

AP Chemistry Chapter 8 Lecture Notes- Basic Bonding

8.1 Chemical Bonds, Lewis Symbols, and the Octet Rule

- The properties of many materials can be understood in terms of their microscopic properties.
- Microscopic properties of molecules include:
 - the connections between atoms and
 - the 3-D shape of the molecule.
- When atoms or ions are strongly attracted to one another, we say that there is a **chemical bond** between them.
 - In chemical bonds, electrons are shared or transferred between atoms.
- Types of chemical bonds include:
 - ionic bonds** (electrostatic forces that hold ions together, e.g., NaCl);
 - covalent bonds** (result from sharing electrons between atoms, e.g., Cl₂);
 - metallic bonds** (refers to metal nuclei floating in a sea of electrons, e.g., Na).

Lewis Symbols

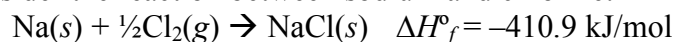
- The electrons involved in bonding are called *valence electrons*.
 - Valence electrons are found in the incomplete, outermost shell of an atom.
- As a pictorial understanding of where the electrons are in an atom, we represent the electrons as dots around the symbol for the element.
 - The number of valence electrons available for bonding are indicated by unpaired dots.
 - These symbols are called **Lewis symbols**, or Lewis electron-dot symbols.
 - We generally place the electrons on four sides of a square around the element's symbol.

The Octet Rule

- Atoms tend to gain, lose, or share electrons until they are surrounded by eight valence electrons; this is known as the **octet rule**.
 - An octet consists of full *s* and *p* subshells.
 - We know that s^2p^6 is a noble gas configuration.
 - We assume that an atom is stable when surrounded by eight electrons (four electron pairs).

8.2 Ionic Bonding

- Consider the reaction between sodium and chlorine:



- The reaction is violently exothermic.

- We infer that the NaCl is more stable than its constituent elements.

- Sodium has lost an electron to become Na⁺ and chlorine has gained the electron to become Cl⁻.
- Note that Na⁺ has an Ne electron configuration and Cl⁻ has an Ar configuration.
- That is, both Na⁺ and Cl⁻ have an octet of electrons.

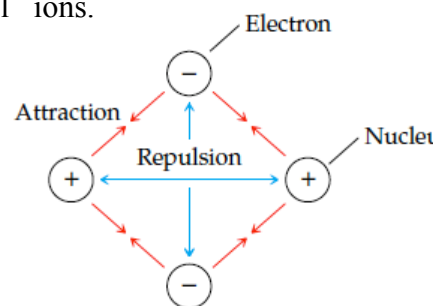
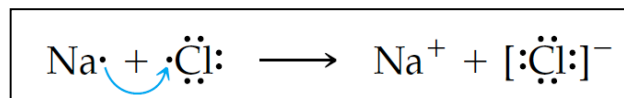
- NaCl forms a very regular structure in which each Na⁺ ion is surrounded by six Cl⁻ ions.

- Similarly each Cl⁻ ion is surrounded by six Na⁺ ions.

- There is a regular arrangement of Na⁺ and Cl⁻.

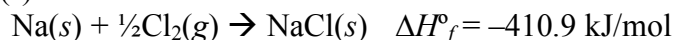
- The ions are packed as closely as possible.

- It is not easy to find a molecular formula to describe the ionic lattice.



Energetics of Ionic Bond Formation

- The heat of formation of NaCl(s) is exothermic:



- Separation of the NaCl into sodium and chloride ions is endothermic:



- The energy required to separate one mole of a solid ionic compound into gaseous ions is called the **lattice energy**, $\Delta H_{\text{lattice}}$.

- Lattice energy depends on the charge on the ions and the size of the ions.

- The stability of the ionic compound comes from the attraction between ions of unlike charge.

- The specific relationship is given by Coulomb's equation:
$$E = k \frac{Q_1 Q_2}{d}$$

$E = k \frac{Q_1 Q_2}{d}$ •where E is the potential energy of the two interacting charged particles, Q_1 and Q_2 are the charges on the particles, d is the distance between their centers, and k is a constant ($8.99 \times 10^9 \text{ J}\cdot\text{m}/\text{C}^2$).
 •As Q_1 and Q_2 increase, E increases; and as d increases, E decreases.

Calculation of Lattice Energies: The Born-Haber Cycle

•The **Born-Haber cycle** is a thermodynamic cycle that analyzes lattice energy precisely.

•Consider a Born-Haber cycle for the formation of $\text{NaCl}(s)$ from $\text{Na}(s)$ and $\text{Cl}_2(g)$.

