{1moleH^+}{moleHCl}\right]\left[\frac{1moleCaCO_3}{2moleH^+}\right]\left[\frac{100.09gCaCO_3}{moleCaCO_3}\right] = gCaCO_3$$

Chapter 4: Reactions in Aqueous Solutions

Water – the stuff of life and of the earth.

Reactions – carried out in water, either as or as adsorbed water onto surfac

Types of common reactions

precipitation
acid-base
oxidation-reduction reactions

Why you must spend taxes to maintain not to build

FITCH Rules

General
G1: Suzuki is Success
G2. Slow me down
G3. Scientific Knowledge is Referential
G4. Watch out for Red Herrings
G5. Chemists are Lazy

Chemistry
C1. It's all about charge
C2. Everybody wants to "be like Mike"
C3. Size Matters
C4. Still Waters Run Deep
C5. Alpha Dogs eat first

$$E_{el} = k\left(\frac{q_1q_2}{r_1 + r_2}\right)$$

Net ionic reaction (excludes spectators)

$$Zn_{(s)} + 2H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

zinc reduces the protons = a **reducing agent**

protons oxidize the zinc = **oxidizing agent**

For convenience split

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e$$

Oxidation (loss of electrons) increase in positive chargec

$$2H^{+}_{(aq)} + 2e \rightarrow H_{2(g)}$$

Reduction (gain of electrons) (reduction in the positive charge on the ion)

VERY IMP #1. - can't really have 1/2 reaction because the electrons need to go somewhere!!!!

They are TOO Charge dense To exist in solution

VERY IMP #2 - no net change in number of electrons (corollary of VERY IMP #1 - if we "lose" 2 need to "gain" 2

OIL - Oxidation is Losing Electrons

RIG Reduction is Gaining Electrons

Writing the Net Ionic Reaction of an Oxidation-Reduction Reaction

Track the **oxidation numbers** of the various species

Oxidation Numbers

1. Ox # of an element in an elementary substance is zero
   Cl_2 Ox# Cl =0; O_2 Ox# = 0
2. Ox # of an element in a monatomic ion = charge of ion
   Cl- Na+
3. Group 1 always +1 = always Ox # +1
   Group 2 always +2 = always Ox# +2

   F = -1 (always) Other halogens = -1 (usually) unless oxyanions
   ClO_4^- Ox# of Cl = +7
   O= -2, with exceptions
   H = +1, with exceptions (when bonded to a nonmetal)

Writing the Net Ionic Reaction of an Oxidation-Reduction Reaction

Track the **oxidation numbers** of the various species

Oxidation Numbers

1. Ox # of an element in an elementary substance is zero
   Cl_2 Ox# Cl =0; O_2 Ox# = 0
2. Ox # of an element in a monatomic ion = charge of ion
   Cl- Na+
3. Group 1 always +1 = always Ox # +1
   Group 2 always +2 = always Ox# +2

   F = -1
   O= -2, with exceptions
   H = +1, with exceptions

4. Sum of Ox #=charge charge on molecule or ion

Oxidation numbers in

PbO	litharge, lead oxide
PbO_2	platternite, lead dioxide
Pb_3O_4	minium, trilead tetraoxide

Oxygen (with some exceptions) is -2

PbO	PbO_2	Pb_3O_4
? + -2= 0	? + 2(-2) = 0	3(?) +4(-2) = 0
? = +2	? = +4	3(?) + -8 = 0

This one is an oddity!!!!
Need to know that lead only adopts +2 or +4 oxidation numbers

2(+2) + (+4) = -8

Writing the Net Ionic Reaction of an Oxidation-Reduction Reaction

Split into Reactions to be balanced

Write oxidation half reaction

A. Balance atoms of element oxidized
b. Balance Rx sides by adding e
c. Balance charge by adding H^+ or OH^-
d. Balance hydrogen by adding H_2O
e. Balance oxygen

Write reduction half reaction

A. Balance atoms of element reduced
b. Balance Rx sides by adding e
c. Balance charge by adding H^+ or OH^-
d. Balance hydrogen by adding H_2O
e. Balance oxygen

Combine and balance electrons

Writing the Net Ionic Reaction of an Oxidation-Reduction Reaction

Reaction to be balanced

Write oxidation half reaction

A. Balance atoms of element oxidized
b. Balance Ox number with e
c. Balance charge by adding H^+ or OH^-
d. Balance hydrogen by adding H_2O
e. Balance oxygen

Write reduction half reaction

A. Balance atoms of element reduced
b. Balance Ox number with e
c. Balance charge by adding H^+ or OH^-
d. Balance hydrogen by adding H_2O
e. Balance oxygen

Protons are added in an acidic solution and hydroxyls are added in a basic solution.

