

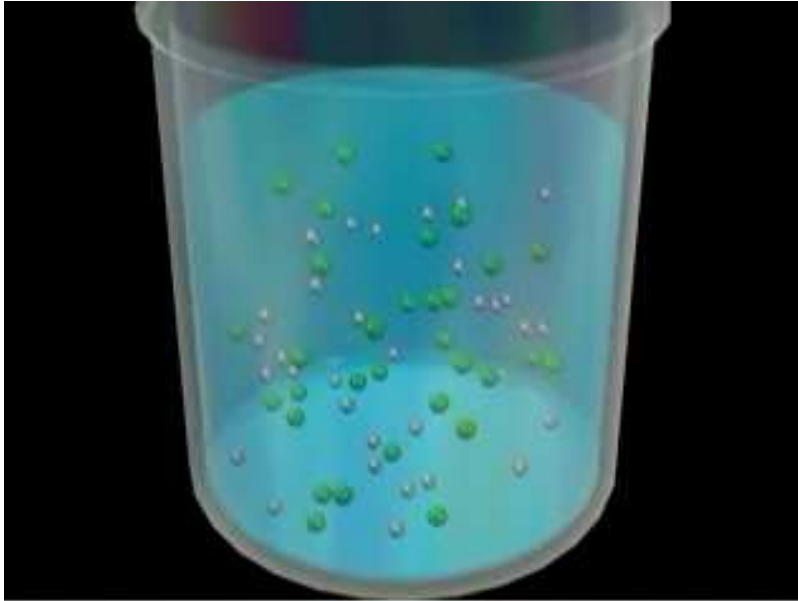
# Chapter 4: Aqueous RXNs and Solutions Stoichiometry:



- Homogeneous mixtures of two or more pure substances.
- The **solvent** is present in greatest abundance.
- All other substances are **solutes**.

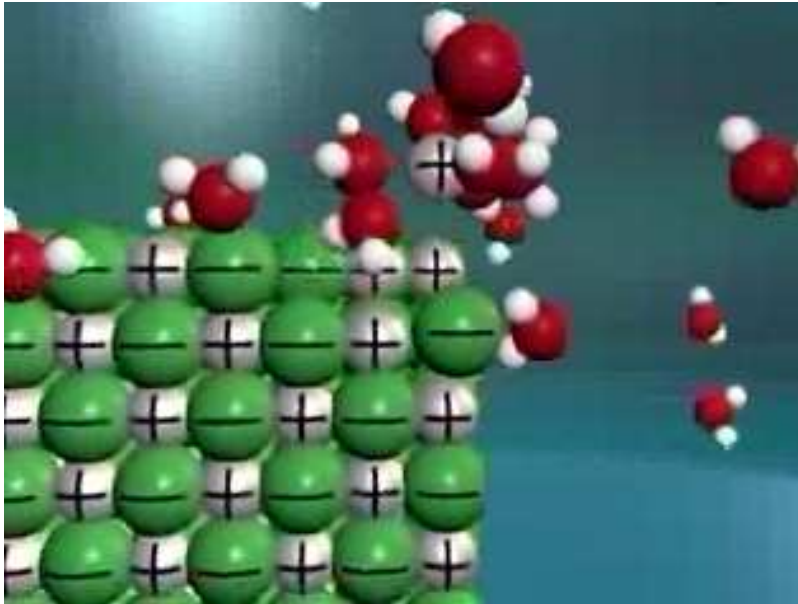
Salt water : water = solvent, salt = solute

# Electrolytes & Non-electrolytes



- Electrolytes **dissociate into ions** when dissolved (dissociates) in water. (Strong or weak)
- A **nonelectrolyte** may dissolve in water, but it does **not dissociate** into ions.

# Dissociation



- When an **ionic** substance dissolves in water, the solvent pulls the individual ions from the crystal and solvates them.
- This process is called **dissociation**.

# Guide for which compounds are soluble in water

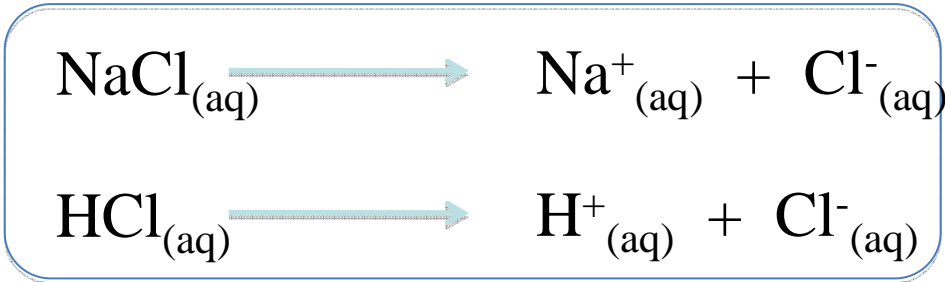
| TABLE 4.1 Solubility Guidelines for Common Ionic Compounds in Water |   |
|---|---|
| Soluble Ionic Compounds   | Important Exceptions  |
| Compounds containing $\text{NO}_3^-$                                | None  |
| <del><math>\text{C}_2\text{H}_3\text{O}_2^-</math></del>            | <del>None</del>   |
| $\text{Cl}^-$   | Compounds of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$  |
| $\text{Br}^-$   | Compounds of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$  |
| $\text{I}^-$  | Compounds of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$  |
| $\text{SO}_4^{2-}$  | Compounds of $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$                            |
| Insoluble Ionic Compounds   | Important Exceptions  |
| Compounds containing $\text{S}^{2-}$                                | Compounds of $\text{NH}_4^+$ , the alkali metal cations, and $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , and $\text{Ba}^{2+}$ |
| $\text{CO}_3^{2-}$  | Compounds of $\text{NH}_4^+$ and the alkali metal cations   |
| $\text{PO}_4^{3-}$  | Compounds of $\text{NH}_4^+$ and the alkali metal cations   |
| $\text{OH}^-$   | Compounds of the alkali metal cations, and $\text{NH}_4^+$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , and $\text{Ba}^{2+}$ |

