General Chemistry

402101-4

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General Chemistry 402101-4

Lecture Chapter	Topics	Chang Chapters	Weekly schedule
1	Introduction to Chemistry	CH1	Week 1
1	States of matter, measurements, precision and accuracy, and significant figures	CH1	Week 2
2	Atoms, quantum numbers and electron configurations	CH2, CH7	Week 3
3	The periodic table: Chemical properties of elements in the periodic table	CH8	Week 4
4	Atomic weight, molecular weight, moles and mass percent calculations.	CH3	Week 5
5	Chemical reactions in solutions: Concentration calculations, chemical equations, and types of chemical reactions	CH3, CH4, CH12	Week 6
6	Chemical equilibrium: Equilibrium constant calculations	CH14	Week 7
	Midterm Exam	Week 8	
6	Chemical equilibrium: Factors affecting chemical equilibrium	CH14	Week 9
7	Acids and bases, and pH calculations	CH15	Week 10
8	Thermochemistry: Introduction to thermodynamics, and calculation of heat capacity	CH6	Week 11
8	Thermochemistry: Enthalpy of reaction calculation	CH6	Week 12
9	Organic chemistry: Hydrocarbons, and alkane nomenclature and reactions	CH24	Week 13
	Review		Week 14

<u>Primary reference:</u> "Chemistry," R. Chang, McGraw-Hill Higher Education

<u>Grading</u>: The midterm exam will account for 20%, final exam for 40%, lab for 30%, and quizzes or scientific activities for 10% of the final grade.

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Matter
Measurements
Significant figures

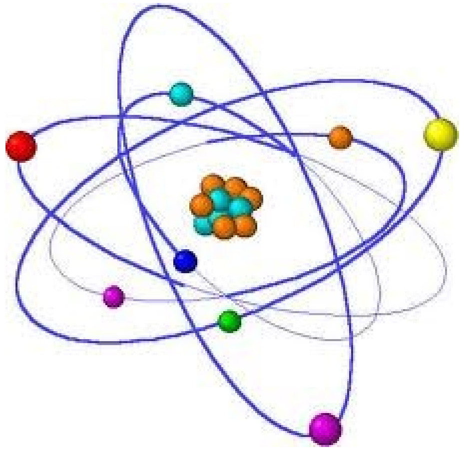
Chapter
1

Chang-chapter1

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

First lecture

What is Chemistry?



The study of matter and the changes it undergoes

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1st lecture

Matter

Any thing that occupies space and has mass

Separation by physical methods

pure substance ← **mixture**

Has a definite composition and distinct properties

A combination of two or more substances in which the substances retain their distinct identities

Separation by chemical methods

element ← **compound**

heterogeneous

homogeneous

cannot be separated into simpler substances by chemical means

composed of two different elements or more chemically united in fixed proportions.

The composition is not uniform

The composition is the same throughout

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NaCl	
Salt water	
Iron	
sugar	
air	
helium	
water	
salad	

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compound **element** **homogeneous mixture** **heterogeneous**

NaCl	compound
Salt water	homogeneous mixture
Iron	element
sugar	compound
air	homogeneous mixture
helium	element
water	compound
salad	heterogeneous mixture

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compound element homogeneous mixture heterogeneous

- **Element**
  - **Atoms**: C, Al, Ti
  - **Molecules**
    - **Polyatomic**: S_8, O_3
    - **Diatomic**: O_2 , H_2
- **Compound**
  - **Molecules containing different two atoms or more**: H_2O, H_2SO_3

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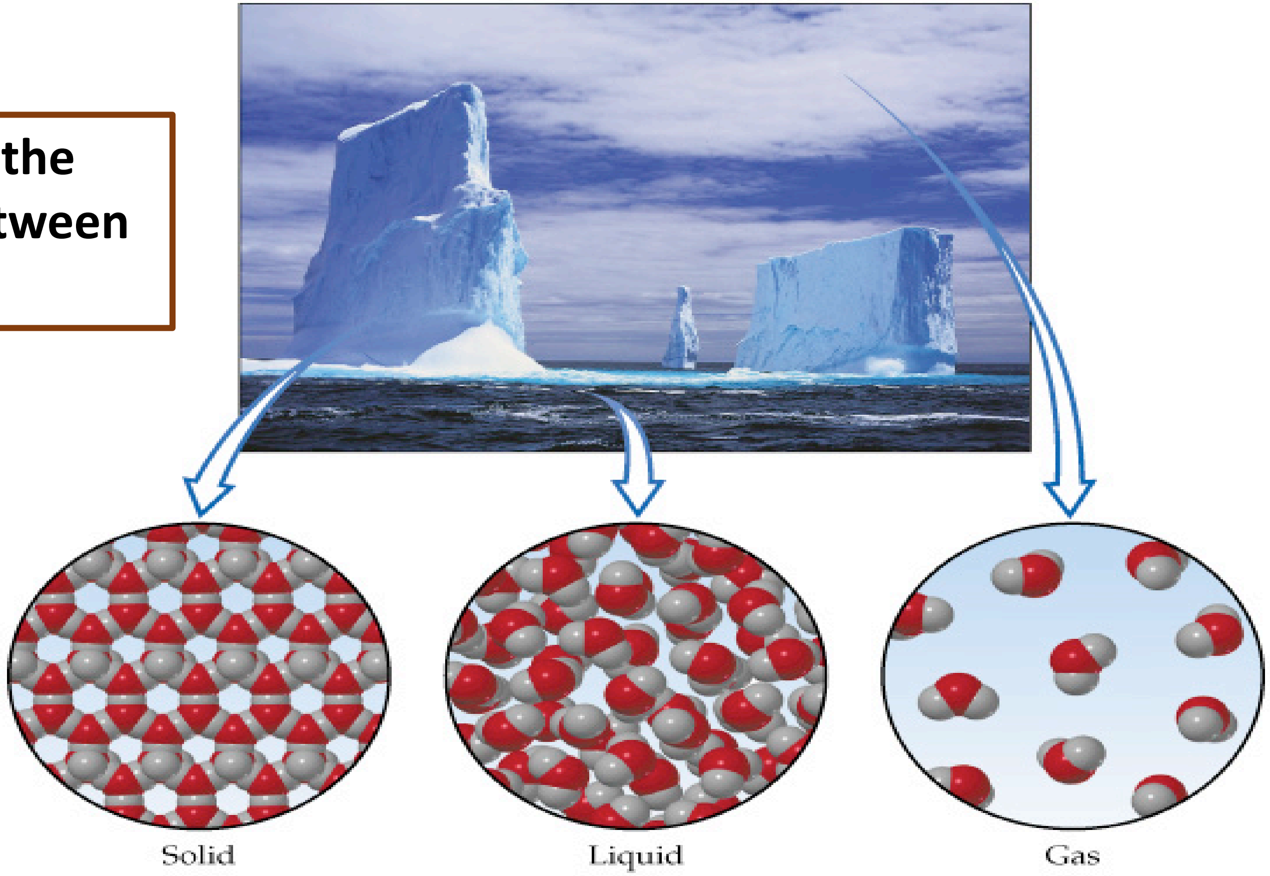
Matter States

The difference between the states is the distance between the molecules.



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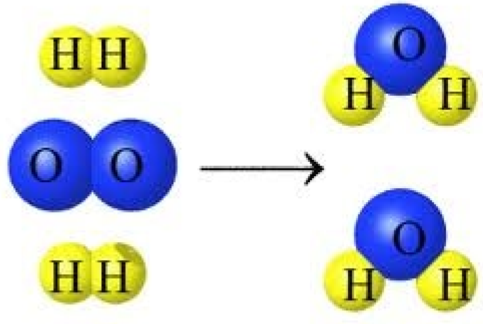
Matter properties

Chemical

is a property when the matter undergoes a chemical change or reaction

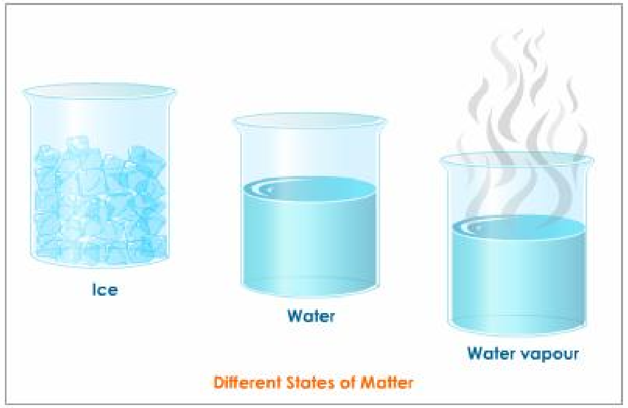


reactivity,
flammability

Physical

Can be measured and observed without changing the composition or identity of a substances

color,
mass,
size



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Measurable properties of matter

Extensive	Intensive
depends on how much matter is being considered	Does not depend on how much matter is being considered
Mass	Density
volume	temperature

How can these properties be measured ?

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Measurement

SI Units

International system of units

Base Quantity	Name of unit	Symbol
Length		
Mass		
Time		
Electrical current		
Temperature		
Amount of substance		
Luminous intensity		

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Measurement

SI Units

International system of units

Base Quantity	Name of unit	Symbol
Length	meter	m
Mass	Kilogram	Kg
Time	Second	s
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	candela	cd

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Prefixes Used with SI Units

Prefix	Symbol	Multiple of Base Unit
Giga	G	1,000,000,000 or 10^9
Mega	M	1,000,000 or 10^6
kilo	k	1,000 or 10^3
deci	d	0.1 or 10^{-1}
centi	c	0.01 or 10^{-2}
milli	m	0.001 or 10^{-3}
micro	m	0.000001 or 10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}
Femto	f	10^{-15}

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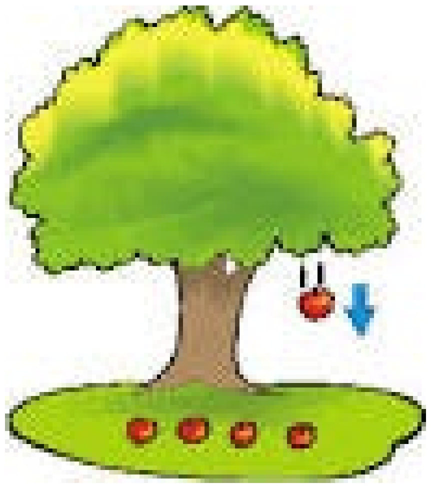
Mass and weight



What is the difference between mass and weight?

Mass: is a measure of amount of matter in an object

$$1\ \text{Kg} = 1000\ \text{g} = 1 \times 10^{3}\ \text{g}$$

Weight: is the force that gravity exerts on an object

Newton (N)

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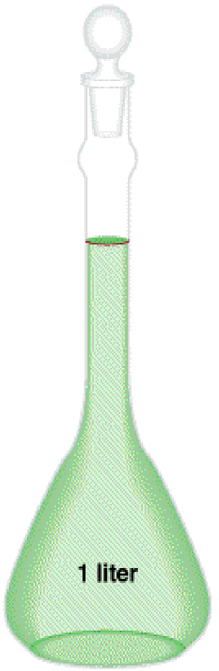
Volume

Volume – SI derived unit for volume is cubic meter (m^3)

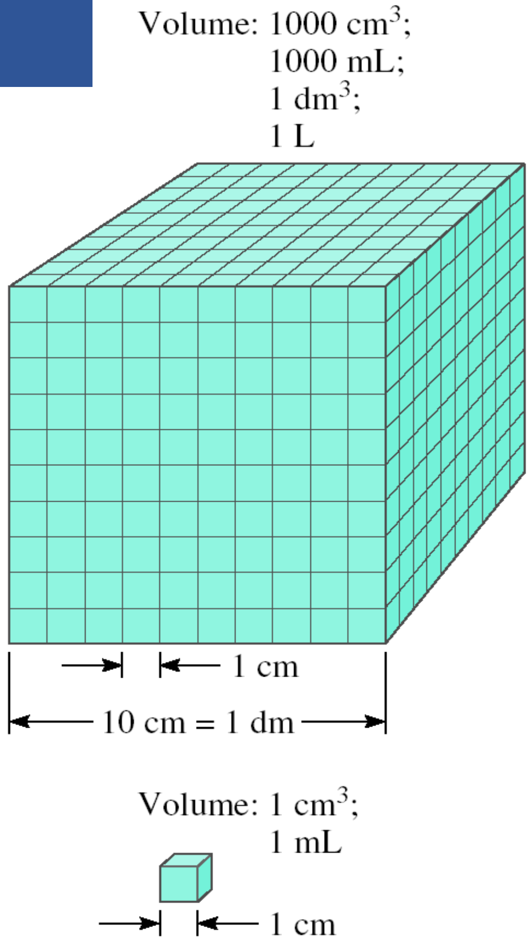


Volumetric flask

$$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$

$$\boxed{1\ cm^3 = 1\ mL}$$

$$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$$



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Dimensional Analysis Method of Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

$$1.63\ \cancel{L} \times \frac{1000\ mL}{1\cancel{L}} = 1630\ mL$$

$$\cancel{1.63\ L \times \frac{1L}{1000\ mL} = 0.001630\ \frac{L^2}{mL}}$$

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Density

Density is defined as <u>the mass per unit volume</u>.

density = mass/volume $\qquad d = \dfrac{m}{V}$ $\qquad$ S.I. units for density = **kg/m^3**

g/cm^3 for solids

g/ml for liquids

g/L for gases

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Density

A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?

$$d = \frac{m}{V}$$

$$m = d \times V$$

$$m = 21.5\ g/\cancel{cm^3} \times 4.49\ \cancel{cm^3} = 96.5\ g$$

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1st lecture

Temperature

Temperature scales

Fahrenheit
°F

$$^{0}F = \left(\frac{9}{5} \times {}^{0}C\right) + 32$$

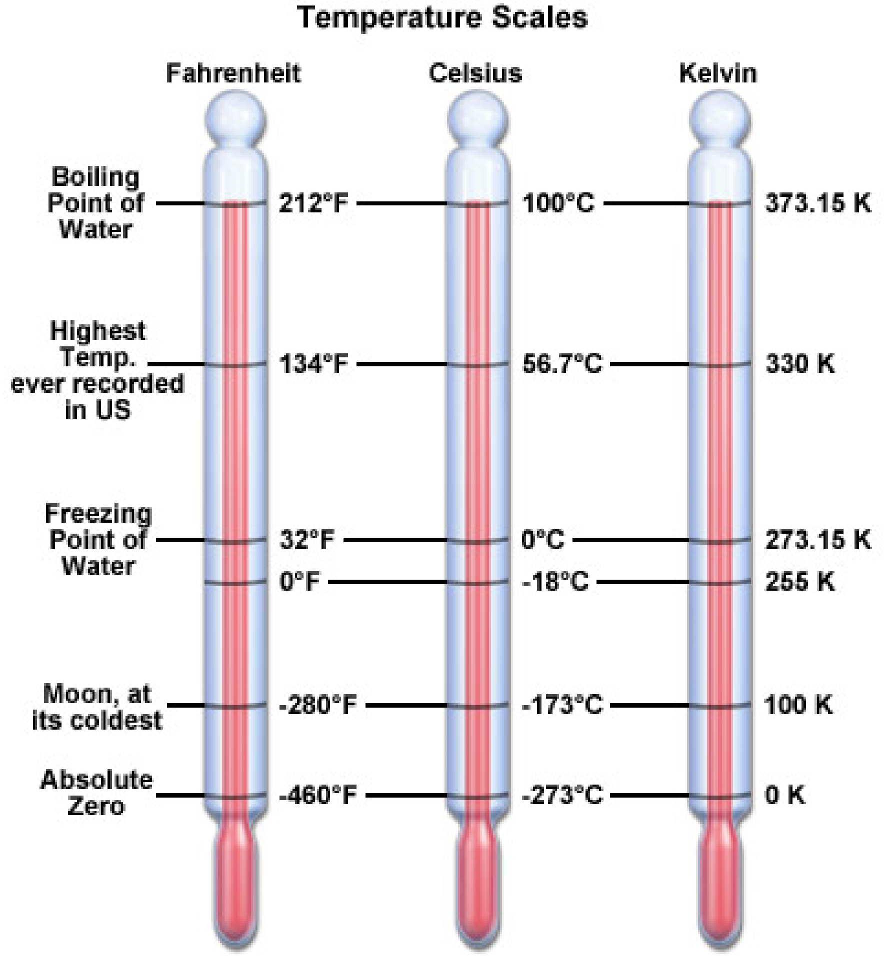
32 ^{0}F = 0 ^{0}C
212 ^{0}F = 100 ^{0}C

Celsius
°C

273 K = 0 ^{0}C
373 K = 100 ^{0}C

Kelvin
K

$$T(\text{in Kelvin}) = T(\text{in Celsius}) + 273.15$$



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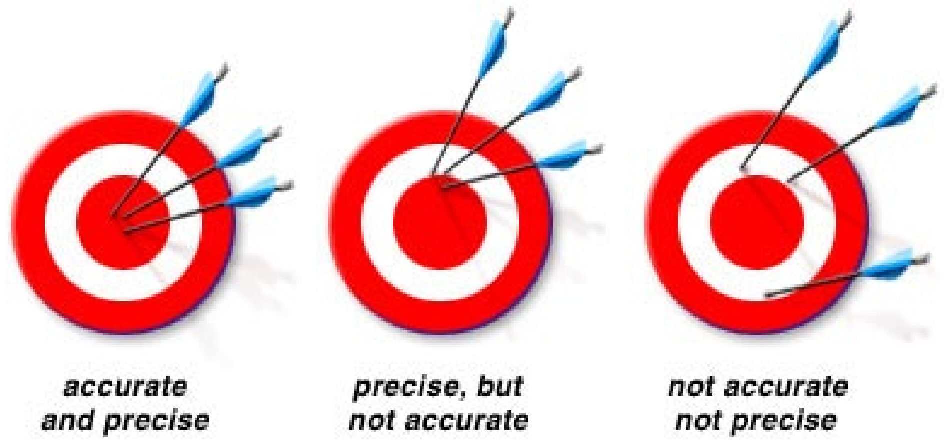
1st lecture

Precision and Accuracy



	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average	1.971 g	1.970 g	2.001 g

The true mass of object= 2.000 g

Precision: How close a set of measurements are to each other (reproducibility).

Accuracy: How close your measurements are to the true value.

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Significant Figures

- Any digit that is not zero is significant

  1.234 kg 4 significant figures

- Zeros between nonzero digits are significant

  606 m 3 significant figures

- Zeros to the left of the first nonzero digit are not significant

  0.08 L 1 significant figure

- If a number is greater than 1, then all zeros to the right of the decimal point are significant

  2.0 mg 2 significant figures

- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

  0.00420 g 3 significant figures

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Scientific Notation

$\boxed{N \times 10^{n}}$

N is a number between 1 and 10

n is a positive or negative integer

The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

6.022×10^{23}

The mass of a single carbon atom in grams:

0.0000000000000000000000199

1.99×10^{-23}

568.762 = 5.68762 $\times 10^{2}$ **(6 SF)**

0.00000772 = 7.72 $\times 10^{-6}$ **(3 SF)**

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How many significant figures are in each of the following measurements?

1) 24 ml
2) 3001 g
3) 0.0320 m^3
4) 6.4 $\times$ 10^4 molecules
5) 560 kg

- 2 significant figures
- 4 significant figures
- 3 significant figures
- 2 significant figures
- 3 significant figures– to clarify use the scientific notation 5.60×10^2 kg

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Tip: start to count the sig. fig. from the left when you see a non zero number until the end of the number.

Significant Figures: Addition & Subtraction

If addition or subtraction:

1- must have same power before addition or subtraction

2- sig. fig. in the answer is as the smaller digits after decimal point

4.31×10^4

$+$

3.9×10^3 (0.39×10^4)

$= 4.70 \times 10^4$ (3 SF)

7.4×10^3

$+$ (1 decimal digit: this has the smallest digit)

0.10×10^3

$= 7.5 \times 10^3$ (2 SF)

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Significant Figures: Multiplication & Division

If multiplication or division:

1- add exponent for multiplication or subtract exponent for division

2- write the answer with the smaller sig. fig.

$$(8.0 \times 10^4) \; . \; (5.00 \times 10^2) = 40 \times 10^6 \text{ or } 4.0 \times 10^7$$

(2 SF) (3 SF) (2 SF) (2 SF)

$$4.51 \text{ x } 3.6666 = 16.53636 \approx 16.5$$

(3 sf) (5 sf) (3 sf)

$$6.8 \div 112.04 = 0.0606926 \approx 0.061$$

(2 sf) (5 sf) (2 sf)

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Significant Figures

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures

The average of three measured lengths; 6.64, 6.68 and 6.70?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 \qquad = \cancel{7}$$

Because 3 is an *exact number*

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Question 1

Which of the following is an example of a physical property?

A) combustibility

B) corrosiveness

C) explosiveness

D) density

E) A and D

Question 2

Which of the following represents the greatest mass?

A) 2.0×10^3 mg

B) 10.0 dg

C) 0.0010 kg

D) 1.0×10^6 µg

E) 3.0×10^{12} pg

Question 3

Convert 240 K and 468 K to the Celsius scale.

A) 513°C and 741°C

B) -59°C and 351°C

C) -18.3°C and 108°C

D) -33°C and 195°C

Question 4

Calculate the volume occupied by 4.50×10^2 g of gold (density = 19.3 g/cm^3).

A) 23.3 cm^3

B) 8.69×10^3 cm

C) 19.3 cm^3

D) 450 cm^3

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Question 5

The melting point of bromine is -7°C. What is this melting point expressed in °F?

A) 45°F

B) -28°F

C) -13°F

D) 19°F

E) None of these is within 3°F of the correct answer.

Question 6

How many significant figures are there in the measurement 3.4080 g?

A) 6

B) 5

C) 4

D) 3

E) 2

Question 7

How many significant figures should you report as the sum of 8.3801 + 2.57?

A) 3

B) 5

C) 7

D) 6

E) 4

Question 8

How many significant figures are there in the number 0.0203610 g?

A) 8

B) 7

C) 6

D) 5

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Question 9

The value of 345 mm is a measure of

A) temperature

B) density

C) volume

D) distance

E) Mass

Question 10

The measurement 0.000 004 3 m, expressed correctly using scientific notation, is

A) 0.43×10^{-5} m

B) 4.3×10^{-6}

C) 4.3×10^{-7}

D) 4.3×10^{-5}

Question 11

A laboratory technician analyzed a sample three times for percent iron and got the following results: 22.43% Fe, 24.98% Fe, and 21.02% Fe. The actual percent iron in the sample was 22.81%. The analyst's

A) precision was poor but the average result was accurate.

