BONDING

Chemistry 20

Reactivity

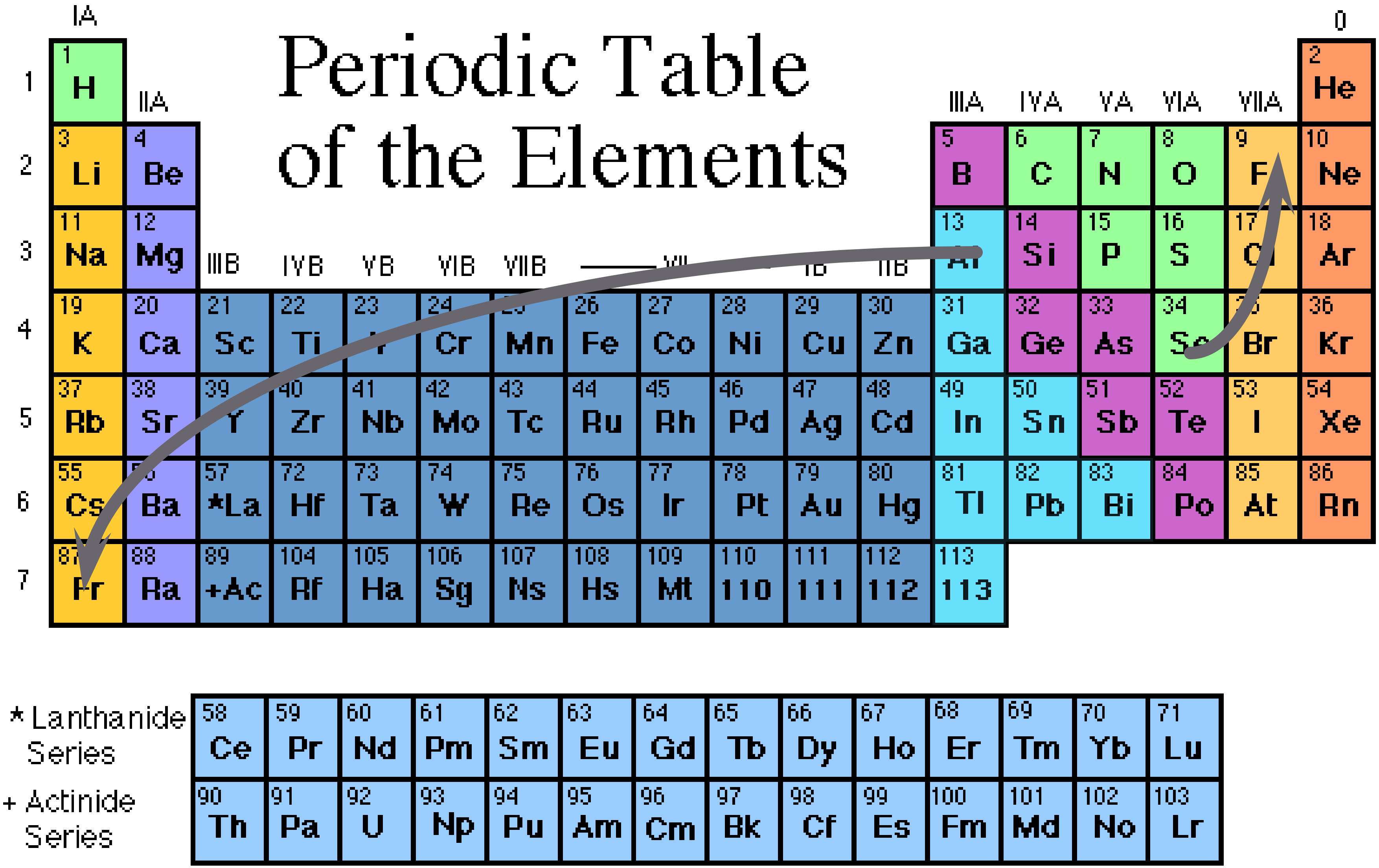
Most reactive metal: Fr ;
Most reactive non-metal: F;
Halogens react vigorously with alkali metals because they transfer electrons.

naclrxn.mov

Reactivity

- Nobel gases are very stable because their orbitals are full
- Reactivity increases down and to the left for metals and up and to the right for non-metals

Periodic Table of the Elements



	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt	110 110	111 111	112 112	113 113					

* Lanthanide Series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
+ Actinide Series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Bonding

A chemical bond is a <u>net electrostatic attractive force between two atoms, ions, or molecules.</u>

Bonding

Atoms try to achieve electron structures of the nearest Nobel Gas

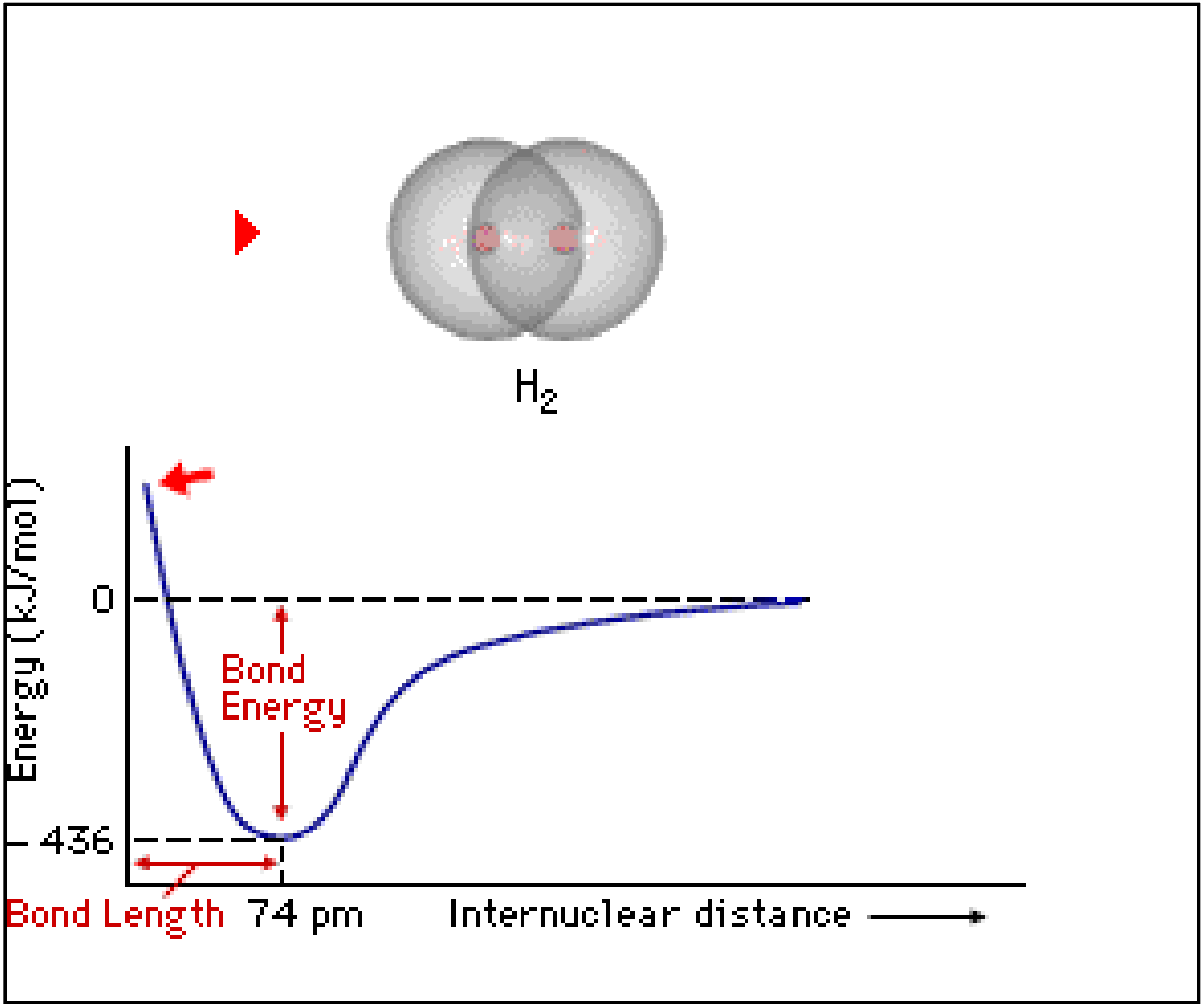
Lose energy-forming a bond is exothermic

Gain energy- breaking a bond is endothermic

Bonding

Energy required to break a bond is called the <u>Bond Energy</u>

Look at the demo on the next slide.



bond length.mov

Ionic Vs. Molecular

solid	solid liquid or gas
any colour	white or colourless
varying solubility	varying solubility
conductive	nonconductive
ionic bond	covalent bond
ions present	no ions present
metal+nonmetal	nonmetals only
mp above 300C	mp below 300C

Bonding

Two main types of bonding forces.

a) intermolecular forces-between different molecules i.e.) water and salt- Strength depends on size, polarity, & electrons

b) Intramolecular forces-within molecules i.e.) hydrogen and oxygen in a water molecule

Bonding

Relative Bond Strength(approx.)

STRONG

Intra: Network-Ionic-Metallic-Covalent

Inter: hydrogen-dipole dipole-LD forces

WEAK

(Watch KhanAcademy.org)

VALENCE ORBITALS & LEWIS DOT DIAGRAMS

Introduction

Bohr said that electrons exist in energy shells or orbitals .

These shells are regions around the nucleus where you would most likely find the electron.

Aufbau's Principle

Electrons enter the lowest energy level first. Valence electrons are electrons in the highest energy level. Valence orbital: Last energy level

Rule 1) Paul’s exclusion Principle

There are four valence shells/suborbitals each with a maximum of two electrons. Hydrogen and Helium are two exceptions with only one shell/suborbital.

