# Ch 7. Chemical Reactions

# Chemical Equations

# Chemical Reactions

Chemical reaction: a process that involves the rearrangement of the ways atoms are grouped together

| Methane | | Oxygen | | Carbon dioxide | | Water |
|---|---|---|---|---|---|---|
| $CH_4$ | $+$ | $O_2$ | $\rightarrow$ | $CO_2$ | $+$ | $H_2O$ |

# Evidence for Chemical Reactions

**Table 6.1** Some Clues That a Chemical Reaction Has Occurred

1. The color changes.
2. A solid forms.
3. Bubbles form.
4. Heat and/or a flame is produced, or heat is absorbed.

# Evidence for Chemical Reactions

What is the clue that a chemical reaction has occurred when when a solution of sodium dichromate is added to a solution of lead nitrate?

a) A gas forms.
b) A solid forms.
c) Bubbles are present.
d) A flame is produced.

# Chemical Equations

Chemical Equations: A way to represent chemical reactions

Reactants → Products

(left) (right)

Methane + Oxygen → Carbon dioxide + Water

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

Reactants: $CH_4$, $O_2$

Products: $CO_2$, $H_2O$

# Chemical Equations

- In a chemical reaction, there is a <u>rearrangement</u> in the way atoms are grouped.
- Atoms are neither destroyed nor created in a chemical reaction (law of conservation of mass).
- Thus, there must be the same number of each type of atom on both sides of the arrow:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Balanced chemical equation

# Chemical Equations

Physical states of compounds are often given in a chemical equation.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

| Symbol | State |
|---|---|
| (*s*) | solid |
| (*l*) | liquid |
| (*g*) | gas |
| (*aq*) | dissolved in water (in aqueous solution) |

# Steps to Writing Balanced Chemical Equations

1. Write unbalanced equation: What are the reactants, the products, correct formulas?

   $CH_4 + O_2 \rightarrow CO_2 + H_2O$

2. Balance the equation: Use coefficients to balance number of each type of atom on both sides.

   (*Don't change formulas of reactants or products!)

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

3. Check coefficients: Do they give balanced number of atoms? ✓ Do they give smallest whole number coefficients possible? ✓

4. Specify physical states.

   $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

# Example: Balancing Chemical Equations

Experimental observation 1: We observe that when hydrogen gas and oxygen gas are mixed together, water forms.

1. Write unbalanced equation for what we observe.

   Which are reactants, which are products, what are physical states?

$$H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(l)$$

# Example, cont'd

2. Balance equation.
   - Don't change formulas of molecules!
   - Balance equation only by adding more reactants and/or products.

$H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(l)$

Count atoms: (2H, 2O) (2H, 1O) Not balanced

$H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$

Count atoms: (2H, 2O) (4H, 2O) Not balanced

$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$

Count atoms: ( 4H, 2O) (4H, 2O) ✓ BALANCED

# Example, continued

Balanced: $2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$

- There are many possible balanced equations:

Multiply by 2: $4H_2\,(g) + 2\,O_2\,(g) \rightarrow 4H_2O\,(l)$

Multiply by 3: $6H_2\,(g) + 3O_2\,(g) \rightarrow 6H_2O\,(l)$

Multiply by 4: $8H_2\,(g) + 4O_2\,(g) \rightarrow 8H_2O\,(l)$

- Best balanced equation is one with smallest whole number coefficients:

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$$

Ex Probs

# Aqueous Reactions

# Reactions in Aqueous Solutions

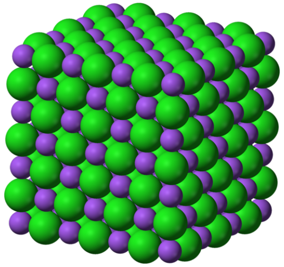
1. Precipitation Reactions: Formation of solid
2. Acid-Base Reactions: Neutralization
3. Redox Reactions: Transfer of electrons

# Electrolytes

Many reactions in water are reactions of electrolytes.

