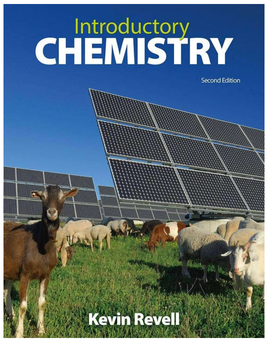
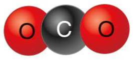
Introductory Chemistry
Chem 103

Chapter 9 – Covalent Bonding and Molecules

Lecture Slides



Covalent Molecules

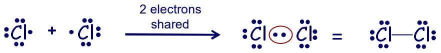


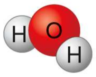
Covalent bonds:

Two atoms share electrons
Occur between nonmetal atoms

Octet rule:

Atoms are stabilized by having 8 electrons in the valence shell





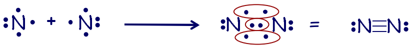
Lewis structures:

Show the arrangement of covalently bonded atoms
Use a dash to represent two shared electrons

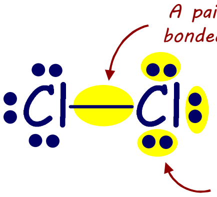
Covalent Double and Triple Bonds

Covalent double bonds:

Covalent triple bonds:



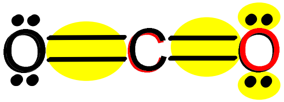
Pairs of Electrons in Compounds



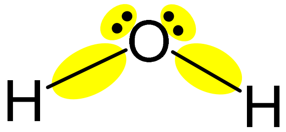
A pair of covalently bonded electrons

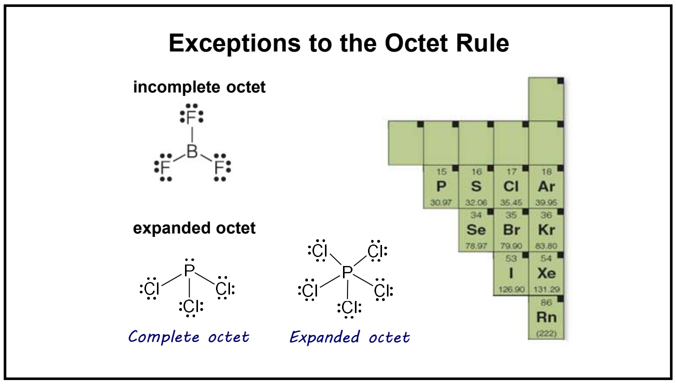
A lone pair of electrons

In Most Molecules, Atoms Follow the Octet Rule



In Most Molecules, Atoms Follow the Octet Rule, Continued





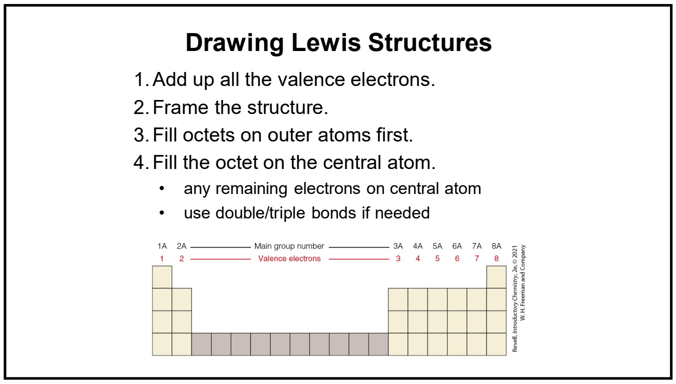
Exceptions to the Octet Rule

incomplete octet

expanded octet

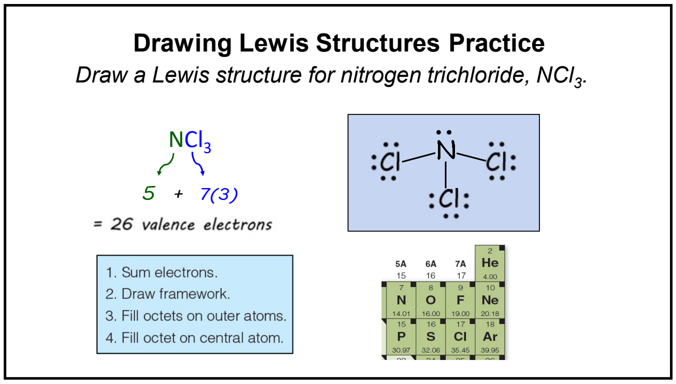
Complete octet

Expanded octet



Drawing Lewis Structures

1. Add up all the valence electrons.
2. Frame the structure.
3. Fill octets on outer atoms first.
4. Fill the octet on the central atom.
   - any remaining electrons on central atom
   - use double/triple bonds if needed



Drawing Lewis Structures Practice

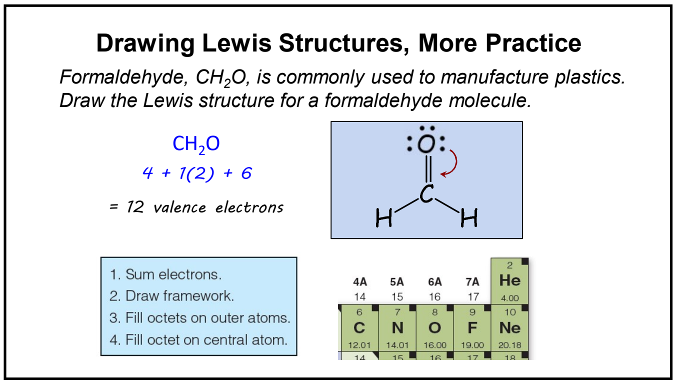
Draw a Lewis structure for nitrogen trichloride, NCl_3.

NCl_3

5 + 7(3)

= 26 valence electrons

1. Sum electrons.
2. Draw framework.
3. Fill octets on outer atoms.
4. Fill octet on central atom.



Drawing Lewis Structures, More Practice

Formaldehyde, CH_2O, is commonly used to manufacture plastics. Draw the Lewis structure for a formaldehyde molecule.