Rule 2

A shell/suborbital can contain 0,1, or 2 electrons. The repulsion of the 3 electron is too great.

Rule 3: Hund’s rule

Single electrons (***bonding electrons***) occupy each of the four valence shells/suborbitals before forming electrons pairs (***lone pairs***).

Rule 4: Octet rule

A maximum of eight electrons can occupy an energy level.

Exception: H & He only have 2 electrons.

NOTE: This rule only works for the first three energy levels

Electron configurations

1) Shells or suborbitals

- s shell --> 2e
- p shells (px, py, pz) --> 6e
- d shells --> 10e (transition)

- part of the 4th energy level/not 3rd

- f shells --> 14 e (rare earths)

- part of the 6th and 7th energy level

Lewis Dot Diagrams

Gilbert Lewis (1875-1946) developed a drawing scheme now called ***Lewis or electron dot diagrams***: A chemical symbol to represent the nucleus and inner energy levels and dots to represent the valence electrons.

Lewis dot diagram rules

1. Write the element symbol to represent the nucleus & inner energy levels. This is called the kernel of the atom. Ie) Na
2. Place a dot to represent each valence electron. Place **one** dot on each side of symbol. Ie) Na•
3. If necessary add a **second** dot to a maximum of 8 (octet rule) ie) $:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Ne}}}:$

Dot diagram rule for ions

4. Remove electrons for cation charge and add electrons for anion charge. Enclose ions in square brackets with charge on outside.

Ie) [Na]+ [: Cl :] –

Dot Diagram Animation

Example #1

Atom or Ion: Ca^{2+}

# of valence electrons: (8e-) 0

Electron Dot diagram: $[Ca]^{2+}$

# of bonding electrons: 0

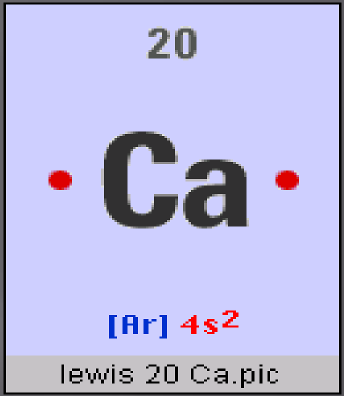
# of lone pairs: 4

Example #2

Atom or Ion: Ca

# of valence electrons: 2e-

Electron Dot diagram:



# of bonding electrons: 2

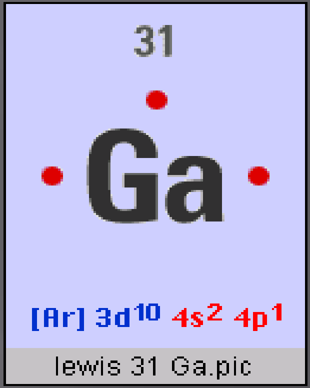
# of lone pairs: 0

Example #3

Atom or Ion: Ga

# of valence electrons: 3e-

Electron Dot diagram:



# of bonding electrons: 3

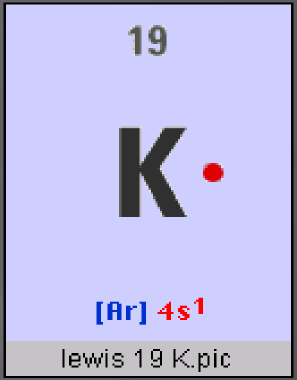
# of lone pairs: 0

Example #4

Atom or Ion: K

# of valence electrons: 1e-

Electron Dot diagram:



# of bonding electrons: 1

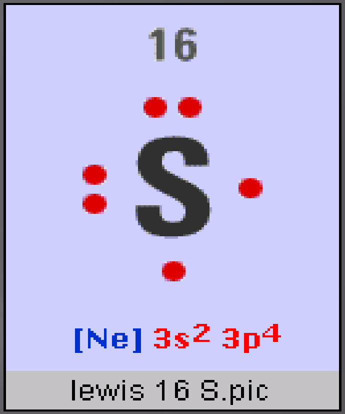
# of lone pairs: 0

Example #5

Atom or Ion: S

# of valence electrons: 6e-

Electron Dot diagram:



# of bonding electrons: 2

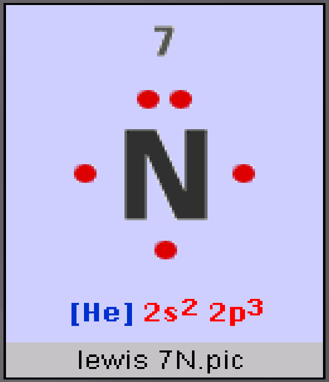
# of lone pairs: 2

Example #6

Atom or Ion: N

# of valence electrons: 5e-

Electron Dot diagram:



# of bonding electrons: 3

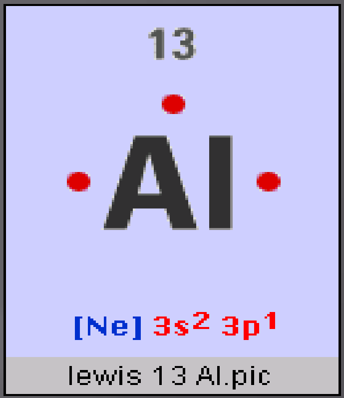
# of lone pairs: 1

Example #7

Atom or Ion: Al

# of valence electrons: 3e-

Electron Dot diagram:



# of bonding electrons: 3

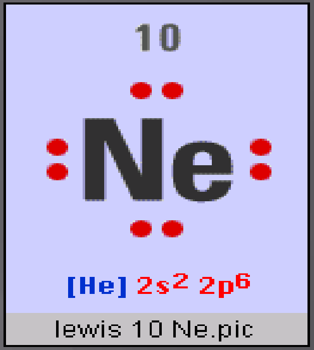
# of lone pairs: 0

Example #8

Atom or Ion: Ne

# of valence electrons: 8e-

Electron Dot diagram:



# of bonding electrons: 0

# of lone pairs: 4

Example #9

Atom or Ion: F-

# of valence electrons: 8e-

Electron Dot diagram: 

$[\,:\overset{..}{\underset{..}{F}}:\,]^-$

# of bonding electrons: 0

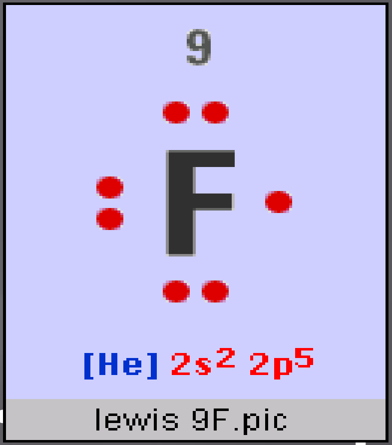
# of lone pairs: 4

Example #10

Atom or Ion: F

# of valence electrons: 7e-

Electron Dot diagram:



# of bonding electrons: 1

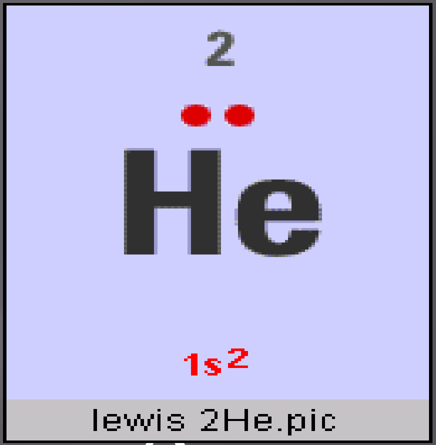
# of lone pairs: 3

Example #11

Atom or Ion: He

# of valence electrons: 2e-

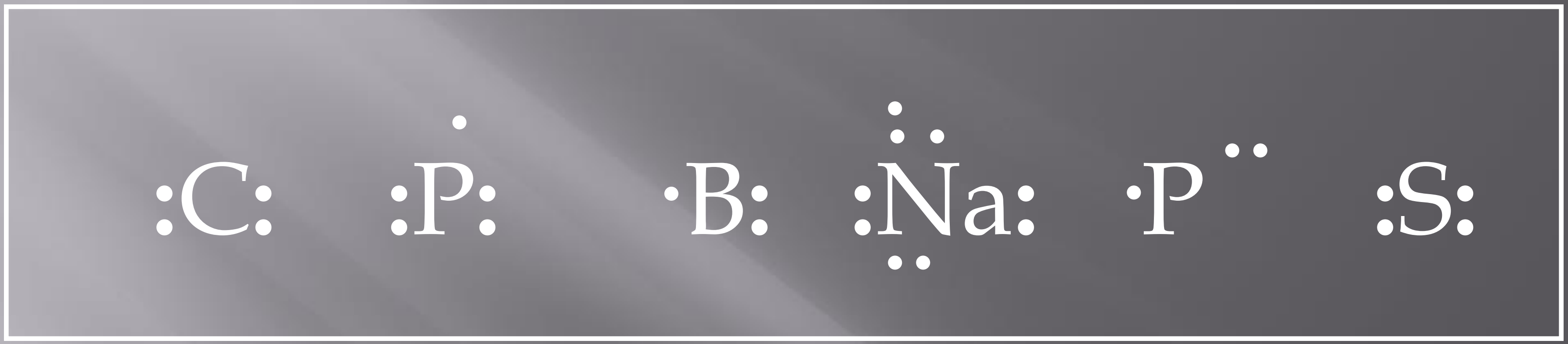
Electron Dot diagram:



# of bonding electrons: 0

# of lone pairs: 1(excp)

Incorrect Lewis Dot diagrams

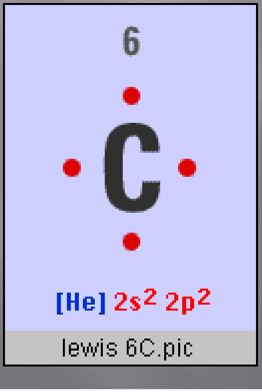


$:\mathrm{C}: \quad :\overset{\cdot}{\mathrm{P}}: \quad \cdot\mathrm{B}: \quad :\overset{\vdots\,\cdot}{\underset{\cdot\cdot}{\mathrm{Na}}}: \quad \cdot\mathrm{P}^{\cdot\cdot} \quad :\mathrm{S}:$

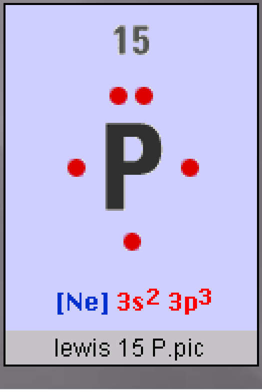
Correct Lewis Dot diagrams

:C: :Ṗ: ·B: :Ṅa: · P :Ṡ:

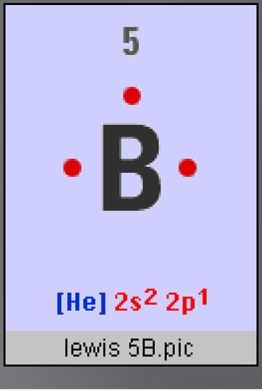
Corrections:



lewis 6C.pic

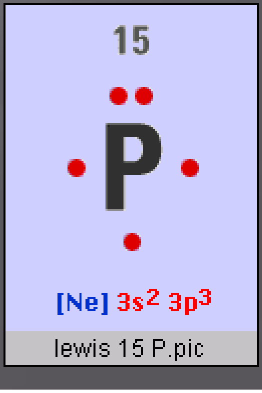


lewis 15 P.pic

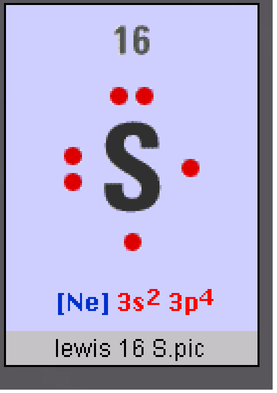


lewis 5B.pic

lewis 11 Na.pic



lewis 15 P.pic



lewis 16 S.pic

Lewis Diagrams for Molecular

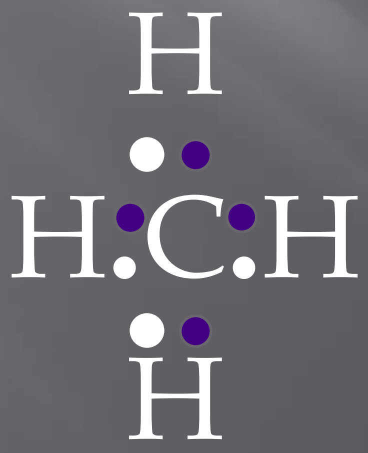
1) Sum the valence electrons for all the atoms in the molecule. Only this number of electrons can be used.
2) Determine which atom can form the most number of bonds - this is the central atom.

Lewis Diagrams for Molecular

3) Remember valence electrons obey the octet rule - except H.

Examples: Cl_2 - 14 val e-

CH_4 - 8 val e-



Example

Lewis diagrams for multiple bonds

- Double covalent bonds have two pairs of electrons shared.
- Triple covalent bonds have three pairs of electrons shared.

Examples of double/triple bond

Double:

O_2 - 12 e- $:\overset{..}{\underset{..}{O}}:\overset{..}{\underset{..}{O}}:$ $:\underset{..}{O}::\overset{..}{O}:$

CO_2

Triple:

N_2 - 10e- $:N:::N:$

Lewis diagrams for Ionic

- Electrons are transfered
- Electrons around the metal are lost
- Charges ions are enclosed in brackets.

Examples of Ionic Lewis

NaCl $[\mathrm{Na}]^+ \, [:\overset{..}{\underset{..}{\mathrm{Cl}}}:]^-$

K_2S $[\mathrm{K}]^+ \, [:\overset{..}{\underset{..}{\mathrm{S}}}:]^{2-} \, [\mathrm{K}]^+$

Animation

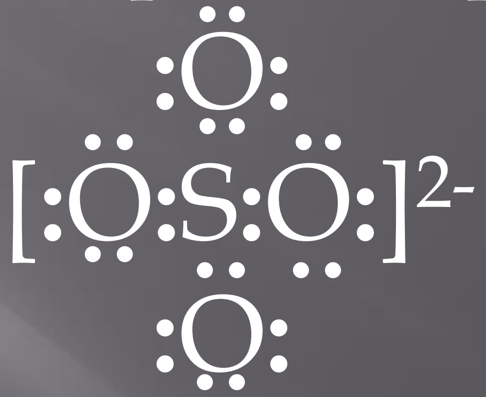
Lewis diagrams for Complex Ions

- Sum the valence electrons for all the atoms in the molecule.
- Add electrons for negative charges and subtract for positive charges
- Arrange the electrons around the atom so the sum is achieved and octet rule followed (hit and miss)

Examples of Complex Ions

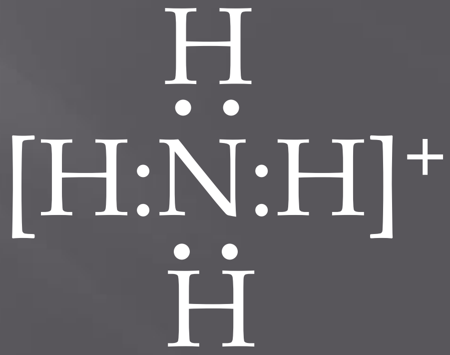
Cyanide - CN^- (10e-) = $[:C:::N:]^-$

Sulphate-SO_4^{2-}(32e-)



$$\begin{array}{c} :\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}: \\ {[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}:S:\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}:]^{2-}} \\ :\underset{\cdot\cdot}{O}: \end{array}$$

Ammonium-NH_4^+(8e-)



$$\begin{array}{c} H \\ \cdot\cdot \\ {[H:N:H]^+} \\ \cdot\cdot \\ H \end{array}$$

IONIC BONDING

A. Review

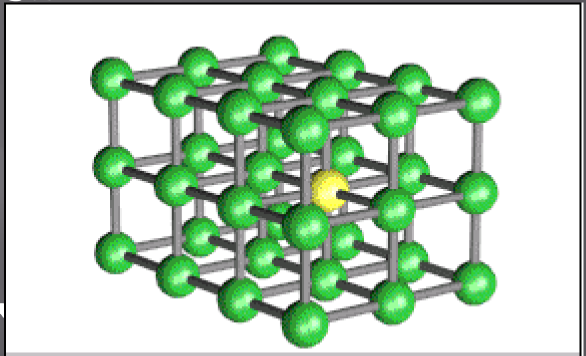
Ions: atom or group of atoms with a positive or negative charge

Metallic Ions: cation with positive charge (low electronegativity)

Non-metallic Ions: anion with negative charge
(high electronegativity)

Ionic Bond

- Ionic Bond: net electrostatic attraction between cation and anion
- Ionic bonds produce a crystal lattice with a repeating pattern
- Ionic compound char (look at the beginning of your notes)



ioniccrystal.mov

Formula Unit

Formula Unit: the lowest whole number ratio of ions in an ionic compound.

ie) NaCl - ratio is 1 to 1; In reality there is a lot more than 1 sodium or 1 ch

nacl ball&stick.mov

Dot diagrams

Electron Dot Diagrams

2 Na + Cl_2 --------->>2 $[Na]^+[:\overset{..}{\underset{..}{Cl}}:]^-$
(brackets show that they are not sharing but transfering)

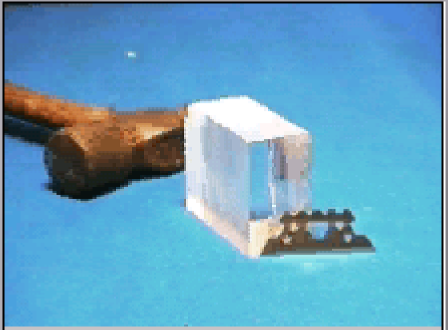
Oxidation-Reduction

- Oxidation - Reduction: (LEO goes GER - Losing Electron Oxidation, Gaining Electrons Reduction) (OIL RIG - Oxidation Is Losing, Reduction Is Gaining)

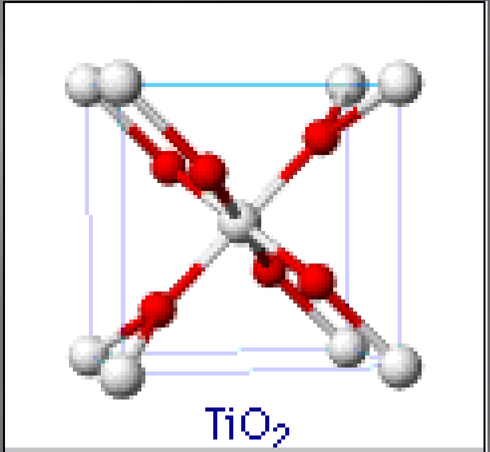
Oxidation - Reduction

Oxidation: loss of electrons(metal)
(The lower the electronegativity, the easier it is oxidized.)
ie) Na(s) => $Na^+ + 1e^-$
Reduction: gaining of electrons (non metal)
ie) $Cl_2 + 2e^- \Rightarrow 2Cl^-$