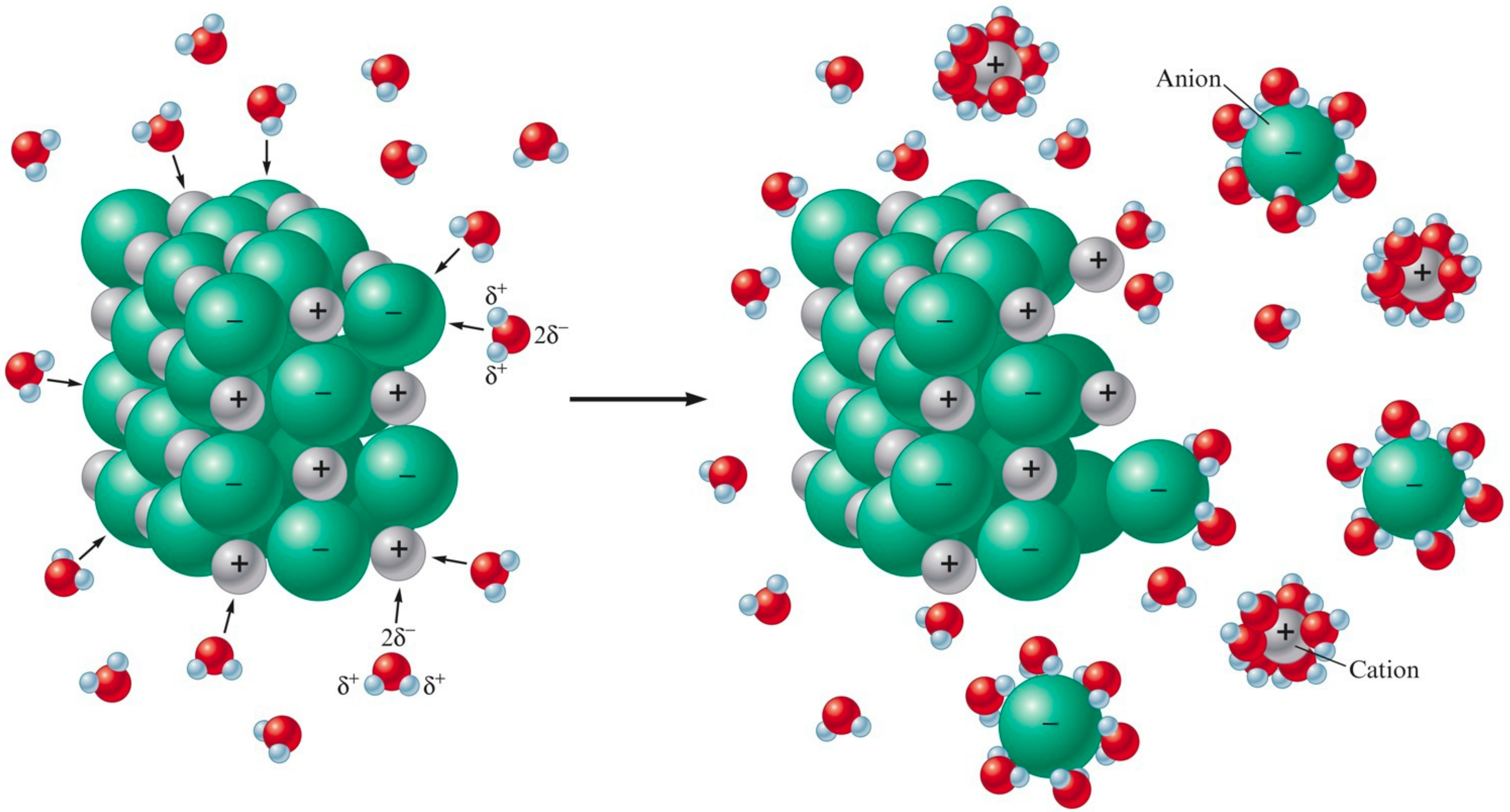
Electrolyte: a substance that dissolves in water to produce separated, free-moving ions, giving an electrically conductive solution

e.g., NaCl dissolved in water.



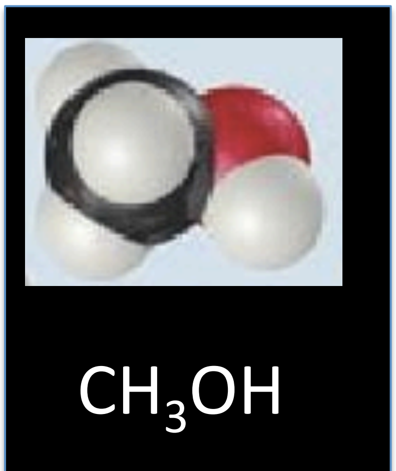
$$NaCl\,(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

