Ch14

Solutions

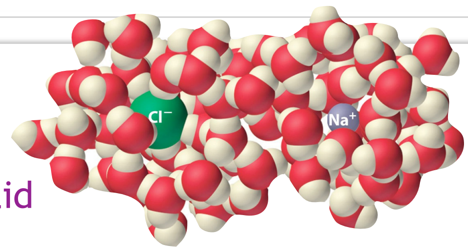
A closer look at mixtures.

Solutions and reactions in solution.

version 1.5

Ch14

Reactions in Solution



- Solubility
  - Why Solids are Solid
  - Making solutions
    - Molecular solvation
    - Ionic solvation — Dissociation
  - Electrolyte solutions
    - Electrolyte & Non-Electrolyte solns
    - Electrolyte strength
- Reactions in solution
  - Double Displacement: AB + CD ⇄ AD + CB
  - Equilibrium
  - Precipitation Reactions
- Representing Aqueous Reactions
  - Molecular solutions
  - Ionic solutions
    - Molecular eqns
    - Complete Ionic eqns
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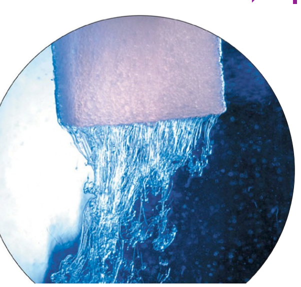

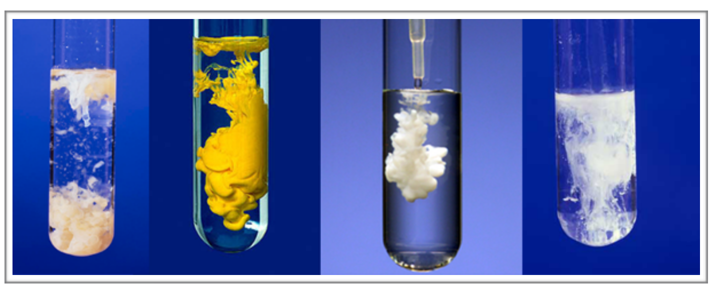

- Other Reaction Types
  - Acid-Base Reactions
    - Neutralization; $H_2O_{(l)}$
  - Gas Evolution Reactions
    - $H_2S_{(g)}$, $CO_{2\,(g)}$, $NH_{3\,(g)}$, NH_4OH, H_2CO_3
- Oxidation & Reduction
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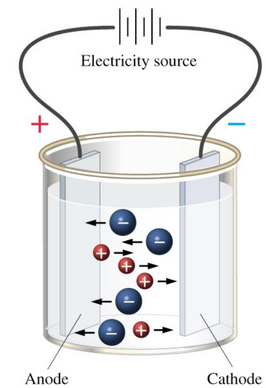
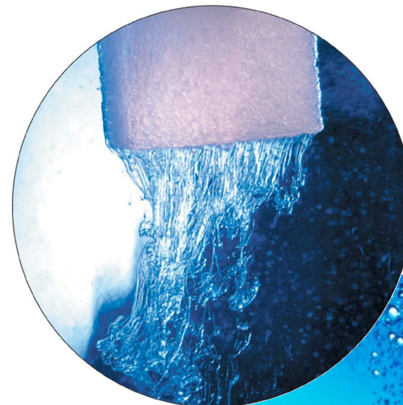


Solutions

- Solutions are homogenous mixtures.
- Mixtures can be liquids, gas, or solid.
- We're going to discuss the structure of mixtures.
- How substances come into mixtures and how substances can be driven out of mixtures.
- How substances in mixtures interact.
  - … and how that interaction facilitates chemical reaction between the mixtures components.



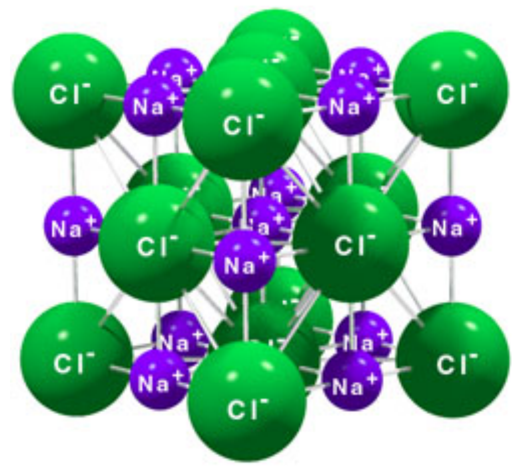
A solution is a homogenous mixture.

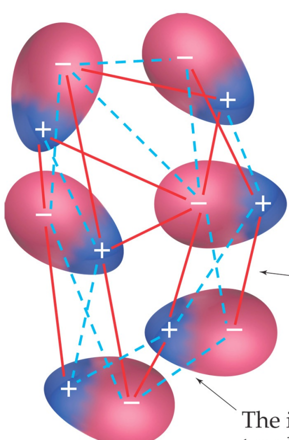
A solvent is the largest component of the mixture.

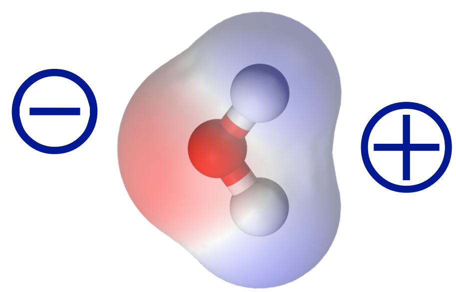
A solute is a smaller components of the mixture.

Why solids are solid.

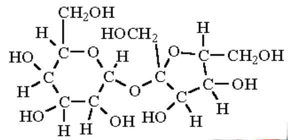
- Intermolecular forces hold solids together.
  - It’s usually about plus being attracted to minus (electrostatic attraction).
  - Molecular Solids are held together by many types of intermolecular forces.
    - The quick story is molecules have a negative end and a positive end.
    - The negative end of one molecule sticks to the positive end of another.
    - We’ll discuss the rest in Chapter 11.
  - Ionic Solids are held together by one type of intermolecular force.
    - It’s a simpler story.
    - The cations stick to a bunch of anions.
    - Those anions stick to more cations.
    - The result is a big clump of particles.

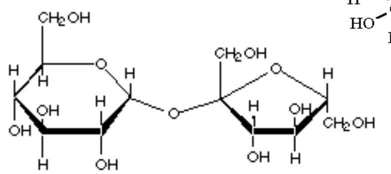




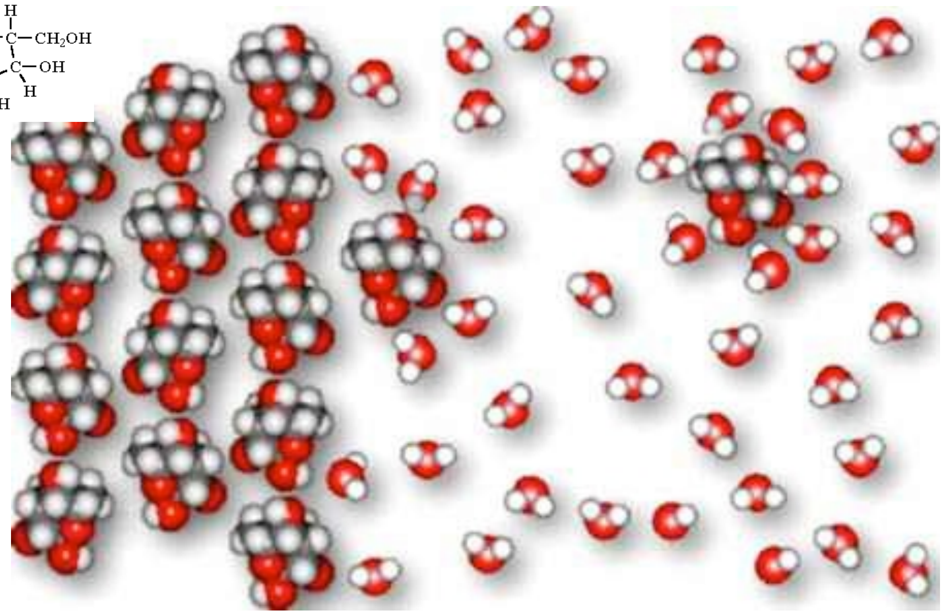


Molecular Solids Dissolve in Water



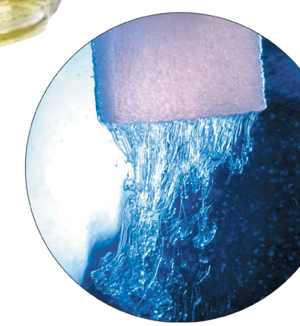


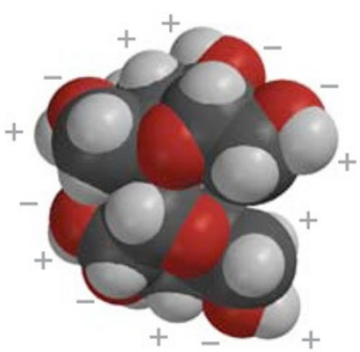
Sucrose
(glucose (α1-->2) fructose)



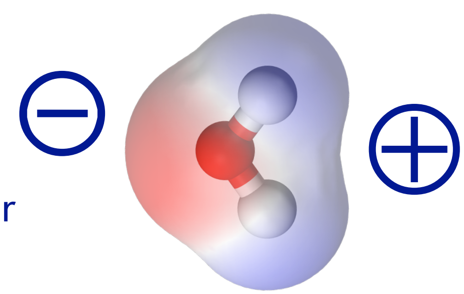
$$C_{12}H_{22}O_{11}\,(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}\,(aq)$$