When in doubt add protons

Example: Balance the following redox equation in an acidic solution

$$Fe^{2+}_{(aq)} + MnO^-_{4(aq)} \rightarrow Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)}$$

Ox #balance

2+ 3+

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

Mass balance

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e$$

+7 +2

$$MnO^-_{4(aq)} \rightarrow Mn^{2+}_{(aq)}$$

Mass balance

ox# balance

-1 = 4(-2) +? +2
8-1 = ?
7 = ?

$$MnO^-_{4(aq)} + 5e \rightarrow Mn^{2+}_{(aq)}$$

Example: Balance the following redox equation in an acidic solution

$$Fe^{2+}_{(aq)} + MnO^{-}_{4(aq)} \rightarrow Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)}$$

2+ 3+

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

Mass balanced

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e$$

+7 +2

$$MnO^{-}_{4(aq)} \rightarrow Mn^{2+}_{(aq)}$$

Mass balanced

$$MnO^{-}_{4(aq)} + 5e \rightarrow Mn^{2+}_{(aq)}$$

-1 + -5 = -6 +2

Charge balance

$$MnO^{-}_{4(aq)} + 5e + 8H^{+}_{aq} \rightarrow Mn^{2+}_{(aq)}$$

Example: Balance the following redox equation in an acidic solution

$$Fe^{2+}_{(aq)} + MnO^{-}_{4(aq)} \rightarrow Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)}$$

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e$$

$$MnO^{-}_{4(aq)} + 5e + 8H^{+}_{(aq)} \rightarrow Mn^{2+}_{(aq)}$$

Balance hydrogen by adding water

$$MnO^{-}_{4(aq)} + 5e + 8H^{+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O$$

Check oxygen balance

4O =4O

=+2

Check total charge balance: (-1)+5(-1)+8(+1)=+2

Example: Balance the following redox equation in an acidic solution

$$Fe^{2+}_{(aq)} + MnO^{-}_{4(aq)} \rightarrow Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)}$$

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e$$

$$MnO^{-}_{4(aq)} + 5e + 8H^{+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O$$

Recombine while balancing electrons

$$5Fe^{2+}_{(aq)} \rightarrow 5Fe^{3+}_{(aq)} + 5e$$

Example: Balance the following redox equation in an acidic solution

$$Fe^{2+}_{(aq)} + MnO^{-}_{4(aq)} \rightarrow Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)}$$

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e$$

$$MnO^{-}_{4(aq)} + \cancel{5e} + 8H^{+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O$$

$$5Fe^{2+}_{(aq)} \rightarrow 5Fe^{3+}_{(aq)} + \cancel{5e}$$

$$MnO^{-}_{4(aq)} + 5Fe^{2+}_{(aq)} + 8H^{+}_{(aq)} \rightarrow 5Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)} + 4H_2O$$

Final equation does not show any electrons because electrons do not exist in solution

$$MnO^{-}_{4(aq)} + 5Fe^{2+}_{(aq)} + 8H^{+}_{(aq)} \rightarrow 5Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)} + 4H_2O$$

Having balanced the equation we can know predict or measure quantities that should react.

What volume of 0.684 M $KMnO_4$ is required to react completely with 27.50 mL of 0.250 M $Fe(NO_3)_2$

$$\left[27.50mLFe(NO_3)_2\right]\left[\frac{1L}{10^3 mL}\right]\left[\frac{0.250moleFe(NO_3)_2}{1L}\right]\left[\frac{1moleFe^{2+}}{1moleFe(NO_3)_2}\right]\left[\frac{1moleMnO_4^-}{5moleFe^{2+}}\right]\left[\frac{1moleKMnO_4}{1moleMnO_4^-}\right]\left[\frac{LKMnO_4}{0.684moleKMnO_4}\right] = 2.01x10^{-3} LKMnO_4$$

$$= 2.01x10^{-3}\ LKMnO_4$$

Vol $KMnO_4$? ← Moles MnO_4^- ← Moles Fe^{2+}

stoichiometry

FITCH Rules

General
- G1: Suzuki is Success
- G2. Slow me down
- G3. Scientific Knowledge is Referential
- G4. Watch out for Red Herrings
- G5. Chemists are Lazy

Chemistry
- C1. It's all about charge
- C2. Everybody wants to "be like Mike"
- C3. Size Matters
- C4. Still Waters Run Deep
- C5. Alpha Dogs eat first

$$E_{el} = k\left(\frac{q_1 q_2}{r_1 + r_2}\right)$$

"A" students work (without solutions manual)
~ 10 problems/night.

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