General Chemistry

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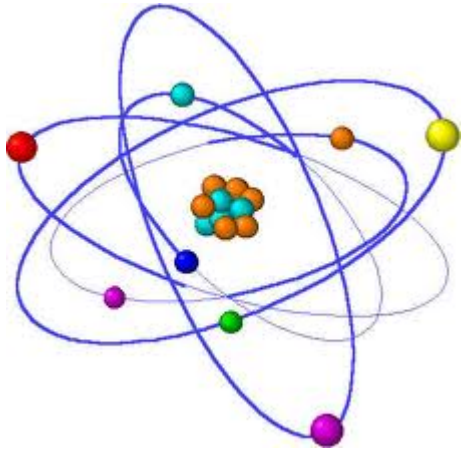
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Matter
Measurements
Significant figures

Chapter
1

Chang-chapter1

COURSE NAME: CHEMISTRY 101
COURSE CODE: 28171101-4



What is Chemistry?

Chemistry is the study of composition, structure, properties, and reactions of matter and the changes it undergoes.

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Matter

Any thing that occupies space and has mass

Separation by physical methods

pure substance

Has a definite composition and distinct properties

Separation by chemical methods

element

cannot be separated into simpler substances by chemical means

compound

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

mixture

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

heterogeneous

The composition is not uniform

homogeneous

The composition is the same throughout

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NaCl	
Salt water	
Iron	
sugar	
air	
helium	
water	
salad	

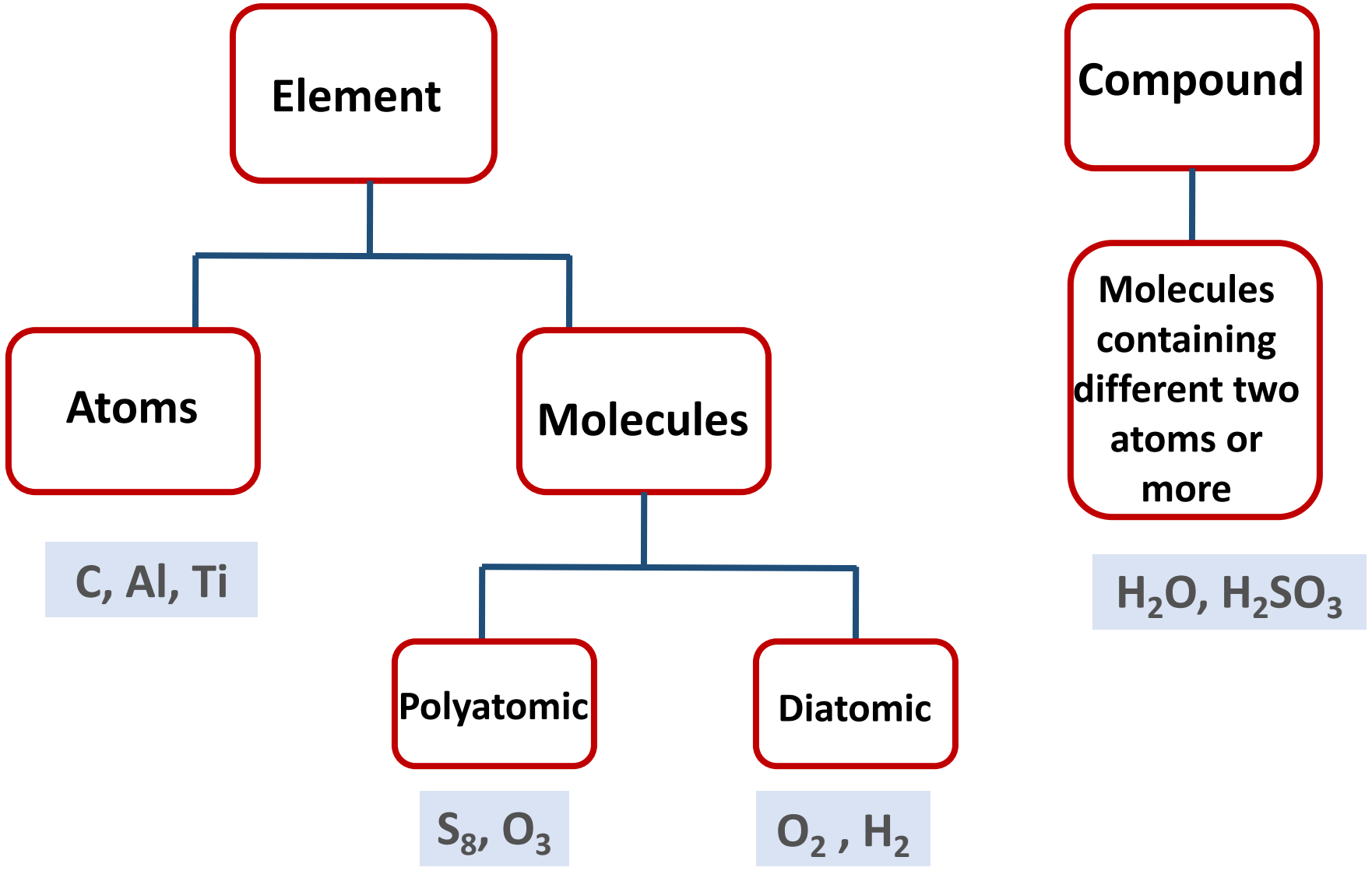
compound **element** **homogeneous mixture** **heterogeneous**

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

compound element homogeneous mixture heterogeneous

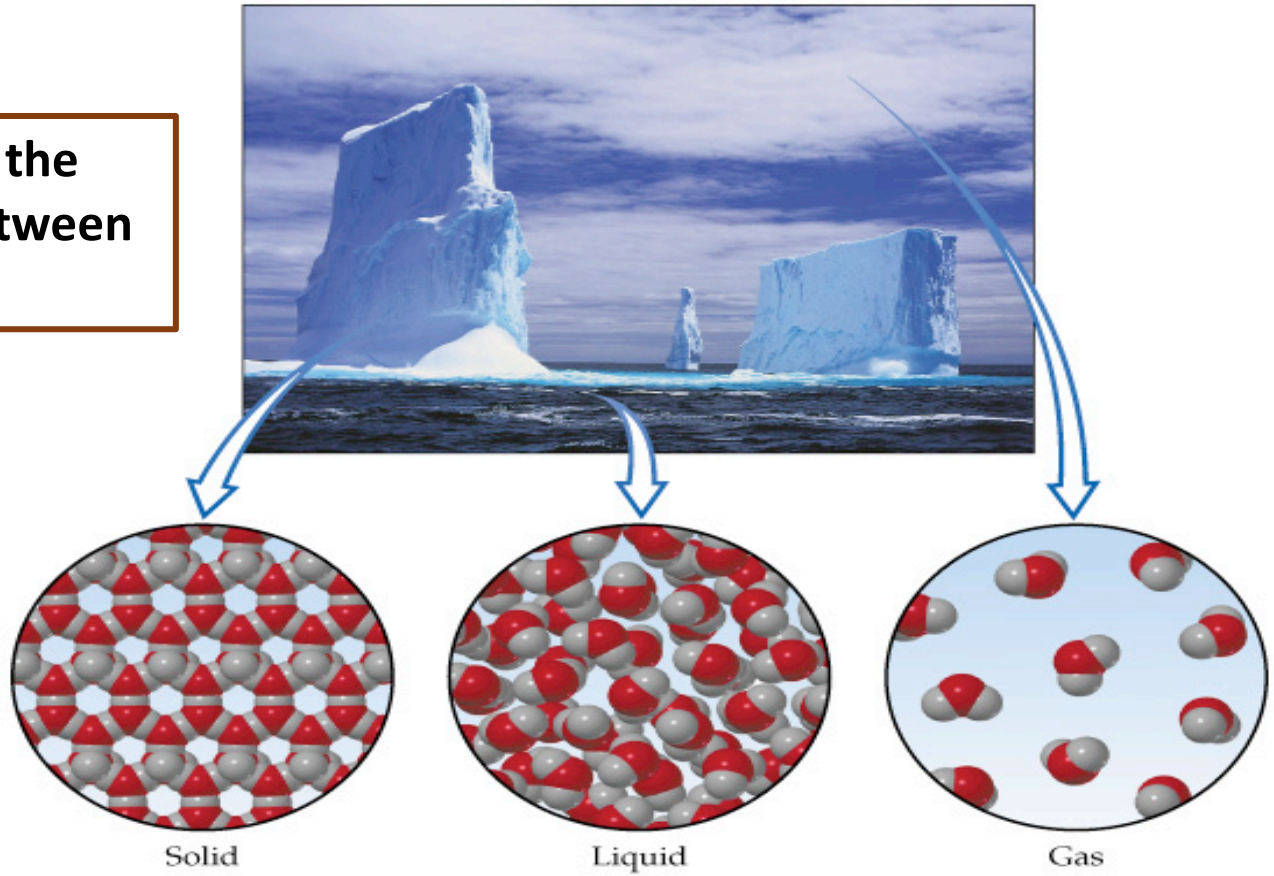
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Matter States

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

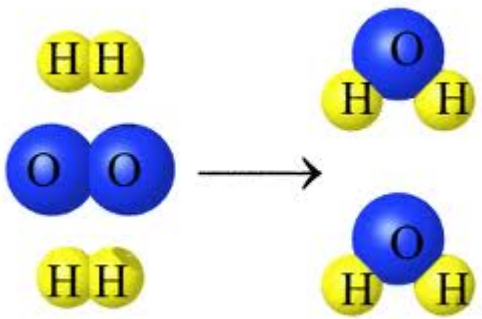


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Matter properties

Chemical

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

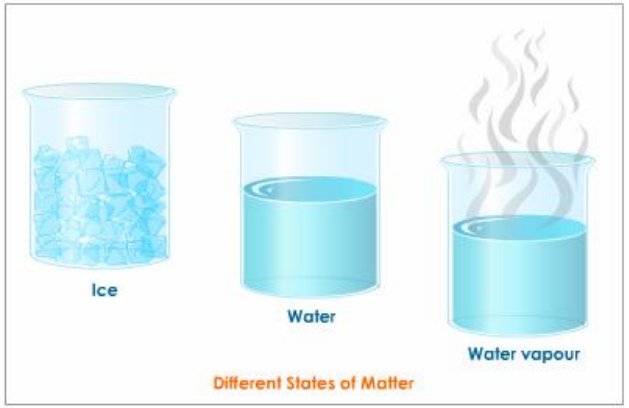


reactivity, flammability

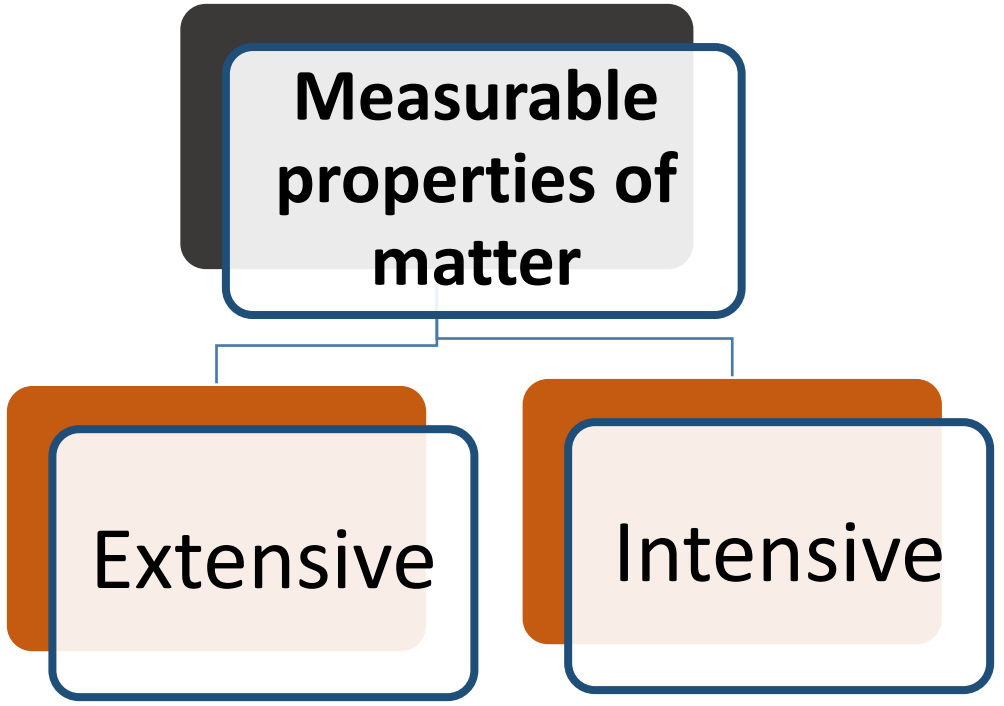
Physical

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

color, mass, size



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depends on how much matter is being considered

Does not depend on how much matter is being considered

Mass
volume

Density
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

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}

Mass and weight



What is the difference between mass and weight?

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

$1\ Kg = 1000\ g = 1 \times 10^3\ g$

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)

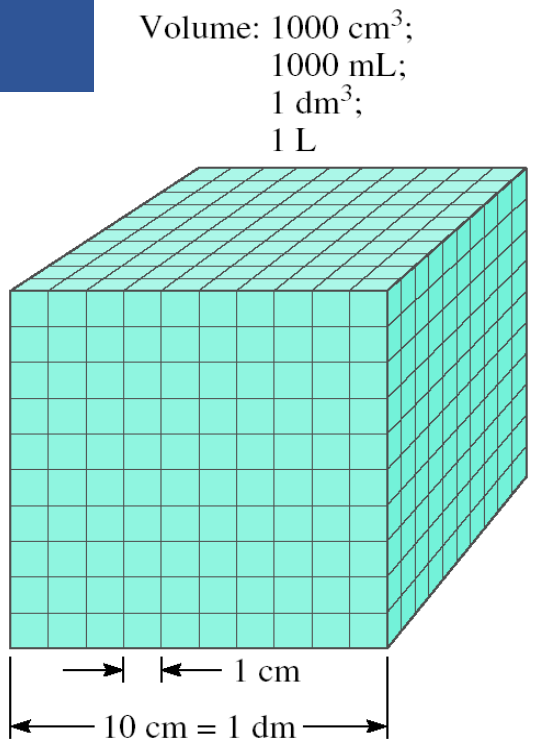


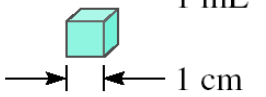
Volumetric flask

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

$1\ cm^3 = 1\ mL$

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





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Dimensional Analysis Method of Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

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

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

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Density

Density is defined as the mass per unit volume.

density = mass/volume $\quad d = \frac{m}{V}$ 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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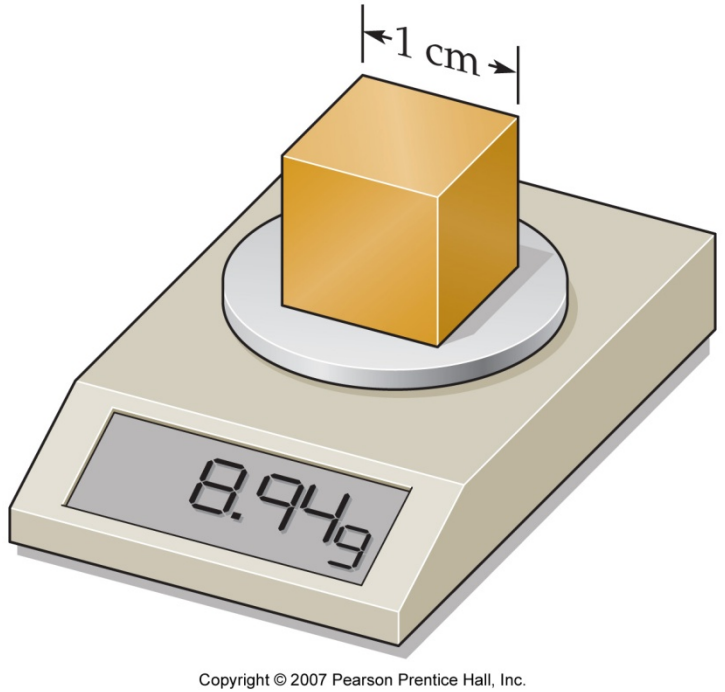
Density

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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Temperature

Temperature scales

- Fahrenheit, ^{o}F
- Celsius, ^{o}C
- Kelvin, K

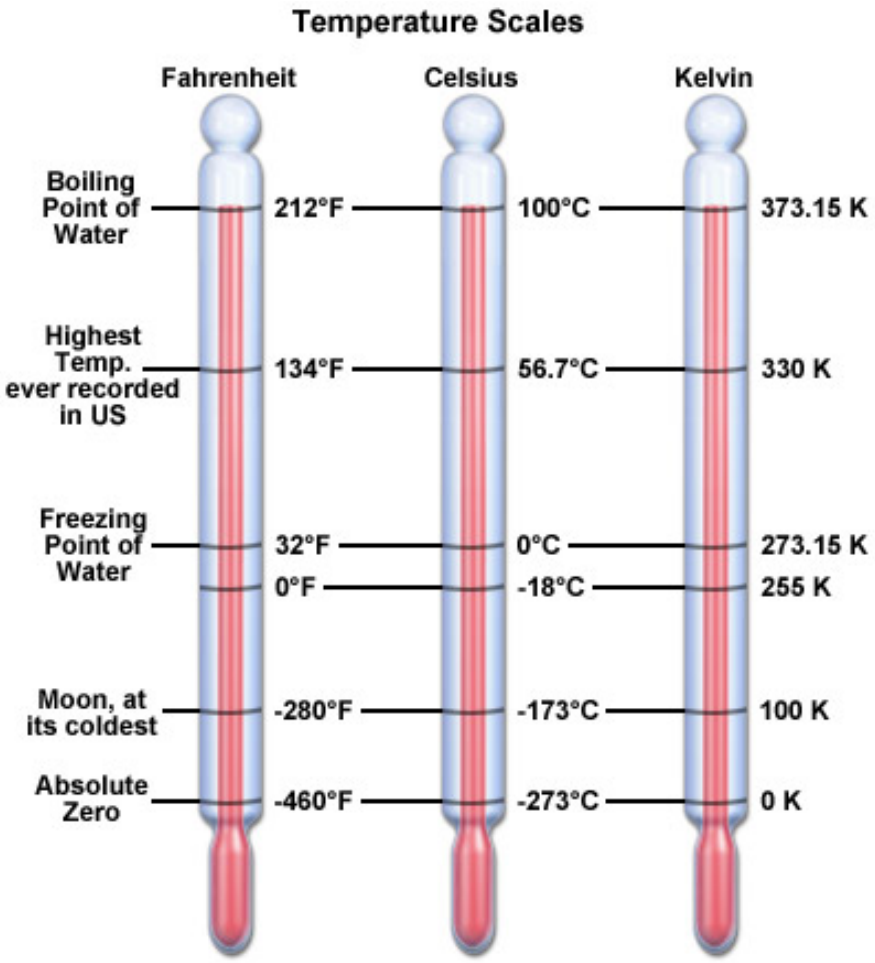
$$^{0}F = \left(\frac{9}{5} \times {}^{0}C\right) + 32$$