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**$\text{NO}_3^-$  and  $\text{NH}_4^+$  ALWAYS SOLUBLE**

# Strong Electrolytes: Soluble Ionic Compds & Strong Acids/Bases

Strong Electrolytes Dissociate COMPLETELY

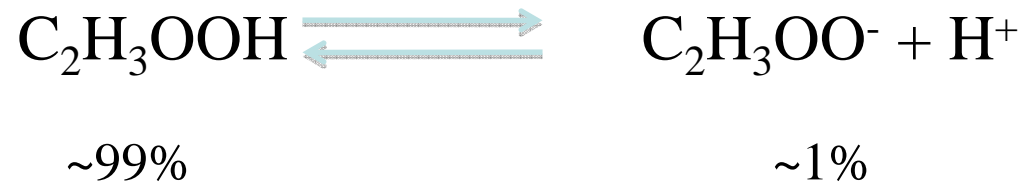


| very common acids | TABLE 4.2 Common Strong Acids and Bases  |  | very common bases |
|-------------------|--|--|-------------------|
|                   | Strong Acids                             | Strong Bases   |                   |
| →                 | Hydrochloric, HCl                        | Group 1A metal hydroxides (LiOH, NaOH, KOH, <del>RbOH, CsOH</del> )  | ←                 |
|                   | Hydrobromic, HBr                         | Heavy group 2A metal hydroxides<br><del>[Ca(OH)<sub>2</sub>, Sr(OH)<sub>2</sub>, Ba(OH)<sub>2</sub>]</del> |                   |
|                   | Hydroiodic, HI                           |  |                   |
|                   | Chloric, HClO <sub>3</sub>               |  |                   |
|                   | Perchloric, HClO <sub>4</sub>            |  |                   |
| →                 | Nitric, HNO <sub>3</sub>                 |  |                   |
| →                 | Sulfuric, H <sub>2</sub> SO <sub>4</sub> |  |                   |

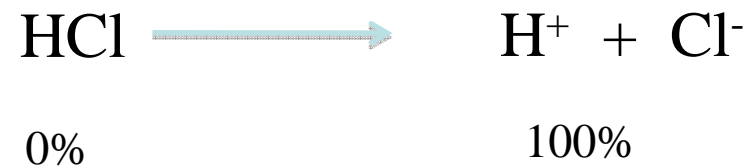
Weak Electrolytes: DO NOT completely dissociate

Weak acids: Acetic acid, (C<sub>2</sub>H<sub>3</sub>OOH)

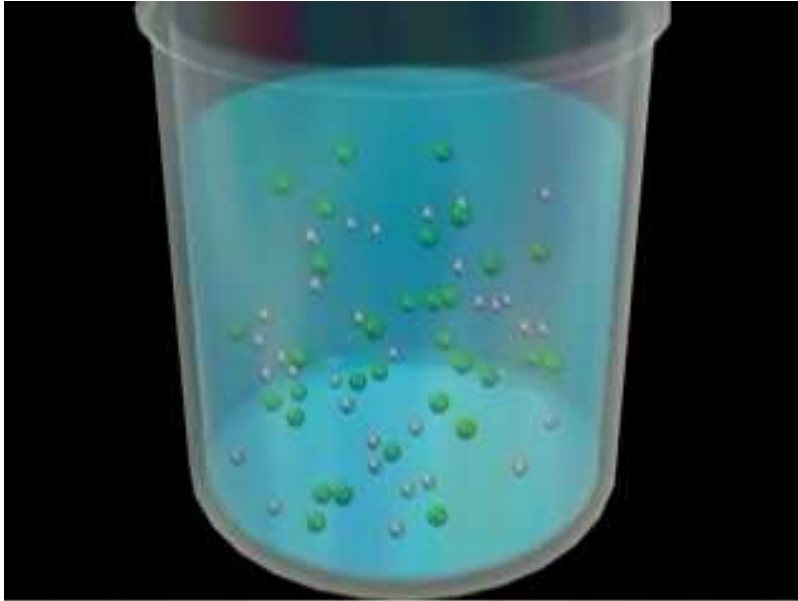
Weak bases: NH<sub>3</sub>



Strong electrolytes dissociate completely:



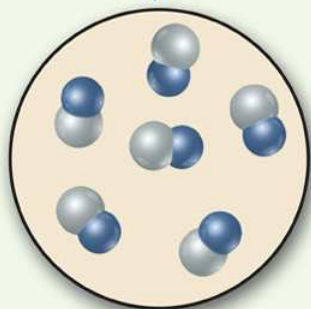
# Electrolytes & Non-electrolytes



- Electrolytes **dissociate into ions** when dissolved (dissociates) in water. (Strong or weak)
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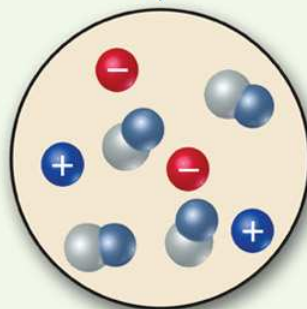
## ELECTROLYTIC PROPERTIES

*One way to differentiate two aqueous solutions is to employ a device that measures their electrical conductivities. The ability of a solution to conduct electricity depends on the number of ions it contains. An electrolyte solution contains ions that serve as charge carriers, causing the bulb to light.*



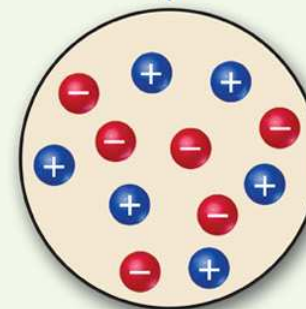
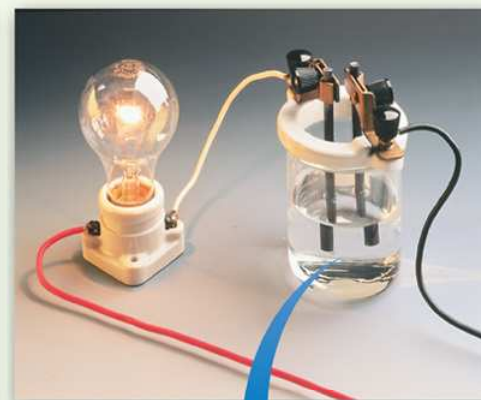
No ions

A nonelectrolyte solution does not contain ions, and the bulb does not light.



Few ions

If the solution contains a small number of ions, the bulb will be only dimly lit.



Many ions

If the solution contains a large number of ions, the bulb will be brightly lit.

# Solution Terminology : Review

Solvent – water



Solute – blue compound

Solutes are either electrolytes or non-electrolytes

Non-electrolytes DO NOT dissociate ( $C_6H_{12}O_6$ )

Electrolytes dissociate ( $NaCl \rightarrow Na^+ + Cl^-$ )

Electrolytes are either weak or strong

Strong = 100% dissociation

Weak ~1% dissociation

More ions in solution = More electrical conductivity

## 5 Types of Reactions in solution:

Precipitation Reactions – formation of a solid

Gas Forming Reactions – forms gas

Acid / Base Reaction – Exchange of protons

Oxidation / Reduction Reaction – e- exchange

Displacement Reaction – solid dissolves & new solid forms

# Precipitation Reactions & Solubility Rules

**PPT RXN:** When one mixes ions that form compounds that are insoluble, a precipitate (solid) is formed.



| Soluble Ionic Compounds |                                    | Important Exceptions   |
|-------------------------|------------------------------------|--|
| Compounds containing    | $\text{NO}_3^-$                    | None   |
|                         | $\text{C}_2\text{H}_3\text{O}_2^-$ | None   |
|                         | $\text{Cl}^-$                      | Compounds of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$                       |
|                         | $\text{Br}^-$                      | Compounds of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$                       |
|                         | $\text{I}^-$                       | Compounds of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$                       |
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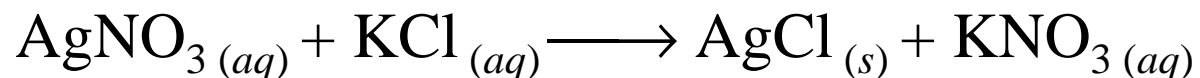
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| Insoluble Ionic Compounds |                    | Important Exceptions  |
|---------------------------|--------------------|---|
| Compounds containing      | $\text{S}^{2-}$    | Compounds of $\text{NH}_4^+$ , the alkali metal cations, and $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , and $\text{Ba}^{2+}$ |
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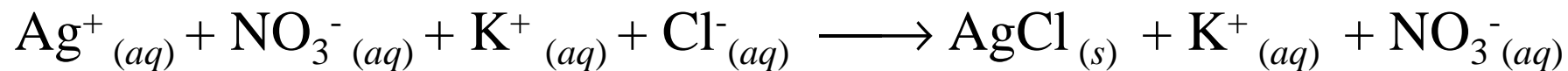
Often used for  
Qualitative  
Analysis

## Molecular Equation for Precipitation RXN:

The molecular equation lists the reactants and products in their molecular form.