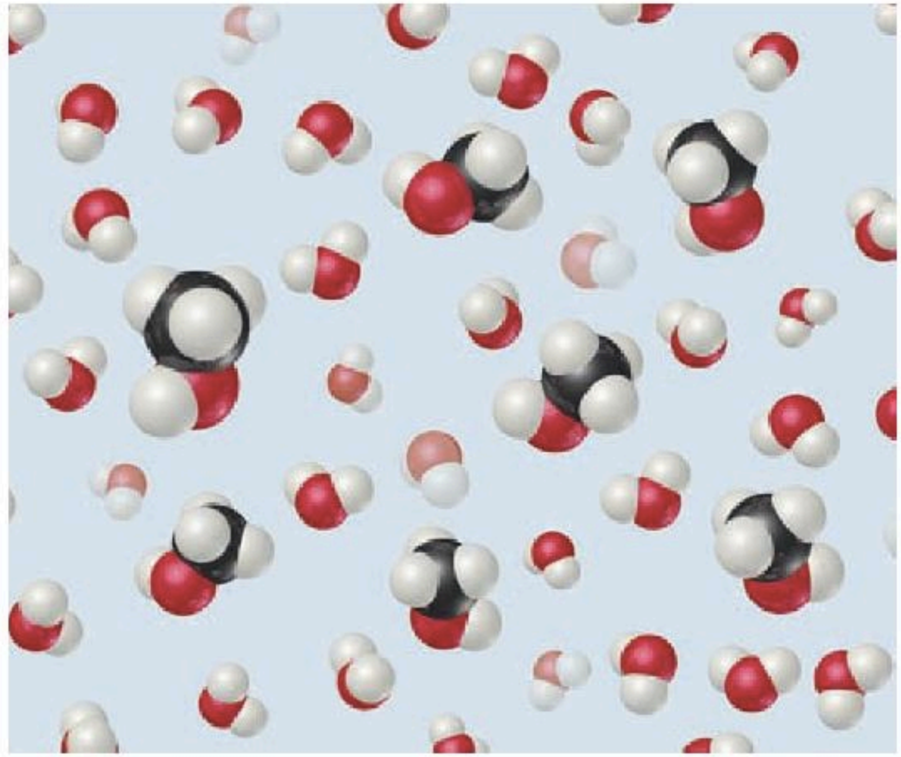
# Electrolyte Dissolved in Water




Each ion is separated through solvation by water (hydration), and the ionic bonds between cations and anions in the solid are broken.