B) accuracy was poor but the precision was good.

C) work was only qualitative.

D) work was precise.

E) C and D.

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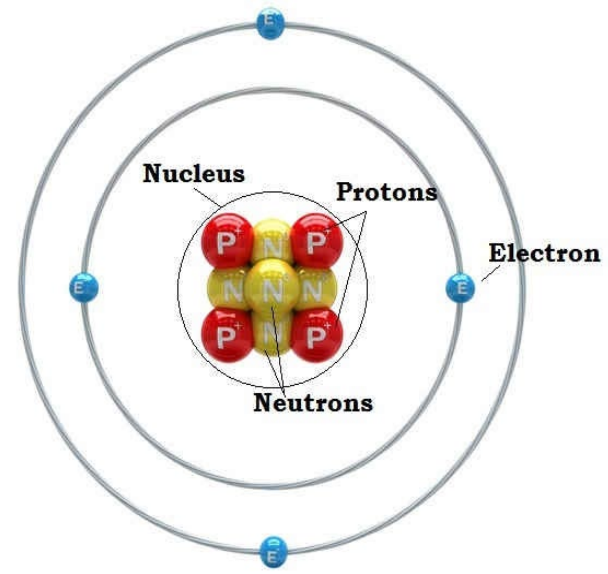


Atoms, Quantum numbers & Electron configurations

Chapter 2

Chang-chapter2,7

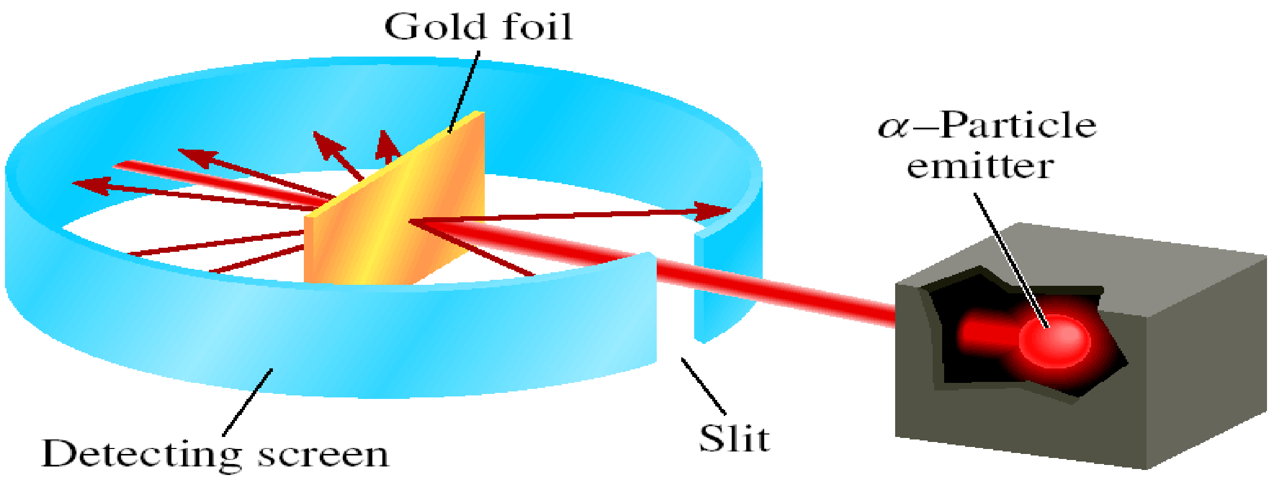
COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

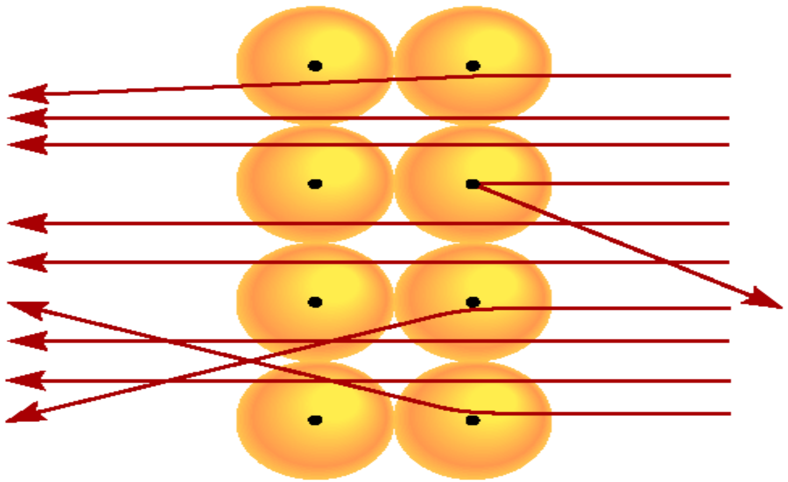
Dalton's Atomic Theory (1808)

1. Elements are composed of extremely small particles called ***atoms***.
2. All ***atoms*** of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. ***Compounds*** are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.
4. A ***chemical reaction*** involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Rutherford's Experiment

(1908 Nobel Prize in Chemistry)





α particle velocity ~ 1.4×10^7 m/s
(~5% speed of light)

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1. atoms positive charge is concentrated in the nucleus
2. proton (p) has opposite (+) charge of electron (-)
3. mass of p is 1840 x mass of e^- (1.67×10^{-24} g)

TABLE 2.1 Mass and Charge of Subatomic Particles

		Charge	
Particle	**Mass (g)**	**Coulomb**	**Charge Unit**
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

*More refined measurements have given us a more accurate value of an electron's mass than Millikan's.

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mass p ≈ mass n ≈ 1840 x mass e^-

Atomic number, Mass number and Isotopes

Atomic number (Z) = number of protons in nucleus

Mass number (A) = number of protons + number of neutrons

= atomic number (Z) + number of neutrons

Isotopes are atoms of the same element (X) with different numbers of neutrons in their nuclei

Mass Number → $^{A}_{Z}X$ ← Element Symbol

Atomic Number →

$^{1}_{1}H$ $^{2}_{1}H$ (D) $^{3}_{1}H$ (T)

$^{235}_{92}U$ $^{238}_{92}U$

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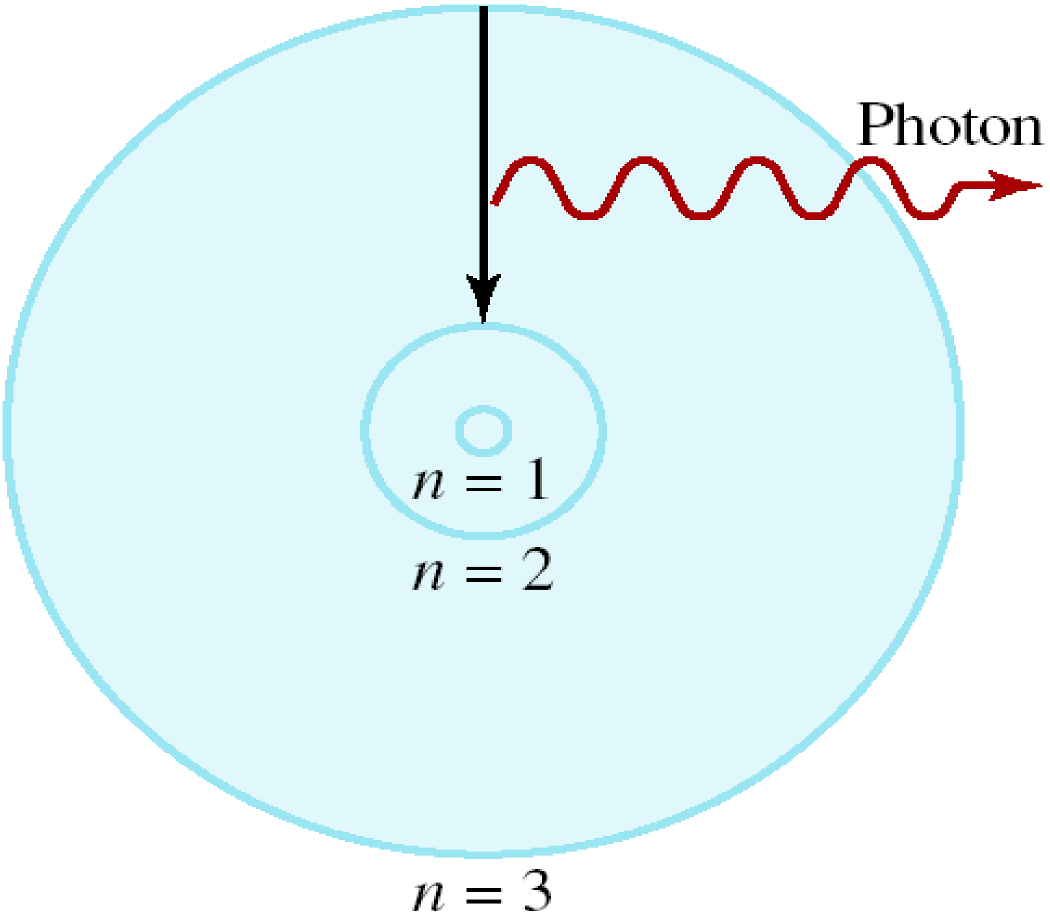
Bohr's Model of the Atom (1913)

1. e^- can only have specific (quantized) energy values
2. light is emitted as e^- moves from one energy level to a lower energy level

$$E_n = -R_H \, (1/n^2)$$

n (principal quantum number) = 1,2,3,…

R_H (Rydberg constant) = 2.18×10^{-18} J

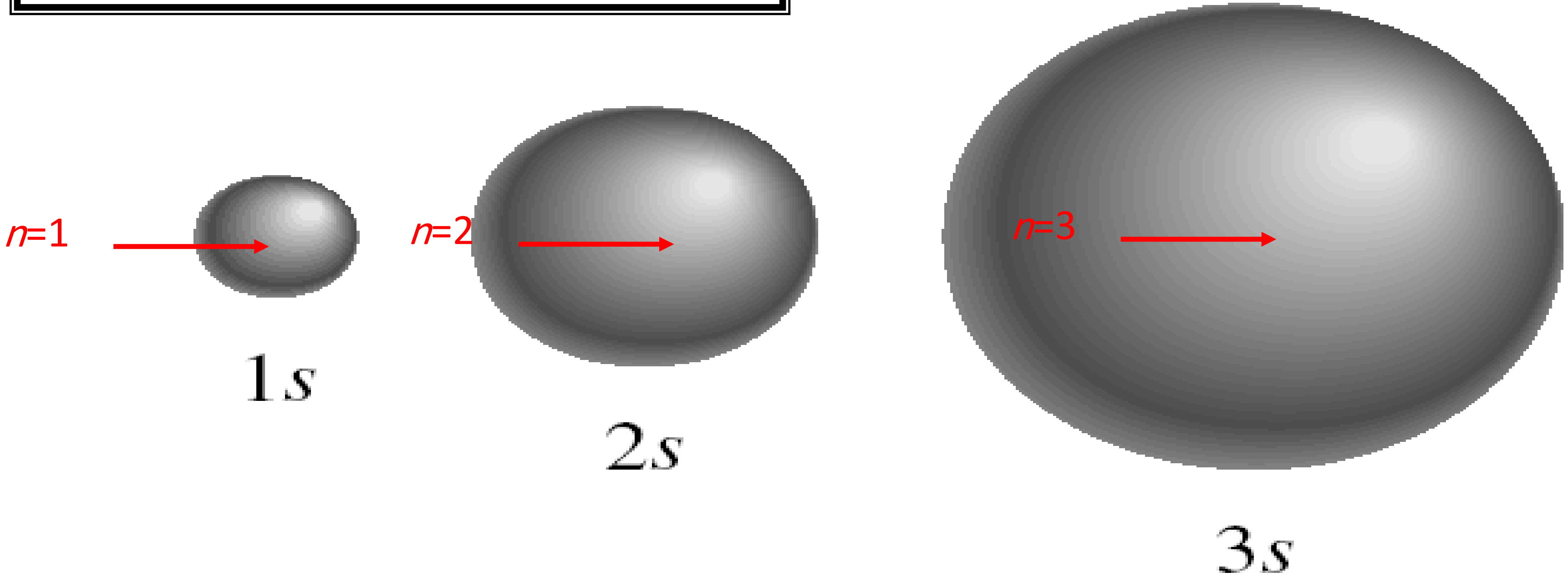


Quantum numbers (n, l, m_l, m_s)

principal quantum number (*n*)

n = 1, 2, 3, 4, ….

distance of e⁻ from the nucleus



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Angular momentum quantum number (l)

for a given value of n, $l = 0, 1, 2, 3, \ldots n-1$

$n = 1$, $l = 0$
$n = 2$, $l = 0$ **or** 1
$n = 3$, $l = 0, 1,$ **or** 2

$l = 0$ s orbital
$l = 1$ p orbital
$l = 2$ d orbital
$l = 3$ f orbital

Shape of the "volume" of space that the e^- occupies

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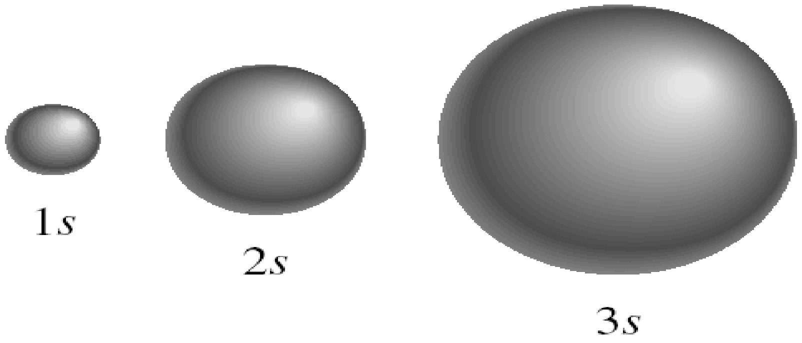
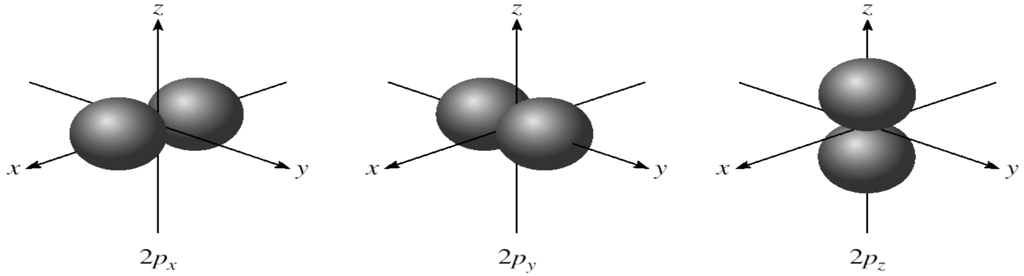
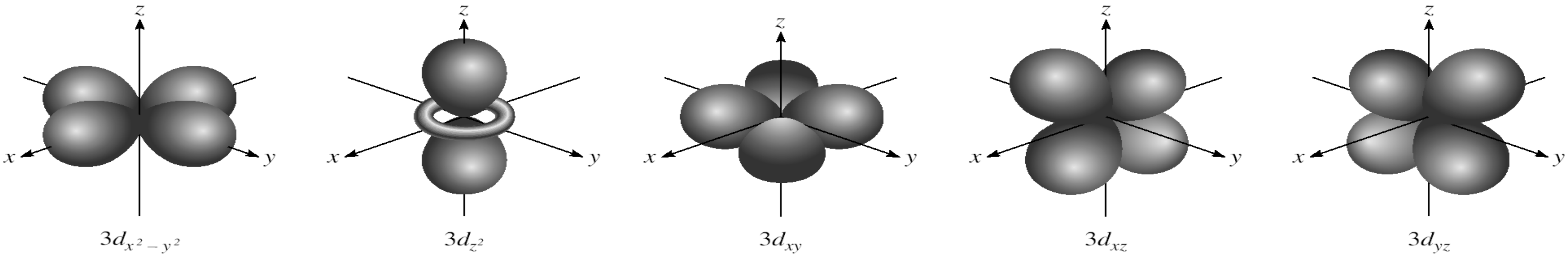
l = 0 (*s* orbitals)



l = 1 (*p* orbitals)





l = 2 (*d* orbitals)

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magnetic quantum number (m_l)

orientation of the orbital in space

for a given value of l

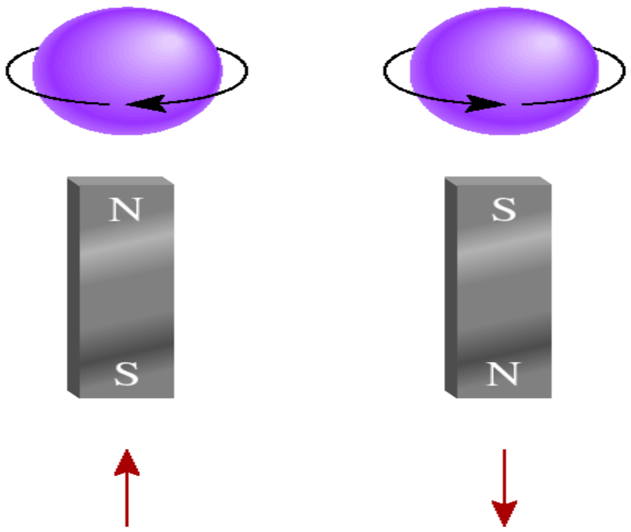
$m_l = -l, \ldots., 0, \ldots. +l$

if $l = 1$ (p orbital), $m_l = -1, 0,$ **or** 1

if $l = 2$ (d orbital), $m_l = -2, -1, 0, 1,$ **or** 2

spin quantum number (m_s)

$m_s = +½$ **or** $-½$



$m_s = +½$ $m_s = -½$

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Pauli exclusion principle - no two electrons in an atom can have the same four quantum numbers.

TABLE 7.2 Quantum Numbers for the First Four Levels of Orbitals in the Hydrogen Atom

n	ℓ	Orbital Designation	m_ℓ	Number of Orbitals
1	0	1s	0	1
2	0	2s	0	1
	1	2p	−1, 0, +1	3
3	0	3s	0	1
	1	3p	−1, 0, 1	3
	2	3d	−2, −1, 0, 1, 2	5
4	0	4s	0	1
	1	4p	−1, 0, 1	3
	2	4d	−2, −1, 0, 1, 2	5
	3	4f	−3, −2, −1, 0, 1, 2, 3	7

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Electron configuration

Electron configuration is how the electrons are distributed among the various atomic orbitals in an atom.

$1s^1$

- 1: principal quantum number *n*
- s: angular momentum quantum number *l*
- ^{1}: number of electrons in the orbital or subshell

Orbital diagram

H $\boxed{\uparrow}$

$1s^1$

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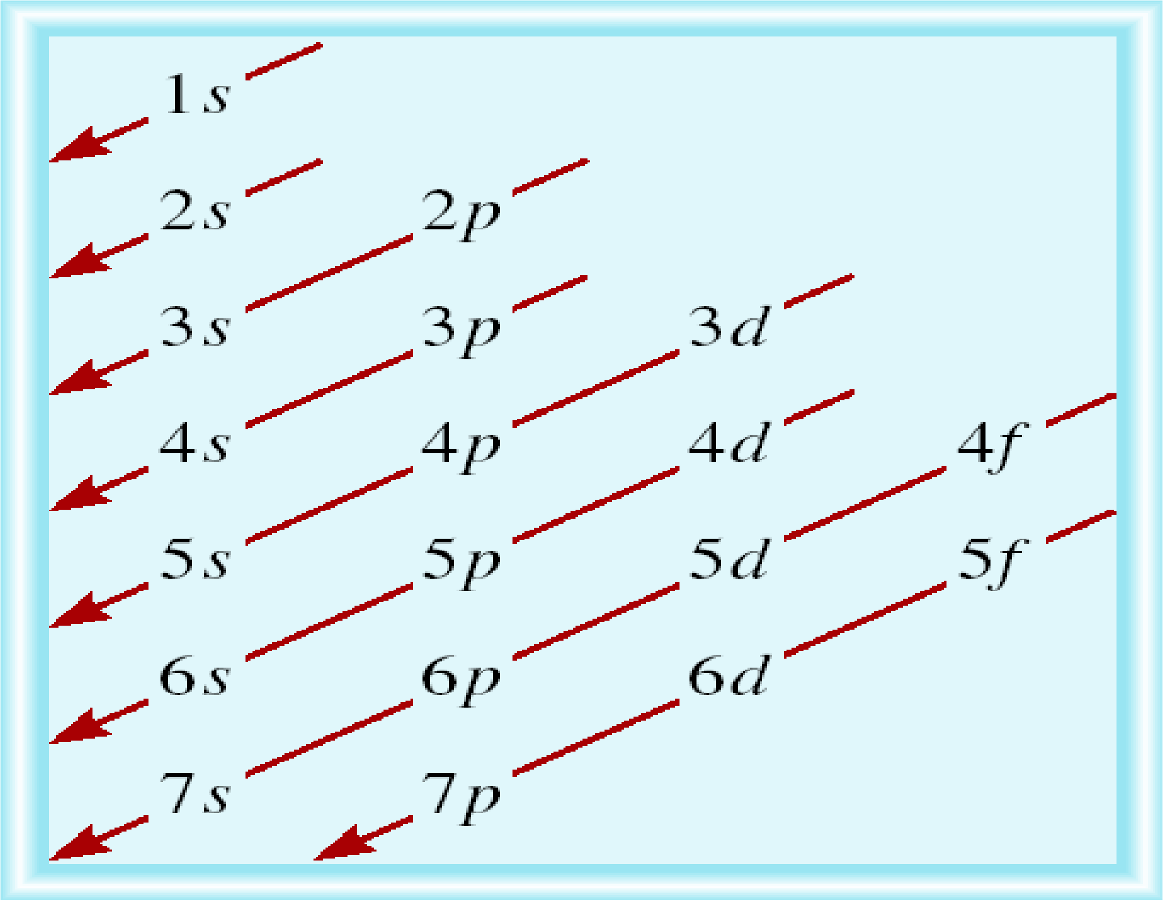
Order of orbitals (filling) in multi-electron atom



