Stoichiometry: amounts of substances in balanced chemical reactions

___ C_8H_{18} + ___ O_2 → ___ CO_2 + ___ H_2O

___ C_8H_{18} ___ O_2 ___ CO_2 ___ H_2O
molecules + molecules → molecules + molecules

___ mol ___ mol ___ mol ___ mol
C_8H_{18} + O_2 → CO_2 + H_2O

How many moles of CO_2 can be produced from the combustion of 8.7 mol octane (C_8H_{18})?
(Use coefficients from balanced chemical equation to make a mole ratio)

How many grams of CO_2 can be produced from the combustion of 4.50 x 10^4 g of octane?

Whenever amounts of more than one reactant are known, you must find the limiting reactant

- Which reactant is consumed first?
- Calculate moles of a single product that each reactant will make
- Reactant that makes fewer product moles is the limiting reactant.
- That's how many product moles can be formed.

$Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$

If 0.30 mol Zn react with 0.52 mol HCl, how many mol H_2 can be formed?

To calculate amount of leftover reactant, first calculate amount of excess reactant that was actually used, then subtract the given amount.

If 7.36 g Zn react with 6.45 g S_8, how many grams ZnS can be produced? What mass of reactant remains?

$8\ Zn + S_8 \longrightarrow 8\ ZnS$

Theoretical yield: calculated product mass from stoichiometry calculation

Actual yield: measured product mass from experiment

$$\textbf{\% Yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\%$$

Most chemical reactions in this course take place in solution (dissolved in water)

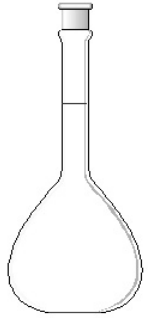
$$\text{Any concentration} = \frac{\text{amount solute}}{\text{amount solution}}$$

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

For instance, a 0.50 M $AgNO_3(aq)$ solution is called a 0.50 "molar" silver nitrate solution

1 liter of this solution contains ____ moles of $AgNO_3$

How do you prepare 250.0 mL of a 0.10 M $AgNO_3(aq)$ solution? (Use the given molarity as a conversion factor between moles solute and liters solution)



volumetric flask

7.5 g $CuCl_2$ are dissolved in water to make 500.0 mL of solution. What is the molar concentration of this solution?

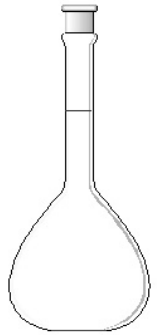
A reaction calls for 0.241 g K_2CO_3. How many mL of 0.125 M $K_2CO_3(aq)$ should be added?

Dilution: solvent is added to make a solution more dilute (less concentrated)

When adding water to a solution, what happens to the number of moles of **solute**?

Dilution equation: $M_1V_1 = M_2V_2$

Concentrated sulfuric acid is 18.0 *M* $H_2SO_4(aq)$. How do you make 10.0 L of 1.50 M $H_2SO_4(aq)$ by dilution?



Molarity converts between…

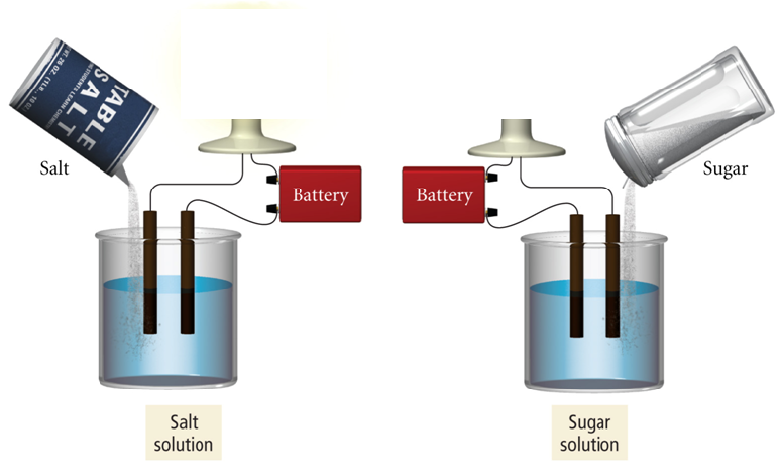
So, molarities can be used in stoichiometry problems along with the mole ratio from the balanced chemical equation.

How many grams of lead(II) iodide can be formed by mixing 1.0 mL of 0.50 M lead(II) nitrate solution with 2.0 mL 0.30 M sodium iodide solution?

