

# Chapter 9

## Chemical Reactions in Aqueous Solutions



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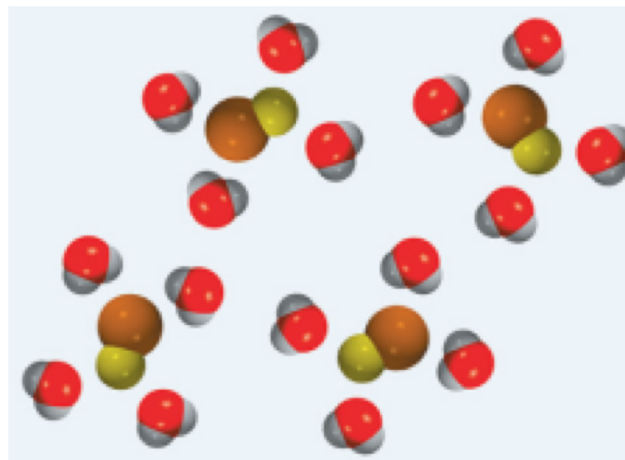
# Solutions: Key Terms

## Solution

- Homogenous mixture of 2 or more substances

## Solvent:

- Component with largest amount
- Water, the “universal solvent”
  - (does not dissolve everything)

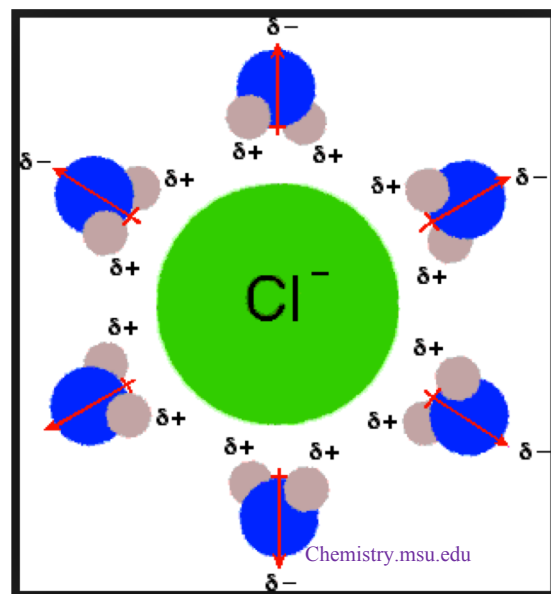


## Solute:

- Components present in smaller amounts than the solvent

## Solvation/dissolving:

- Solvent molecules surround & support solute molecules or ions
- Solvent is NOT part of chemical reactions



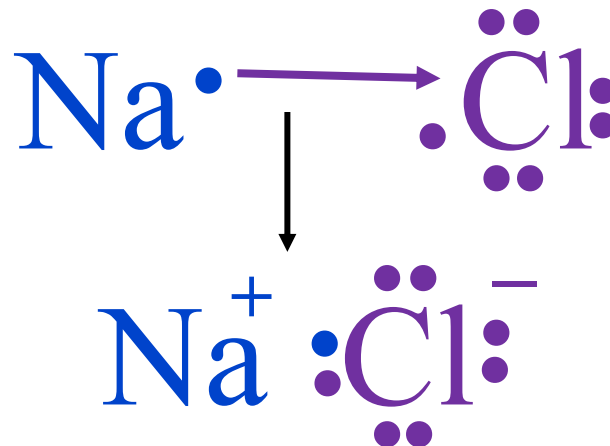
# Solubility Depends on Bond Type

**Electronegativity** – measure of atom's attraction for electrons

- Electronegativity difference determines bonding type

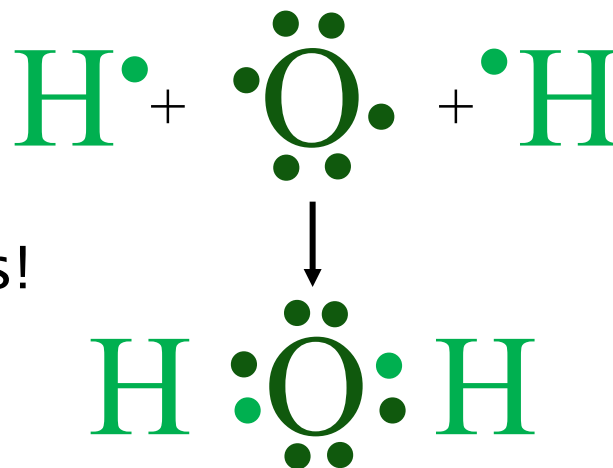
## Ionic:

- Large electronegativity difference
- Metal + nonmetal
- Transfer of electrons
  - Involves ions – have charges!



## Covalent:

- Moderate to small electroneg. diff.
- Often two nonmetals
- Electrons are shared
  - Atoms remain neutral – not ions!
- Two kinds, polar and nonpolar



# Covalent Bonds: Nonpolar vs Polar

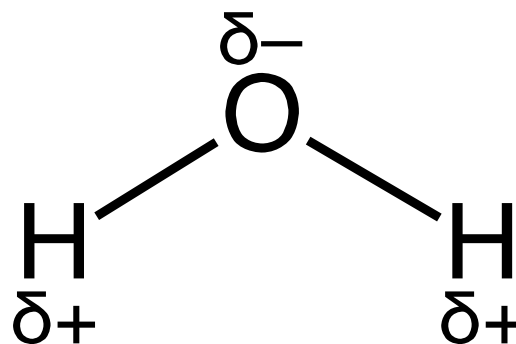
**Nonpolar – electrons shared equally**

- very small/no electroneg. difference
- no partial charges

**Polar – electrons shared unequally**

- moderate electroneg. difference
- **results in partially charged atoms ( $\delta^+$  or  $\delta^-$ )**
- water is polar

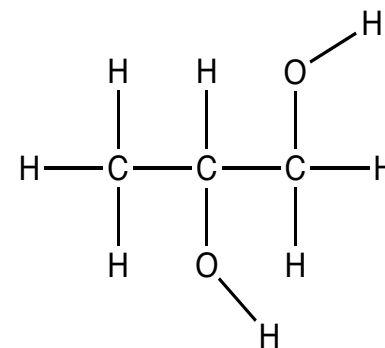
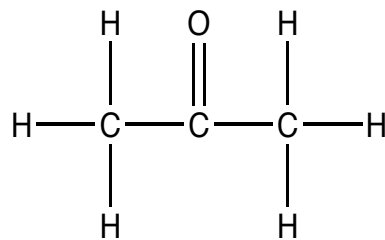
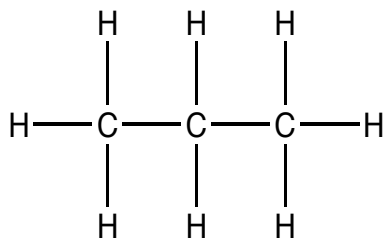
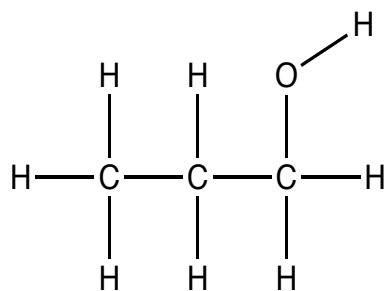
**Like Dissolves Like!**





Rank the following in order of increasing water solubility.

Remember that water is polar & can hydrogen bond  
Greater polarity & hydrogen bonding ability increase  
water solubility.



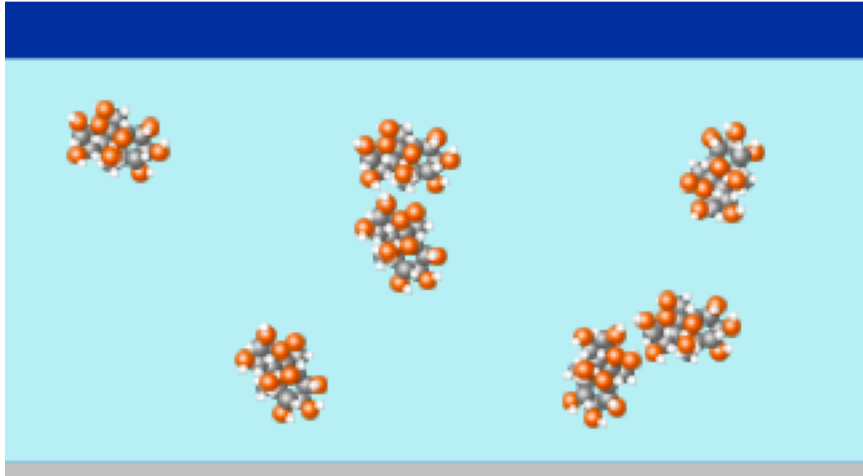
# Dissolution vs Dissociation

## Ionic & Covalent Compounds Behave Differently

Sugar vs Salt

# Dissolution vs Dissociation

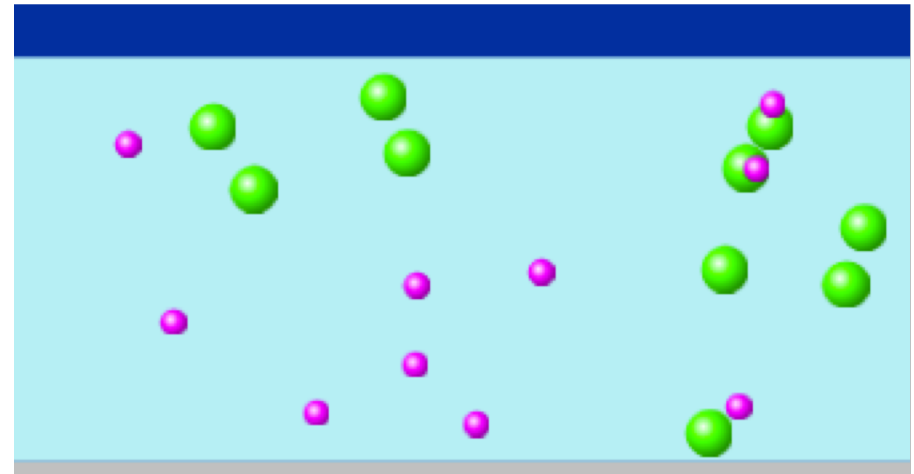
## Covalent Molecule



Dissolved sugar:

- molecule remains together
- NOT dissociation!

