

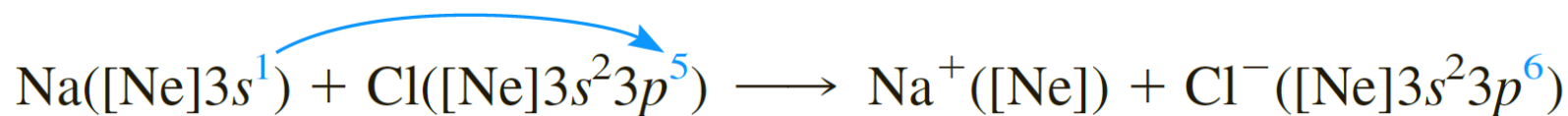
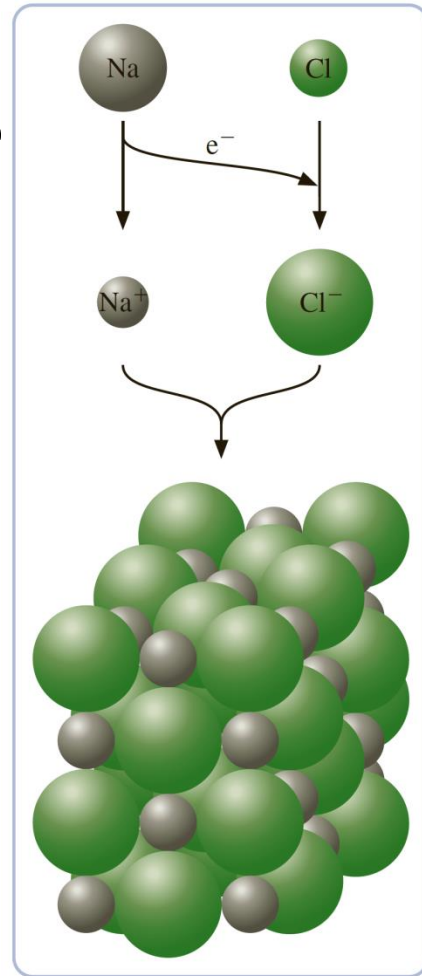
EBBING - GAMMON

General
Chemistry
ELEVENTH EDITION

Ionic and Covalent Bonding

➤ Ionic Bonds 9.1 Describing Ionic Bonds

- ✓ An **ionic bond** is a chemical bond formed by the electrostatic attraction between positive and negative ions.
- ✓ The bond forms between two atoms when one or more electrons are transferred from the valence shell of one atom to the valence shell of the other.
- ✓ The atom that loses electrons becomes a *cation* (positive ion), and the atom that gains electrons becomes an *anion* (negative ion).
- ✓ As a result of the electron transfer, ions are formed, each of which has a noble-gas configuration.



Low ionization E High electron affinity

➤ Lewis Electron-Dot Symbols

✓ is a symbol in which the electrons in the valence shell of an atom or ion are represented by dots placed around the letter symbol of the element

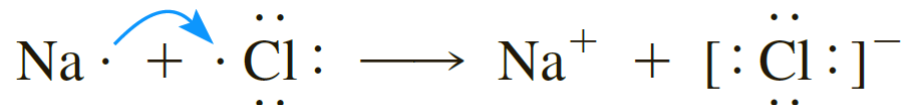
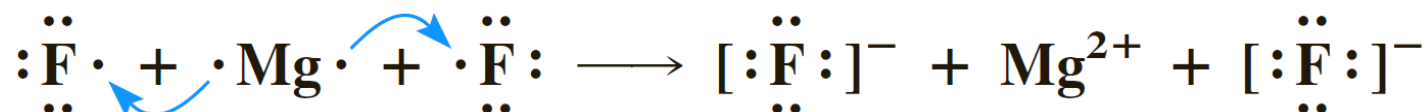


Table 9.1 Lewis Electron-Dot Symbols for Atoms of the Second and Third Periods

	1A ns^1	2A ns^2	3A ns^2np^1	4A ns^2np^2	5A ns^2np^3	6A ns^2np^4	7A ns^2np^5	8A ns^2np^6
Period								
Second	Li·	·Be·	·B·	·C·	:N·	:O·	:F·	:Ne:
Third	Na·	·Mg·	·Al·	·Si·	:P·	:S·	:Cl·	:Ar:

(Q) Use Lewis electron-dot symbols to represent the transfer of electrons from magnesium to fluorine atoms to form ions with noble-gas configurations



➤ Energy Involved in Ionic Bonding

✓ Formation of an ionic bond between a sodium atom and a chlorine atom:



✓ The overall energy is $(496 - 349) = 147 \text{ kJ/mol}$

→ the process requires more energy to remove an electron from the sodium atom than is gained when the electron is added to the chlorine atom.

→ formation of ions from the atoms is not in itself energetically favorable.

BUT When positive and negative ions bond → energy is released to make the overall process favorable.

Coulomb's law
$$E = \frac{kQ_1Q_2}{r}$$

Coulomb's law states that *the potential energy obtained in bringing two charges Q_1 and Q_2 , initially far apart, up to a distance r apart is directly proportional to the product of the charges and inversely proportional to the distance between them.*

$$E = \frac{kQ_1Q_2}{r}$$

$$k = 8.99 \times 10^9 \text{ J}\cdot\text{m}/\text{C}^2$$

The charge on Na^+ is $+e$ and that on Cl^- is $-e$.