Sugar dissolving in water

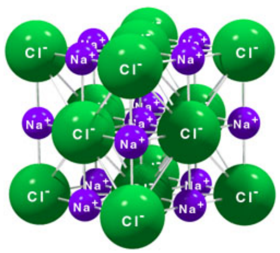


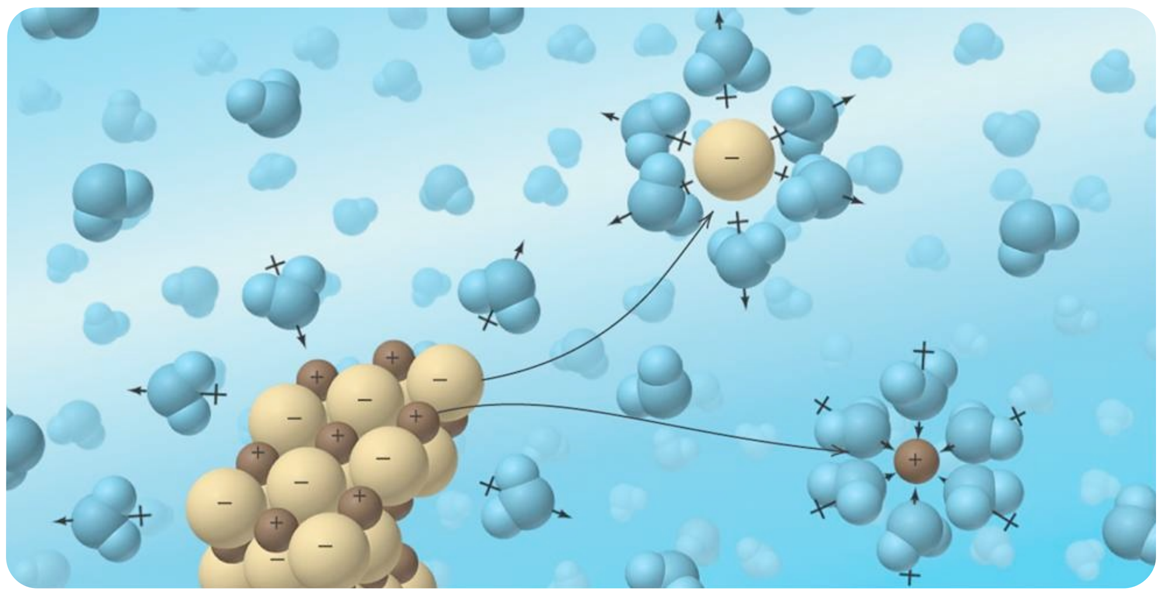


- Sugar dissolves in water.
- The molecules remain intact.
- Water molecules get in between sugar molecules.
- The result is a mixture of sugar and water.
- Mostly water.



Ionic Solids Dissolve in Water

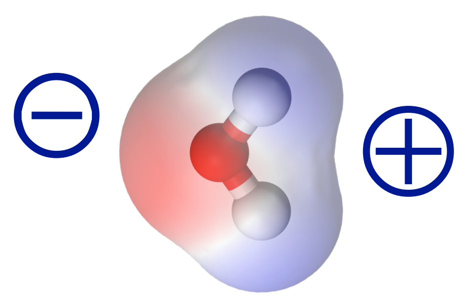




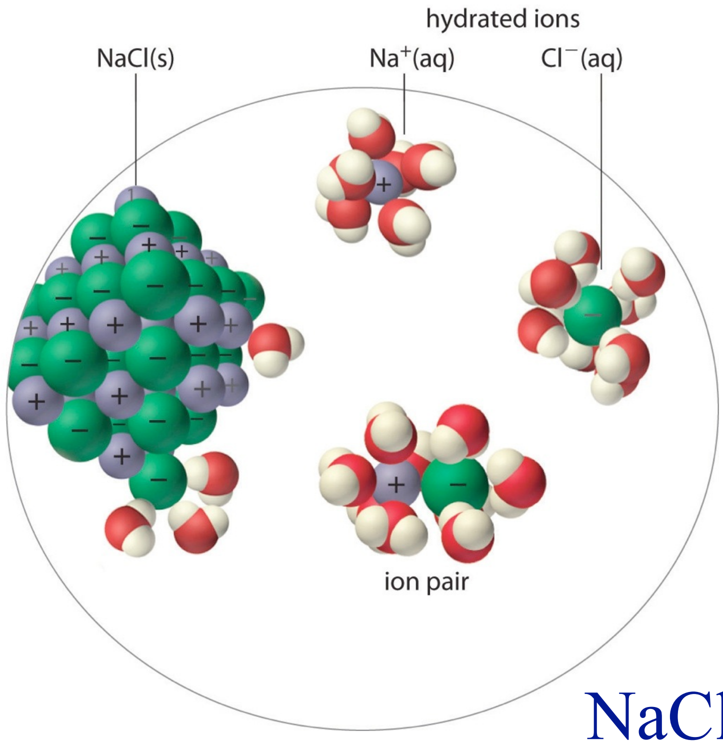
$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

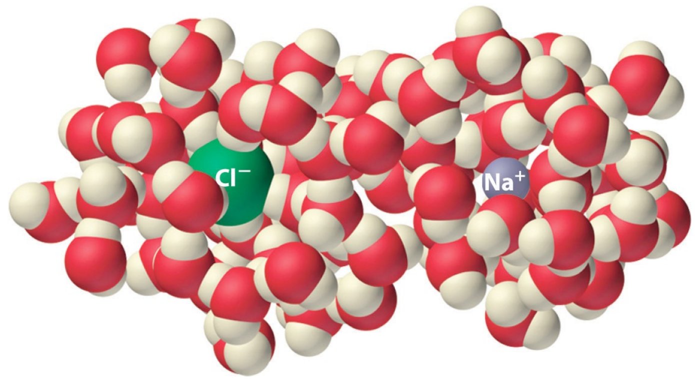
Dissociation of sodium chloride in water

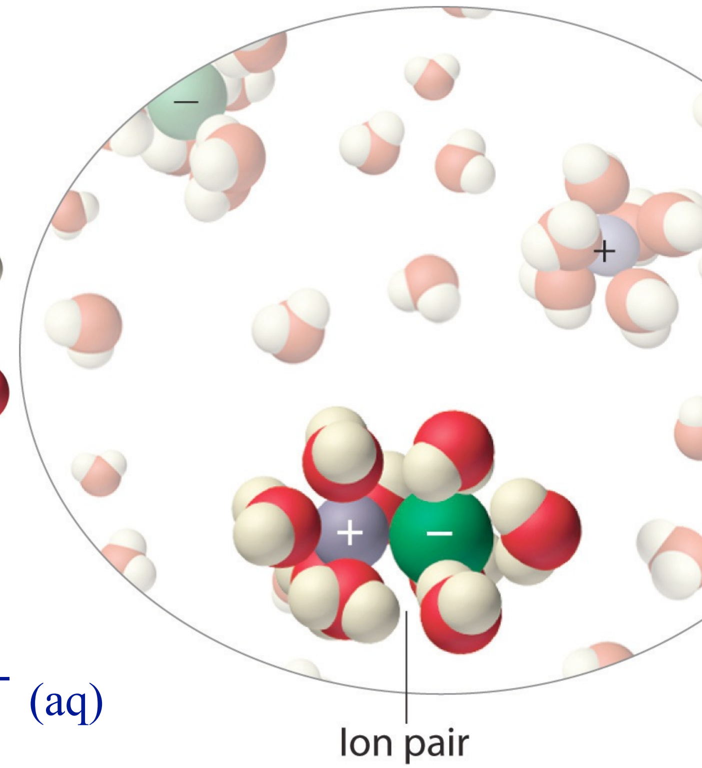
- Salt dissolves in water.
- The the ions separate.
- Water molecules get in between the ions.
- The result is a mixture of ions and water.
- Mostly water.
- Ions separating in solution is a process called dissociation.



Dissociation is often Reversible







$$NaCl_{(aq)} \underset{}{\overset{H_2O}{\rightleftarrows}} Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Dissociation of sodium chloride in water

- Dissolved ions in solution can find other dissolved ions.
- If the attraction between those ions is strong, they can re-associate.
- These dissolved ions form ion pairs.
- The ion pair is not a solid, it's still dissolved in solution.
- Ions that dissociate and re-associate in solution are a kind of reversible reaction.

Electrolytes & Acids in Solution

- Substances that dissociate in water are electrolytes.
- Those that do not dissociate in water are non-electrolytes.
- Electrolytic solutions contain dissociated ions.
- Substances that release H^+ are acids.
- Substances that accept H^+ are bases.
- Equilibrium is the state of a reversible reaction where the forward and reverse reactions are happening at the same rate.
- At equilibrium the ratio of products to reactants is constant.
- Different materials will have different product to reactant ratios.
- Electrolytic solutions conduct electricity.
- The more ions, the better it conducts.
- Electrical conductivity can be used to test the equilibrium ratio of dissociated ions to associated acids and electrolytes.
- Acids and electrolytes that favor the dissociated state are called strong.
- Acids and electrolytes that favor the associated state in water are called weak.