## Ionic Salt



Dissolved salt:

- ions separate
- dissociation!

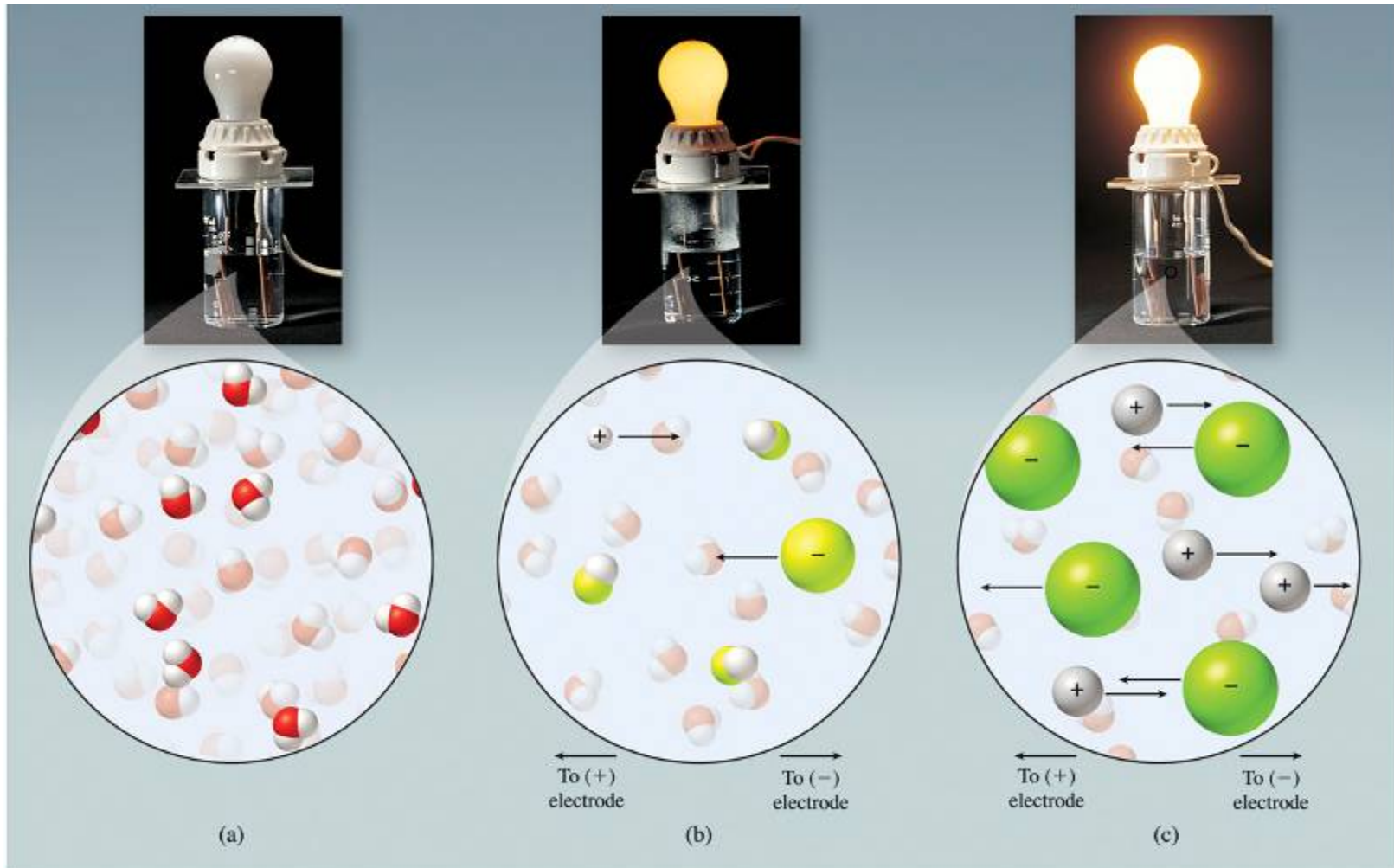
Extent of dissociation/ionization impacts many properties  
– e.g. acid base strength, conductivity, freezing point

# Conductivity of Electrolytes in Aqueous Solutions

Non-electrolyte  
No ionization  
Ex: sugar

Weak electrolyte  
Some ionization  
Ex: acetic acid

Strong electrolyte  
Full ionization  
Ex: NaCl



# Concentration of Solutions: Molarity

**Molarity ( $M$ ) = moles solute/L solution**

- Units: mol/L (molar, M)
- **Conversion factor between moles solute & volume of solution.**

How to prepare 2 liters of a 1.0M solution of NaCl:

1. Calculate mass of NaCl needed.

$$\frac{1\text{mol}_{\text{NaCl}}}{1\text{L}_{\text{NaCl}}} \times \frac{58.5\text{g}_{\text{NaCl}}}{1\text{mol}_{\text{NaCl}}} \times \frac{2\text{L}_{\text{NaCl}}}{1} = 117\text{g}_{\text{NaCl}}$$

2. Weigh out mass of NaCl.

3. Pour NaCl into volumetric flask.

4. Add water until the water reaches the 2L mark.



# Molality

**Molality ( $m$ ) = moles solute / kg solvent**

Units: mol/kg (molal, m)

Based on mass, not volume, therefore:

**NOT TEMPERATURE DEPENDENT**

How to prepare ~2 kg of a 1.0m solution of NaCl:

1. Calculate mass of NaCl needed.

$$\frac{1 \text{ mol}_{\text{NaCl}}}{\text{kg solv.}} \times \frac{58.5 \text{ g}_{\text{NaCl}}}{1 \text{ mol}_{\text{NaCl}}} \times 2 \text{ kg solv.} = 117 \text{ g}_{\text{NaCl}}$$

2. Weigh out mass of NaCl.

3. Place NaCl into container.

4. Tare (zero) the container. Add water until you obtain the desired mass.

# Molarity Calculations:

1. What is the concentration, in moles/L, if 22.8 g of potassium dichromate ( $\text{K}_2\text{Cr}_2\text{O}_7$ ) is dissolved to make 500.0 mL of solution?

A: 0.155 M

2. How many grams of potassium dichromate ( $K_2Cr_2O_7$ ) are needed to prepare 125 mL of a 1.83 M solution?

A: 67.3 g



3. What volume of a 2.50 M solution of potassium dichromate ( $\text{K}_2\text{Cr}_2\text{O}_7$ ) would contain 350.0 g of  $\text{K}_2\text{Cr}_2\text{O}_7$ ?

**A: 476 mL**

# Dilution Calculations

Water is added to a small amount of stock solution to make a less concentrated solution.

Addition of solvent does not change the mass or moles of solute in a solution but does change the solution concentration.



