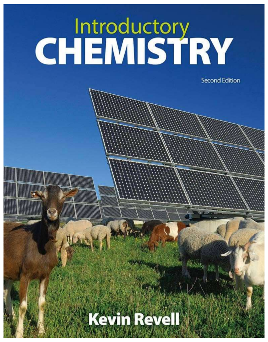
Introductory Chemistry
Chem 103

Chapter 4 – Light and Electronic Structure

Lecture Slides



The Electromagnetic Spectrum

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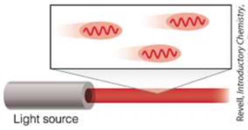
What is Light?

electromagnetic radiation

- a form of energy
- travels in waves
- exists in increments called **photons**



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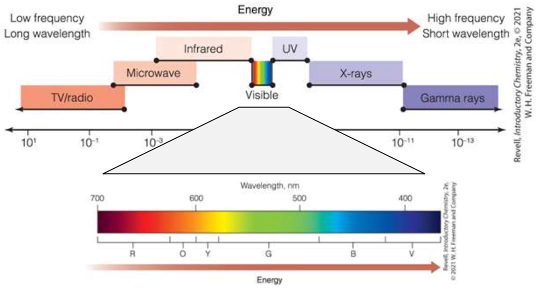
The Electromagnetic Spectrum



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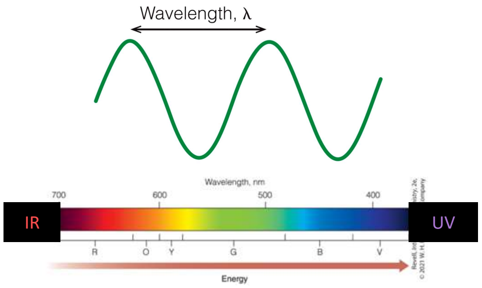
Wavelength



Describing Electromagnetic Waves

wavelength (λ) – The length of one wave

frequency (ν) – The number of waves per second

1 wave/second = 1 hertz (Hz)

10,000 Hz   10,000/s   10,000 s^{-1}

Describing Electromagnetic Waves, Continued

↓ wavelength
↑ frequency
} *inversely related*

$$c = \lambda\nu$$

speed of light = wavelength x frequency

$$\frac{m}{s} = m \times \frac{1}{s}$$

c = speed of light = 3.00 × 10^8 m/s

Example of Describing Electromagnetic Waves

A beam of green light has a wavelength of 500 nm.
What is the frequency of this light?

$$c = \lambda\nu$$

$c = 3.00 \times 10^8 \ m/s$

$\lambda = 500 \ nm = 500 \times 10^{-9} \ m$

$1 \ nm = 10^{-9} \ m$

$\nu = ?$

$$\frac{c}{\lambda} = \nu$$

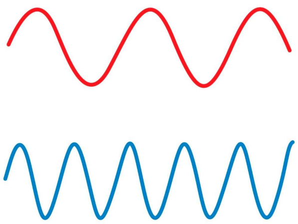
$$\frac{3.00 \times 10^8 \ \cancel{m}/s}{500 \times 10^{-9} \ \cancel{m}} = \nu$$

units: $1/s = Hz$

$6 \times 10^{14} \ Hz = \nu$

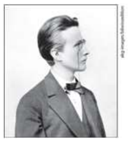
Frequency and Wavelength

The energy of light depends on its frequency and wavelength.



longer wavelength
lower frequency
lower energy

shorter wavelength
higher frequency
higher energy



Energy of a photon:

$$\boxed{E = h\nu}$$

energy

frequency

Planck's constant

$= 6.63 \times 10^{-34}\ J{\cdot}s$

$$\nu = c/\lambda$$

$$\boxed{E = hc/\lambda}$$

Example of Photon Energy

A photon has a frequency of 7.50 × 10^{14} Hz. What is the wavelength of this light? What color is this light? What is the energy of the photon?

$$\boxed{c = \lambda\nu}$$

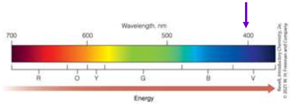
$$\frac{c}{\nu} = \lambda$$

$$\frac{3.00 \times 10^{8}\ m/\cancel{s}}{7.50 \times 10^{14}/\cancel{s}} = \lambda$$

$$4.00 \times 10^{-7}\ m = \lambda$$

$$= 400\ nm$$

violet

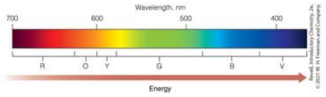


$$\boxed{E = h\nu}$$

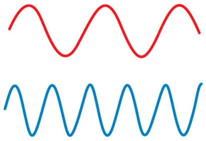
$$E = (6.63 \times 10^{-34}\ J{\cdot}\cancel{s})(7.50 \times 10^{14}/\cancel{s})$$

$$E = 4.97 \times 10^{-19}\ J$$

Summary of Electromagnetic Waves



- Light is a form of electromagnetic radiation
- We describe light by its
  - frequency (ν)
  - wavelength (λ)
  - energy (E)
- $c = \lambda\nu$
- $E = h\nu = hc/\lambda$



Color, Line Spectra, and the Bohr Model

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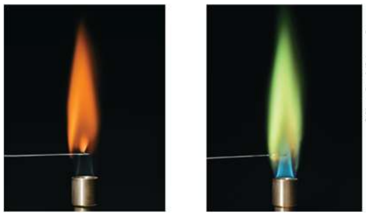
Flame Tests

observe colors emitted by different metal ions



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Gas lamps also produce unique colors:

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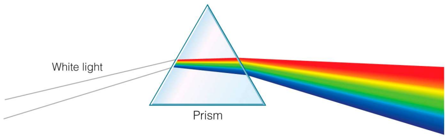
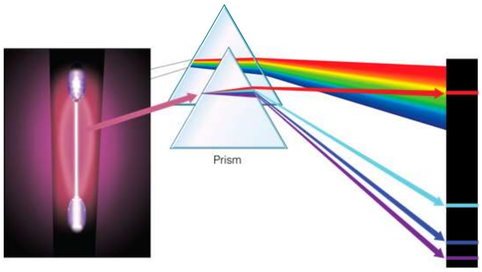
Line Spectra

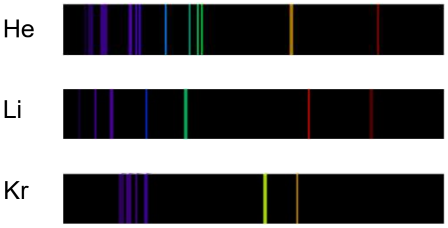


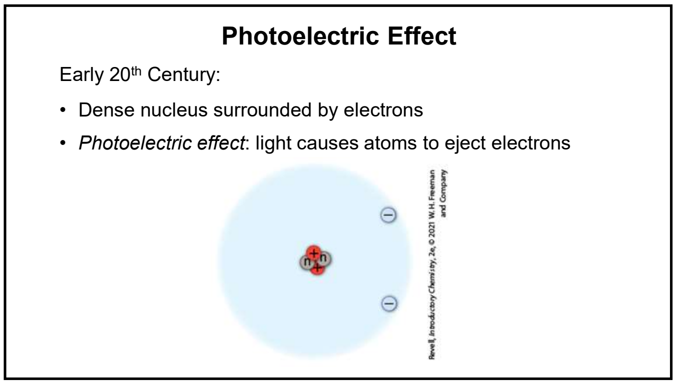
Line Spectra, Continued



Examples of Line Spectra

Each element produces a unique line spectrum.





Photoelectric Effect

Early 20th Century:

- Dense nucleus surrounded by electrons
- *Photoelectric effect*: light causes atoms to eject electrons

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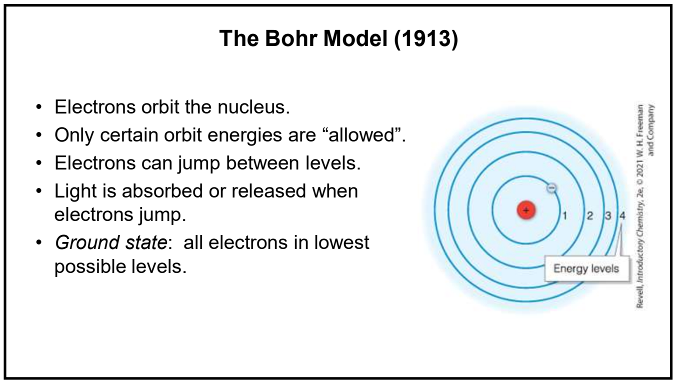

The Bohr Model (1913)

- Electrons orbit the nucleus.
- Only certain orbit energies are "allowed".
- Electrons can jump between levels.
- Light is absorbed or released when electrons jump.
- *Ground state*: all electrons in lowest possible levels.

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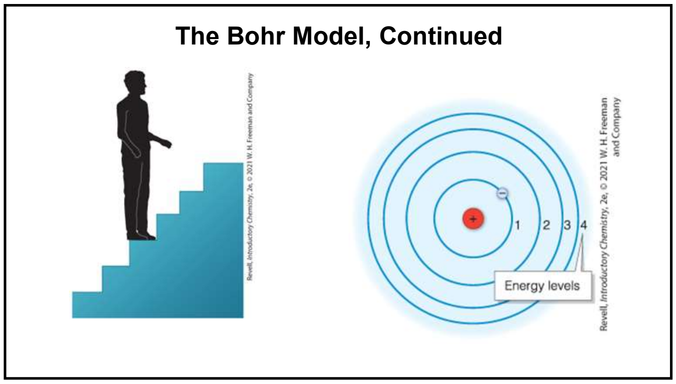

The Bohr Model, Continued

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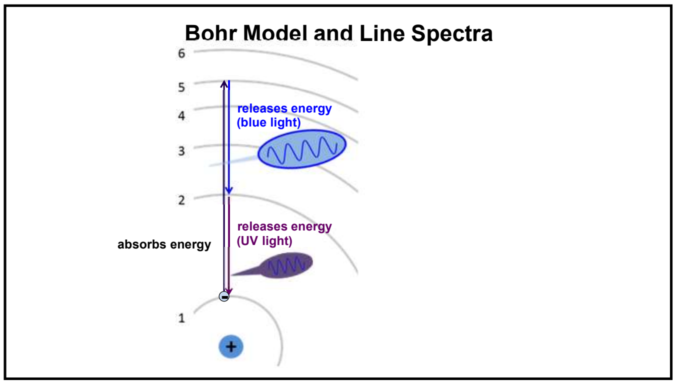
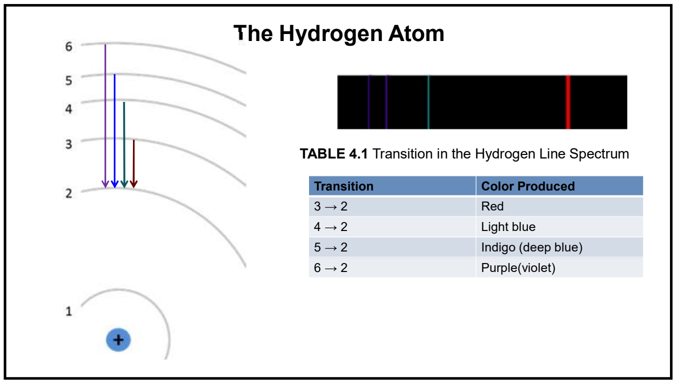