# Nonelectrolyte Dissolved in Water

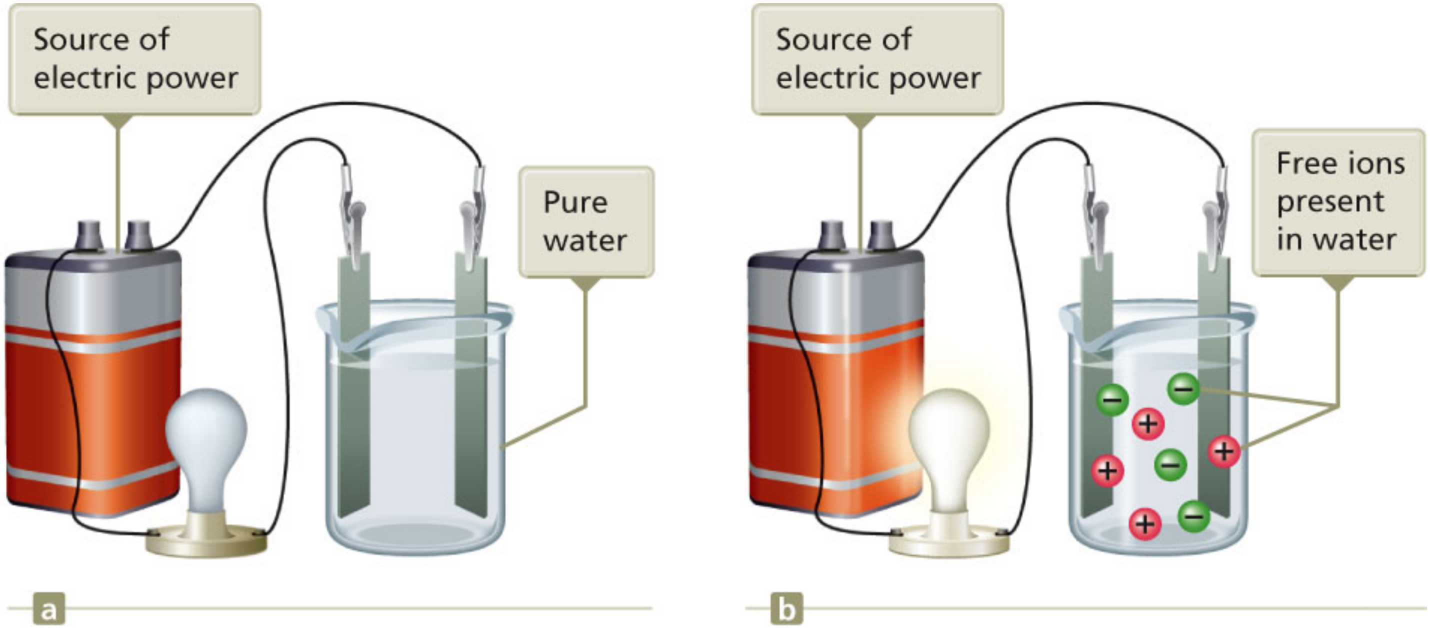


Solvation of whole molecule by water, but NO ionization
(thus <u>soluble in water</u>, but NOT an electrolyte)

# Electrolytes Conduct Electricity

- A solution of <u>free-moving, separated ions</u> can conduct electricity.




Pure water (Nonelectrolyte)

Aqueous NaCl solution (Electrolyte)

# Precipitation Reactions

Precipitation Reactions

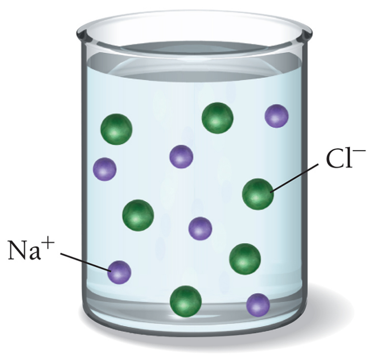
# Precipitation Reactions

- Precipitation Reactions: reactions that result in the formation of an insoluble product (the precipitate)

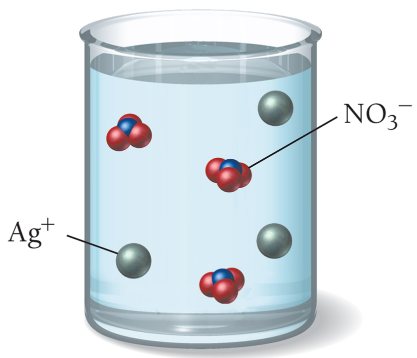
- *Soluble* – solid dissolves in solution; physical state (*aq*)
- *Insoluble* – solid does not dissolve in solution; physical state (*s*)

# Solable vs Insoluble

## Soluble

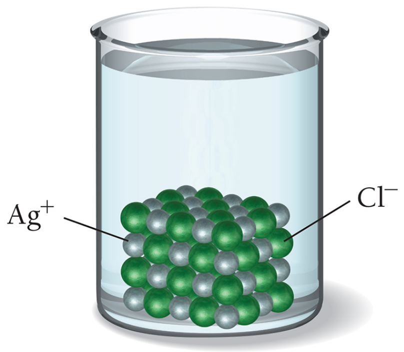



A sodium chloride solution contains independent $Na^{+}$ and $Cl^{-}$ ions.




A silver nitrate solution contains independent $Ag^{+}$ and $NO_3^{-}$ ions.

## Insoluble




When silver chloride is added to water, it remains as solid AgCl—it does not dissolve into independent ions.

# Precipitation Reactions

Precipitations occur when certain pairs of oppositely-charged ions attract each other so strongly that they form an insoluble solid.