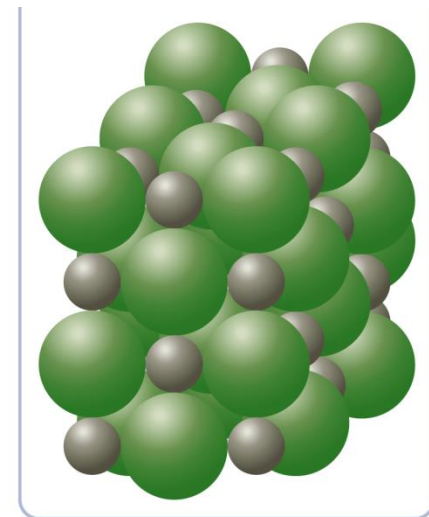
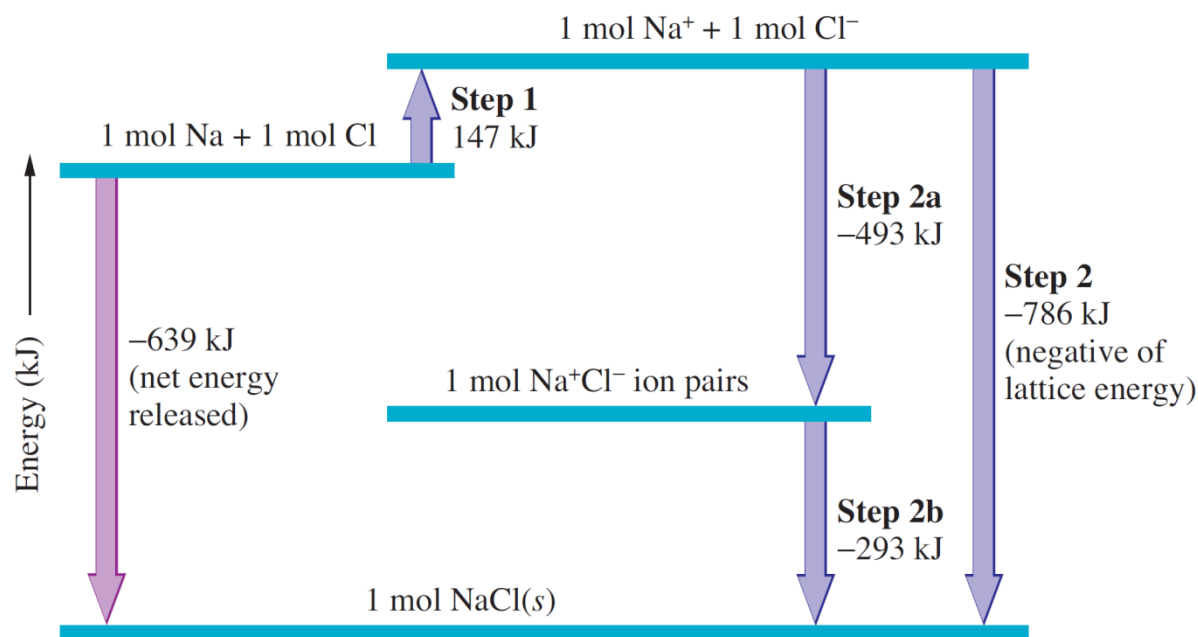
$$e = 1.602 \times 10^{-19} \text{ C}$$

r = distance between Na^+ and Cl^- = 282 pm, or $2.82 \times 10^{-10} \text{ m}$.

$$E = \frac{-(8.99 \times 10^9 \text{ J}\cdot\text{m}/\text{C}^2) \times (1.602 \times 10^{-19} \text{ C})^2}{2.82 \times 10^{-10} \text{ m}} = -8.18 \times 10^{-19} \text{ J}$$

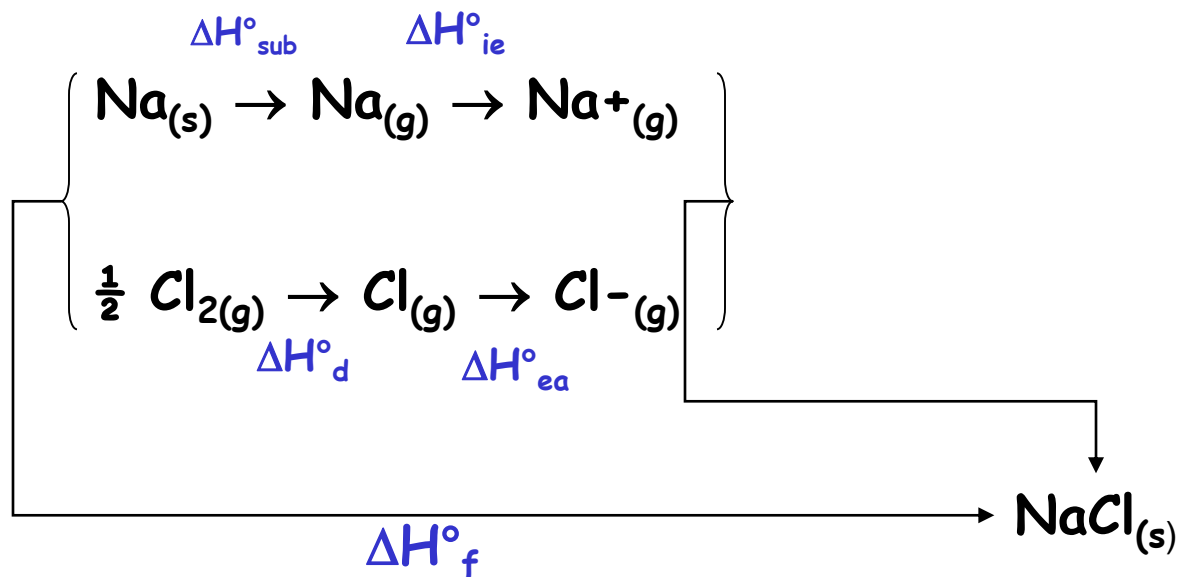
- ✓ The minus sign means energy is released
- ✓ This energy is for the formation of one ion pair
- ✓ Multiplying by Avogadro's number, $6.02 \times 10^{23} \rightarrow -493 \text{ kJ/mol}$

- ✓ The **lattice energy** is *the change in energy that occurs when an ionic solid is separated into isolated ions in the gas phase*. For sodium chloride, the process is



- ✓ The negative sign shows that there has been a net decrease in energy, which you expect when stable bonding has occurred.
- ✓ Ionic bond forms between elements if the ionization energy of one is sufficiently small and the electron affinity of the other is sufficiently large

The Born-Haber Cycle for NaCl (Energy diagram)



$$\Delta H^\circ_{\text{f}} = \Delta H^\circ_{\text{sub}} + \Delta H^\circ_{\text{ie}} + \frac{1}{2} \Delta H^\circ_{\text{d}} + \Delta H^\circ_{\text{ea}} + U_o \quad U_o = \text{Lattice Energy}$$

$$-411 = 109 + 496 + \frac{1}{2} (242) + (-349) + U_o$$

$$U_o = -788 \text{ kJ/mol}$$

You must use the correct stoichiometry and signs to obtain the correct lattice energy.

➤ Properties of Ionic Substances

- ✓ Strong ionic bonds (strong electrostatic interaction) → high-melting points of ionic solids.

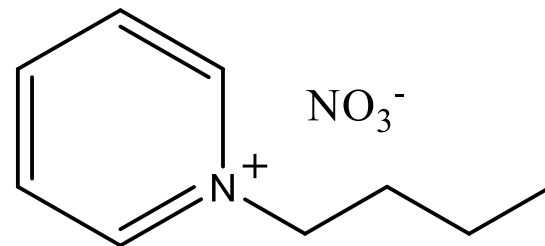
m.p of MgO (2800 °C) > m.p NaCl (801 °C)

charges (Mg²⁺ and O²⁻),

charges (Na⁺ and Cl⁻),

$$E = \frac{kQ_1Q_2}{r}$$

- ✓ The liquid melt from an ionic solid consists of ions, and so the liquid melts conducts an electric current.
- ✓ Ionic liquids have low m.p (RT) because of the cations are large and non-spherical.