$Pb(NO_3)_2 + 2\ NaI \longrightarrow PbI_2 + 2\ NaNO_3$

Electrolyte: solute that causes solution to conduct electricity

Nonelectrolyte: solute that does not caused solution to conduct electricity

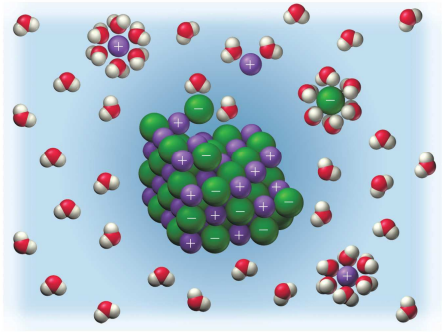


Solute	**Electrolyte?**	**Nonelectrolyte?**
Deionized water		
NaCl		
$C_{12}H_{22}O_{11}$ (sugar)		
NH_4Cl		
C_3H_6O		

Electrolytes **dissociate into ions** when dissolving

$NaCl(s) \rightarrow$

Dissolution of an Ionic Compound



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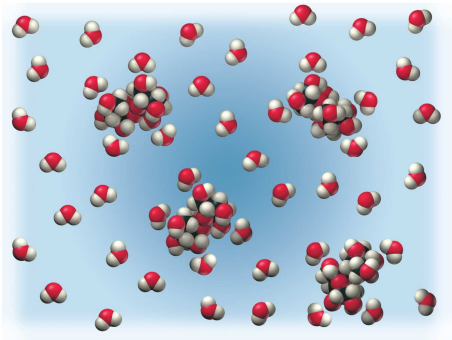
Nonelectrolytes **remain as neutral molecules** when dissolving

Sugar Solution



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Acids are molecular substances that dissociate when dissolved to release H^+ ions.

Strong electrolytes dissociate completely into ions:

$NaCl(s) \rightarrow$

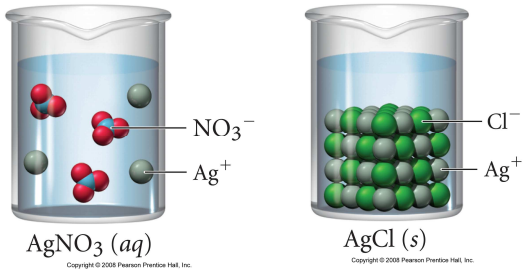
$HCl(g) \rightarrow$

Weak electrolytes do not completely dissociate into ions:

$HCH_3CO_2(aq) \rightleftharpoons H^+(aq) + CH_3CO_2^-(aq)$

Not all ionic compounds dissolve in water!

- The ones that **do** are strong electrolytes
- The ones that **do not dissolve** remain solid in their ionic lattice when added to water

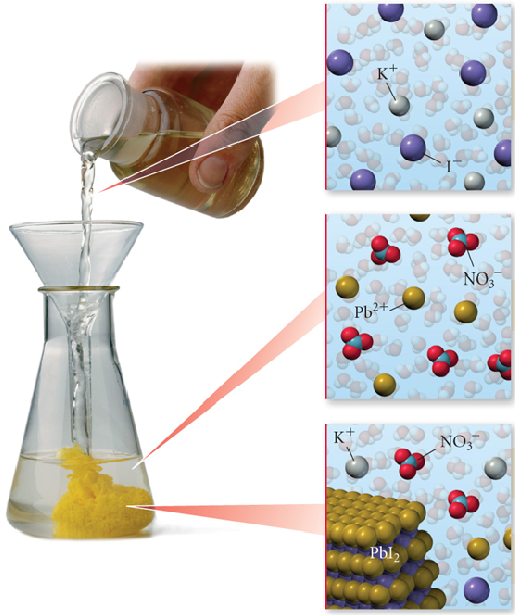


Solubility Rules for Ionic Compounds
The following table will be given on the exam **without** the formulas in parentheses.

Compounds Containing the Following Ions Are Mostly Soluble	**Exceptions**
Li^+, Na^+, K^+, NH_4^+	None
nitrate (NO_3^-), acetate ($C_2H_3O_2^-$)	None
chloride (Cl^-), bromide (Br^-), iodide (I^-)	When any of these ions pairs with Ag^+, Hg_2^{2+}, or Pb^{2+}, the compound is insoluble
sulfate (SO_4^{2-})	When sulfate pairs with Sr^{2+}, Ba^{2+}, Pb^{2+}, or Ca^{2+} the compound is insoluble
Compounds Containing the Following Ions Are Mostly Insoluble	**Exceptions**
hydroxide (OH^-), sulfide (S^{2-})	When either of these ions pairs with Li^+, Na^+, K^+, or NH_4^+, the compound is soluble
	When sulfide (S^{2-}) pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the compound is soluble
	When hydroxide (OH^-) pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the compound is slightly soluble (for many purposes, these may be considered insoluble)
carbonate (CO_3^{2-}), phosphate (PO_4^{3-})	When either of these ions pairs with Li^+, Na^+, K^+, or NH_4^+, the compound is soluble

Precipitation reaction: two ionic compounds trade their ions to produce a solid, insoluble product

Write the balanced chemical equation with phase labels for reaction of solutions of potassium iodide and lead(II) nitrate.



If all products are soluble, **NO REACTION** occurs!