Electrolytes:
 eg: HCl, KNO_3, NaCl, CH_3COOH, HF

Acids:
 eg: HCl, CH_3COOH, HF, NH_4^+

Bases:
 eg: Cl^{1-}, $CH3COO^{1-}$, F^{1-}, NH_3

$HCl_{(aq)} \rightleftarrows H^+_{(aq)} + Cl^{1-}_{(aq)}$

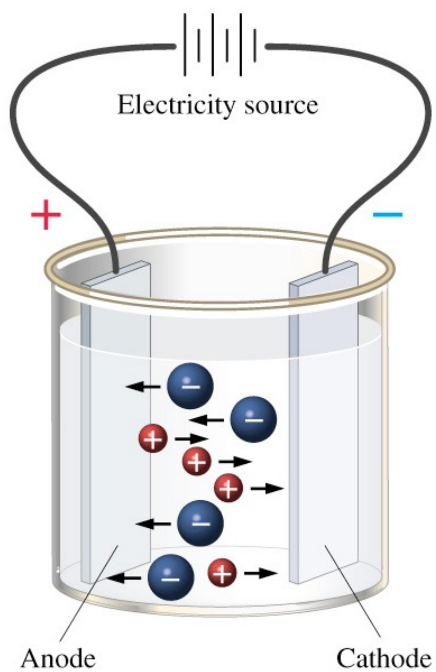
$KNO_{3\,(aq)} \rightleftarrows K^+_{(aq)} + NO_3^{1-}{}_{(aq)}$

$NH_4^+{}_{(aq)} \rightleftarrows H^+_{(aq)} + NH_{3\,(aq)}$

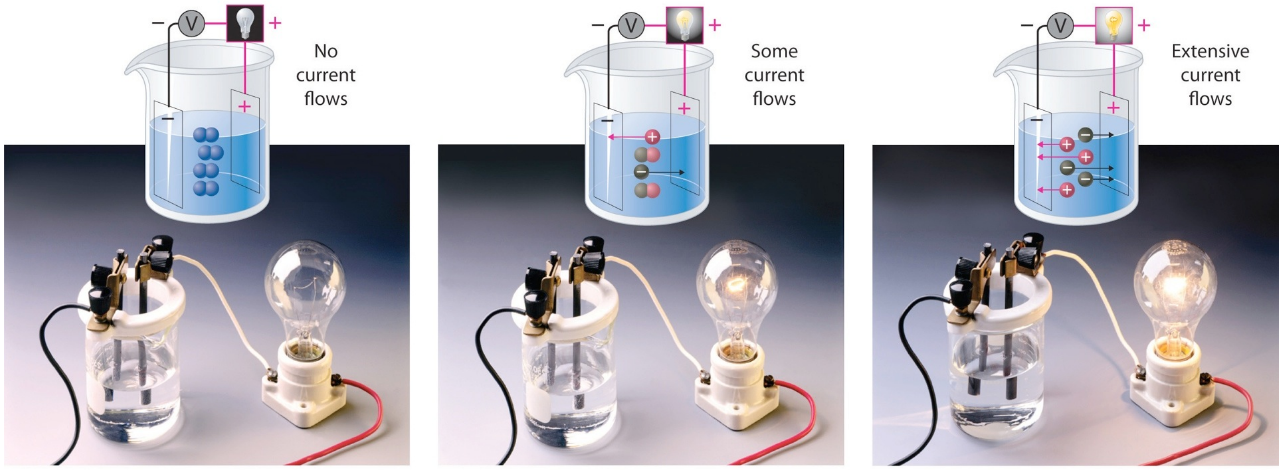
$NaCl_{(aq)} \rightleftarrows Na^+_{(aq)} + Cl^{1-}_{(aq)}$

$CH_3COOH_{(aq)} \rightleftarrows CHCOO^{1-}_{(aq)} + H^+_{(aq)}$

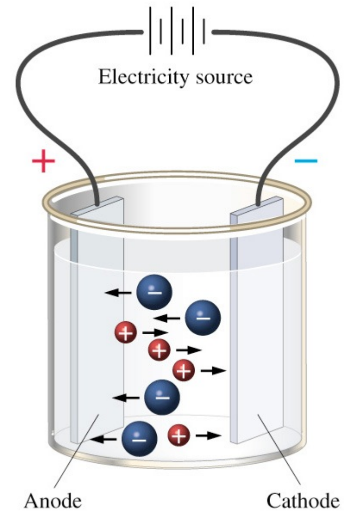
$HF_{(aq)} \rightleftarrows H^+_{(aq)} + F^{1-}_{(aq)}$



Electrolyte Strength



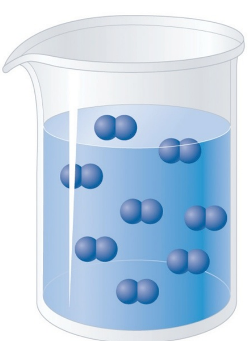
(a) Nonelectrolyte (b) Weak electrolyte (c) Strong electrolyte



Nonelectrolytes

- Molecular Substances
- Insoluble Ionic Salts

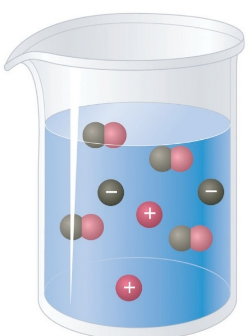
eg Sugar, AgCl, NO_2



Weak Electrolytes

- Weak Acids
- Weak Bases
- Partially soluble Ionic Salts

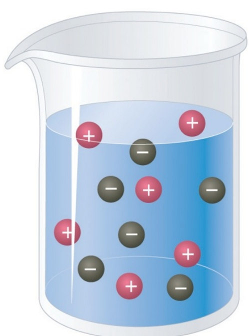
eg HOAc, HF (aq)



Strong Electrolytes

- Strong Acids
- Strong Bases
- Soluble Ionic Salts

eg HCl (aq), NaCl, H_2SO_4



$$CH_3COOH_{(aq)} \rightleftarrows CHCOO^-_{(aq)} + H^+_{(aq)}$$

4 of 100 molecules dissociate

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

100 of 100 dissociate

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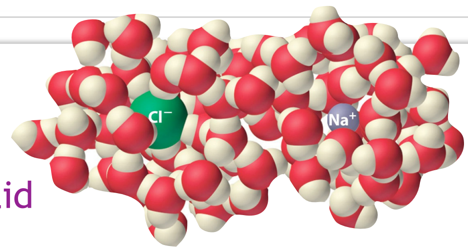
Reactions in Solution



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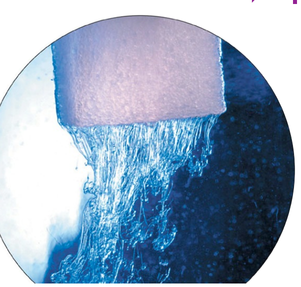

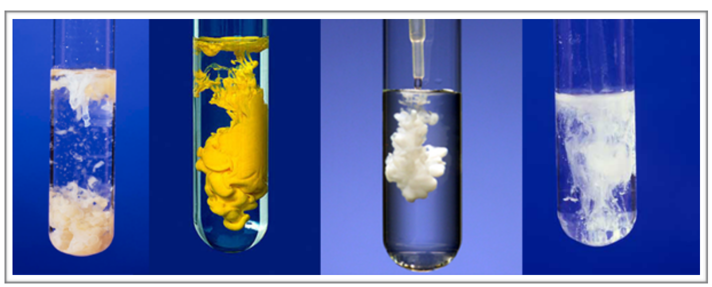

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Oxidation & Reduction
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  - How oxidation occurs
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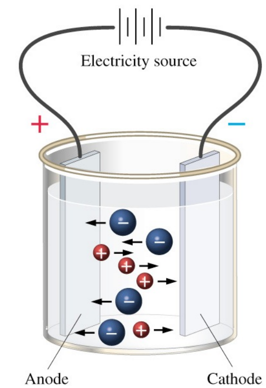


Reactions in Solution

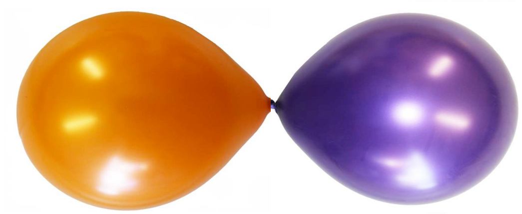
- If you dissolve more than one electrolyte in solution, you get a mixture of ions.
- The ions bump into each other and apart again, trading partners and just bouncing around the solution.
- That's not exciting.

$$KI_{(aq)} \rightleftarrows K^{+}_{(aq)} + I^{1-}_{(aq)}$$

$$NaNO_{3\,(aq)} \rightleftarrows Na^{+}_{(aq)} + NO_3^{1-}{}_{(aq)}$$

$$KI_{(aq)} + NaNO_{3\,(aq)} \rightleftarrows K^{+}_{(aq)} + I^{1-}_{(aq)} + Na^{+}_{(aq)} + NO_3^{1-}{}_{(aq)} \rightleftarrows NaI_{(aq)} + KNO_{3\,(aq)}$$

- But those ions sometimes pair up to form things that are non-electrolytes.
- When they do an irreversible reaction occurs.



$$Pb(NO_3)_2 \rightleftarrows Pb^{2+}_{(aq)} + NO_3^{1-}{}_{(aq)}$$

$$KI_{(aq)} + Pb(NO_3)_{2\,(aq)} \rightleftarrows K^{+}_{(aq)} + I^{1-}_{(aq)} + Pb^{2+}_{(aq)} + NO_3^{1-}{}_{(aq)}$$

$$KI_{(aq)} + Pb(NO_3)_{2\,(aq)} \rightleftarrows K^{+}_{(aq)} + I^{1-}_{(aq)} + Pb^{2+}_{(aq)} + NO_3^{1-}{}_{(aq)} \rightarrow PbI_{2\,(s)}\downarrow + KNO_{3\,(aq)}$$

- This removes dissociated ions from equilibrium. Which pulls more substrate ions into the dissociated state.
- And drives the reaction to complete formation of the non-electrolyte product.
- Possible non-electrolytes that can drive the reaction include:
    - insoluble solids (precipitates)
    - volatile gases (NH_3, CO_2, H_2S)
    - water (H_2O)

Double Displacement Reactions

- We call this class of reaction, where two electrolytes react in solution, a double displacement reaction.

$$AB + CD \rightarrow AD + CB$$

- It's only a reaction if a product is a non-electrolyte.