$$M_1V_1 = M_2V_2$$

$$(\text{mol/L})(\text{L}) = (\text{mol/L})(\text{L})$$

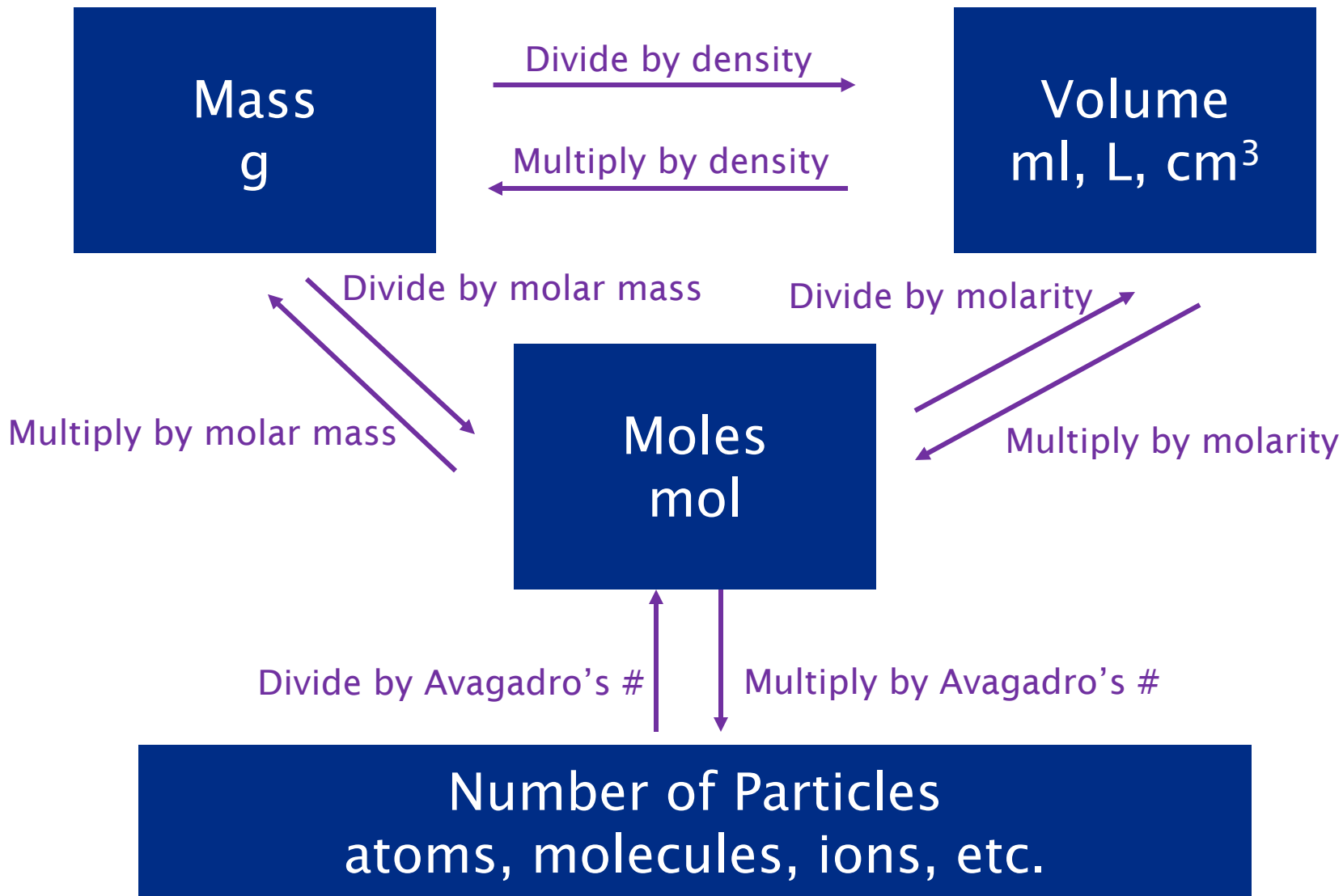


Calculate the volume of 1.0M stock solution needed to make 2000.0mL of a 0.12M solution of HCl.

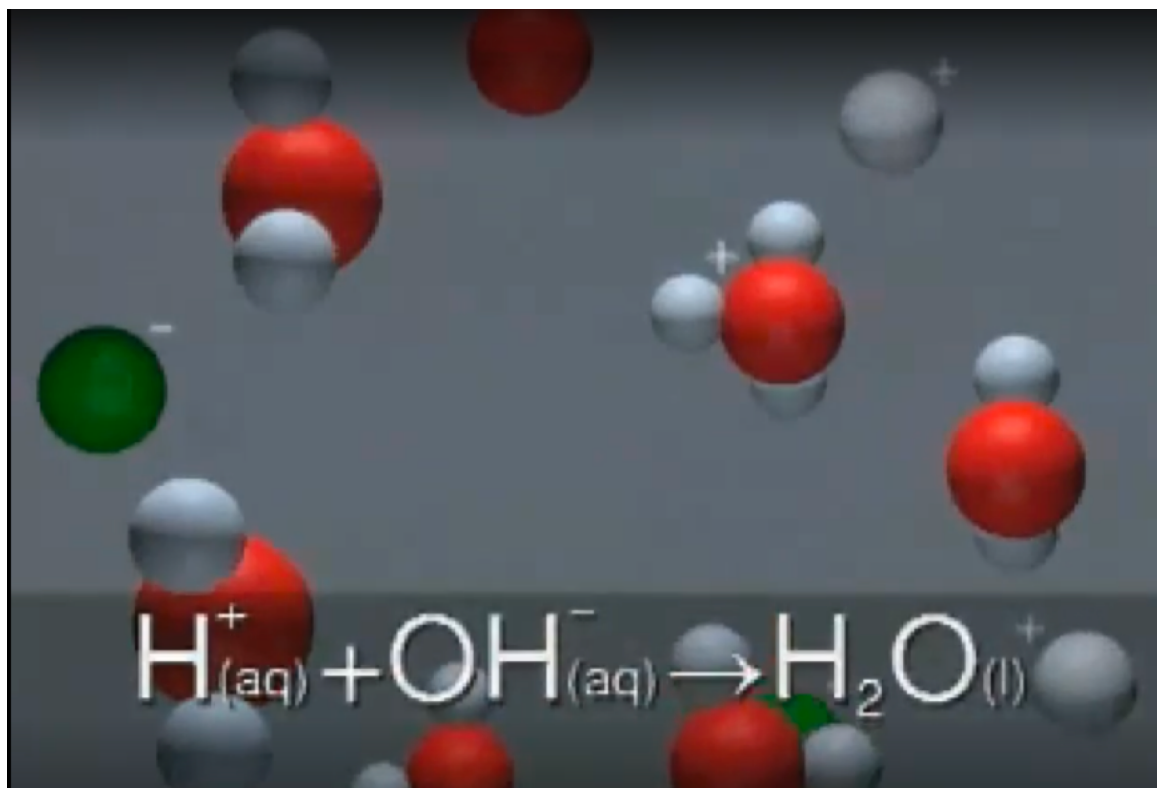
A: 240 mL

**This equation is ONLY for dilutions.  
Never use this equation for titrations.**

# Conversion Relationships



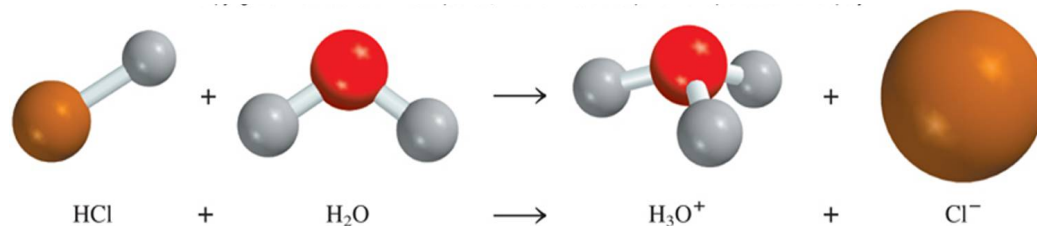
# Acid–Base Reactions



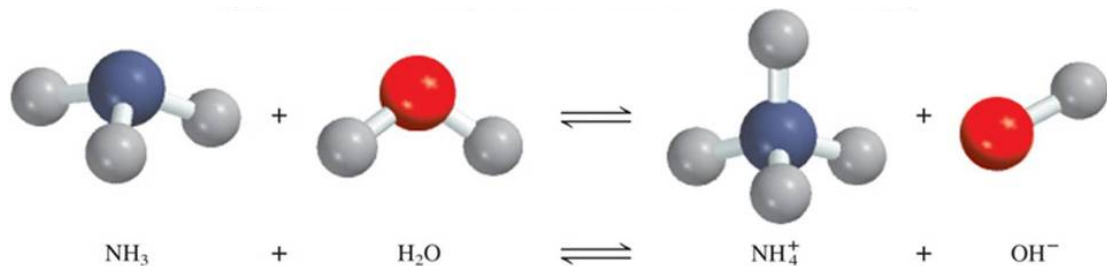
# Acid–Base Reactions

## Arrhenius Acids and Bases

**Acid:** Compound that ionizes in water to form a solution of  $\text{H}^+$  ions ( $\text{H}_3\text{O}^+$ ) and anions.



**Base:** Compound that ionizes in water to form a solution of  $\text{OH}^-$  ions and cations



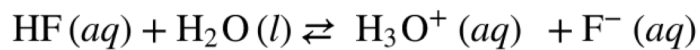
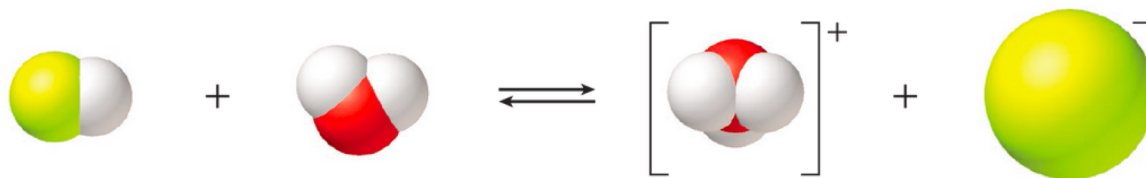
**Neutralization:** Reaction between Arrhenius acid & base  
 $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$  & cation + anion → salt

# Acid–Base Reactions

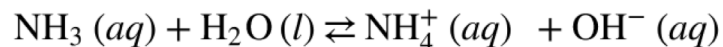
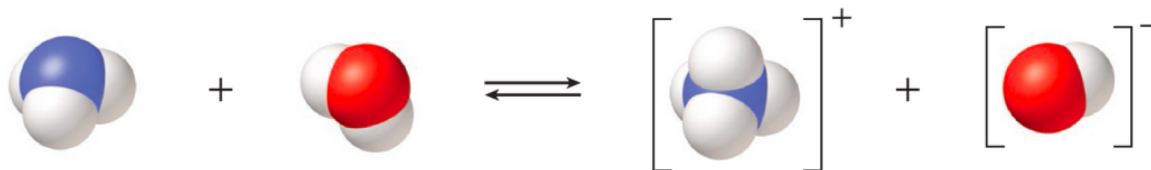
## Brønsted Acids and Bases

A hydrogen atom that has lost its electron is often referred to as a **proton** (because that is all that is left).