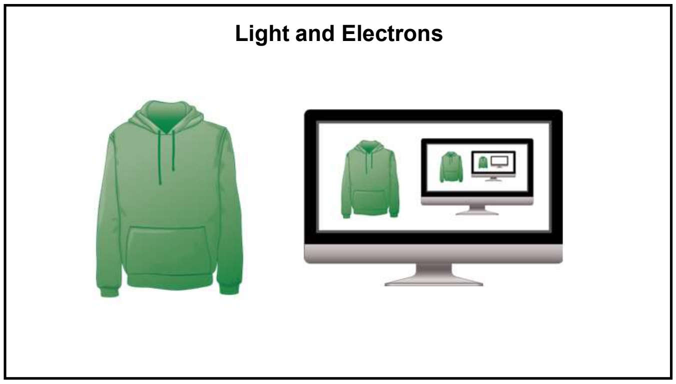
Bohr Model and Line Spectra



The Hydrogen Atom

TABLE 4.1 Transition in the Hydrogen Line Spectrum

Transition	Color Produced
3 → 2	Red
4 → 2	Light blue
5 → 2	Indigo (deep blue)
6 → 2	Purple(violet)



Light and Electrons

Sources of Light

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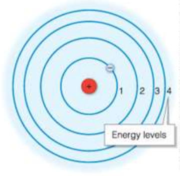
Do You Give Off Light?

Monty Rakusen/Getty Images

Summary of the Bohr Model

Explained
- The hydrogen line spectrum
- Some properties of main group elements

Did not explain
- More complex line spectra
- Properties of the transition elements



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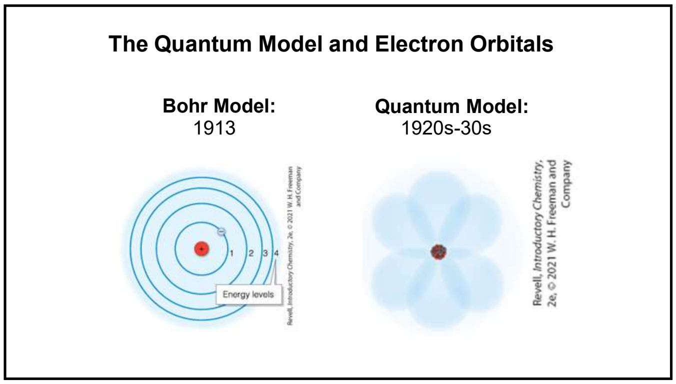

The Quantum Model and Electron Orbitals

Bohr Model:
1913

Quantum Model:
1920s-30s

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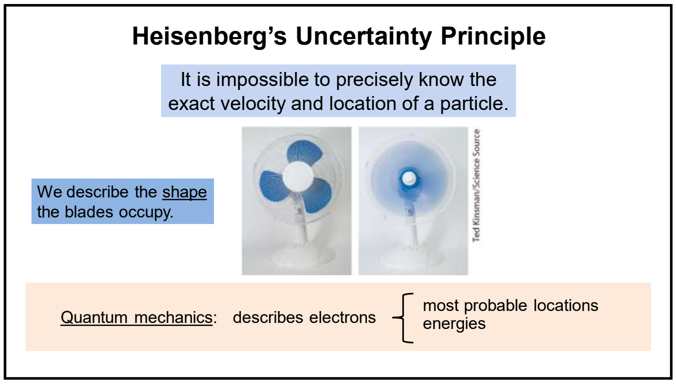

Heisenberg's Uncertainty Principle

It is impossible to precisely know the exact velocity and location of a particle.

We describe the <u>shape</u> the blades occupy.

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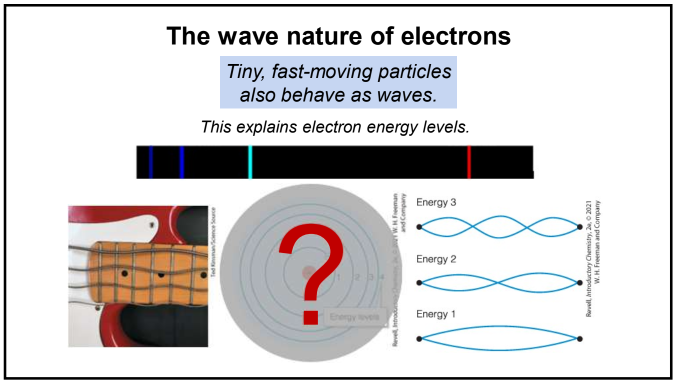
<u>Quantum mechanics</u>: describes electrons {
- most probable locations
- energies



The wave nature of electrons

Tiny, fast-moving particles also behave as waves.

This explains electron energy levels.

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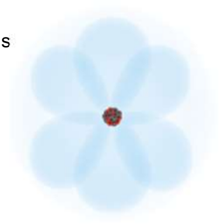
The Quantum Model

Main Ideas:

- uncertainty principle
- wave nature of electrons

QM describes electrons by

- energy
- probable locations



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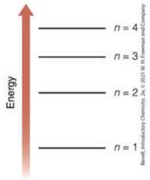
Energy Levels and Sublevels, Part 1

1. Electrons occupy different energy levels.
    - Level is identified by its **principal quantum number**, *n* (1, 2, 3…)
    - Higher energy levels can hold more electrons



Level	Electron Capacity
1	2
2	8
3	18
4	32

Energy Levels and Sublevels, Part 2

2. Each energy level contains one or more **sublevels**.

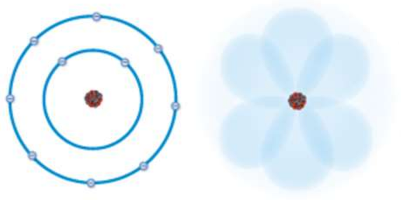
Sublevel
s
p
d
f

Energy Levels and Sublevels, Part 3

3. Each sublevel contains one or more **orbitals.**

Sublevel	Number of Orbitals
s	1
p	3
d	5
f	7

The Bohr Model and the Quantum Model



Bohr model
Electrons orbit like planets

Quantum model
Electrons behave like waves that occupy different regions

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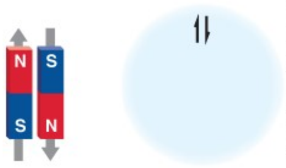
Energy Levels and Sublevels, Part 4

4. Each orbital holds up to two electrons.
   - Electrons have a magnetic field, called spin.
   - Electrons with opposite spins pair together.



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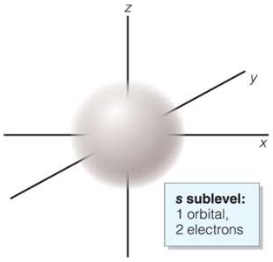
Energy Levels and Sublevels, Summary

1. Electrons occupy different energy levels.
2. Each level contains sublevels.
3. Each sublevel contains orbitals.
4. Each orbital holds up to two electrons.

Sublevel	Number of Orbitals	Electron Capacity
s	1	2
p	3	6
d	5	10
f	7	14

Level 1: *s* only



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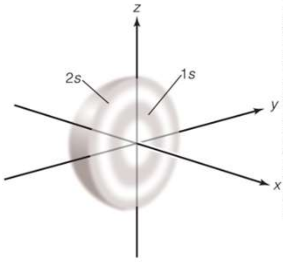
Level 2: *s* + *p*, part 1



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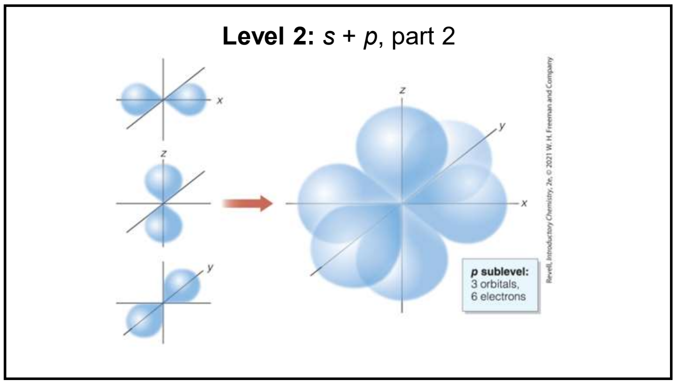

Level 2: *s* + *p*, part 2

***p* sublevel:**
3 orbitals,
6 electrons

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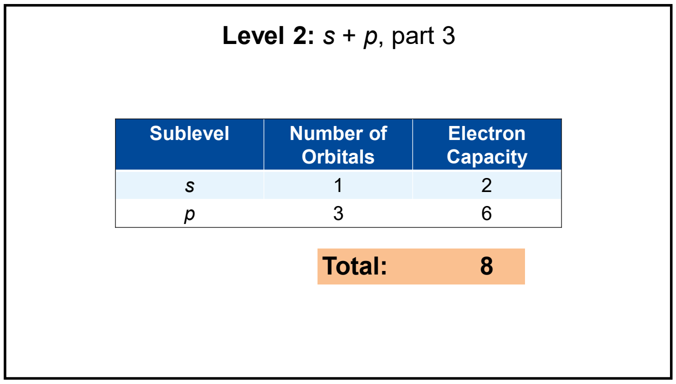
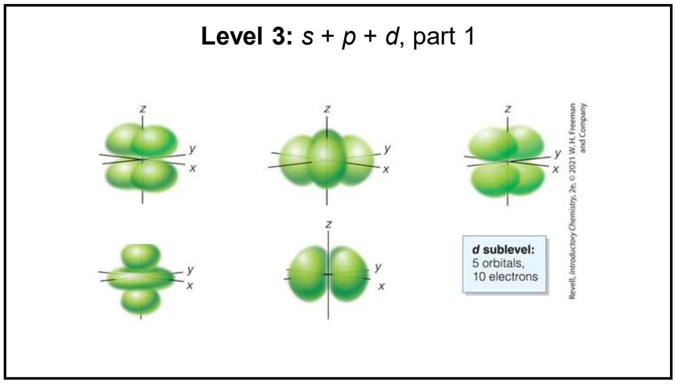