$32\ ^{0}F = 0\ ^{0}C$

$212\ ^{0}F = 100\ ^{0}C$

$273\ K = 0\ ^{0}C$

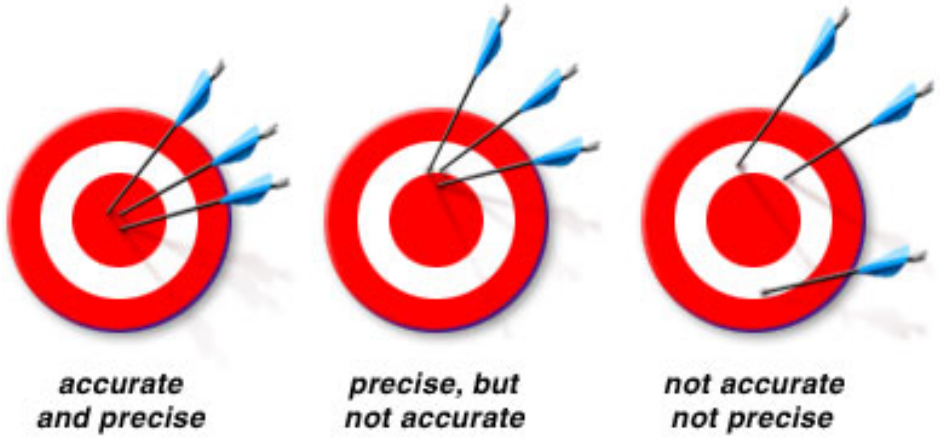
$373\ K = 100\ ^{0}C$



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

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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

$N \times 10^{n}$

n is a positive or negative integer

N is a number between 1 and 10

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:

0.000000000000000000000199

1.99×10^{-23}

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

0.00000772 = 7.72×10^{-6} (3 SF)

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

1) 24 ml
   - 2 significant figures
2) 3001 g
   - 4 significant figures
3) 0.0320 m^3
   - 3 significant figures
4) 6.4×10^4 molecules
   - 2 significant figures
5) 560 kg
   - 3 significant figures– to clarify use the scientific notation 5.60×10^2 kg

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

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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

$$4.31 \times 10^4$$
$$+$$
$$3.9 \times 10^3 \quad (0.39 \times 10^4)$$
$$= \quad 4.70 \times 10^4 \text{ (3 SF)}$$

$$7.4 \times 10^3$$
$$+$$
$$0.10 \times 10^3$$

(1 decimal digit: this has the smallest digit)

$$= \quad 7.5 \times 10^3 \text{ (2 SF)}$$

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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.

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

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

$$4.51 \times 3.6666 = 16.53636 \approx 16.5$$

(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?

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

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.

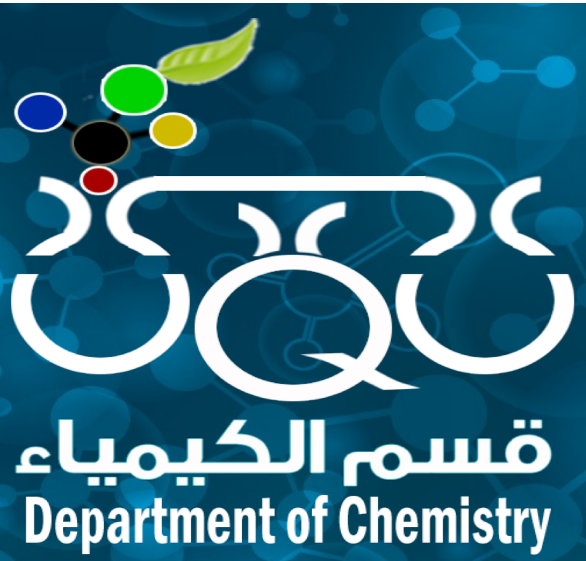
C) work was only qualitative.

D) work was precise.

E) C and D.

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Atomic Weight
Molecular Weight
Moles Calculations

Chapter
2

Chang-chapter2

COURSE NAME: CHEMISTRY 101
COURSE CODE: 28171101-4

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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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

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 of CH_4 = $6.07\ g_{(CH4)} \times \left(\frac{1\ mol_{(CH4)}}{16.04\ g_{(CH4)}}\right) = 0.378\ mol_{(CH4)}$

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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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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

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

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

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

Mass percent of O $= \frac{1 \times 16.00 \text{ g/mol}}{46.07 \text{ g/mol}} \times 100 = 34.73\ \%$ (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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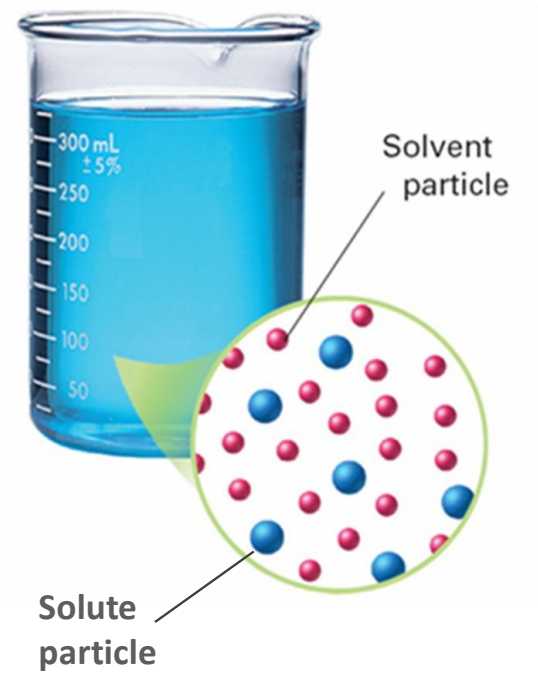
Chemical Reactions in Solutions & Concentrations

Chapter 3

Chang-chapter3,4,12

COURSE NAME: CHEMISTRY 101
COURSE CODE: 28171101-4

Solutions



Solution: a homogeneous mixture of two or more substances

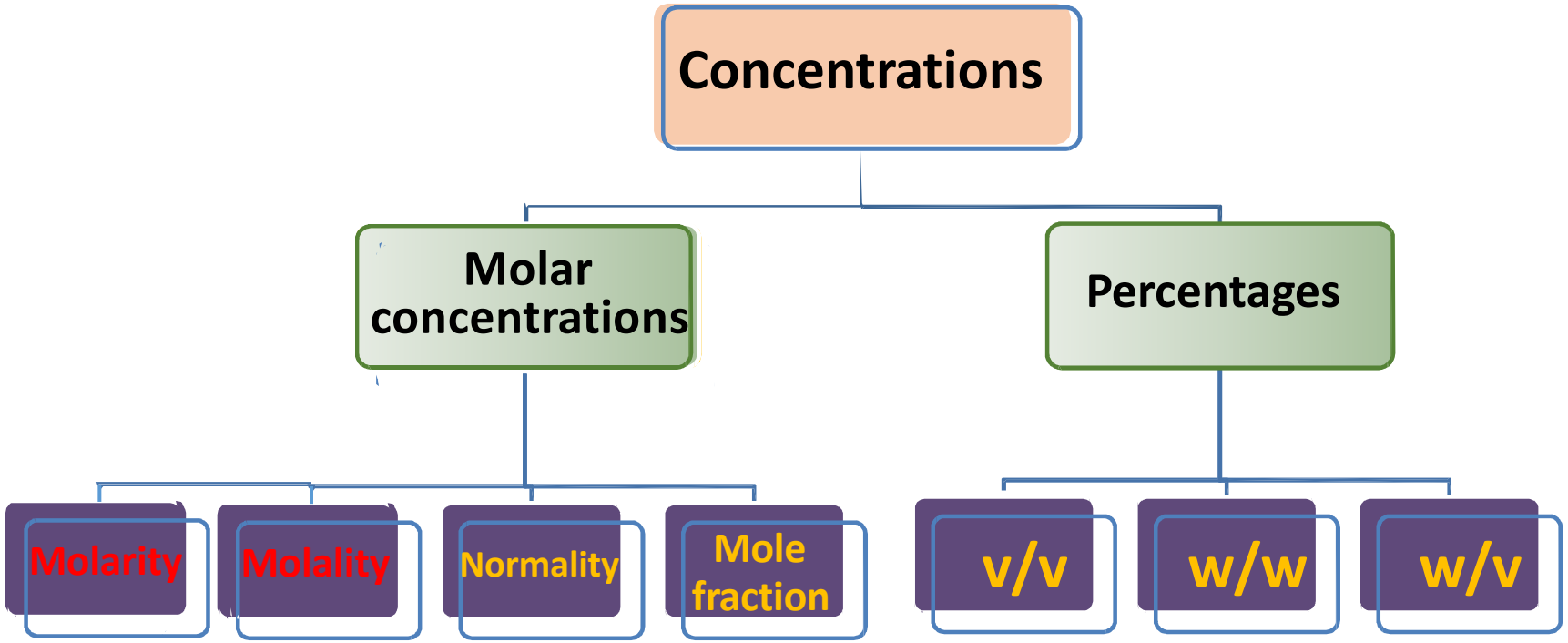
Solute: a substance that is being dissolved (smaller amount)

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.



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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 ⟶ **Products**

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

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

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

l = liquid, g = gas, s = solid

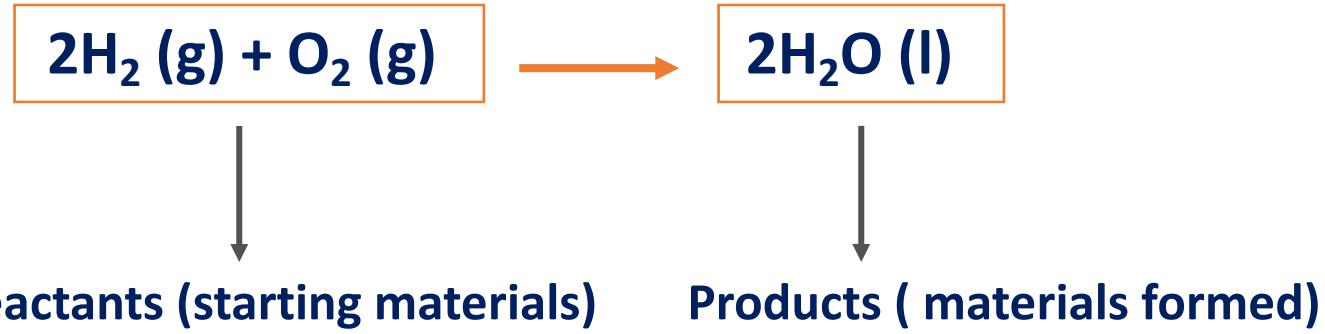
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Chemical Equations

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)



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Balancing Chemical Equations

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

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

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

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

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

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) $S_8 + O_2 \longrightarrow SO_2$

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

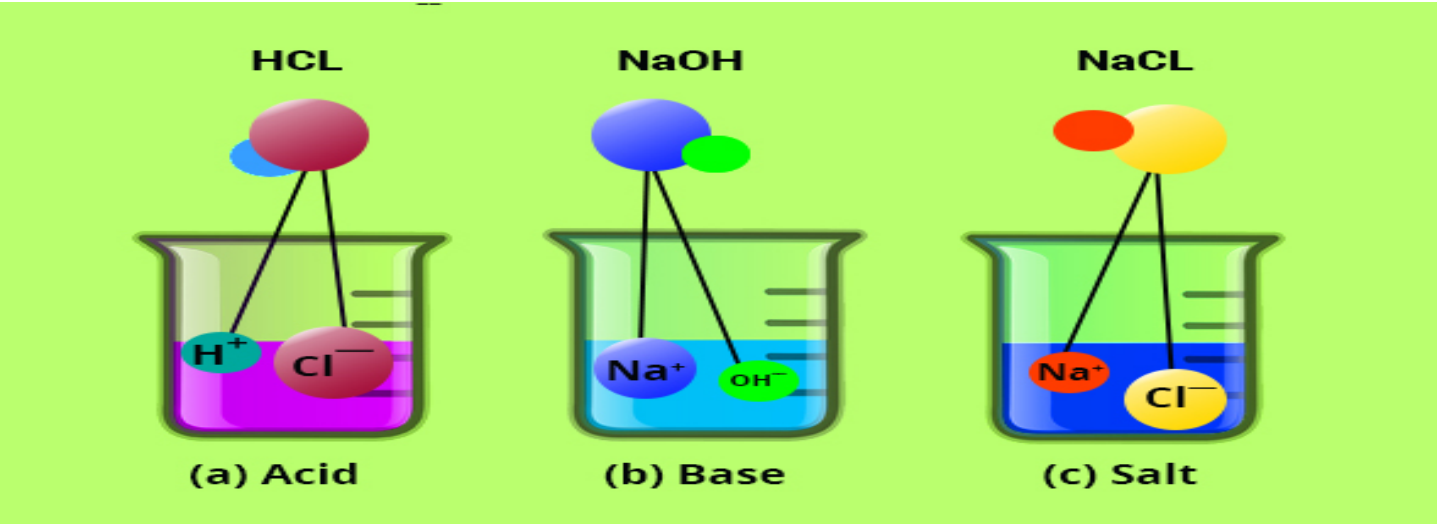
3) Precipitation Reactions

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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

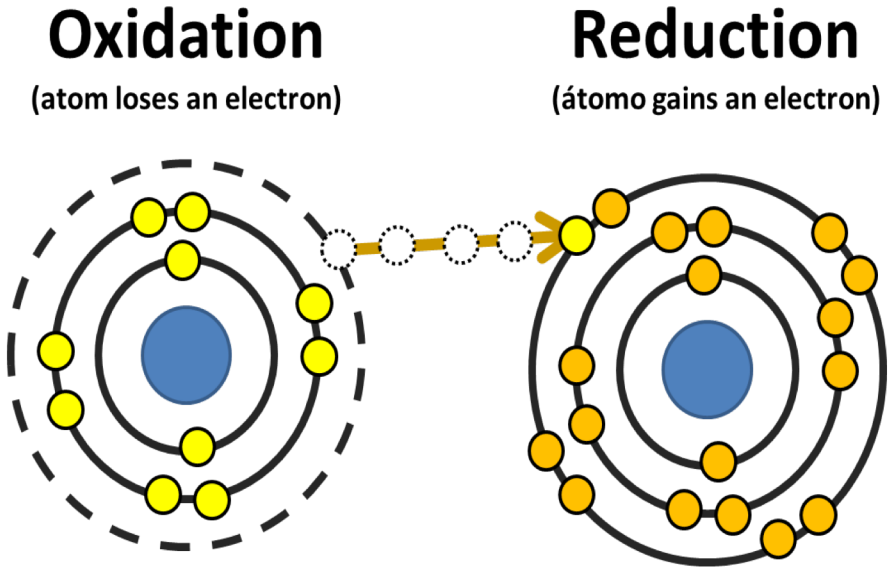
Redox reactions are electron transfer reactions

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

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

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

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

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



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) 2.5
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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Thermochemistry

Chapter

4

Chang-chapter6

COURSE NAME: CHEMISTRY 101
COURSE CODE: 28171101-4

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

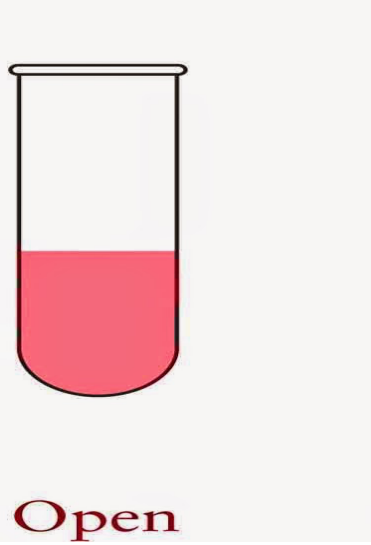
can exchange mass and energy

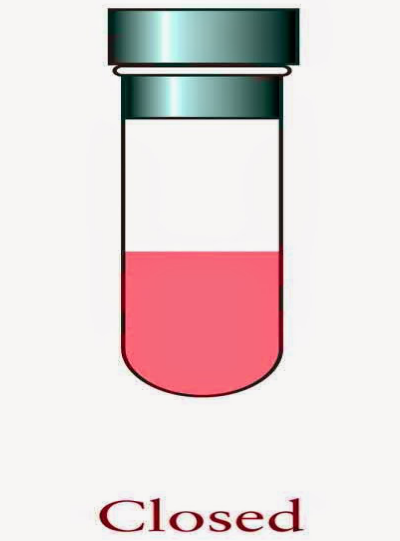
Closed system

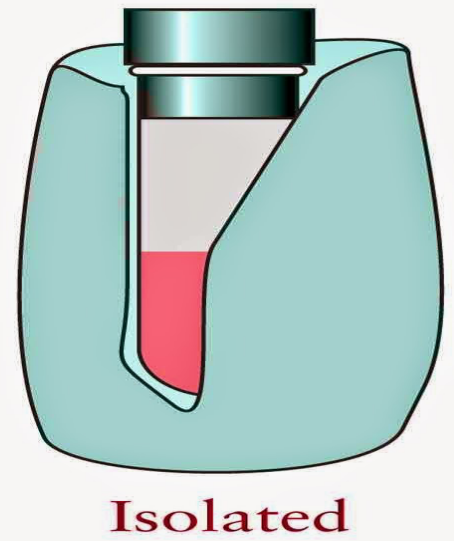
allows the transfer of energy (heat) but not mass