**Acid:** Proton donor



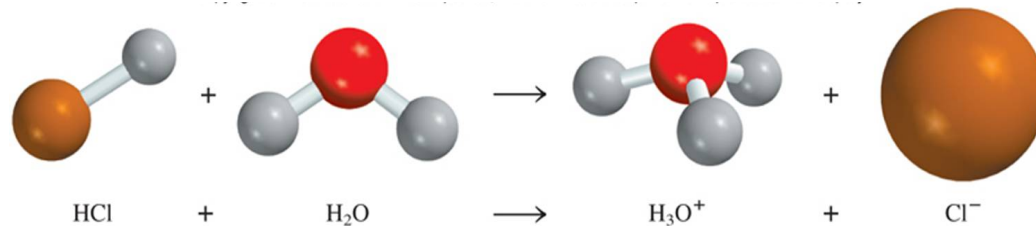
**Base:** proton acceptor



# Strength of Acids and Bases

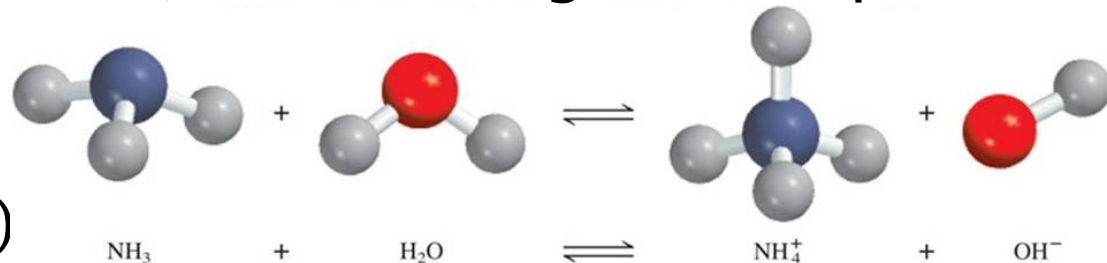
## Strong acids and bases

- Completely ionized in water to give either  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$
- Good conductors of electricity.
- Directional arrow ( $\rightarrow$ ) indicates dissociation is complete



## Weak acids and bases

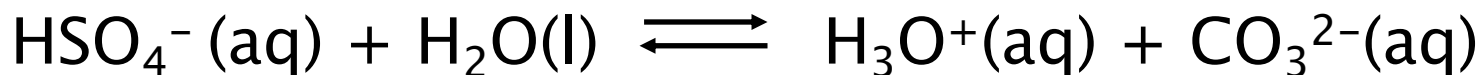
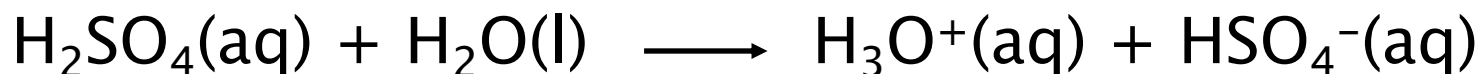
- Partial ionization in water, most of original compound remains
- Poor conductors
- Double arrow ( $\leftrightarrow$ ,  $\rightleftharpoons$ ) indicates dissociation is incomplete



# Polyprotic Acids

## Acids that have more than one ionizable proton

- Monoprotic – one ionizable proton
- Diprotic – two ionizable protons
- Polyprotic – multiple ionizable protons
- Ionize in successive steps

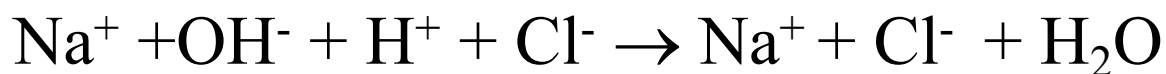
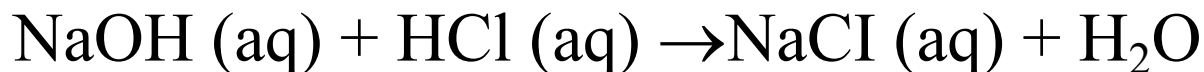


- In this case, the first proton dissociates completely
  - strong
- The second proton is only partially removed
  - weak

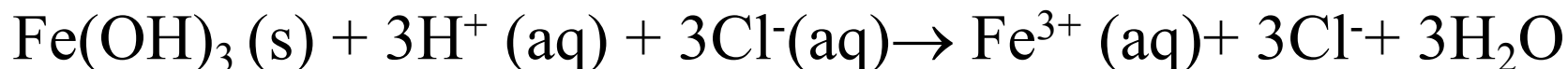
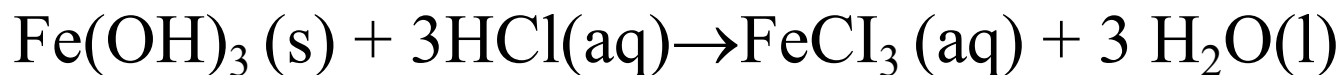


# Acid-Base Neutralizations

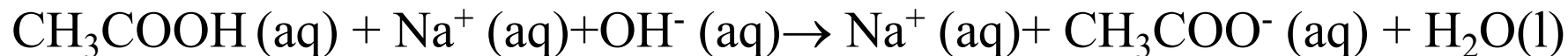
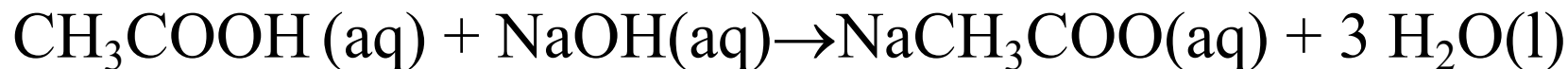
## Strong acid with strong base



## Solid (weak) base with strong acid



## Weak acid with strong base



# Common Acids and Bases

## Strong Acids

HCl, HI, HBr

HNO<sub>3</sub>

H<sub>2</sub>SO<sub>4</sub>

HClO<sub>4</sub>

## Weak Acids

HF, HNO<sub>2</sub>, H<sub>3</sub>PO<sub>4</sub>

Acetic acid (CH<sub>3</sub>COOH)

Organic acids tend to be weak

## Strong Bases

NaOH, KOH,  
LiOH

Ba(OH)<sub>2</sub>

## Weak Bases

Other hydroxides  
(don't dissolve)

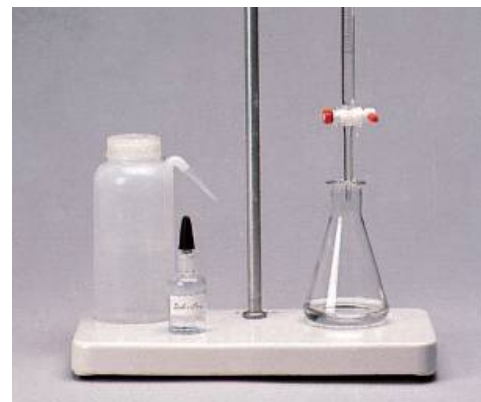
Ammonia: NH<sub>3</sub>

**REMEMBER: FULL IONIZATION = STRONG**

# Simple Acid–Base Titration

## Used to determine acid/base [concentration]

1. To determine [acid], use known basic soln  
To determine [base], use known acidic soln
2. React solution of known concentration w/ measured volume of unknown solution
3. Add known solution dropwise until **endpoint** of reaction
  - Ratio of reactants equals that in balanced equation.
  - **For acid/base: moles  $H^+$  = moles  $OH^-$**
  - Use an indicator to determine endpoint
4. Record volume of known solution needed to reach endpoint
5. Calculate molarity of unknown solution based on initial volume of unknown soln & molarity & volume needed for known soln.