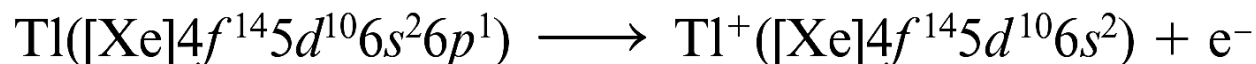
9.2 Electron Configurations of Ions

➤ Ions of the Main-Group Elements

Table 9.2 Ionization Energies of Na, Mg, and Al (in kJ/mol)*

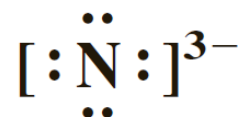
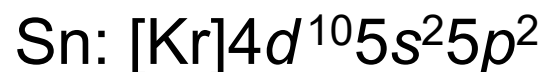
Element	Successive Ionization Energies			
	First	Second	Third	Fourth
Na	496	4,562	6,910	9,543
Mg	738	1,451	7,733	10,542
Al	578	1,817	2,745	11,577

- ✓ Valence electrons are easily removed
- ✓ Much higher energy is needed to remove further electrons.
- No compounds are found with ions having charges greater than the group number. **e.g : Na²⁺, Mg³⁺, Al⁴⁺ (Doesn't exist)**
- ✓ Boron (Group 3A) doesn't form ionic compounds with B³⁺ ions, the bonding is normally covalent.
- ✓ The remaining elements of Group 3A do form compounds containing 3⁺ ions because of decreasing ionization energy.
- ✓ Thallium in 3A, Period 6, has compounds with 1⁺ ions and compounds with 3⁺ ions



- ✓ The first three elements of Group 4A (C, Si, and Ge) are metalloids and usually form covalent rather than ionic bonds.
- ✓ Tin (Sn) and lead (Pb) (group 4A) commonly form ionic compounds with 2^+ ions.
- ✓ Tin forms tin(II) chloride, SnCl_2 , which is an ionic compound and tin(IV) chloride SnCl_4 which is a covalent compound.
- ✓ Bi (group 5A) forms ionic Bi^{3+} cpds and covalent Bi^{5+} cpds.
- ✓ Anions of Groups 5A to 7A gain electrons (large EA) to form noble-gas or pseudo-noble-gas configurations.
- ✓ Hydrogen forms compounds of the 1^- ion, H^- (hydride ion).
- ✓ Although the electron affinity of nitrogen ($2s^2 2p^3$) = 0 N^{3-} ion ($2s^2 2p^6$) is stable in the presence of Li^+ (Li_3N) and other alkaline earth elements ions (Mg_3N_2).

(Q) Write the electron configuration and the Lewis symbol for N^{3-} .



➤ Transition-Metal Ions

- ✓ M^{2+} is a common oxidation state as two electrons are removed from the outer ***ns*** shell. Fe: [Ar] $4s^2 3d^6$

Fe²⁺ : [Ar] 3d⁶ loses 4s electrons first

$\text{Fe}^{3+} : [\text{Ar}] 3d^5$ then loses $3d$ electrons

(Q) What are the correct electron configurations for Cu & Cu²⁺ ?

A. $[\text{Ar}] 4s^2 3d^9$, $[\text{Ar}] 3d^9$

B. $[\text{Ar}] 3d^{10}4s^1$, $[\text{Ar}] 4s^13d^8$

C. $[\text{Ar}] 3d^{10}4s^1$, $[\text{Ar}] 3d^9$

D. $[\text{Ar}] 4s^2 3d^9$, $[\text{Ar}] 3d^{10} 4s^1$

E. [K] $4s^23d^9$, [Ar] $3d^9$

(Q) What are the correct electron configurations for zirconium(II) and zirconium(IV) ions?

A. $[\text{Kr}] 5d^2$ $[\text{Kr}] 4d^1$

B. $[\text{Ar}] 4d^2$ $[\text{Ar}] 5s^2$

C. $[\text{Kr}] 5s^2 4d^2$ $[\text{Kr}]$

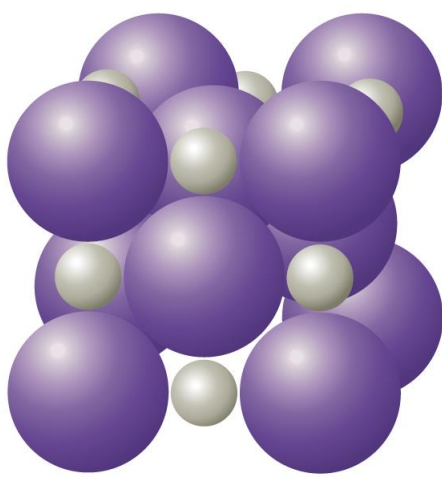
D. $[\text{Kr}] 5s^2 4d^6$ $[\text{Kr}] 4d^6$

E. [Rb] $4d^2$ [Rb]

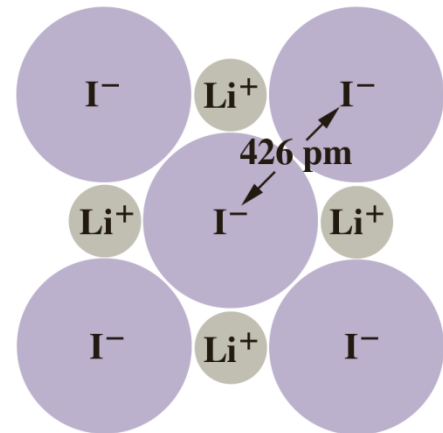
I A																										VIII A							
1 H Hydrogen 1.008																										2 He Helium 4.002602							
3 Li Lithium 6.94		4 Be Beryllium 9.007937												5 B Boron 10.81		6 C Carbon 12.011		7 N Nitrogen 14.007		8 O Oxygen 15.999		9 F Fluorine 18.99847363		10 Ne Neon 20.1797									
11 Na Sodium 22.98976928		12 Mg Magnesium 24.305		3 IIIB		4 IVB		5 VB		6 VIB		7 VIIB		8 VIIIB		9 VIIIB		10 IIIB		11 IB		12 IIB											
13 Al Aluminum 26.9815385		14 Si Silicon 28.085		15 P Phosphorus 30.973761998		16 S Sulfur 32.06		17 Cl Chlorine 35.45		18 Ar Argon 39.948																							
19 K Potassium 39.0983		20 Ca Calcium 40.078		21 Sc Scandium 44.955909		22 Ti Titanium 47.867		23 V Vanadium 50.9415		24 Cr Chromium 51.9961		25 Mn Manganese 54.938044		26 Fe Iron 55.845		27 Co Cobalt 58.933194		28 Ni Nickel 58.6934		29 Cu Copper 63.546		30 Zn Zinc 65.38											
31 Ga Gallium 69.723		32 Ge Germanium 72.630		33 As Arsenic 74.921595		34 Se Selenium 78.971		35 Br Bromine 79.904		36 Kr Krypton 83.798																							
37 Rb Rubidium 85.4678		38 Sr Strontium 87.62		39 Y Yttrium 88.90584		40 Zr Zirconium 91.224		41 Nb Niobium 92.90637		42 Mo Molybdenum 95.94		43 Tc Technetium (98)		44 Ru Ruthenium 101.07		45 Rh Rhodium 102.90550		46 Pd Palladium 106.42		47 Ag Silver 107.8682		48 Cd Cadmium 112.411											
49 In Indium 114.818		50 Sn Tin 118.710		51 Sb Antimony 121.757		52 Te Tellurium 127.60		53 I Iodine 126.90547		54 Xe Xenon 131.293																							
55 Cs Caesium 132.90545196		56 Ba Barium 137.327		57 - 71 Lanthanoids		72 Hf Hafnium 178.49		73 Ta Tantalum 180.94788		74 W Tungsten 183.84		75 Re Rhenium 186.207		76 Os Osmium 190.23		77 Ir Iridium 192.222		78 Pt Platinum 195.084		79 Au Gold 196.966569		80 Hg Mercury 200.59											
81 Tl Thallium 204.3833		82 Pb Lead 207.2		83 Bi Bismuth 208.9804		84 Po Polonium 209		85 At Astatine 210		86 Rn Radon 222																							