Isolated system

doesn't allow transfer of either mass or energy







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Examples







Thermodynamics

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

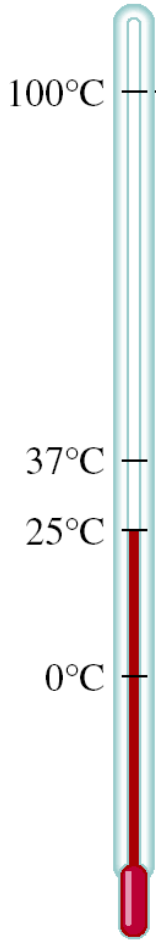
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Heat (q)

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

Temperature is a measure of the thermal energy

Temperature ≠ 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)

$$w = F \times d$$

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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

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 ^oC, 1 atm)

$$N_2\ (g) + 3H_2\ (g) \longrightarrow 2NH_3\ (g)\ \Delta H^o = -92.38\ kJ$$

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

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

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

$$H_2O\,(s) \longrightarrow H_2O\,(l) \qquad \Delta H = 6.01\ kJ/mol$$

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

$$H_2O\,(l) \longrightarrow H_2O\,(s) \qquad \Delta H = -6.01\ kJ/mol$$

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

$$2H_2O\,(s) \longrightarrow 2H_2O\,(l) \qquad \Delta H = 2 \times 6.01 = 12.0\ kJ$$

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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



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

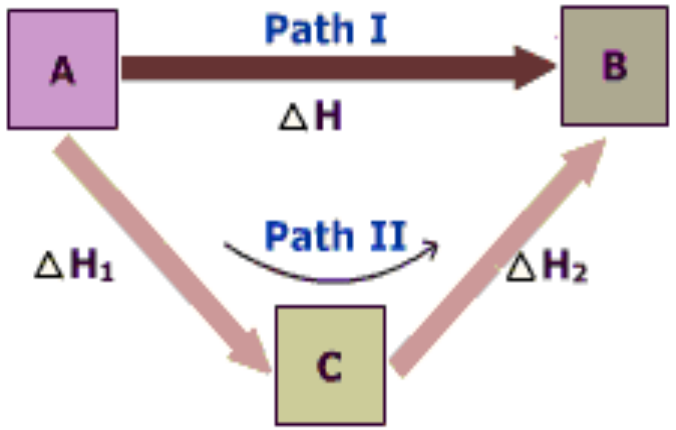
$$\Delta H^o = [(\Delta H_f^o\,(Al_2O_3)) + (2\times \Delta H_f^o\,(Fe))] - [(2\times \Delta H_f^o\,(Al)) + (\Delta H_f^o\,(Fe_2O_3))]$$

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

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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.

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

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

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.

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

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

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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- $CO\ (g) + \frac{1}{2}O_2\ (g) \longrightarrow CO_2\ (g)$ $\quad \Delta H^o = -283.0$ kJ

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

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

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

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

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.°C. Determine the energy, in calories, necessary to raise the temperature of a 55.5 g piece of aluminum from 23.0 to 48.6°C.

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.°C. Determine the energy, in calories, necessary to raise the temperature of a 55.5 g piece of aluminum from 23.0 to 48.6°C.

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:
$\Delta H^{o}f\ CH_4(g) = -74.8$ kJ/mol;
$\Delta H^{o}f\ CO_2(g) = -393.5$ kJ/mol;
$\Delta H^{o}f\ H_2O(l) = -285.5$ kJ/mol.
$CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(l)$

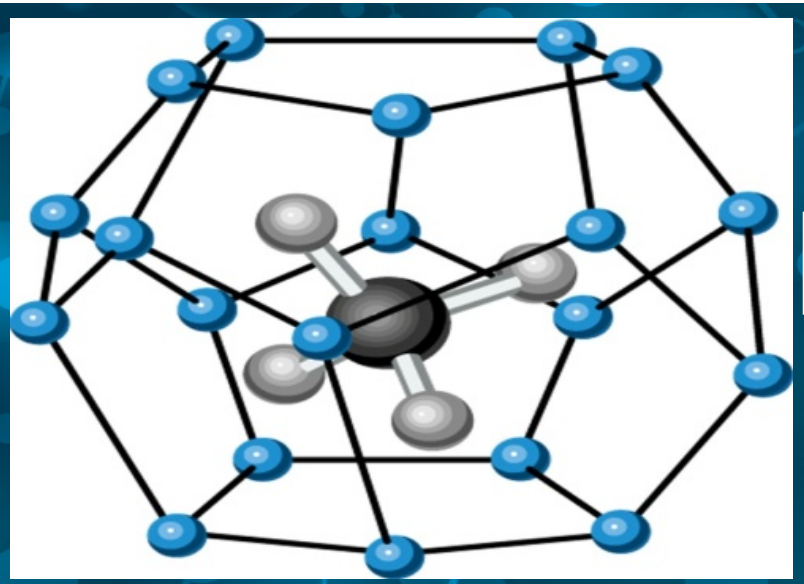
A. -604.2 kJ
B. 889.7 kJ
C. -997.7 kJ
D. -889.7 kJ
E. None of the above

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) $\Delta H^{o}_{rxn} = -1411$ kJ;
C(s) + O2(g) ----> CO2(g) $\Delta H^{o}_{f} = -393.5$ kJ;
H2(g) + ½O2(g) ----> H2O(l) $\Delta H^{o}_{f} = -285.8$ 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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Electrochemistry

Chapter

5

Chang-chapter24

COURSE NAME: CHEMISTRY 101
COURSE CODE: 28171101-4

Electrochemistry:

study of the interchange between chemical change and electrical work

Electrochemical cells:

systems utilizing a redox reaction to produce or use electrical energy

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Redox reactions: electron transfer processes

Oxidation: loss of 1 or more electrons (e^-)

Reduction: gain of 1 or more electrons (e^-)

Oxidation numbers: imaginary charges (Balancing redox reactions)

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Types of cells

Voltaic (galvanic) cells:

a spontaneous reaction generates electrical energy

Electrolytic cells:

absorb free energy from an electrical source to drive a nonspontaneous reaction

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Common Components

Electrodes:

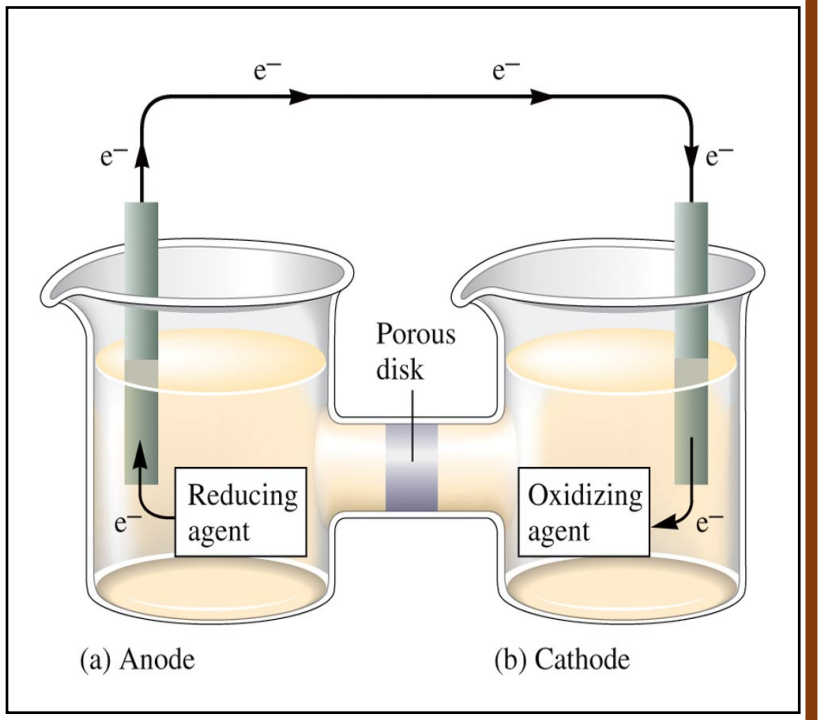
conduct electricity between cell and surroundings

Electrolyte:

mixture of ions involved in reaction or carrying charge

Salt bridge:

completes circuit (provides charge balance)



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Electrodes

Anode:

Oxidation occurs at the anode

Cathode:

Reduction occurs at the cathode

Active electrodes: participate in redox

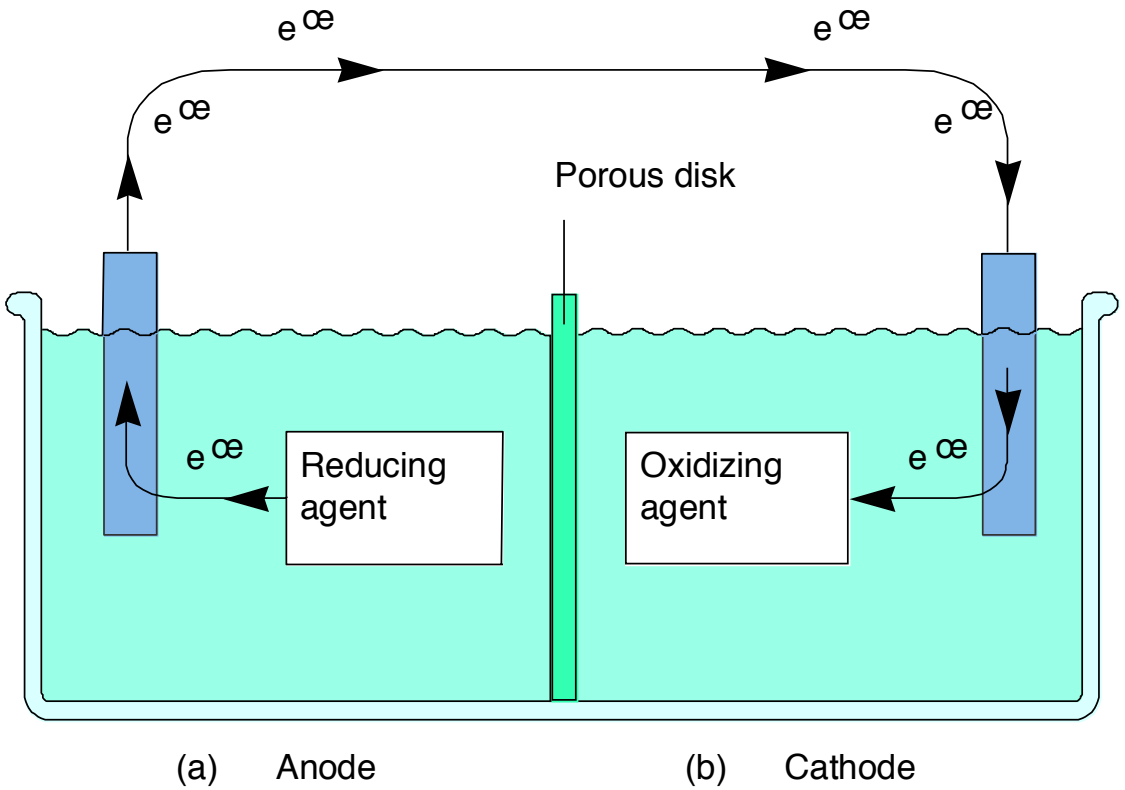
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Voltaic (Galvanic) Cells

A device in which chemical energy is changed to electrical energy.

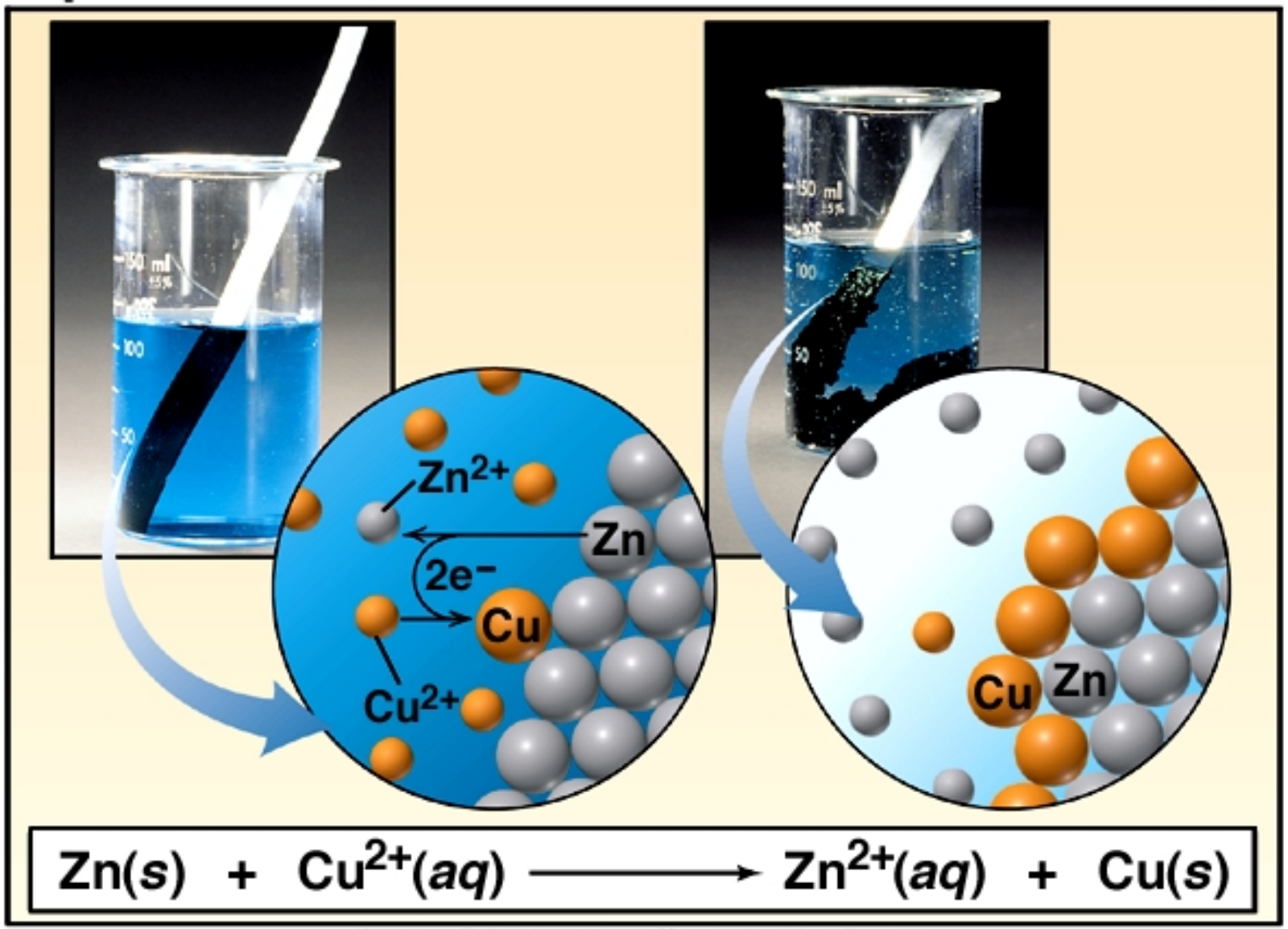
Uses a spontaneous reaction.



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Martin S. Silberberg, *Chemistry: The Molecular Nature of Matter and Change*, 2nd Edition. Copyright © The McGraw-Hill Companies, Inc. All rights reserved.

Spontaneous Reaction between Zn and Cu^{2+}



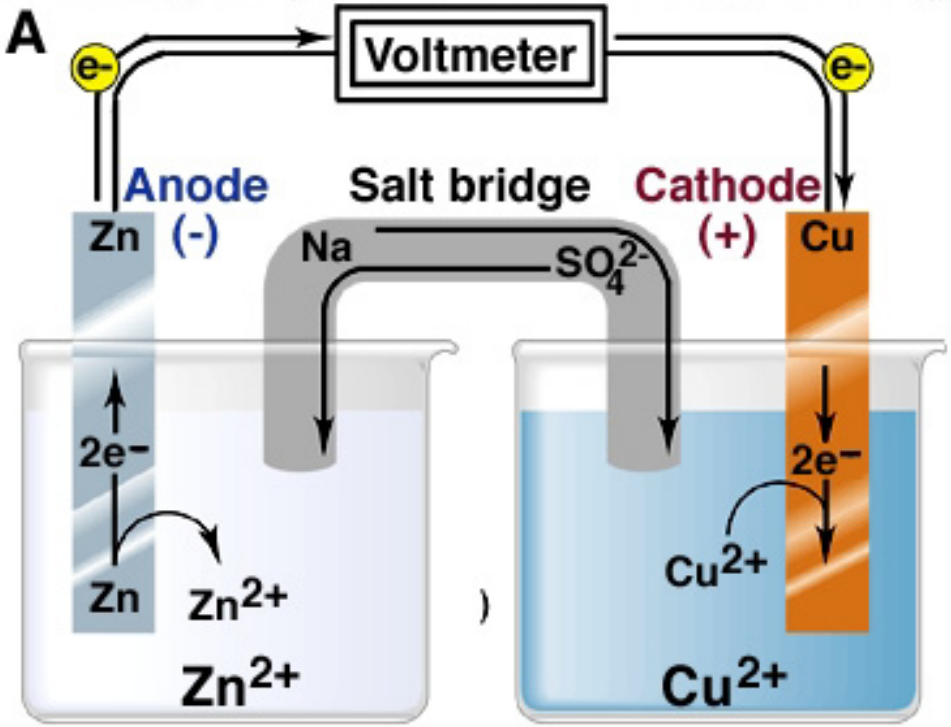
$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

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A



B



Oxidation half-reaction

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

Reduction half-reaction

$$2e^- + Cu^{2+}(aq) \longrightarrow Cu(s)$$

Overall (cell) reaction

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

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Zinc-Copper Reaction Voltaic Cell

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Cell Potential

- A galvanic cell consists of an oxidizing agent (in cathode half-cell) and a reducing agent (in anode half-cell).
- Electrons flows through a wire from the anode half-cell to the cathode half-cell.
- The driving force that allows electrons to flow is called the electromotive force (emf) or the cell potential (E_{cell}).
  - The unit of electrical potential is *volt* (V).
    - 1 V = 1 J/C of charge transferred.