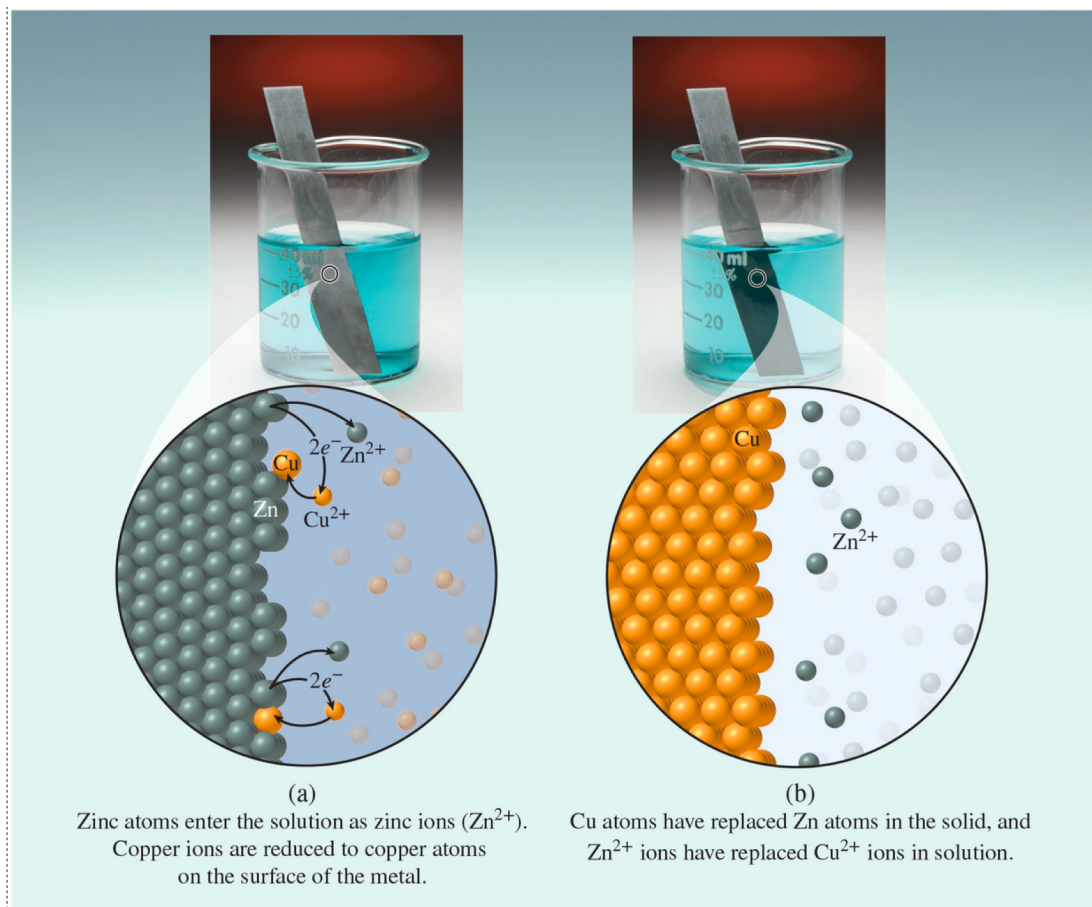
# Simple Acid–Base Titration

25.00–mL of 0.200 M ( $\text{H}_2\text{SO}_4$ ) is titrated with 12.32 ml of a NaOH solution. What is the molarity of the NaOH solution?

- 1) Write the balanced equation
- 2) Determine # moles used of known solution ( $\text{H}_2\text{SO}_4$ )
- 3) Use balanced equation to get moles of unknown (NaOH)
- 4) Divide by volume of unknown to get molarity

$$M_{\text{NaOH}} = 0.812\text{M}$$

# Redox Reactions



# Oxidation–Reduction Reactions

## Oxidation–reduction reactions (REDOX reactions):

- Occur when electrons are transferred from one reactant to another during a chemical reaction.
  - There is a change in oxidation number for both substances

**Oxidation State/oxidation number:** Theoretical charge on atom

**Oxidation** – the process where the oxidation number increases.  
– Electrons are **lost** from the substance

**Reduction** – the process where the oxidation number decreases.  
– Electrons are **gained** by the substance

Oxidation and reduction always accompany each other;  
Neither can occur alone

# LEO the lion says GER



**LEO**

**L**ose  
**E**lectrons  
**O**xidation

**GER**

**G**ain  
**E**lectrons  
**R**eduction

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## OIL RIG



**O**xidation  
**I**s  
**L**oss

**R**eduction  
**I**s  
**G**ain

# Oxidation Number Rules

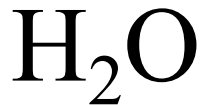
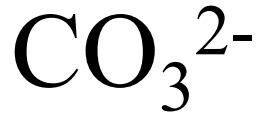
**The rule earlier in the list always takes precedence.**

- 1.) Overall Ox # for a compound is zero
- 2.) Ox # = 0 for an element (not in a compound)  
Ox # = ionic charge for an ion
- 3.) Ox # = +1 for IA elements & H (note: if w/metal H is -1)
- 4.) Ox # = +2 for 2A elements
- 5.) Ox # = -2 for oxygen (usually)
- 6.) Ox # = -1 for 7A elements (If both elements are in 7A, the one higher in the list is -1)
- 7.) Ox # = -2 for 6A elements other than oxygen
- 8.) Ox # = -3 for 5A elements (very shaky!!!)



# Determining Oxidation Numbers

Determine the oxidation number of each element in:



If compound contains polyatomic ions, separate into ions before determining Ox. state of each element.



## Is it REDOX?

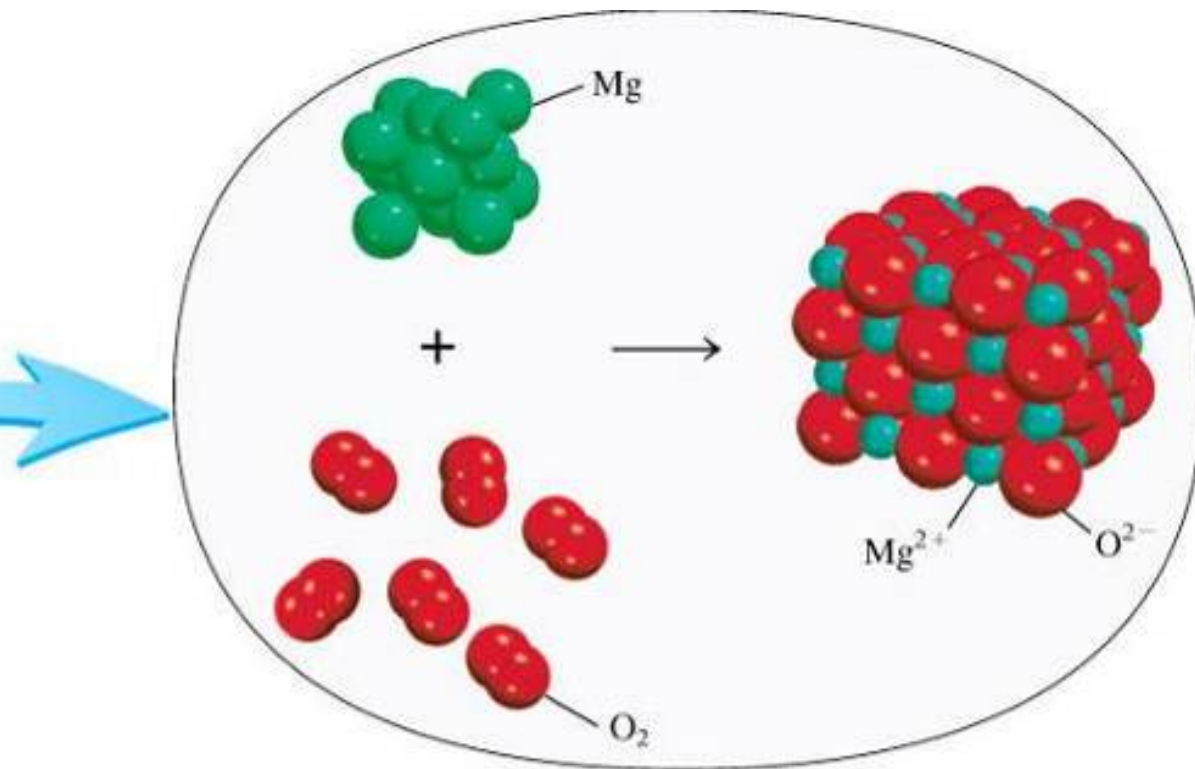


# Redox Reaction: Half-reactions

Oxidation half-reaction:  $\text{Mg (s)} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$

Reduction half-reaction:  $1/2\text{O}_2(\text{g}) + 2\text{e}^- \rightarrow \text{O}^{2-}$

Sum of half-reactions:  $\text{Mg (s)} + 1/2\text{O}_2(\text{g}) \rightarrow \text{MgO(s)}$



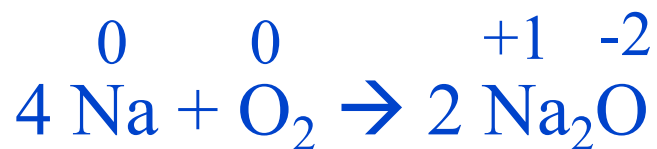
# Oxidizing and Reducing Agents

## Oxidizing agent: reactant that promotes oxidation

- Oxidation = loss of electrons
- Oxidizing agent takes  $e^-$  from other species  $\rightarrow$  is reduced!
- Characteristic of nonmetals: ex: fluorine, oxygen.
- High electron affinity: easily gains electrons

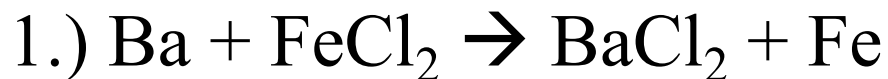
## Reducing agent: reactant that promotes reduction

- Reduction = gain in electrons
- Reducing agent loses  $e^-$   $\rightarrow$  is oxidized!
- Characteristic of an active metal, such as sodium.
- Low ionization energy: easily loses electrons



Na oxidized; is reducing agent  
O reduced; is oxidizing agent

# Writing Half Reactions & Determining Oxidizing & Reducing Agents



oxidation:

reduction:



oxidation:

reduction:

# Balancing Simple Redox Reactions Using Charge

Balance the following redox reaction:



Process:

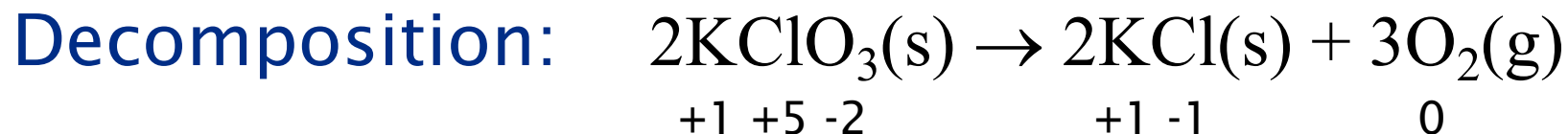
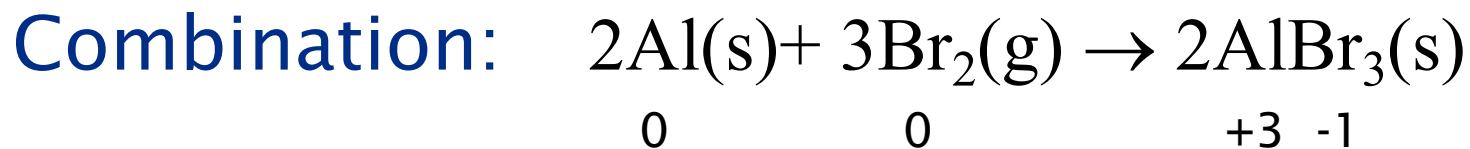
1. Write out the half-reactions
2. Add coefficients to get an equal number electrons in products & reactants
3. Add the two half-reactions and cancel the electrons



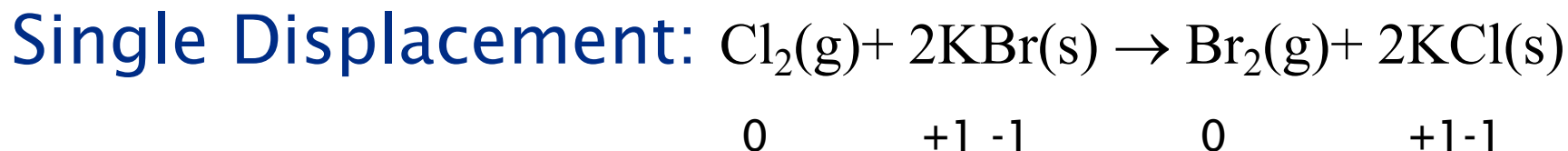
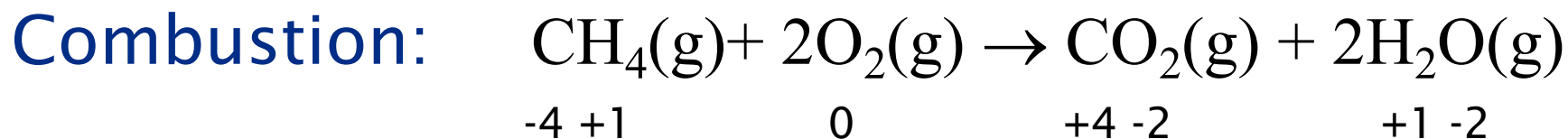
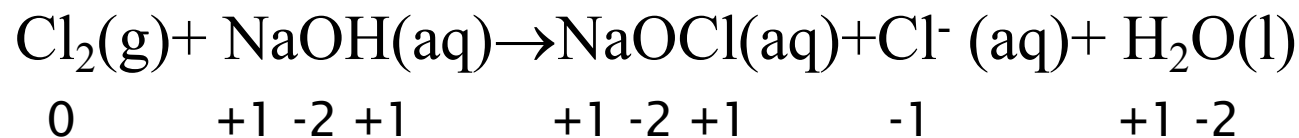
# Elemental Oxidation Numbers

1 1A											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A	
1 <b>H</b> +1 -1											5 <b>B</b> +3	6 <b>C</b> +4 +2 -4	7 <b>N</b> +5 +4 +3 +2 +1 -3	8 <b>O</b> +2 -1 -2	9 <b>F</b> -1	10 <b>Ne</b>	
3 <b>Li</b> +1	4 <b>Be</b> +2											13 <b>Al</b> +3	14 <b>Si</b> +4 -4	15 <b>P</b> +5 +3 -3	16 <b>S</b> +6 +4 +2 -2	17 <b>Cl</b> +7 +6 +5 +4 +3 +1 -1	18 <b>Ar</b>
11 <b>Na</b> +1	12 <b>Mg</b> +2	3 3B	4 4B	5 5B	6 6B	7 7B	8	9	10	11 1B	12 2B	13 <b>Ga</b> +3	14 <b>Ge</b> +4 -4	15 <b>As</b> +5 +3 -3	16 <b>Se</b> +6 +4 -2	17 <b>Br</b> +5 +3 +1 -1	18 <b>Kr</b> +4 +2
19 <b>K</b> +1	20 <b>Ca</b> +2	21 <b>Sc</b> +3	22 <b>Ti</b> +4 +3 +2	23 <b>V</b> +5 +4 +3 +2	24 <b>Cr</b> +6 +5 +4 +3 +2	25 <b>Mn</b> +7 +6 +4 +3 +2	26 <b>Fe</b> +3 +2	27 <b>Co</b> +3 +2	28 <b>Ni</b> +2	29 <b>Cu</b> +2 +1	30 <b>Zn</b> +2	31 <b>Ga</b> +3	32 <b>Ge</b> +4 -4	33 <b>As</b> +5 +3 -3	34 <b>Se</b> +6 +4 -2	35 <b>Br</b> +5 +3 +1 -1	36 <b>Kr</b> +4 +2
37 <b>Rb</b> +1	38 <b>Sr</b> +2	39 <b>Y</b> +3	40 <b>Zr</b> +4	41 <b>Nb</b> +5 +4	42 <b>Mo</b> +6 +4 +3	43 <b>Tc</b> +7 +6 +4	44 <b>Ru</b> +8 +6 +4 +3	45 <b>Rh</b> +4 +3 +2	46 <b>Pd</b> +4 +2	47 <b>Ag</b> +1	48 <b>Cd</b> +2	49 <b>In</b> +3	50 <b>Sn</b> +4 +2	51 <b>Sb</b> +5 +3 -3	52 <b>Te</b> +6 +4 -2	53 <b>I</b> +7 +5 +1 -1	54 <b>Xe</b> +6 +4 +2
55 <b>Cs</b> +1	56 <b>Ba</b> +2	57 <b>La</b> +3	72 <b>Hf</b> +4	73 <b>Ta</b> +5	74 <b>W</b> +6 +4	75 <b>Re</b> +7 +6 +4	76 <b>Os</b> +8 +4	77 <b>Ir</b> +4 +3	78 <b>Pt</b> +4 +2	79 <b>Au</b> +3 +1	80 <b>Hg</b> +2 +1	81 <b>Tl</b> +3 +1	82 <b>Pb</b> +4 +2	83 <b>Bi</b> +5 +3	84 <b>Po</b> +2	85 <b>At</b> -1	86 <b>Rn</b>

# Types of Redox Reactions



Disproportionation (e.g. bleach production):



**No need to memorize – you can always just figure out the oxidation numbers!**



# Redox Titration

Determine the concentration (M) of a potassium permanganate ( $\text{KMnO}_4$ ) solution if 25.32 mLs are needed to react completely with 7.24 g  $\text{Na}_2\text{C}_2\text{O}_4$  (s).