➤ 9.3 Ionic Radii

Determining the iodide ion radius in the lithium iodide (LiI) crystal







a A three-dimensional view of the crystal.



Ionic radius of I^-
 $= 426 / 2 = 213 \text{ pm}$

Exercise 9.6
arrange the following ions in order of increasing ionic radius: Sr^{2+} , Mg^{2+} , Ca^{2+} .

$Mg^{2+} < Ca^{2+} < Sr^{2+}$

 Na [He] $2s^2 2p^6 3s^1$	 Na^+ [He] $2s^2 2p^6$
 Cl [Ne] $3s^2 3p^5$	 Cl^- [Ne] $3s^2 3p^6$

✓ Ionic radii increase down any column because of the addition of electron shells.

Table 9.3 Ionic Radii (in pm) of Some Main-Group Elements					
Period	1A	2A	3A	6A	7A
2	Li ⁺	Be ²⁺		O ²⁻	F ⁻
	60	31		140	136
3	Na ⁺	Mg ²⁺	Al ³⁺	S ²⁻	Cl ⁻
	95	65	50	184	181
4	K ⁺	Ca ²⁺	Ga ³⁺	Se ²⁻	Br ⁻
	133	99	62	198	195
5	Rb ⁺	Sr ²⁺	In ³⁺	Te ²⁻	I ⁻
	148	113	81	221	216
6	Cs ⁺	Ba ²⁺	Tl ³⁺		
	169	135	95		

1 IA												18 VIIIA					
1 H Hydrogen 1.008	2 IIA										2 He Helium 4.002602						
3 Li Lithium 6.94	4 Be Beryllium 9.0121831											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998403163	10 Ne Neon 20.1797
11 Na Sodium 22.98976928	12 Mg Magnesium 24.305											13 Al Aluminium 26.9815385	14 Si Silicon 28.085	15 P Phosphorus 30.973761998	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.948
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955908	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938044	26 Fe Iron 55.845	27 Co Cobalt 58.933194	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.630	33 As Arsenic 74.921595	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90584	40 Zr Zirconium 91.224	41 Nb Niobium 92.90637	42 Mo Molybdenum 95.95	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.90550	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	52 Te Tellurium 127.60	53 I Iodine 126.90447	54 Xe Xenon 131.293

➤ Pattern across a period

Cation	Na ⁺	Mg ²⁺	Al ³⁺	Anion	S ²⁻	Cl ⁻
Radius (pm)	95	> 65	> 50	Radius (pm)	184	181

- ✓ All of these cations have Ne configuration 1s²2s²2p⁶ but different nuclear charges (they are isoelectronic).
- ✓ **Isoelectronic** refers to different species having the same number and configuration of electrons

9.47 Arrange the following in order of increasing ionic radius:

As³⁻, Se²⁻, Br⁻

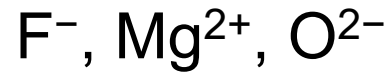
Br⁻ < Se²⁻ < As³⁻

Or As³⁻ > Se²⁻ > Br⁻

✓ **Within an isoelectronic series, the radius of increases as the atomic number decreases**

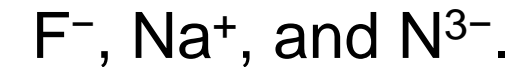
Or $\text{As}^{3-} > \text{Se}^{2-} > \text{Br}^-$ number decreases																			
1 IA																		18 VIIIA	
1 H Hydrogen 1.008	2 He Helium 4.002602																		
3 Li Lithium 6.94	4 Be Beryllium 9.0121831																		
11 Na Sodium 22.98976928	12 Mg Magnesium 24.305																		
		13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA													
		5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998403163	10 Ne Neon 20.1797												
		13 Al Aluminium 26.9815385	14 Si Silicon 28.085	15 P Phosphorus 30.973761998	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.948												
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(Q) arrange the following ions in order of decreasing ionic radius:



isoelectronic series $\rightarrow \text{Mg}^{2+} (Z=12) < \text{F}^- (Z=9) < \text{O}^{2-} (Z=8)$

9.49 Arrange the following in order of increasing ionic radius:



isoelectronic series $\rightarrow \text{Na}^+ (Z=11) < \text{F}^- (Z=9) < \text{N}^{3-} (Z=7)$

9.48 Which has the larger radius, N^{3-} or P^{3-} ? **P^{3-}**
NOT isoelectronic

1 IA												18 VIIIA						
1 H Hydrogen 1.008																		2 He Helium 4.002602
3 Li Lithium 6.94	4 Be Beryllium 9.0121831											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998403163	10 Ne Neon 20.1797	
11 Na Sodium 22.98976928	12 Mg Magnesium 24.305	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 Al Aluminium 26.9815385	14 Si Silicon 28.085	15 P Phosphorus 30.973761998	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.948	
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955908	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938044	26 Fe Iron 55.845	27 Co Cobalt 58.933194	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.630	33 As Arsenic 74.921595	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 83.798	
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➤ Covalent Bonds

- ✓ a chemical bond formed by the sharing of a pair of electrons between atoms.

9.4 Describing Covalent Bonds

- ✓ The distance between nuclei at minimum energy is called the *bond length* of H_2 .

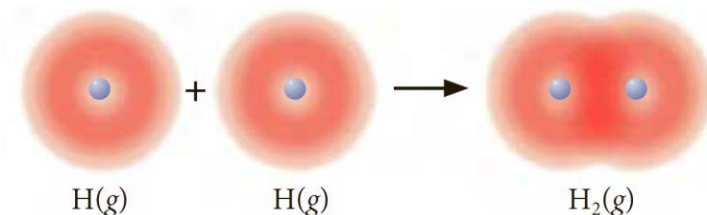


Figure 9.10 ▲

The electron probability distribution for the H_2 molecule

The electron density (shown in red) occupies the space around both atoms.