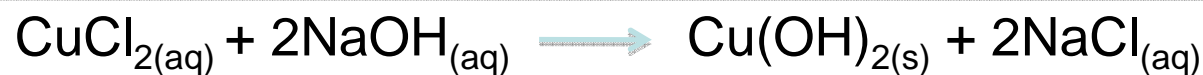


In the ionic equation all strong electrolytes (strong acids, strong bases, and soluble ionic salts) are dissociated into their ions.



# Precipitation Demonstrations

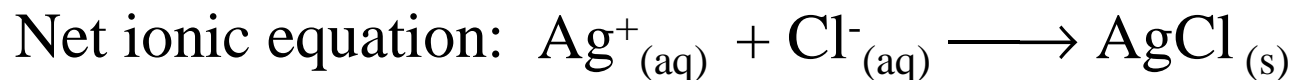
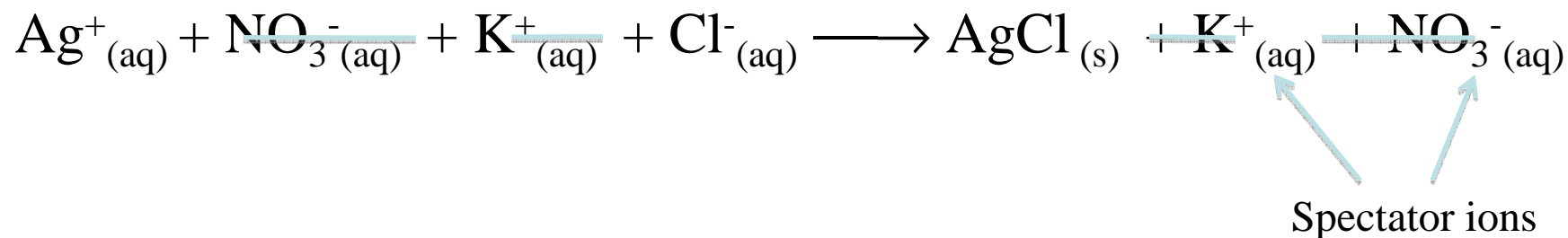
|  | NaOH <sub>(aq)</sub>   | KI <sub>(aq)</sub>  |
|--|------------------------|---------------------|
| CuCl <sub>2</sub>  | Cu(OH) <sub>2(s)</sub> | No RXN              |
| Pb(C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> ) <sub>2</sub> | Pb(OH) <sub>2(s)</sub> | PbI <sub>2(s)</sub> |
| NiCl <sub>2</sub>  | Ni(OH) <sub>2(s)</sub> | No RXN              |



## Net Ionic Equation

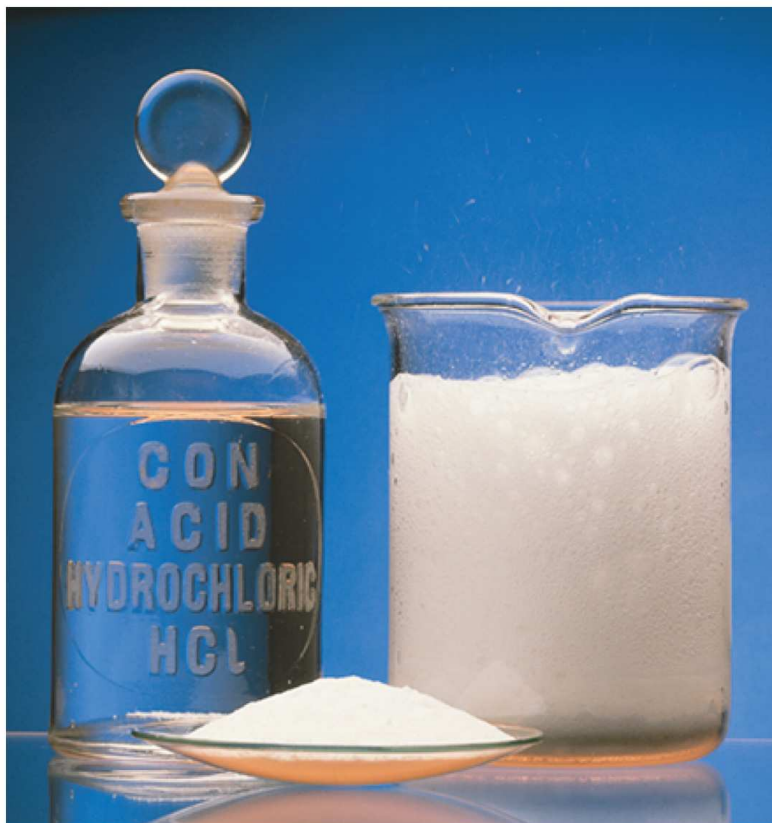
To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.