$NaCl(aq) + AgNO_3(aq) \rightarrow$

$AgCl(s) + NaNO_3(aq)$


An exchange reaction/ double displacement reaction

# Ionic Equations

- Molecular (Formula) Equation

  $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

- Total Ionic Equation

  $\underline{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \underline{NO_3^-(aq)} \rightarrow$

  $AgCl(s) + \underline{Na^+(aq)} + \underline{NO_3^-(aq)}$

  Underlined = Spectator ions

- Net Ionic Equation

  $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

# Solubility Rules for Ionic Compounds in Water

| | |
|---|---|
| **SOLUBLE** | |
| $Li^+$, $Na^+$, $K^+$, $NH_4^+$ | |
| $NO_3^-$, $C_2H_3O_2^-$ | |
| $Cl^-$, $Br^-$, $I^-$ | Except when paired with $Ag^+$, $Hg_2^{2+}$, $Pb^{2+}$ |
| $SO_4^{2-}$ | Except when paired with $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$, $Pb^{2+}$, $Hg_2^{2+}$ |
| **INSOLUBLE** | |
| $OH^-$ | Except when paired with $Li^+$, $Na^+$, $K^+$, $NH_4^+$. *Hydroxides of $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$ are slightly soluble. |
| $S^{2-}$ | Except when paired with $Li^+$, $Na^+$, $K^+$, $NH_4^+$, $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$ |
| $CO_3^{2-}$, $PO_4^{3-}$ | Except when paired with $Li^+$, $Na^+$, $K^+$, $NH_4^+$. |

Ex Probs

# Steps to Writing Molecular Equations for Double Displacement Reactions

1. Exchange cation/anion partners to get product ion pairs.
2. Write correct formulas for products.
3. Balance equation.
4. Look at solubility rules and assign physical states.

Ex Probs

# Steps to Writing Total Ionic Equations

1. Separate all aqueous compounds into ions; keep all solid compounds together.
2. Take into account both coefficients and subscripts of each chemical formula to get the correct moles of ions.
3. Write physical states for each species!

# Acid-Base Neutralization Reactions

# Acids and Bases

- Acids
  - Sour taste
  - Eg. Acetic acid (vinegar), citric acid (lemons)

- Bases
  - Bitter taste, slippery feel
  - Eg. Drano (for unclogging drains)

# Arrhenius Definition of Acids and Bases

## Based on $H^+$ or $OH^-$ Production in Water

- **Acid:** Substance that produces $H^+$ ions (protons) when dissolved in water
  - $HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$
  - $H_2SO_4\ (aq) \rightarrow H^+\ (aq) + HSO_4^-\ (aq)$

- **Base:** Substance that produces $OH^-$ ions (hydroxide) in water
  - $NaOH\ (aq) \rightarrow Na^+ + OH^-$

# Bronsted-Lowry Definition of Acids and Bases

## Based on Proton Transfer:

- Acids: Proton donors (Must have a proton to donate)

  $HCl \rightarrow H^+ + Cl^-$

- Bases: Proton acceptors (Must have a nonbonding electron pair to accept proton)

  $NH_3 + H^+ \rightarrow NH_4^+$

# List of Common Acids and Bases

| | |
|---|---|
| Strong Acids<br>HCl hydrochloric acid<br>HBr hydrobromic acid<br>HI hydroiodic acid<br>$HNO_3$ nitric acid<br>$H_2SO_4$ sulfuric acid<br>$HClO_4$ perchloric acid | Strong Bases:<br>Metal Hydroxides of Group 1A cations<br>LiOH, NaOH, KOH, etc.<br>*and*<br>Heavier Group 2A cations:<br>$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$ |
| Common Weak Acids<br>$HC_2H_3O_2$ or $CH_3COOH$ acetic acid<br>$H_2CO_3$ carbonic acid | Common Weak Base<br>$NH_3$ ammonia |

# Acid-Base Reaction

- Neutralization reaction: the reaction of an acid with a base

# Acid-Base Reaction: Acid + Strong Base

When an acid and a strong base ($OH^-$ base) react, they form water and a salt (ionic compound).

Molecular Equation:

$HCl\ (aq) + NaOH\ (aq) \rightarrow \underline{H_2O}\ (l) + \underline{NaCl}\ (aq)$

water salt

- Total Ionic Equation:

$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow$

$H_2O\ (l) + Na^+\ (aq) + Cl^-\ (aq)$

- Net Ionic Equation:

$H^+(aq) + OH^-(aq) \rightarrow H_2O\ (l)$ Formation of Water

# Acid-Base Reaction: Gas-Forming

When an acid and certain bases (eg. $CO_3^{2-}$, $HCO_3^{-}$, $S^{2-}$) react, gases are formed.