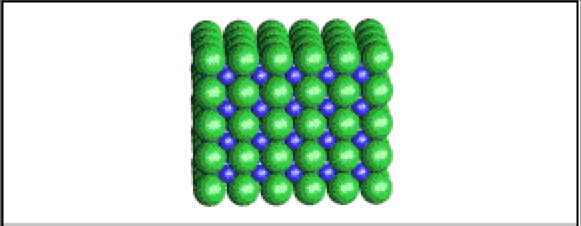
Ionic Crystals



kbr crystal.mov



tio2crystal.mov



csclvsnacl.mov

METALLIC BONDING

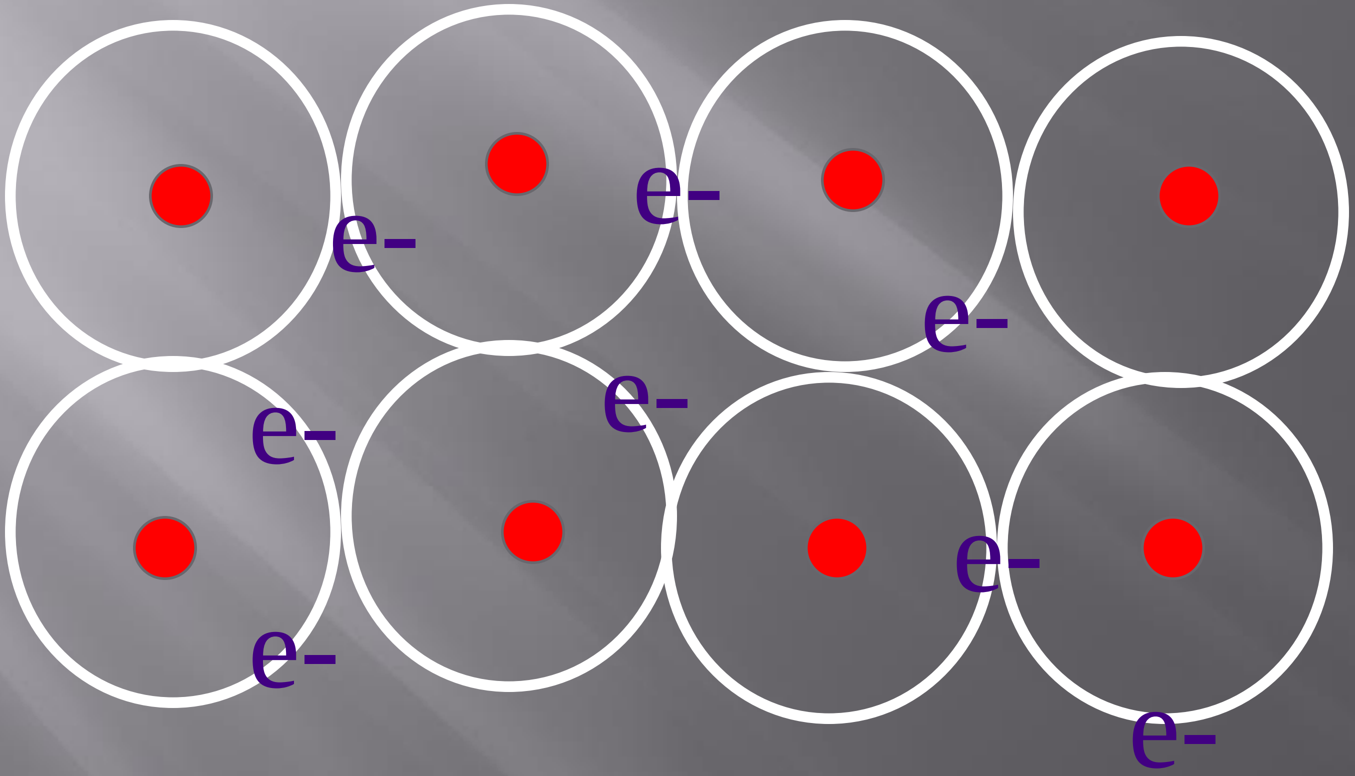
Definition

Metal: a network of very loosely bound valence electrons around fixed positively charged metal ions

Properties

- positive ions are momentarily formed(ionic) while electrons are shared(covalent)
- Great conductors of heat and electricity(electrons free to move)

Diagram



NETWORK COVALENT

Definition

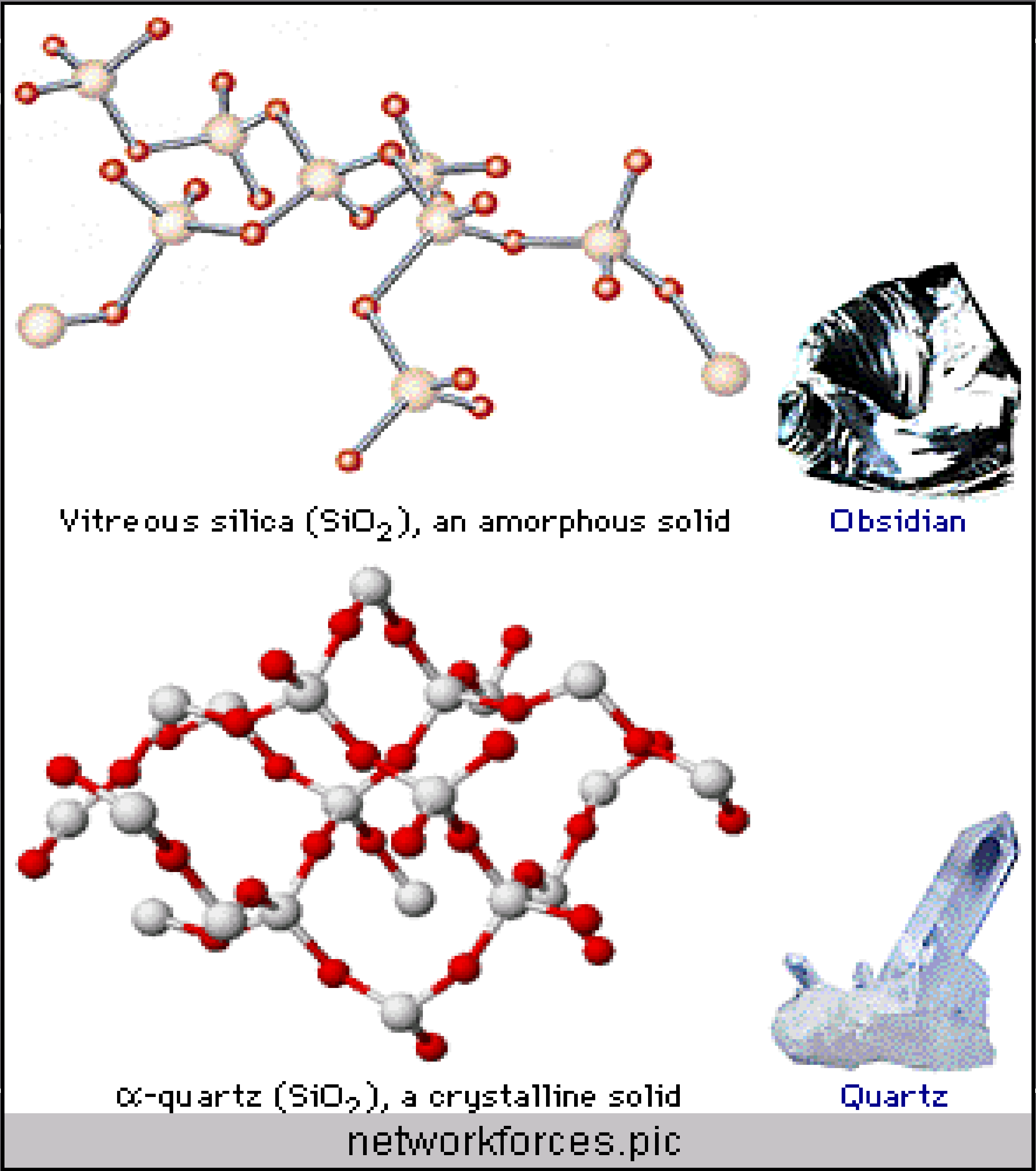
A network of covalent bonds between group Si & C (group 14) where electrons do not move freely.

Properties

Hardest inflexible bond known

High melting point - diamonds vaporize at 3500 C.