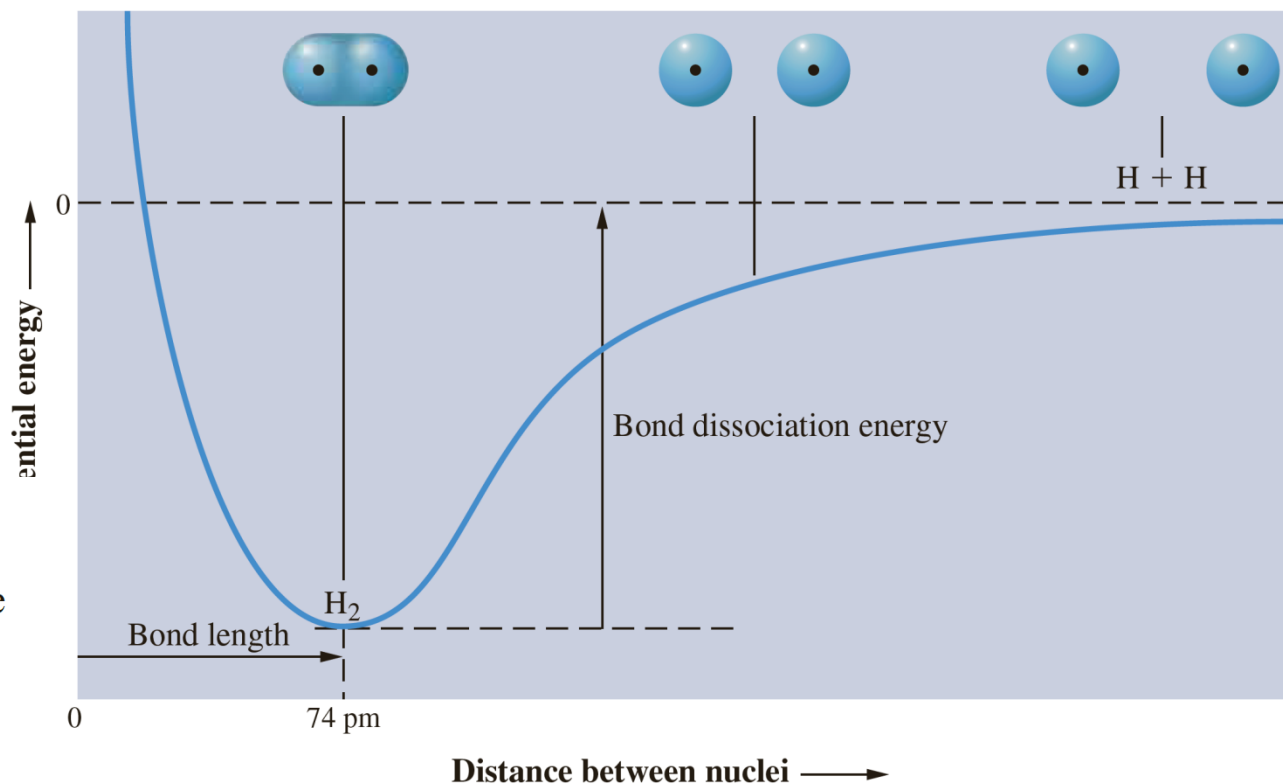
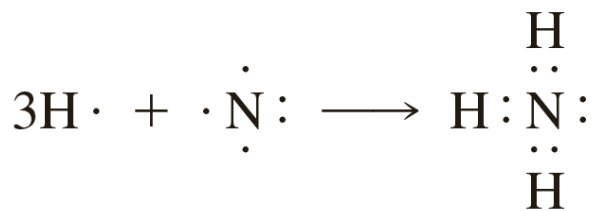
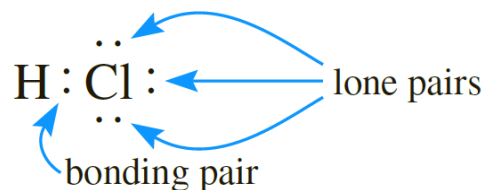
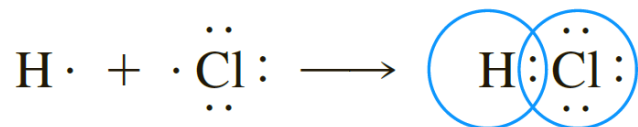
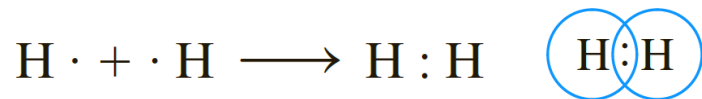


Figure 9.11 ◀

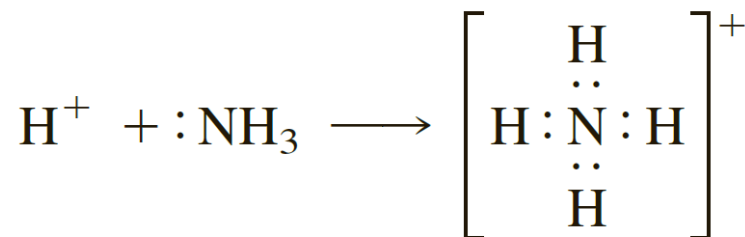
Potential-energy curve for H_2 The stable molecule occurs at the bond distance corresponding to the minimum in the potential-energy curve.

➤ Lewis Formulas



➤ Coordinate Covalent Bonds

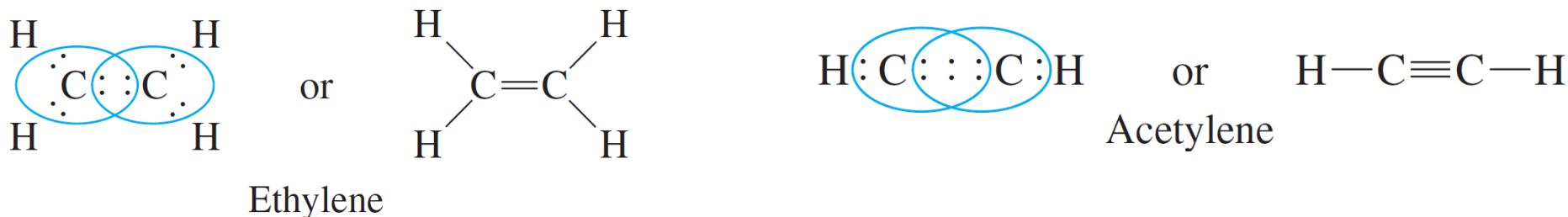
✓ *is a bond formed when both electrons of the bond are donated by one atom*