- Molecular Equation:

$2HCl\,(aq) + K_2CO_3(aq) \rightarrow 2KCl(aq) + [H_2CO_3(aq)]$

$\downarrow$

$\underline{H_2O(l)} + \underline{CO_2(g)}$

water gas

- Total Ionic Equation:

$2H^+(aq) + 2Cl^-\,(aq) + 2K^+ + CO_3^{2-}(aq) \rightarrow 2K^+\,(aq) + 2Cl^-\,(aq) + H_2O(l) + CO_2(g)$

- Net Ionic Equation:

$2H^+(aq) + CO_3^{2-}(aq) \rightarrow [H_2CO_3(aq)] \rightarrow H_2O(l) + CO_2(g)$

Ex Probs

# Categories of Electrolytes

1. Strong electrolytes: Very good conductors of electricity. Completely ionized in water.

- Soluble ionic compounds: NaCl, $K_2CO_3$

  $K_2CO_3(s) \rightarrow 2K^+(aq) + CO_3^{2-}(aq)$

- Strong acids: HCl, $HNO_3$, $H_2SO_4$

  $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

- Strong bases: NaOH, KOH, $Ca(OH)_2$

  $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

# Categories of Electrolytes

2. Weak electrolytes: Poor conductors of electricity. A small degree of ionization in water.
   - "Insoluble" (minimally soluble) ionic compounds: $BaCO_3$
   - Weak acids: acetic acid ($CH_3COOH$)
   - Weak bases: ammonia ($NH_3$)

# Categories of Electrolytes

3. Nonelectrolytes

Do not conduct electricity. Do not produce ions in water (even if soluble).

- Covalent compounds, like sugars (glucose $C_6H_{12}O_6$ ), alcohols (ethanol $C_2H_5OH$)

# Question

Match each of the following to a picture below:

$H_2SO_4$, $HC_2H_3O_2$, methanol ($CH_3OH$)