The only things left in the equation are those things that change (i.e., react) during the course of the reaction.



note: balanced equation has total charge equal on both sides

# Acids



Substances that increase the concentration of  $\text{H}^+$  when dissolved in water (Arrhenius).



There are only seven strong acids:

- Hydrochloric acid (HCl)
- Hydrobromic acid (HBr)
- Hydroiodic acid (HI)
- Nitric acid ( $\text{HNO}_3$ )
- Sulfuric acid ( $\text{H}_2\text{SO}_4$ )
- Chloric acid ( $\text{HClO}_3$ )
- Perchloric acid ( $\text{HClO}_4$ )

# Bases



Substances that increase the concentration of  $\text{OH}^-$  when dissolved in water (Arrhenius).

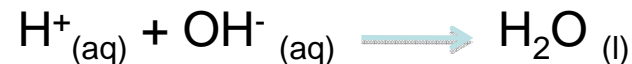
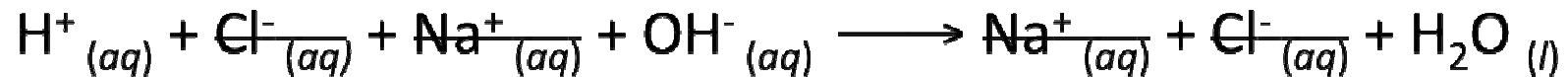
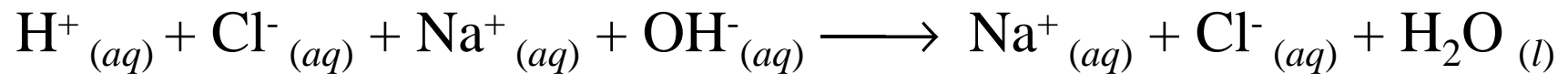


The strong bases are the soluble salts of hydroxide ion:

- Alkali metals (Li, Na, K, Rb, Cs)
- Calcium
- Strontium
- Barium

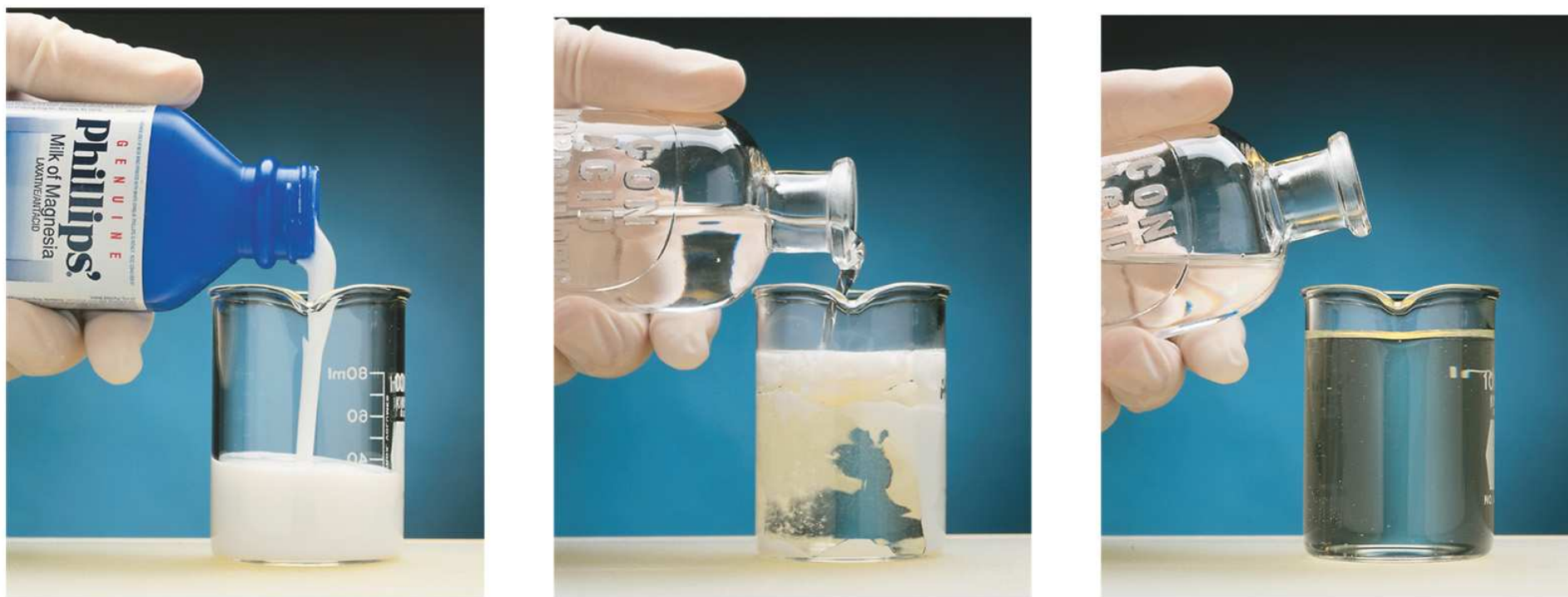
# Neutralization Reactions

Generally, when solutions of an acid and a base are combined, the products are a salt and water.