1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

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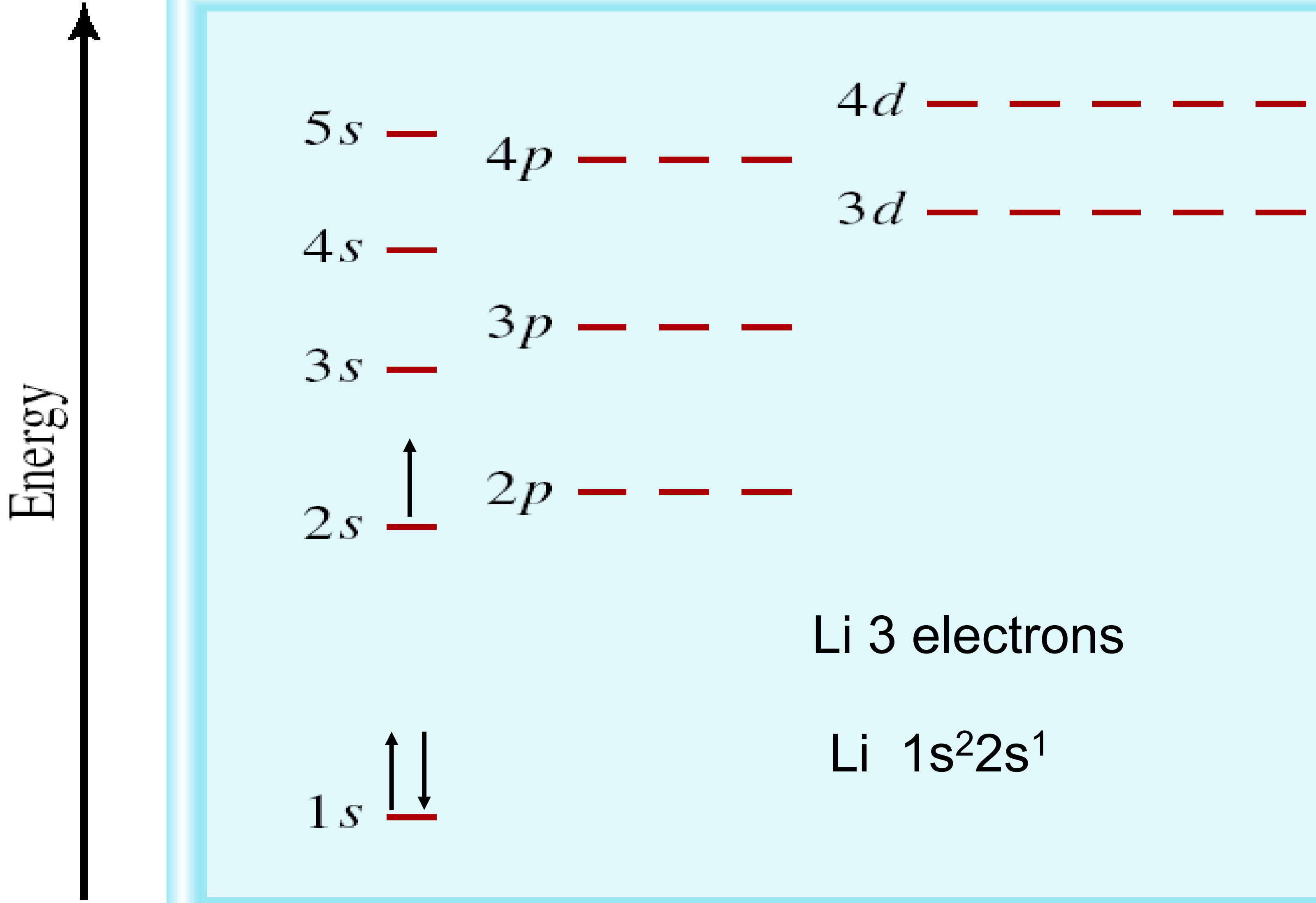
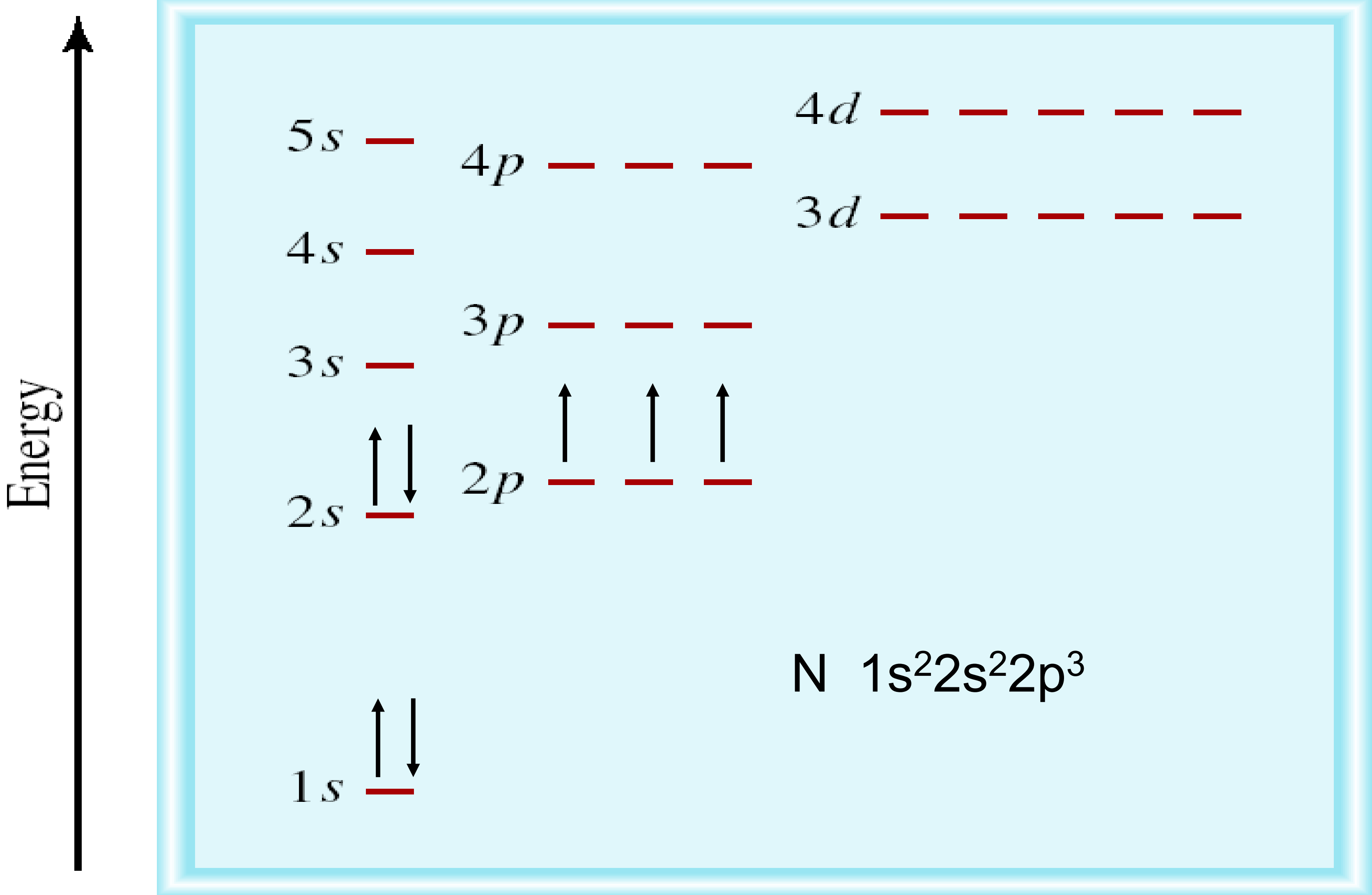
"Fill up" electrons in lowest energy orbitals first (*Aufbau principle*)



Li 3 electrons

Li $1s^22s^1$

The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins (*Hund's rule*).



What is the electron configuration of Mg?

Mg 12 electrons

$1s < 2s < 2p < 3s < 3p < 4s$

$1s^2 2s^2 2p^6 3s^2$

$2 + 2 + 6 + 2 = 12$ electrons

[Ne] $1s^2 2s^2 2p^6$

Abbreviated as [Ne]$3s^2$

What are the possible quantum numbers for the last (outermost) electron in Cl?

Cl 17 electrons

$1s < 2s < 2p < 3s < 3p < 4s$

$1s^2 2s^2 2p^6 3s^2 3p^5$

$2 + 2 + 6 + 2 + 5 = 17$ electrons

Last electron added to 3p orbital

$n = 3$ $\quad l = 1$ $\quad m_l = -1, 0, \text{or } +1$ $\quad m_s = ½ \text{ or } -½$

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Questions

Chose the correct answer

1. Protons are located in the nucleus of the atom. A proton has
a) No charge
b) A negative charge
c) A positive and a negative charge
d) A positive charge

2. The atomic number of an atom is
a) The mass of the atom
b) The number of protons added to the number of neutrons
c) The number of protons
d) Negatively charged

3. Neutrons are in the nucleus of the atom. A neutron has
a) A positive charge
b) No charge
c) A negative charge
d) Twice as much positive charge as a proton

4. The atoms of the same element can have different isotopes. An isotope of an atom
a) Is an atom with a different number of protons
b) Is an atom with a different number of neutrons
c) Is an atom with a different number of electrons
d) Has a different atomic number

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5. Which one of the following sets of four quantum numbers that most likely describe the last electron of the **Zn atom** **(Zn atomic number is 30)** ?

a) $n = 3,\ l = 2,\ m_l = 2,\ ms = -\frac{1}{2}$
b) $n = 3,\ l = 1,\ m_l = 1,\ ms = +\frac{1}{2}$
c) $n = 3,\ l = 3,\ m_l = 2,\ ms = -\frac{1}{2}$
d) $n = 4,\ l = 2,\ m_l = 0,\ ms = +\frac{1}{2}$
e) $n = 4,\ l = 3,\ m_l = 3,\ ms = -\frac{1}{2}$

6. Which one of the following sets of quantum numbers can correctly represent a **3*p*** orbital?

a.	b.	c.	d.	e.
$n = 3$	$n = 1$	$n = 3$	$n = 3$	$n = 3$
$l = 1$	$l = 3$	$l = 2$	$l = 1$	$l = 0$
$m_l = 2$	$m_l = 3$	$m_l = 1$	$m_l = -1$	$m_l = 1$

A) d
B) b
C) c
D) a
E) e

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7. True or false?

1. Electrons are found in the nucleus of an atom. False

2. Neutrons and electrons are attracted to one another. False

3.The first energy level of atom is closest to the nucleus. True

8. Fill-in-the-blank

1.Different atoms of the same element can have a different number of ________neutrons

2.When an atom loses an electron, it forms a ___________ positive ion.

3.When an atom gains an electron, it forms a ___________ negative ion.

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Choose the correct answer:

9. The electronic configuration of Aluminum (Al atomic number = 13) is:

a) [Ne] $2s^2 2p^1$
b) [Ne] $2s^1 2p^2$
c) [Ne] $3s^2 3p^1$
d) [Ne] $3s^1 3p^2$

10. The electronic configuration of Sodium (Na atomic number = 11) is:

a) $1s^2 2s^2 2p^6 3s^1$
b) $1s^2 2s^2 2p^5 3s^2$
c) $1s^2 2s^2 2p^7 3s^0$
d) None of the previous

11- The most favorable electronic configuration of Fe^{3+} (Fe atomic number = 26) is:

a) $[Ar]4s^0 3d^5$
b) $[Ar]4s^1 3d^4$
c) $[Ar]4s^2 3d^3$
d) $[Ar]4s^2 3d^5$

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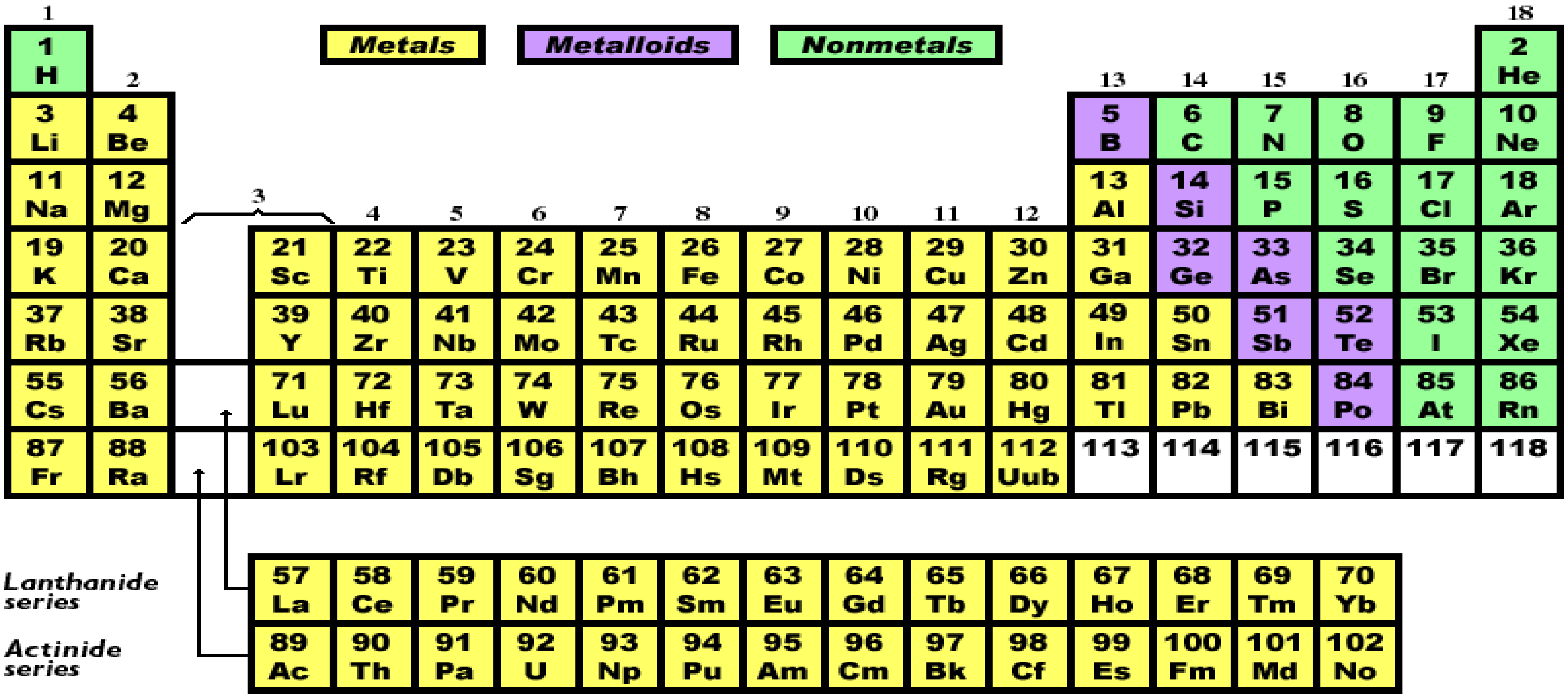
Periodic Table

Chapter 3

Chang-chapter8

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Metals, Nonmetals and Metaloids



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Metals, Nonmetals and Metaloids

metals
- Tend to lose one electron or more
- 1A, 2A, 3A (except H, B)

nonmetals
- Tend to accept one electron or more
- On the right of zigzag line

metalloids
- Between metals and nonmetals
- Around the zigzag line

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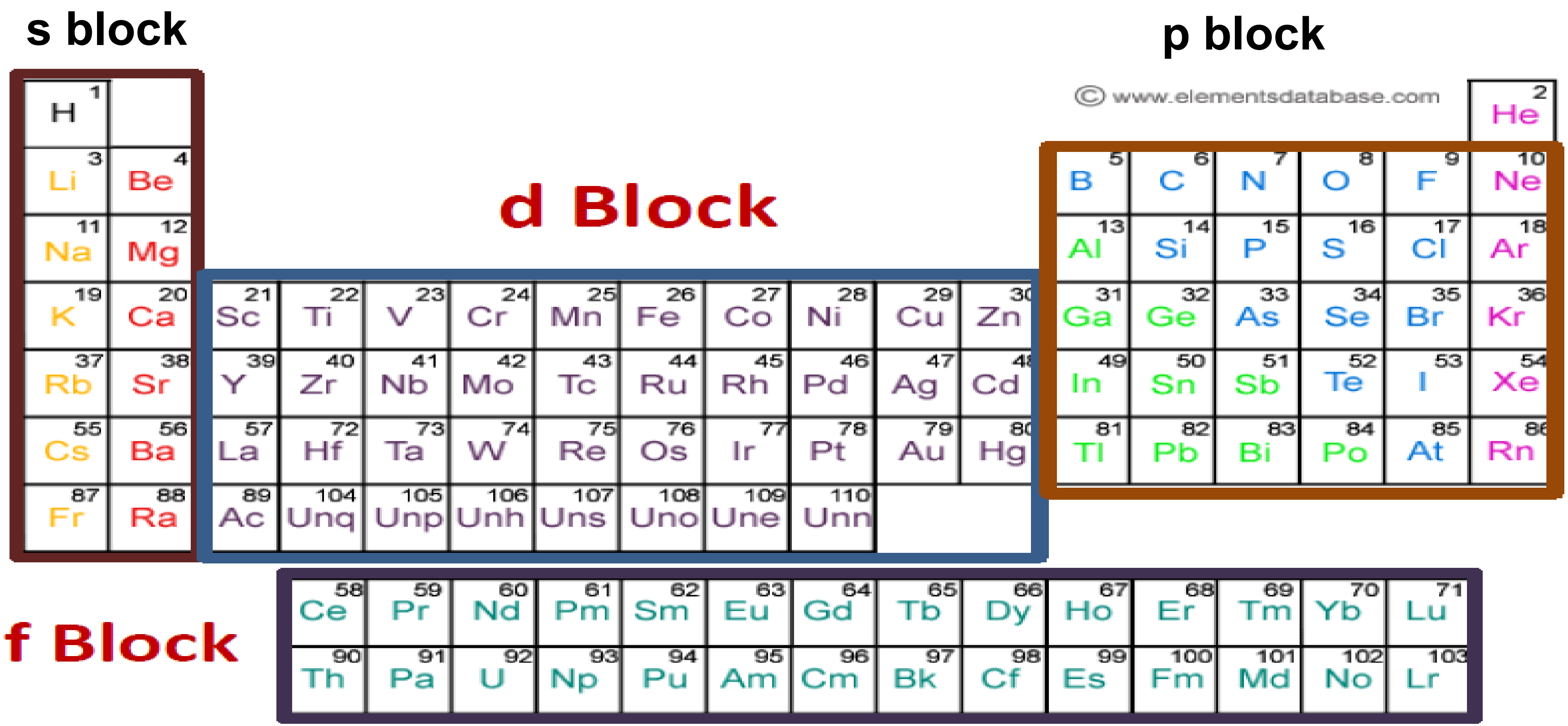
Blocks in Periodic Table



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Chemical Properties of Elements in Periodic Table

Atomic Radius

Ionization Energy

Electronic Affinity

Electronegativity

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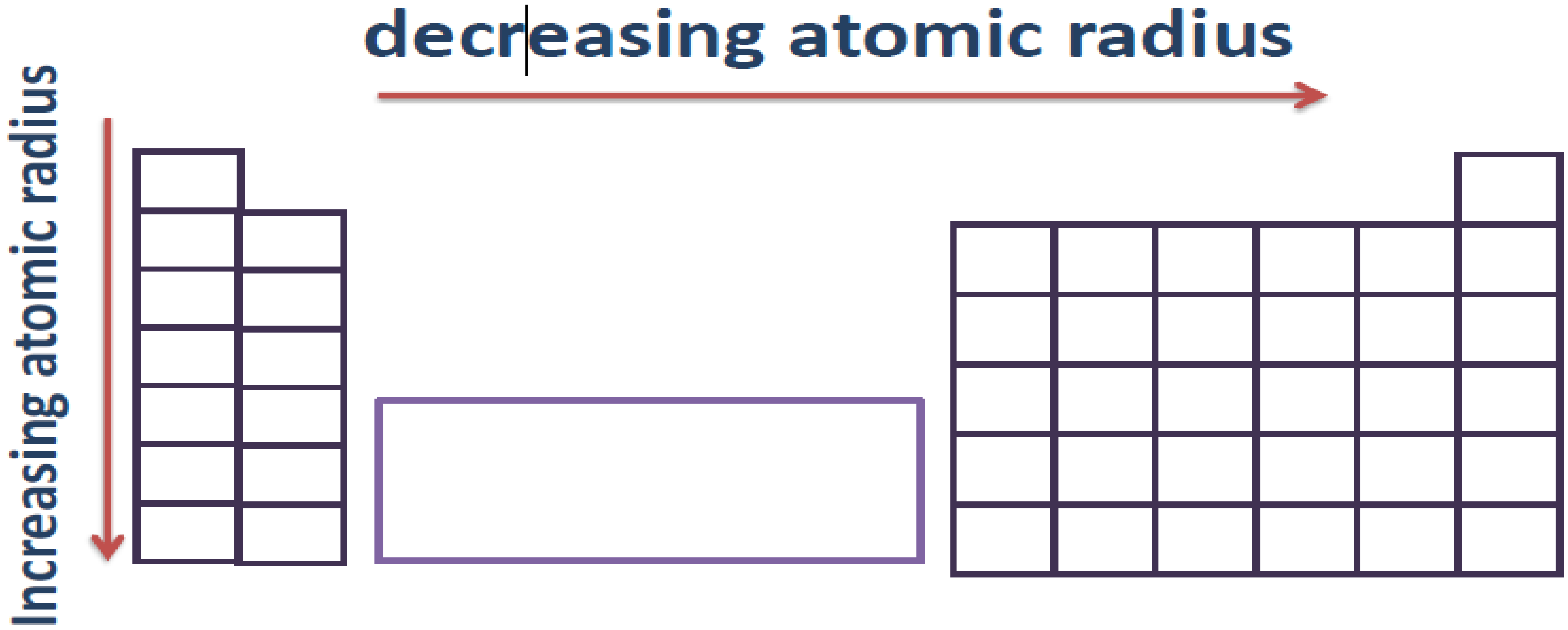
Atomic Radius