$$KI_{(s)} + NaNO_{3\,(s)} \rightleftarrows KNO_{3\,(aq)} + NaI_{(aq)}$$ ← no reaction (write "N/R")

$$KI_{(s)} + Pb(NO_3)_{2\,(s)} \rightarrow PbI_{2\,(s)}\downarrow + KNO_{3\,(aq)}$$ ← a reaction because PbI_2 is not soluble in water

- When there is a reaction you can show it three different ways:

Molecular Equation: $KI_{(aq)} + Pb(NO_3)_{2\,(aq)} \rightarrow PbI_{2\,(s)}\downarrow + KNO_{3\,(aq)}$

Complete Ionic Equation: $K^{+}_{(aq)} + I_{(aq)} + Pb^{2+}_{(aq)} + NO_3^{1-}{}_{(aq)} \rightarrow PbI_{2\,(s)}\downarrow + K^{+}_{(aq)} + NO_3^{1-}{}_{(aq)}$

Net Ionic Equation: $I^{1-}_{(aq)} + Pb^{2+}_{(aq)} \rightarrow PbI_{2\,(s)}\downarrow$

Remove the spectator Ions

- When there is no reaction you show it this way:

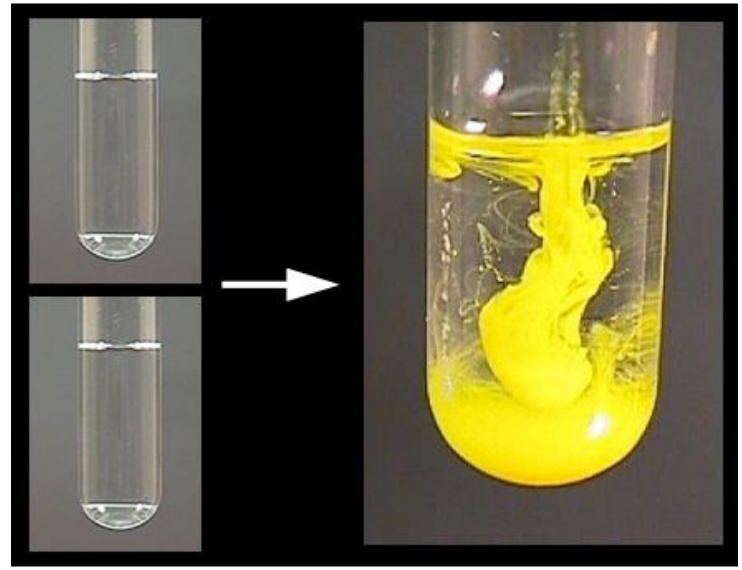
$$KI_{(s)} + Pb(NO_3)_{2\,(s)} \rightarrow N/R$$

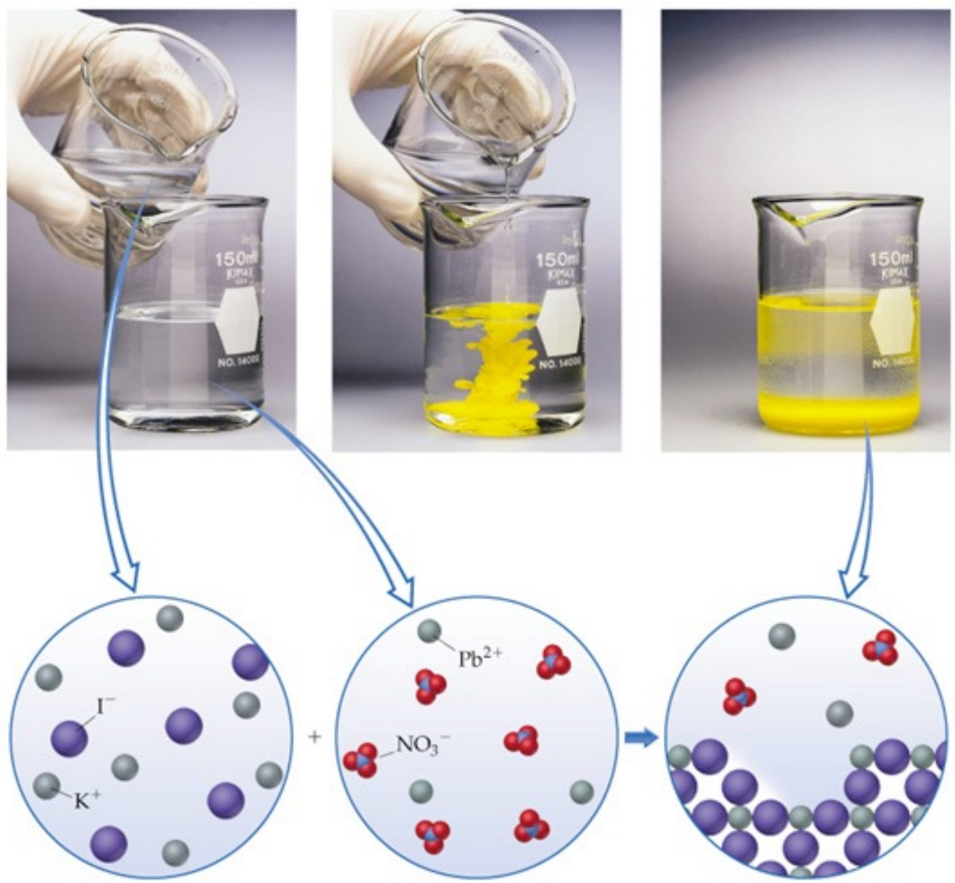
- How do you know if there's a reaction? (non-electrolytes)

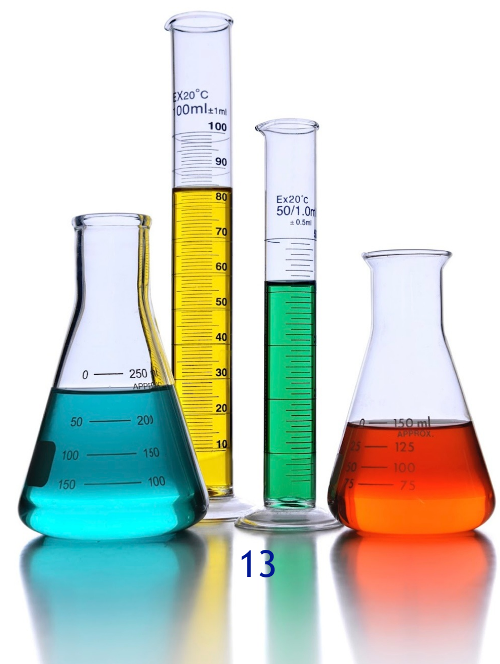
If one of the following products form, you know a reaction occurred:
(a) An insoluble solid (precipitate) (b) a Gas (c) Water

Solubility & Precipitation

- Different materials have different solubility properties.
- If an insoluble material forms in solution, it precipitates or falls out of solution.







Finding the Net Equation

Aqueous solutions of magnesium chloride and lead (II) acetate, are mixed, a bright yellow solid appears in the solution. What happened?

Magnesium Chloride$_{(aq)}$ + Lead(II) Acetate$_{(aq)}$ → ?

$$AB + CD \rightarrow AD + CB$$

$$MgCl_{2(aq)} + Pb(OAc)_{2(aq)} \rightleftarrows Mg(OAc)_{2(aq)} + PbCl_{2(aq)}$$

$$MgCl_{2(aq)} + Pb(OAc)_{2(aq)} \rightarrow Mg(OAc)_{2(aq)} + PbCl_{2(s)} \downarrow$$

Molecular Equation

$$\cancel{Mg^{2+}}_{(aq)} + Cl^{1-}_{(aq)} + Pb^{2+}_{(aq)} + \cancel{OAc^{1-}}_{(aq)} \rightarrow \cancel{Mg^{2+}}_{(aq)} + \cancel{OAc^{1-}}_{(aq)} + PbCl_{2(s)} \downarrow$$

Complete Ionic Equation

Spectator ions appear on both sides of the arrow.

$$Cl^{1-}_{(aq)} + Pb^{2+}_{(aq)} \rightarrow PbCl_{2(s)} \downarrow$$

Net Ionic Equation

$$2\ Cl^{1-}_{(aq)} + Pb^{2+}_{(aq)} \rightarrow PbCl_{2(s)} \downarrow$$

Balanced Net Ionic Equation

What forms a precipitate?

Check each step, **in order**.

Solubility Rules you are responsible for.

	Solubility Rules	Soluble (no precipitate)	Insoluble (forms precipitate)
Step 1 ANIONS	Acetates (OAc^{1-} or $CH3COO^{1-}$) Nitrates (NO_3^{1-})	Always	Never
Step 2 CATIONS	Ammonium (NH_4^{1+}) Alkali metal (Na^{1+}, Li^{1+}, K^{1+}...) Acids (the ones we learned)	Always	Never
Step 3 ANIONS	Carbonates (CO_3^{2-}) Phosphates (PO_4^{3-})	Never	Always
Step 4 has exceptions ANIONS	Halogens (Cl^{1-}, Br^{1-}, I^{1-}, F^{1-})	Usually	Except: $Ag+$, Hg_2^{2+} or Pb^{2+}
	Sulfates (SO_4^{2-})	Usually	Hg_2^{2+} or Pb^{2+} Sr^{2+}, Ba^{2+}
	Sulfides (S^{2-}) Hydroxy Salts (OH^{1-})	Except: Sr^{2+}, Ba^{2+}, Ca^{2+}	Usually

Hg_2^{2+}
mercury (I) ion

Hg^{2+}
mercury (II) ion

If you remember 1-3 you'll be good 85% of the time
If you remember 1-3 and 4 you'll be good 95%
Remembering the exceptions isn't that hard
— there's only six ions that cause exceptions
and lead, mercury, and silver are the most commonly encountered ones.

Is it soluble?