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Example: $Fe^{3+}(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + Fe^{2+}(aq)$

- Half-Reactions:
  - $Fe^{3+} + e^- \rightarrow Fe^{2+}$ $\mathcal{E}° = 0.77$ V
  - $Cu^{2+} + 2e^- \rightarrow Cu$ $\mathcal{E}° = 0.34$ V
- To balance the cell reaction and calculate the cell potential, we must reverse reaction 2.
  - $Cu \rightarrow Cu^{2+} + 2e^-$ $-\mathcal{E}° = -0.34$ V
- Each Cu atom produces two electrons but each Fe^{3+} ion accepts only one electron, therefore reaction 1 must be multiplied by 2.
  - $2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$ $\mathcal{E}° = 0.77$ V

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Calculating E^0_{cell}

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

$2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$; $\mathcal{E}° = 0.77$ V (cathode)

$Cu \rightarrow Cu^{2+} + 2e^-$; $-\mathcal{E}° = -0.34$ V (anode)

- Balanced Cell Reaction:

  $Cu + 2Fe^{3+} \rightarrow Cu^{2+} + 2Fe^{2+}$

- Cell Potential: $\mathcal{E}$

$$\mathcal{E}°_{cell} = \mathcal{E}°(cathode) - \mathcal{E}°(anode)$$

$$\mathcal{E}°_{cell} = 0.77\text{ V} - 0.34\text{ V} = 0.43\text{ V}$$

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Nernst Equation

Under nonstandard conditions

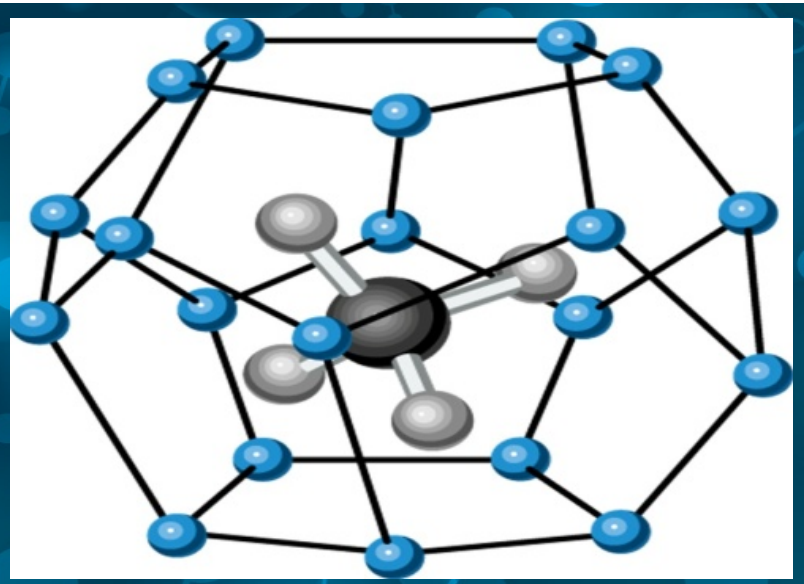
$$\Delta G = \Delta G^0 + RT\ln Q$$

$$-nFE = -nFE^0 + RT\ln Q$$

$$E_{cell} = E^0 - \frac{RT}{nF}\ln Q$$

$$E_{cell}^{298K} = E^0 - \frac{0.0592}{n}\ln Q$$

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Gases

Chapter

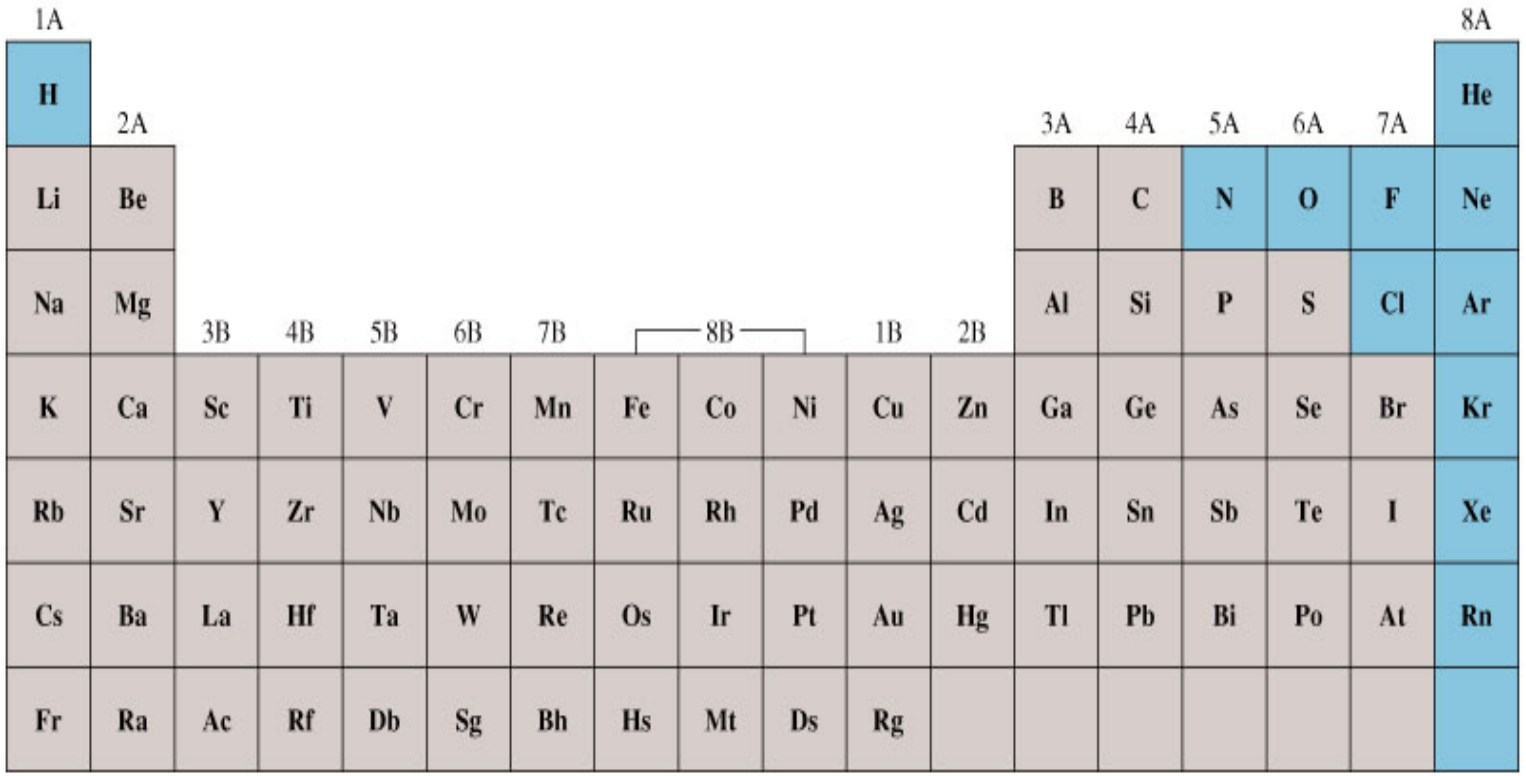
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Chang-chapter24

COURSE NAME: CHEMISTRY 101
COURSE CODE: 28171101-4

Substances that exist as gases

Elements that exist as gases at 25°C and 1 atmosphere



1A												3A	4A	5A	6A	7A	8A
H	2A																He
Li	Be											B	C	N	O	F	Ne
Na	Mg	3B	4B	5B	6B	7B		8B		1B	2B	Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

The noble gases (the Group 8A elements) are monatomic species, whereas all other gases exist as diatomic molecules. Ozone (O_3) is also gas.

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TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

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Physical Characteristics of Gases

- Gases assume the volume and shape of their containers.
- Gases are the most compressible state of matter.
- Gases will mix evenly and completely when confined to the same container.
- Gases have much lower densities than liquids and solids.

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Pressure of Gases and its Units



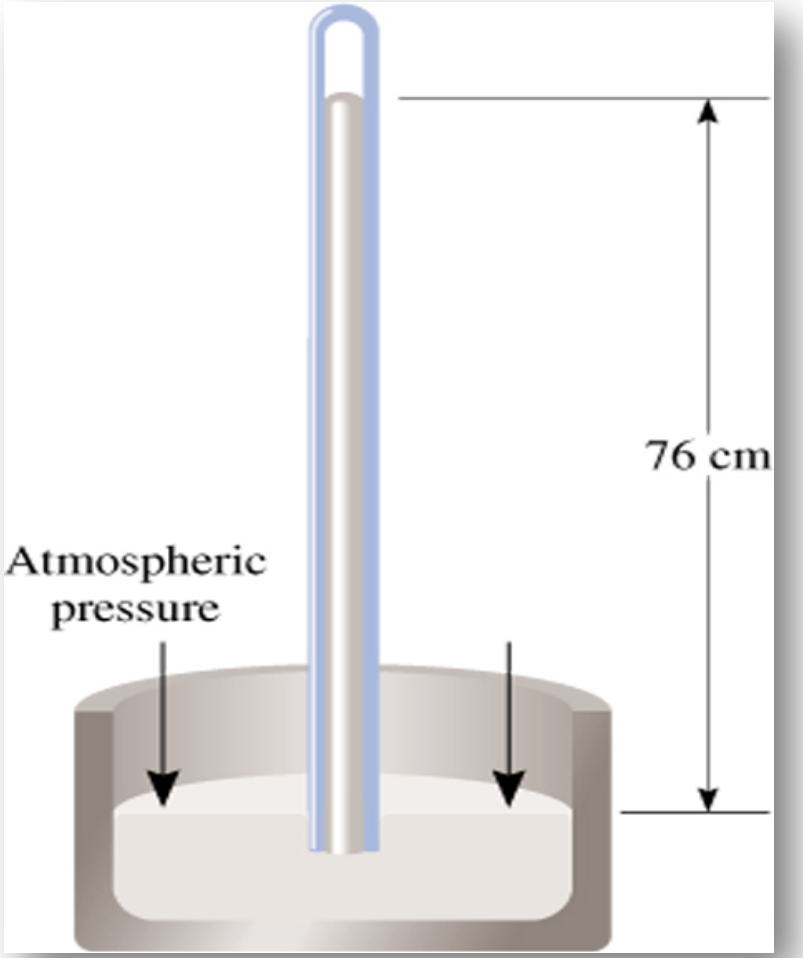
- Pressure is defined as the force applied per unit are

$$\textbf{Pressure} = \frac{\textbf{Force}}{\textbf{Area}} = \textbf{N/m}^2$$

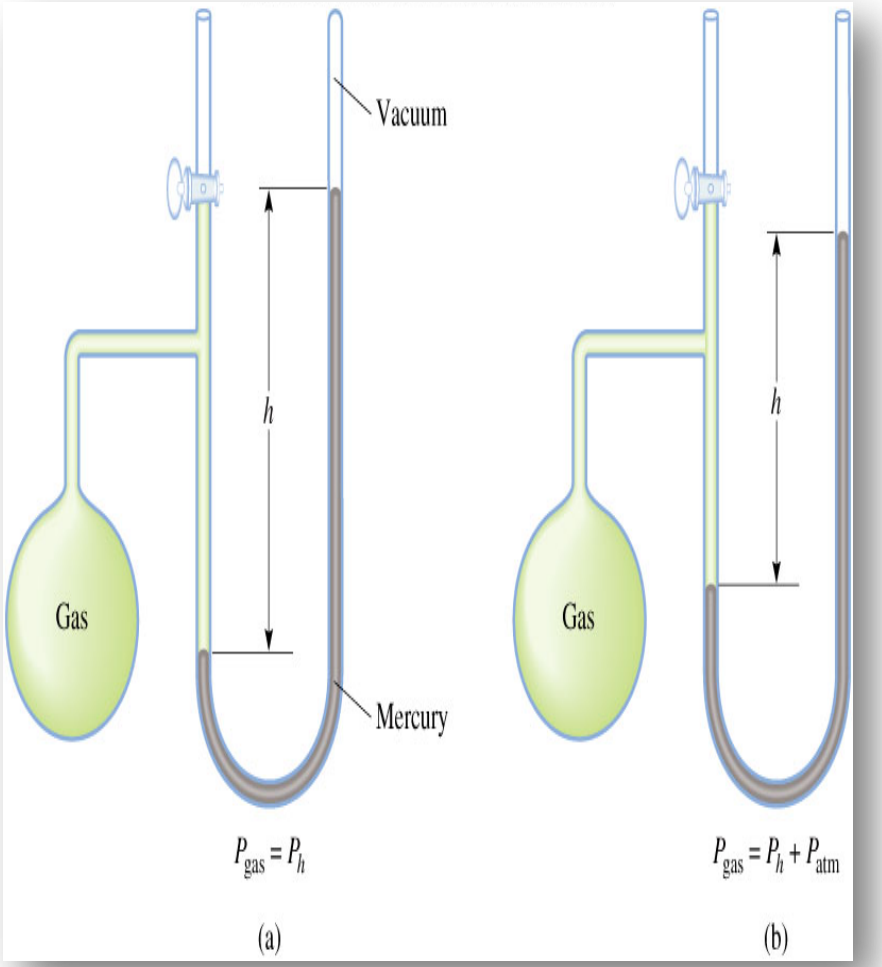
- The SI unit of pressure is Pascal (Pa) define as one Newton per square meter **(1Pa = N/m^2)**

- **Standard atmospheric pressure (1 atm)** is equal to the pressure that supports a column of mercury exactly 760 mm (or 76 cm) high at 0^oC at sea level.

- Measured using a Barometer! - A device that can weigh the atmosphere above us!



Barometer



Manometer

a Barometer! - A device that can weigh the atmosphere above us!

A simple manometer, a device for measuring the pressure of a gas in a container

Common Units of Pressure

Unit	Atmospheric Pressure	Scientific Field Used
Pascal (Pa)	1.01325×10^5 Pa	SI unit; physics, chemistry
kilopascal (kPa)	1.01325×10^2 kPa	
Bar	1.01325 bar	Meteorology, chemistry
Atmosphere (atm)	1 atm	Chemistry
Millimeters of mercury (mmHg)	760 mmHg	Chemistry, medicine biology
Torr	760 torr	Chemistry
Pounds per square inch (psi or lbs./in^2)	14.7 lb/in^2	Engineering

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Common Units of Pressure

Remember the conversions for pressure:

760 mm Hg = 760 torr

1 atm = 760 mm Hg

760 mm Hg = 101,325 Pa (**1.01325×10^5 Pa**)

Convert 2.3 atm to torr:

Example ……

$$2.0 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 1520 \text{ torr}$$

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Example 5.1

DCM, Worked Examples

The pressure outside a jet plane flying at high altitude falls considerably below standard atmospheric pressure. Therefore, the air inside the cabin must be pressurized to protect the passengers. What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

Strategy Because 1 atm = 760 mmHg, the following conversion factor is needed to obtain the pressure in atmospheres

$$\frac{1 \text{ atm}}{760 \text{ mmHg}}$$

Solution The pressure in the cabin is given by

$$\text{pressure} = 688 \cancel{\text{mmHg}} \times \frac{1 \text{ atm}}{760 \cancel{\text{mmHg}}}$$

$$= 0.905 \text{ atm}$$

Practice Exercise Convert 749 mmHg to atmospheres.

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Example 5.2

DCM, Worked Examples

The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

Strategy Here we are asked to convert mmHg to kPa. Because

$$1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa} = 760 \text{ mmHg}$$

the conversion factor we need is

$$\frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}}$$

Solution The pressure in kPa is

$$\begin{aligned} \text{pressure} &= 732 \cancel{\text{mmHg}} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \cancel{\text{mmHg}}} \\ &= 9.76 \times 10^4 \text{ Pa} \\ &= 97.6 \text{ kPa} \end{aligned}$$

Practice Exercise Convert 295 mmHg to kilopascals.

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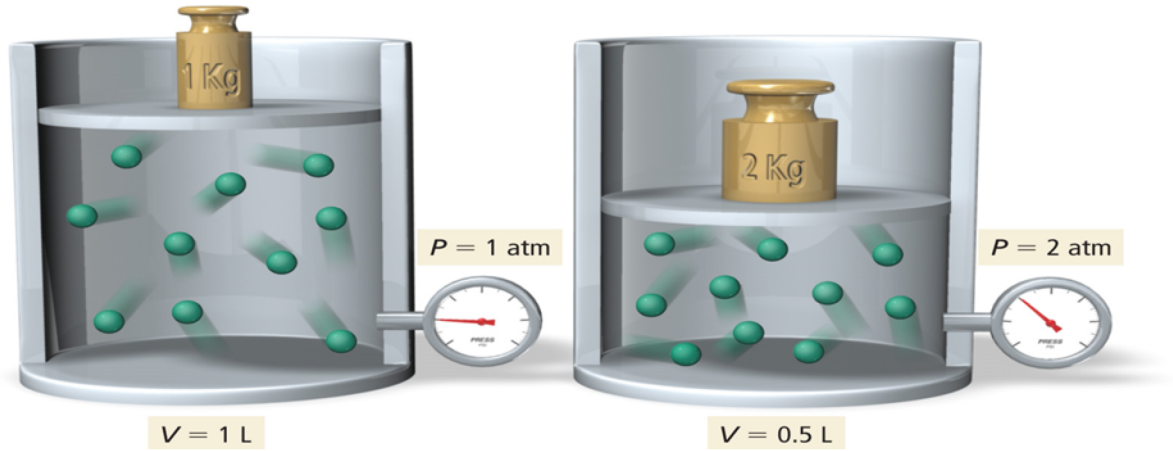
The gas laws

- **Boyle's Law , V - P relationship**
- **Charles Law , V - T- relationship**
- **Avogadro's Law ,V and Amount**

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Boyle's Law : P - V relationship





- **Pressure is inversely proportional to Volume**

- $P_1V_1 = k$ $\qquad\qquad$ $P_2V_2 = k'$

$k = k'$ Then : $P_1V_1 = P_2V_2$

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Example A cylinder with a movable piston has a volume of 7.25 L at 4.52 atm. What is the volume at 1.21 atm?