Diagram
overhe



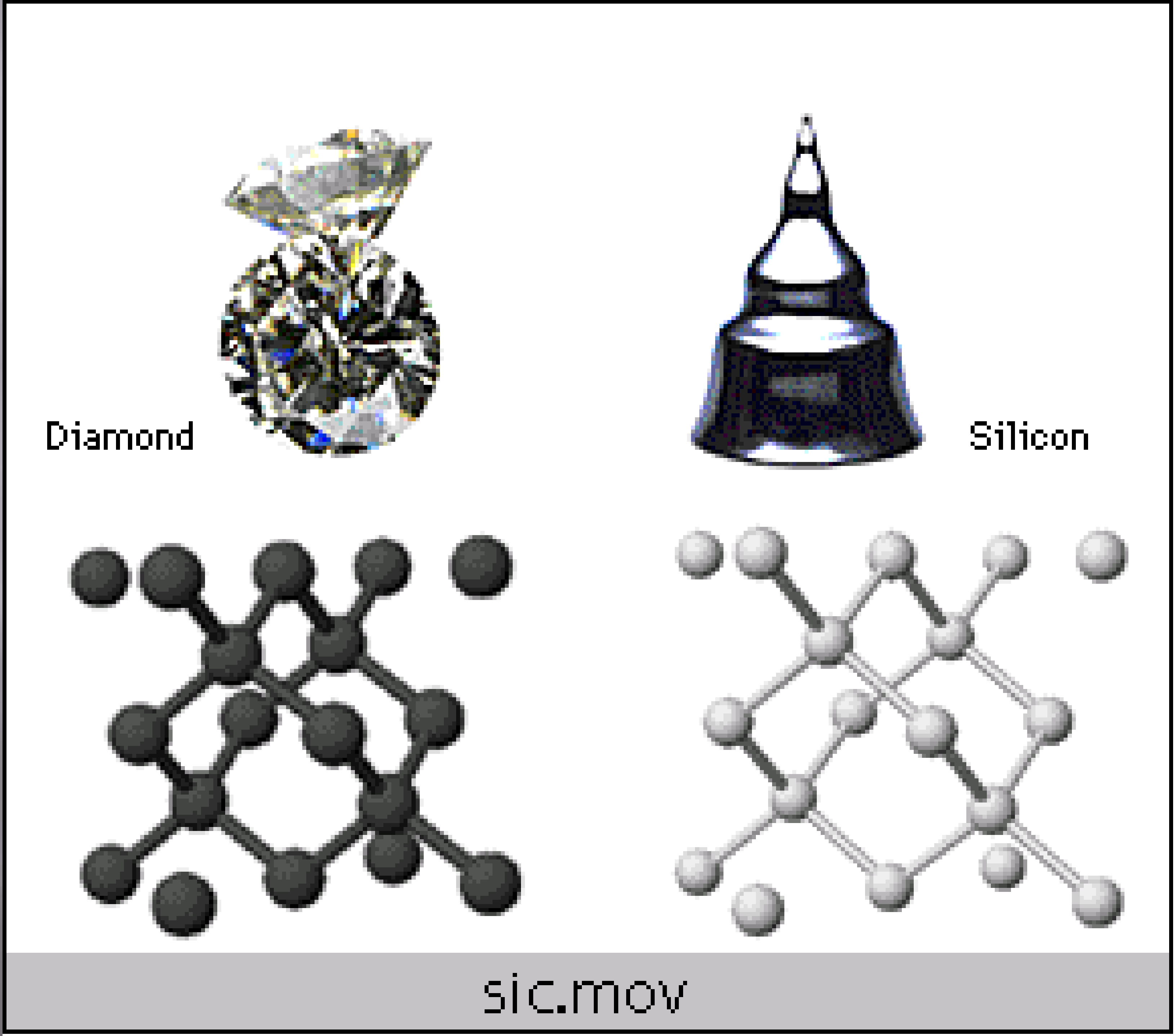
Vitreous silica (SiO_2), an amorphous solid

Obsidian

α-quartz (SiO_2), a crystalline solid

Quartz

networkforces.pic


Diamond
Silicon
sic.mov

COVALENT BONDS

Introduction

Summary: occur between non metals by sharing of electrons

Form an octet. Can be single, double, or triple bonds

Bonding Capacity

Bonding Capacity: The maximum number of bonds an atom can form. Bonds can be represented by dashes.

Hydrogen

Atom	H
# of valence	1
# of bonding	1
Bonding Cap	1 H-(one single)

Carbon

Atom	C
# of valence	4
# of bonding	4
Bonding cap	4 −C≡ (triple single)

=C= (2 doubles) −C− (4 singles)

−C= (double 2 singles)

Nitrogen

Atom N

# of valence 5

# of bonding 3

Bonding cap 3 N≡ (1 triple)

-N= (one double, 1 single)

-N- (3 singles)

|

Oxygen

Atom	O	
# of valence	6	
# of bonding	2	
Bonding cap	2	O= (1 double) -O- (2 singles)

Fluorine

Atom	F	
# of valence	7	
# of bonding	1	
Bonding cap	1	F- (1 single)

Neon

Atom	Ne
# of valence	8
# of bonding	0
Bonding cap	0

NO BONDS = INERT

Structural Diagrams

Replace the shared pairs with a line (bond) and erase the lone pairs

Molecular formula: written with symbols only

Diatomic

Molecular Cl_2

Lewis $:\overset{\cdot\cdot}{\underset{\cdot\cdot}{Cl}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{Cl}}:$

Stuctural: Cl-Cl

One/two lone pair

Molecular	NH_3	H_2O
Lewis	·· H:N:H ·· H	·· H:O: ·· H
Structural	H-N-H | H	H-O | H

Double Bonds

Molecular O_2

Lewis $\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}::\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}$

Structural $O=O$

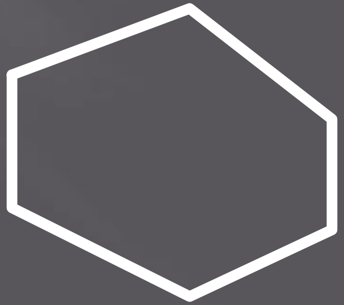
Triple Bonds

Molecular	N_2
Lewis	:N:::N:
Structural	N=N

Cyclic carbon compounds

Molecular cyclohexane C_6H_{12}

H
··
Lewis: ⋱ C ⋰
··
H:C:H H H:C:H
·· ··
H:C:H H H:C:H
··
⋰ C ⋱
··
Structural: H



Lewis Problems

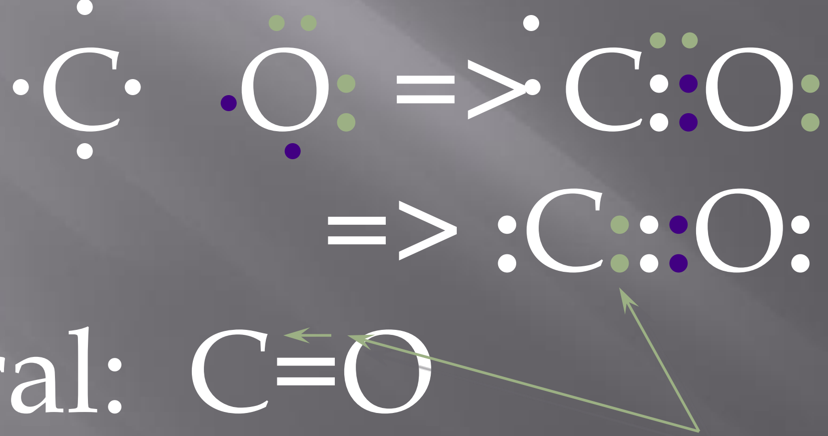
1) Coordinate covalent bonds:

one atom supplies both electrons (lone pair) to form a covalent bond.

Examples

a) carbon monoxide

Molecular: CO - 10 e-



Lewis: ·C· ·O: => C::O:

=> :C:::O:

Structural: C=O

coordinate covalent bond

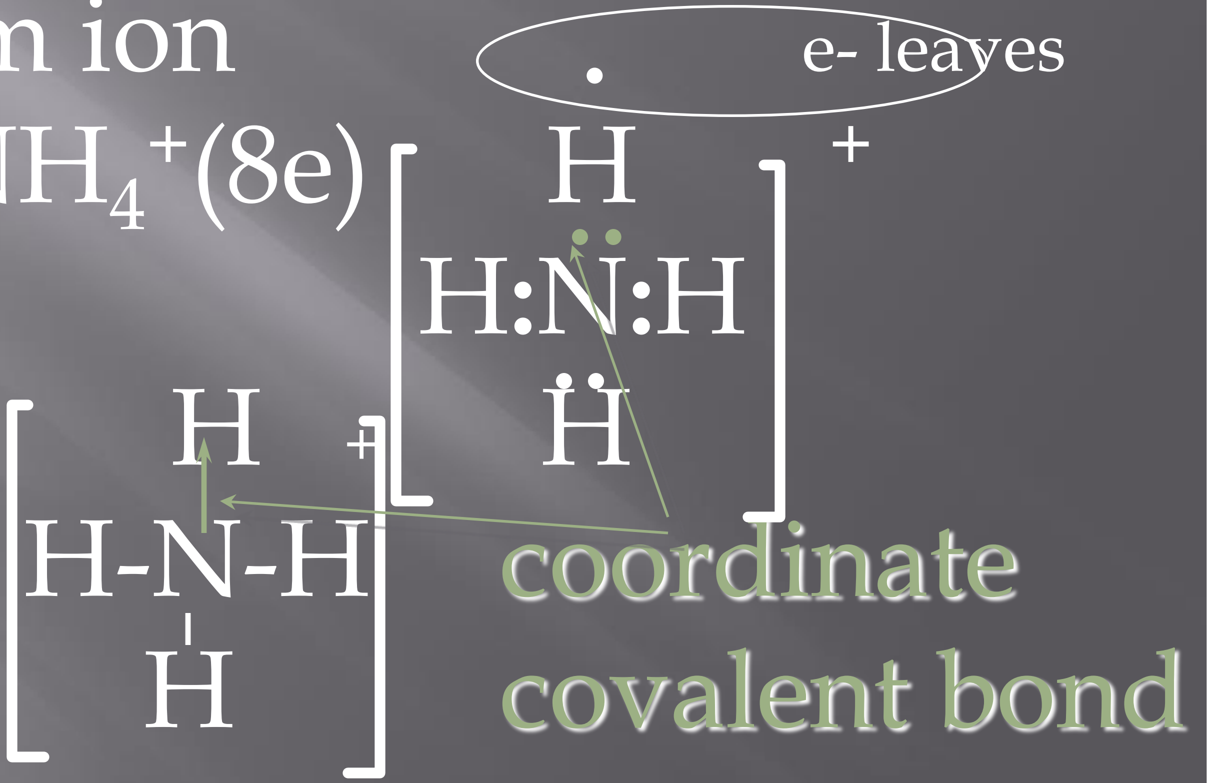
Examples

ammonium ion

Formula:NH_4^+(8e)

Lewis:

Structural:



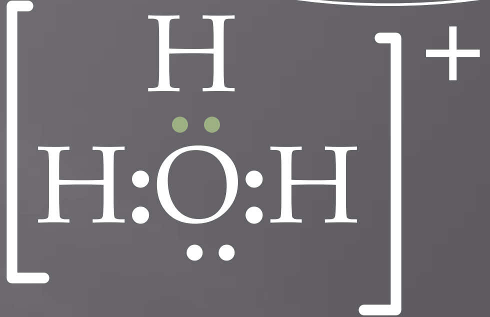
coordinate covalent bond

hydroxide ion

Formula: H_3O^+(8e)

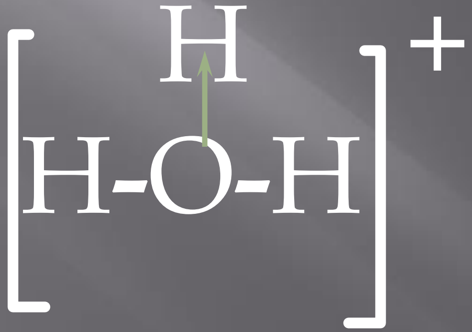
• e- leaves

Lewis:



$$\left[\begin{array}{c} H \\ H:\overset{..}{\underset{..}{O}}:H \end{array}\right]^+$$

Structural



$$\left[\begin{array}{c} H \\ \uparrow \\ H-O-H \end{array}\right]^+$$

coordinate covalent bond

Problems with Lewis con’t

Resonance: when two or more electron dot diagrams can be written

Hybrid: single diagram that combines the resonance's.