A

$CH_3OH$

B

$H_2SO_4$

C

$CH_3COOH$

Ex Probs

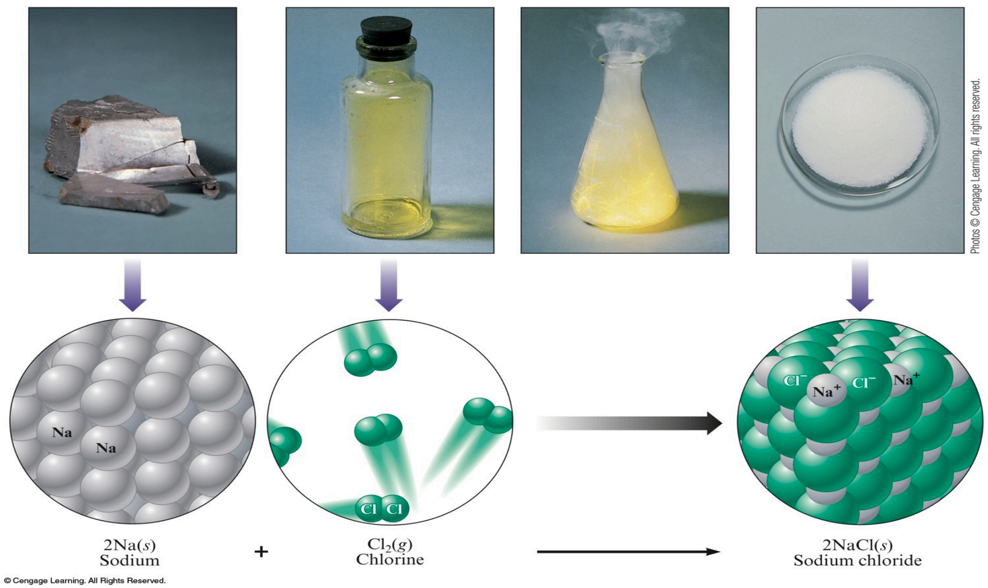
# Oxidation-Reduction (Redox) Reactions

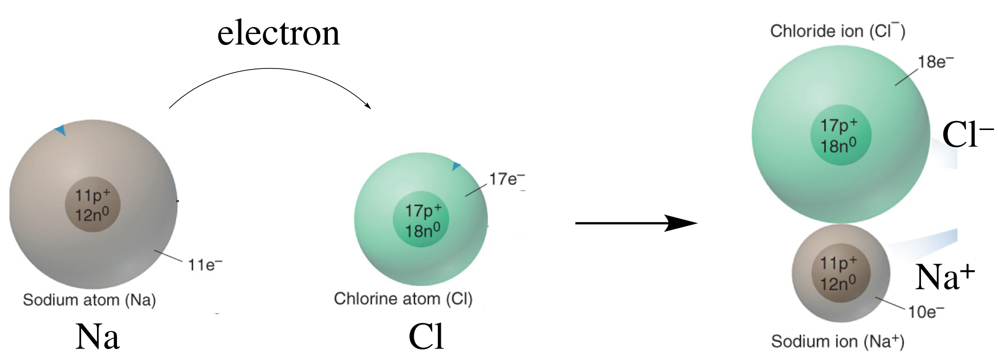
# Oxidation-Reduction Reactions

**Oxidation-Reduction (Redox) Reaction:**
Reactions in which one or more electrons are transferred between reaction partners

# Oxidation-Reduction Reaction

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



# Recognizing Redox Reactions

Redox reactions are those in which:

1. A metal reacts with a nonmetal.
2. A substance reacts with elemental oxygen:

   Combustion (reaction with $O_2$, often with release of heat and light; "burning")

   $$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

3. More generally, one substance transfers electrons to another substance.

# Recognizing Redox Reactions

Which of these are redox reactions?

a) $2Li(s) + Cl_2(g) \rightarrow 2\ LiCl(s)$
b) $2Al(s) + 3Sn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Sn(s)$
c) $Pb(NO_3)_2(aq) + 2LiCl(aq) \rightarrow PbCl_2(s) + 2LiNO_3(aq)$
d) $C(s) + O_2(g) \rightarrow CO_2(g)$

# Oxidation-Reduction Reactions

- Oxidation: Loss of electrons
- Reduction: Gain of electrons

OIL RIG or LEO says GER

- In $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$:

Na was oxidized (lost electron to make $Na^+$)

Cl was reduced (gained electron to make $Cl^–$)

# Redox Reactions

Redox reactions occur in:

- Formation of ionic compounds

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

- Formation of covalent compounds

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

# Oxidation Numbers (Oxidation States)

- Oxidation number (state): A number assigned to an element in a compound, which represents the number of electrons lost or gained by the atom of that element

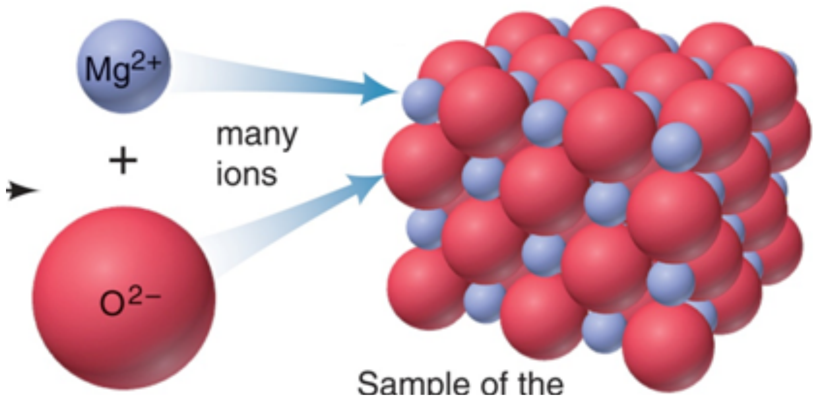
- Helps keep track of electrons in a redox reaction (a book-keeping system)

# Oxidation Numbers: Ionic Compounds

In Ionic Compounds: Oxidation number reflects transfer of electrons.

→ Oxidation Number = Ion charge for ions

Eg. MgO



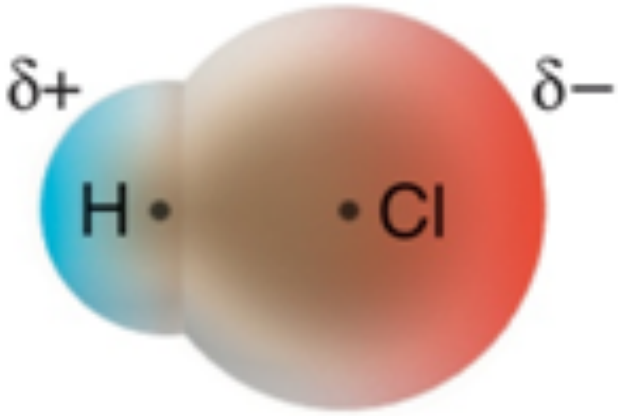

$Mg^{2+}$ ox # = +2

$O^{2-}$ ox # = -2

# Oxidation Numbers: Molecular Compounds

In Molecular Compounds: Oxidation number reflects <u>shifting</u> of electrons, unequal sharing of electrons

→ Ox number = Oxidation number the atom would have if it were an ion.