Given: V_1 =7.25 L, P_1 = 4.52 atm, P_2 = 1.21 atm

Find: V_2, L

Concept Plan: $V_1, P_1, P_2 \Rightarrow V_2$

Relationships: $P_1 \cdot V_1 = P_2 \cdot V_2$

$$V_2 = \frac{P_1 \bullet V_1}{P_2}$$

Solution:

$$V_2 = \frac{P_1 \bullet V_1}{P_2}$$

$$= \frac{(4.52 \cancel{\text{atm}}) \bullet (7.25\ \text{L})}{(1.21 \cancel{\text{atm}})} = 27.1\ \text{L}$$

Check: **since P and V are inversely proportional, when the pressure decreases ~4x, the volume should increase ~4x, and it does**

A balloon is put in a bell jar and the pressure is reduced from 782 torr to 0.500 atm. If the volume of the balloon is now 2780 mL, what was it originally?

Given:	V_2 =2780 mL, P_1 = 782 torr, P_2 = 0.500 atm
Find:	V_1, mL
Concept Plan:	$V_1, P_1, P_2 \Rightarrow V_2$ $V_1 = \frac{P_2 \bullet V_2}{P_1}$
Relationships:	$P_1 \cdot V_1 = P_2 \cdot V_2$, 1 atm = 760 torr (exactly)
Solution:	$782 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.03 \text{ atm}$ $V_1 = \frac{P_2 \bullet V_2}{P_1}$ $= \frac{(0.500 \text{ atm}) \bullet (2780 \text{ mL})}{(1.03 \text{ atm})} = 1350 \text{ mL}$
Check:	**since P and V are inversely proportional, when the pressure decreases ~2x, the volume should increase ~2x, and it does**

Charles' Law



- volume is directly proportional to temperature
  - constant P and amount of gas

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- as T increases, V also increases
- Kelvin T = Celsius T + 273

- V = constant x T
  - if T measured in Kelvin

A gas has a volume of 2.57 L at 0.00°C. What was the temperature (in both K and °C) at 2.80 L?

Given: $V_2 = 2.57$ L, $t_2 = 0.00$°C, $V_1 = 2.80$ L

Find: t_1, K and °C

Concept Plan:

$V_1, V_2, T_2 \Rightarrow T_1$

Relationships: $T(K) = t(°C) + 273.15$, $\frac{V_1}{T_1} = \frac{V_2}{T_2} \rightarrow T_1 = T_2 \bullet \frac{V_1}{V_2}$

Solution:

$T_2 = 0.00 + 273.15$

$T_2 = 273.15\ K$

$$T_1 = \frac{T_2 \bullet V_1}{V_2}$$

$$= \frac{(273.15\ K) \bullet (2.80\ \cancel{L})}{(2.57\ \cancel{L})} = 297.6\ K$$

$t_1 = T_1 - 273.15$

$t_1 = 297.6 - 273.15$

$t_1 = 24\ °C$

Check: since T and V are directly proportional, when the volume decreases, the temperature should decrease, and it does

EXAMPLE

- **Gas occupy 6L at 37^0C what will be its volume when its temperature decreased to the half?**

$$V_1=6L\ ,\ T_1=37^0C$$

$$V_2=???\ ,\ T_2=½\ T_1$$

$$\frac{V_1}{T_1}=\frac{V_2}{T_2} \qquad V_2=\frac{V_1}{T_1}(1/2T_1)$$

$$V_2=½V_1$$

$$V_2=½\ (6)=3L$$

Avogadro's Law



- volume directly proportional to the number of gas molecules
  - **V** = constant x **n**
  - constant P and T
  - more gas molecules = larger volume
- count number of gas molecules by **moles**
- equal volumes of gases contain equal numbers of molecules
  - ***the gas doesn't matter***

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

A 0.225 mol sample of He has a volume of 4.65 L. How many moles must be added to give 6.48 L?

Given:	V_1 =4.65 L, , n_1 = 0.225 mol, V_2 = 6.48 L
Find:	n_2, and added moles
Concept Plan:	$V_1, V_2, n_1 \Rightarrow n_2$
Relationships:	mol added = $n_2 - n_1$, $\frac{V_1}{n_1} = \frac{V_2}{n_2}$
Solution:	$n_2 = \frac{n_1 \bullet V_2}{V_1} = \frac{(0.225\ \text{mol}) \bullet (6.48\ \cancel{\text{L}})}{(4.65\ \cancel{\text{L}})} = 0.314\ \text{mol}$ $\text{moles added} = 0.314 - 0.225$ $\text{moles added} = 0.089\ \text{mol}$
Check:	since n and V are directly proportional, when the volume increases, the moles should increase, and it does

Ideal Gas Law

- By combing the gas laws we can write a general equation
- **R** is called the **gas constant**
- the value of **R** depends on the units of P and V
  - we will use 0.08206 $\frac{\text{atm} \bullet \text{L}}{\text{mol} \bullet \text{K}}$ and convert P to atm and V to L
  - the other gas laws are found in the ideal gas law if two variables are kept constant
  - allows us to find one of the variables if we know the other 3

$$PV = nRT$$

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Example 5.3

DCM, Worked Examples

Sulfur hexafluoride (SF_6) is a colorless, odorless, very unreactive gas. Calculate the pressure (in atm) exerted by 1.82 moles of the gas in a steel vessel of volume 5.43 L at 69.5°C.

Strategy The problem gives the amount of the gas and its volume and temperature. Is the gas undergoing a change in any of its properties? What equation should we use to solve for the pressure? What temperature unit should we use?

Solution Because no changes in gas properties occur, we can use the ideal gas equation to calculate the pressure. Rearranging Equation (5.8), we write

$$P = \frac{nRT}{V}$$

$$= \frac{(1.82\ \text{mol})(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})(69.5 + 273)\ \text{K}}{5.43\ \text{L}}$$

$$= 9.42\ \text{atm}$$

Practice Exercise Calculate the volume (in liters) occupied by 2.12 moles of nitric oxide (NO) at 6.54 atm and 76°C.

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Standard Conditions

- since the volume of a gas varies with pressure and temperature, chemists have agreed on a set of conditions to report our measurements so that comparison is easy – we call these **standard conditions**
  - STP
- standard pressure = 1 atm
- standard temperature = 273 K = 0°C
- One mole of a gas occupy **22.41** L at STP

Example

- **Calculate the (volume in liters occupied by 7.40g of NH_3 at STP**
- **Solution 2**
- **$n_{NH3} = 7.4 / 17 = 0.435$ mol**
- **V = nRT/P**
- **V = 0.435 (0.0821) 273/1 = 9.74 L**
- **Solution 2**
- **1 mole occupy 22.4 L at STP**
- **0.435 mole x >>>>>>>>>>> V = 0.435 X 22.4=9.74 L**

EXAMPLE

- What is the volume of 2g O_2 gas at STP

$$PV = nRT$$

$$V = nRT/P$$

$$V = 2 \times 0.0821 \times 273/32 \times 1$$

$$V = 1.4\ L$$

- What is the volume of 2g O_2 gas at 4 atm and 35^0C

$$PV = nRT$$

$$V = nRT/P$$

$$T = 35 + 273 = 308\ K$$

$$V = 2 \times 0.0821 \times 308/\ 32 \times 4$$

$$V = 0.395\ L$$

Combined Gas Law

- When we introduced Boyle's, Charles's, and Gay-Lussac's laws, we assumed that one of the variables remained constant.
- Experimentally, all three (temperature, pressure, and volume) usually change.
- By combining all three laws, we obtain the ***combined gas law***:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

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Example 5.5

DCM, Worked Examples

An inflated helium balloon with a volume of 0.55 L at sea level (1.0 atm) is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm. Assuming that the temperature remains constant, what is the final volume of the balloon?

Strategy The amount of gas inside the balloon and its temperature remain constant, but both the pressure and the volume change. What gas law do you need?

Solution We start with Equation (5.9)

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$ and $T_1 = T_2$,

$$P_1V_1 = P_2V_2$$

which is Boyle's law [see Equation (5.2)]. The given information is tabulated:

Initial Conditions	Final Conditions
$P_1 = 1.0$ atm	$P_2 = 0.40$ atm
$V_1 = 0.55$ L	$V_2 = ?$

Therefore,

$$V_2 = V_1 \times \frac{P_1}{P_2}$$

$$= 0.55 \text{ L} \times \frac{1.0 \text{ atm}}{0.40 \text{ atm}}$$

$$= 1.4 \text{ L}$$

Check When pressure applied on the balloon is reduced (at constant temperature), the helium gas expands and the balloon's volume increases. The final volume is greater than the initial volume, so the answer is reasonable.

Practice Exercise A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. Calculate the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL.

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Gas Density

$$\text{mass} \times \frac{1\,\text{mol}}{\text{molar mass}} = \text{moles} \quad \therefore \quad \text{moles} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{density} = \frac{\text{mass in grams}}{\text{volume in liters}}$$

$$P \times V = n \times R \times T$$

$$P \times V = \frac{\text{mass}}{\text{molar mass}} \times R \times T$$

$$\frac{\text{mass}}{V} = \text{density} = \frac{P \times (\text{molar mass})}{R \times T}$$

- **density is directly proportional to molar mass**

Example 5.8

DCM, Worked Examples

Calculate the density of carbon dioxide (CO_2) in grams per liter (g/L) at 0.990 atm and 55°C.

Strategy We need Equation (5.11) to calculate gas density. Is sufficient information provided in the problem? What temperature unit should be used?

Solution To use Equation (5.11), we convert temperature to kelvins ($T = 273 + 55 = 328$ K) and use 44.01 g for the molar mass of CO_2:

$$d = \frac{PM}{RT}$$

$$= \frac{(0.990 \text{ atm})(44.01 \text{ g/mol})}{(0.0821 \text{ L } \cancel{\text{atm}}/\text{K } \cancel{\text{mol}})(328 \text{ K})} = 1.62 \text{ g/L}$$

Alternatively, we can solve for the density by writing

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Assuming that we have 1 mole of CO_2, the mass is 44.01 g. The volume of the gas can be obtained from the ideal gas equation

$$V = \frac{nRT}{P}$$

$$= \frac{(1 \text{ mol})(0.0821 \text{ L } \cancel{\text{atm}}/\text{K } \cancel{\text{mol}})(328 \text{ K})}{0.990 \text{ atm}}$$

$$= 27.2 \text{ L}$$

Therefore, the density of CO_2 is given by

$$d = \frac{44.01 \text{ g}}{27.2 \text{ L}} = 1.62 \text{ g/L}$$

Comment In units of grams per milliliter, the gas density is 1.62×10^{-3} g/mL, which is a very small number. In comparison, the density of water is 1.0 g/mL and that of gold is 19.3 g/cm^3.

Practice Exercise What is the density (in g/L) of uranium hexafluoride (UF_6) at 779 mmHg and 62°C?

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Molar Mass of a Gas

- **From density calculations**
- $\mathcal{M}$ = **dRT/ P**

- **From number of moles calculations**
- **n = mass /** $\mathcal{M}$
- **PV= nRT**
- **PV= (mass /** $\mathcal{M}$ **) RT**

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Example 5.9

DCM, Worked Examples

A chemist has synthesized a greenish-yellow gaseous compound of chlorine and oxygen and finds that its density is 7.71 g/L at 36°C and 2.88 atm. Calculate the molar mass of the compound and determine its molecular formula.

Strategy Because Equations (5.11) and (5.12) are rearrangements of each other, we can calculate the molar mass of a gas if we know its density, temperature, and pressure. The molecular formula of the compound must be consistent with its molar mass. What temperature unit should we use?

Solution From Equation (5.12)

$$\mathcal{M} = \frac{dRT}{P}$$
$$= \frac{(7.71\ \text{g/L})(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})(36 + 273)\ \text{K}}{2.88\ \text{atm}}$$
$$= 67.9\ \text{g/mol}$$

Alternatively, we can solve for the molar mass by writing

$$\text{molar mass of compound} = \frac{\text{mass of compound}}{\text{moles of compound}}$$

From the given density we know there are 7.71 g of the gas in 1 L. The number of moles of the gas in this volume can be obtained from the ideal gas equation

$$n = \frac{PV}{RT}$$
$$= \frac{(2.88\ \text{atm})(1.00\ \text{L})}{(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})(309\ \text{K})}$$
$$= 0.1135\ \text{mol}$$

Therefore, the molar mass is given by

$$\mathcal{M} = \frac{7.71\ \text{g}}{0.1135\ \text{mol}} = 67.9\ \text{g/mol}$$

We can determine the molecular formula of the compound by trial and error, using only the knowledge of the molar masses of chlorine (35.45 g) and oxygen (16.00 g). We know that a compound containing one Cl atom and one O atom would have a molar mass of 51.45 g, which is too low, while the molar mass of a compound made up of two Cl atoms and one O atom is 86.90 g, which is too high. Thus, the compound must contain one Cl atom and two O atoms and have the formula ClO_2, which has a molar mass of 67.45 g.

Practice Exercise The density of a gaseous organic compound is 3.38 g/L at 40°C and 1.97 atm. What is its molar mass?

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Example

- A chemist synthesized a greensh-yellow gaseous compound of chlorine and oxygen and find that its denity is 7.7g/L at 36°C and 2.88 atm. Calculate the molar mass and determine its molecular formula.
- **Molar mass = dRT/ P**
- $\mathcal{M}$ = 7.7g/L ×0.0821×(36+273)/2.88 = 67.9 g/mol
- Mass of empirical formula (ClO)= 35.45+16= 51.45
- Ratio = **Molar mass** / Mass of empirical formula = 67.9/51.45= 1.3
- molecular formula. ClO_2

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Example 5.10

DCM, Worked Examples

Chemical analysis of a gaseous compound showed that it contained 33.0 percent silicon (Si) and 67.0 percent fluorine (F) by mass. At 35°C, 0.210 L of the compound exerted a pressure of 1.70 atm. If the mass of 0.210 L of the compound was 2.38 g, calculate the molecular formula of the compound.

Strategy This problem can be divided into two parts. First, it asks for the empirical formula of the compound from the percent by mass of Si and F. Second, the information provided enables us to calculate the molar mass of the compound and hence determine its molecular formula. What is the relationship between empirical molar mass and molar mass calculated from the molecular formula?

Solution We follow the procedure in Example 3.9 (p. 88) to calculate the empirical formula by assuming that we have 100 g of the compound, so the percentages are converted to grams. The number of moles of Si and F are given by

$$n_{Si} = 33.0 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} = 1.17 \text{ mol Si}$$

$$n_{F} = 67.0 \text{ g F} \times \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 3.53 \text{ mol F}$$

Therefore, the empirical formula is $Si_{1.17}F_{3.53}$, or, dividing by the smaller subscript (1.17), we obtain SiF_3.

To calculate the molar mass of the compound, we need first to calculate the number of moles contained in 2.38 g of the compound. From the ideal gas equation

$$n = \frac{PV}{RT}$$

$$= \frac{(1.70 \text{ atm})(0.210 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(308 \text{ K})} = 0.0141 \text{ mol}$$

Because there are 2.38 g in 0.0141 mole of the compound, the mass in 1 mole, or the molar mass, is given by

$$\mathcal{M} = \frac{2.38 \text{ g}}{0.0141 \text{ mol}} = 169 \text{ g/mol}$$

The molar mass of the empirical formula SiF_3 is 85.09 g. Recall that the ratio (molar mass/empirical molar mass) is always an integer (169/85.09 ≈ 2). Therefore, the molecular formula of the compound must be $(SiF_3)_2$ or Si_2F_6.

Practice Exercise A gaseous compound is 78.14 percent boron and 21.86 percent hydrogen. At 27°C, 74.3 mL of the gas exerted a pressure of 1.12 atm. If the mass of the gas was 0.0934 g, what is its molecular formula?