Examples

a) Ozone - $O_{3(g)}$ - 18 e-

Lewis: $\overset{..}{\underset{..}{\mathrm{O}}}::\underset{..}{\mathrm{O}}:\overset{..}{\underset{..}{\mathrm{O}}}:$ or $:\overset{..}{\underset{..}{\mathrm{O}}}:\underset{..}{\mathrm{O}}::\overset{..}{\underset{..}{\mathrm{O}}}$

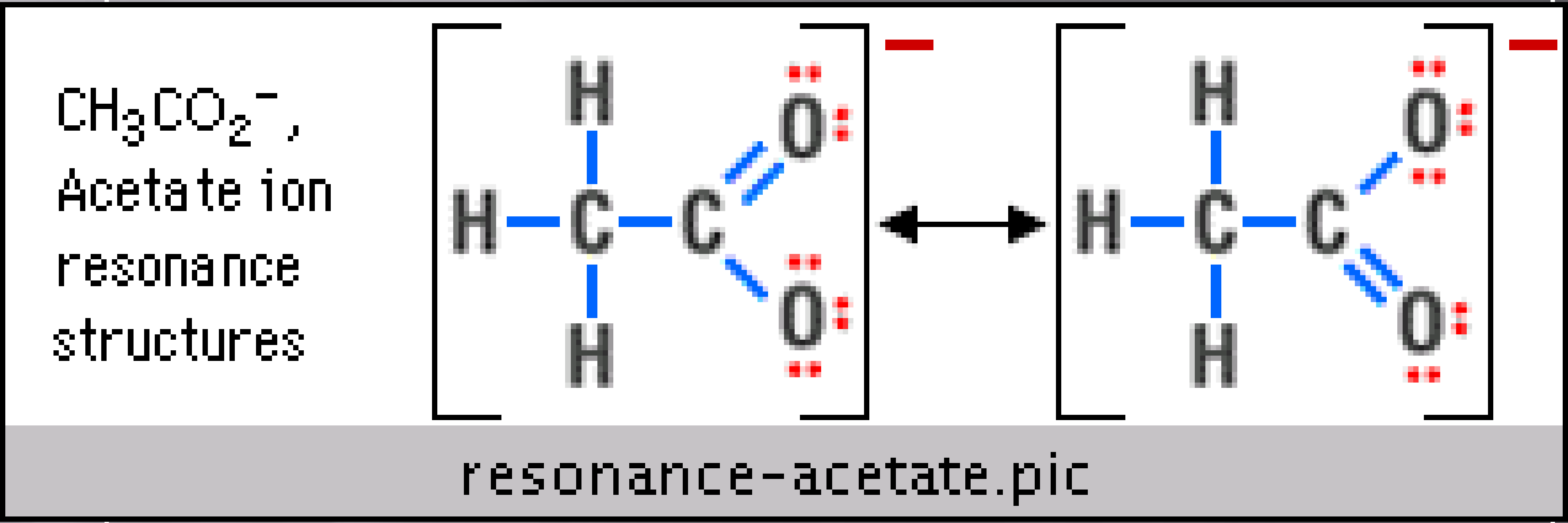
Structural: O=O-O or O-O=O

Hybrid: $\mathrm{O}\,\overset{-}{\scriptstyle -}\,\mathrm{O}\,\overset{-}{\scriptstyle -}\,\mathrm{O}$