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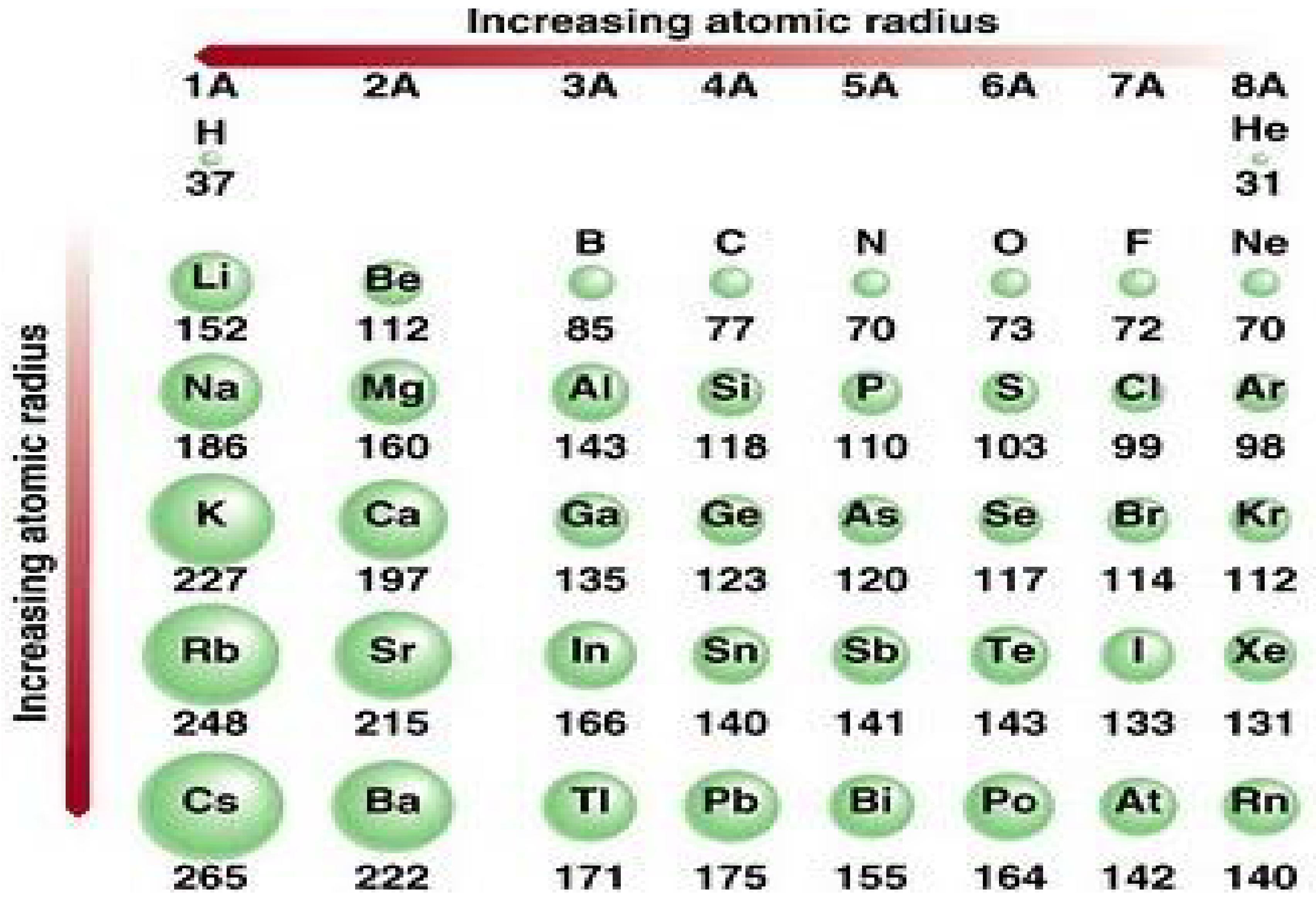
Atomic Radius



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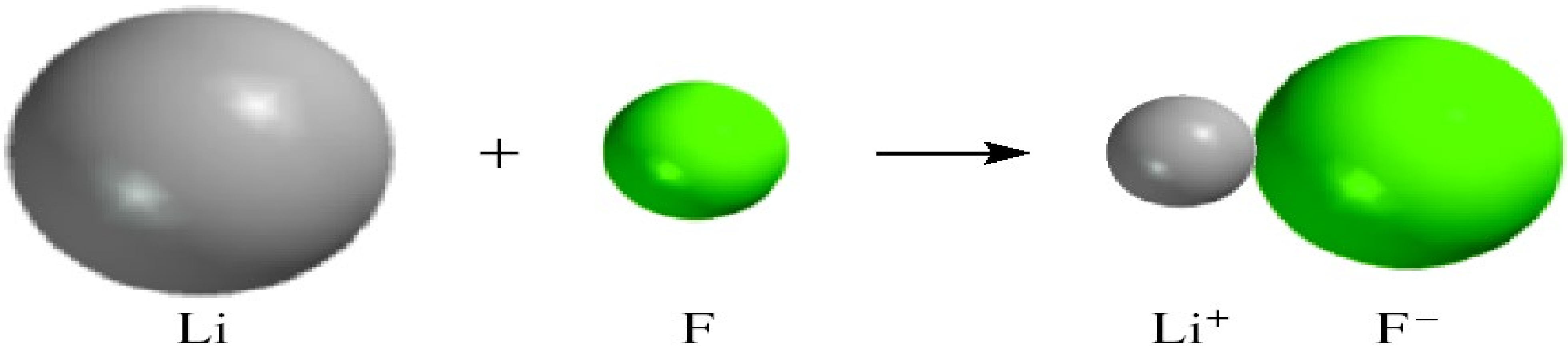
Atomic Radius



Cation is always **smaller** than atom from which it is formed.
Anion is always **larger** than atom from which it is formed.

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Ionization Energy

The minimum energy required to remove an electron from a gaseous atom in its ground state

M (g) + energy = M^+ + e^-

First ionization $$M_{(g)} + IE_1 \longrightarrow M^+{}_{(g)} + e^-$$

Second ionization $$M^+{}_{(g)} + IE_2 \longrightarrow M^{2+}{}_{(g)} + e^-$$

Third ionization $$M^{2+}{}_{(g)} + IE_3 \longrightarrow M^{3+}{}_{(g)} + e^-$$

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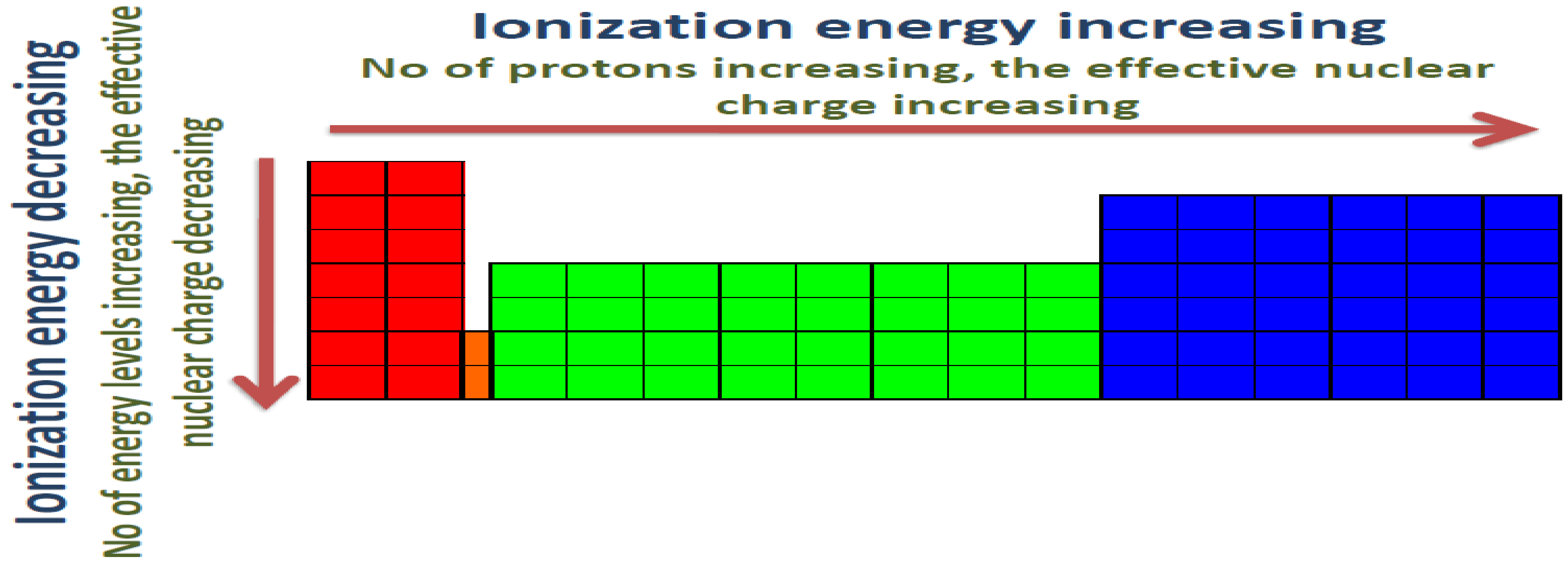
$$IE_1 < IE_2 < IE_3$$

Ionization Energy



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Electronic Affinity

Electronic affinity is the negative of the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)}$$

$F_{(g)} + e^- \longrightarrow X^-_{(g)}$ $\Delta H = -328$ kJ/mol EA = +328 kJ/mol

$O_{(g)} + e^- \longrightarrow O^-_{(g)}$ $\Delta H = -141$ kJ/mol EA = +141 kJ/mol

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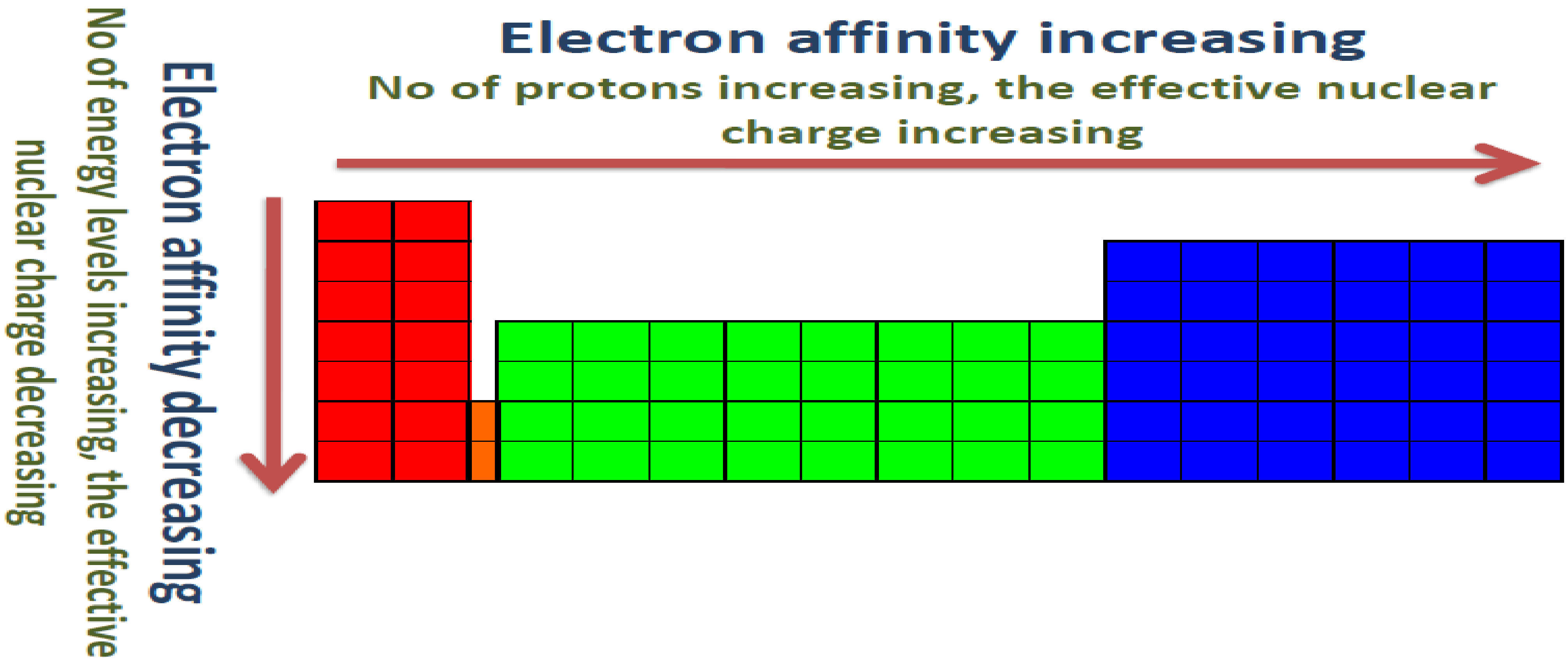
Electronic Affinity



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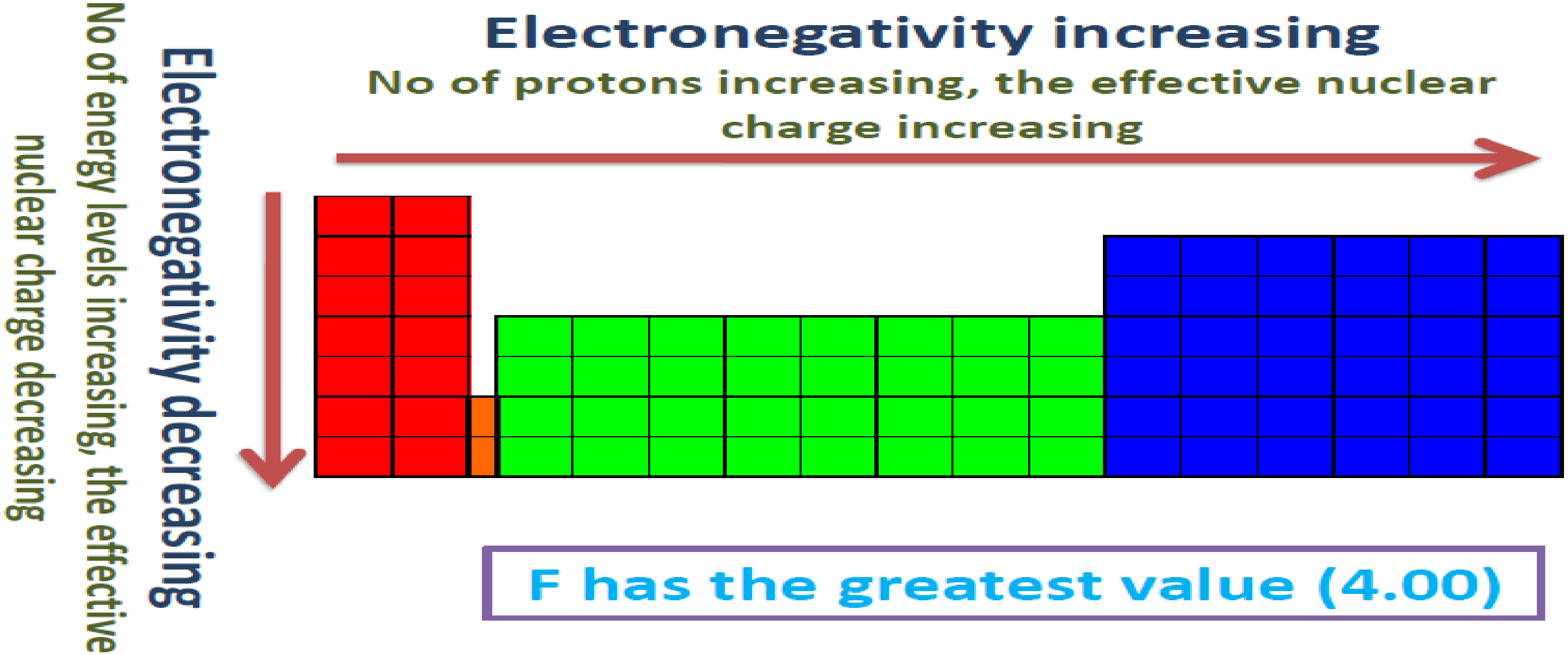
Electronegativity

The ability of an atom to attract toward itself the electrons in a chemical bond



F has the greatest value (4.00)

Questions

Choose the correct answer:

1- Tend to accept an electron or more:

a) Metals
b) Nonmetals
c) Metaloids
d) None of the previous

2- The minimum energy required to remove an electron from a gaseous atom in its ground state

a) Atomic radius
b) Ionization energy
c) Electronic affinity
d) Electronegativity

3- The ability of an atom to attract toward itself the electrons in a chemical bond:

a) Atomic radius
b) Ionization energy
c) Electronic affinity
d) Electronegativity

4- First ionization energy is ……… second ionization energy.

a) equals to
b) higher than
c) lower than
d) None of the previous

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Questions

Choose the correct answer:

5- The negative of the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion:

a) Atomic radius
b) Ionization energy
c) Electronic affinity
d) Electronegativity

6- Cation is always .... atom from which it is formed.

a) smaller than
b) larger than
c) equal
d) none of the previous

7- Atoms lose electrons so that .... has a noble-gas outer electron configuration.

a) electrons
b) cation
c) anions
d) atoms

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Atomic Weight
Molecular Weight
Moles Calculations

Chapter
4

Chang-chapter3

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Atomic Mass

The mass of an atom in atomic mass units (amu)

6 ← Atomic number

C

12.01 ← Atomic mass

The atomic mass of elements is relative to a standard atom ^{12}C (6 protons, 6 neutrons)

Molar Mass (Atomic weight Aw)

The mass of an element atoms per one mole (g/mol)

= Atomic Mass numerically

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4th lecture

Mole (mol)

The amount of a substance that contains as many elementary particles (atoms, molecules or ions), where each mole has number of 6.022×10^{23} particles.

1 mole= 6.022×10^{23} particles = Avogadro's number N_a

1 mol Al = 6.02×10^{23} atoms

1 mol CO_2 = 6.02×10^{23} molecules

1 mol NaCl = 6.02×10^{23} Na^+ ions = 6.02×10^{23} Cl^- ions

The number of atoms in exactly 12 g of ^{12}C is one mole

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Molar Mass (Atomic weight A_w):

mass (weight) of 1 mole of atoms in grams

1 mol C atoms = 12.01 g A_w of C = 12.01* g/mol

1 mol Cl atoms = 35.45 g A_w of Cl = 35.45* g/mol

1 mol Fe atoms = 55.85 g A_w of Fe = 55.85* g/mol

*(get from periodic table)

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Think: What is the difference between the mass and weight?

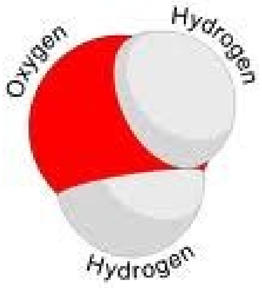
Molar Mass (Molecular weight M_w):

The sum of atomic weights of 1 mol of the molecule

M_w of 1 mol of H_2O = 2 (A_w of H) + A_w of O

= (2× 1.008) + 16

= 18.02 g/mol



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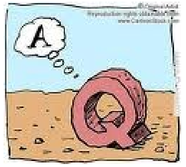

What are the molecular weights of the following:

C_2H_6

N_2O_4

$C_8H_{18}O_4N_2S$

$Al_2(CO_3)_3$

$MgSO_4.7H_2O$

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Number of moles (n)

$$n = \frac{wt(g)}{Mw(g/mol)}$$

Remember: No. of particles = No. of moles × Avogadro's number

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Example

Methane (CH_4) is the principal component of the natural gas. How many moles of methane are present in 6.07 g of CH_4?

M_w of CH_4 = 12.01 + (4× 1.008) = 16.04 g/mol

M_w = 16.04 g/mol

$$n \text{ of } CH_4 = 6.07\ \cancel{g}_{\,(CH4)} \times \left(\frac{1mol_{(CH4)}}{16.04\ \cancel{g}_{\,(CH4)}}\right) = 0.378\ mol_{(CH4)}$$

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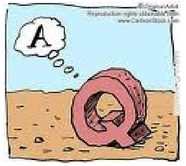
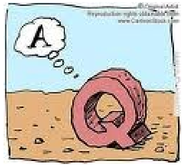
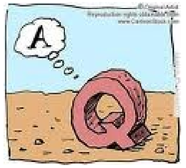
Learning check



What is the number of moles in 21.5 g $CaCO_3$?



What is the mass in grams of 0.6 mol C_4H_{10}?



How many atoms of Cu are present in 35.4 g of Cu?

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4th lecture

Percent Composition of Compounds

Mass percent (weight percent) of each element in a compound.

$$\%x = \frac{n \times A_w(x)}{Mw} \times 100$$

n is number of atoms of each element in the compound

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Example

Calculate the mass percent of each element in ethanol (C_2H_5OH) ?

$$\%x = \frac{n \times A_w(x)}{Mw} \times 100$$

Mass of 1 mol (molar mass) of C_2H_5OH = 24.02+6.048+16.00= 46.07 g/mol

$$\text{Mass percent of C} = \frac{2 \times 12.01 \text{ g/mol}}{46.07 \text{ g/mol}} \times 100 = \underline{52.14}\ \% \quad \text{(4 sf)}$$

$$\text{Mass percent of H} = \frac{6 \times 1.008 \text{ g/mol}}{46.07 \text{ g/mol}} \times 100 = \underline{13.13}\ \% \quad \text{(4 sf)}$$

$$\text{Mass percent of O} = \frac{1 \times 16.00 \text{ g/mol}}{46.07 \text{ g/mol}} \times 100 = \underline{34.73}\ \% \quad \text{(4 sf)}$$

Total mass = 52.14 +13.13 + 34.73 =100%

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Percent composition

Determining the Formula of a Compound:

empirical formula → *molecular formula*

CH_2O $C_6H_{12}O_6$

Molecular formula = (Empirical formula)$_x$

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A ***molecular formula*** shows the exact number of atoms of each element in the smallest unit of a substance

An ***empirical formula*** shows the simplest whole-number ratio of the atoms in a substance

molecular	**empirical**
H_2O	H_2O
$C_6H_{12}O_6$	CH_2O
O_3	O
N_2H_4	NH_2

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Question 1

Determine the number of moles of aluminum in 0.2154 kg of Al.

A) 1.297 x 1023 mol
B) 5.811 x 103 mol
C) 7.984 mol
D) 0.1253 mol
E) 7.984 x 10-3 mol

Question 2

How many phosphorus atoms are there in 2.57 g of P?

A) 4.79 x 1025
B) 1.55 x 1024
C) 5.00 x 1022
D) 8.30 x 10-2
E) 2.57

Question 3

One mole of H2

A) contains 6.0 x 10^{23} H atoms
B) contains 6.0 x 10^{23} H_2 molecules
C) contains 1 g of H_2
D) is equivalent to 6.02 x 10^{23} g of H_2
E) None of the above

Question 4

How many oxygen atoms are present in 5.2 g of O2?