- ✓ KNO_3
- ✓ $(NH_4)_3P$
- ✓ $MnCl_2$
- ✗ $PbCl_2$
- ✓ $HClO_3$
- ✓ $CuCH_3CO_2$
- ✓ $Ca(OAc)_2$
- ✗ $CaCO_3$

Always:	Acetates Nitrates
	Ammonium Alkali metal Acids
	Carbonates Phosphates

Usually:	
	Halogens
	Sulfates
	Sulfides Hydroxy Salts

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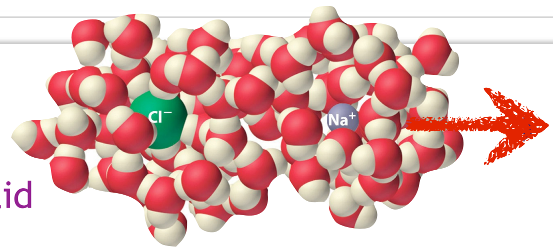
Reactions in Solution



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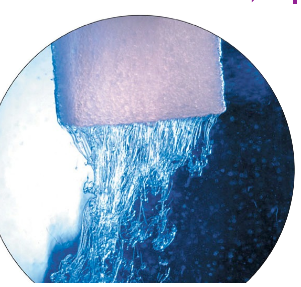

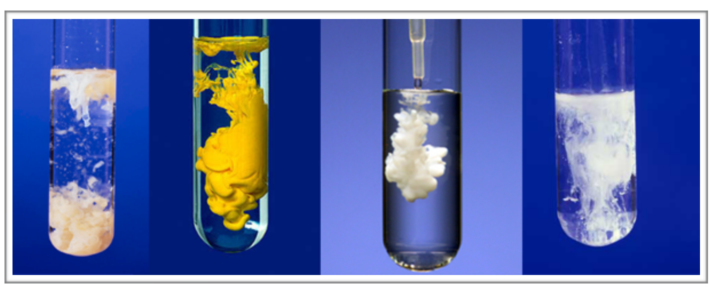

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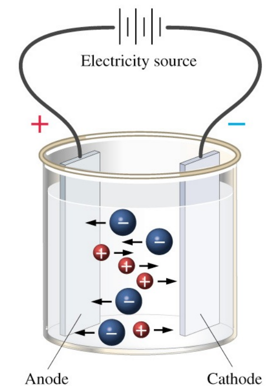


Acid-Base Reactions

- Acids and bases have multiple definitions.
- For now:
  - An **acid** is any substance which dissociates to release H+ (aq).
  - A **base** is any substance which reacts with H+ (aq).

    (You will explore other definitions in Chem 220.)

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

Acid — Base — Water — Salt

- **Acid-base reactions** are reactions between an acid and a base.
- **Neutralization reactions** are irreversible reactions between an acid and a base.
- Neutralization reactions produce water.
- The irreversible production of water can drive equilibrium forward, the same as precipitate formation.

$$HCl_{(aq)} + NaOH_{(aq)} \rightleftarrows H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)}$$

Molecular Equation

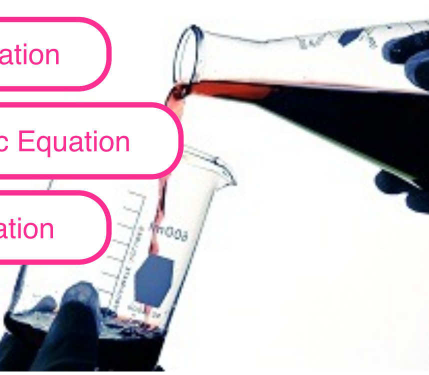
$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$$

Complete Ionic Equation

$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + Cl^-_{(aq)} + Na^+_{(aq)}$$

Net Ionic Equation

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$



Gas Formation Reactions

- Volatile gases like $CO_{2\,(g)}$, $H_2S_{(g)}$ and $NH_{3\,(g)}$ that form immediately bubble off.
- The gases escape, their formation is irreversible.
- Sometimes the double displacement reaction forms an unstable compound that decomposes into the gases. Example:

$$H_2CO_{3\,(aq)} \rightarrow H_2O_{(l)} + CO_{2\,(g)} \qquad NH_4OH_{(aq)} \rightarrow H_2O_{(l)} + NH_{3\,(g)}$$

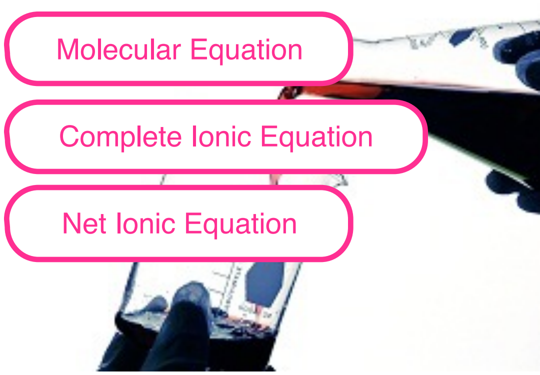
- If a double displacement reaction forms $CO_{2\,(g)}$, $H_2S_{(g)}$, or $NH_{3\,(g)}$ gases this irreversible reaction will drive equilibrium forward.
- If a double displacement reaction forms $H_2CO_{3\,(aq)}$ or $NH_4Cl_{(aq)}$ these decompose to gases and drive equilibrium forward.
- Examples:

$$HCl_{(aq)} + Na_2S_{(aq)} \rightleftarrows H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + S^{2-}_{(aq)}$$

$$HCl_{(aq)} + Na_2S_{(aq)} \rightarrow H_2S_{(g)}\uparrow + NaCl_{(aq)}$$

$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + S^{2-}_{(aq)} \rightarrow H_2S_{(g)}\uparrow + Cl^-_{(aq)} + Na^+_{(aq)}$$

$$H^+_{(aq)} + S^{2-}_{(aq)} \rightarrow H_2S_{(g)}\uparrow$$



Molecular Equation

Complete Ionic Equation

Net Ionic Equation

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- If a double displacement reaction forms $H_2CO_{3\,(aq)}$ or $NH_4Cl_{(aq)}$ these decompose to gases and drive equilibrium forward.

$$H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightleftarrows H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + Na^+_{(aq)} + HCO_3^-{}_{(aq)}$$

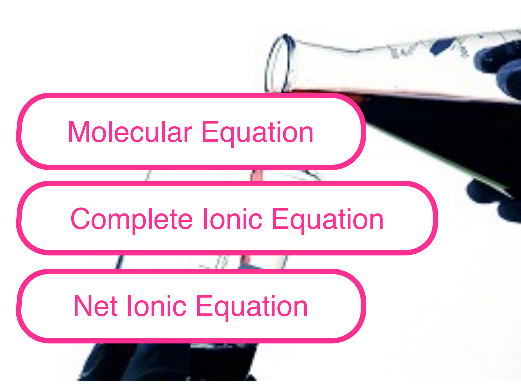
$$H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightarrow SO_4^{2-}{}_{(aq)} + Na^+_{(aq)} + H_2CO_{3(aq)}$$

$$H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightarrow Na_2CO_{3(aq)} + H_2CO_{3(aq)}$$

$$H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightarrow Na_2SO_{4\,(aq)} + H_2O_{(l)} + CO_{2(g)}\uparrow$$

$$H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + Na^+_{(aq)} + HCO_3^-{}_{(aq)} \rightarrow SO_4^{2-}{}_{(aq)} + Na^+_{(aq)} + H_2O_{(l)} + CO_{2(g)}\uparrow$$

$$H^+_{(aq)} + HCO_3^-{}_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}\uparrow$$



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- If a double displacement reaction forms $H_2CO_{3\,(aq)}$ or $NH_4Cl_{\,(aq)}$ these decompose to gases and drive equilibrium forward.

$$NaOH_{(aq)} + NH_4Cl_{(aq)} \rightleftarrows Na^+_{(aq)} + OH^-_{(aq)} + NH_4^+_{(aq)} + Cl^-_{(aq)}$$

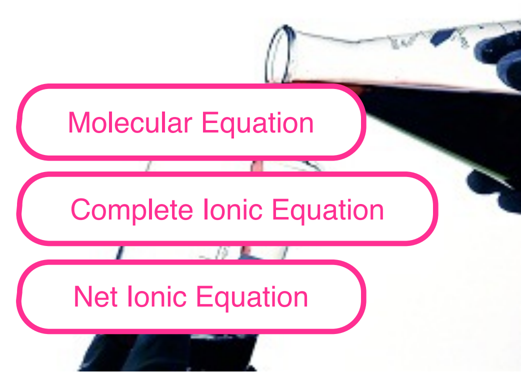
$$NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + NH_4OH_{(aq)}$$

$$NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NaCl + NH_4OH_{(aq)}$$

$$NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NaCl_{\,(aq)} + H_2O_{(l)} + NH_{3(g)} \uparrow$$

$$Na^+_{(aq)} + OH^-_{(aq)} + NH_4^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)} + NH_{3(g)} \uparrow$$

$$OH^-_{(aq)} + NH_4^+_{(aq)} \rightarrow H_2O_{(l)} + NH_{3(g)} \uparrow$$



Molecular Equation

Complete Ionic Equation

Net Ionic Equation

Double Displacement Reactions

- If I mix two electrolytes (AB & CD), I can look at the two possible double displacement products (AD & CB) to predict if a reaction will occur.
- If either of the two products forms irreversibly, a reaction will occur.
  - Irreversible reactions include precipitation formation, neutralization and gas formation.
- For each pair of possible products below, did a reaction occur?

AB + CD → AD + CB

$(NH_4)_2SO_4$ + H_2CO_3

$Fe(OAc)_3$ + $Mn(NO_3)_2$

H_2SO_4 + $Hg_2(NO_3)_2$

CO_2 + $MnCl_2$

C_2H_4O + AgCl

$HBrO_3$ + H_2S

H_2O + KBr

NH_4OH + $MgCO_3$

Predict the products...