•The direct route is:



•Alternatively, we can form:

•sodium gas (108 kJ; *endothermic*), then

•chlorine atoms (122 kJ; *endothermic*), then

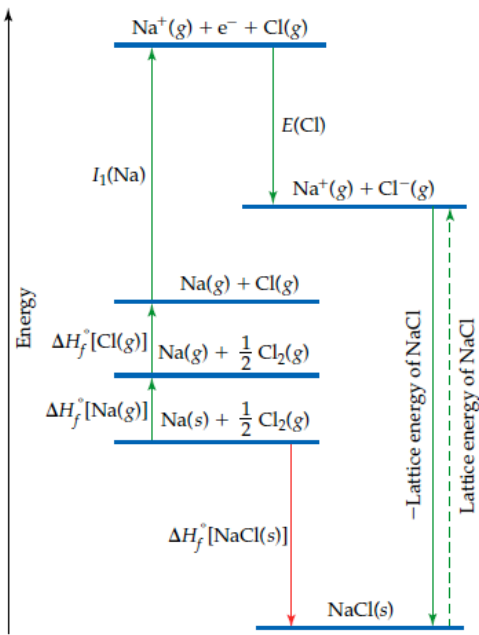
•sodium ions (ionization energy for Na, 496 kJ; *endothermic*), then

•chloride ions (electron affinity for Cl, -349 kJ ; *exothermic*), then

•form the ionic lattice (*exothermic*).

•The sum of the above enthalpies is -411 kJ .

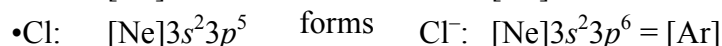
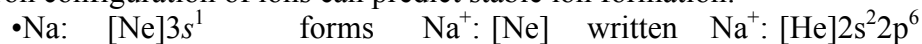
•The lattice energy is $+788 \text{ kJ}$.



Electron Configuration of Ions of the Representative Elements

•These are derived from the electron configuration of elements with the required number of electrons added or removed from the most accessible orbital.

•Electron configuration of ions can predict stable ion formation:



Transition-Metal Ions

•Lattice energies compensate for the loss of up to three electrons.

•We often encounter cations with charges of 1+, 2+ or 3+ in ionic compounds.

•However, transition metals cannot attain a noble gas conformation (>3 electrons beyond a noble gas core).

•Transition metals tend to lose the valence shell electrons first and then as many d electrons as are required to reach the desired charge on the ion.

•Thus, electrons are removed from $4s$ **before** the $3d$, etc..

Polyatomic Ions

•Polyatomic ions are formed when there is an overall charge on a compound containing covalent bonds.

•Examples are NH_4^+ and CO_3^{2-} .

•In polyatomic ions, two or more atoms are bound together by predominantly covalent bonds. The stable grouping carries a charge.

8.3 Covalent Bonding

•The majority of chemical substances do not have characteristics of ionic compounds.

•A chemical bond formed by sharing a pair of electrons is called a *covalent bond*.

•Both atoms acquire noble-gas electronic configurations.

•This is the “glue” that binds atoms together.

Lewis Structures

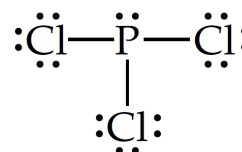
•Formation of covalent bonds can be represented using Lewis symbols.

•The structures are called **Lewis structures**.

•We show each electron pair shared between atoms as a line and show unshared electron pairs as dots.

•Each pair of shared electrons constitutes one chemical bond.

•Example: $\text{H}\cdot + \text{H}\cdot \rightarrow \text{H}:\text{H}$ has electrons on a line connecting the two H nuclei (H–H).



Multiple Bonds

- It is possible for more than one pair of electrons to be shared between two atoms (e.g., **multiple bonding**).
- One shared pair of electrons is a **single bond** (e.g., H₂).
- Two shared pairs of electrons form a **double bond** (e.g., O₂). $:\ddot{\text{N}}\cdot + \cdot\ddot{\text{N}}: \longrightarrow :\text{N}::\text{N}: \quad (\text{or } :\text{N}\equiv\text{N}:)$
- Three shared pairs of electrons form a **triple bond** (e.g., N₂).
- **Bond length** is the distance between the nuclei of the atoms in a bond.
- Generally, bond distances decrease as we move from single through double to triple bonds.