A) 5.4 x 10-25 atoms
B) 9.8 x 1022 atoms
C) 2.0 x 1023 atoms
D) 3.1 x 1024 atoms
E) 6.3 x 1024 atoms

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Question 5

How many protons and neutrons are in sulfur-33?

A) 2 protons, 16 neutrons

B) 16 protons, 31 neutrons

C) 16 protons, 17 neutrons

D) 15 protons, 16 neutrons

Question 6

What is the mass of 5.45 x 10-3 mol of glucose, C6 H12O6?

A) 0.158 g
B) 982 g
C) 3.31 x 104 g
D) 0.982 g
E) None of the above.

Question 7

Determine the mass percent of iron in Fe4[Fe(CN)6] 3.

A) 45% Fe
B) 26% Fe
C) 33% Fe
D) 58% Fe
E) None of the above.

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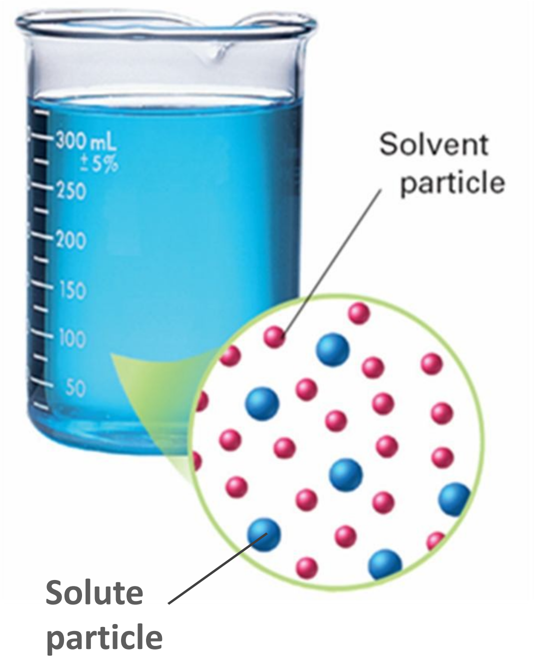
Chemical Reactions in Solutions & Concentrations

Chapter 5

Chang-chapter3,4,12

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Solutions



Solution: a homogeneous mixture of two or more substances

Solute: a substance that is being dissolved (smaller amount)

Solvent: a substance which dissolves a solute (larger amount)

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Concentrations

The *concentration* of a solution is the amount of solute present in a given quantity of a solvent or solution.

- **Concentrations**
  - **Molar concentrations**
    - Molarity
    - Molality
    - Normality
    - Mole fraction
  - **Percentages**
    - v/v
    - w/w
    - w/v

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Molarity

The number of moles of solute dissolved in one liter of solution.

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

What is the unit of molarity?
What is the relationship between weight and molarity?

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Example

A solution has a volume of 2.0 L and contains 36.0 g of glucose ($C_6H_{12}O_6$). If the molar mass of glucose is 180 g/mol, what is the molarity of the solution?

No. of mol of glucose = wt (g) / Mw (g/mol) = 36.0 g/ 180 g/mol = 0.2 mol

M = n (mol) / V (L) = 0.2 mol /2.0 L = 0.1 mol/L

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Molality

The number of moles of solute dissolved in one kilogram of solvent

Molality *(m)*

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}}$$

Molarity *(M)*

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

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Example

What is the molality of a 5.86 *M* ethanol (C_2H_5OH) solution whose density is 0.927 g/mL?

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}}$$

Assume 1 L of solution:

5.86 moles ethanol = 270 g ethanol

927 g of solution (1000 mL x 0.927 g/mL)

mass of solvent = mass of solution – mass of solute

= 927 g – 270 g = 657 g = 0.657 kg

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{5.86 \text{ moles } C_2H_5OH}{0.657 \text{ kg solvent}} = 8.92\ m$$

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Learning check

What is the concentration of a solution in mol/L when 80 g of calcium carbonate, $CaCO_3$, is dissolved in 2 L of solution?

How many liters of 0.25 M NaCl solution must be measured to obtain 0.1 mol of NaCl?

A student needs to prepare 250 ml of 0.1 M of $Cd(NO_3)_2$ solution. How many grams of cadmium nitrate are required?

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Chemical Reactions

Reactants $\longrightarrow$ **Products**

A process in which one or more substances is changed into one or more new substances.

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

$$2HgO\,(s) \longrightarrow 2Hg\,(l) + O_2\,(g)$$

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Chemical Equations

It is a way to represent the chemical reaction.

It shows us:

- The chemical symbols of reactants and products
- The physical states of reactants and products– (s), (l), (g), (aq)
- Balanced equation (same number of atoms on each side)

$$\boxed{2H_2\,(g) + O_2\,(g)} \longrightarrow \boxed{2H_2O\,(l)}$$

Reactants (starting materials) Products (materials formed)

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Balancing Chemical Equations

The number of atoms of each element must be the same on both sides of the equation.

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

Reactants	Products
2 C	1 C
6 H	2 H
2 O	3 O

$$C_2H_6 + 7/2\,O_2 \longrightarrow 2CO_2 + 3H_2O$$

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

Reactants	Products
4 C	4 C
12 H	12 H
14 O	14 O

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Balance the following equations:

(a) $C + O_2 \longrightarrow CO$

(b) $CO + O_2 \longrightarrow CO_2$

(c) $H_2 + Br_2 \longrightarrow HBr$

(d) $K + H_2O \longrightarrow KOH + H_2$

(e) $Mg + O_2 \longrightarrow MgO$

(f) $O_3 \longrightarrow O_2$

(g) $H_2O_2 \longrightarrow H_2O + O_2$

(h) $N_2 + H_2 \longrightarrow NH_3$

(i) $Zn + AgCl \longrightarrow ZnCl_2 + Ag$

(j) $S_8 + O_2 \longrightarrow SO_2$

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Stoichiometry

The quantitative study of reactants and products in a chemical reaction

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

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Type of Chemical Reactions in Aqueous Solutions

1) Acid-Base Reactions

2) Oxidation-Reduction Reactions

3) Precipitation Reactions

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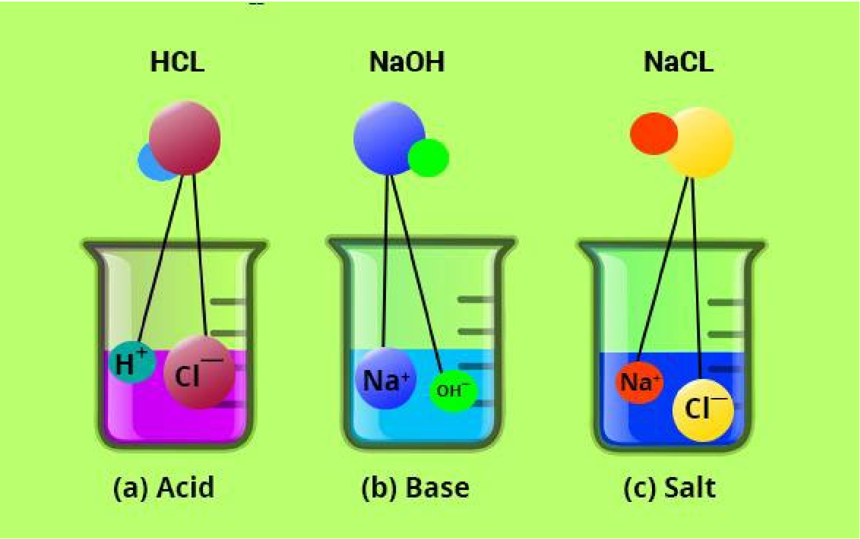
I. Acid-Base Reactions

acid + base → salt + water

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



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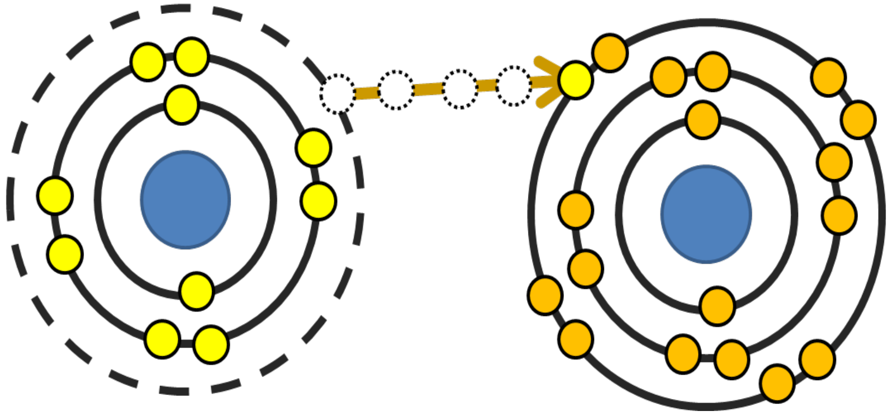
II. Oxidation-Reduction Reactions

Redox reactions are electron transfer reactions

$Mg + 2Ag^{+} \rightarrow Mg^{2+} + 2Ag$



Half-reactions:

$Mg\ (s) \rightarrow Mg^{2+} + 2e$

$2Ag^{+} + 2e \rightarrow 2Ag$

Oxidation Reactions : half-reaction that involves a loss of electrons

Reduction Reactions : half-reaction that involves a gain of electrons

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III. Precipitation Reactions

A precipitate is an insoluble solid that separates from the solutions

$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$

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

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

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Question 1

Molarity is the number of …… of solute

dissolved

Solution

a) Grams

b) Milliliter

c) Second

d) moles

Question 2

Molality is the number of moles of ……. dissolved in 1kg solvent

a) Solvent
b) Solute
c) Solution
d) acid

Question 3

Molarity is the number of moles of solute

dissolved **1 …… of the Solution**

a) Grams

b) Liter

c) Second

d) moles

Question 4

A solution has a volume of 2.0 L and contains 36.0 g of glucose (C6H12O6). If the molar mass of glucose is 180 g/mol, what is the molarity of the solution

a) 1.0
b) 1.00
c) 0.1
d) 0.01

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Question 5

How many liters of 0.25 M NaCl solution must be measured to obtain 0.1 mol of NaCl

A) 1
B) 2
C) 0.4
D) 3.5

Question 6

What is the concentration of a solution in mol/L when 80 g of calcium carbonate, $Ca(CO_{3)2}$, is dissolved in 2 L of solution? (Molecular weight of Ca(CO3)2 = 100g/mol

A) 0.4
B) 4
C) 0.004
D) 1

Question 7

A student needs to prepare 250 ml of 0.1 M of $Cd(NO_3)_2$ solution. How many grams of cadmium nitrate are required? (Molecular weight of $Cd(NO_3)_2$ = 236 g/mol

A) 5.9
B) 5.1
C) 5.4
D) 5.6

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Chemical Equilibrium

Chapter 6

Chang-chapter14

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Equilibrium

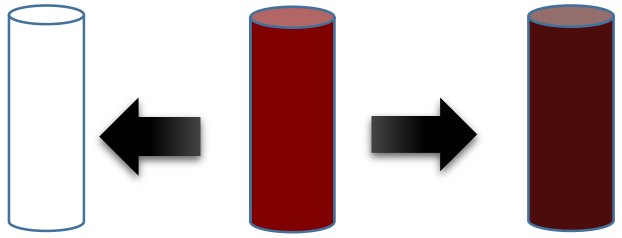
Equilibrium is a state in which there are no observable changes as time goes by

Chemical equilibrium is achieved when:

- the rates of the forward and reverse reactions are equal and
- the concentrations of the reactants and products remain constant

Chemical equilibrium

$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$



N_2O_4 Colorless | **Equilibrium state** | **NO_2 Dark-brown**

Physical equilibrium

$$H_2O\,(l) \rightleftharpoons H_2O\,(g)$$

Physical equilibrium is between two states of the same substance

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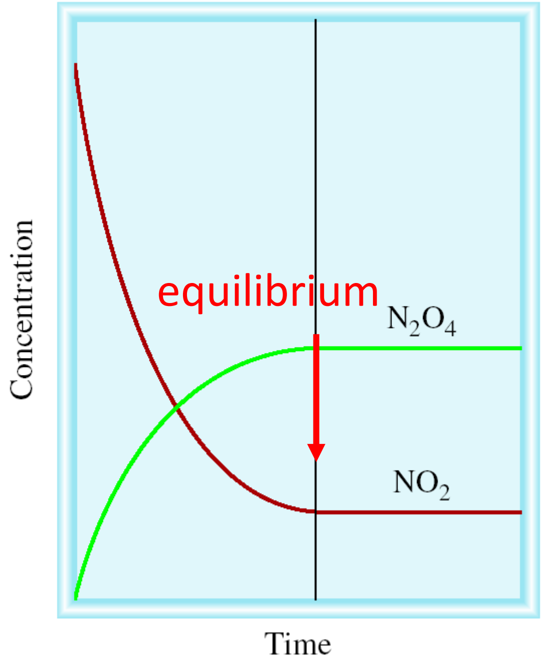
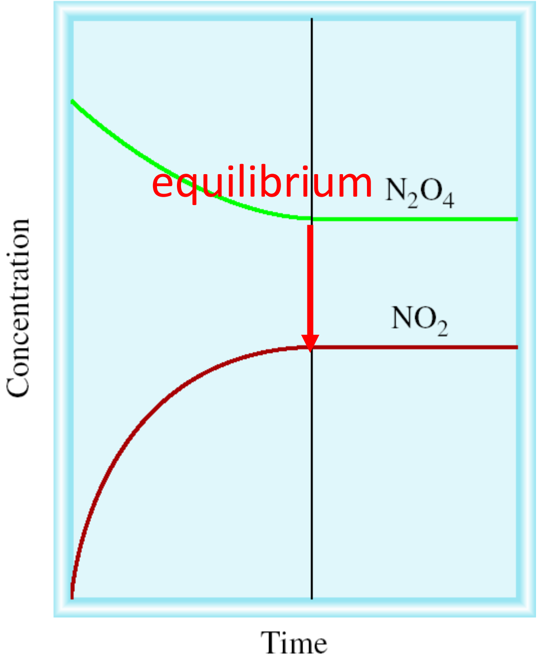
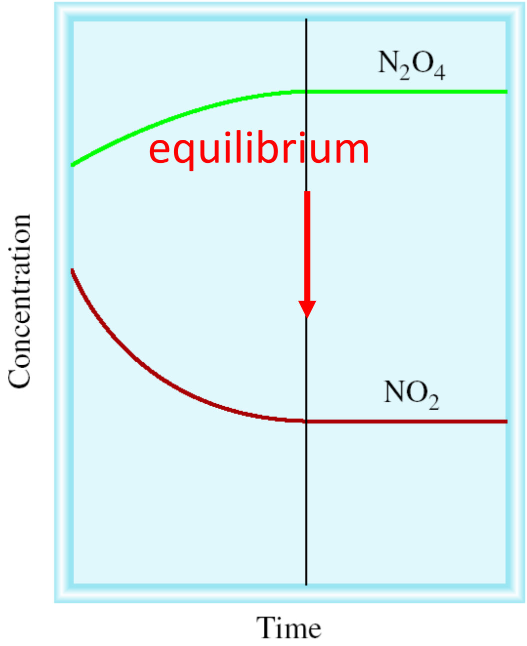
$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$



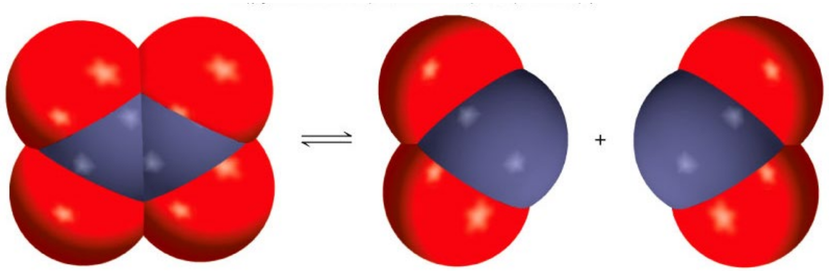
Start with NO_2



Start with N_2O_4



Start with NO_2 & N_2O_4



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Equilibrium Constant K

$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} \qquad K_p = \frac{P_{NO_2}^{\,2}}{P_{N_2O_4}}$$

$$K = \frac{[NO_2]^2}{[N_2O_4]} = 4.63 \times 10^{-3}$$

$$aA + bB \rightleftharpoons cC + dD$$

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Law of Mass Action

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Equilibrium Position

K>>1

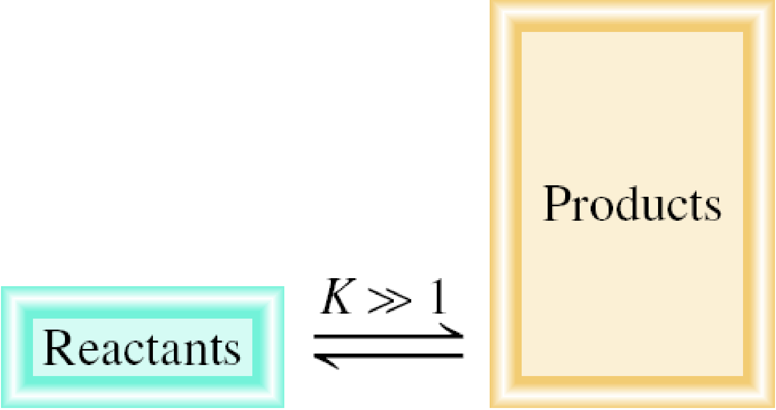
Products are favored at equilibrium

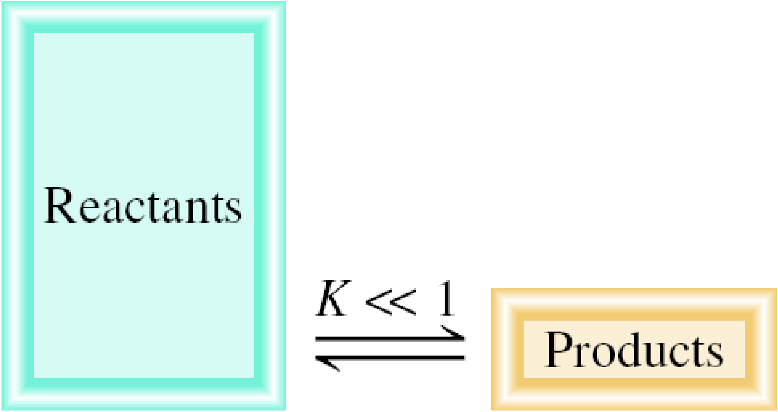
(the equilibrium lie to the right)

K<<1

Reactants are favored at equilibrium

(the equilibrium lie to the left)





Relation between K_c and K_p

$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} \qquad K_p = \frac{P^2_{NO_2}}{P_{N_2O_4}}$$

In most cases

$$K_c \neq K_p$$

$$aA\,(g) + bB\,(g) \rightleftharpoons cC\,(g) + dD\,(g)$$

$$K_p = K_c(RT)^{\Delta n}$$

Δn = moles of gaseous products – moles of gaseous reactants

$= (c + d) - (a + b)$

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Homogeneous Equilibrium

Homogenous equilibrium applies to reactions in which all reacting species **are in the same phase.**

$$CH_3COOH\,(aq) + H_2O\,(l) \rightleftharpoons CH_3COO^-\,(aq) + H_3O^+\,(aq)$$

$$K_c' = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH][H_2O]} \qquad [H_2O] = \text{constant}$$

$$K_c = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = K_c'[H_2O]$$

General practice **not** to include units for the equilibrium constant.

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The equilibrium concentrations for the reaction between carbon monoxide and molecular chlorine to form $COCl_2$ (*g*) at 74^{0}C are [CO] = 0.012 *M*, [Cl_2] = 0.054 *M*, and [$COCl_2$] = 0.14 *M*. Calculate the equilibrium constants K_c and K_p.

$$CO\,(g) + Cl_2\,(g) \rightleftharpoons COCl_2\,(g)$$

$$K_c = \frac{[COCl_2]}{[CO][Cl_2]} = \frac{0.14}{0.012 \times 0.054} = 220$$

$$K_p = K_c(RT)^{\Delta n}$$

$\Delta n = 1 - 2 = -1$ $\quad$ $R = 0.0821$ $\quad$ $T = 273 + 74 = 347$ K

$$K_p = 220 \times (0.0821 \times 347)^{-1} = 7.7$$

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The equilibrium constant K_p for the reaction: $2NO_2\,(g) \rightleftharpoons 2NO\,(g) + O_2\,(g)$

is 158 at 1000K. What is the equilibrium pressure of O_2 if the P_{NO} = 0.400 atm and P_{NO} = 0.270 atm?

$$K_p = \frac{P_{NO}^2\,P_{O_2}}{P_{NO_2}^2}$$

$$P_{O_2} = K_p\,\frac{P_{NO_2}^2}{P_{NO}^2}$$

$$P_{O_2} = 158 \times (0.400)^2/(0.270)^2 = 347\ \text{atm}$$

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Heterogeneous Equilibrium

Heterogenous equilibrium applies to reactions in which reactants and products **are in different phases**

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

$$K'_c = \frac{[CaO][CO_2]}{[CaCO_3]}$$

$[CaCO_3]$ = constant
$[CaO]$ = constant

$$\mathbf{K_c = [CO_2]} = K'_c \times \frac{[CaCO_3]}{[CaO]}$$

$$K_p = P_{CO_2}$$

The concentration of **solids** and **pure liquids** are not included in the expression for the equilibrium constant.