$NaCl + Mn(NO_3)_3 \rightarrow$ N/R

$NaCl + AgNO_3 \rightarrow AgCl_{(s)}\downarrow + NaNO_{3\,(aq)}$

$K_2CO_3 + Ca(NO_3)_2 \rightarrow KNO_{3\,(aq)} + CaCO_{3\,(s)}\downarrow$

$K_2CO_3 + NaCl \rightarrow$ N/R

$K_2CO_3 + HBr \rightarrow KBr_{(aq)} + H_2O_{(l)} + CO_{2(g)}\uparrow$

$FeCl_3 + Hg_2(OAc)_2 \rightarrow Hg_2Cl_{2\,(s)}\downarrow + Fe(OAc)_{3\,(aq)}$

$TiCl_4 + NH_4NO_3 \rightarrow$ N/R

$NH_4OH + H_2SO_4 \rightarrow (NH_4)_2SO_{4\,(aq)} + H_2O_{(l)}$

Always:	Acetates Nitrates
	Ammonium Alkali metal Acids
	Carbonates Phosphates

Usually:	
	Halogens
	Sulfates
	Sulfides Hydroxy Salts

(a) Precipitate Product (b) Gas Product (c) Water as a Product

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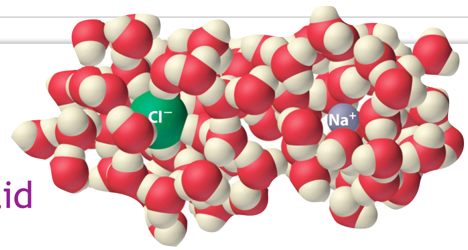
Reactions in Solution



- Solubility
  - Why Solids are Solid
  - Making solutions
    - Molecular solvation
    - Ionic solvation — Dissociation
  - Electrolyte solutions
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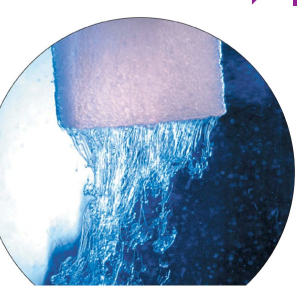
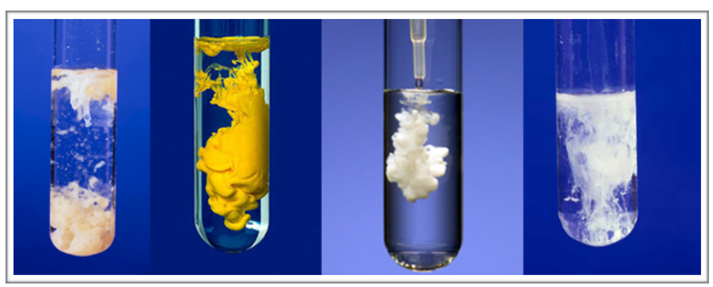
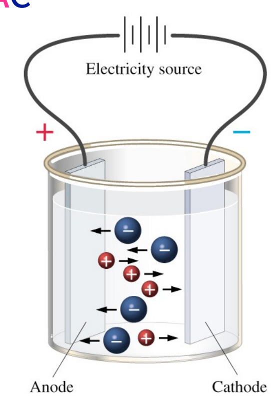




- Other Reaction Types
  - Acid-Base Reactions
    - Neutralization; $H_2O_{(l)}$
  - Gas Evolution Reactions
    - $H_2S_{(g)}$, $CO_{2\,(g)}$, $NH_{3\,(g)}$, NH_4OH, H_2CO_3
- Oxidation & Reduction
  - Single Displacement: A + BC ⇄ B + AC
    - How oxidation occurs
    - Oxidation Numbers
  - Red-Ox Reactions
    - Half Reactions
    - Metal Activity
  - Combustion Reactions



Oxidation & Reduction

- If an atom gains electrons, it's said to be reduced.

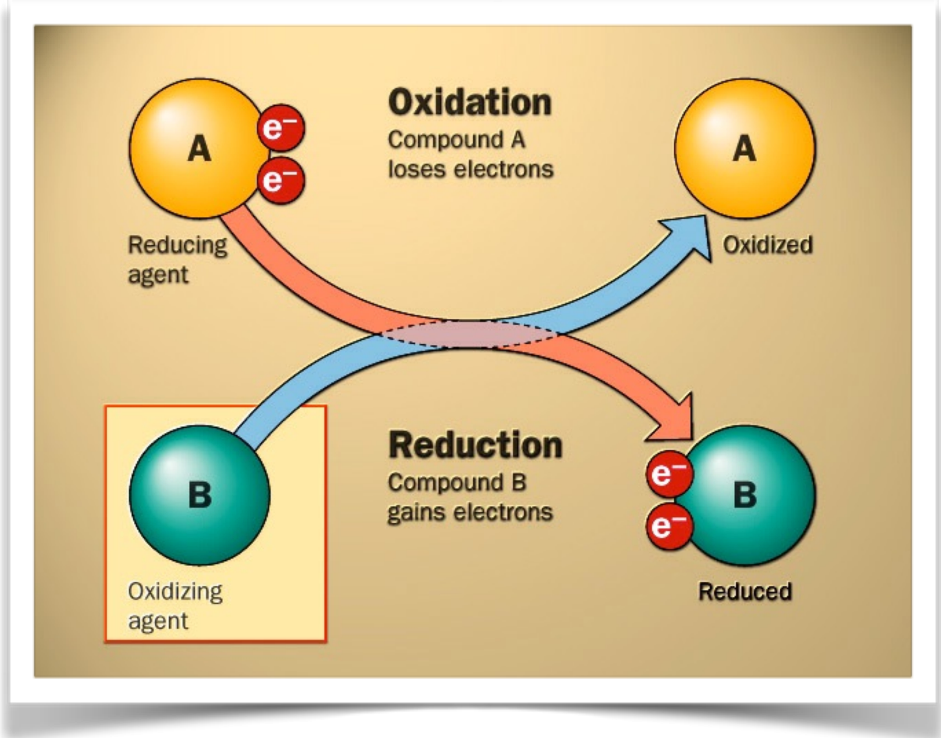
  Example: $Fe^{3+} \rightarrow Fe^{0}$

  $Cl^{0} \rightarrow Cl^{1-}$

- If an atom looses electrons, it's said to be oxidized.

  Example $Fe^{0} \rightarrow Fe^{3+}$

  $Cl^{1-} \rightarrow Cl^{0}$



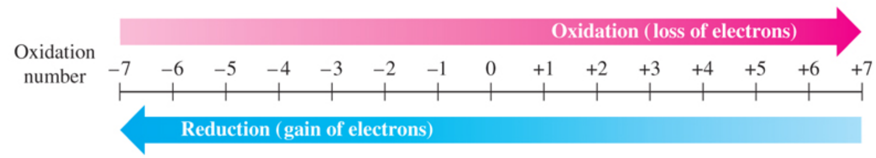
- Chemical reactions where electrons are transferred from one atom to another are called oxidation-reduction reactions.

  Example: $Fe + HCl \rightarrow FeCl_3 + H_2$

- It can be tricky to figure out which atoms gained or lost electrons in a reaction.
- In the above reaction:
  - Iron was oxidized.
  - Chlorine neither gained nor lost electrons.
  - Hydrogen was reduced.
- To help us explore oxidation-reduction reactions we assign oxidation numbers to each atom in the solution.

Oxidation Numbers

- Every atom in solution has an oxidation number.
  - If the number goes up, the species has been oxidized.
  - If the number goes down it's been reduced.
- Oxidation numbers can be positive or negative.
- The sum of the oxidation numbers in a molecule or ion equals it's charge.
- Finding oxidation Numbers:
  - Elements in their natural state are always oxidation number 0.
    - Fe, Au, Ne, H_2, Cl_2, P_4, S_8 are all oxidation number 0.
  - Monatomic Ions have an oxidation number equal to their charge.
    - Na+ is 1, Mg^{2+} is 2, Ca^{2+} is 2, S^{2-} is -2, N^{3-} is -3
  - Elements in a compound or molecule...
    - Fluorine is the king. He is always oxidation number -1.
    - Hydrogen is the wild card. He's usually:
      - +1 when bonded to non-metals
      - -1 when bonded to metals
    - Oxygen is next. Unless trumped by fluorine, oxygen is usually -2 (exception: in peroxides he's -1)
    - Other elements get priority in order of their proximity to Fluorine:
      - elements in row 7A get -1, 6A get -2, 5A get -3
    - It's like musical chairs, the last element get's what ever is left over.



CH_2O H_2O $MgCl_2$ NaH CO_3^{2-} $NaClO_4$ PO_4^{3-} ClO_2^- HSO_4^- P_4 PH_3 $Cr_2O_7^{2-}$

Chlorine's oxidation number?

0	Cl_2
-1	NaCl (Na^+ Cl^{1-})
+1	ClO^{1-}
+5	$HClO_3$
+7	$ClOF_5$
+9	Cl_2O_9

Identifying Red-Ox Reaction

- When an atoms oxidation number goes up in a reaction, it's been oxidized (lost electrons).
- When an atoms oxidation number goes down in a reaction, it's been reduced (gained electrons).
- For underlined atom in each reaction below, determine if it's been oxidized, reduced, or neither.