➤ Octet Rule

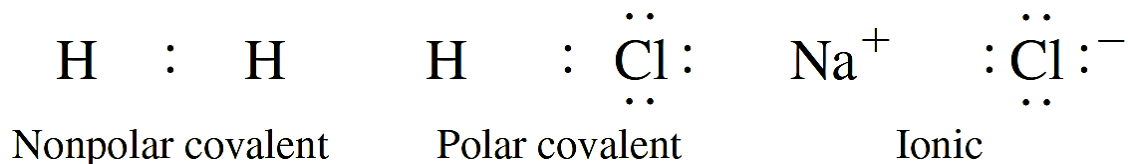
- ✓ *The tendency of atoms in molecules to have eight electrons in their valence shells (two for hydrogen atoms)*

➤ Multiple Bonds

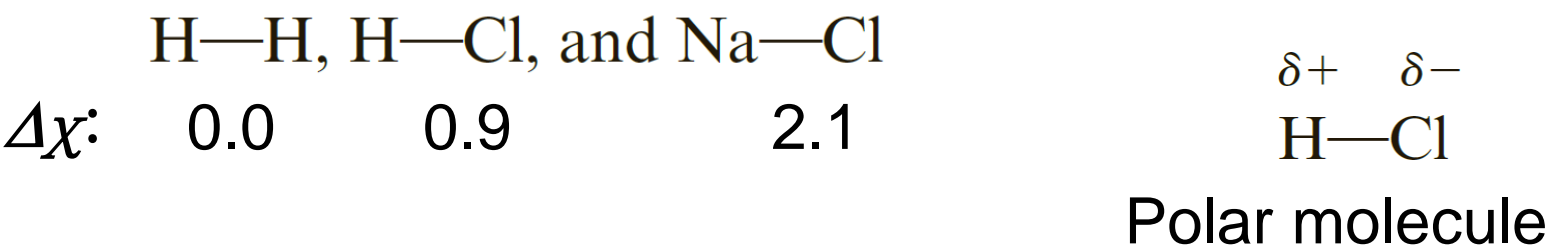


9.5 Polar Covalent Bonds (Polar Bonds)

- ✓ *is a covalent bond in which the bonding electrons spend more time near one atom than the other.*



- **Electronegativity** is a measure of the ability of an atom in a molecule to draw bonding electrons to itself.
- ✓ Mulliken electronegativity (χ):
$$\chi = \frac{I.E. + E.A.}{2}$$
- ✓ F has large $E.A.$ and large $I.E.$ → large electronegativity
- ✓ Li has small $E.A.$ and small $I.E.$ → small electronegativity
- ✓ Pauling's electronegativity (χ): depends on bond enthalpies
- ✓ *Electronegativity increases from left to right and decreases from top to bottom in the periodic table.*
- ✓ Metals are the least electronegative elements (they are *electropositive*) and nonmetals the most electronegative.
- ✓ The absolute value of the difference in electronegativity of two bonded atoms gives a rough measure of the **polarity** of a bond



➤ Writing Lewis Electron-Dot Formulas

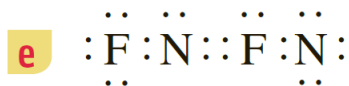
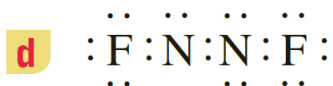
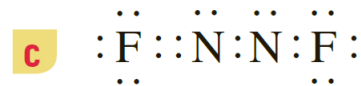
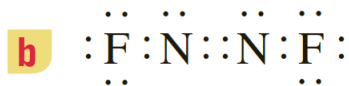
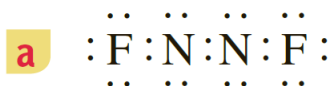
These will be done in class:

H_2O , NF_3 , CCl_2F_2 , CO_2 , SCl_2 , POCl_3 , COCl_2 , HSO_3Cl ,
 CO_3^{2-} , NH_4^+ , BF_4^- , H_3O^+ , ClO_2^- .



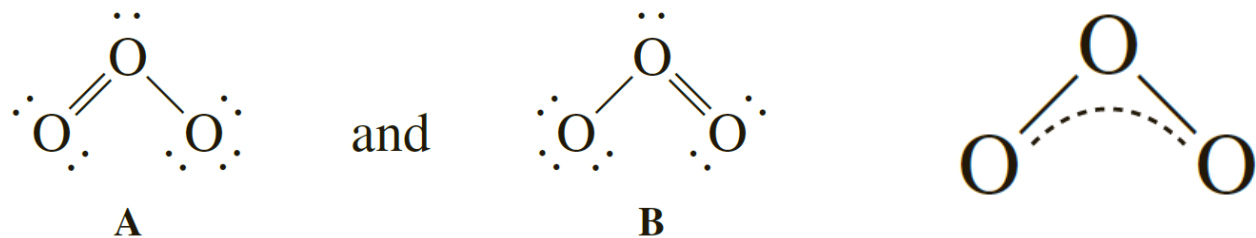
CONCEPT CHECK 9.2

Each of the following may seem, at first glance, to be plausible electron-dot formulas for the molecule N_2F_2 . Most, however, are incorrect for some reason. What concepts or rules apply to each, either to cast it aside or to keep it as the correct formula?



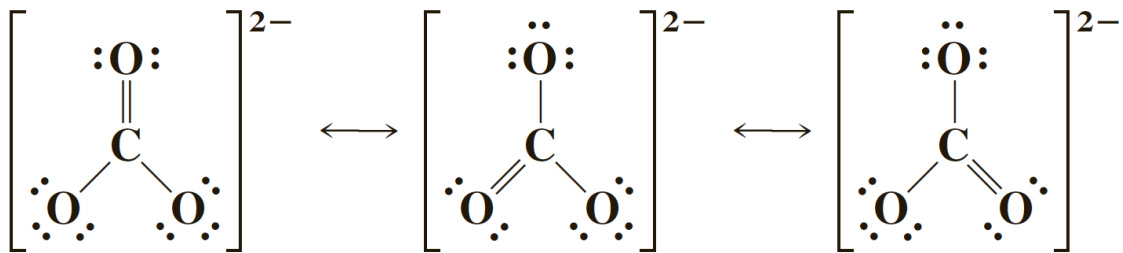
9.7 Delocalized Bonding: Resonance

Ozone (O₃)



- ✓ The lengths of the two oxygen–oxygen bonds (that is, the distances between the atomic nuclei) are both 128 pm.
- ✓ **delocalized bonding**