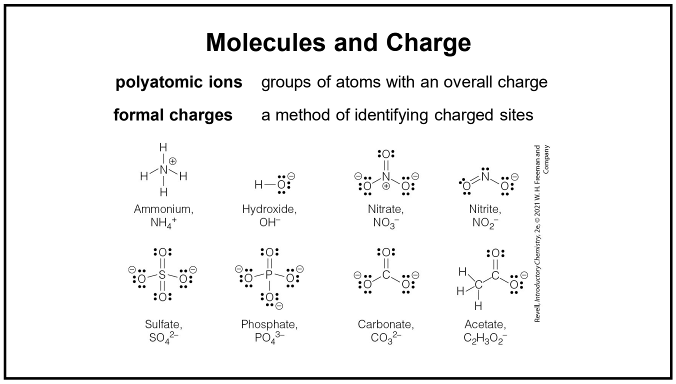
CH_2O

4 + 1(2) + 6

= 12 valence electrons

1. Sum electrons.
2. Draw framework.
3. Fill octets on outer atoms.
4. Fill octet on central atom.

4A 14	5A 15	6A 16	7A 17	2 He 4.00
6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18



Molecules and Charge

polyatomic ions	groups of atoms with an overall charge
formal charges	a method of identifying charged sites

Ammonium, NH_4^+

Hydroxide, OH^-

Nitrate, NO_3^-

Nitrite, NO_2^-

Sulfate, SO_4^{2-}

Phosphate, PO_4^{3-}

Carbonate, CO_3^{2-}

Acetate, $C_2H_3O_2^-$

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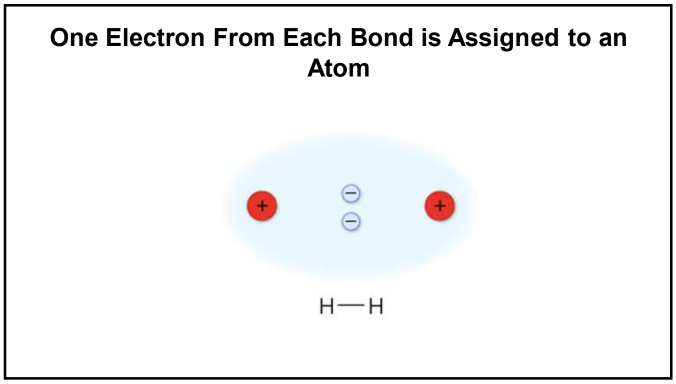
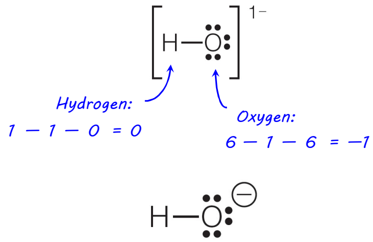


One Electron From Each Bond is Assigned to an Atom

Calculating Formal Charge Practice

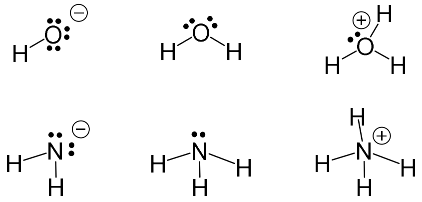
Formal charge = Valence electrons in the neutral atom − Number of covalent bonds − Number of unshared electrons



Hydrogen:
1 − 1 − 0 = 0

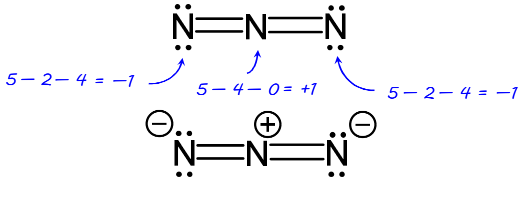
Oxygen:
6 − 1 − 6 = −1

Oxygen and Nitrogen Atoms Often Have Formal Charges



Calculating Formal Charge, More Practice

Automotive air bags contain sodium azide, NaN_3. The Lewis structure for the azide ion (without charges) is shown. Calculate the formal charge on each atom in this structure. What is the overall charge of the azide ion?



5 − 2 − 4 = −1

5 − 4 − 0 = +1

5 − 2 − 4 = −1

Azide Ion: N_3^-

Drawing Lewis Structures for Polyatomic Ions

- Similar to neutral molecules
- Consider charge when finding the number of valence electrons

How many valence electrons are in a hydroxide ion?

OH^-

$6 + 1 + 1$ = 8 valence electrons

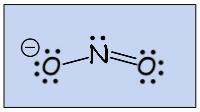
Lewis Structures for Polyatomic Ions Practice

Draw a Lewis structure for the nitrite ion, NO_2^-. Show all nonzero formal charges.

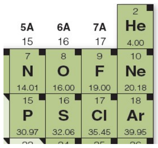
NO_2^-

$5 + 6(2) + 1$

= 18 valence electrons



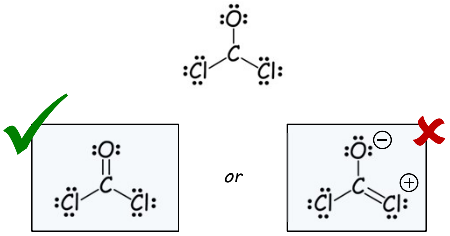
1. Sum electrons.
2. Draw framework.
3. Fill octets on outer atoms.
4. Fill octet on central atom.



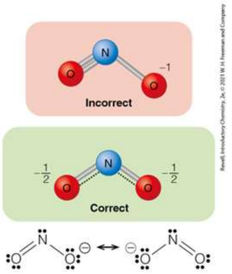
5A 15	6A 16	7A 17	2 He 4.00
7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95

Identifying the Best Lewis Structure

What is the best structure for phosgene, $COCl_2$?



Resonance Structures



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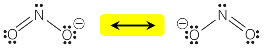
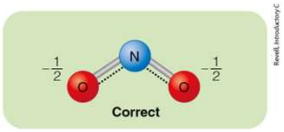
Resonance Structures, Continued

- a set of structures that show how electrons are distributed.
- used when a single Lewis structure is insufficient.