Level 2: *s* + *p*, part 3

Sublevel	Number of Orbitals	Electron Capacity
s	1	2
p	3	6

Total: 8



Level 3: *s* + *p* + *d*, part 1

***d* sublevel:**
5 orbitals,
10 electrons

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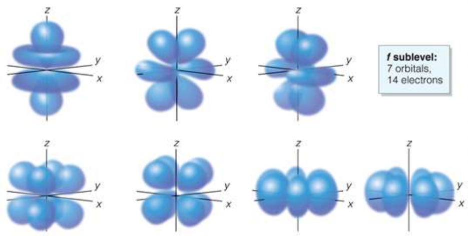
Level 3: *s* + *p* + *d*, part 2

Sublevel	Number of Orbitals	Electron Capacity
s	1	2
p	3	6
d	5	10

Total: 18

Level 4: *s* + *p* + *d* + *f*, part 1



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Level 4: *s* + *p* + *d* + *f*, part 2

Sublevel	Number of Orbitals	Electron Capacity
s	1	2
p	3	6
d	5	10
f	7	14

Total: 32

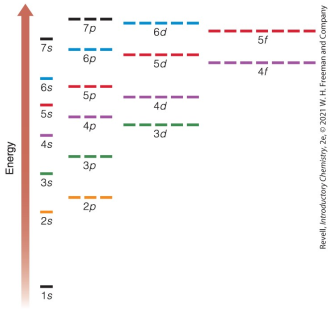
Summary of Atomic Energy Levels

TABLE 4.4 Energy Levels, Sublevels, and Electron Capacity

Energy Level	1	2	3	4
Sublevels				*f* (14 e^-)
			d (10 e^-)	*d* (10 e^-)
		p (6 e^-)	*p* (6 e^-)	*p* (6 e^-)
	s (2 e^-)	*s* (2 e^-)	*s* (2 e^-)	*s* (2 e^-)
Electron Capacity	2	8	18	32

Note : the symbol e^- means electron.

Energy Differences Between Levels



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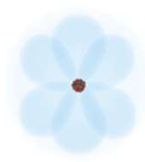
Describing Electron Configuration

Quantum Model:
Energy levels – 1, 2, 3…
Energy sublevels – *s*, *p*, *d*, *f*



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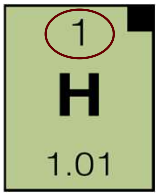
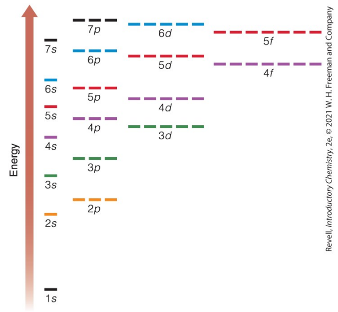


Filling the Energy Levels



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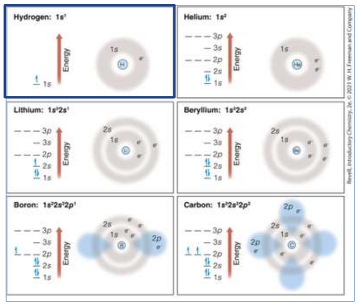
Hydrogen:



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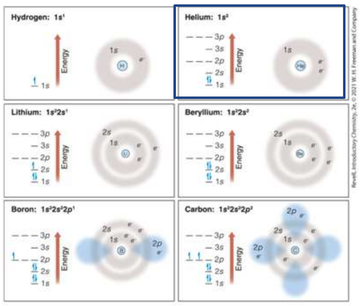
Helium:



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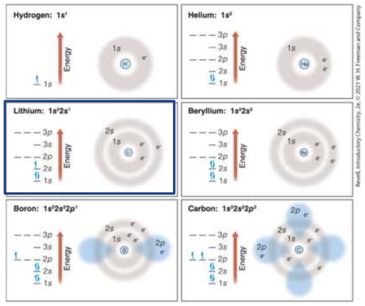
Lithium:



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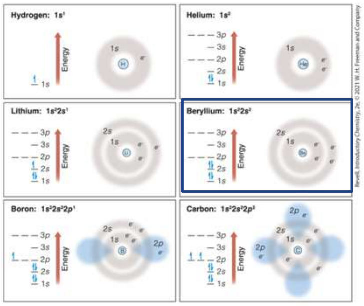
Beryllium:



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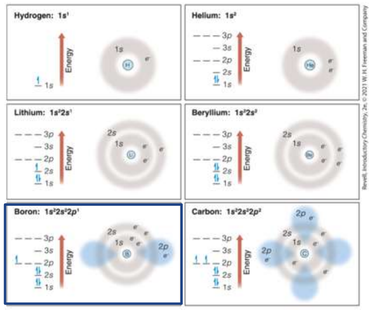
Boron:



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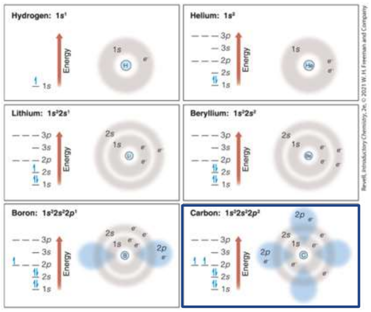
***Hund's Rule*:**
If empty orbitals of the same energy are available, electrons singly occupy orbitals rather than pairing together.

Carbon:



Electron Configurations of Row 2 Elements

3 Li 6.94	4 Be 9.01

5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18

Li: $1s^22s^1$

Be: $1s^22s^2$

B: $1s^22s^22p^1$

C: $1s^22s^22p^2$

N: $1s^22s^22p^3$

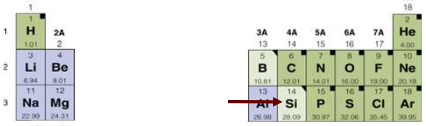
O: $1s^22s^22p^4$

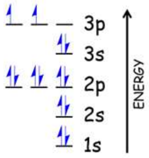
F: $1s^22s^22p^5$

Ne: $1s^22s^22p^6$

Example for Silicon

What is the electron configuration of silicon?

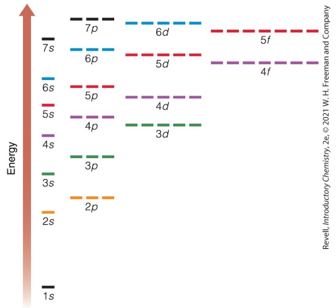




14 e^- total

$1s^22s^22p^63s^23p^2$

Energy Diagram and Writing Electron Configurations



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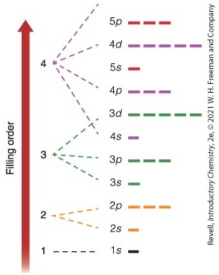
Describing Electron Configuration, Part 2

valence level: The highest occupied electron energy level

- Up to 8 electrons in valence level



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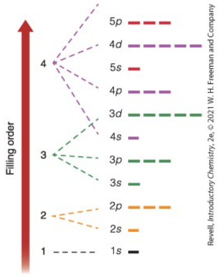
Describing Electron Configurations, Part 3

Argon: (18 e^-)

$$1s^22s^22p^6\underbrace{3s^23p^6}_{\text{valence level}}$$

Potassium: (19 e^-)

$$1s^22s^22p^63s^23p^6\underbrace{4s^1}_{\text{valence level}}$$



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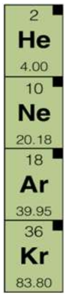
Noble Gases have Filled Valences



$\underbrace{1s^2}$

$1s^2\underbrace{2s^22p^6}$

$1s^22s^22p^6\underbrace{3s^23p^6}$

$1s^22s^22p^63s^23p^6\underbrace{4s^2}3d^{10}\underbrace{4p^6}$

Octet Rule:
An atom is stabilized by having its highest-occupied (valence) energy level filled.

Electron Configurations for Larger Atoms

	inner electrons	Noble gas notation
Sodium:	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Phosphorous:	$1s^22s^22p^63s^23p^3$	$[Ne]3s^23p^3$
Chlorine:	$1s^22s^22p^63s^23p^5$	$[Ne]3s^23p^5$

$1s^22s^22p^6 = [Ne]$

Electron Configurations for Larger Atoms, Continued

not involved with bonding

$$\text{Titanium: } \overbrace{[Ar]}^{\text{inner electrons}}\underbrace{4s^23d^2}_{\text{outer electrons}}$$

valence
+ d, f sublevels
essential to bonding

Example of Writing an Electron Configuration

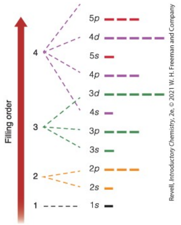
Write the electron configuration for selenium using the noble gas shorthand. Identify the inner electrons, the outer electrons, and the valence electrons.

34
Se
78.97

$$\underbrace{1s^2 2s^2 2p^6 3s^2 3p^6}_{[Ar]} 4s^2 3d^{10} 4p^4$$

$$\overbrace{[Ar]}^{\text{inner}}\overbrace{\boxed{4s^2}\,3d^{10}\,\boxed{4p^4}}^{\text{outer}}$$

valence



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Example, Electron Configuration for Ions - Sodium

11
Na
22.99

What is the electron configuration of a sodium atom?
What is the electron configuration of a sodium ion with a +1 charge?

species	Symbol	full configuration	noble-gas shorthand
sodium atom	Na	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
sodium ion (+1 charge)	Na^+	$1s^2 2s^2 2p^6$	$[He]2s^2 2p^6$ or $[Ne]$

Example, Electron Configuration for Ions - Oxygen

8
O
16.00

What is the electron configuration of an oxide ion, which is an oxygen ion with a charge of −2?

species	symbol	full configuration	noble-gas shorthand
oxygen atom	O	$1s^2 2s^2 2p^4$	$[He]2s^2 2p^4$
oxide ion (−2 charge)	O^{2-}	$1s^2 2s^2 2p^6$	$[He]2s^2 2p^6$ or $[Ne]$

Many ions form noble gas configurations

O: $1s^22s^22p^4$ Na: $1s^22s^22p^63s^1$

O^{2-}: $1s^22s^22p^6$ Na^+: $1s^22s^22p^6$

Ne: $1s^22s^22p^6$

These are ***isoelectronic***

Electron Configuration
and the
Periodic Table

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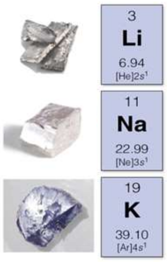
Group 1A Electron Configurations



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Lithium (3 electrons): [He]$2s^1$

Sodium (11 electrons): [Ne]$3s^1$

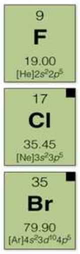
Potassium (19 electrons): [Ar]$4s^1$

Group 7A Electron Configurations

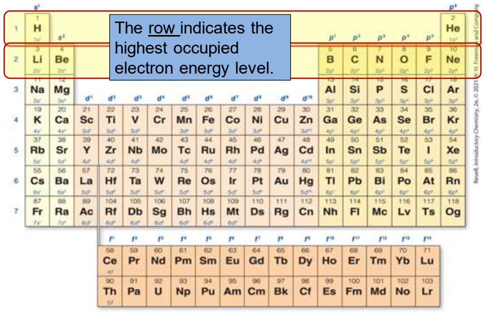
Fluorine: [He]$2s^22p^5$

Chlorine: [Ne]$3s^23p^5$

Bromine: [Ar]$4s^23d^{10}4p^5$

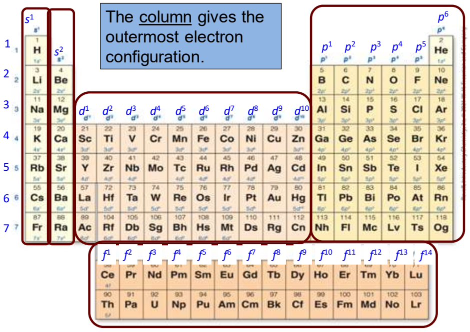


Row and Energy Level



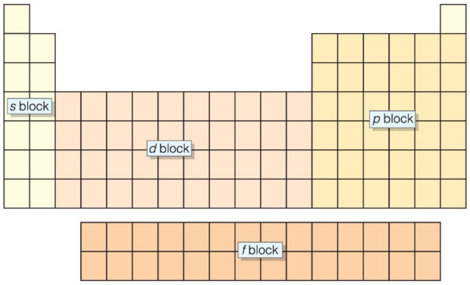
The row indicates the highest occupied electron energy level.