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Consider the following equilibrium at 295 K:

$$NH_4HS\ (s) \rightleftharpoons NH_3\ (g) + H_2S\ (g)$$

The partial pressure of each gas is 0.265 atm. Calculate K_p and K_c for the reaction?

$$K_p = P_{NH_3}P_{H_2S} = 0.265 \times 0.265 = 0.0702$$

$$K_p = K_c(RT)^{\Delta n}$$

$$K_c = K_p(RT)^{-\Delta n}$$

$$\Delta n = 2 - 0 = 2 \qquad T = 295\ K$$

$$K_c = 0.0702 \times (0.0821 \times 295)^{-2} = 1.20 \times 10^{-4}$$

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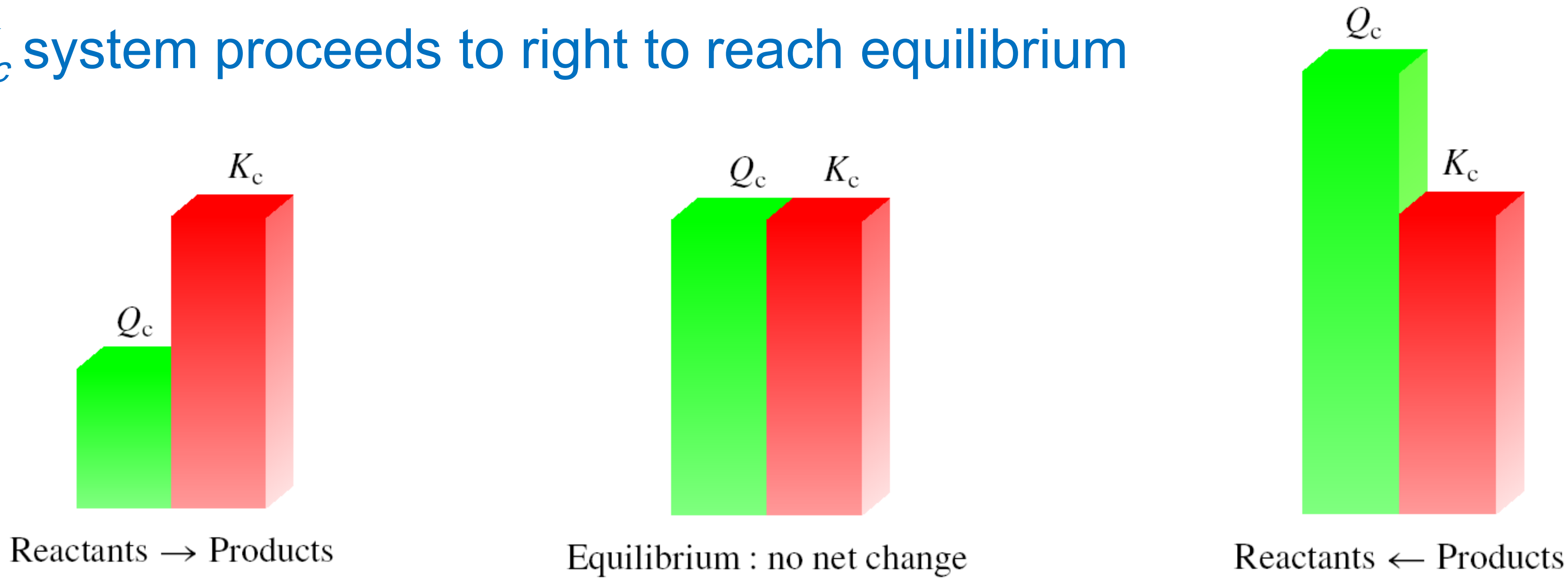
Reaction Quotient Q_c

The ***reaction quotient*** ***(Q_c)*** is calculated by substituting the initial concentrations of the reactants and products into the equilibrium constant (K_c) expression.

- $Q_c > K_c$ system proceeds to left to reach equilibrium
- $Q_c = K_c$ the system is at equilibrium
- $Q_c < K_c$ system proceeds to right to reach equilibrium



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- Find the value of Q and determine which side of the reaction is favored. Given K_{eq}=0.5

  A (aq) + B (aq) ⇌ D (aq)

  [A]= 0.1 M [B]= 0.2 M [C]=0.1 M

$$\mathbf{Qc = \frac{[C]}{[A][B]} = \frac{0.1}{(0.1)\,(0.2)} = 5}$$

Qc = 5 … *Q* is larger than K_{eq} so the reaction shifts left, favors the reactants.

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Equilibrium Constant Calculations

- **If a reaction can be expressed as the sum of two or more reactions, the equilibrium constant for the overall reaction is given by the <u>product of the equilibrium constants</u> of the individual reactions.**

$$\cancel{A} + B \rightleftharpoons \cancel{C} + \cancel{D} \qquad K'_c$$

$$\cancel{C} + \cancel{D} \rightleftharpoons E + F \qquad K''_c$$

$$A + B \rightleftharpoons E + F \qquad K_c$$

$$K'_c = \frac{[C][D]}{[A][B]} \qquad K''_c = \frac{[E][F]}{[C][D]}$$

$$K_c = \frac{[E][F]}{[A][B]}$$

$$K_c = K'_c \times K''_c$$

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Equilibrium Constant Calculations

- **When the equation for a reversible reaction is written in the opposite direction, the equilibrium constant becomes <u>the reciprocal of the original equilibrium constant</u>.**

$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$

$$K = \frac{[NO_2]^2}{[N_2O_4]} = 4.63 \times 10^{-3}$$

$$2NO_2\,(g) \rightleftharpoons N_2O_4\,(g)$$

$$K' = \frac{[N_2O_4]}{[NO_2]^2} = \frac{1}{K} = 216$$

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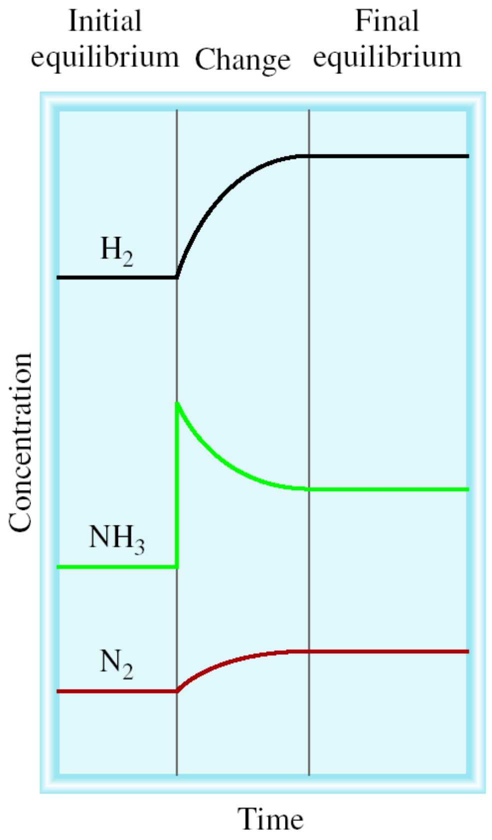
Le Châtelier's Principle

If an external stress is applied to a system at equilibrium, the system adjusts in such a way that the stress is partially offset as the system reaches a new equilibrium position.

I. Changes in Concentration

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

Equilibrium shifts left to offset stress ← Add NH_3



Changes in Concentration continued

Add   Add

$$aA + bB \rightleftharpoons cC + dD$$

Remove   Remove

Change	Shifts the Equilibrium
Increase concentration of product(s)	left
Decrease concentration of product(s)	right
Increase concentration of reactant(s)	right
Decrease concentration of reactant(s)	left

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Le Châtelier's Principle

II. Changes in Volume and Pressure

$$A\,(g) + B\,(g) \rightleftharpoons C\,(g)$$

Change	**Shifts the Equilibrium**
Increase pressure	Side with fewest moles of gas
Decrease pressure	Side with most moles of gas
Increase volume	Side with most moles of gas
Decrease volume	Side with fewest moles of gas

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Le Châtelier's Principle

III. Temperature Changes

- **Consider heat as a product in exothermic reactions**

A + B = AB + Heat

- **Add heat → Shift to reactants**
- **Remove heat → Shift to products**

Consider heat as a reactant in endothermic reactions

A + B + heat = AB

Add heat → Shift to products

Remove heat → Shift to reactants

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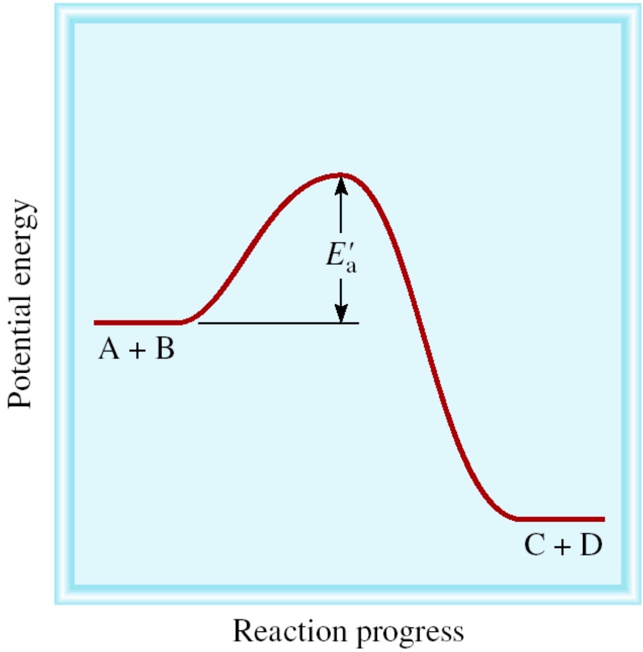
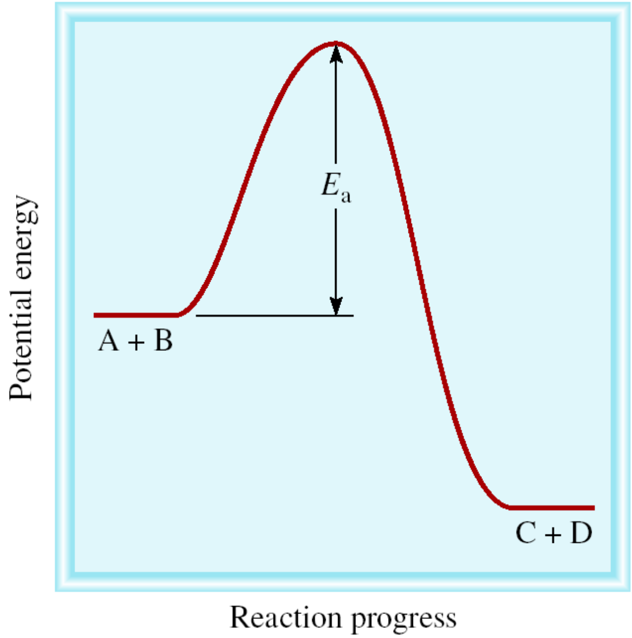
Le Châtelier's Principle

- **Adding a Catalyst**
  - does not change *K*
  - does not shift the position of an equilibrium system
  - system will reach equilibrium sooner





Catalyst lowers E_a for **both** forward and reverse reactions.

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Le Châtelier's Principle - Summary

Change	Shift Equilibrium	Change Equilibrium Constant
Concentration	yes	no
Pressure	yes*	no
Volume	yes*	no
Temperature	yes	yes
Catalyst	no	no

*Dependent on relative moles of gaseous reactants and products

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Question 1

Which equilibrium in gaseous phase would be unaffected by an increase in pressure:

(a) N_2O_4 ->$2NO_2$
(b) $N_2 + O_2$ ->$2NO$
(c) N2 + $3H_2$ ->$2NH_3$
(d) CO + ½ O_2 ->O_2 + CO_2

Question 2

For the equilibrium ,
$2NO_2$(g) -> N_2O_4(g) + 14.6 kcal
An increase of temperature will:

(a) Favour the formation of N_2O_4
(b) Favour the decomposition of N_2O_4
(c) Not affect the equilibrium
(d) Stop the reaction

Question 3

The equilibrium constant (K_c) for the reaction is
$2SO_3$(g) -> $2SO_2$(g) + O_2(g)
system as described by the above equation is:

(a) $[SO_2]^2/[SO_3]$
(b) $[SO_2]^2[O_2]/[SO_3]^2$
(c) $[SO_3]^2/[SO_3]^2[O_2]$
(d) $[SO_2][O_2]$

Question 4

At equilibrium, __________.

(a) the rates of the forward and reverse reactions are equal
(b) the rate constants of the forward and reverse reactions are equal
(c) all chemical reactions have ceased
(d) the value of the equilibrium constant is 1

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Question 5

The value of Keq for the following reaction is 0.25:

SO_2 (g) + NO_2 (g) -> SO_3 (g) + NO (g)

The value of Keq at the same temperature for the reaction below is __________.

$2SO_2$ (g) + $2NO_2$ (g) -> $2SO_3$ (g) + 2NO (g)

(a) 0.062
(b) 16
(c) 0.25
(d) 0.50

Question 6

Consider the reaction: 2 SO_2(g) + O_2(g) $\leftrightarrow$ 2 SO_3(g). If, at equilibrium at a certain temperature, $[SO_2]$ = 1.50 M, $[O_2]$ = 0.120 M, and $[SO_3]$ = 1.25 M, what is the value of the equilibrium constant?

(a) 5.79 (b) 6.94
(c) 8.68 (d) 0.14

Question 7

What is the correct equilibrium constant expression for the following reaction? 2 Cu(s) + O_2(g) $\rightarrow$ 2 CuO(s)

(a) Keq = $1/[O_2]^2$
(b) Keq = $[CuO]^2/[Cu]^2[O_2]$
(c) Keq = $[O_2]$
(d) Keq = $1/[O_2]$

Question 8

What is the relationship of the equilibrium constants for the following two reactions?

(1) 2 NO_2(g) $\leftrightarrow$ N_2O_4(g); K_1

(2) N_2O_4(g) $\leftrightarrow$ 2 NO_2(g) K_2

(a) K_1= 1/ K_2 (b) K_2= 1/ K_1
(c) K_1= K_2 (d) both a and b are correct

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Question 9

Consider the following endothermic reaction:
$H_2(g) + I_2(g) \leftrightarrow 2\ HI(g)$. If the temperature is increased,

(a) more HI will be produced
(b) some HI will decompose, forming H_2 and I_2
(c) the magnitude of the equilibrium constant will decrease
(d) the pressure in the container will increase

Question 10

Consider the following reaction at equilibrium:
$NO_2(g) + CO(g) \leftrightarrow NO(g) + CO_2(g)$. Suppose the volume of the system is decreased at constant temperature, what change will this cause in the system?

(a) A shift to produce more NO
(b) A shift to produce more CO
(c) A shift to produce more NO_2
(d) No shift will occur

Question 11

Which of these four factors can change the value of the equilibrium constant?

(a) catalyst
(b) pressure
(c) concentration
(d) temperature

Question 12

Which general rule helps predict the shift in direction of an equilibrium reaction?

(a) Le Chatelier's principle
(b) Haber process
(c) Equilibrium constant
(d) Bosch theory

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Acids & Bases pH Calculations

Chapter 7

Chang-chapter15

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Acids & Bases

Definition of acids and bases

Arrhenius concept

Brønsted-Lowry concept

Lewis concept

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1- Arrhenius Concept

An acid is a compound that releases H^+ ions in water
A base is a compound that releases OH^- in water.

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

$$NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

Limitations: Some bases do not contain OH^-

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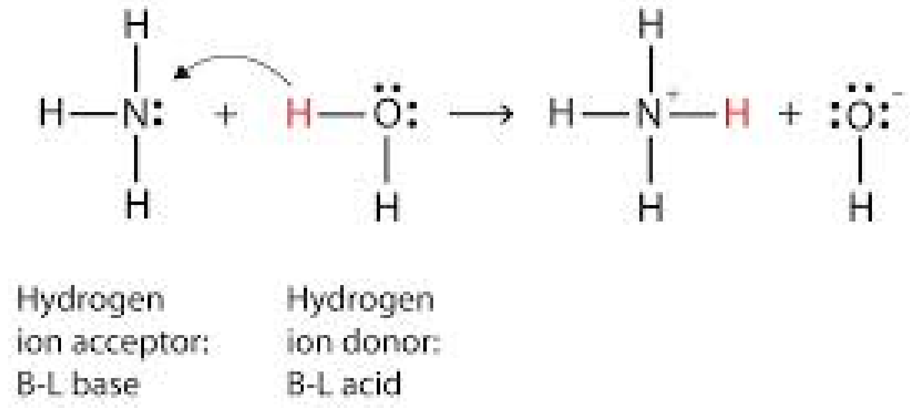
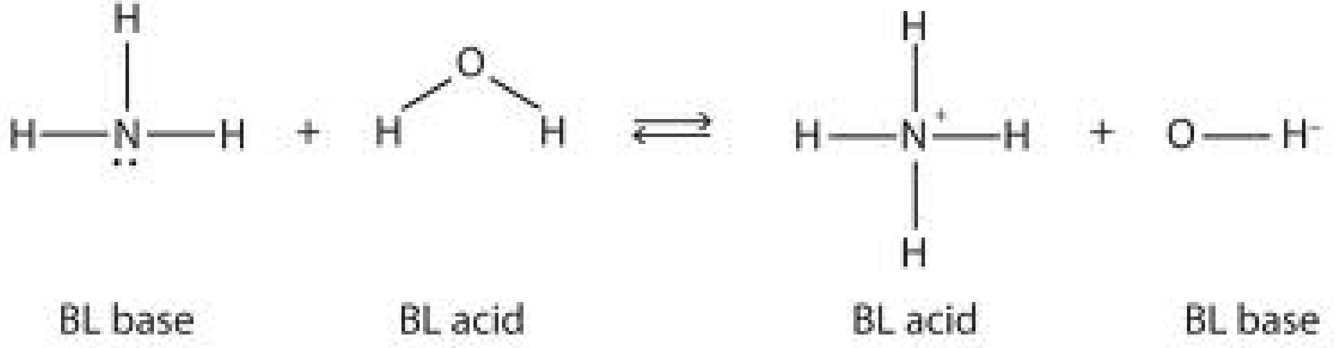
2- Brønsted-Lowry Concept

An acid is any molecule or ion that can donate a proton H^+. A base is any molecule or ion can accept a proton.



- proton-transfer reaction



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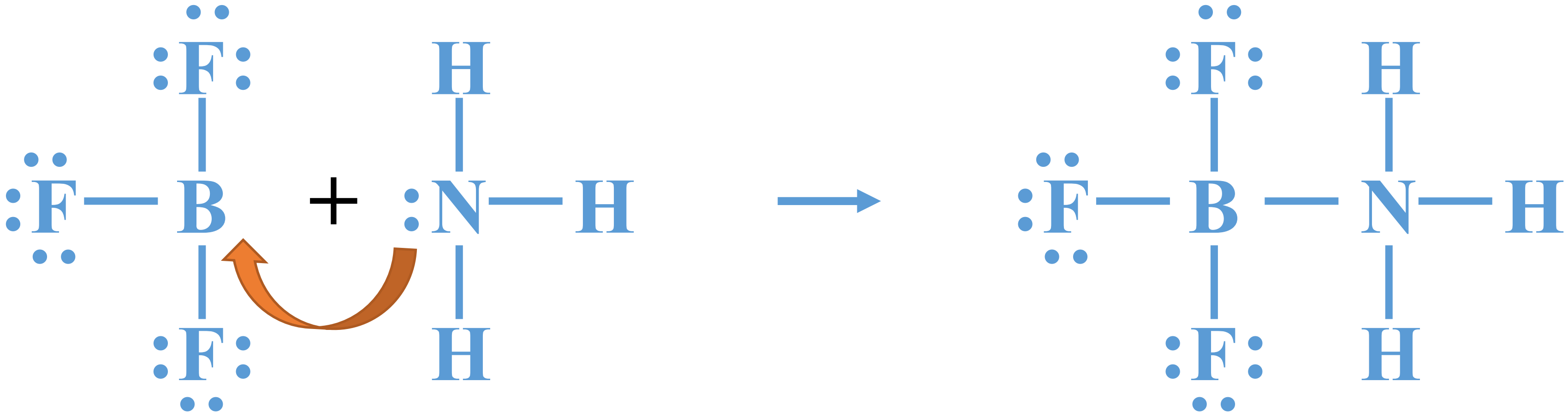
3- Lewis Concept

An acid as an electron pair acceptor and a base as an electron pair donor.



Another examples: hydration of $AlCl_3$, BCl_3, OH^-

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Strength of Acids and Bases

A strong acid or base ionizes completely in water

Strong Acids	Strong bases
HCl	$LiOH$
HBr	$NaOH$
HI	KOH
HNO_3	$Ca(OH)_2$
H_2SO_4	$Sr(OH)_2$
$HClO_4$	$Ba(OH)_2$

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Weak Acids and Bases

A weak acid or base ionizes only to a limited extent in water

Examples: CH_3COOH, NH_3

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Acid or Base Ionization Constant

It is a measure of the strength of acid or base.

The ionization constant has the same equilibrium expression.