<u>Iron</u> rusting to <u>Iron</u> (III) oxide. — Oxidized

$2\ AgCl_{(s)} + \underline{H}_{2\ (g)} \rightarrow 2\ \underline{H}^{1+}_{(aq)} + 2\ Ag_{(s)} + 2\ Cl^{1-}_{(aq)}$ — Oxidized

$\underline{Mn}O_4{}^{1-}{}_{(aq)} + I^{1-}{}_{(aq)} \rightarrow \underline{Mn}^{2+}{}_{(aq)} + I_{2\ (s)}$ — Reduced

$\underline{Na}_3PO_{4\ (aq)} + H_2SO_{4\ (aq)} \rightarrow H_3PO_{4\ (aq)} + \underline{Na}_2SO_{4\ (aq)}$ — Neither

Precipitating <u>gold</u> metal from <u>gold</u> ions in sea water. — Reduced

Oxidation-Reduction Reactions

- Atoms that gain electrons (negative charges) are reduced.
- Atoms that loose electrons are oxidized.
- Electrons always end up somewhere. If something in the reaction is getting oxidized, something else is getting reduced.
- Red-ox processes are *not* an equilibrium processes — someone wins; someone looses; end of story. No trade-backs.
- You can drive equilibrium with red-ox processes, just like you drive it with other precipitation, gas formation or water formation.
- Metals can be oxidized by acids and salts (rust is an example).
- Metal oxidation often occurs by a single displacement mechanism.

$$A + BC \rightarrow AC + B$$

$$Zn_{(s)} + HBr_{(aq)} \rightarrow ZnBr_{2(aq)} + H_{2(g)}$$

$$Zn_{(s)} + H^{1+}_{(aq)} + Br^{1-}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Br^{1-}_{(aq)} + H_{2(g)}$$

oxidation number	0	+1	-1	+2	-1	0

Zn is oxidized (0 goes to +2)
Hydrogen is Reduced (+1 goes to 0)
Bromine is neither.

Oxidation-Reduction Reactions

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$$Mn_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Mn(NO_3)_{2(aq)} + Pb_{(s)}$$

$$Cu_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow N/R$$

- If the reaction will occur?

Oxidation-Reduction Half Reactions

- How do we know if the reaction happens? Look at the complete ionic equation.

$$Mn_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Mn(NO_3)_{2(aq)} + Pb_{(s)}$$ Molecular Equation

$$Mn_{(s)} + Pb^{2+} + NO_3^{1-} \rightarrow Mn^{2+} + NO_3^{1-} + Pb_{(s)}$$ Complete Ionic Equation

- Remove the spectator ions to see the net ionic equation.

$$Mn_{(s)} + Pb^{2+} \rightarrow Mn^{2+} + Pb_{(s)}$$ Net Ionic Equation

- There are two half reactions which make up the net ionic equation.

$$Mn_{(s)} \rightarrow Mn^{2+} + 2e^-$$
$$Pb^{2+} + 2e^- \rightarrow Pb_{(s)}$$
Half Reaction Equations

- The two half reactions show that we're looking at a competition for electrons. It's basically a tug of war.
- You can turn around one equation to compare them side to side.
  We need to decide who's gonna win the fight over those two electrons.

$$Mn^{2+} + 2e^- \rightarrow Mn_{(s)}$$
$$Pb^{2+} + 2e^- \rightarrow Pb_{(s)}$$

- We could look up numbers for whose is better at holding electrons,
  or we could just reference a list of "who beat's who" — the activity series.

The Activity Series

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^{+}(aq) + e^{-}$
Potassium	$K(s) \longrightarrow K^{+}(aq) + e^{-}$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^{-}$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^{-}$
Sodium	$Na(s) \longrightarrow Na^{+}(aq) + e^{-}$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^{-}$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^{-}$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^{-}$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^{-}$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^{-}$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^{-}$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^{-}$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^{-}$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^{-}$
Hydrogen	$H_2(g) \longrightarrow 2H^{+}(aq) + 2e^{-}$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$
Silver	$Ag(s) \longrightarrow Ag^{+}(aq) + e^{-}$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^{-}$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^{-}$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^{-}$

- Which metal (oxidation zero) is more "active"?
- We look at the half reactions.
- **An atom of an element in the activity series will displace an atom of an element below it from one of its compounds.**

The Activity Series

<u>Metals</u>

K
Ca
Na
Mg
Al
Zn
Fe
Ni
Sn
Pb
H
Cu
Ag
Au

<u>Grouped Metals</u>

K Ca Na Mg

Ga Al Zn

Fe Co Ni

Sn Pb

H

Cu Ag Au

- Which metal (oxidation zero) is more "active"?
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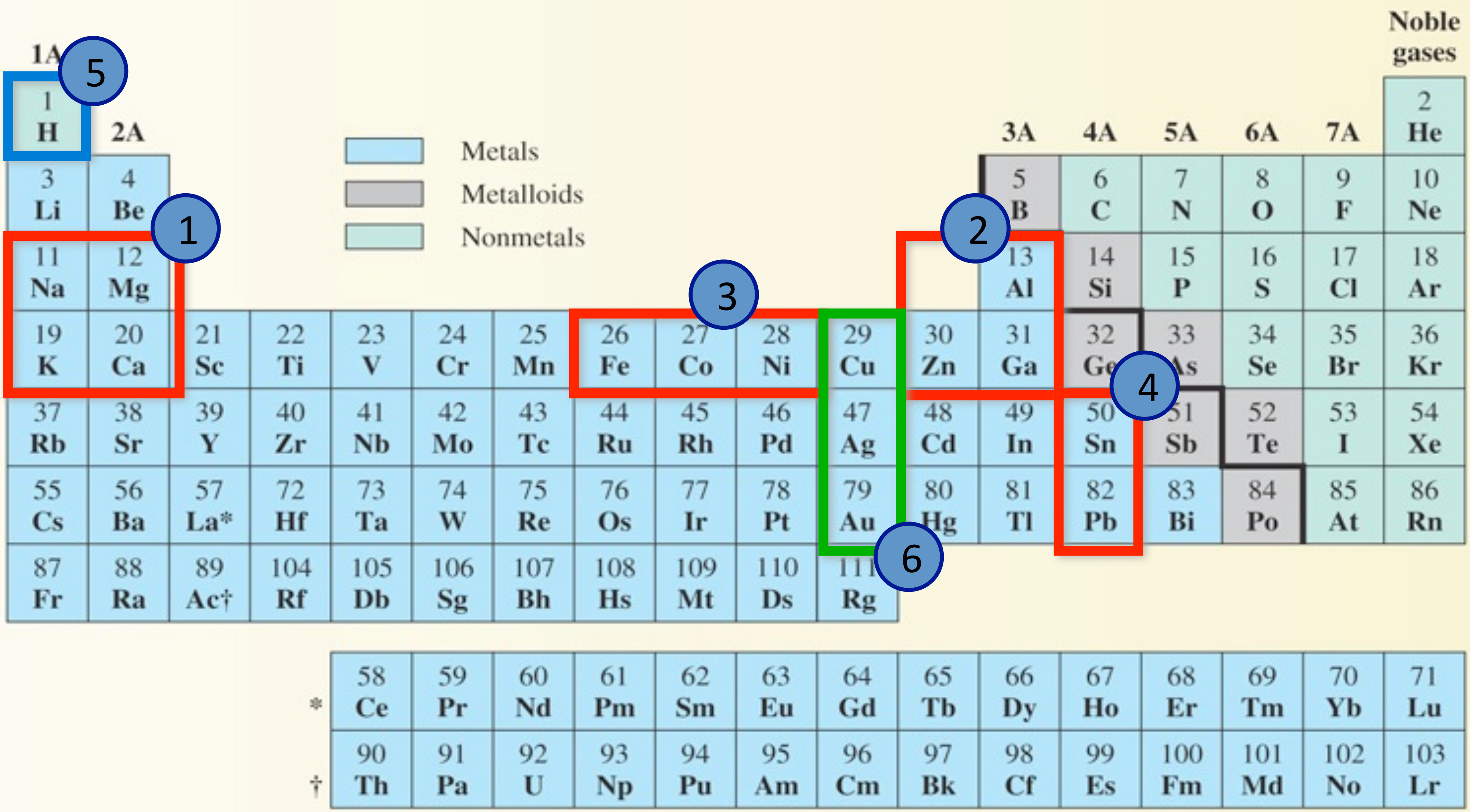
Metal (cation) Activity Series



Oxidation & Reduction

- How do we know which metal gives up it's electrons? Check "activity." The more active ion is the one more likely to turn into a cation (give up it's electrons).
- Which is more active (more likely to loose it's electrons)?

Sodium or Iron?

Al or Co?

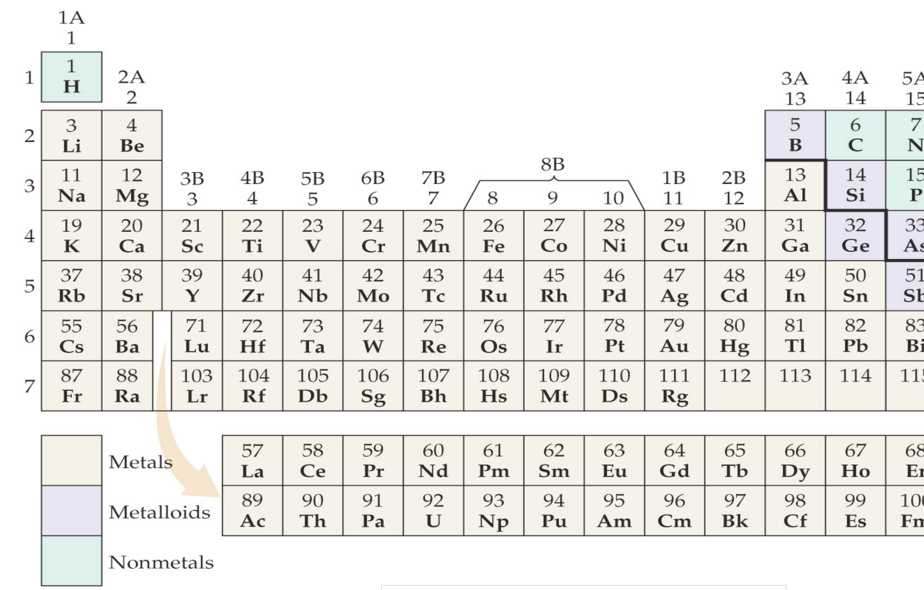
H_2 or Mg?