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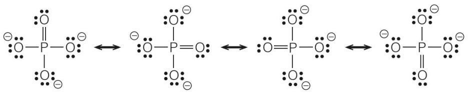
Ions With Resonance Structures Spread Charges Over Multiple Atoms

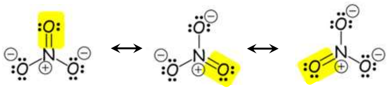
NO_3^-	SO_4^{2-}
NO_2^-	CO_3^{2-}
PO_4^{3-}	$C_2H_3O_2^-$



Only 2nd bonds and lone pairs change in resonance structures.

Using Resonance Structures to Calculate Formal Charge

The nitrate ion (NO_3^-) has three major resonance structures. Draw each structure. Based on these structures, what is the average charge on each oxygen atom?



charge on each oxygen: $-\frac{2}{3}$

Shapes of Molecules

CO_2

gas

H_2O

liquid

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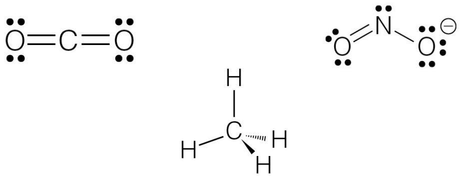
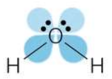
Predicting Molecular Shapes

Valence Shell Electron Pair Repulsion

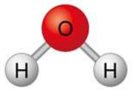
VSEPR



Predicting Molecular Shapes, Continued



Electronic geometry
Arrangement of electrons
around the central atom



Molecular geometry
Shape caused by the
arrangement of atoms

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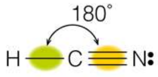
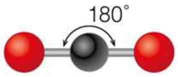
Two Electron Sets: Linear



Electronic geometry
Linear



Molecular geometry
Linear

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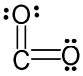
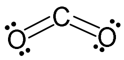
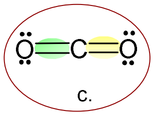
double and triple
bonds count as 1 "set"

Geometric Stability

Which geometry is most stable?

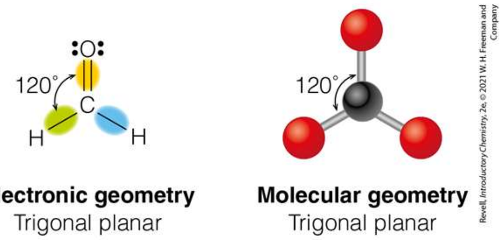

a.


b.


c.

2 sets of electrons: Linear

Three Electron Sets: Trigonal Planar



Electronic geometry
Trigonal planar

Molecular geometry
Trigonal planar

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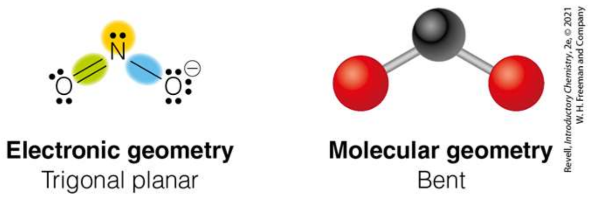
Three Electron Sets: Trigonal Planar, Continued



Electronic geometry
Trigonal planar

Molecular geometry
Bent

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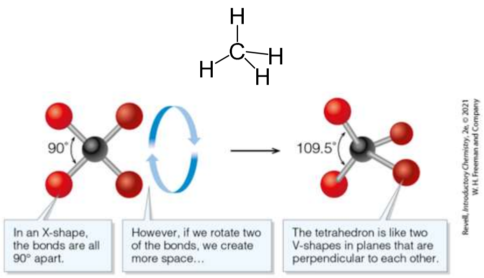
Four Electron Sets: Tetrahedral, Part 1



In an X-shape, the bonds are all 90° apart.

However, if we rotate two of the bonds, we create more space…

The tetrahedron is like two V-shapes in planes that are perpendicular to each other.

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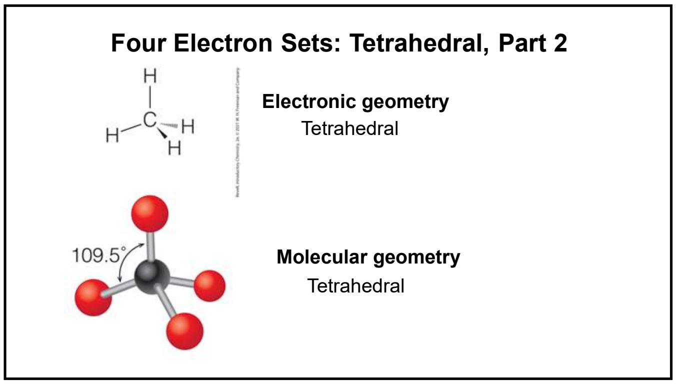
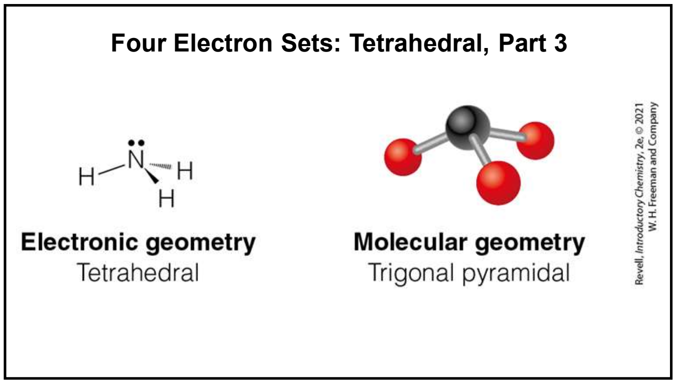


Four Electron Sets: Tetrahedral, Part 2

Electronic geometry
Tetrahedral

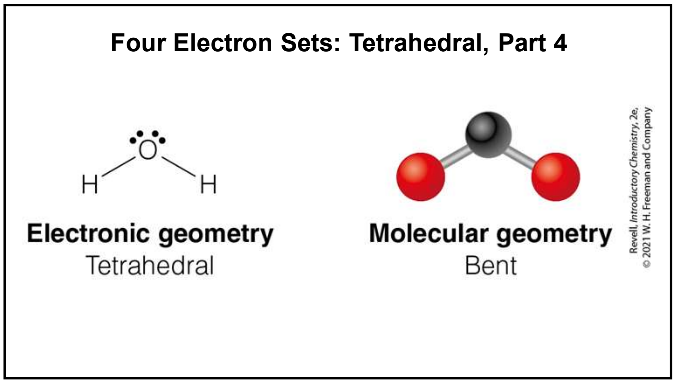
Molecular geometry
Tetrahedral



Four Electron Sets: Tetrahedral, Part 3

Electronic geometry
Tetrahedral

Molecular geometry
Trigonal pyramidal

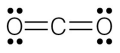
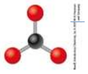
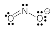
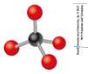