1.5 bond

Examples

acetate ion (CH_3COO^-) ^-24e

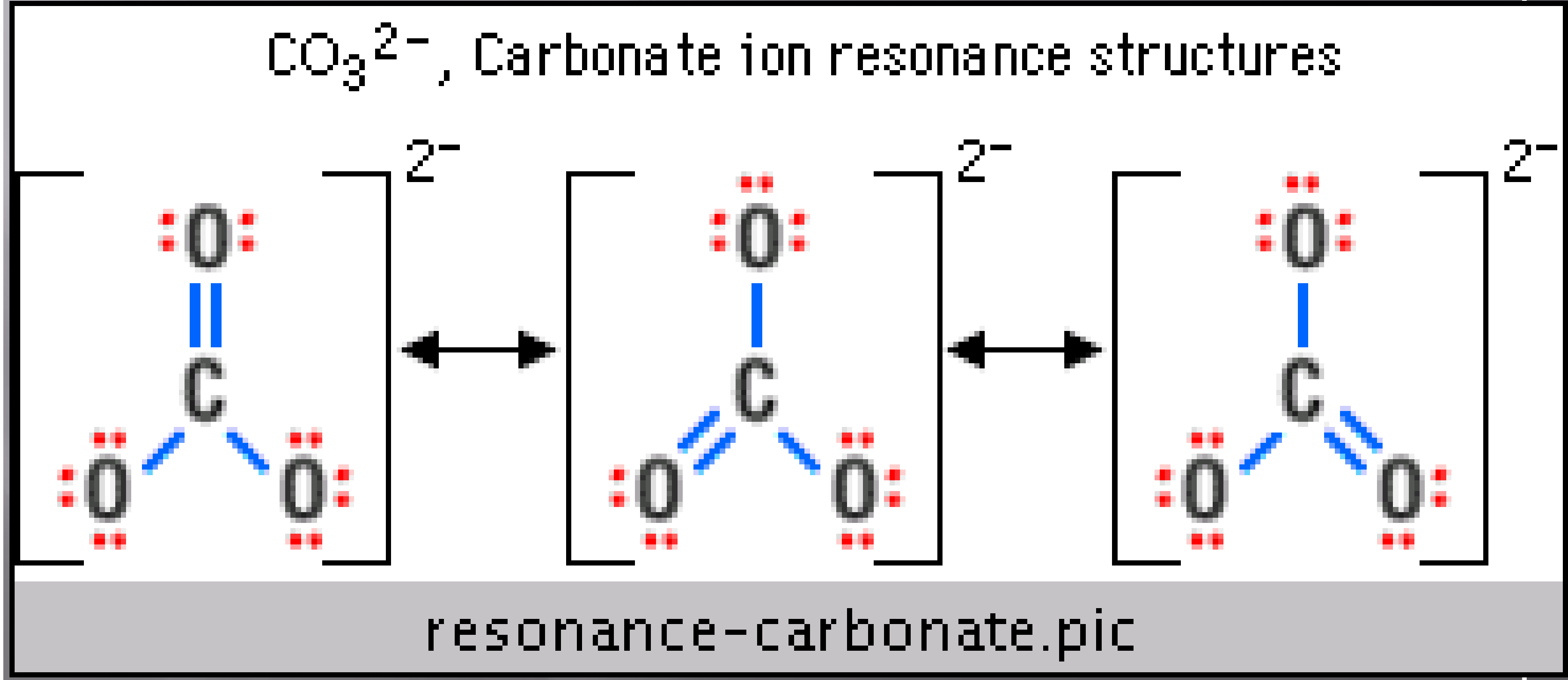


$CH_3CO_2^-$, Acetate ion resonance structures

resonance-acetate.pic

Examples

carbonate ion CO_3^{2-} $^{-}$24e



CO_3^{2-}, Carbonate ion resonance structures

resonance-carbonate.pic

More problems with Lewis

Exceptions to the octet rule

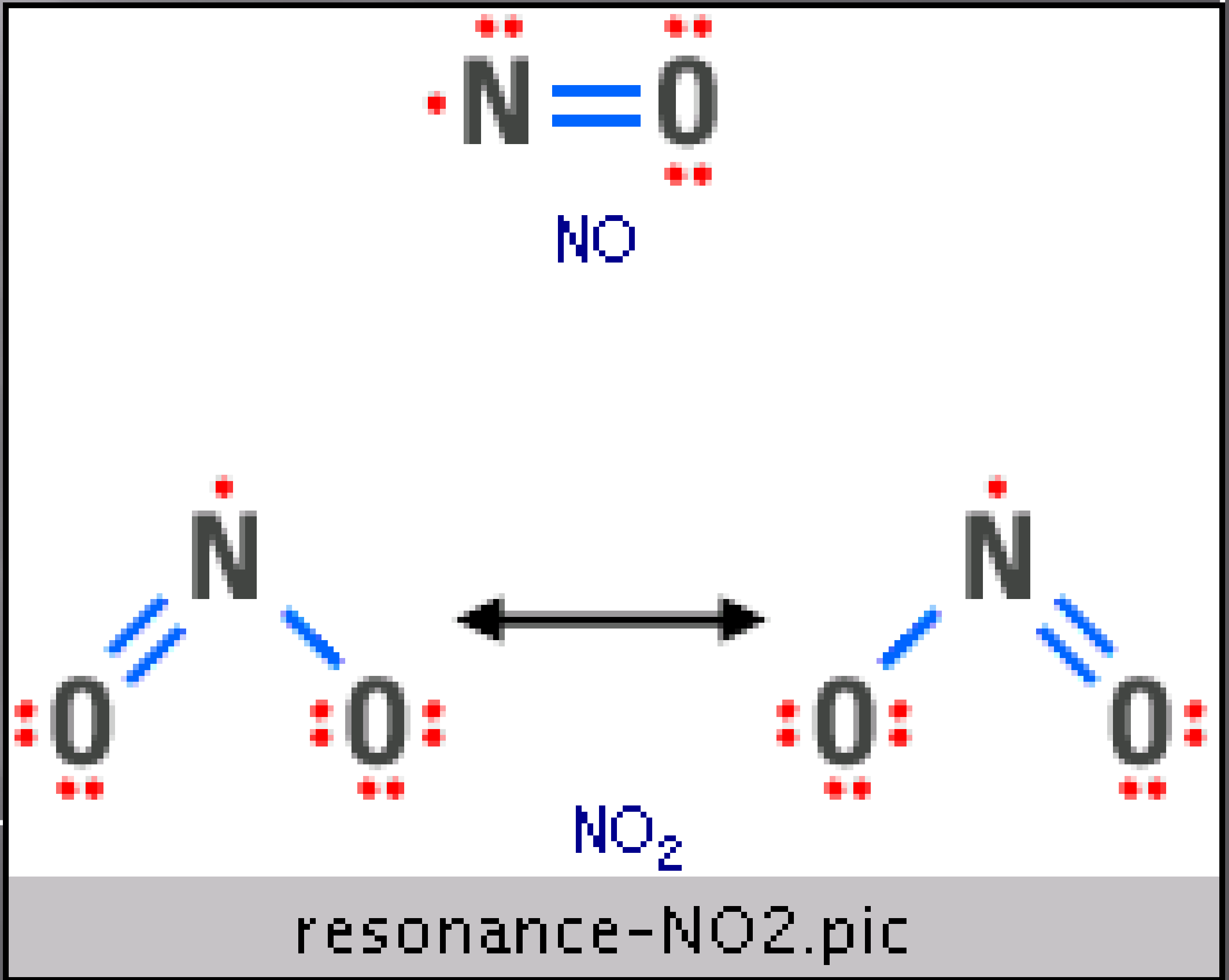
- More or less than 8 electrons

  -P, S & Se can have 10 or 12 electrons around the atoms

  -Be & B can have 4 or 6 electrons around their atoms
- Less than than 2 electrons - NO_2

Examples

nitrogen dioxide (NO_2)



Summary

Lewis diagrams *tell us:*

1) electrons involved with bonding

2) type & number of bonds

Lewis diagrams *do not tell us:*

1) what an atom looks like (ie- inner energy levels)

2) where the electrons are (ie- random location

Summary

- Bonds are formed by *sharing* of electrons
- Two electrons are involved in each *covalent* bond
- diagrams and formulas usually satisfy the *octet* rule

VSEPR THEORY

Stereochemistry

- *the study of the molecular shapes*
- VESPR: The theory of Valence-Shell Electron-Pair Repulsion used to predict molecular shapes

VESPR

Assumptions:

- all molecules have a 3-D shape
- number of valence electrons, number of bonding electrons and the central must be known to determine the shape

VSEPR

- Electron pairs repel each other; they try to get the maximum distance from each other
- lone pairs repel each other more than the bonding pairs of electrons

VSEPR Rules

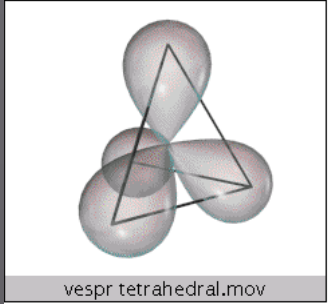
1) DRAW *Covalent bonds in the same plane with solid lines (-); covalent bonds between atoms behind the paper with dotted lines(--); covalent bonds in front of the paper with a wedge* (▲)

Rules

2)

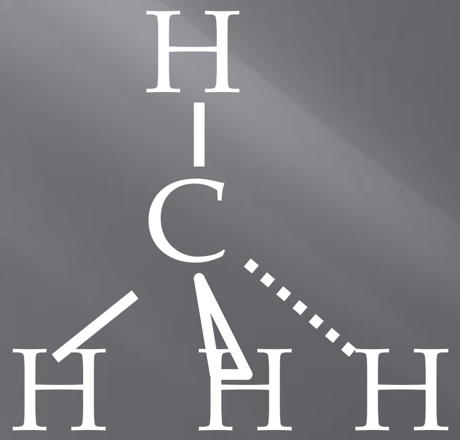
3) *Minimize the bonds in front and behind the central atom*

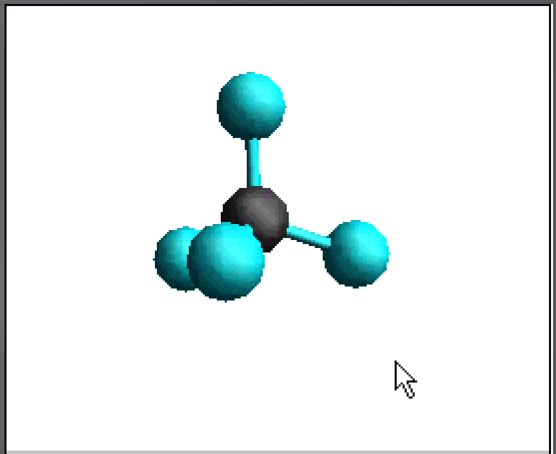
Shapes


vespr tetrahedral.mov

1) Tetrahedral: 4 bonding pairs, 0 lone pairs, single bonds, group 14

i.e.) CH_4



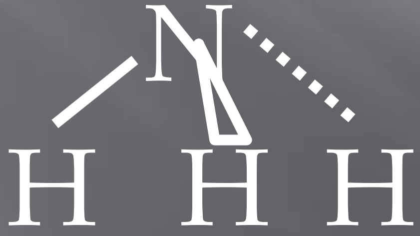

pyramidal.mov

properties: non-polar if pendant atoms are all the same; 109.5^0

Shapes

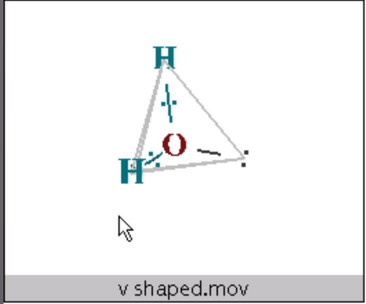
2) Pyramidal: 3 bonding pairs, 1 lone pairs , single bonds, group 15

i.e.) NH_3



properties: always polar; 120^0 bond angle

Shapes



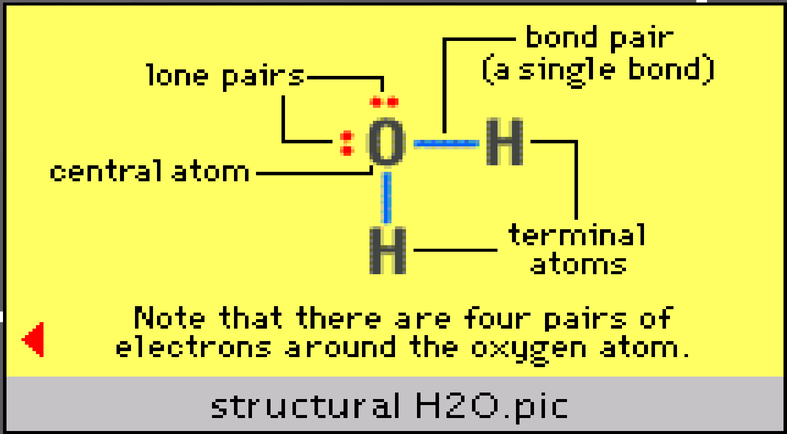
v shaped.mov

3) V shaped(bent): 2 bonding pairs, 2 lone pairs , single bonds, group 16

i.e.)H_2O

properties: polar

104.5^0

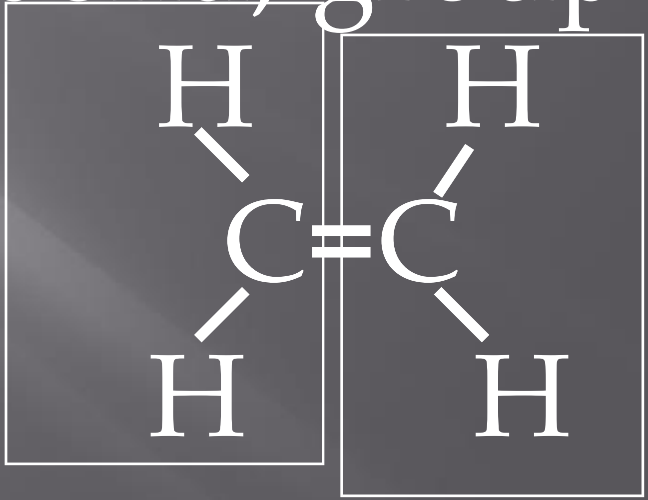


Note that there are four pairs of electrons around the oxygen atom.

structural H2O.pic

Shapes

4) Trigonal planer: 4 bonding pairs, 1 double bond, group 14 central atom



i.e.) C_2H_4

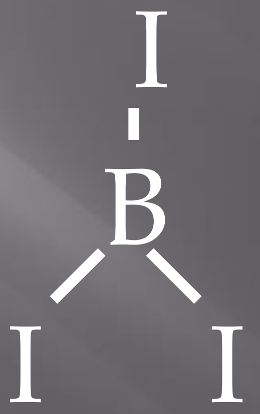
2 trigonals

properties: nonpolar if pendant atoms are the same; 120^0

Shapes

4) Trigonal planer: 3 bonding pairs, 0 lone pairs around atom(group 13)

i.e.) BI_3



(no octet)

properties: non polar if pendant atoms are the same; 120^0

SHAPES

5) Linear- (diatomic): 1 bonding pair, 0 or 3 lone pairs around the atom

i.e.) HCl H--Cl - polar

Cl_2 Cl--Cl - non polar

Even more Shapes

5) Linear-tetratomic: 4 bonding pairs, 1 triple bond around the atom (group 14 atoms)

ie) C_2H_4 H-C≡C-H

properties: non polar if pendant atoms are all the same.