Hydrogen or Gold?

Sodium or Zinc?

Pb or Cu?

Nickel or Calcium?



Oxidation & Reduction

- How do we know which metal gives up it's electrons? Check "activity." The more active ion is the one more likely to turn into a cation (give up it's electrons).
- Which reactions will occur?

A + BC ➞ AC + B

$Na_{(s)} + FeBr_{3(aq)} \rightarrow ?$

Na more active than Fe? Yes.

$Na_{(s)} + FeBr_{3(aq)} \rightarrow Fe_{(s)} + NaBr_{(aq)}$

$Fe_{(s)} + Zn(ClO_3)_{2(aq)} \rightarrow ?$

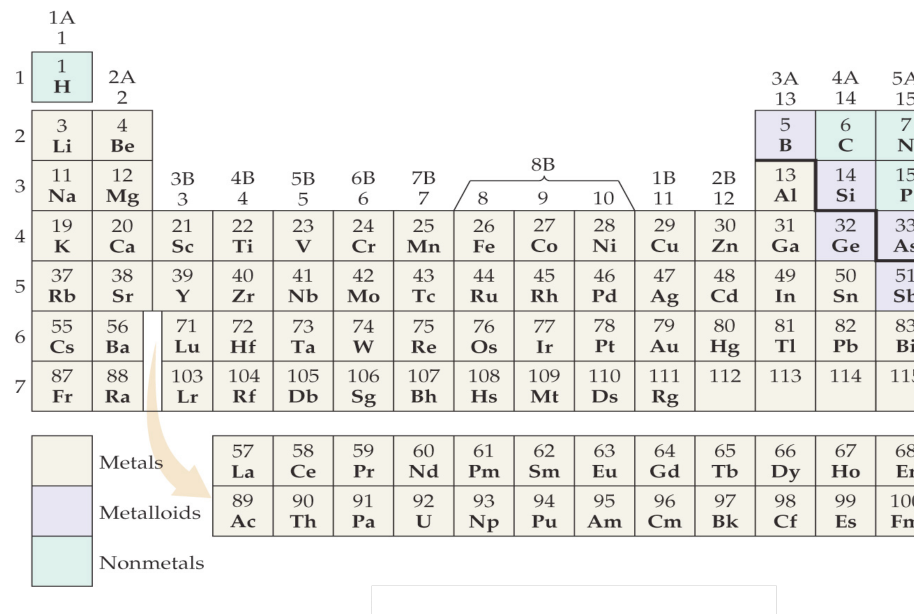
Fe more active than Zn? No.

$Fe_{(s)} + Zn(ClO_3)_{2(aq)} \rightarrow$ N/R

$Sn_{(s)} + HNO_{3(aq)} \rightarrow ?$

Sn more active than H? Yes.

$Sn_{(s)} + HNO_{3(aq)} \rightarrow H_{2(g)} \uparrow + Sn(NO_3)_{4(aq)}$



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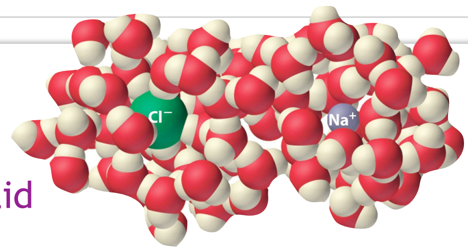
Reactions in Solution



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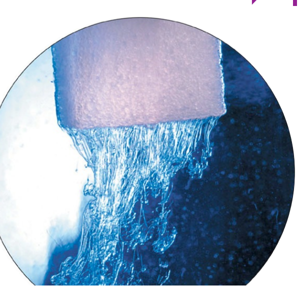

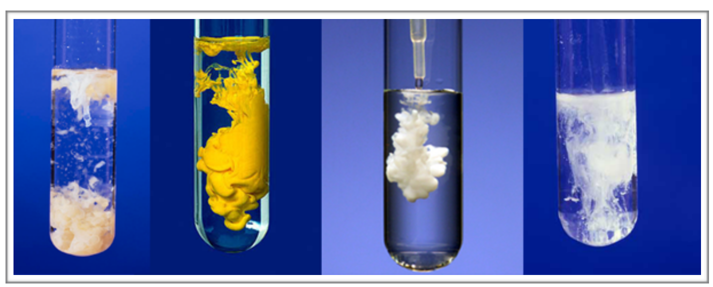

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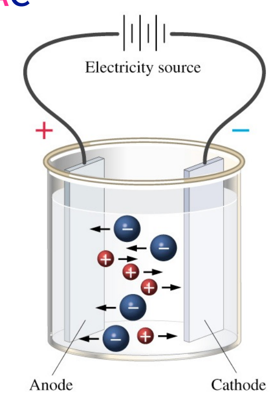
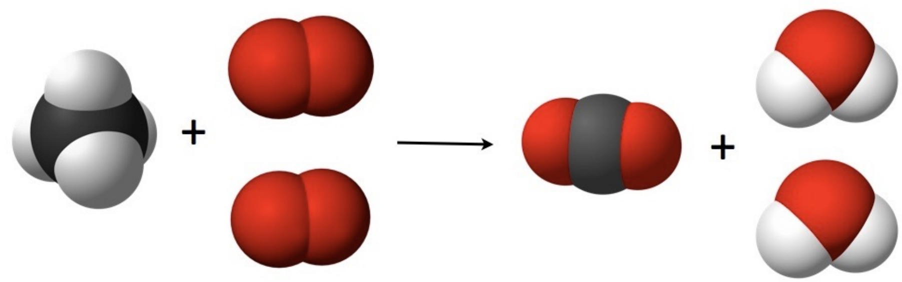


Combustion Reactions

- Burning something is causing it to combust.
- Combustion reactions are reacting any substance with oxygen to form the most stable binary compounds of it's elements and oxygen.
- The most common products are CO_2 and H_2O. Other common products are NO_2 and P_2O_5.
- Combustion reactions are red-ox reactions, in which oxygen is reduced.
- The driving force in combustion reactions is oxygens fierce demand for electrons. Harnessing that property of oxygen is what gave us the internal combustion engine and is at the heart of most of fuels humans use.



$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

$$X + O_2 \rightarrow H_2O + CO_2 + NO_2 + P_2O_5 + \ldots$$

Reaction Types

Considering…

- Kinetics (what could be formed?)
  - Double Displacement
  - Single Displacement
- Driving force (will it happen?)
  - Precipitation Reactions
  - Acid Base Reactions
  - Gas Evolution Reactions
  - Reduction-Oxidation Reactions
    - Metal Activity
    - Combustion

$$AB + CD \rightarrow AD + CB$$

$$A + BC \rightarrow B + AC$$

… you can predict if two substances will react and what products it will likely produce.

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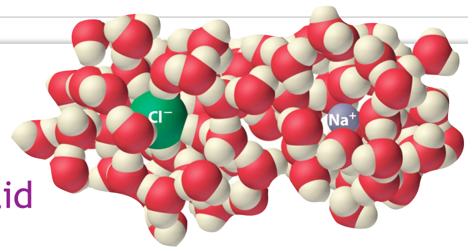
Reactions in Solution



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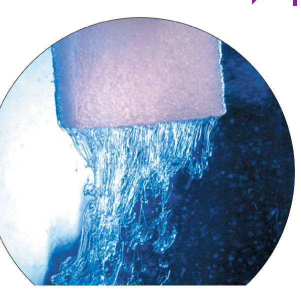
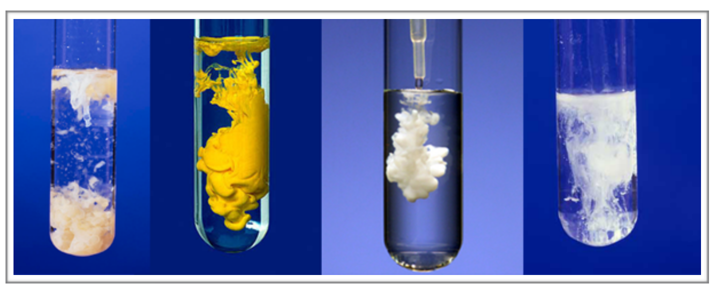
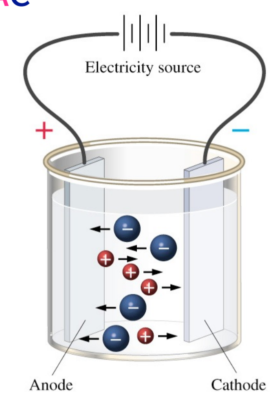
    - Electrolyte strength
- ## Reactions in solution
  - Double Displacement: $AB + CD \rightleftarrows AD + CB$
  - Equilibrium
  - Precipitation Reactions
- ## Representing Aqueous Reactions
  - Molecular solutions
  - Ionic solutions
    - Molecular eqns
    - Complete Ionic eqns
    - Net Ionic eqns





- ## Other Reaction Types
  - Acid-Base Reactions
    - Neutralization; $H_2O_{(l)}$
  - Gas Evolution Reactions
    - $H_2S_{(g)}$, $CO_{2\,(g)}$, $NH_{3\,(g)}$, NH_4OH, H_2CO_3
- ## Oxidation & Reduction
  - Single Displacement: $A + BC \rightleftarrows B + AC$
    - How oxidation occurs
    - Oxidation Numbers
  - Red-Ox Reactions
    - Half Reactions
    - Metal Activity
  - Combustion Reactions



Questions?