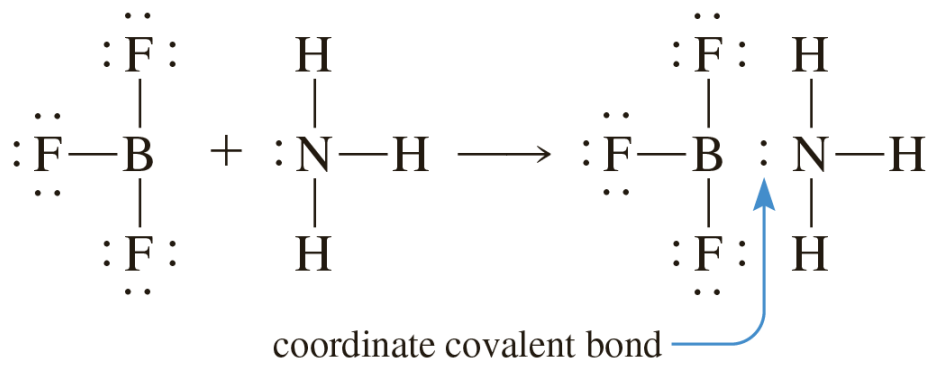
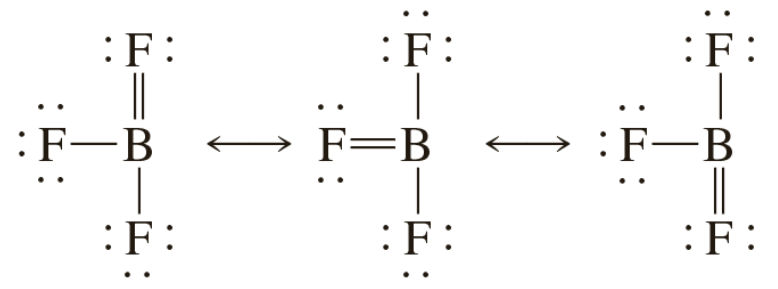
CO₃²⁻



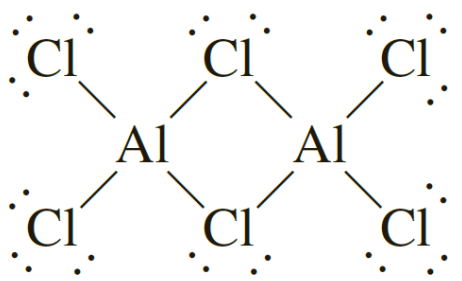
NO₃⁻

9.8 Exceptions to the Octet Rule

These will be done in class:

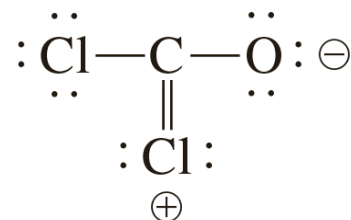
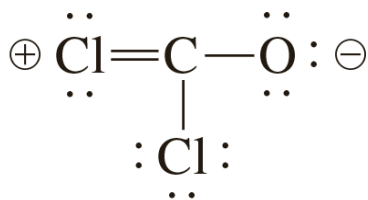
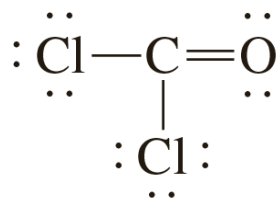
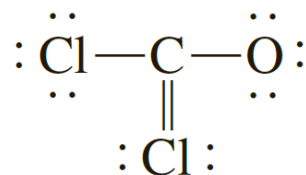
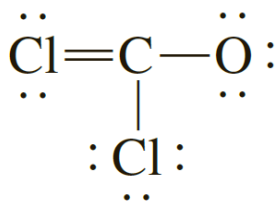
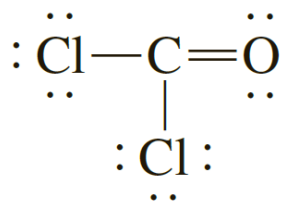


AlCl_3 @ RT & at melting point (very low 192°C)



two of the Cl atoms are in **bridge positions**

9.9 Formal Charge and Lewis Formulas



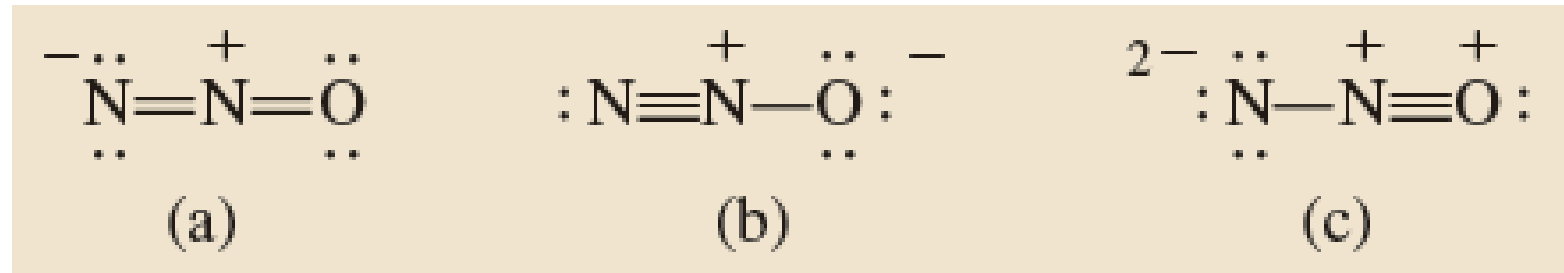
RULE A Whenever you can write several Lewis formulas for a molecule, choose the one having the lowest magnitudes of formal charges.

RULE B When two proposed Lewis formulas for a molecule have the same magnitudes of formal charges, choose the one having the negative formal charge on the more electronegative atom.

RULE C When possible, choose Lewis formulas that do not have like charges on adjacent atoms.

(Q) Write the Lewis formula that best describes the charge distribution in the sulfuric acid molecule, H₂SO₄, according to the rules of formal charge. (HNO₃, H₃PO₄, HCN)

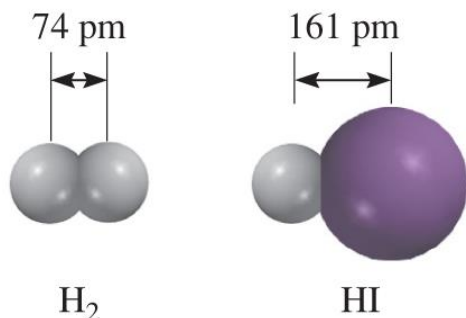
(Q) Draw three resonance structures for the molecule nitrous oxide, N_2O (the atomic arrangement is NNO)



Structure (b) is the most important one because the negative charge is on the more electronegative oxygen atom.

Structure (c) is the least important one because it has a larger separation of formal charges. Also, the positive charge is on the more electronegative oxygen atom.

➤ 9.10 Bond Length and Bond Order



covalent radius:

Covalent radius of an atom X = half of the covalent bond length of a homonuclear X-X single bond.