Molecular equation: balanced chemical equation showing neutral formulas for all reactants and products.

$$Ca(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow CaSO_4(s) + 2\ NaNO_3(aq)$$

Complete ionic equation: all strong electrolytes are written as separate ions with their own coefficients and phase labels. (Pure solids, liquids, gases, weak electrolytes, and nonelectrolytes written as molecules.)

Spectator ions do not participate in the chemical reaction (shown the same on both sides)

Net ionic equation: omits spectator ions - only shows the chemical change that has occurred

(What happens in the net ionic equation when it's considered "no reaction?" Try it for the combination of NaI and $CaNO_3$)

Acid: produces H^+ ions in aqueous solutions
Base: produces OH^- ions in aqueous solutions

Strong acids dissociate completely (they are strong electrolytes) - there are 6 common ones:

- HCl
- HBr
- HI
- HNO_3
- H_2SO_4
- $HClO_4$

Weak acids are other formulas that start with H such as:

- HF
- $HC_2H_3O_2$

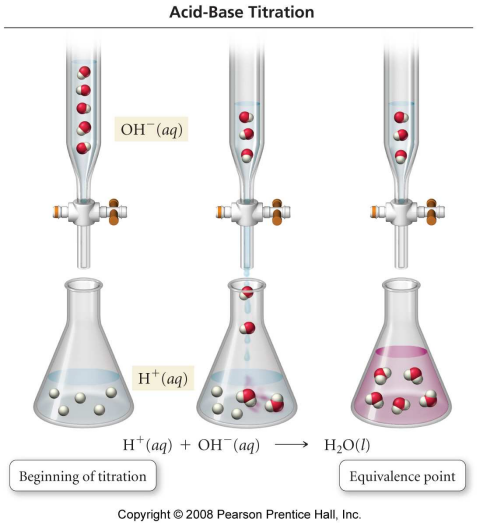
Strong bases are the soluble or slightly soluble hydroxides: NaOH, LiOH, KOH, $Ca(OH)_2$, $Ba(OH)_2$

Acid-base neutralization reactions are the combination of an acid and a base to produce water and a salt

Write an equation for the acid-base neutralization reaction of nitric acid and calcium hydroxide

Titration: controlled addition of one reactant (with known concentration) to another to determine an unknown concentration

$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$
Net: $H^+ + OH^- \rightarrow H_2O$



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Equivalence point: reaction is complete - equal moles of both reactants completely react with each other

Acid-base indicator: changes color to signal endpoint (phenolphthalein changes pink in excess OH^-)

25.0 mL of HCl(*aq*) of unknown concentration were titrated with 12.45 mL of 0.100 M NaOH(*aq*). What is the molarity of the HCl(*aq*)?

$HCl(aq) + NaOH(aq) \rightarrow$

Sulfides, sulfites, and carbonates form a gaseous product when reacting with an acid

$HCl(aq) + Na_2S(aq) \rightarrow$

$HCl(aq) + Na_2SO_3(aq) \rightarrow H_2SO_3(aq) + NaCl(aq)$

$HCl(aq) + Na_2CO_3(aq) \rightarrow$

Oxidation-reduction (redox) reaction: electrons are transferred from one reactant to the other

Oxidation numbers: keep track of electrons in a reaction

Rules for Assigning Oxidation Numbers (must be memorized)

Rule	Applies to	Statement
1	Elements	The oxidation number of an atom in an element is zero
2	Monatomic ions	The oxidation number of an atom in a monatomic ion equals the charge on the ion.
3	Oxygen	The oxidation number of oxygen is −2 in most of its compounds. (An exception is O in H_2O_2 and other peroxides, where the oxidation number is −1.)
4	Hydrogen	The oxidation number of hydrogen is +1 in most of its compounds. (The oxidation number of hydrogen is −1 in binary compounds with a metal, such as CaH_2.)
5	Halogens	The oxidation number of fluorine is −1 in all of its compounds. Each of the other halogens (Cl, Br, I) has an oxidation number of −1 in binary compounds, except when the other element is another halogen above it in the periodic table or the other element is oxygen.
6	Compounds and ions	The sum of the oxidation numbers of the atoms in a compound is zero. The sum of the oxidation numbers of the atoms in a polyatomic ion equals the charge on the ion.

What is the oxidation number of sulfur in Na_2SO_4?

Oxidation: loss of electrons (oxidation # increases)
Reduction: gain of electrons (oxidation # decreases)

Oxidizing agent: the reactant that is itself reduced
Reducing agent: the reactant that is itself oxidized

Identify the elements that are oxidized and reduced, and identify the oxidizing agent and reducing agent.

$2\ Na(s) + Cl_2(g) \longrightarrow 2\ NaCl(s)$

Combustion reaction: an oxidation reduction reaction with O_2 as a reactant.

Carbon-containing compounds undergo combustion to form CO_2 and H_2O

Write the equation for the combustion of C_6H_6

Metals undergo combustion to form metal oxides.