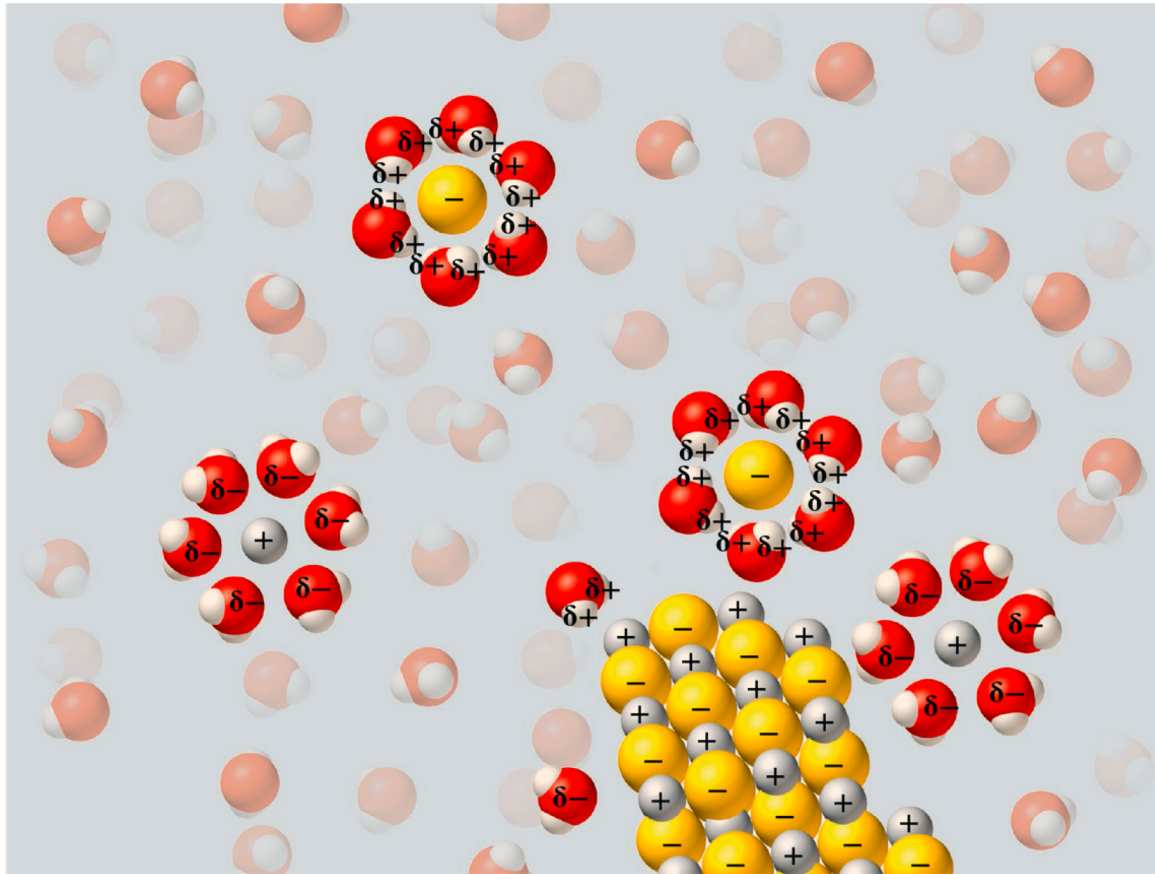


1.) Determine amount in moles of  $\text{KMnO}_4$ . (Stoichiometry!)

2.) Determine concentration of  $\text{KMnO}_4$ . (Molarity!)

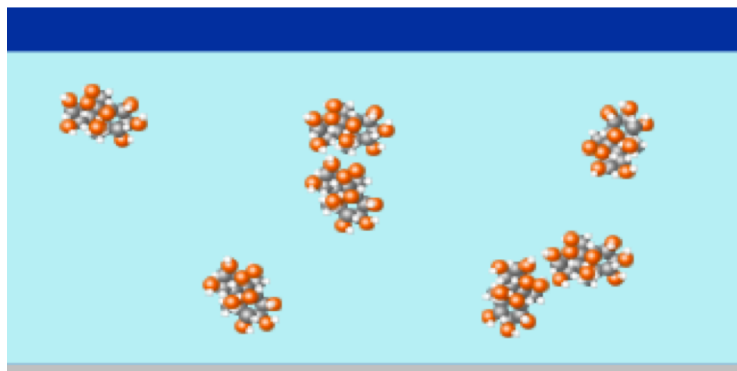
**A: 0.854 M  $\text{KMnO}_4$**

# Ion Concentration & Ionic Equations



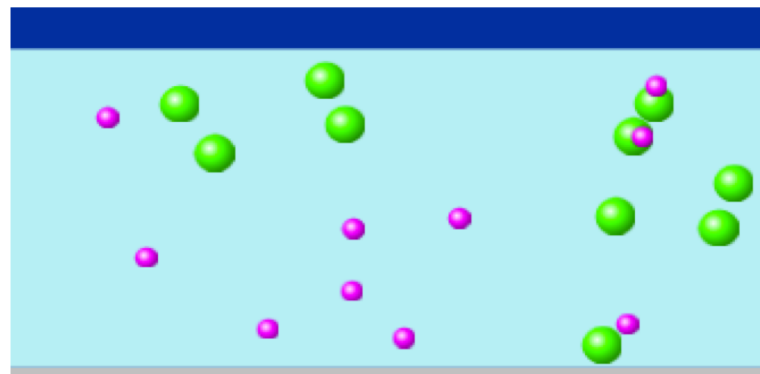
# Total Ion Concentration Can Be Higher Than Concentration of Compound in Solution

## Covalent Molecule



Concentration:  
7 molecules/jar

## Ionic Salt, ex. NaCl



Concentration:  
10 NaCl units/jar  
 $10 \text{ Na}^+ + 10 \text{ Cl}^- = 20 \text{ ions/jar}$

# Calculating Ion Concentrations in Solution

What are the concentrations of aluminum ion, sulfate ion & nitrate ion in a solution that is 1.2 M aluminum sulfate and 1.0 M aluminum nitrate?

A: 3.6 M  $\text{SO}_4^{2-}$ ; 3.0 M  $\text{NO}_3^-$ ;  
3.4 M  $\text{Al}^{3+}$

1. Write down how the salts dissociate in water.
2. Multiply concentration of each material by the number of ions it puts into the solution.
3. Add up ions if there is more than 1 source.

# Mathematical Solubility Problems

What is the molarity of a sodium phosphate solution if adding  $\text{AgNO}_3$  (aq) to 75.0 mL produces 0.205 g  $\text{Ag}_3\text{PO}_4$  (s)?

1. Produce the balanced chemical equation.
2. Determine the amount of  $\text{Ag}_3\text{PO}_4$ (s) in moles.  
MM = 418.5754 g/mol
3. Use moles of  $\text{Ag}_3\text{PO}_4$ (s) to determine the moles  $\text{Na}_3\text{PO}_4$ .
4. Calculate the molarity (mol/L) of  $\text{Na}_3\text{PO}_4$ .

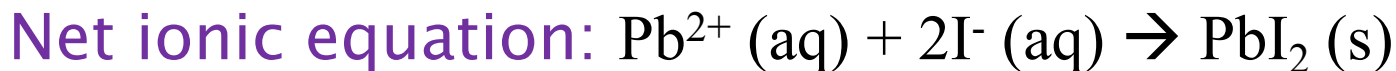
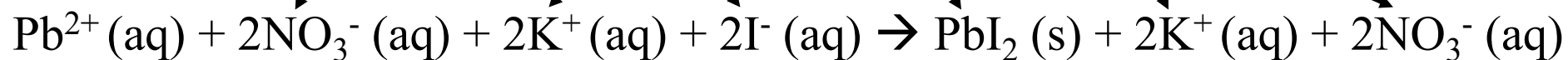
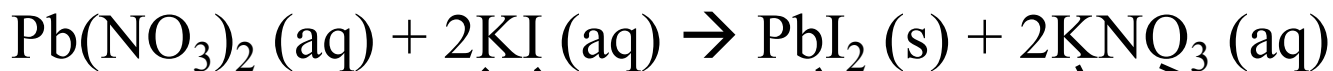
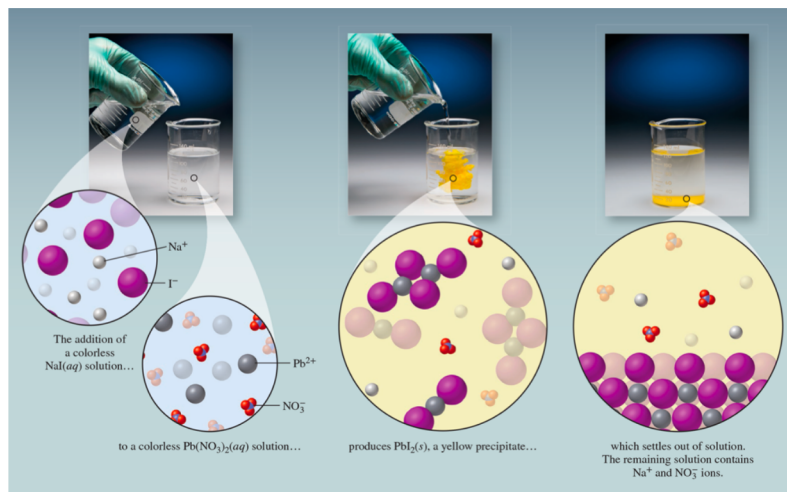
A:  $6.53 \times 10^{-3} \text{ M}$ <sup>41</sup>

# Types of Ionic Reactions: Precipitation Reactions

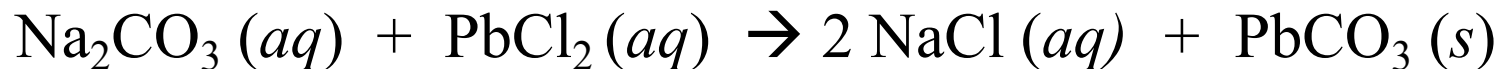
**Precipitation:** Ions in solution combine to form an insoluble solid salt

**Precipitate:** Solid salt that is formed

**Spectator Ions:** Ions that do not react in solution; they remain as ions



# Writing Net Ionic Equations



## Steps:

1. Write ions (all ions = total ionic equation)
2. Cross off species that are the same on both sides
3. Write net ionic equations using only species that have changed (changed: aq on one side, s, l, or g on the other)



# Solubility Rules for ions

Solubility rules classify compounds into those that are usually **soluble** and those that are usually **insoluble**.