Eg. HCl

H ox # = +1

Cl ox # = -1

# Rules for Assigning Oxidation Numbers

1. For an atom in its elemental form: ox # = 0
   (eg. Na, $O_2$, $Cl_2$)

2. For a monoatomic ion: ox # = ion charge

3. Sum of ox #s for the atoms in a compound = 0.

4. Sum of ox #s for the atoms in a polyatomic ion = ion charge

# Rules for Assigning Oxidation Numbers

## Rules for Specific Elements

| **Element** | **Oxidation Number** |
|---|---|
| Group 1A | +1 |
| Group 2A | +2 |
| H | +1 if bonded to nonmetals (-1 if bonded to metals) |
| Halogens (7A) | -1 (Exception: Halogens other than F are positive when bonded to O) |
| O | -2 (Exceptions: -1 in peroxides) |

Other elements can have a range of ox. #s (eg. +4 to -4 for C).

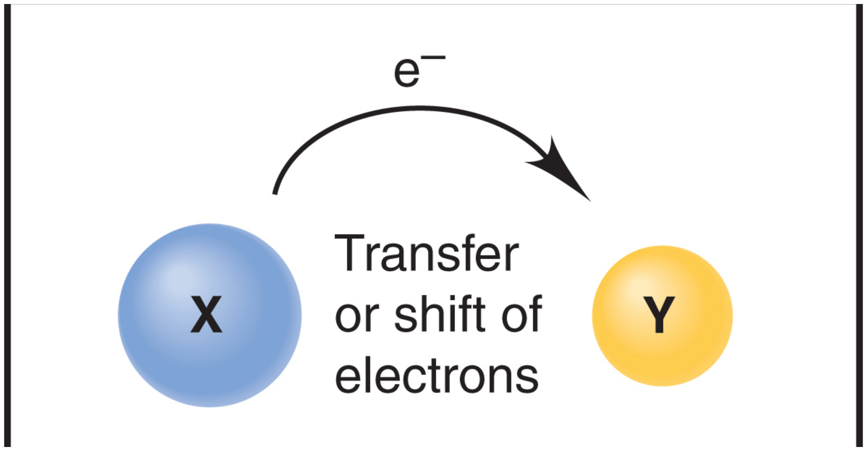
Ex probs

# Oxidation-Reduction Definitions

- Oxidation: Loss of electrons
- Reduction: Gain of electrons
- Oxidizing agent: the substance being reduced
- Reducing agent: the substance being oxidized

# Oxidation-Reduction Definitions

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




<u>X = Na</u>

- Na loses electrons → is oxidized (ox # increases, becomes more positive)
- Na(s) is reducing agent.

<u>Y = Cl</u>

- Cl gains electrons → is reduced (ox # decreases, becomes more negative)
- $Cl_2(g)$ is oxidizing agent.

Ex probs

# Classification of Reactions by What Atoms Do

# Some Examples of Redox Reactions

| Type of Reaction | Generic Equation |
|---|---|
| synthesis or combination | $A + B \longrightarrow AB$ |
| decomposition | $AB \longrightarrow A + B$ |
| single-displacement | $A + BC \longrightarrow AC + B$ |
| double-displacement | $AB + CD \longrightarrow AD + CB$ |

# Combination (Synthesis)

$A + B \rightarrow AB$

- $2Na(s) + Cl_2(g) \rightarrow NaCl(s)$
- $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$
- $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

# Decomposition

$AB \rightarrow A + B$

- $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
- $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
- $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

# Single Displacement

$A + BC \rightarrow AB + C$

- $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$
- $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
- $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

# Double Displacement

$AB + CD \rightarrow AD + CB$

- $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
- $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$

# Question

Combination (synthesis), decomposition, single-displacement, or double-displacement reaction?

- $NaNO_3(aq) + KCl(aq) \rightarrow NaCl(s) + KNO_3(aq)$
- $2Al(s) + 3Br_2(l) \rightarrow 2AlBr_3(s)$
- $Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow 2H_2O(l) + Ca(NO_3)_2(aq)$
- $CuCl_2(aq) \rightarrow Cu(s) + Cl_2(g)$
- $2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$