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Gas stoichiometry

- **Example**
- Calculate the volume of O_2(in L) required for the complete combustion of 7.64 L of (C_2H_2) measured at the same T & P

$$2\ C_2H_{2\ (g)} + 5O_{2\ (g)} \rightarrow 4CO_{2\ (g)} + 2H_2O_{\ (l)}$$

➢From Avogadro low v= Rn

➢Volume of $O_2 = 7.64\ L \times 5L\ O_2 / 2L\ C_2H_2 = 19.1\ L$

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Example

$$2NaN_3\ (S) \rightarrow 2\ Na\ (s) + 3N_2\ (g)$$

Calculate the volume of N_2 generate at 80°C and 823 mmHg by the decomposition of 60 g of NaN_3

- **n of N_2 = (60/65.02) × 3/2= 1.38**

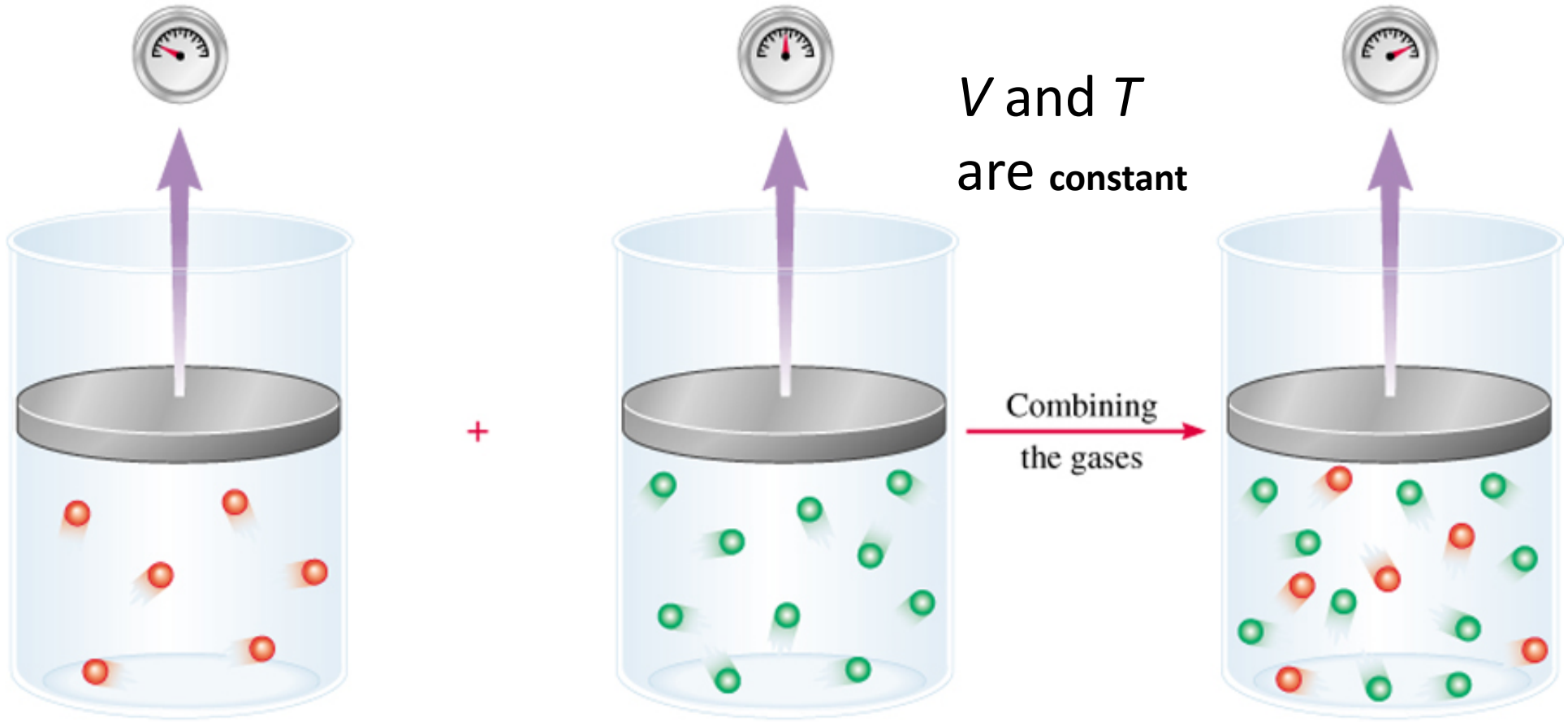
- **PV=nRT → V=nRT/P**

- **V= 1.38 × 0.0821 ×(80+273)/ (823/760)**

  = 36.9 L

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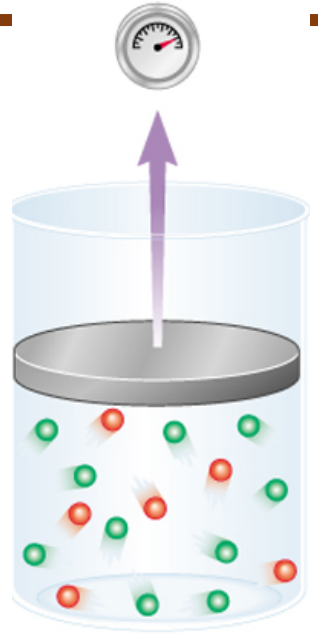
Dalton's Law of Partial Pressures



P_A P_B $P_{total} = P_A + P_B$

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Consider a case in which two gases, A and B, are in a container of volume V.



$$P_A = \frac{n_A RT}{V}$$ n_A is the number of moles of A

$$P_B = \frac{n_B RT}{V}$$ n_B is the number of moles of B

$$P_T = P_A + P_B$$

$$X_A = \frac{n_A}{n_A + n_B}$$

$$X_B = \frac{n_B}{n_A + n_B}$$

$$P_A = X_A P_T$$

$$P_B = X_B P_T$$

$$P_i = X_i P_T$$

mole fraction (X_i) = $\frac{n_i}{n_T}$

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Example 5.14

DCM, Worked Examples

A mixture of gases contains 4.46 moles of neon (Ne), 0.74 mole of argon (Ar), and 2.15 moles of xenon (Xe). Calculate the partial pressures of the gases if the total pressure is 2.00 atm at a certain temperature.

Strategy What is the relationship between the partial pressure of a gas and the total gas pressure? How do we calculate the mole fraction of a gas?

Solution According to Equation (5.14), the partial pressure of Ne (P_{Ne}) is equal to the product of its mole fraction (X_{Ne}) and the total pressure (P_T)

need to find

$$P_{Ne} = X_{Ne}P_T$$

want to calculate — given

Using Equation (5.13), we calculate the mole fraction of Ne as follows:

$$X_{Ne} = \frac{n_{Ne}}{n_{Ne} + n_{Ar} + n_{Xe}} = \frac{4.46\ \text{mol}}{4.46\ \text{mol} + 0.74\ \text{mol} + 2.15\ \text{mol}} = 0.607$$

Therefore

$$\begin{aligned} P_{Ne} &= X_{Ne}P_T \\ &= 0.607 \times 2.00\ \text{atm} \\ &= 1.21\ \text{atm} \end{aligned}$$

Similarly,

$$\begin{aligned} P_{Ar} &= X_{Ar}P_T \\ &= 0.10 \times 2.00\ \text{atm} \\ &= 0.20\ \text{atm} \end{aligned}$$

and

$$\begin{aligned} P_{Xe} &= X_{Xe}P_T \\ &= 0.293 \times 2.00\ \text{atm} \\ &= 0.586\ \text{atm} \end{aligned}$$

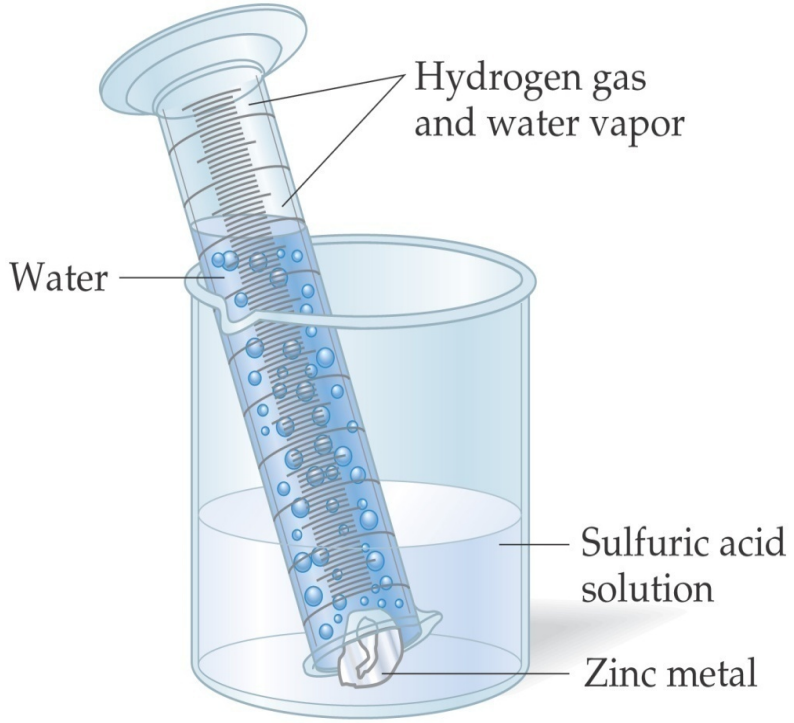
Check Make sure that the sum of the partial pressures is equal to the given total pressure; that is, (1.21 + 0.20 + 0.586) atm = 2.00 atm.

Practice Exercise A sample of natural gas contains 8.24 moles of methane (CH_4), 0.421 mole of ethane (C_2H_6), and 0.116 mole of propane (C_3H_8). If the total pressure of the gases is 1.37 atm, what are the partial pressures of the gases?

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Collecting a Gas Over Water

- We can measure the volume of a gas by displacement.
- By collecting the gas in a graduated cylinder, we can measure the amount of gas produced.
- The gas collected is referred to as “wet” gas since it also contains water vapor.



$$P_T = P_{O_2} + P_{H_2O}$$

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Example 5.15

DCM, Worked Examples

Oxygen gas generated by the decomposition of potassium chlorate is collected as shown in Figure 5.15. The volume of oxygen collected at 24°C and atmospheric pressure of 762 mmHg is 128 mL. Calculate the mass (in grams) of oxygen gas obtained. The pressure of the water vapor at 24°C is 22.4 mmHg.

Strategy To solve for the mass of O_2 generated, we must first calculate the partial pressure of O_2 in the mixture. What gas law do we need? How do we convert pressure of O_2 gas to mass of O_2 in grams?

Solution From Dalton's law of partial pressures we know that

$$P_T = P_{O_2} + P_{H_2O}$$

Therefore

$$\begin{aligned} P_{O_2} &= P_T - P_{H_2O} \\ &= 762 \text{ mmHg} - 22.4 \text{ mmHg} \\ &= 740 \text{ mmHg} \end{aligned}$$

From the ideal gas equation we write

$$PV = nRT = \frac{m}{\mathcal{M}}RT$$

where m and $\mathcal{M}$ are the mass of O_2 collected and the molar mass of O_2, respectively. Rearranging the equation we obtain

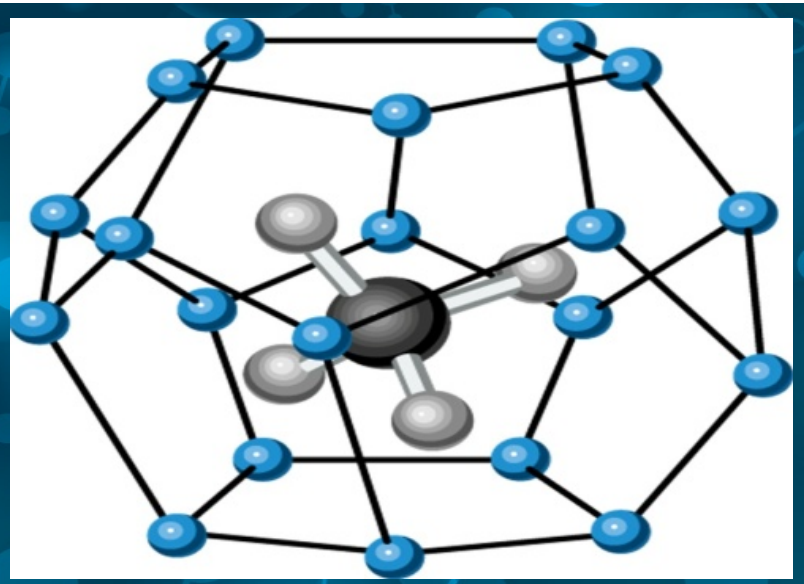
$$m = \frac{PV\mathcal{M}}{RT} = \frac{(740/760)\text{atm}(0.128 \text{ L})(32.00 \text{ g/mol})}{(0.0821 \text{ L} \cancel{\text{atm}}/\text{K} \cancel{\text{mol}})(273 + 24) \text{ K}}$$

$$= 0.164 \text{ g}$$

Check The density of the oxygen gas is (0.164 g/0.128 L), or 1.28 g/L, which is a reasonable value for gases under atmospheric conditions (see Example 5.8).

Practice Exercise Hydrogen gas generated when calcium metal reacts with water is collected as shown in Figure 5.15. The volume of gas collected at 30°C and pressure of 988 mmHg is 641 mL. What is the mass (in grams) of the hydrogen gas obtained? The pressure of water vapor at 30°C is 31.82 mmHg.

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States of Matter: Liquids and Solids

Chapter

7

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COURSE CODE: 28171101-4

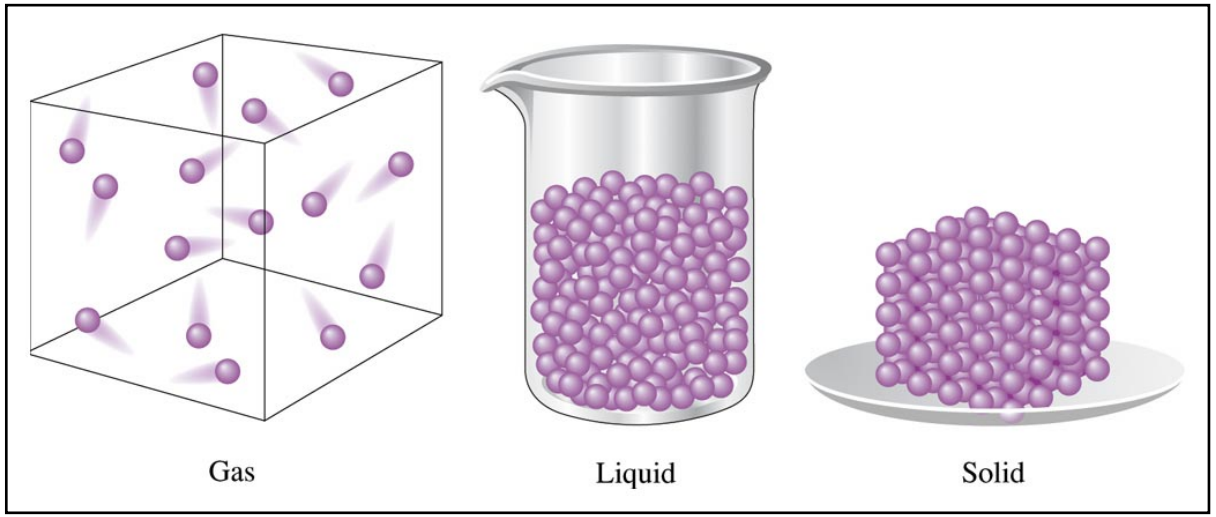
States of Matter: Liquids and Solids

- Comparison of **gases**, **liquids**, and **solids**.

– **Gases** are compressible fluids. Their molecules are widely separated.

– **Liquids** are relatively incompressible fluids. Their molecules are more tightly packed.

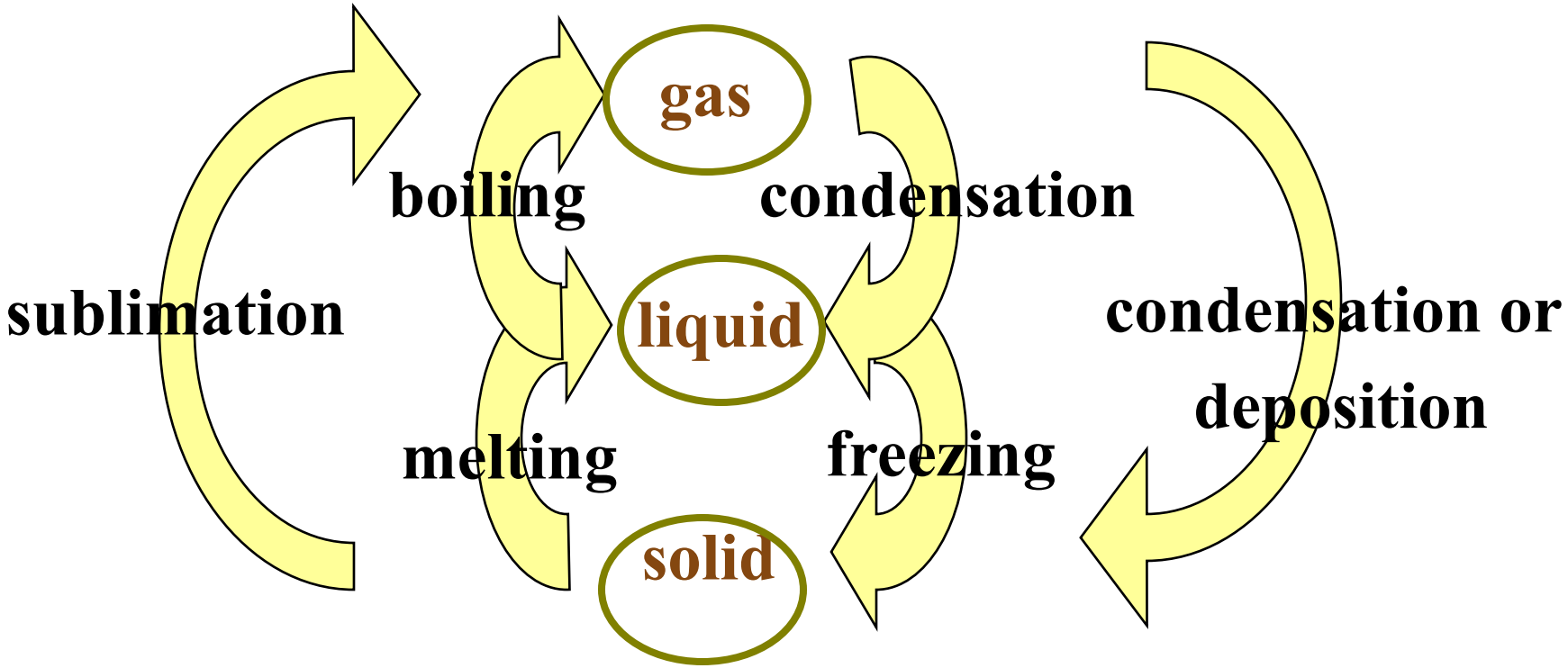
– **Solids** are nearly incompressible and rigid. Their molecules or ions are in close contact and do not move.