**Soluble**

Ammonium ( $\text{NH}_4^+$ )  
Hydrogen ( $\text{H}^+$ )  
Alkali metals (group 1A)  
Nitrate ( $\text{NO}_3^-$ )  
Perchlorate ( $\text{ClO}_4^-$ ) & Chlorate ( $\text{ClO}_3^-$ )  
Acetate ( $\text{CH}_3\text{COO}^-$ )

**Always soluble**

**Usually Soluble**

Halides ( $\text{F}^-$ ,  $\text{Cl}^-$ ,  $\text{Br}^-$ , &  $\text{I}^-$ )

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Sulfate ( $\text{SO}_4^{2-}$ )

**Exceptions (insoluble if with):**  
 **$\text{Pb}^{2+}$ ,  $\text{Hg}_2^{2+}$ ,  $\text{Ag}^+$**

**$\text{Pb}^{2+}$ ,  $\text{Hg}_2^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Ba}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$**

**Sparingly Soluble (Insoluble)**

Sulfide ( $\text{S}^{2-}$ )  
Hydroxide ( $\text{OH}^-$ )  
Oxide ( $\text{O}^{2-}$ )  
Carbonate ( $\text{CO}_3^{2-}$ )  
Phosphate ( $\text{PO}_4^{3-}$ )  
Chromate ( $\text{CrO}_4^{2-}$ )

**Exceptions:**  
**soluble if with any of the cations listed in the always soluble box**



# Determining Products of Precipitation Reactions



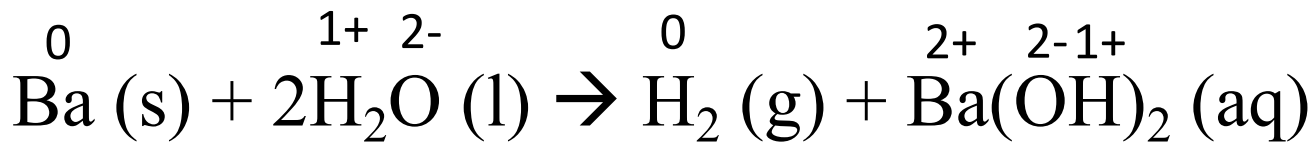
1. Divide reactant compounds into cations & anions
2. Match cation from one salt with anion from other salt  
**Note: Always keep the metal on the left in salts!**
3. Balance charges in salts to generate formulas
4. Write balanced equation
5. Use solubility rules to predict solubility of products



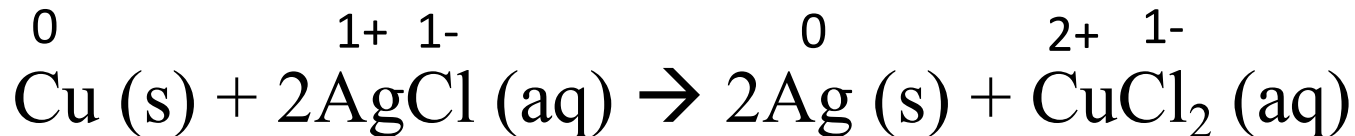
# Single Displacement Reactions



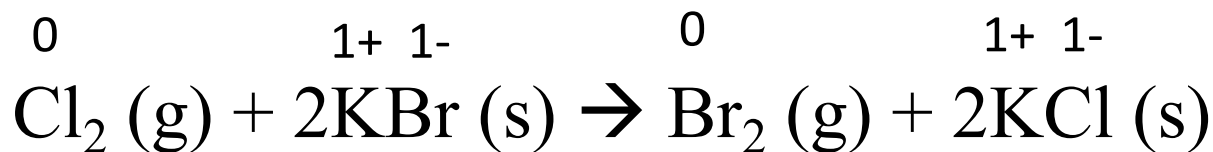
## Hydrogen Displacement



## Metal Displacement



## Halogen Displacement: $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$



Relative Activities with Water & Acid

Reducing strength increases

$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$	
$\text{K} \rightarrow \text{K}^+ + \text{e}^-$	React with cold water to produce $\text{H}_2$
$\text{Ba} \rightarrow \text{Ba}^{2+} + 2\text{e}^-$	
$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$	
$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$	
$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$	
$\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$	
$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$	React with steam to produce $\text{H}_2$
$\text{Cr} \rightarrow \text{Cr}^{3+} + 3\text{e}^-$	
$\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$	
$\text{Cd} \rightarrow \text{Cd}^{2+} + 2\text{e}^-$	
$\text{Co} \rightarrow \text{Co}^{2+} + 2\text{e}^-$	
$\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$	React with acids to produce $\text{H}_2$
$\text{Sn} \rightarrow \text{Sn}^{2+} + 2\text{e}^-$	
$\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$	
$\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$	
$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$	
$\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$	Do not react with water or acids to produce $\text{H}_2$
$\text{Hg} \rightarrow \text{Hg}^{2+} + 2\text{e}^-$	
$\text{Pt} \rightarrow \text{Pt}^{2+} + 2\text{e}^-$	
$\text{Au} \rightarrow \text{Au}^{3+} + 3\text{e}^-$	



Silver Plating



Production of Bromine Gas

# Activity Series

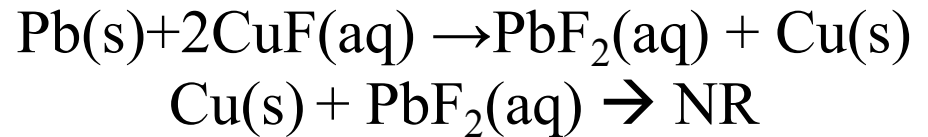
Reducing strength increases

$\text{Li} \rightarrow \text{Li}^+ + e^-$	
$\text{K} \rightarrow \text{K}^+ + e^-$	React with cold water to produce $\text{H}_2$
$\text{Ba} \rightarrow \text{Ba}^{2+} + 2e^-$	
$\text{Ca} \rightarrow \text{Ca}^{2+} + 2e^-$	*Also react w/ steam & acid
$\text{Na} \rightarrow \text{Na}^+ + e^-$	
$\text{Mg} \rightarrow \text{Mg}^{2+} + 2e^-$	
$\text{Al} \rightarrow \text{Al}^{3+} + 3e^-$	
$\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$	React with steam to produce $\text{H}_2$
$\text{Cr} \rightarrow \text{Cr}^{3+} + 3e^-$	
$\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^-$	*Also react w/ acid
$\text{Cd} \rightarrow \text{Cd}^{2+} + 2e^-$	
$\text{Co} \rightarrow \text{Co}^{2+} + 2e^-$	
$\text{Ni} \rightarrow \text{Ni}^{2+} + 2e^-$	React with acids to produce $\text{H}_2$
$\text{Sn} \rightarrow \text{Sn}^{2+} + 2e^-$	
$\text{Pb} \rightarrow \text{Pb}^{2+} + 2e^-$	
$\text{H}_2 \rightarrow 2\text{H}^+ + 2e^-$	
$\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^-$	
$\text{Ag} \rightarrow \text{Ag}^+ + e^-$	
$\text{Hg} \rightarrow \text{Hg}^{2+} + 2e^-$	Do not react with water or acids to produce $\text{H}_2$
$\text{Pt} \rightarrow \text{Pt}^{2+} + 2e^-$	
$\text{Au} \rightarrow \text{Au}^{3+} + 3e^-$	

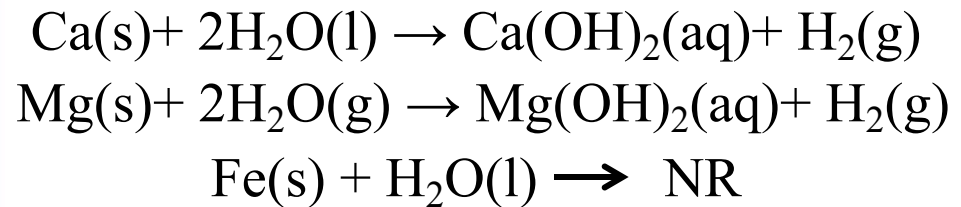
## Reactivity of 2 metals

Higher metal replaces lower metal

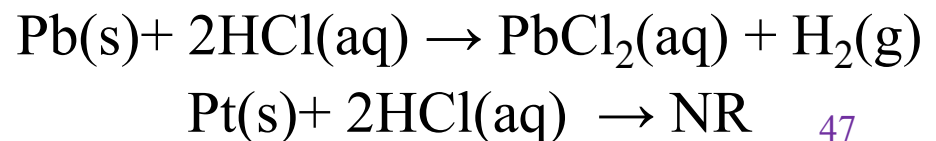
Higher metal becomes cation  
Lower metal will be free metal



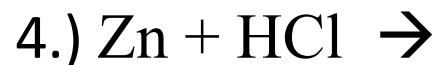
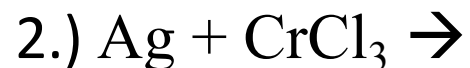
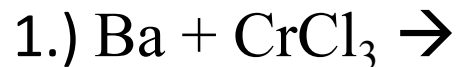
## Reactivity with water



## Reactivity with acid



# Determining Products of Single Displacement Reactions



**Must Know These Diatomic Molecules:**

**$\text{H}_2, \text{F}_2, \text{Cl}_2, \text{Br}_2, \text{I}_2, \text{N}_2, \text{O}_2$**

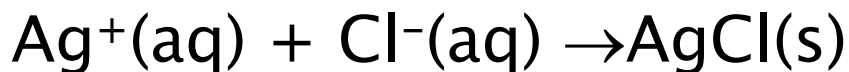
# Solution Based Experiments: Gravimetric Analysis

What is the mass % Cl in a sample of unknown composition?

1. Dissolve known mass of an unknown sample in water.



2. React unknown with  $Ag^+$  to form a precipitate.



3. Filter, dry, & weigh precipitate.

1.0882g  $AgCl(s)$  recovered

4. Use stoichiometry to determine moles and mass of chlorine, then determine %Cl.



**A: 47.54%Cl**