Column and Electron Configuration

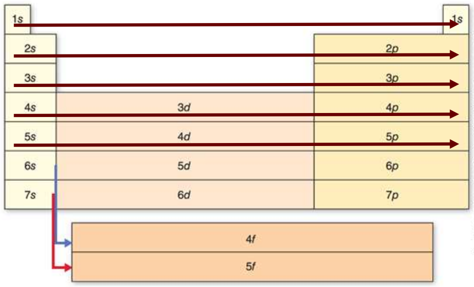


The column gives the outermost electron configuration.

Four Blocks of the Periodic Table



Organization of the Periodic Table



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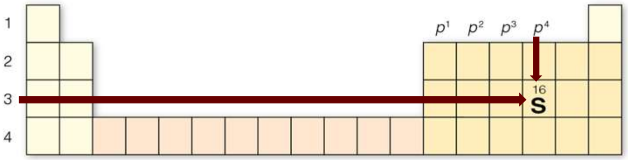
Sulfur Electron Configuration

What is the outermost electron configuration for sulfur?

$3p^4$



Highest-Energy Occupied Sublevel

Write the configuration for the highest-energy occupied sublevel for potassium, phosphorus, and iron.

K: $4s^1$

P: $3p^3$

Fe: $3d^6$

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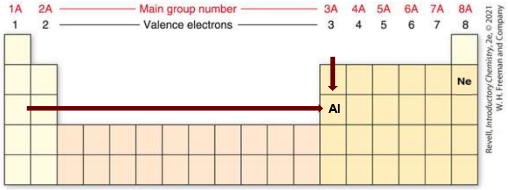
Electron Configuration of Aluminum

Write the electron configuration for aluminum.

How many valence electrons does aluminum have?

$[Ne]3s^2 3p^1$

3 valence electrons



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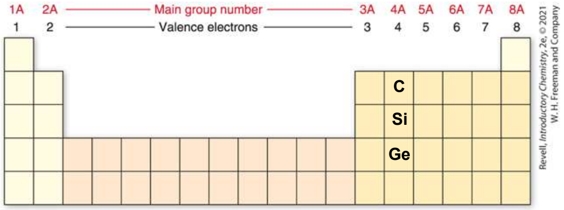
Valence Electrons



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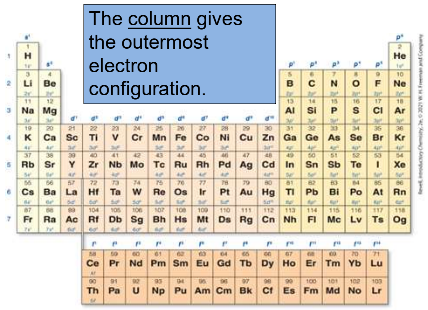
Summary of Periodic Table Organization

The row indicates the highest occupied electron energy level.

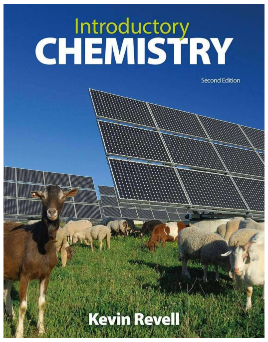
The column gives the outermost electron configuration.



Introductory Chemistry
Chem 103

Chapter 5 – Chemical Bonds and Compounds

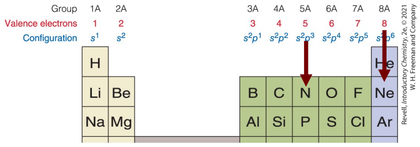
Lecture Slides



Lewis Symbols and the Octet Rule

Valence electrons

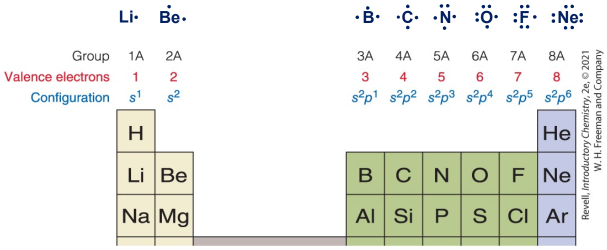
- electrons in highest occupied energy level
- *s* and *p* sublevels
- generally up to 8 electrons



Lewis Symbols Show Valence Electrons

Lewis dot symbols

- Represent valence electrons as dots around atomic symbol



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The Octet Rule

Octet Rule: An atom is stabilized by having its valence energy level filled.

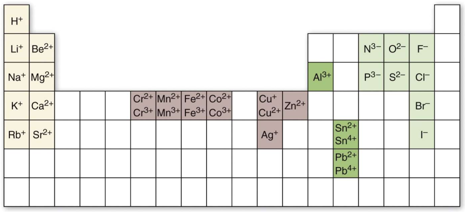
Noble gases fulfill the octet rule.

Other atoms fulfill the octet rule by:

- gaining or losing electrons (ions).
- sharing electrons.

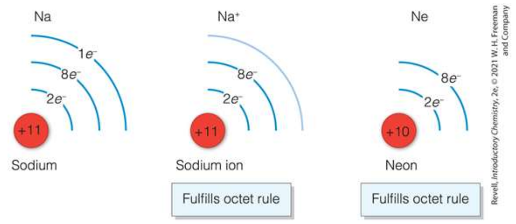
Ions

Atoms or groups of atoms that have an overall charge.

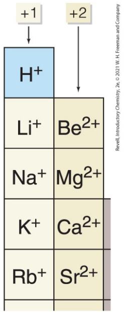


Cations – positively charged ions

Main group metals fulfill the octet rule by forming cations



Lewis Structures Show Cation Formation



$$\dot{Na} \xrightarrow{\text{loss of one electron}} Na^+ + e^-$$

$[Ne]3s^1$ $[Ne]$

$$\dot{Mg}\bullet \xrightarrow{\text{loss of two electrons}} Mg^{2+} + 2e^-$$

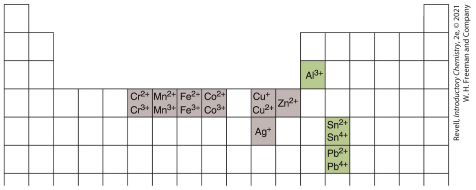
$[Ne]3s^2$ $[Ne]$

Transition metals also form cations.

Typical charges are +1, +2, +3, or +4

Some metals form multiple charged ions.

p-block metals also do this.



Naming Cations

Metal cations have the same name as the neutral metal.

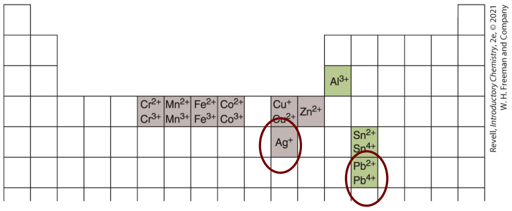
Na^+ sodium
Mg^{2+} magnesium

Atom	Ion	Older Name	Modern Name
Iron	Fe^{2+}	ferrous	iron(II)
	Fe^{3+}	ferric	iron(III)
Copper	Cu^+	cuprous	copper(I)
	Cu^{2+}	cupric	copper(II)

Practice Naming Cations

Name the following cations:

Ag^+	Pb^{2+}	Pb^{4+}
silver	*lead(II)*	*lead(IV)*



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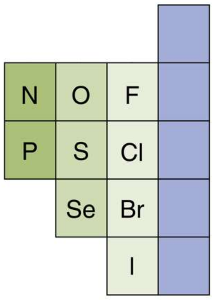
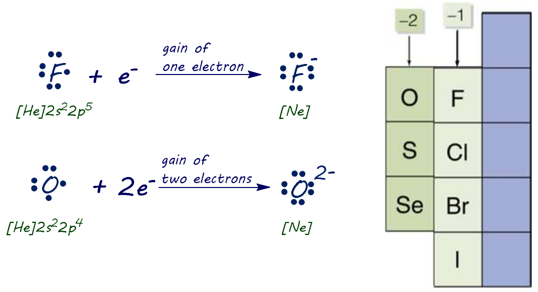
Anions – negatively charged ions



N	O	F	
P	S	Cl	
	Se	Br	
		I	

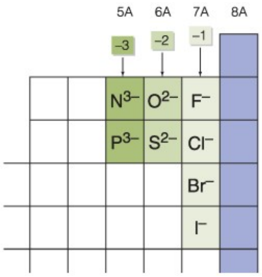
Anions Fulfill the Octet Rule, Part 1

Most nonmetals fulfill the octet rule by gaining electrons.



Anions Fulfill the Octet Rule, Part 2

Most nonmetals fulfill the octet rule by gaining electrons.