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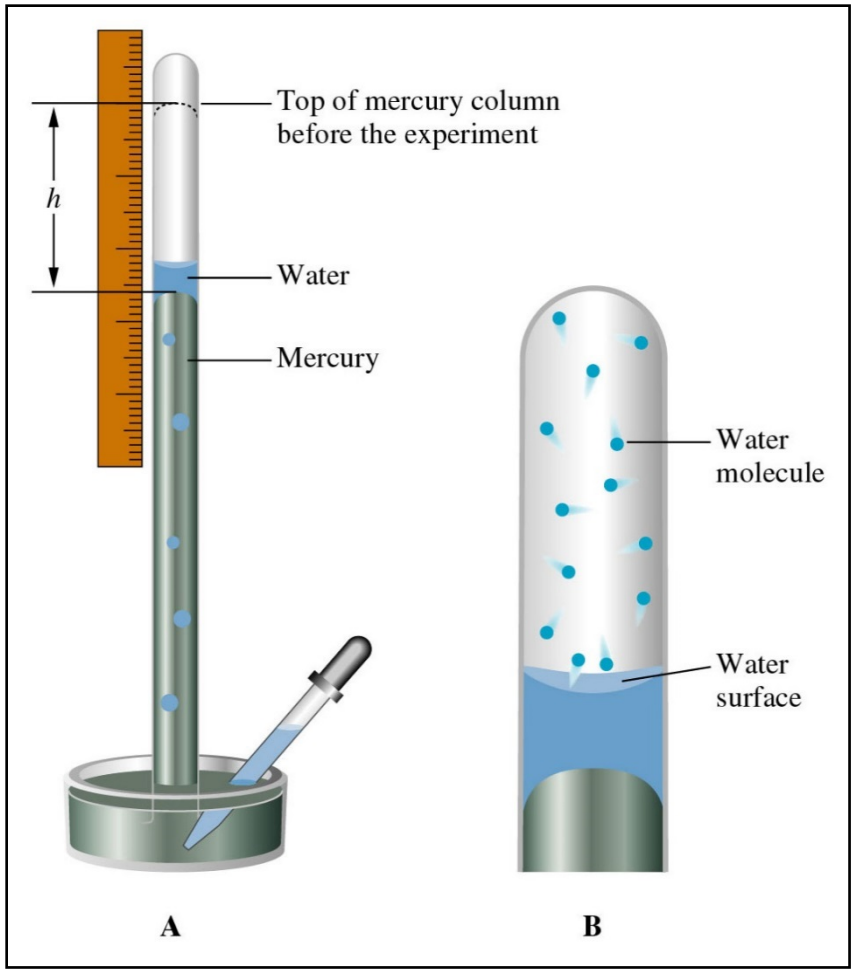
Changes of State

- A change of state or phase transition is a change of a substance from one state to another.



Vapor Pressure

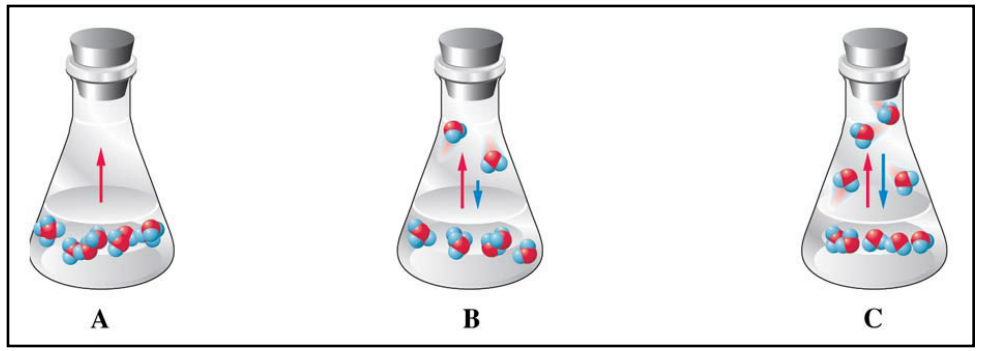
- Liquids are continuously vaporizing.
  - If a liquid is in a closed vessel with space above it, a partial pressure of the vapor state builds up in this space.
  - The **vapor pressure** of a liquid is the partial pressure of the vapor over the liquid, measured at equilibrium at a given temperature.



Measurement of the vapor pressure of water.

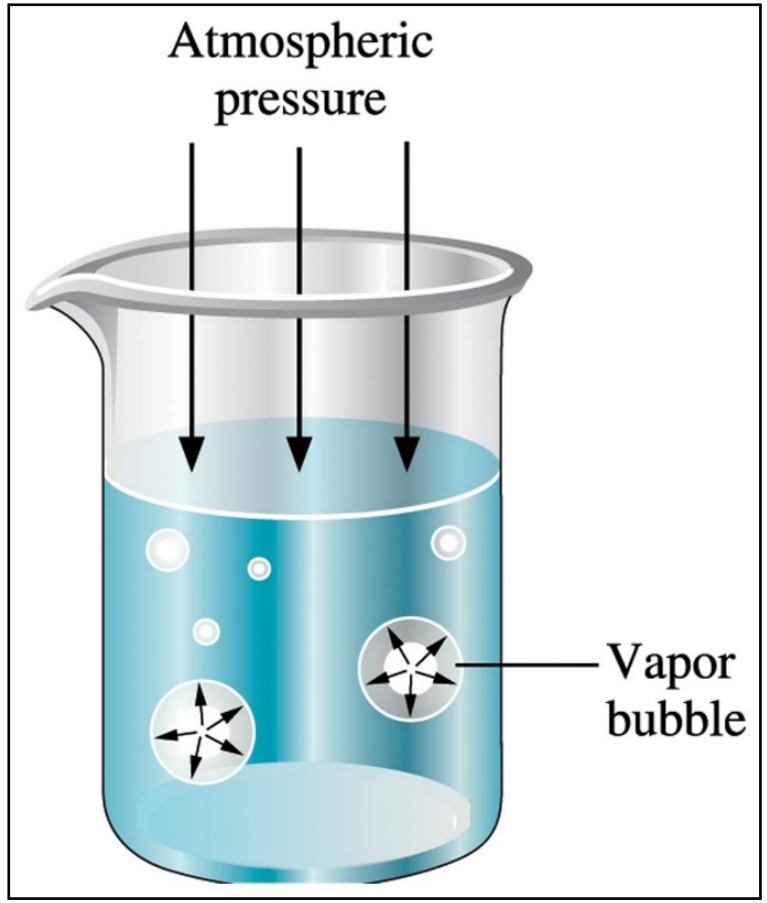
Vapor Pressure

- The vapor pressure of a liquid **depends on its temperature**.
  - As the temperature increases, the kinetic energy of the molecular motion becomes greater, and vapor pressure increases.
  - Liquids and solids with relatively high vapor pressures at normal temperatures are said to be **volatile**.



Boiling Point

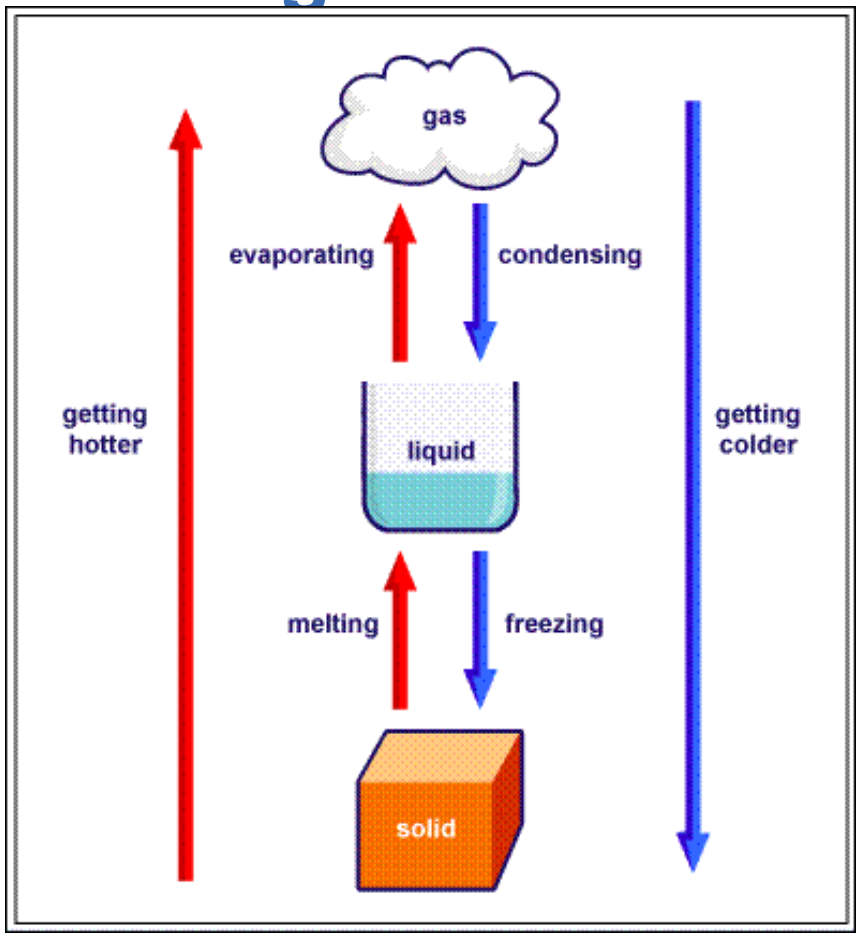
- The temperature at which the vapor pressure of a liquid equals the pressure exerted on the liquid is called the **boiling point**.



- As the temperature of a liquid increases, the vapor pressure increases until it reaches atmospheric pressure.
- At this point, stable bubbles of vapor form within the liquid. This is called **boiling**.
- The **normal boiling point** is the boiling point at 1 atm.

Freezing Point

- The temperature at which a pure liquid changes to a crystalline solid, or **freezes**, is called the **freezing point**.
  - The **melting point** is identical to the freezing point and is defined as the temperature at which a solid becomes a liquid.
  - Unlike boiling points, melting points are affected significantly by only large pressure changes.



Heat of Phase Transition

- To melt a pure substance at its melting point requires an extra boost of energy to overcome **lattice energies**.
  - The heat needed to melt 1 mol of a pure substance is called the **heat of fusion** and denoted ΔH_{fus}.
  - For ice, the heat of fusion is 6.01 kJ/mol.

$$H_2O(s) \rightarrow H_2O(l); \qquad \Delta H_{fus} = 6.01\,kJ$$

A Problem to Consider

- The heat of vaporization of ammonia is 23.4 kJ/mol. How much heat is required to vaporize 1.00 kg of ammonia?
  - First, we must determine the number of moles of ammonia in 1.00 kg (1000 g).

$$1.00\times10^{3}\ \text{g NH}_3 \times \frac{1\ \text{mol NH}_3}{17.0\ \text{g NH}_3} = 58.8\ \text{mol NH}_3$$

A Problem to Consider

- The heat of vaporization of ammonia is 23.4 kJ/mol. How much heat is required to vaporize 1.00 kg of ammonia?
  - Then we can determine the heat required for vaporization.

$$58.8\ \text{mol}\ NH_3 \times 23.4\ \text{kJ/mol} = 1.38 \times 10^3\ \text{kJ}$$

Intermolecular Forces; Explaining Liquid Properties

- **Viscosity** is the resistance to flow exhibited by all liquids and gases.
  - Viscosity can be illustrated by measuring the time required for a steel ball to fall through a column of the liquid.
  - Even without such measurements, you know that syrup has a greater viscosity than water.
  - In comparisons of substances, **as intermolecular forces increase, viscosity usually increases.**

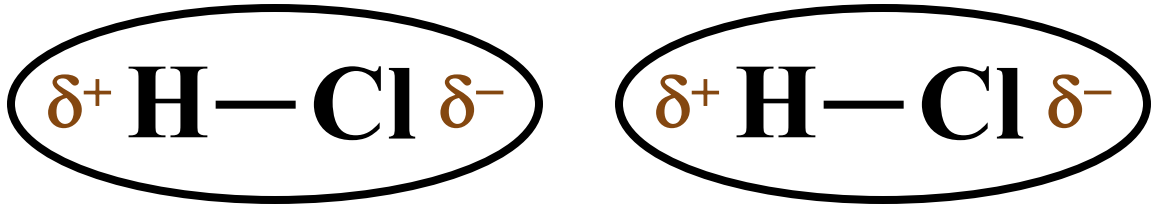
Intermolecular Forces; Explaining Liquid Properties

- Many of the physical properties of liquids (and certain solids) can be explained in terms of **intermolecular forces, the forces of attraction between molecules**.

  - Three types of forces are known to exist between neutral molecules.
    1. **Dipole-dipole forces**
    2. **London (or dispersion) forces**
    3. **Hydrogen bonding**

Dipole-Dipole Forces

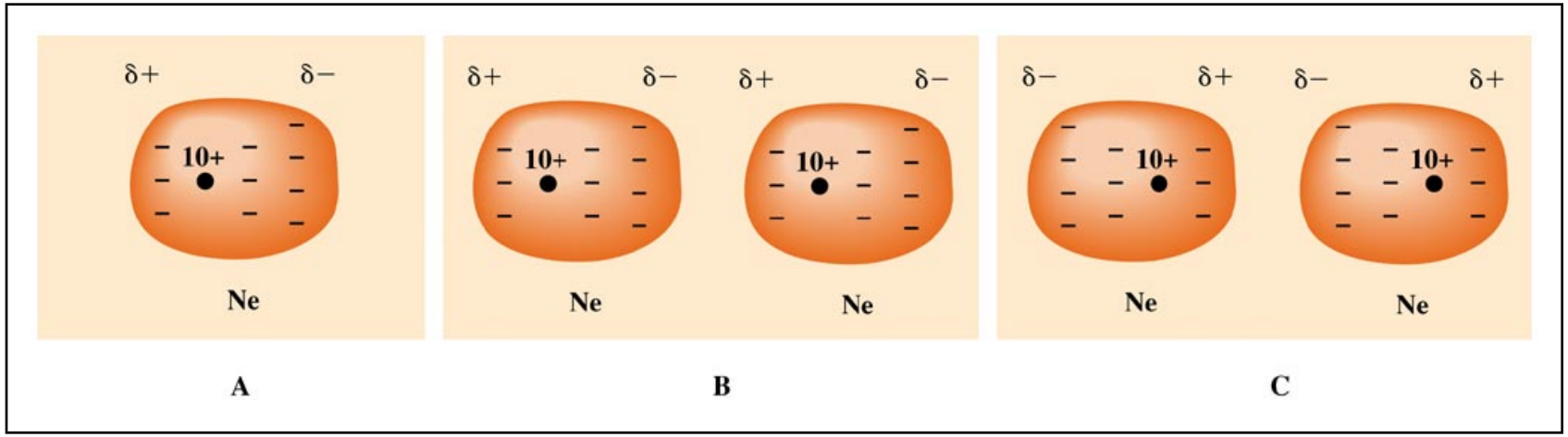
- Polar molecules can attract one another through **dipole-dipole forces**.
  - The **dipole-dipole force** is an attractive intermolecular force resulting from the tendency of **polar molecules** to align themselves positive end to negative end.



shows the alignment of polar molecules.

London Forces

- **London forces** are the weak attractive forces resulting from **instantaneous dipoles** that occur due to the **distortion of the electron cloud** surrounding a molecule.
  - London forces **increase with molecular weight**. The larger a molecule, the more easily it can be distorted to give an instantaneous dipole.
  - **All** covalent molecules exhibit some London force.
  - Figure illustrates the effect of London forces.



Van der Waals Forces and the Properties of Liquids

- In summary, intermolecular forces play a large role in many of the physical properties of liquids and gases. These include:
  - vapor pressure
  - boiling point
  - viscosity

Van der Waals Forces and the Properties of Liquids

- The **vapor pressure** of a liquid depends on intermolecular forces. When the intermolecular forces in a liquid are strong, you expect the vapor pressure to be low.

Van der Waals Forces and the Properties of Liquids

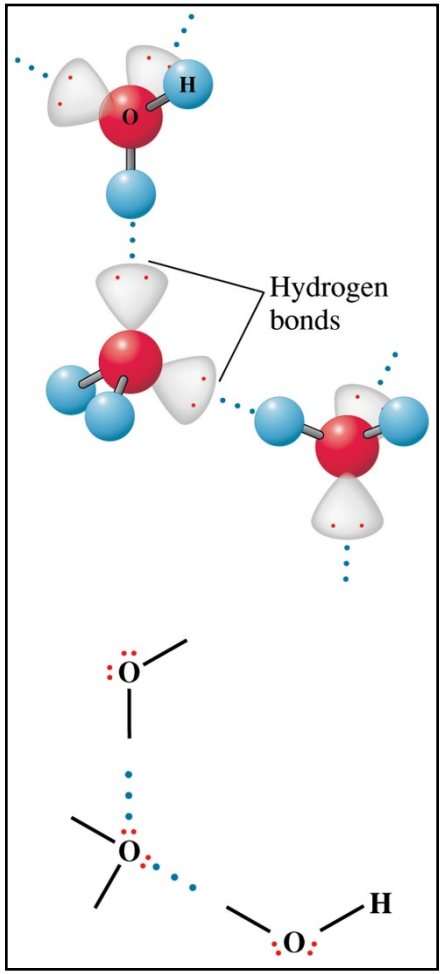
- **Viscosity** increases with increasing intermolecular forces because increasing these forces increases the resistance to flow.
  - Other factors, such as the possibility of molecules tangling together, affect viscosity.
  - Liquids with long molecules that tangle together are expected to have high viscosities.