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

$$K_a = \frac{[CH_3COO^-]\,[H_3O^+]}{[CH_3COOH]}$$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + HO^-$$

$$K_b = \frac{[NH_4^+]\,[HO^-]}{[NH_3]}$$

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Self-ionization of water

Water acts either as an acid or a base

$$H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$$

$$K_w = [H_3O^+][OH^-]$$

Or

$$K_w = [H^+][OH^-]$$

K_w = water dissociation constant

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Self-ionization of water

$$K_w = [H^+][OH^-]$$

$$K_w = 1.0\times 10^{-14} \quad \text{at } 25\ ^{o}C$$

$$[H^+] = [OH^-] = \sqrt{1.0\times 10^{-14}} = 1.0\times 10^{-7}$$

At 25°C, you observe the following conditions.

an acidic solution, $[H^+] > [OH^-]$
a neutral solution, $[H^+] = [OH^-]$
a basic solution, $[H^+] < [OH^-]$

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pH of Solutions

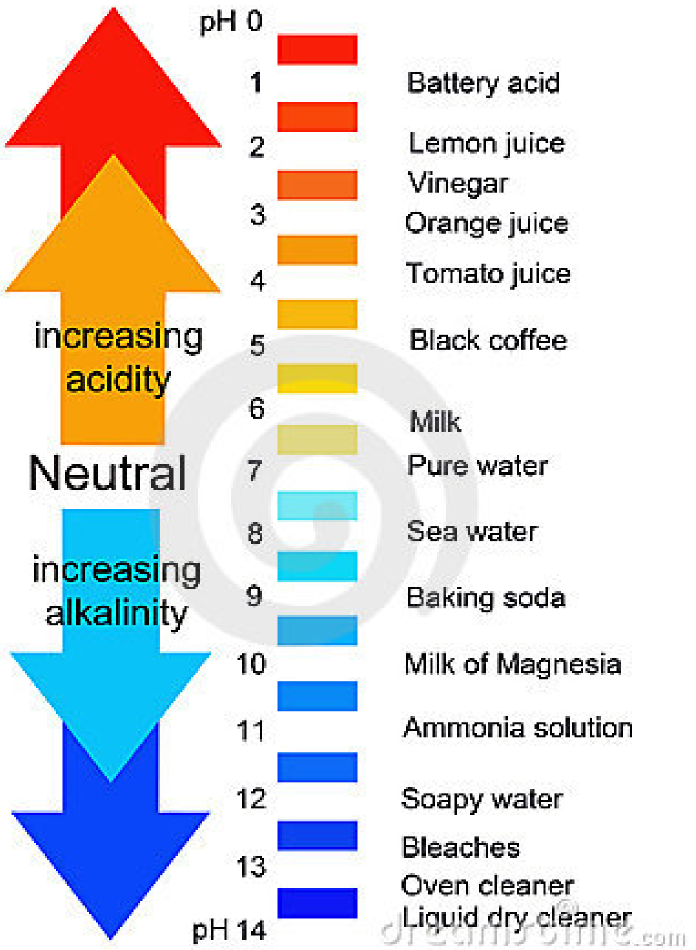
The pH of a solution is defined as the negative logarithm of the molar hydrogen-ion concentration

$$pH = -\log[H^+]$$

$$[H^+]=10^{-pH}$$

$$pH + pOH = 14.00$$

In a **neutral solution**, whose hydrogen-ion concentration is 1.0 x 10^{-7}, the **pH = 7.00**



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pH of Solutions

At 25°C, you observe the following conditions

In an acidic solution, $[H^+] > 1.0 \times 10^{-7}$ M, pH<7

In a neutral solution, $[H^+] = 1.0 \times 10^{-7}$ M, pH=7

In a basic solution, $[H^+] < 1.0 \times 10^{-7}$ M, pH>7

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Example

For a solution in which the hydrogen-ion concentration is 1.0 x 10-3, the pH is:

$$pH = -\log(1.0 \times 10^{-3}) = 3.00$$

Note that the number of decimal places in the pH equals the number of significant figures in the hydrogen-ion concentration

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Examples

The hydrogen ion concentration of a fruit juice is 3.3×10^{-2} M. What is the pH of the juice? Is it acidic or basic?

$$pH = -\log(3.3\times10^{-2}) = -(-1.48) = 1.48$$

If a solution has pH of 5.50, calculate its $[OH^-]$

$$14 = pH + pOH$$

$$pOH = 14.00 - 5.50 = 8.50$$

$$pOH = -\log[OH^-]$$

$$\log[OH^-] = -8.50$$

$$[OH^-] = 10^{-8.50} = 3.2 \times 10^{-9} M$$

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pH of Strong Acids and Bases

Dissociation of a strong base:

$$NaOH \longrightarrow Na^+ + OH^-$$

complete dissociation of a base
and no base in the form of NaOH will be left in solution

$$pOH = -\log[OH^-]$$

$$pH = 14 - pOH = 14 + \log[OH^-]$$

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Example

An ammonia solution has a hydroxide-ion concentration of 1.9 x 10-3 M. What is the pH of the solution?

You first calculate the pOH:

$$pOH = -\log(1.9 \times 10^{-3}) = 2.72$$

Then the pH is:

$$pH = 14.00 - 2.72 = 11.28$$

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pH of Weak Acids and Bases

Dissociation of weak acids ($\approx Ka < 10^{-4}$)

Examples: K_a (HF)=7.1 x 10^{-4} , K_a (HCOOH)=1.7 x 10^{-4}

$$\underset{c-x}{HA} \longleftrightarrow \underset{x}{A^-} + \underset{x}{H^+}$$

$$K_a = \frac{[A^-][H^+]}{[HA]} = \frac{x^2}{c-x} = \frac{x^2}{c}$$

c-x = concentration of an acid at equilibrium
x = concentration of products at equilibrium
c = concentration of an acid at the beginning

$c >> x$
for diluted weak acids

$$[H^+] = x = (K_a\, c)^{1/2}$$

$$\boxed{pH = -\log [H^+] = -\log (K_a\, c)^{1/2}}$$

$$\boxed{pK_a = -\log K_a}$$

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7th lecture

Question 1

The solution with the lowest pH is

A. 1.0M HF
B. 1.0M HCN
C. 1.0M HCOOH
D. 1.0M CH_3COOH

Question 2

As the $[H_3O^+]$ in a solution decreases, the $[OH^-]$

A. increases and the pH increases
B. increases and the pH decreases
C. decreases and the pH increases
D. decreases and the pH decreases

Question 3

The value of pK_w at 25°C is

A. 1.0×10^{-14}
B. 1.0×10^{-7}
C. 7.00
D. 14.00

Question 4

Which of the following describes the relationship between $[H_3O^+]$ and $[OH^-]$?

A. $[H_3O^+][OH^-] = 14.00$
B. $[H_3O^+] + [OH^-] = 14.00$
C. $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$
D. $[H_3O^+] + [OH^-] = 1.0 \times 10^{-14}$

Question 5

What is the pOH of 0.1 M NaOH?

A. 1
B. 0.0032
C. 0.40
D. 13.60

Question 6

The pH of a solution for which $[OH^-] = 1.0 \times 10^{-6}$ is

A. 1.00
B. 8.00
C. 6.00
D. –6.00

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Question 7

The ionization of water at room temperature is represented by

A. $H_2O = 2H^+ + O^{2-}$

B. $2H_2O = 2H_2 + O_2$

C. $2H_2O = H_2 + 2OH^-$

D. $2H_2O = H_3O^+ + OH^-$

Question 8

According to the Bronsted-Lowry theory, a base is a(n)

A. proton donor

B. proton acceptor

C. electron donor

D. electron acceptor

Question 9

the pH of 1.0 M acetic acid (Ka is 1.86 x10^{-5} at 20 °C).

A. 1.37

B. 2.37

C. 3.73

D. 4.73

Question 10

Addition of HCl to water causes

A. both $[H_3O^+]$ and $[OH^-]$ to increase

B. both $[H_3O^+]$ and $[OH^-]$ to decrease

C. $[H_3O^+]$ to increase and $[OH^-]$ to decrease

D. $[H_3O^+]$ to decrease and $[OH^-]$ to increase

Question 11

Which of the following statements concerning Arrhenius acids and Arrhenius bases is correct?

A. In the pure state, Arrhenius acids are covalent compounds.

B. In the pure state, Arrhenius bases are ionic compounds

C. Dissociation is the process by which Arrhenius acids produce H^+ ions in solution

D. Arrhenius bases are also called hydroxide bases

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Thermochemistry

Chapter
8

Chang-chapter6

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Energy

Energy is the capacity to do work.

- ***Thermal energy*** is the energy associated with the random motion of atoms and molecules
- ***Chemical energy*** is the energy stored within the bonds of chemical substances
- ***Nuclear energy*** is the energy stored within the collection of neutrons and protons in the atom
- ***Potential energy*** is the energy available by virtue of an object's position

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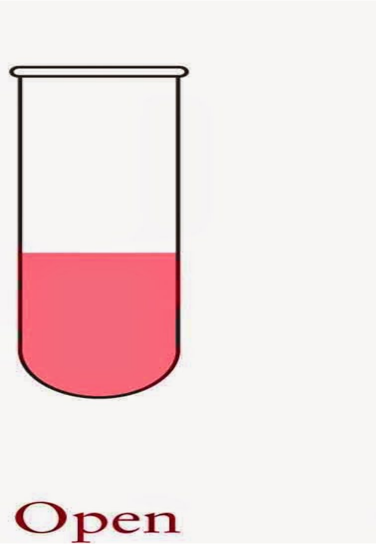
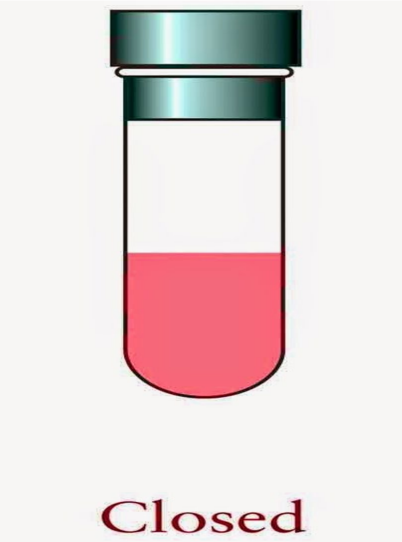
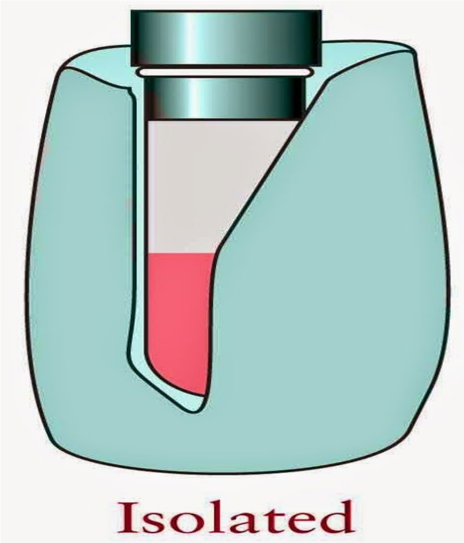
Kinds of Systems

Open system	Closed system	Isolated system
can exchange mass and energy	allows the transfer of energy (heat) but not mass	doesn't allow transfer of either mass or energy







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Examples

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Thermodynamics

Thermodynamics is the scientific study of the interconversion of heat and other kinds of energy

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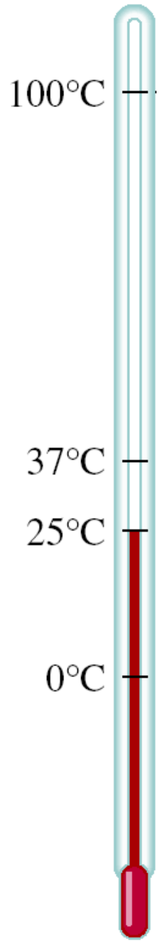
Heat (q)

Heat is the transfer of thermal energy between two bodies that are at different temperatures.

Temperature is a measure of the thermal energy



Temperature $\neq$ Thermal Energy

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First Law of Thermodynamics

First Law: Energy of the Universe is Constant

$$E = q + w$$

q = heat. Transferred between two bodies

w = work. Force acting over a distance (F x d)

w = F x d

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Thermodynamic State Functions

- Thermodynamic State Functions: Thermodynamic properties that are dependent on the state of the system only regardless of the pathway. Examples: (Energy, pressure, volume, temperature)

$$\Delta E = E_{final} - E_{initial}$$

$$\Delta P = P_{final} - P_{initial}$$

$$\Delta V = V_{final} - V_{initial}$$

$$\Delta T = T_{final} - T_{initial}$$

- Other variables will be dependent on pathway (Examples: q and w). These are Path Functions. The pathway from one state to the other must be defined.

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Thermochemistry

Thermochemistry is the study of **heat change** in chemical reactions.

Exothermic process is any process that gives off heat – transfers thermal energy from the system to the surroundings.

$$2H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(l) + \text{energy}$$

$$H_2O\,(g) \longrightarrow H_2O\,(l) + \text{energy}$$

Endothermic process is any process in which heat has to be supplied to the system from the surroundings.

$$\text{energy} + 2HgO\,(s) \longrightarrow 2Hg\,(l) + O_2\,(g)$$

$$\text{energy} + H_2O\,(s) \longrightarrow H_2O\,(l)$$

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Enthalpy of Chemical Reactions

Definition of Enthalpy

- Thermodynamic Definition of Enthalpy (H):

$$H = E + PV$$

E = energy of the system

P = pressure of the system

V = volume of the system

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Changes in Enthalpy (ΔH)

- Consider the following expression for a chemical process:

$$\Delta H = H_{products} - H_{reactants}$$

If $\Delta H > 0$, then $q_p > 0$. (+) The reaction is endothermic

If $\Delta H < 0$, then $q_p < 0$. (-) The reaction is exothermic

$$\Delta H = q_p$$

q_p : heat at constant pressure

Calorimetry: the measurement of heat change

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Kinds of Processes
(chemical reactions or physical changes)

Endothermic processes

H_2O (s)

Heat absorbed from system to surroundings

$q = +$

H_2O (l)

Exothermic processes

CH_4 (g) + $2O_2$ (g)

Heat released from system to surroundings

$q = -$

CO_2 (g) + $2H_2O$ (l)

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Standard Enthalpy (Heat) of reaction (ΔH^o_{rxn})

Enthalpy change at standard conditions (25 °C, 1 atm)

$$N_{2\,(g)} + 3H_{2\,(g)} \longrightarrow 2NH_{3\,(g)} \quad \Delta H^o = -92.38\ kJ$$

Thermochemical reaction

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Standard Heat of formation (ΔH_f^o)

The heat change that results when 1 mol of the compound is formed from standard state of its elements

The standard enthalpy of formation of any element in its most stable form is zero.

ΔH^0 (C, diamond) = 1.90 kJ/mol

What is ΔH_f^o of O_2 (g), Hg(l), C(graphite)?

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Thermochemical Equations

$$CH_4\ (g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2H_2O\ (l) \quad \Delta H = -890.4\ \text{kJ/mol}$$

- It shows the physical states of all products and reactants
- Balanced
- It shows Heat of reaction kJ

$$H_2O\ (s) \longrightarrow H_2O\ (l) \quad \Delta H = 6.01\ \text{kJ/mol}$$

- If you reverse a reaction, the sign of ΔH changes

$$H_2O\ (l) \longrightarrow H_2O\ (s) \quad \Delta H = -6.01\ \text{kJ/mol}$$

- If you multiply both sides of the equation by a factor n, then ΔH must change by the same factor n.

$$2H_2O\ (s) \longrightarrow 2H_2O\ (l) \quad \Delta H = 2 \times 6.01 = 12.0\ \text{kJ}$$

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How to calculate ΔH_r^o

1- Direct method 2- indirect method

1- Direct method: by standard heat of formation

$$\Delta H^o = \sum n \, \Delta H_f^o \text{ (products)} - \sum n \, \Delta H_f^o \text{ (reactants)}$$

n = no. of moles in the balanced thermochemical equation

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Example

Calculate the enthalpy of the following reaction:

$$2Al_{(s)} + Fe_2O_{3\,(s)} \longrightarrow Al_2O_{3\,(s)} + 2Fe_{(l)}$$

ΔH_f^o *of* Fe_2O_3, Al_2O_3 and Fe(l) = - 822.2, - 1669.8 and 12.40 kJ/mol

$$\Delta H^o = \sum n\, \Delta H_f^o\,(\text{products}) - \sum n\, \Delta H_f^o\,(\text{reactants})$$

$$\Delta H^o = [(\Delta H_f^o\,(Al_2O_3)) + (2\times \Delta H_f^o\,(Fe))] - [(2\times\Delta H_f^o\,(Al)) + (\Delta H_f^o\,(Fe_2O_3))]$$

$$\Delta H^o = [\,(-1669.8) + (2\times 12.40)] - [2\times(0) + (-822.2)] = -822.8\ \text{kJ}$$

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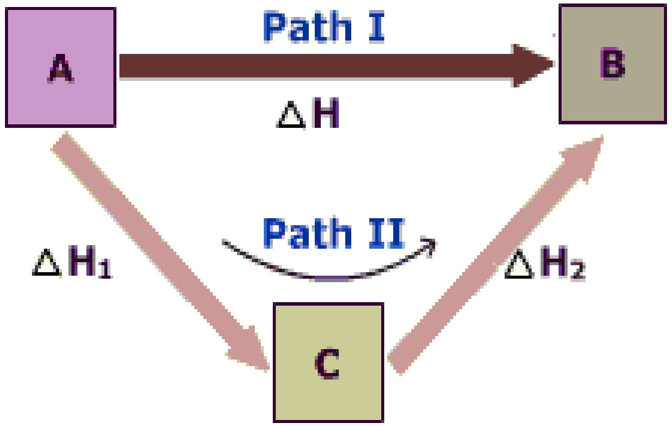
2- indirect method :(Hess's Law)

ΔH for a process involving the transformation of reactants into products is not dependent on pathway. Therefore, we can pick any pathway to calculate ΔH for a reaction.



When reactants are converted to products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps.

ΔH is a state function

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Hess' Law: Details

- Once can always reverse the direction of a reaction when making a combined reaction. When you do this, the sign of ΔH changes.

$$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H = 68\ kJ$$

$$2NO_2(g) \longrightarrow N_2(g) + 2O_2(g) \quad \Delta H = -68\ kJ$$

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Hess' Law: Details (cont.)

- If the coefficients of a reaction are multiplied by a constant, the value of ΔH is also multiplied by the same integer.

$$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H = 68\ kJ$$

$$2N_2(g) + 4O_2(g) \longrightarrow 4NO_2(g) \qquad \Delta H = 136\ kJ$$

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2- Hess's Law

Calculate the enthalpy of the following reaction:

$Fe_2O_3\,(s) + 3CO\,(g) \longrightarrow 2Fe\,(s) + 3CO_2\,(g)$

1- $CO\,(g) + ½O_2\,(g) \longrightarrow CO_2\,(g)$ $\quad \Delta H^o = -\ 283.0$ kJ

2- $2Fe\,(s) + 3/2O_2\,(g) \longrightarrow Fe_2O_3\,(s)$ $\quad \Delta H^o = -\ 822.2$ kJ

$3CO\,(g) + \cancel{3/2O_2}\,(g) \longrightarrow 3CO_2\,(g)$ $\quad \Delta H^o = 3\times -\ 283.0 = -\ 849$ kJ

$Fe_2O_3\,(s) \longrightarrow 2Fe\,(s) + \cancel{3/2O_2}\,(g)$ $\quad \Delta H^o = +\ 822.2$ kJ

$Fe_2O_3\,(s) + 3CO\,(g) \longrightarrow 2Fe\,(s) + 3CO_2\,(g)$ $\quad \Delta H^o = -26.7$ kJ

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Question 1

An exothermic reaction causes the surroundings to:

A. become basic
B. decrease in temperature
C. condense
D. increase in temperature
E. decrease in pressure

Question 2

How much heat is evolved when 320 g of SO_2 is burned according to the chemical equation shown below?

$2\ SO_2(g) + O_2(g) \longrightarrow 2\ SO_3(g)\quad \Delta H^o_{rxn} = -198\ kJ$

A. 5.04 x 10-2 kJ
B. 9.9 x 102 kJ
C. 207 kJ
D. 5.0 x 102 kJ
E. None of the above

Question 3

The specific heat of aluminum is 0.214 cal/g.oC. Determine the energy, in calories, necessary to raise the temperature of a 55.5 g piece of aluminum from 23.0 to 48.6oC.

A. 109 cal
B. 273 cal
C. 577 cal
D. 347 cal
E. 304 cal

Question 4

Energy is the ability to do work and can be:

A. converted to one form to another
B. can be created and destroyed
C. used within a system without consequences
D. none of the above

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Question 5

To which one of the following reactions, occurring at 25°C, does the symbol ΔH^o_f [H_2SO_4(l)] refer?

A. H2(g) + S(s) + 2 O2(g) ----> H2SO4(l)
B. H2SO4(l) ----> H2(g) + S(s) + 2 O2(g)
C. H2(g) + S(g) + 2 O2(g) ----> H2SO4(l)
D. H2SO4(l) ----> 2 H(g) + S(s) + 4 O(g)
E. 2 H(g) + S(g) + 4 O(g) ----> H2SO4(l)

Question 6

Given: SO2(g) + ½O2(g) ----> SO3(g) ΔH^o_{rxn} = -99 kJ, what is the enthalpy change for the following reaction? 2 SO3(g) ----> O2(g) + 2 SO2(g)

A. 99 kJ
B. -99 kJ
C. 49.5 kJ
D. -198 kJ
E. 198 kJ

Question 7

The specific heat of aluminum is 0.214 cal/g.oC. Determine the energy, in calories, necessary to raise the temperature of a 55.5 g piece of aluminum from 23.0 to 48.6oC.

A. 109 cal
B. 273 cal
C. 577 cal
D. 347 cal
E. 304 cal

Question 8

Standard enthalpy of reactions can be calculated from standard enthalpies of formation of reactants.