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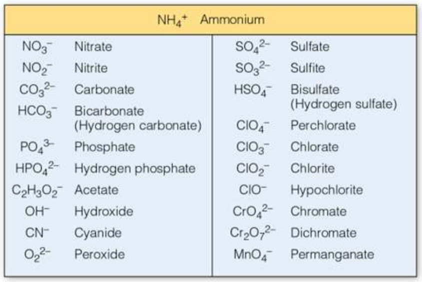
Naming Anions: change ending to *-ide*

Atom	Anion Symbol	Anion Name
chlorine	Cl^{-}	chloride
oxygen	O^{2-}	oxide
sulfur	S^{2-}	sulfide
nitrogen	N^{3-}	nitride

Polyatomic ions: groups of atoms with a charge, part 1



NH_4^+ Ammonium			
NO_3^-	Nitrate	SO_4^{2-}	Sulfate
NO_2^-	Nitrite	SO_3^{2-}	Sulfite
CO_3^{2-}	Carbonate	HSO_4^-	Bisulfate (Hydrogen sulfate)
HCO_3^-	Bicarbonate (Hydrogen carbonate)	ClO_4^-	Perchlorate
PO_4^{3-}	Phosphate	ClO_3^-	Chlorate
HPO_4^{2-}	Hydrogen phosphate	ClO_2^-	Chlorite
$C_2H_3O_2^-$	Acetate	ClO^-	Hypochlorite
OH^-	Hydroxide	CrO_4^{2-}	Chromate
CN^-	Cyanide	$Cr_2O_7^{2-}$	Dichromate
O_2^{2-}	Peroxide	MnO_4^-	Permanganate

Polyatomic ions: groups of atoms with a charge, part 2

Oxyanions – contain oxygen

Usually named as element root + *-ate*

$CO_3^{2–}$	carbonate
$PO_4^{3–}$	phosphate

Polyatomic ions: groups of atoms with a charge, part 3

More than one oxyanion:

-ate	more oxygen atoms
-ite	fewer oxygen atoms

$NO_3^–$	nitrate
$NO_2^–$	nitrite

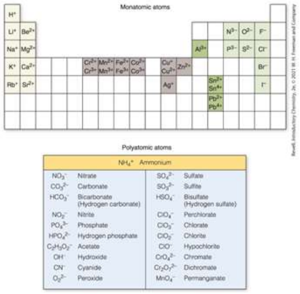
Polyatomic ions: groups of atoms with a charge, part 4

More than one oxyanion:

-ate more oxygen atoms
-ite fewer oxygen atoms

ClO_4^-	perchlorate
ClO_3^-	chlorate
ClO_2^-	chlorite
ClO^-	hypochlorite

Ions to Know



NH_4^+	Ammonium		
NO_3^-	Nitrate	SO_4^{2-}	Sulfate
CO_3^{2-}	Carbonate	SO_3^{2-}	Sulfite
HCO_3^-	Bicarbonate (Hydrogen carbonate)	HSO_4^-	Bisulfate (Hydrogen sulfate)
NO_2^-	Nitrite	ClO_4^-	Perchlorate
PO_4^{3-}	Phosphate	ClO_3^-	Chlorate
HPO_4^{2-}	Hydrogen phosphate	ClO_2^-	Chlorite
$C_2H_3O_2^-$	Acetate	ClO^-	Hypochlorite
OH^-	Hydroxide	CrO_4^{2-}	Chromate
CN^-	Cyanide	$Cr_2O_7^{2-}$	Dichromate
O_2^{2-}	Peroxide	MnO_4^-	Permanganate

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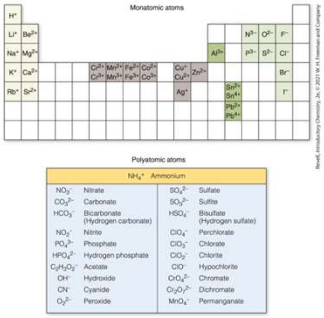
Ionic Bonds and Compounds



NH_4^+	Ammonium		
NO_3^-	Nitrate	SO_4^{2-}	Sulfate
CO_3^{2-}	Carbonate	SO_3^{2-}	Sulfite
HCO_3^-	Bicarbonate (Hydrogen carbonate)	HSO_4^-	Bisulfate (Hydrogen sulfate)
NO_2^-	Nitrite	ClO_4^-	Perchlorate
PO_4^{3-}	Phosphate	ClO_3^-	Chlorate
HPO_4^{2-}	Hydrogen phosphate	ClO_2^-	Chlorite
$C_2H_3O_2^-$	Acetate	ClO^-	Hypochlorite
OH^-	Hydroxide	CrO_4^{2-}	Chromate
CN^-	Cyanide	$Cr_2O_7^{2-}$	Dichromate
O_2^{2-}	Peroxide	MnO_4^-	Permanganate

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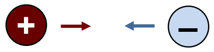
Ionic Bonds and Compounds, Continued



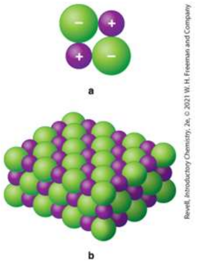
ionic bond – an attraction between oppositely charged ions

ionic compound – composed of charged ions

Metal cations and nonmetal anions form ionic compounds.

Ionic Compound Structure

ionic lattice – an array of positive and negative ions.



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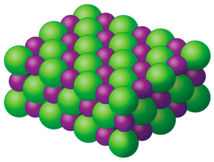
Chemical Formulas

Show the type and amount of each element present

Empirical formula: The smallest whole-number ratio of atoms

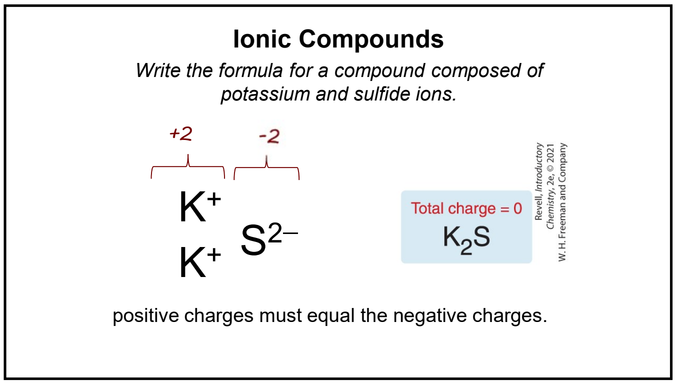
Formula unit: The smallest number of ions necessary to form a compound



cation

NaCl

anion



Ionic Compounds

Write the formula for a compound composed of potassium and sulfide ions.

$$\underbrace{K^+ \; K^+}_{+2} \quad \underbrace{S^{2-}}_{-2}$$

Total charge = 0

K_2S

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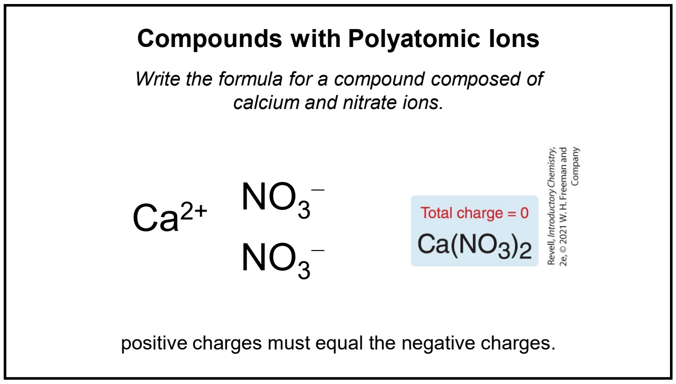
positive charges must equal the negative charges.



Compounds with Polyatomic Ions

Write the formula for a compound composed of calcium and nitrate ions.

Ca^{2+} NO_3^- NO_3^-

Total charge = 0

$Ca(NO_3)_2$

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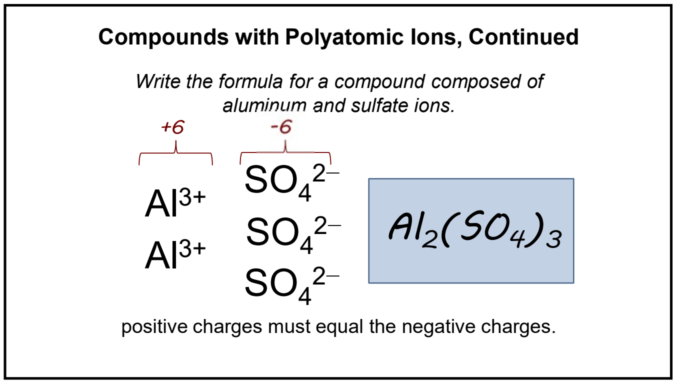
positive charges must equal the negative charges.



Compounds with Polyatomic Ions, Continued

Write the formula for a compound composed of aluminum and sulfate ions.

$$\underbrace{Al^{3+} \; Al^{3+}}_{+6} \quad \underbrace{SO_4^{2-} \; SO_4^{2-} \; SO_4^{2-}}_{-6}$$

$Al_2(SO_4)_3$

positive charges must equal the negative charges.

Naming Ionic Compounds, Part 1

cation anion

NaCl sodium chloride

$MgCl_2$ magnesium chloride

$\underbrace{Mg}\underbrace{SO_4}$ magnesium sulfate

Naming Ionic Compounds, Part 2

cation anion

Cu^+

CuCl copper(I) chloride

$CuCl_2$ copper(II) chloride

Cu^{2+}

Example, Naming Ionic Compounds

1. Name the compound $Fe(NO_2)_2$.

Fe^{2+} NO_2^-

NO_2^-

iron(II) nitrite

iron(II)

nitrite

2. Write the empirical formula for ammonium sulfide.

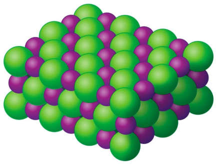
NH_4^+ S^{2-}

NH_4^+

$\boxed{(NH_4)_2S}$

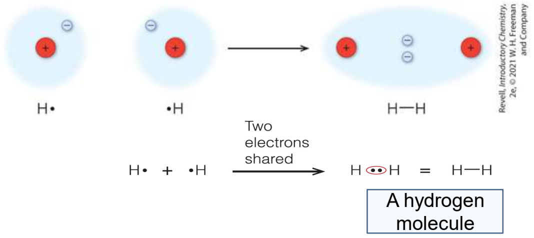
Summary, Ionic Compounds

- Ionic bonds occur between oppositely charged ions
- In ionic compounds, total charge = 0
- Named as “cation anion”
- Formula ⇔ Name



Covalent Bonding, Part 1

covalent bond – electrons shared between two atoms



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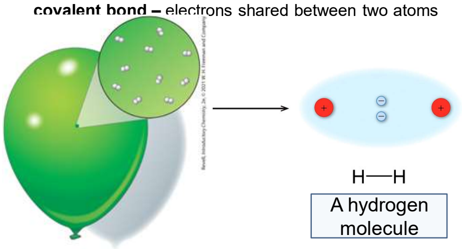
By sharing electrons, each hydrogen completes its valence level.

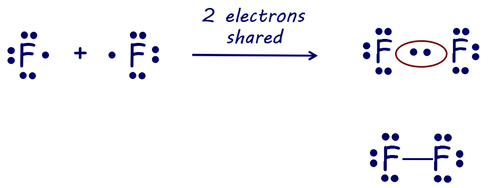
Covalent Bonding, Part 2

covalent bond – electrons shared between two atoms



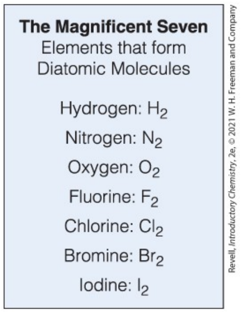
By sharing electrons, each hydrogen completes its valence level.