If covalent radius of (C = 76 pm) & (Cl = 102 pm) \rightarrow bond length of C-Cl = $(76 + 102) = 178$ pm

chloromethane, CH_3Cl , 178.4 pm;

tetrachloromethane, CCl_4 , 176.6 pm;

Average Bond Lengths of Some Common Single, Double, and Triple Bonds

Bond Type	Bond Length (pm)
C—H	107
C—O	143
C=O	121
C—C	154
C=C	133
C \equiv C	120
C—N	143
C=N	138
C \equiv N	116
N—O	136
N=O	122
O—H	96

Bond lengths:

Triple bond < Double Bond < Single Bond

3

2

1

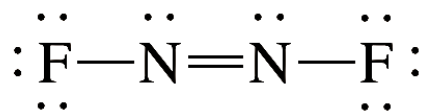
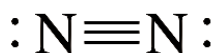
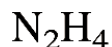
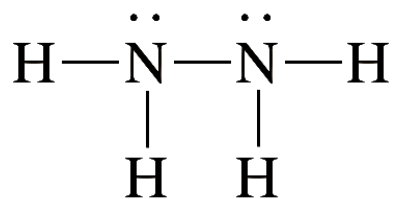
➤ Trends for atomic radii

1. Within a period, the covalent radius tends to decrease with increasing atomic number.
2. Within a group, the covalent radius tends to increase with period number.

(Q) Consider the molecules N_2H_4 , N_2 , and N_2F_2 .

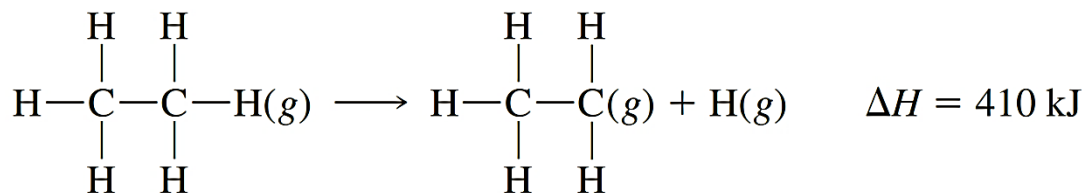
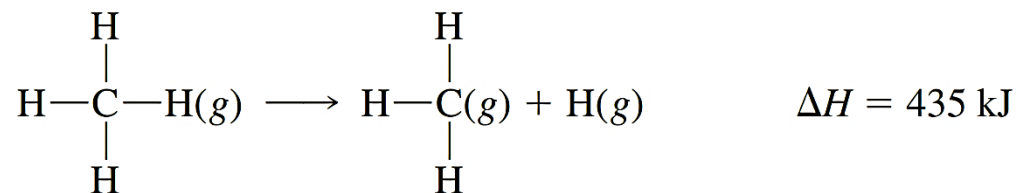
Which molecule has the shortest nitrogen–nitrogen bond?

Which has the longest nitrogen–nitrogen bond?



9.11 Bond Enthalpy (BE)

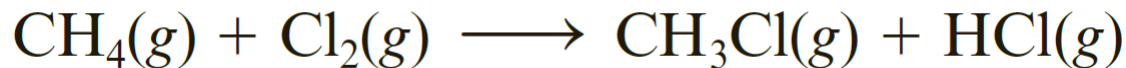
“bond enthalpy” and “bond energy” are often used interchangeably



$$\rightarrow BE(\text{C}-\text{H}) = \frac{1}{4} \times 1662 \text{ kJ} = 416 \text{ kJ}$$

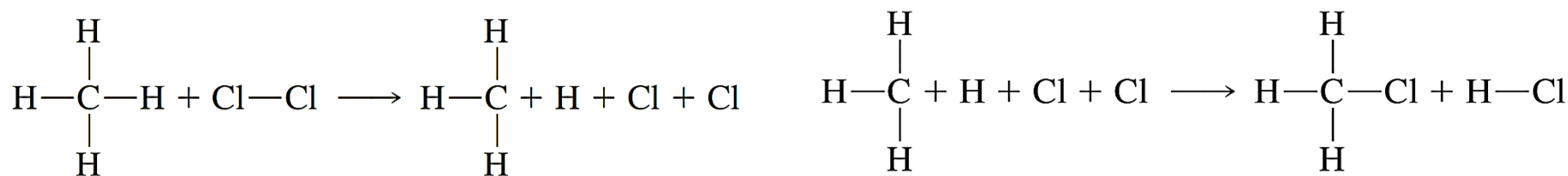
- ✓ Because it takes energy to break a bond, bond enthalpies are always positive numbers.
- ✓ Bond enthalpy is a measure of the strength of a bond:
the larger the bond enthalpy, the stronger the chemical bond

(Q) Use bond enthalpies to estimate the enthalpy change for the following reaction:



Given that bond enthalpies (kJ/mol) for:

(C-H) = 413, (Cl-Cl) = 242, (C-Cl) = 328, (H-Cl) = 431,



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$$\begin{aligned} \Delta H &\approx BE(\text{C}-\text{H}) + BE(\text{Cl}-\text{Cl}) - BE(\text{C}-\text{Cl}) - BE(\text{H}-\text{Cl}) \\ &= (413 + 242 - 328 - 431) \text{ kJ} \\ &= -104 \text{ kJ} \end{aligned}$$

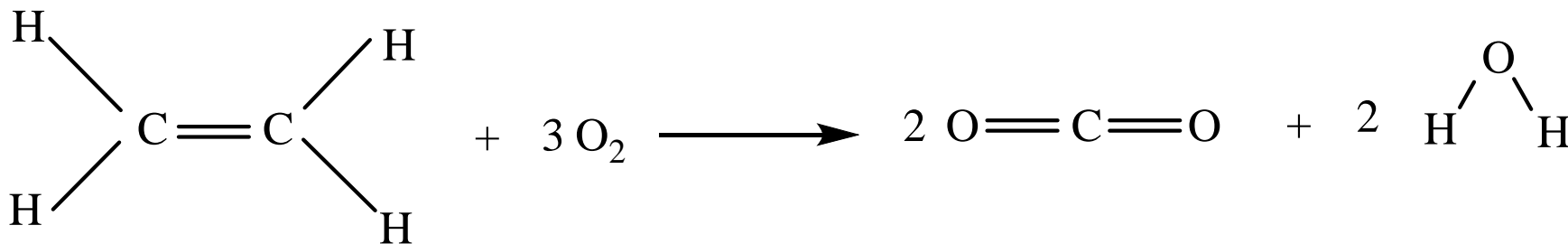
In general, the enthalpy of reaction is (approximately) equal to the sum of the bond enthalpies for bonds broken minus the sum of the bond enthalpies for bonds formed.

Exercise 9.18 Use bond enthalpies to estimate the enthalpy change for the combustion of ethylene, C_2H_4 , according to the equation



Given that bond enthalpies (kJ/mol) for:

(C=C) = 614, (C-H) = 413, (O=O) = 498, (C=O) = 804, (O-H) = 463



$$\begin{aligned} \Delta H &= \{[614 + (4 \times 413) + (3 \times 498)] - [(4 \times 804) + (4 \times 463)]\} \text{ kJ} \\ &= -1308 \text{ kJ} \end{aligned}$$