Hydrogen Bonding

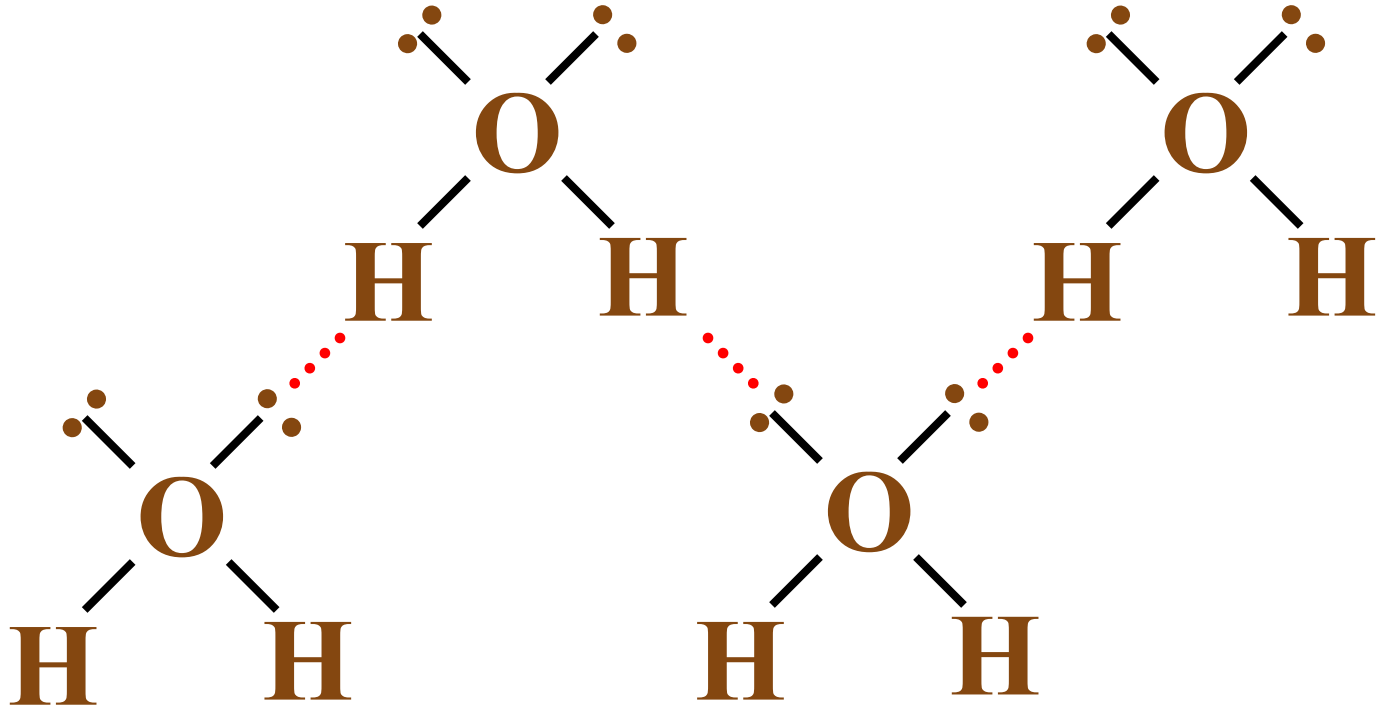
- **Hydrogen bonding** is a force that exists between a hydrogen atom covalently bonded to a very electronegative atom, X, and a lone pair of electrons on a very electronegative atom, Y.
  - To exhibit hydrogen bonding, one of the following three structures must be present.



  - Only N, O, and F are electronegative enough to leave the hydrogen nucleus exposed.



Hydrogen Bonding



Hydrogen Bonding

- Molecules exhibiting hydrogen bonding have **abnormally high boiling points** compared to molecules with similar van der Waals forces.
  - For example, water has the highest boiling point of the Group VI hydrides.
  - Similar trends are seen in the Group V and VII hydrides.

Hydrogen Bonding

- A hydrogen atom bonded to an electronegative atom appears to be special.
  - The electrons in the O-H bond are drawn to the O atom, leaving the dense positive charge of the hydrogen nucleus exposed.
  - It's the strong attraction of this exposed nucleus for the lone pair on an adjacent molecule that accounts for the strong attraction.
  - A similar mechanism explains the attractions in HF and NH_3.

Solid State

- A **solid** is a nearly incompressible state of matter with a well-defined shape. The units making up the solid are in **close contact** and in **fixed positions**.
  - Solids are characterized by the **type of force** holding the structural units together.
  - In some cases, these forces are **intermolecular**, but in others they are **chemical bonds** (metallic, ionic, or covalent).

From this point of view, there are four types of solids.

Molecular (Van der Waals forces) Metallic (Metallic bond) Ionic (Ionic bond) Covalent (Covalent bond)

Table 11.4

Types of Solids

Type of Solid	Structural Units	Attractive Forces Between Structural Units	Examples
Molecular	Atoms or molecules	Intermolecular forces	Ne, H_2O, CO_2
Metallic	Atoms (positive cores surrounded by electron "sea")	Metallic bonding (extreme delocalized bond)	Fe, Cu, Ag
Ionic	Ions	Ionic bonding	CsCl, NaCl, ZnS
Covalent network	Atoms	Covalent bonding	Diamond, graphite, asbestos

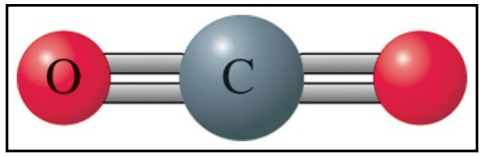
Table 11.5

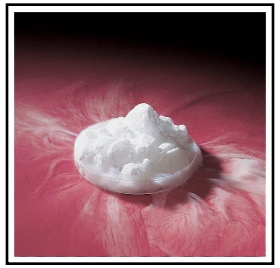
Properties of the Different Types of Solids

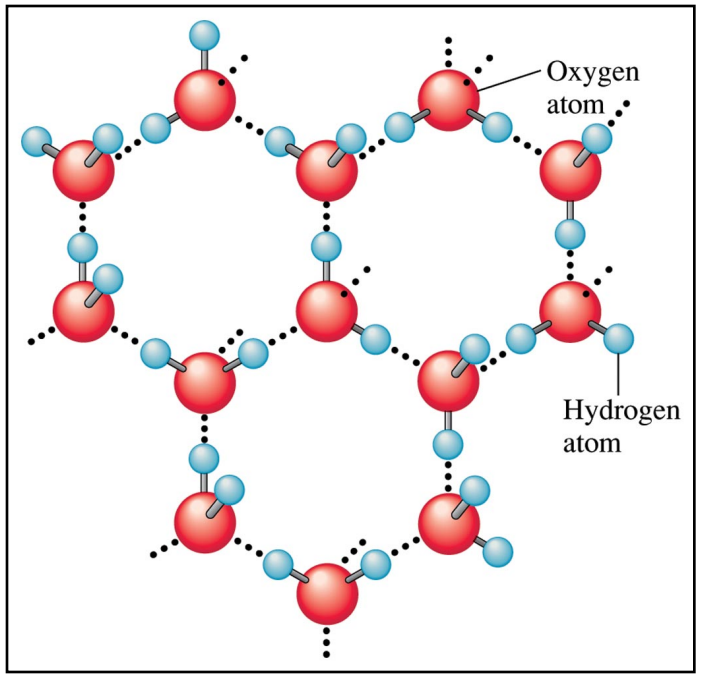
Type of Solid	Melting Point	Hardness and Brittleness	Electrical Conductivity
Molecular	Low	Soft and brittle	Nonconducting
Metallic	Variable	Variable hardness; malleable	Conducting
Ionic	High to very high	Hard and brittle	Nonconducting solid (conducting liquid)
Covalent network	Very high	Very hard	Usually nonconducting

Types of Solids

- A **molecular solid** is a solid that consists of atoms or molecules held together by intermolecular forces.
  - Many solids are of this type.
  - Examples include solid neon, solid water (ice), and solid carbon dioxide (dry ice).







Types of Solids

- A **metallic solid** is a solid that consists of positive cores of atoms held together by a surrounding “sea” of electrons (metallic bonding).
  - In this kind of bonding, positively charged atomic cores are surrounded by delocalized electrons.
  - Examples include iron, copper, and silver.



Types of Solids

- An **ionic solid** is a solid that consists of cations and anions held together by electrical attraction of opposite charges (ionic bond).
  - Examples include cesium chloride, sodium chloride, and zinc sulfide (but ZnS has considerable covalent character).

Types of Solids

- A **covalent network solid** is a solid that consists of atoms held together in large networks or chains by covalent bonds.
  - Examples include carbon, in its forms as diamond or graphite, asbestos, and silicon carbide.
  - Table 11.5 summarizes these four types of solids.



Physical Properties

- Many physical properties of a solid can be attributed to its structure.
  - **Melting Point and Structure**
    - For a solid to melt, the forces holding the structural units together must be overcome.
    - For a **molecular solid**, these are weak intermolecular attractions.
    - Thus, **molecular solids tend to have low melting points** (below 300^{o}C).

Physical Properties

- Many physical properties of a solid can be attributed to its structure.
  - **Melting Point and Structure**
    - For **ionic solids** and **covalent network solids** to melt, chemical bonds must be broken.
    - For that reason, **their melting points are relatively high**.

Table 11.1

Melting Points and Boiling Points (at 1 atm) of Several Substances

Name	Type of Solid*	Melting Point, °C	Boiling Point, °C
Neon, Ne	Molecular	−249	−246
Hydrogen sulfide, H_2S	Molecular	−86	−61
Chloroform, $CHCl_3$	Molecular	−64	62
Water, H_2O	Molecular	0	100
Acetic acid, $HC_2H_3O_2$	Molecular	17	118
Mercury, Hg	Metallic	−39	357
Sodium, Na	Metallic	98	883
Tungsten, W	Metallic	3410	5660
Cesium chloride, CsCl	Ionic	645	1290
Sodium chloride, NaCl	Ionic	801	1413
Magnesium oxide, MgO	Ionic	2800	3600
Quartz, SiO_2	Covalent network	1610	2230
Diamond, C	Covalent network	3550	4827

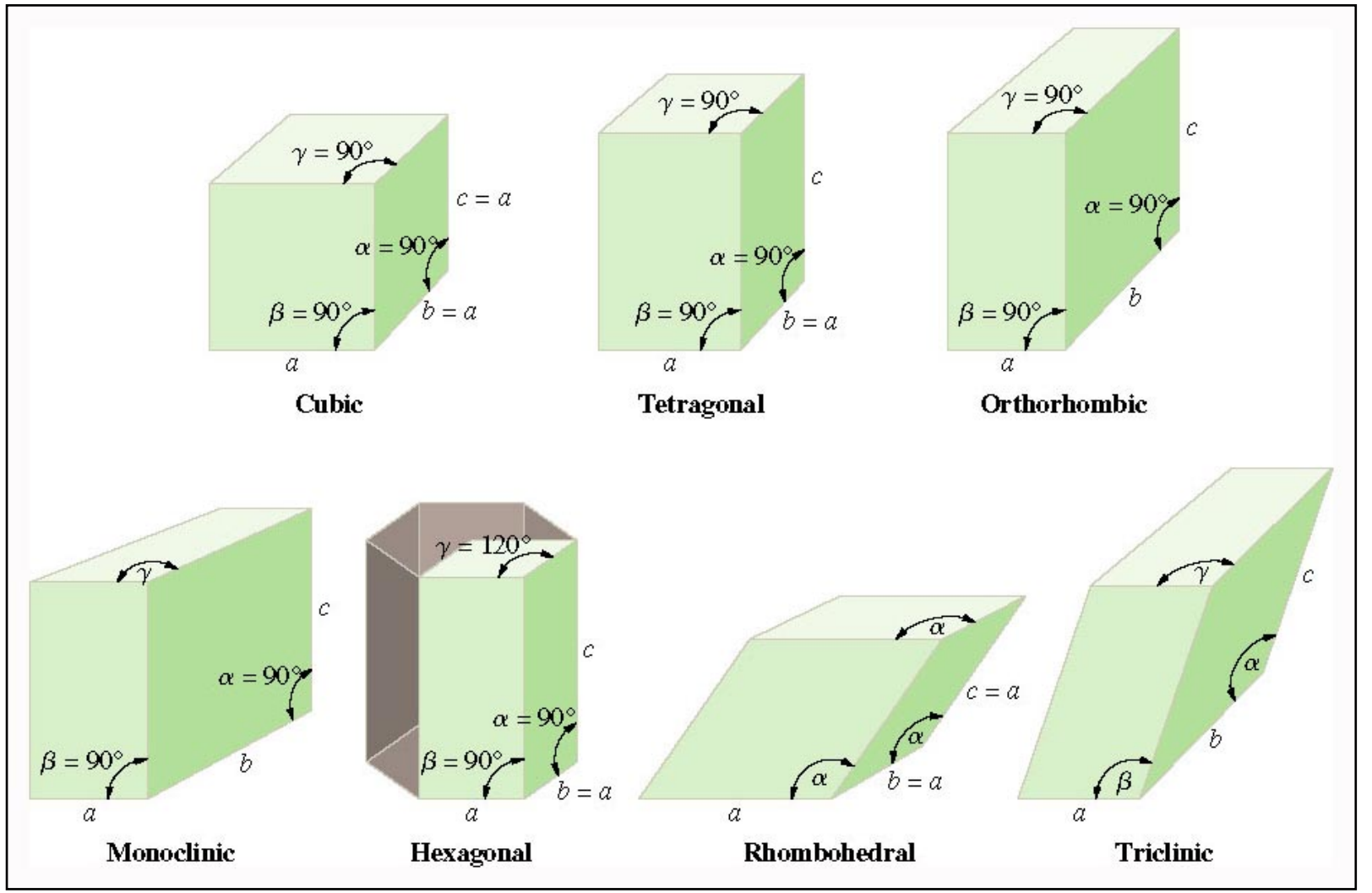
*Types of solids are discussed in Section 11.6.

Crystalline Solids; Crystal Lattices and Unit Cells

- Solids can be crystalline or amorphous.
  - A **crystalline solid** is composed of one or more crystals; each crystal has a well-defined, ordered structure in three dimensions.
    - Examples include sodium chloride and sucrose.
  - An **amorphous solid** has a disordered structure. It lacks the well-defined arrangement of basic units found in a crystal.
    - Glass is an amorphous solid.

Crystal Lattices

- A **crystal lattice** is the geometric arrangement of lattice points in a crystal.
  - These seven systems can have more than one possible crystal lattice.
  - A “primitive” lattice has lattice points only at the corners of each cell.

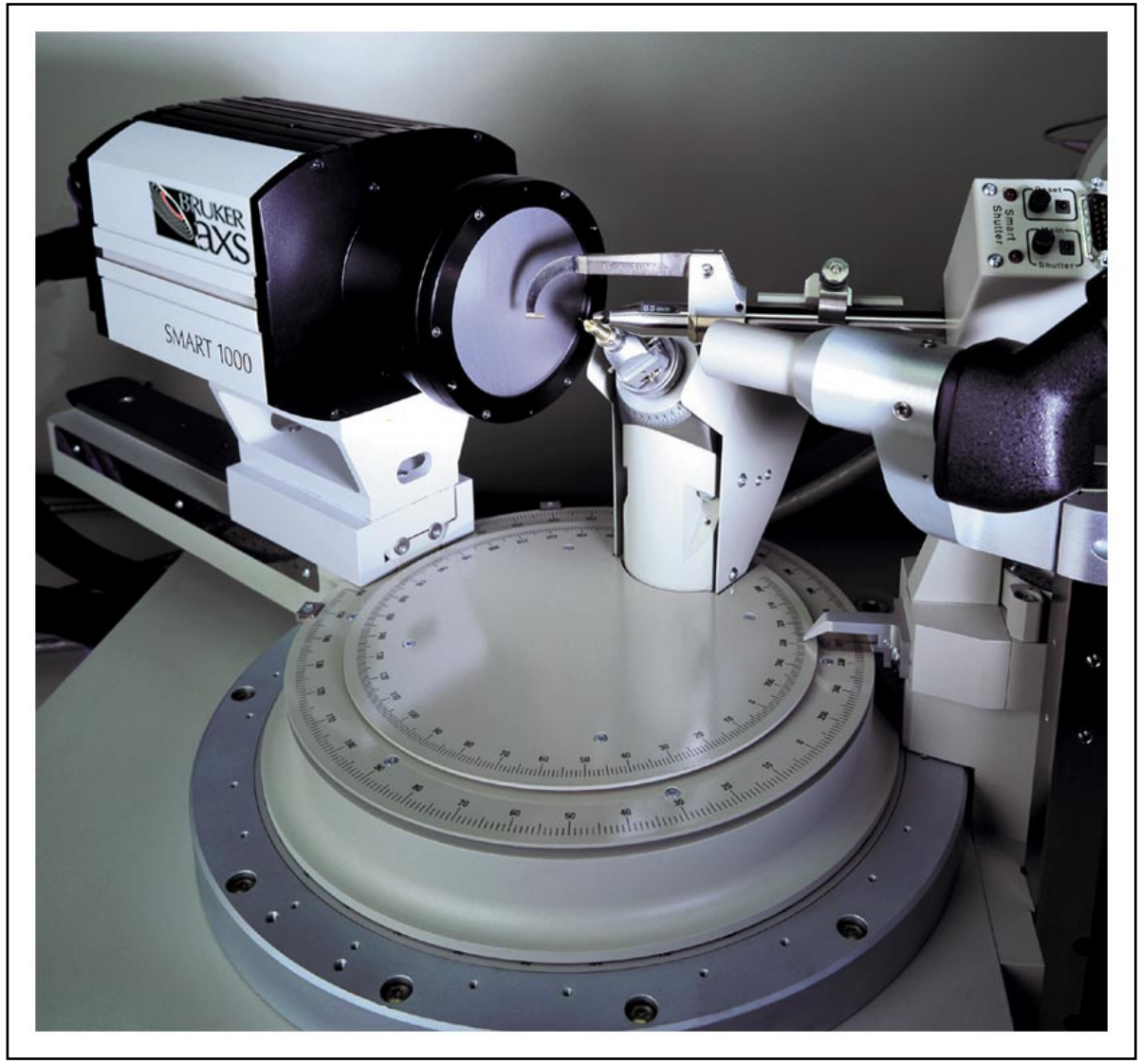

γ = 90°
c = a
α = 90°
β = 90°
b = a
a
Cubic
γ = 90°
c
α = 90°
β = 90°
b = a
a
Tetragonal
γ = 90°
c
α = 90°
β = 90°
b
a
Orthorhombic
γ
c
α = 90°
β = 90°
b
a
Monoclinic
γ = 120°
c
α = 90°
β = 90°
b = a
a
Hexagonal
α
c = a
α
α
b = a
a
Rhombohedral
γ
c
α
b
β
a
Triclinic

Calculations Involving Unit Cell Dimensions

- **X-ray diffraction** is a method for determining the structure and dimensions of a unit cell in a crystalline compound.
  - Once the dimensions and structure are known, the volume and mass of a **single atom** in the crystal can be calculated.
  - The determination of the mass of a single atom gave us one of the first accurate determinations of Avogadro's number.

Determination of Crystal Lattice by X-Ray Diffraction

- When x-rays are reflected from the planes of a crystal, they show a diffraction pattern that can be recorded on photographic film.
  - Analysis of these diffraction patterns allows the determination of the positions of the atoms in the unit cell of the solid.
  - The following figures illustrate how the diffraction of the x-rays occurs.



Automated X-ray Diffractometer

Automated X-ray Diffractometer