Covalent Bonding, Part 3



Lewis structure

Seven Elements Form Diatomic Molecules



The Magnificent Seven
Elements that form
Diatomic Molecules

Hydrogen: H_2
Nitrogen: N_2
Oxygen: O_2
Fluorine: F_2
Chlorine: Cl_2
Bromine: Br_2
Iodine: I_2

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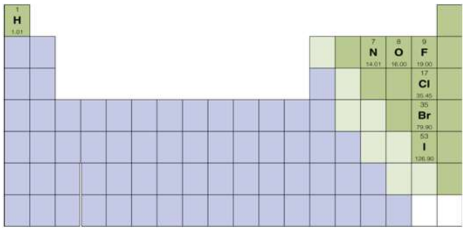
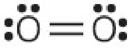
Seven Elements Form Diatomic Molecules, Continued



Double and Triple Bonds in Lewis Structures

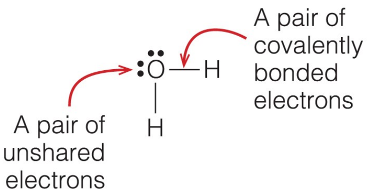


double covalent bond

triple covalent bond

Covalent Compounds

Covalent compounds fulfill the octet rule by sharing electrons.



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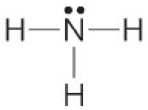
Electrons in Lewis Structures

In this structure, how many electrons does the nitrogen atom share through covalent bonds? How many of the valence nitrogen electrons are not shared? Does this nitrogen atom have a complete octet?

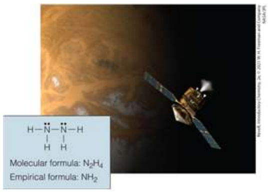


Nitrogen has 6 shared electrons
and 2 unshared electrons

8 electrons – a complete octet

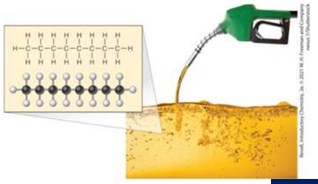
Covalent Compounds, Continued

molecular formula – gives the number of atoms in the molecule



Empirical Formula: NH_2
Molecular Formula: N_2H_4

Covalent Compound Structures



Covalent compounds often have complex structures.

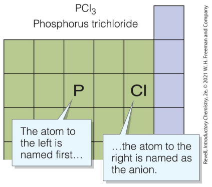
Compound name	Formula
Phosphorus monoxide	PO
Diphosphorus trioxide	P_2O_3
Diphosphorus tetroxide	P_2O_4
Tetraphosphorus decoxide	P_4O_{10}

Naming Binary Covalent Compounds

Atoms	Prefix
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

** omit for first element*

PCl_3 phosphorus trichloride
PCl_5 phosphorus pentachloride



Using Greek Prefixes

“pent” or “penta”

PCl_5	phosphorus pentachloride
P_2O_5	diphosphorus pentoxide

Remove “a” if anion begins with a vowel.

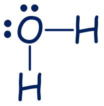
Practice Naming Covalent Compounds

Nitrogen and oxygen form two covalent compounds, NO_2 and N_2O_4. Name each of these compounds.

NO_2	nitrogen dioxide
N_2O_4	dinitrogen tetroxide

Summary of Covalent Compounds

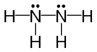
- In covalent bonds, atoms share electrons
- Covalent bonds form between nonmetals
- Most covalent compounds form discrete molecules
- We describe molecules using
  - Lewis structures
  - Molecular formulas
- Naming binary covalent compounds
  - Leftmost element first
  - Second element named as anion
  - Prefixes indicate the number of atoms present



Distinguishing Ionic and Covalent Compounds

To fulfill the octet rule, atoms
- gain or lose electrons (ions)
- share electrons (covalent bonds)

Covalent compounds
- share electrons
- between nonmetal atoms
- usually form molecules
- *molecular formula*



Ionic compounds
- oppositely-charged ions
- don't form molecules
- *formula unit* or *empirical formula*

Na^+ Cl^- Na^+ Cl^-
Cl^- Na^+ Cl^- Na^+
Na^+ Cl^- Na^+ Cl^-

Properties of Ionic and Covalent Compounds

Limestone ($CaCO_3$)

Olive Oil

Left to right: © drewrawcliffe/depositphotos.com; Luis Carlos Jimenez del rio/Shutterstock

Identifying and Naming Compounds

Covalent compounds
- all nonmetals

Ionic compounds
- metal + nonmetal
- contains polyatomic ions

Identify these compounds as ionic or covalent, and name each one:

MgF_2
ionic
magnesium fluoride

P_2O_4
covalent
diphosphorus tetroxide

$Fe(NO_3)_3$
ionic
iron(III) nitrate

SCl_6
covalent
sulfur hexachloride

Aqueous Solutions:
How Ionic and Covalent Compounds Differ

aqueous solution A homogeneous mixture, in which the main component is water

soluble Able to dissolve

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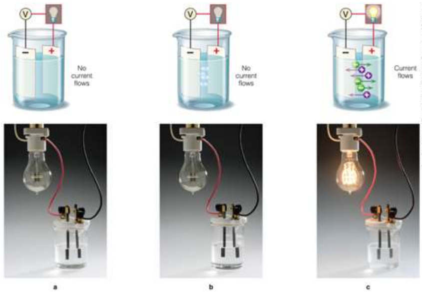
Electrolyte Solutions Conduct Electricity



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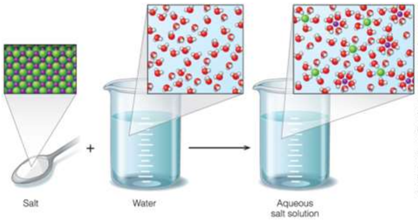
dissociation Ions are pulled apart in an aqueous solution



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Acids

covalent compounds that produce H^+ ions in aqueous solution

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

$$HNO_3 \longrightarrow H^+ + NO_3^-$$

Common Acids

Formula	Name	Formula	Name
HF	hydrofluoric acid	HNO_3	nitric acid
HCl	hydrochloric acid	HNO_2	nitrous acid
HBr	hydrobromic acid	H_2SO_4	sulfuric acid
HI	hydroiodic acid	H_3PO_4	phosphoric acid
H_2CO_3	carbonic acid	$HC_2H_3O_2$	acetic acid

Binary Acids

HF	hydrofluoric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid

Oxyacids

form H^+ and oxyanion

1. *-ate → -ic acid*

NO_3^-	nitrate	HNO_3	nitric acid
CO_3^{2-}	carbonate	H_2CO_3	carbonic acid
SO_4^{2-}	sulfate	H_2SO_4	sulfuric acid
PO_4^{3-}	phosphate	H_3PO_4	phosphoric acid

Oxyacids, Continued

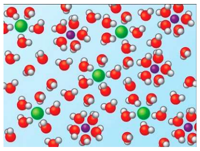
form H^+ and oxyanion

2. *-ite* → *-ous acid*

NO_2^-	nitrite	HNO_2	nitrous acid
ClO_2^-	chlorite	$HClO_2$	chlorous acid

Summary, Electrolytes

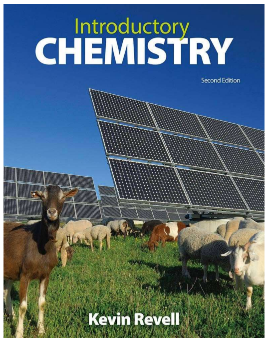
electrolytes
- ionic compounds
- acids (form H^+ ions in water)



Introductory Chemistry
Chem 103

Chapter 6 – Chemical Reactions

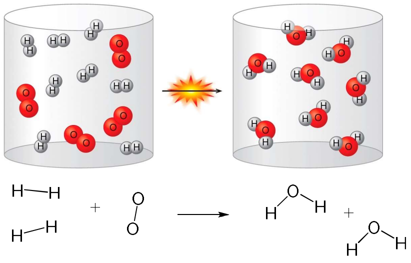
Lecture Slides



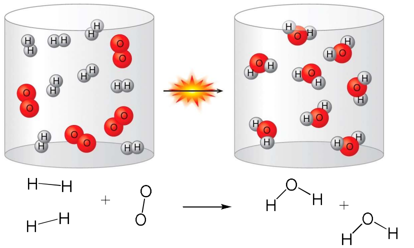
Chemical Equations

$$\underbrace{CaCO_3}_{\text{reactant}} \longrightarrow \underbrace{CaO + CO_2}_{\text{products}}$$

Chemical Equations Show Ratios of Substances



Chemical Equations Show Ratios of Substances, Continued



$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

subscript *coefficient*

The Ratios In a Chemical Reaction Are Constant

$2\ H_2$	+	O_2	$\longrightarrow$	$2\ H_2O$
2 molecules		1 molecule		2 molecules
2,000,000 molecules		1,000,000 molecules		2,000,000 molecules

*In a **balanced equation**, the number and type of each atom are the same on both sides of the arrow.*

Properly balanced – smallest whole-number ratio

Balancing Equations

*In a **balanced equation**, the number and type of each atom are the same on both sides of the arrow.*

$$Fe + O_2 \longrightarrow Fe_2O_3$$

Practice Balancing Equations

$$4\,Fe + 3\,O_2 \longrightarrow 2\,Fe_2O_3$$

☒	4 ~~1~~	Fe	4 ~~2~~	Fe
☒	6 ~~2~~	O	6 ~~3~~	O

1. Identify number and type on each side.
2. Add coefficients to balance atoms.
3. Do not change subscripts.

Practice Balancing Equations, Continued

$$Al_2O_3 + 3\,C + 3\,Cl_2 \longrightarrow 2\,AlCl_3 + 3\,CO$$

Al - 2
O - 3
C - ~~1~~3
Cl - ~~2~~6

Al - ~~1~~2
O - ~~1~~3
C - ~~1~~3
Cl - ~~3~~6

☒

Balance elemental forms last.