Four Electron Sets: Tetrahedral, Part 4

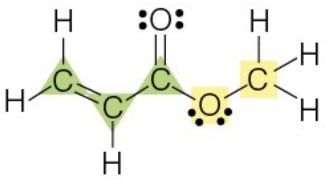
Electronic geometry
Tetrahedral

Molecular geometry
Bent

Electronic and Molecular Geometry

Electron sets	Electronic geometry	Model	Bonding sets	Lone pairs	Molecular geometry	Examples
2	Linear		2	0	Linear	
3	Trigonal Planar		3	0	Trigonal Planar	
			2	1	Bent	
4	Tetrahedral		4	0	Tetrahedral	
			3	1	Trigonal pyramidal	
			2	2	Bent	

Electronic and Molecular Geometry Practice

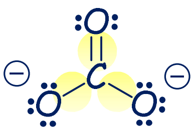


🟩 Trigonal planar

🟨 Tetrahedral

Electronic and Molecular Geometry, More Practice

Lithium carbonate is a simple ionic compound that is widely used to treat bipolar disorder. What is the molecular geometry of the carbonate ion?



Electronic geometry:
Trigonal planar

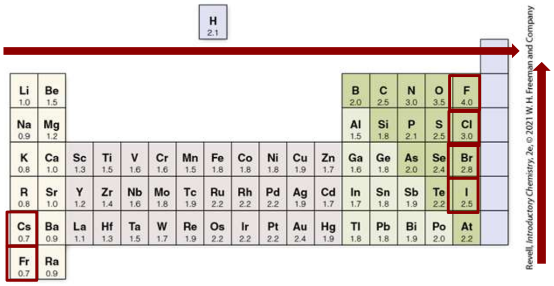
Molecular geometry:
Trigonal planar

Polar Bonds and Molecules

Polar covalent bond atoms do not share the electrons evenly

Electronegativity

how strongly atoms pull bonded electrons



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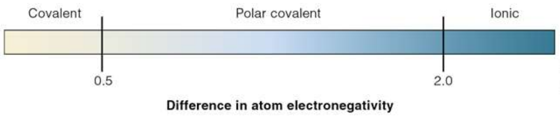
Comparing Covalent, Polar Covalent and Ionic Bonds

F—F H—F Na^+ $:F^-$



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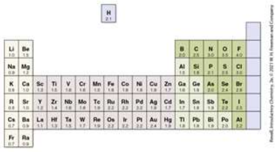
Covalent: < 0.5

Polar Covalent: 0.5 – 2.0

Ionic: > 2.0



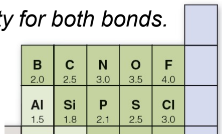
An Analogy for Polar Covalent Bonds

H–F


Kevin Revell

Classifying Bonds Practice

Which bond is more polar, a C-O bond or an F-S bond? Show the direction of polarity for both bonds.



B	C	N	O	F
2.0	2.5	3.0	3.5	4.0
Al	Si	P	S	Cl
1.5	1.8	2.1	2.5	3.0

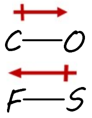
C: 2.5
O: 3.5

Difference = 1.0
Polar Covalent

F: 4.0
S: 2.5

Difference = 1.5
Polar Covalent

F-S is more polar



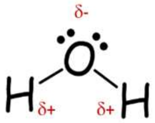
$\delta+$ $\delta-$
C—O

$\delta-$ $\delta+$
F—S

Molecules with dipoles

Molecular dipole an overall polarity in a molecule

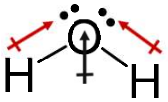
net dipole
dipole



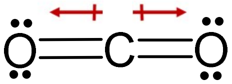
polar covalent bonds
shape


Rich Carey/Shutterstock

Identifying Molecules with a Net Dipole

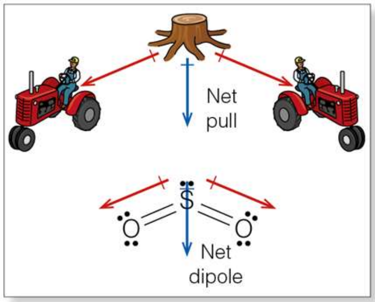


Net Dipole



No Net Dipole

Identifying Molecules with a Net Dipole, Continued



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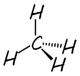
Identifying Molecules with a Net Dipole Practice

Which of these have a net dipole?

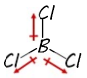
CH_4



Non-polar bonds

No dipole

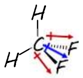
BCl_3



Polar bonds

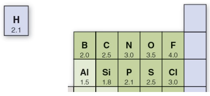
No net dipole

CH_2F_2



Polar bonds

Net dipole



H 2.1

B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0

How Dipoles Affect Properties – A Preview

CO_2 Linear

No net dipole

H_2O Bent

Net dipole

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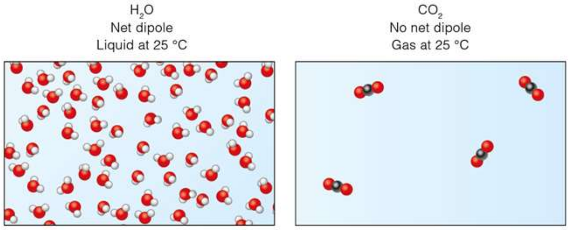
How Dipoles Affect Properties – A Preview, Continued



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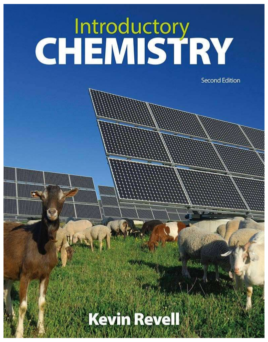
Introductory Chemistry
Chem 103

Chapter 10 – Solids, Liquids, Gases

Lecture Slides

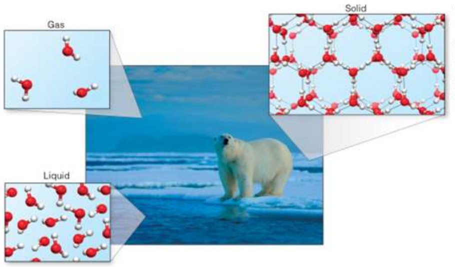


Interactions between Particles

	Atomic/Molecular Arrangement	Macroscopic Properties
Solid	Particles are close together and held in a fixed place.	Definite shape and volume
Liquid	Particles are close together but move freely past each other.	Definite volume; Adopts the shape of the container.
Gas	Particles are far apart and have very little interaction.	Adopts shape and volume of container

Phase change

A transition from one state of matter to another.