More Shapes

Linear-(triatomic): 4 bonding pairs,

2 double bonds around the atom

(group 14 atoms)

i.e.)CO_2 O=C=O

properties: usually non polar

Still More Shapes

6) Combinations:

i.e.)CH_3OH

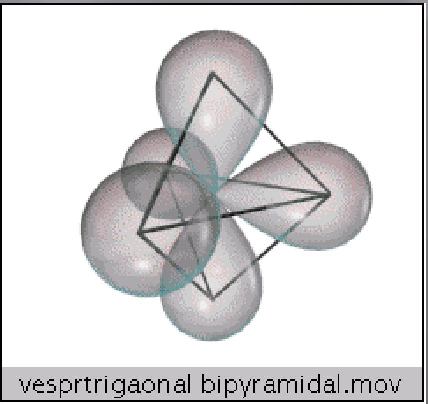
V shaped and tetrahedral

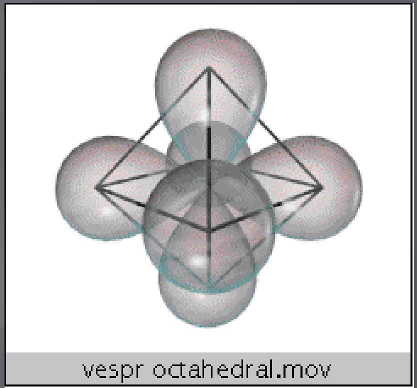
7) Others:

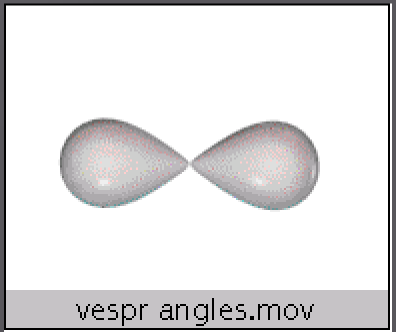
a) Trigonal bipyramidal - PCl_5

b) octahedral - SF_6

c) pentagonal bipyramidal - IF_7







Electronegativity

Electronegativity

measure of an atoms attraction for electrons (number on your periodic table from 0 to 4.0)

Principles by Linus Pauling

1. As you go left and down the metals electrons are held more loosely. Therefore the element loses electrons and the electronegativity decreases.

LEFT- LOW- LOOSE

Pauling's Principles

2. As you go right and up the nonmetals are held more tightly. Therefore the element gains electrons and the electronegativity increases.

RIGHT-HIGH-TIGHT

Pauling's Principles

3. Noble Gas Exception: they hold their electrons most tightly but have no electronegativity because they have an octet (stable)

Electronegativity and Bond Type

2. If the electronegative difference between two atoms is equal to or greater than 1.7, the bond is ***usually ionic.***

Electronegativity and Bond Type

2. If the electronegative difference between two atoms is than 1.7, the bond is ***usually polar covalent.***

Electronegativity and Bond Type

2. If the electronegative difference between two atoms is than 0, the bond is ***usually non-polar covalent or metallic.***

Electronegativity and Polarity

1. polar covalent bond: a bond that forms from the unequal sharing of bonding electrons.

Electronegativity and Polarity

2. polar molecules & dipoles.

Polar molecule: a molecule that has a slightly uneven charge distribution, with charged regions.

Electric poles

Charged regions represented by delta signs

Bond Dipole

When one end of a bond is positively charged and one end is negativly charged due to the differing electronegativities of the atoms sharing the electrons

polarHCl.pic

Drawing bond dipoles

Step 1) Determine the electronegative difference by subtracting the electronegativities and looking at VESPR diagrams.

Step 2)

Use Delta signs to indicate small charged regions. (not ionic)

Step 3)

Draw a dipole with an arrow over the top the molecule, with the tip of the tip pointing towards the negative pole

----->

H - Cl

dipoleHF.pic

Non polar covalent Bond

Bond that forms from the equal sharing of bonding electrons. (electro cancel)

3.0 3.0 3.5 3.5

ie)Cl_2: Cl - Cl CO_2: O=C=O

dipoleCO2.pic

netdipolefff.pic

netdipoleffh.pic

dipoleCCl4.pic

dipoleCHCl3.pic

Conclusion

Polarity is based on the net electronegativity and the shape of the compound

bondingofwater2.mov and bendingofwater.mov

INTERMOLECULAR FORCES

Intermolecular Bonding

- The size of the molecules and the polarity determine the strength of intermolecular forces

Intermolecular forces

Electron and nuclear attractions between molecules cause intermolecular bonding

Intermolecular Forces

- Intermolecular forces affect properties like solubility, melting point, boiling point etc….
- REMEMBER: polar substances dissolve polar substances

Intermolecular Forces

Overview of types of forces

I) 2 Van der Waal forces

1. London Dispersion (all atoms)
2. dipoledipole

II) Hydrogen forces (bonds)

London Dispersion

I) The first 2 are called Van der Waal

1) ***<u>London Dispersion (LD)</u>***

(temporary dipoles): The electrostatic attraction between **<u>all</u>** molecules due to temporary dipoles (*the temporary formation of polar regions)* These are the *<u>weakest</u>* intermolecular forces

LD - Forces

LD forces hold molecules together when they are in solid and liquid states. The diatomic gases have very weak LD forces, while iodine has stronger LD forces. (solid at 25C)

LD Forces

The larger the atom, the stronger the LD forces, because there are more electrons - more temporary poles.

In LD forces there is no net shift of electrons, there LD forces are not permanent and the molecule is still nonpolar

LD Forces

- LD forces are stronger the larger the atom because there are more electrons - more temporary poles.
- In LD forces there are no *net* shift of electrons, therefore LD forces are *not permanent* and the molecule is still *nonpolar*

LD Forces

- If two molecules have the same number electrons they are called *isoelectronic*
- ie) Ar and HCl are isoelectronic because they both have 18e^-. HCl has a higher boiling point because of dipole dipole forces.

Examples

The dotted lines represent the forces

LD forces

Boiling point: are indirect measurement of Van der Waal's forces. The more electrons the higher the force.

dipole dipole

2) *dipole dipole (DD)*: electrostatic attraction between two **polar** molecules

Examples

- Examples of dipole dipole force
- The dotted lines represent the forces

Hydrogen Forces (bonds)

Special, relatively strong dipole dipole forces between H of one molecule and O, F and N of another molecule

H- Bonds

These bonds form because of the *electrostatic attraction between hydrogen & N, O or F*.

H - Bonds

The small size of the hydrogen atom means it has a strong exposed positive nucleus and this positive pole exerts a stron attraction to the negative pole of a nearby molecule.

H - Bonds

V - shaped molecules with hydrogen bonds, like water, have stronger hydrogen bonds than any other molecule, like ethanol.

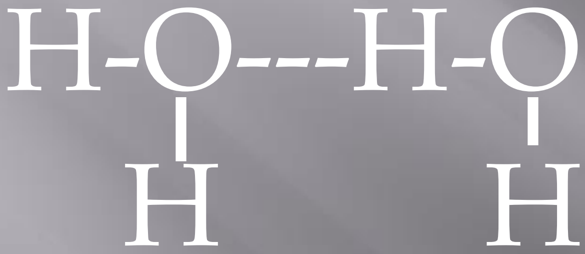
H- Bonds

- This is because they are smaller - easy access to the hydrogen. Water, hydrogen fluoride and ammonia have very strong hydrogen bonds.

Examples

- H-F---H-F hydrogen bonds are represented by the dotted lins because they are between different molecules.

Examples



$$\text{H-}\underset{\substack{| \\ \text{H}}}{\text{O}}\text{---H-}\underset{\substack{| \\ \text{H}}}{\text{O}}$$

NOTE: There are lots of hydrogen bonds that can form because water is small and has two hydrogens to bond with

Ethanol Example

Physical Properties

1) The melting point and boiling point (temperature) increase as the inter molecular forces increase due to increased energy that is needed to break the intermolecular bonds

Intermolecular Forces Physical Properties

To determine mp, need to look at

a) Type of forces involved

Relative Strengh Scale

LD	DD	HB	Ionic	Network
1	5	10	50	100

metallic mp may vary from -39 (Hg) to3410(W) but are generally weaker than ionic

Determining mp

b) Number of electrons - the more electrons, the greater the intermolecular forces.

Exceptions: water, hydrogen fluoride and ammonia because of the strong H - Bonds.

determining mp

c) ***Number and strength of the intermolceular forces*** involved (LD, DD, HB) - the more forces involved the higher the melting point.

Physical properties

To determine melting point you need to look at

1) Number & type of intermolecular forces (water has the highest HB)
2) Number of total electrons (more = high LD forces)

Determining Strengh

Comp	# of e	mp	bp	Force
H_2O	10	0	100	H,DD,LD
H_2S	18	-85.5	-60.7	DD,LD
H_2Se	36	-60.4	-41.5	DD,LD
H_2Te	54	-48.9	-2.2	DD,LD

2) SOLUBILITY

LIKE DISOLVES LIKE

Increased solubility occurs between substances that are both polar (have dipole dipole forces or hydrogen bonds)

3) Shape and Stability of some substances

Expansion of liquid water upon freezing. When water changes to ice, hydrogen bonds leave hexagonal holes making ice less dense than liquid water.