Strategies for Balancing Equations

balance polyatomic ions

$$Ni(NO_3)_2 + 2\,NaOH \longrightarrow Ni(OH)_2 + 2\,NaNO_3$$

nitrate: NO_3^-

hydroxide: OH^-

Strategies for Balancing Equations, Continued

use a fractional coefficient for diatomic molecules

$$\left(C_2H_6 + \frac{7}{2}\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O \right) x\ 2$$

need 7 oxygen atoms!

$$2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O$$

Equations with Phase Notations

phase notations: show phase or state of reaction components

$$CaCO_3\,(s) \longrightarrow CaO\,(s) + CO_2\,(g)$$

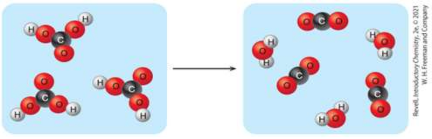
TABLE 6.1 Phase Symbols

Symbol	Meaning
(*s*)	Solid
(*l*)	Liquid
(*g*)	Gas
(*aq*)	Aqueous solution (dissolved in water)

Aqueous Solutions

***(aq)* –** indicates the substance is dissolved in water

$$H_2CO_3\,(aq) \longrightarrow H_2O\,(l) + CO_2\,(g)$$



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Chemical Equations Can Show Changes of State

$$H_2O\,(l) \longrightarrow H_2O\,(s)$$

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Classifying Reactions, Part 1

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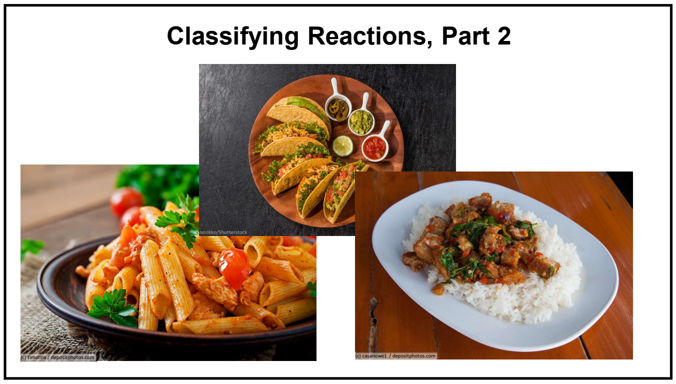

Classifying Reactions, Part 2

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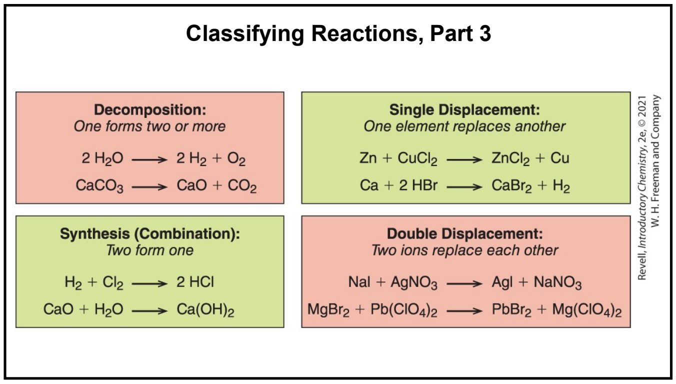

Classifying Reactions, Part 3

Decomposition:
One forms two or more

$2\ H_2O \longrightarrow 2\ H_2 + O_2$

$CaCO_3 \longrightarrow CaO + CO_2$

Single Displacement:
One element replaces another

$Zn + CuCl_2 \longrightarrow ZnCl_2 + Cu$

$Ca + 2\ HBr \longrightarrow CaBr_2 + H_2$

Synthesis (Combination):
Two form one

$H_2 + Cl_2 \longrightarrow 2\ HCl$

$CaO + H_2O \longrightarrow Ca(OH)_2$

Double Displacement:
Two ions replace each other

$NaI + AgNO_3 \longrightarrow AgI + NaNO_3$

$MgBr_2 + Pb(ClO_4)_2 \longrightarrow PbBr_2 + Mg(ClO_4)_2$

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Decomposition Reactions

$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$

Decomposition:
One forms two or more

Synthesis Reactions

$CaO\,(s) + H_2O\,(l) \rightarrow Ca(OH)_2\,(s)$

$H_2\,(g) + Cl_2\,(g) \rightarrow 2\,HCl\,(g)$

Synthesis (Combination):
Two form one

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Single Displacement Reactions

$Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$

$Sn\,(s) + 2\,HCl\,(aq) \rightarrow SnCl_2\,(aq) + H_2\,(g)$

Macmillan Learning

Single Displacement:
One element replaces another

Double Displacement Reactions

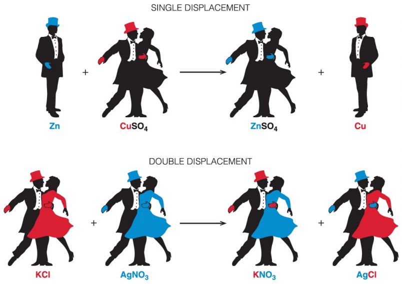
$KCl\ (aq) + AgNO_3\ (aq) \rightarrow KNO_3\ (aq) + AgCl\ (s)$

The anions “swap” positions

Macmillan Learning

Double Displacement
Two ions replace each other

Single and Double Displacement Reactions



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Classifying Reactions Summary

Decomposition:
One forms two or more

$2\ H_2O \longrightarrow 2\ H_2 + O_2$

$CaCO_3 \longrightarrow CaO + CO_2$

Single Displacement:
One element replaces another

$Zn + CuCl_2 \longrightarrow ZnCl_2 + Cu$

$Ca + 2\ HBr \longrightarrow CaBr_2 + H_2$

Synthesis (Combination):
Two form one

$H_2 + Cl_2 \longrightarrow 2\ HCl$

$CaO + H_2O \longrightarrow Ca(OH)_2$

Double Displacement:
Two ions replace each other

$NaI + AgNO_3 \longrightarrow AgI + NaNO_3$

$MgBr_2 + Pb(ClO_4)_2 \longrightarrow PbBr_2 + Mg(ClO_4)_2$

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Reactions between Metals and Nonmetals

Metal + Nonmetal → Ionic Compound

$\underbrace{\text{metal cation} + \text{nonmetal anion}}$

$2\,Al + 3\,Br_2 \rightarrow 2\,AlBr_3$


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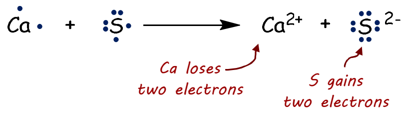
oxidation – loss of electrons

reduction – gain of electrons

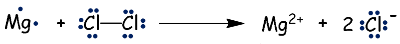
Reactions between Metals and Nonmetals Example 1

$Ca\ (s) + S\ (s) \rightarrow CaS\ (s)$



Reactions between Metals and Nonmetals Example 2

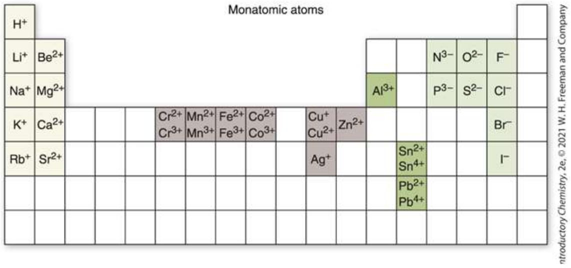
$Mg\ (s) + Cl_2\ (g) \rightarrow MgCl_2\ (s)$



Mg loses two electrons

each Cl gains one electron

Metals and Nonmetals Form Specific, Stable Ions.



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Reactions Between Metals and Nonmetals Practice

What compound is formed when aluminum metal reacts with chlorine gas? Write a balanced equation for this reaction.

Al^{3+} Cl^- Cl^- Cl^-

$AlCl_3$

$$2\,Al + 3\,Cl_2 \rightarrow 2\,AlCl_3$$

Reactions Between Metals and Nonmetals, More Practice

When tin metal reacts with bromine, it is oxidized to the tin(IV) ion, while bromine is reduced to form bromide ions. Write a balanced equation for this reaction.

Sn^{4+} Br^-

$SnBr_4$

$$Sn + 2\,Br_2 \rightarrow SnBr_4$$

Combustion Reactions

reactions in which oxygen gas combines with elements or compounds to produce oxides.

$Sn + O_2 \rightarrow SnO_2$

tin(IV) oxide - ionic

$C + O_2 \rightarrow CO_2$

carbon dioxide - covalent

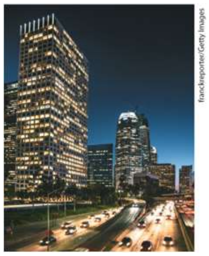
$S + O_2 \rightarrow SO_2$

sulfur dioxide - covalent

Hydrocarbons compounds composed of hydrogen and carbon

TABLE 6.2 Common Hydrocarbons

Formula	Name	Use
CH_4	Methane	Natural gas
C_2H_2	Acetylene	Torches for cutting and welding
C_2H_4	Ethylene	Manufacture of plastic
C_3H_8	Propane	Natural gas component; used for heating and power
C_4H_{10}	Butane	Lighter fluid
C_6H_6	Benzene	Solvent; precursor for many pharmaceutical compounds
C_8H_{18}	Octane	Component of gasoline



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Combustion of Hydrocarbons

hydrocarbon + oxygen → carbon dioxide + water

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

$$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

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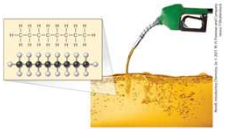

The Combustion of Sulfur Produces Sulfur Oxides

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Combustion Reactions Practice

Write a balanced equation for the combustion of calcium metal.

$2\ Ca + O_2 \longrightarrow 2\ CaO$

Ca^{2+} O^{2-}

CaO

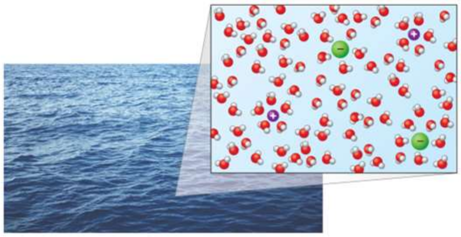
Combustion Reactions, More Practice

Write a balanced equation for the combustion of propane gas, a common fuel used for home heating, cooking, etc. The formula for propane is C_3H_8.

$C_3H_8 + 5\ O_2 \longrightarrow 3\ CO_2 + 4\ H_2O$

Reactions in Aqueous Solution

Ionic compounds **dissociate** when dissolved in water.



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Comparing Molecular and Ionic Equations

molecular equation – shows ions together as compounds

$KBr_{(s)} \rightarrow KBr_{(aq)}$

ionic equation – shows dissociated ions as separate species

$KBr_{(s)} \rightarrow K^+_{(aq)} + Br^-_{(aq)}$

Writing Ionic Equations Practice

Show this process as an ionic equation:

$Mg(NO_3)_{2\,(s)} \rightarrow Mg(NO_3)_{2\,(aq)}$

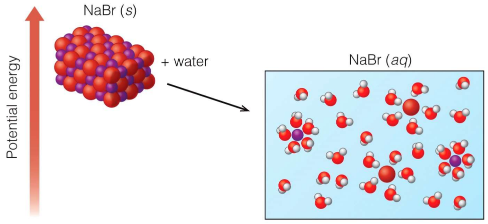
$Mg(NO_3)_{2\,(s)} \rightarrow Mg^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)}$

Predicting Solubility

Many ionic compounds are **insoluble** in water.