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The Forces Between Particles Influence Physical Properties

Stronger forces between particles - higher melting and boiling points.

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Solids and Liquids

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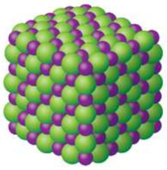
Ionic Substances

Lattices: rigid frameworks of atoms, molecules or ions.



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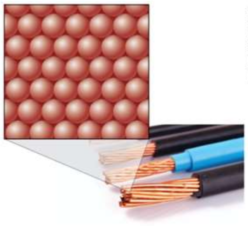
Compound	Melting Point (°C)
$NaCl$	801
KCl	770
$MgCl_2$	714
CaO	2,572
Al_2O_3	2,072

Metallic Substances

- Form lattices of tightly packed atoms.
- Electrons move easily between atoms.
- Shapes of metals are easily altered.
  - Malleable
  - Ductile



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Metallic Substances, Continued

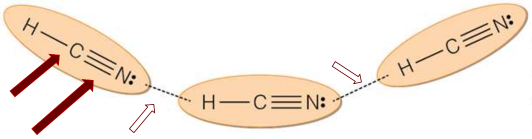

Cultura Creative (RF)/Alamy

Element	Melting Point (°C)
Lead	327
Aluminum	660
Gold	1,064
Copper	1,085
Iron	1,538

Molecular Substances

Forces <u>within</u> molecules: covalent bonds

Forces <u>between</u> molecules: intermolecular forces



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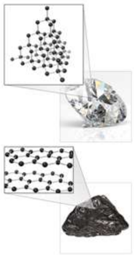
Covalent Networks and Polymers

covalent networks: lattices of covalent bonds that form giant molecules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Covalent Networks and Polymers, Continued

polymers: contain long chains of covalently-bonded atoms

V: Courtesy of Kevin Revell

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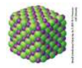
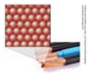
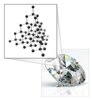
Solids and Liquids Summary



- Ionic Substances
- Metallic Substances



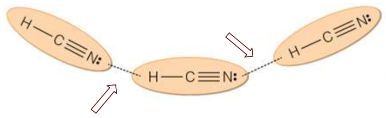
- Molecular Substances
- Covalent Networks and Polymers



Forces Within and Between Molecules

Forces within molecules: covalent bonds

Forces between molecules: intermolecular forces



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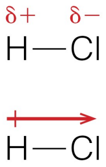
Forces **Between** Molecules

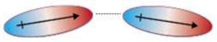
intermolecular forces

1. Dipole-dipole Interactions
2. Hydrogen bonds
3. Dispersion forces

Dipole-Dipole Interactions, Part 1

Attractions between polar covalent molecules:

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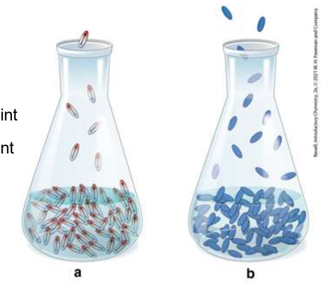
Dipole-Dipole Interactions, Part 2

Dipole

Higher Melting Point

Higher Boiling Point



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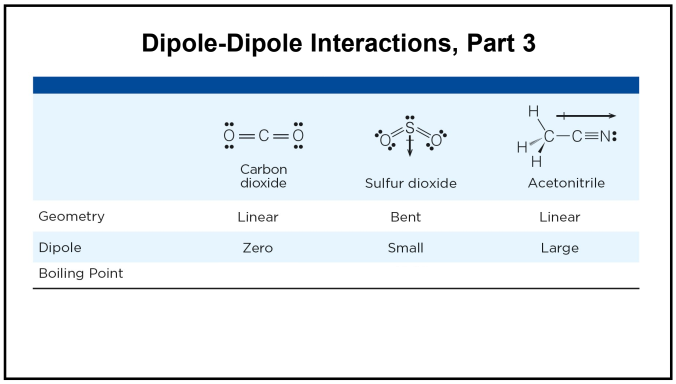
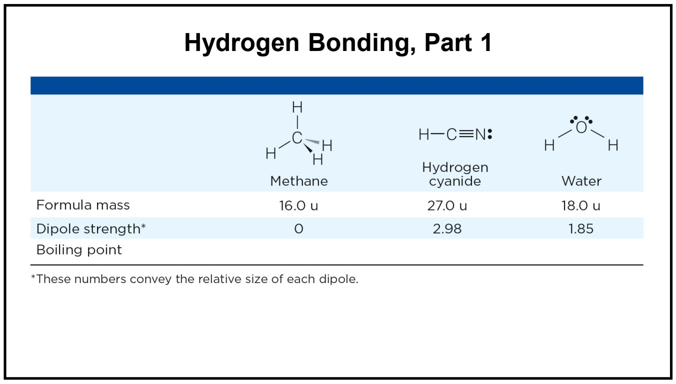
No dipole



Dipole-Dipole Interactions, Part 3

	Carbon dioxide	Sulfur dioxide	Acetonitrile
Geometry	Linear	Bent	Linear
Dipole	Zero	Small	Large
Boiling Point			



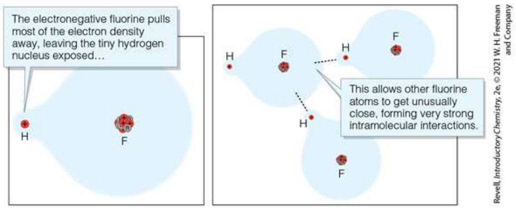
Hydrogen Bonding, Part 1

	Methane	Hydrogen cyanide	Water
Formula mass	16.0 u	27.0 u	18.0 u
Dipole strength*	0	2.98	1.85
Boiling point			

*These numbers convey the relative size of each dipole.