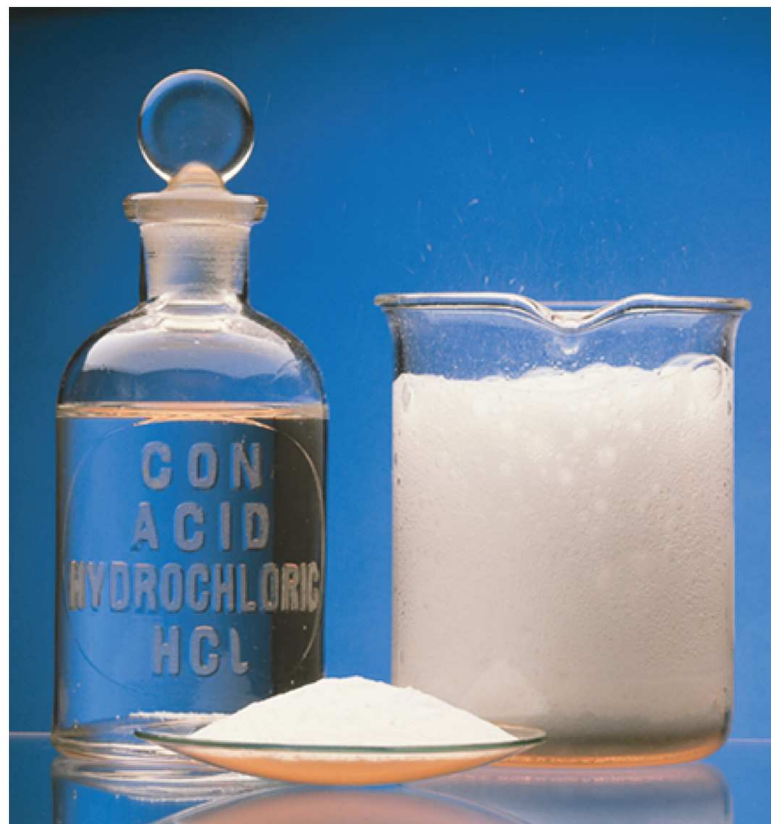
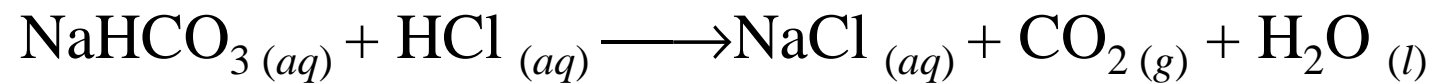


# Neutralization Reactions

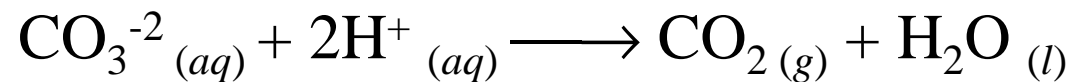
Neutralization of acid with milk of magnesia.



## Gas-Forming Neutralization Reactions



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# Acid / Base Summary

Acids generate  $\text{H}^+$  ions when dissolved in water

Bases generate  $\text{OH}^-$  ions when dissolved in water

Neutralization

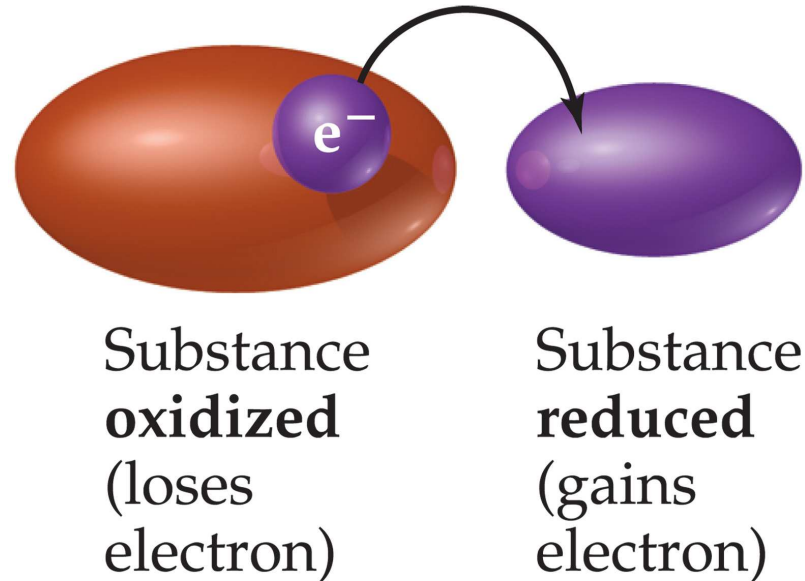
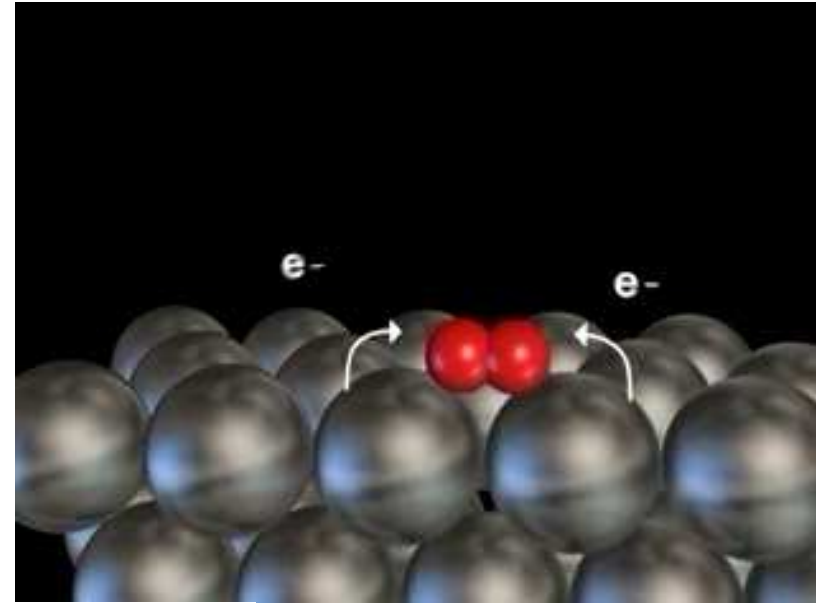