Turtle Rock Scientific/Science Source

Predicting Solubility, Continued



Factors affecting solubility

- Charge on Ions
- Size of Ions
- How tightly ions pack together

Soluble	Insoluble
NaCl (Na$^+$ and Cl$^-$)	Fe_2O_3 (Fe^{3+} and O^{2-})
KNO_3 (K$^+$ and NO_3^-)	PbS (Pb^{2+} and S^{2-})
NH_4Br (NH_4^+ and Br$^-$)	$BaCO_3$ (Ba^{2+} and CO_3^{2-})

Solubility Rules:

- Halogens (F^-, Br^-, Cl^-, I^-) are soluble
  - – Unless bonded to Ag^+ or Pb^{2+}

Soluble	**Insoluble**
KF	AgF
$ZnCl_2$	AgCl
$FeBr_2$	$PbBr_2$
CuI	PbI_2

Solubility Rules, Continued

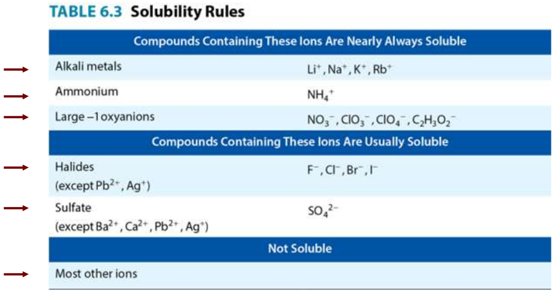
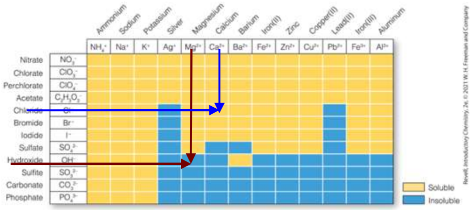


TABLE 6.3 Solubility Rules

Compounds Containing These Ions Are Nearly Always Soluble	
→ Alkali metals	Li^+, Na^+, K^+, Rb^+
→ Ammonium	NH_4^+
→ Large –1 oxyanions	NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$
Compounds Containing These Ions Are Usually Soluble	
→ Halides (except Pb^{2+}, Ag^+)	F^-, Cl^-, Br^-, I^-
→ Sulfate (except Ba^{2+}, Ca^{2+}, Pb^{2+}, Ag^+)	SO_4^{2-}
Not Soluble	
→ Most other ions	

Solubility Tables



		Ammonium	Sodium	Potassium	Silver	Magnesium	Calcium	Barium	Iron(II)	Zinc	Copper(II)	Lead(II)	Iron(III)	Aluminum
		NH_4^+	Na^+	K^+	Ag^+	Mg^{2+}	Ca^{2+}	Ba^{2+}	Fe^{2+}	Zn^{2+}	Cu^{2+}	Pb^{2+}	Fe^{3+}	Al^{3+}
Nitrate	NO_3^-	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨
Chlorate	ClO_3^-	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨
Perchlorate	ClO_4^-	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨
Acetate	$C_2H_3O_2^-$	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨	🟨
Chloride	Cl^-	🟨	🟨	🟨	🟦	🟨	🟨	🟨	🟨	🟨	🟨	🟦	🟨	🟨
Bromide	Br^-	🟨	🟨	🟨	🟦	🟨	🟨	🟨	🟨	🟨	🟨	🟦	🟨	🟨
Iodide	I^-	🟨	🟨	🟨	🟦	🟨	🟨	🟨	🟨	🟨	🟨	🟦	🟨	🟨
Sulfate	SO_4^{2-}	🟨	🟨	🟨	🟦	🟨	🟦	🟦	🟨	🟨	🟨	🟦	🟨	🟨
Hydroxide	OH^-	🟨	🟨	🟨	🟦	🟦	🟦	🟨	🟦	🟦	🟦	🟦	🟦	🟦
Sulfite	SO_3^{2-}	🟨	🟨	🟨	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦
Carbonate	CO_3^{2-}	🟨	🟨	🟨	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦
Phosphate	PO_4^{3-}	🟨	🟨	🟨	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦	🟦

🟨 Soluble
🟦 Insoluble

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Ex.: $CaCl_2$ *soluble* $Mg(OH)_2$ *insoluble*

Determine Solubility

Determine whether the following compounds are soluble or insoluble in water:

Na_3PO_4	$AlCl_3$	$CaCO_3$
soluble	soluble	insoluble

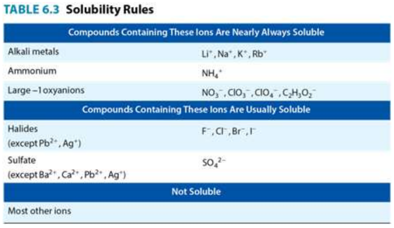


TABLE 6.3 Solubility Rules

Compounds Containing These Ions Are Nearly Always Soluble	
Alkali metals	Li^+, Na^+, K^+, Rb^+
Ammonium	NH_4^+
Large −1 oxyanions	NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$
Compounds Containing These Ions Are Usually Soluble	
Halides (except Pb^{2+}, Ag^+)	F^-, Cl^-, Br^-, I^-
Sulfate (except Ba^{2+}, Ca^{2+}, Pb^{2+}, Ag^+)	SO_4^{2-}
Not Soluble	
Most other ions	

Precipitation Reactions

precipitation reaction	two aqueous solutions produce an insoluble product
precipitate	the solid product formed in the reaction

$$Pb(NO_3)_2\ (aq) + 2\ NaI\ (aq) \rightarrow PbI_2\ (s) + 2\ NaNO_3\ (aq)$$

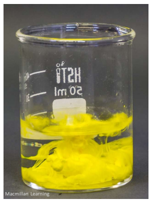
$Pb(NO_3)_2$ *(aq)*

NaI *(aq)*

Precipitation Reactions Are Double Displacement Reactions

$$Pb(NO_3)_2\ (aq) + 2\ NaI\ (aq) \rightarrow PbI_2\ (s) + 2\ NaNO_3\ (aq)$$

The anions "swap" positions

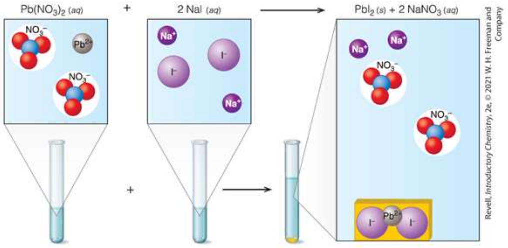


How Precipitation Reactions Occur

$Pb^{2+}(aq) + 2\ NO_3^-(aq) + 2\ Na^+(aq) + 2\ I^-(aq) \rightarrow 2\ Na^+(aq) + 2\ NO_3^-(aq) + PbI_2(s)$

spectator ions   spectator ions

Driving force - formation of the solid



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Comparing Complete and Net Ionic Equations

Complete ionic equation

shows all ions present

$Pb^{2+}(aq) + 2\ NO_3^-(aq) + 2\ Na^+(aq) + 2\ I^-(aq) \rightarrow 2\ Na^+(aq) + 2\ NO_3^-(aq) + PbI_2(s)$

spectator ions   spectator ions

Net ionic equation

Only include ions involved in the precipitation

$Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$

Writing Precipitation Reactions

Three ways to show a precipitation reaction:

Molecular Equation

shows neutral compounds

$Pb(NO_3)_2\ (aq) + 2\ KCl\ (aq) \rightarrow PbCl_2\ (s) + 2\ KNO_3\ (aq)$

Complete Ionic Equation

shows all ions present

$Pb^{2+}(aq) + 2\ NO_3^-(aq) + 2\ K^+(aq) + 2\ Cl^-(aq) \rightarrow PbCl_2\ (s) + 2\ K^+(aq) + 2\ NO_3^-(aq)$

Net Ionic Equation

Omits spectator ions; only shows ions that react.

$Pb^{2+}(aq) + 2\ Cl^-(aq) \rightarrow PbCl_2\ (s)$

Use solubility rules to predict precipitation reactions.

Precipitation Reactions Practice

When aqueous silver acetate is combined with aqueous barium chloride, a white precipitate forms. Write balanced complete ionic, net ionic, and molecular equations to show the reaction that takes place. Include phase symbols.

silver acetate solution: $Ag^{+}\,(aq) + C_2H_3O_2^{-}\,(aq)$

barium chloride solution: $Ba^{2+}\,(aq) + 2\ Cl^{-}\,(aq)$

$AgCl\ (s)$

Complete ionic equation

$2\ Ag^{+}(aq) + 2\ C_2H_3O_2^{-}(aq) + Ba^{2+}(aq) + 2\ Cl^{-}(aq) \rightarrow Ba^{2+}(aq) + 2\ C_2H_3O_2^{-}(aq) + 2\ AgCl(s)$

Net ionic equation

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

Molecular equation

$2\ AgC_2H_3O_2\ (aq) + BaCl_2\ (aq) \rightarrow Ba(C_2H_3O_2)_2\ (aq) + 2\ AgCl\ (s)$

Summary of Precipitation Reactions

- Soluble ionic compounds dissociate in water.
- Some ionic compounds are insoluble in water.
- Solubility rules predict the solubility of compounds.
- Precipitation reaction: two solutions combine to produce an insoluble product.
- We describe reactions in solution using
  - molecular equations
  - complete ionic equations
  - net ionic equations

Reactions in Aqueous Solution

acids compounds that produce H^{+} ions in aqueous solution

TABLE 6.4 Common Acids

Formula	**Name**
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
H_2CO_3	Carbonic acid
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic acid

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

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

Reactions in Aqueous Solution, Continued

bases compounds that produce OH^- ions in aqueous solution

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

TABLE 6.5 Common Hydroxide Base

Formula	**Name**
$LiOH$	Lithium hydroxide
$NaOH$	Sodium hydroxide
KOH	Potassium hydroxide
$Ba(OH)_2$	Barium hydroxide

Neutralization Reactions

Acids and bases undergo **neutralization reactions.**

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

acid + base → water + salt

Ex.: hydrochloric acid reacts with sodium hydroxide

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

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

Ex.: nitric acid reacts with lithium hydroxide

$HNO_3\ (aq) + LiOH\ (aq) \rightarrow H_2O\ (l) + LiNO_3\ (aq)$

a "salt"

Neutralization Reactions, Continued

Acid-base neutralization is a **double displacement reaction**.

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

acid + base → water + salt

The formation of water is the driving force for the reaction.

Acid-Base Reactions Practice

Write a balanced equation to show the reaction of sulfuric acid with sodium hydroxide. Include phase symbols.

acid + base → water + salt

$\underline{H}_2SO_4 + 2\ Na\underline{OH} \rightarrow 2\ H_2O + Na_2SO_4$

$\boxed{H_2SO_4\ (aq) + 2\ NaOH\ (aq) \rightarrow 2\ H_2O\ (l) + Na_2SO_4\ (aq)}$