Hydrogen Bonding, Part 2

A strong intermolecular force between molecules containing H-F, H-O, or H-N bonds.



The electronegative fluorine pulls most of the electron density away, leaving the tiny hydrogen nucleus exposed…

This allows other fluorine atoms to get unusually close, forming very strong intramolecular interactions.

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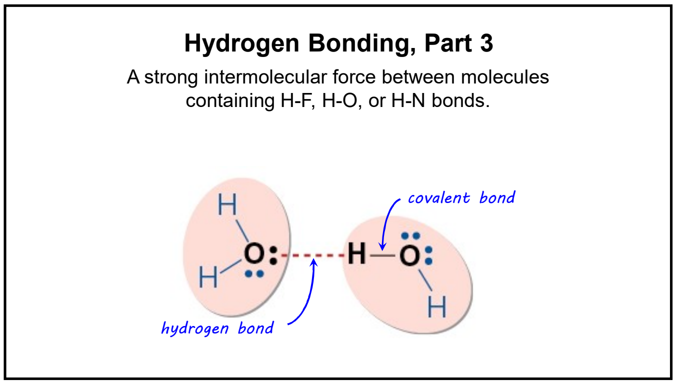
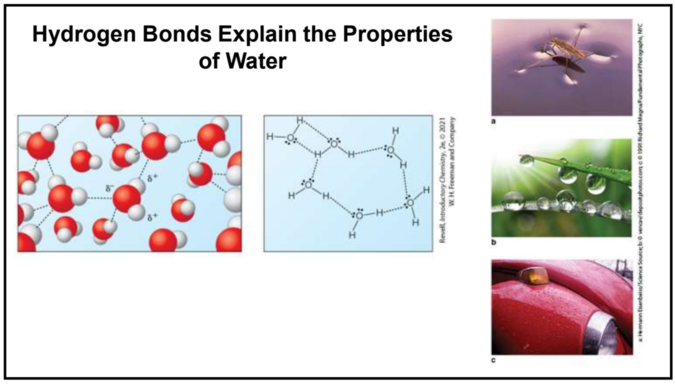


Hydrogen Bonding, Part 3

A strong intermolecular force between molecules containing H-F, H-O, or H-N bonds.



Hydrogen Bonds Explain the Properties of Water

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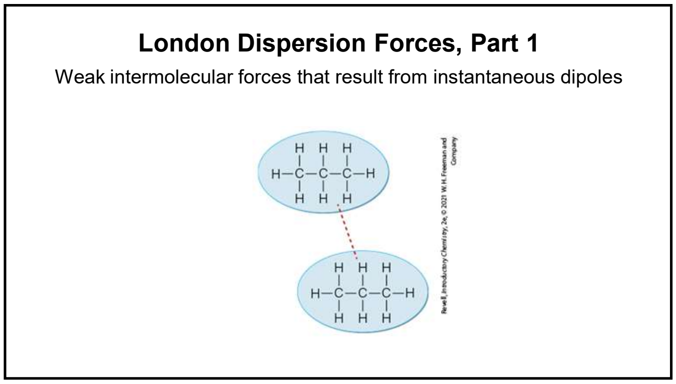

London Dispersion Forces, Part 1

Weak intermolecular forces that result from instantaneous dipoles

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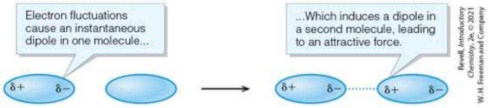
London Dispersion Forces, Part 2

Weak intermolecular forces that result from instantaneous dipoles



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Summary of Intermolecular Forces

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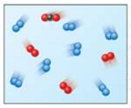
Type	Description	Strength
Hydrogen bonding	molecules with H-F, H-O, or H-N bonds	strongest
Dipole-dipole forces	molecules with net dipole	
London dispersion forces	all covalent molecules	weakest

Describing Gases

particles are spaced far apart
very little interaction between particles

Philip and Karen Smith/Getty Images



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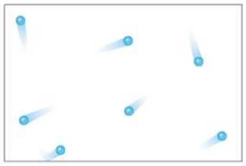
Ideal Gas

1. Volume of particles is much less than container.
2. Particles have no attraction for each other.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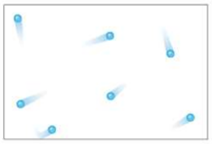
Temperature
Volume
Pressure

Pressure

The force that gases exert on their surroundings.

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$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Measuring Pressure


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Measuring Pressure – Barometers

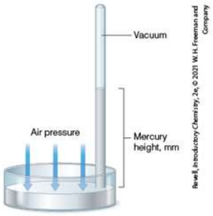
Barometer: a device used to measure atmospheric pressure

Millimeters of mercury (mm Hg)

1 mm Hg = 1 torr

Average air pressure at sea level:

760 mm Hg
760 torr
Standard pressure



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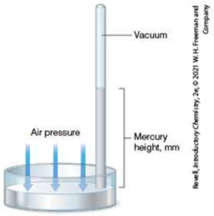
Measuring Pressure – Barometers Continued



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Measuring Pressure – Gauge Pressure

Gauge Pressure: The difference between the compressed gas pressure and the atmospheric pressure.

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Measuring Pressure – Conversion Factors

1 **atmosphere** (atm) = 760 mm Hg (torr)

1 atm = 14.70 pounds per square inch (psi)

1 atm = 101.3 kilopascals (kPa)

1 atm = 1.013 bar

The Gas Laws, Part 1

- Boyle's Law
- Charles's Law
- Combined Gas Law

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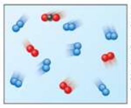

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Boyle's Law

The pressure and volume of a gas are inversely related.