Gas forming neutralization reactions (often involving  $\text{CO}_3^{2-}$ )

# Oxidation-Reduction Reactions

An **oxidation** occurs when an atom or ion *loses* electrons.

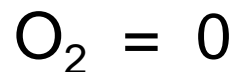
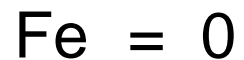
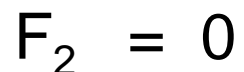
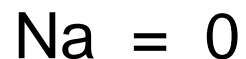
A **reduction** occurs when an atom or ion *gains* electrons.



# Oxidation Number - RULES

To determine if an oxidation-reduction reaction has occurred, we assign an **oxidation number** to each element in a neutral compound or charged entity.

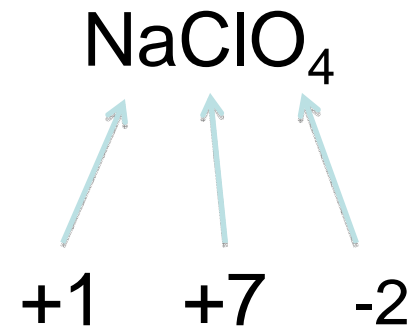
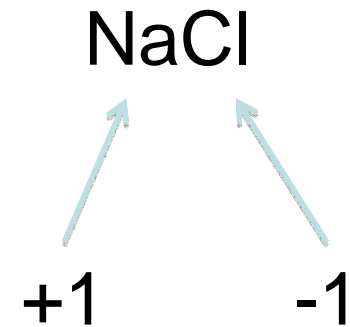
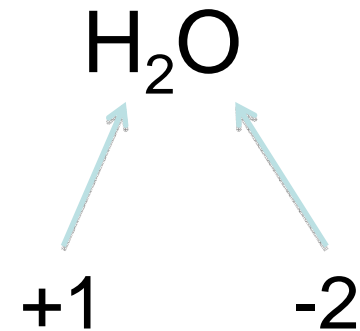
Elements in their elemental form have an oxidation number of 0.  
The oxidation number of a monatomic ion is the same as its charge.



# Oxidation Number - RULES

Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.

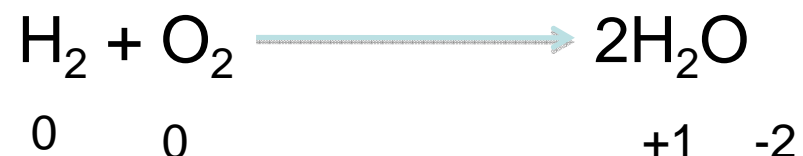
- Oxygen has an oxidation number of  $-2$ , except in the peroxide ion in which it has an oxidation number of  $-1$ .
- Hydrogen is  $-1$  when bonded to a metal,  $+1$  when bonded to a nonmetal.
- Halogens usually have an oxidation number of  $-1$ ; but can have positive oxidation numbers, when in oxyanions.



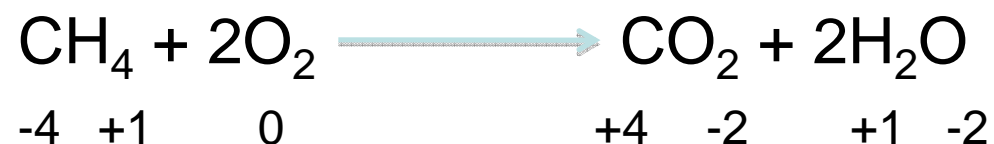
# Oxidation / Reduction Reactions

REDOX RXNS: There is a change in oxidation # from reactant to products. One species will be oxidized and another will be reduced. (the oxidation # of one will increase and another will decrease)