A. True
B. False

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Question 9

Calculate ΔH^o_{rxn} for the combustion reaction of CH_4 shown below given the following:
ΔH^of $CH_4(g)$ = -74.8 kJ/mol;
ΔH^of $CO_2(g)$ = -393.5 kJ/mol;
ΔH^of $H_2O(l)$ = -285.5 kJ/mol.
$CH_4(g) + 2\ O_2(g) \text{ ----> } CO_2(g) + 2\ H_2O(l)$

A. -604.2 kJ
B. 889.7 kJ
C. -997.7 kJ
D. -889.7 kJ
E. None of the above

Question 10

Find the standard enthalpy of formation of ethylene, C2H4(g), given the following data:
C2H4(g) + 3 O2(g) ----> 2 CO2(g) + 2 H2O(l)
ΔH^o_{rxn} = -1411 kJ;
C(s) + O2(g) ----> CO2(g) ΔH^o_f = -393.5 kJ;
H2(g) + ½O2(g) ----> H2O(l) ΔH^o_f = -285.8 kJ

A. 731 kJ
B. 2.77 x 103 kJ
C. 1.41 x 103 kJ
D. 87 kJ
E. 52 kJ

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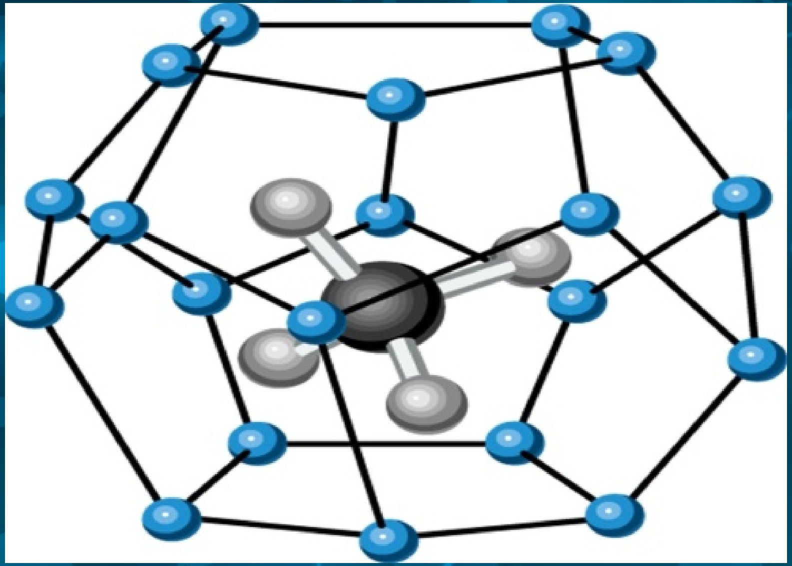

ORGANIC CHEMISTRY

Chapter 9

Chang-chapter24

COURSE NAME: CHEMISTRY 101
COURSE CODE: 402101-4

Organic Chemistry

- The study of the compounds of carbon
- Over 10 million compounds have been identified
  - about 1000 new ones are identified each day!
- C is a small atom
  - it forms single, double, and triple bonds
  - it is intermediate in electronegativity (2.5)
  - it forms strong bonds with C, H, O, N, and some metals

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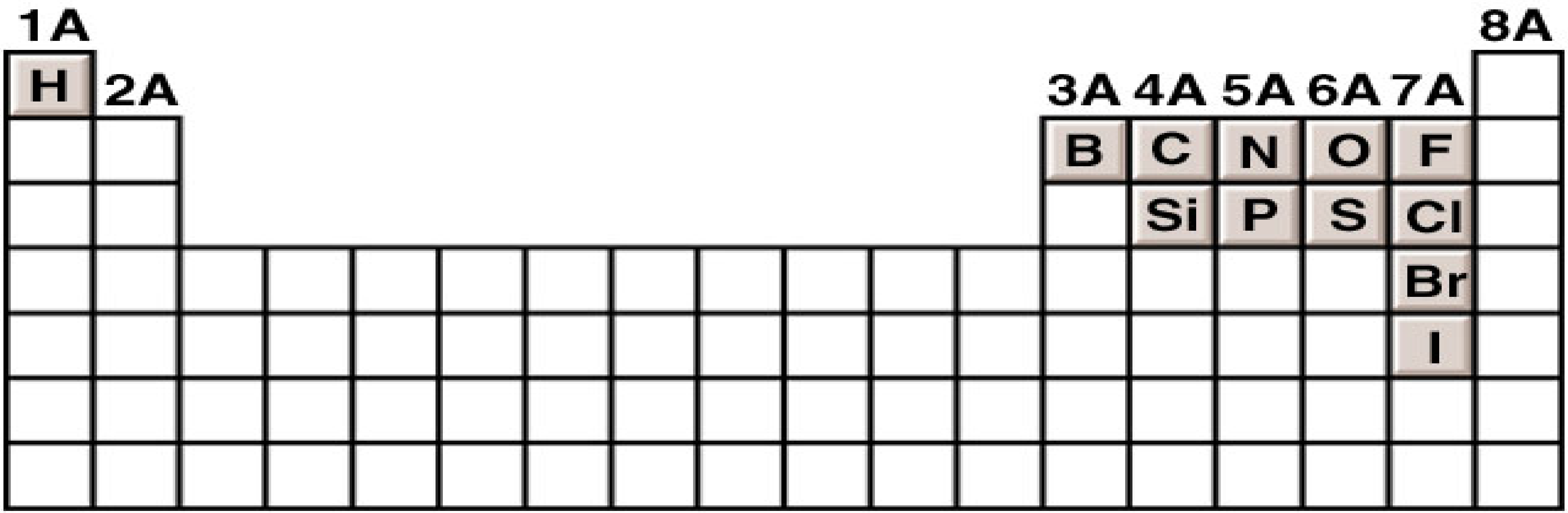
Common Elements in Organic Compounds



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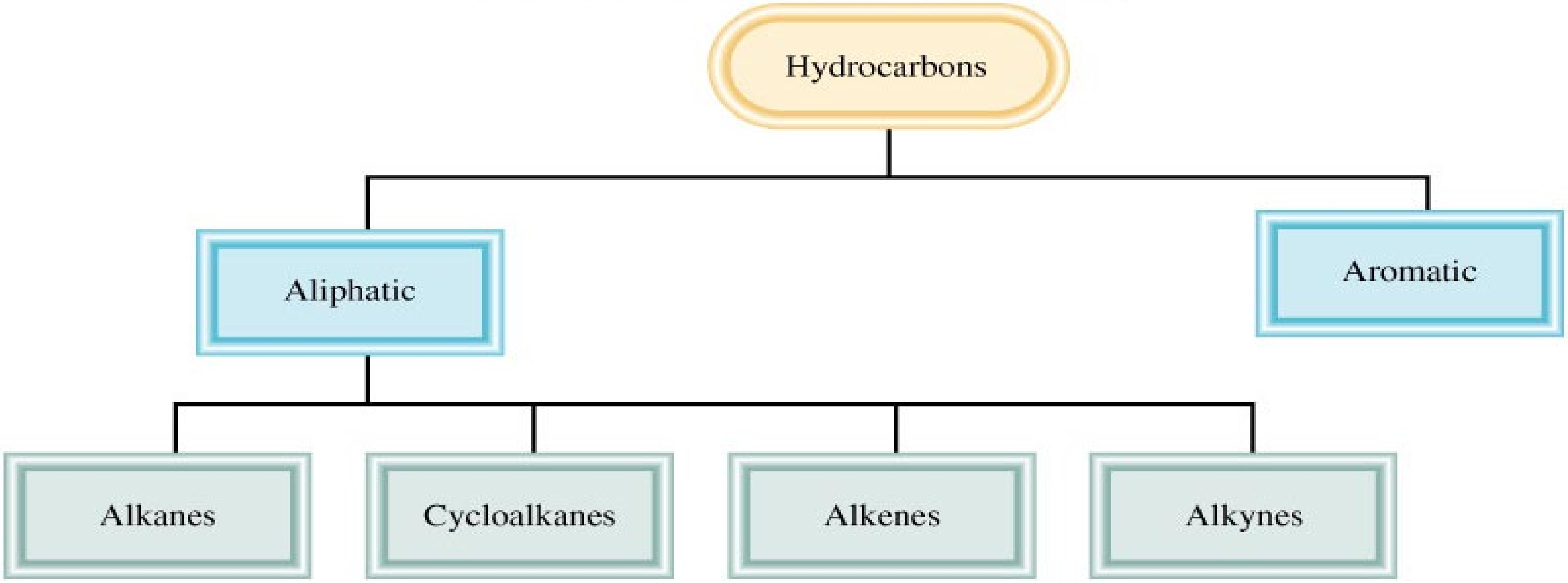
Classification of Hydrocarbons



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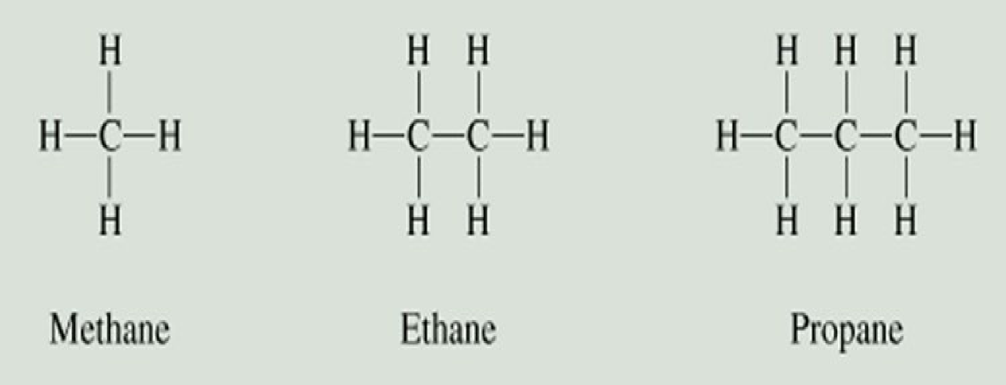
9th lecture

Alkanes

Alkanes have the general formula C_nH_{2n+2} where n = 1,2,3,…

1. only single covalent bonds
2. **saturated hydrocarbons** because they contain the **maximum** number of hydrogen atoms that can bond with the number of carbon atoms in the molecule

# of carbons	boiling point range	Use
1-4	<20 °C	fuel (gasses such as methane, propane, butane)
5-6	30-60	solvents (petroleum ether)
6-7	60-90	solvents (ligroin)
6-12	85-200	fuel (gasoline)
12-15	200-300	fuel (kerosene)
15-18	300-400	fuel (heating oil)
16-24	>400	lubricating oil, asphalt



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9th lecture

Alkane Nomenclature

The First 10 Straight-Chain Alkanes

Name of Hydrocarbon	Molecular Formula	Number of Carbon Atoms	Melting Point (°C)	Boiling Point (°C)
Methane	CH_4	1	−182.5	−161.6
Ethane	$CH_3—CH_3$	2	−183.3	−88.6
Propane	$CH_3—CH_2—CH_3$	3	−189.7	−42.1
Butane	$CH_3—(CH_2)_2—CH_3$	4	−138.3	−0.5
Pentane	$CH_3—(CH_2)_3—CH_3$	5	−129.8	36.1
Hexane	$CH_3—(CH_2)_4—CH_3$	6	−95.3	68.7
Heptane	$CH_3—(CH_2)_5—CH_3$	7	−90.6	98.4
Octane	$CH_3—(CH_2)_6—CH_3$	8	−56.8	125.7
Nonane	$CH_3—(CH_2)_7—CH_3$	9	−53.5	150.8
Decane	$CH_3—(CH_2)_8—CH_3$	10	−29.7	174.0

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Each member C_3 - C_{10} differs by one CH_2 unit. This is called a homologous series.
Methane to butane are gases at normal pressures.
Pentane to decane are liquids at normal pressures.

Alkane Nomenclature

1. The parent name of the hydrocarbon is that given to the longest continuous chain of carbon atoms in the molecule.

$$\underset{1}{CH_3}-\underset{2}{CH_2}-\underset{3}{CH_2}-\underset{4}{\overset{\overset{\displaystyle CH_3}{|}}{CH}}-\underset{5}{CH_2}-\underset{6}{CH_2}-\underset{7}{CH_3}$$

4-methylheptane

2. Alkyl substituents: An alkane less one hydrogen atom is an alkyl group.

drop the -ane and add -yl.

$-CH_3$: methyl

$-C_2H_5$: ethyl

Common Alkyl Groups

Name	Formula
Methyl	$-CH_3$
Ethyl	$-CH_2-CH_3$
n-Propyl	$-CH_2-CH_2-CH_3$
n-Butyl	$-CH_2-CH_2-CH_2-CH_3$
Isopropyl	$-\overset{\overset{\displaystyle CH_3}{\vert}}{\underset{\underset{\displaystyle CH_3}{\vert}}{C}}-H$
t-Butyl*	$-\overset{\overset{\displaystyle CH_3}{\vert}}{\underset{\underset{\displaystyle CH_3}{\vert}}{C}}-CH_3$

*The letter *t* stands for tertiary.

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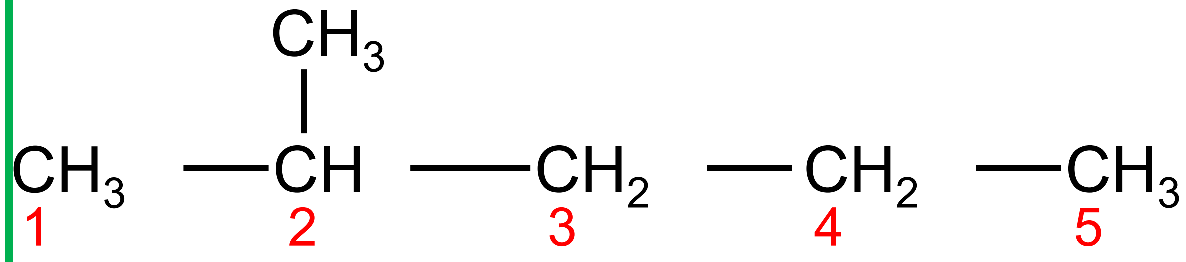
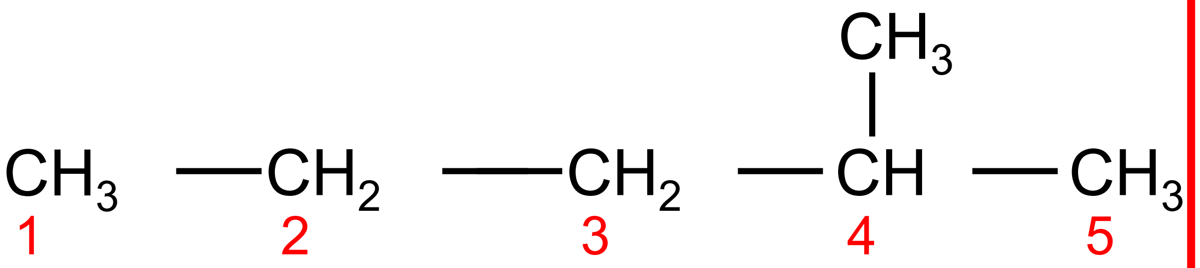
Alkane Nomenclature

3. When one or more hydrogen atoms are replaced by other groups, the name of the compound must indicate the locations of carbon atoms where replacements are made. Number in the direction that gives the smaller numbers for the locations of the branches.



2-methylpentane



4-methylpentane

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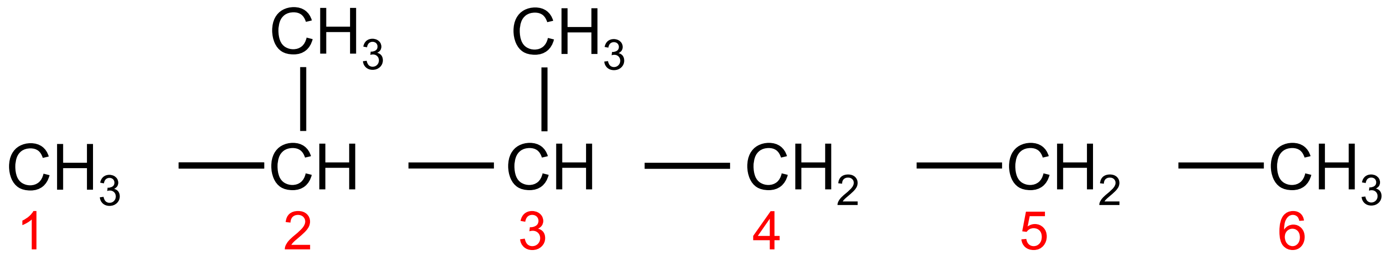
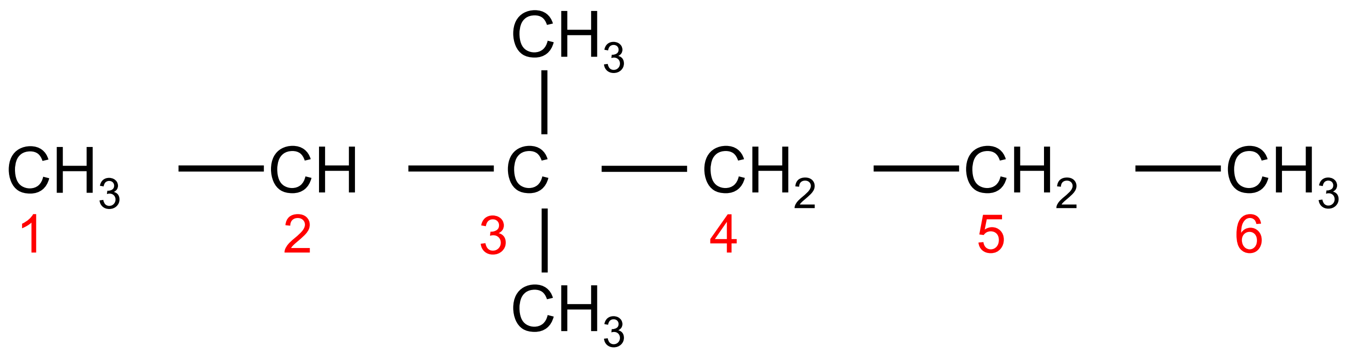
Alkane Nomenclature

4. Use prefixes *di-, tri-, tetra-*, when there is more than one alkyl branch of the same kind.



2,3-dimethylhexane



3,3-dimethylhexane

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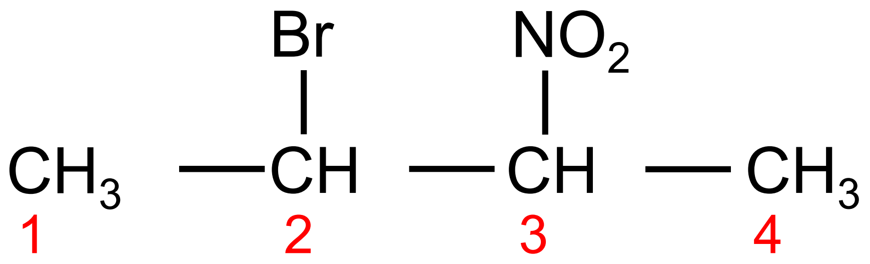
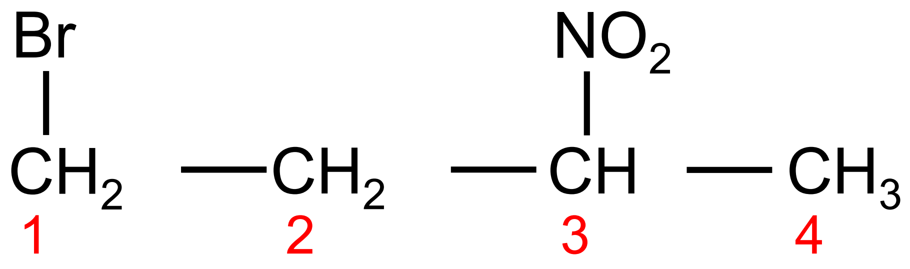
Alkane Nomenclature

5. Use previous rules for other types of substituents.



2-bromo-3-nitrobutane



1-bromo-3-nitrobutane

TABLE

Names of Common Substituent Groups

Functional Group	Name
$—NH_2$	Amino
$—F$	Fluoro
$—Cl$	Chloro
$—Br$	Bromo
$—I$	Iodo
$—NO_2$	Nitro
$—CH{=}CH_2$	Vinyl

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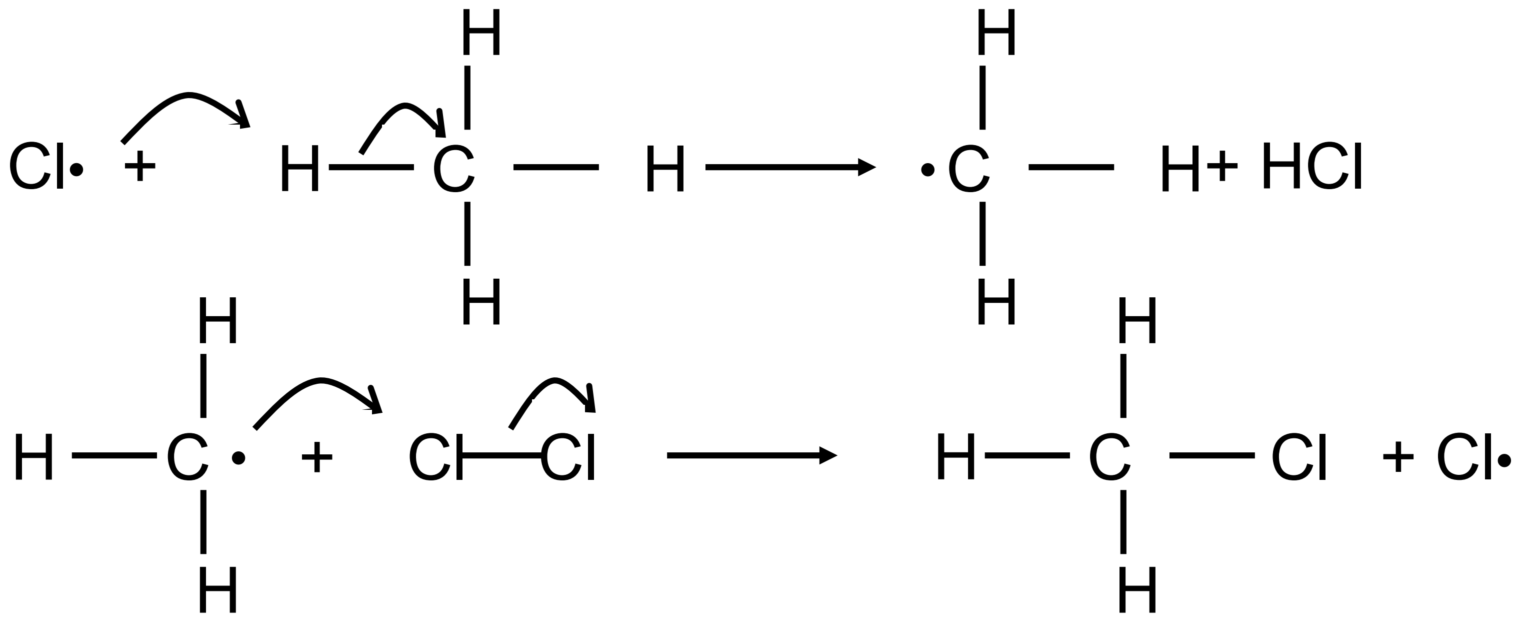
Alkane Reactions

Combustion $CH_4\ (g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2H_2O\ (l) \quad \Delta H^0 = -890.4\ \text{kJ}$

Halogenation $CH_4\ (g) + Cl_2\ (g) \xrightarrow{\text{light}} CH_3Cl\ (g) + HCl\ (g)$

$$Cl_2 + \text{energy} \longrightarrow Cl\cdot + Cl\cdot$$



Questions

1- Organic compounds must contain

A) Oxygen

B) Nitrogen

C) Hydrogen

D) Carbon

2- Which formula represents a saturated hydrocarbon?

A) C_2H_2

B) C_3H_8

C) C_3H_6

D) C_2H_4

3- How many carbon atoms are present per molecule in the compound 3-methyl-4-ethyloctane? How many of those are present on the side chains (branches) only?

A) 11 total; 3 on branches

B) 15 total; 7 on branches

C) 12 total; 3 on branches

D) 15 total; 2 on branches

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5- How many hydrogen atoms would be part of one molecule of pentane?

A) 5

B) 8

C) 10

D) 12

6- $C_2H_4 + Br_2$ = ? What reaction occurs when the above chemicals react?

A) substitution

B) Addition

C) Elimination

D) hydrolysis

7- The general formula for the alkane series is :

A) C_nH_n

B) C_nH_{2n}

C) C_nH_{2n+2}

D) C_nH_{2n-2}

8- A compound with the formula C_6H_6 is :

A) hexane

B) pentene

C) 3-methylButane

D) Benzene

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