$P\uparrow \quad V\downarrow$

$PV = \text{constant}$

$P_1V_1 = P_2V_2$

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Boyle's Law Practice

A commercial compressor stores 2.8 liters of air at a pressure of 150 psi. If this air is allowed to expand until the pressure is equal to 15 psi (just over atmospheric pressure), what volume will the air occupy?

$$P_1V_1 = P_2V_2$$

$P_1 = 150$ psi

$V_1 = 2.8$ L

$P_2 = 15$ psi

$V_2 = ?$

$$V_2 = \frac{P_1V_1}{P_2}$$

$$= \frac{(150\ \cancel{psi})(2.8\ L)}{(15\ \cancel{psi})} \boxed{= 28\ L}$$

Charles's Law

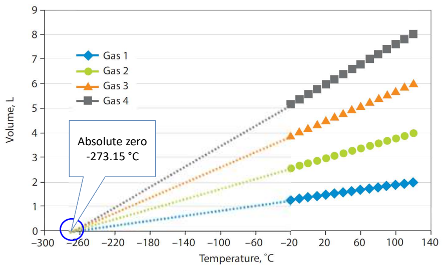
At constant pressure, the volume of a gas is directly proportional to its temperature.

$$T\uparrow \quad V\uparrow$$

$$V \propto T$$

$$\frac{V}{T} = \text{constant}$$

Using Charles's Law to Find Absolute Zero



The Kelvin Scale

Absolute zero

−273.15 °C

0 K

kelvin = °C + 273.15

Working to the nearest degree:

kelvin = °C + 273

Solving Problems with Charles's Law

$$\frac{V}{T} = \text{constant}$$

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

Charles's Law Practice

A balloon has a volume of 3.2 liters at room temperature (25 °C). The gas inside the balloon is then heated to 100 °C. What is the new volume of the balloon?

$V_1 = 3.2\ L$

$V_2 = ?$

$T_1 = 25\,°C + 273 = 298\ K$

$T_2 = 100\,°C + 273 = 373\ K$

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

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

$$= \frac{(3.2\ L)(373\ \cancel{K})}{(298\ \cancel{K})} \boxed{= 4.0\ L}$$

The Combined Gas Law

Photo credit: Carolyn Franks/Alamy

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

The Combined Gas Law Practice

A gas with a temperature of 280 K, a pressure of 200 kPa, and a volume of 25.8 L is compressed to 15.8 L, causing the pressure to increase to 350 kPa. What is the temperature of the gas under the new conditions?

P_1 = 200 kPa

V_1 = 25.8 L

T_1 = 280 K

P_2 = 350 kPa

V_2 = 15.8 L

T_2 = ?

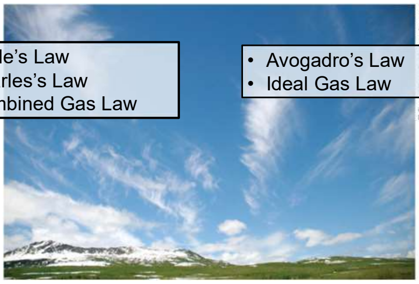
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$= \frac{(350\ \cancel{kPa})(15.8\ \cancel{L})(280\ K)}{(200\ \cancel{kPa})(25.8\ \cancel{L})}$$

$$= \boxed{300\ K}$$

The Gas Laws, Part 2



- Boyle’s Law
- Charles’s Law
- Combined Gas Law

- Avogadro’s Law
- Ideal Gas Law

Avogadro's Law

If temperature and pressure are constant, the volume of a gas is proportional to the number of moles of gas present.

$$V \propto n$$

at Standard Temperature and Pressure (STP)…

T = 0 °C (273 K)
P = 1.0 atm

…1 mole of gas occupies 22.4 Liters

The Ideal Gas Law

$$PV = nRT$$

- R = 0.0821 L·atm/mol·K
- *T* must be in kelvins
- *P*, *V* units must match gas constant

The Ideal Gas Law Practice

What volume does 1.00 mole of gas occupy at a temperature of 0.00 °C and a pressure of 1.00 atmospheres?

$$PV = nRT$$

$$T = 0.00\ °C$$
$$= 273.15\ K$$

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

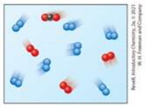
$$= \frac{(1.00\ \cancel{mol})(0.0821\ L\cdot\cancel{atm}/\cancel{mol}\cdot\cancel{K})(273.15}{1.00\ \cancel{atm}}$$

$$\boxed{= 22.4\ L}$$

Mixtures of Gases: Partial Pressure

partial pressure: The pressure caused by one gas in a mixture.

Adding up all partial pressures gives the total pressure.



Air:
78% nitrogen
21% oxygen

Partial Pressure Practice

If a 40.0-L cylinder is filled with 5.00 moles of nitrogen, 2.00 moles of oxygen, and 3.00 moles of carbon dioxide at a temperature of 400 K, what is the pressure inside the cylinder?

$$P_{N2} = \frac{nRT}{V} = \frac{(5.00\ \text{mol})(0.0821\ \text{L·atm/mol·K})(400\ \text{K})}{40.0\ \text{L}} = 4.11\ \text{atm}$$

$$P_{O2} = \frac{nRT}{V} = \frac{(2.00\ \text{mol})(0.0821\ \text{L·atm/mol·K})(400\ \text{K})}{40.0\ \text{L}} = 1.64\ \text{atm}$$

$$P_{CO2} = \frac{nRT}{V} = \frac{(3.00\ \text{mol})(0.0821\ \text{L·atm/mol·K})(400\ \text{K})}{40.0\ \text{L}} = 2.46\ \text{atm}$$

$$P_{Total} = P_{N2} + P_{O2} + P_{CO2}$$

$$= 4.11\ \text{atm} + 1.64\ \text{atm} + 2.46\ \text{atm}\ \boxed{= 8.21\ \text{atm}}$$