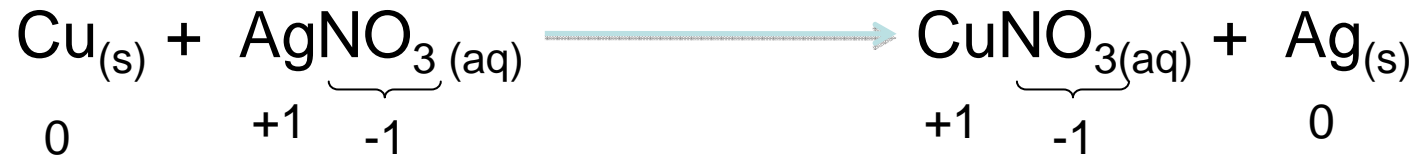
What are the oxidation numbers for H and O as reactant and then as products?



What are the oxidation numbers for C and O as reactant and then as products?



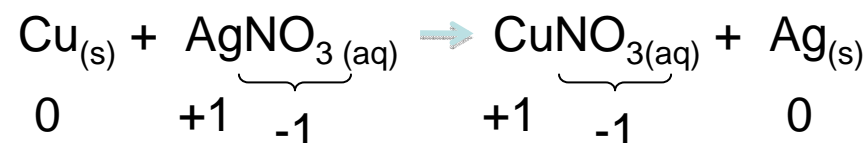
## More examples of REDOX reactions



# Activity Series

| Metal     | Oxidation Reaction                                       |
|-----------|--|
| Lithium   | $\text{Li}(s) \longrightarrow \text{Li}^+(aq) + e^-$     |
| Potassium | $\text{K}(s) \longrightarrow \text{K}^+(aq) + e^-$       |
| Barium    | $\text{Ba}(s) \longrightarrow \text{Ba}^{2+}(aq) + 2e^-$ |
| Calcium   | $\text{Ca}(s) \longrightarrow \text{Ca}^{2+}(aq) + 2e^-$ |
| Sodium    | $\text{Na}(s) \longrightarrow \text{Na}^+(aq) + e^-$     |
| Magnesium | $\text{Mg}(s) \longrightarrow \text{Mg}^{2+}(aq) + 2e^-$ |
| Aluminum  | $\text{Al}(s) \longrightarrow \text{Al}^{3+}(aq) + 3e^-$ |
| Manganese | $\text{Mn}(s) \longrightarrow \text{Mn}^{2+}(aq) + 2e^-$ |
| Zinc      | $\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^-$ |
| Chromium  | $\text{Cr}(s) \longrightarrow \text{Cr}^{3+}(aq) + 3e^-$ |
| Iron      | $\text{Fe}(s) \longrightarrow \text{Fe}^{2+}(aq) + 2e^-$ |
| Cobalt    | $\text{Co}(s) \longrightarrow \text{Co}^{2+}(aq) + 2e^-$ |
| Nickel    | $\text{Ni}(s) \longrightarrow \text{Ni}^{2+}(aq) + 2e^-$ |
| Tin       | $\text{Sn}(s) \longrightarrow \text{Sn}^{2+}(aq) + 2e^-$ |
| Lead      | $\text{Pb}(s) \longrightarrow \text{Pb}^{2+}(aq) + 2e^-$ |
| Hydrogen  | $\text{H}_2(g) \longrightarrow 2\text{H}^+(aq) + 2e^-$   |
| Copper    | $\text{Cu}(s) \longrightarrow \text{Cu}^{2+}(aq) + 2e^-$ |
| Silver    | $\text{Ag}(s) \longrightarrow \text{Ag}^+(aq) + e^-$     |
| Mercury   | $\text{Hg}(l) \longrightarrow \text{Hg}^{2+}(aq) + 2e^-$ |
| Platinum  | $\text{Pt}(s) \longrightarrow \text{Pt}^{2+}(aq) + 2e^-$ |
| Gold      | $\text{Au}(s) \longrightarrow \text{Au}^{3+}(aq) + 3e^-$ |

Easily oxidized



NOT easily oxidized

Any metal can be oxidized by a metal ion lower than it on the table

## Molarity (M)

Molarity is one way to measure the concentration of a solution.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

Example: 6.00 g NaCl dissolved in 0.100L of water, what is the molar concentration?

$$6.00 \text{ g NaCl} \left( \frac{1 \text{ mol}}{58.5 \text{ g}} \right) / (0.1\text{L}) = 1.03 \text{ mol} / \text{L} = \mathbf{1.03 \text{ M}}$$

# Dilution Calculations



(Initial Concentration)(Initial Volume) = (Final concentration)(Final Volume)

$$C_i V_i = C_f V_f$$

# Dilution Example

If I have a 1M NaCl stock solution, and I want to dilute it to make 100 mL of 0.2M NaCl, how much of the stock solution should I use?

$$C_i V_i = C_f V_f$$

$$V_i = C_f V_f / C_i$$

$$V_i = (0.2 \text{ M NaCl}) (100 \text{ mL}) / (1 \text{ M NaCl})$$

$$V_i = 20 \text{ mL of stock}$$