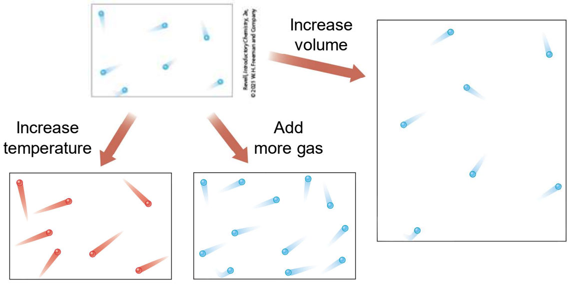
Partial Pressure, More Practice

If a 40.0-L cylinder is filled with 5.00 moles of nitrogen, 2.00 moles of oxygen, and 3.00 moles of carbon dioxide at a temperature of 400 K, what is the pressure inside the cylinder?

$$n_{total} = 5.00\ \text{moles} + 2.00\ \text{moles} + 3.00\ \text{moles} = 10.00\ \text{moles total}$$

$$P_{total} = \frac{nRT}{V} = \frac{(10.00\ \text{mol})(0.0821\ \text{L·atm/mol·K})(400\ \text{K})}{40.0\ \text{L}}\ \boxed{= 8.21\ \text{atm}}$$

A Molecular View of the Gas Laws



Increase volume

Increase temperature

Add more gas

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Diffusion

The spread of particles through random motion.

Lighter particles diffuse more quickly.

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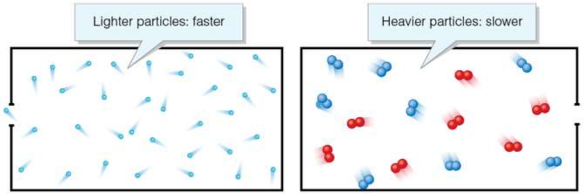
Effusion

The process of a gas escaping from a container.



Lighter particles: faster

Heavier particles: slower

Helium

Air

Particles leak out more quickly

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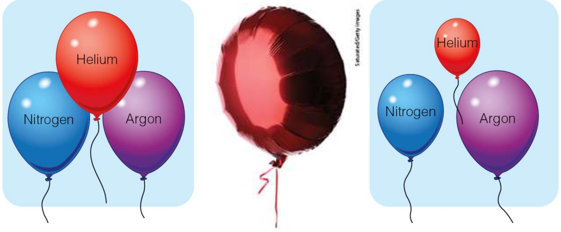
Effusion, Continued



Saturated/Getty Images

Gas Stoichiometry, Part 1

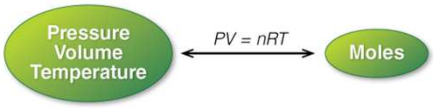
$C_6H_{12}O_6\,(s) \rightarrow 2\,C_2H_6O\,(l) + 2\,CO_2\,(g)$

Scott Eisen/Bloomberg via Getty Images

Stoichiometry

Gas Laws

Gas Stoichiometry, Part 2



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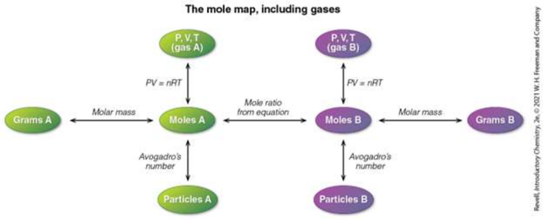
Gas Stoichiometry, Part 3



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Gas Stoichiometry Practice

In the fermentation of glucose, how many moles of carbon dioxide are produced for each kilogram of glucose that reacts? If the reaction takes place in a sealed container and the gas occupies a volume of 8.10 liters at a temperature of 21 °C, find the pressure of the carbon dioxide gas inside the container.

$$C_6H_{12}O_6\ (s) \rightarrow 2\ C_2H_6O\ (l) + 2\ CO_2\ (g)$$

g $C_6H_{12}O_6$ ⇨ Moles $C_6H_{12}O_6$ ⇨ Moles CO_2

$$1{,}000\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ \cancel{mol\ C_6H_{12}O_6}}{180.18\ \cancel{g\ C_6H_{12}O_6}} \times \frac{2\ mol\ CO_2}{1\ \cancel{mol\ C_6H_{12}O_6}} = 11.10\ mol\ CO_2$$

Moles CO_2 ⇨ Pressure CO_2

$$P = \frac{nRT}{V} = \frac{(11.10\ \cancel{mol}\ CO_2)(0.0821\ \cancel{L}\cdot atm/\cancel{mol\cdot K})(294\ \cancel{K})}{8.10\ \cancel{L}} \boxed{= 33.1\ atm\ CO_2}$$

$$T = 21°C + 273 = 294\ K$$

Gas Stoichiometry, More Practice

Natural gas burns cleanly in air, according to this equation:

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

If 13.1 liters of CH_4 burn at a pressure of 1.00 atmosphere and a temperature of 290 K, what mass of carbon dioxide gas is produced?

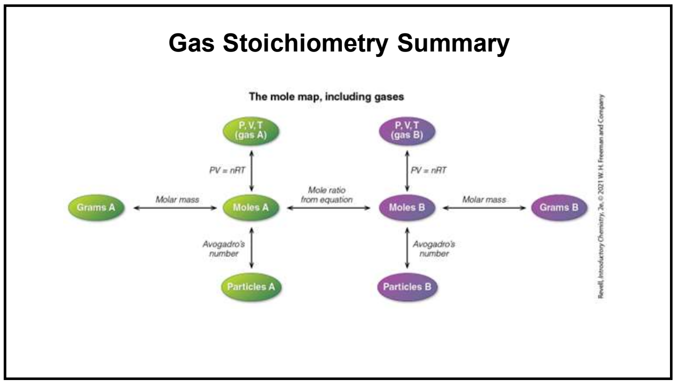
Vitaliy Maselko/Alamy

Pressure, Volume CH_4 ⇨ Moles CH_4

$$n = \frac{PV}{RT} = \frac{(1.00\ \cancel{atm})(13.1\ \cancel{L})}{(0.0821\ \cancel{L\cdot atm}/mol\cdot\cancel{K})(290\ \cancel{K})} = 0.550\ moles\ CH_4$$

Moles CH_4 ⇨ Moles CO_2 ⇨ Grams CO_2

$$0.550\ \cancel{mol\ CH_4} \times \frac{1\ \cancel{mol\ CO_2}}{1\ \cancel{mol\ CH_4}} \times \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} \boxed{= 24.2\ g\ CO_2}$$



Gas Stoichiometry Summary

The mole map, including gases

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