8.4 Bond Polarity and Electronegativity

- The electron pairs shared between two different atoms are usually unequally shared.
- **Bond polarity** describes the sharing of the electrons in a covalent bond.
 - There are two extremes:
 - In a **nonpolar covalent bond** the electrons are shared equally.
 - An example is bonding between identical atoms (example: Cl₂).
 - In a **polar covalent bond**, one of the atoms exerts a greater attraction for bonding electrons than the other (example: HCl)
 - If the difference is large enough, an ionic bond forms (example: NaCl).

Electronegativity

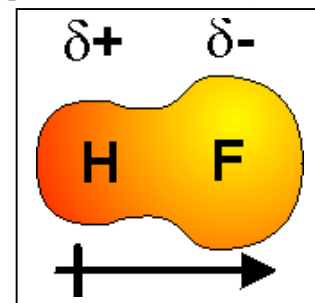
- The ability of an atom *in a molecule* to attract electrons to itself is its **electronegativity**.
- The electronegativity of an element is related to its ionization energy and electron affinity.
- The Pauling electronegativity scale ranges from 0.7 (Cs) to 4.0 (F).
- Electronegativity increases across a period and decreases down a group.

Electronegativity and Bond Polarity

- Electronegativity differences close to zero result in nonpolar covalent bonds.
 - The electrons are equally or almost equally shared.
- The greater the difference in electronegativity is between two atoms, the more polar the bond (polar covalent bonds) is.
- There is no sharp distinction between bonding types.

Dipole Moments

- Molecules like HF have centers of positive and negative charge that do not coincide.
- These are **polar molecules**.
- We indicate the polarity of molecules in two ways:
 - The positive end (or pole) in a polar bond may be represented with a “ $\delta+$ ” and the negative pole with a “ $\delta-$ ”.
- We can also place an arrow over the line representing the bond.
 - The arrow points toward the more electronegative element and shows the shift in electron density toward that atom.
- We can quantify the polarity of the molecule.
 - When charges are separated by a distance, a **dipole** is produced.
 - The **dipole moment** is the quantitative measure of the magnitude of the dipole (μ)
$$\mu = Q r$$
 - The magnitude of the dipole moment is given in *debyes* (D).



Bond Types and Nomenclature

- Previously, we used two different approaches to naming binary compounds.
 - One is for ionic compounds and another is for molecular compounds.
 - In both systems the less electronegative element is given first.
 - The other element follows with the ending *-ide*.
- Both approaches are sometimes used with the same substance!
 - Metals with higher oxidation numbers tend to be molecular rather than ionic.
 - for example: TiO₂
 - The names titanium(IV) oxide and titanium dioxide are used but titanium dioxide is more commonly used.

8.5 Drawing Lewis Structures

- Here are some simple guidelines for drawing Lewis structures:
 - Add up all of the valence electrons on all atoms.
 - For an anion, add electrons equal to the negative charge.
 - For a cation, subtract electrons equal to the positive charge.
 - Identify the central atom.
 - When a central atom has other atoms bound to it, the central atom is usually written first.
 - Example: In CO_3^{2-} the central atom is carbon.
 - Place the central atom in the center of the molecule and add all other atoms around it.
 - Place one bond (two electrons) between each pair of atoms.
 - Complete the octets for all atoms connected to the central atom (exception: hydrogen can only have two electrons).
 - Complete the octet for the central atom; use multiple bonds if necessary.

Formal Charge

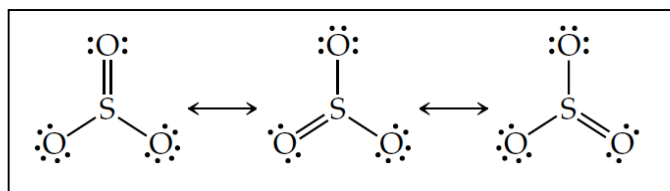
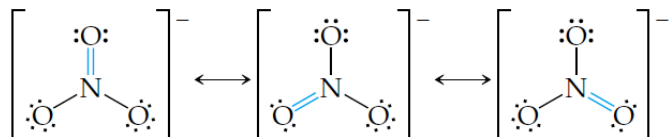
- Sometimes it is possible to draw more than one Lewis structure with the octet rule obeyed for all the atoms.
- To determine which structure is most reasonable, we use formal charge.
- The **formal charge** of an atom is the charge that an atom (in a molecule) would have if all of the atoms had the same electronegativity.
- To calculate formal charge, electrons are assigned as follows:
 - All nonbonding (unshared) electrons are assigned to the atom on which they are found.
 - Half of the bonding electrons are assigned to each atom in a bond.
 - Formal charge is the number of valence electrons in the isolated atom, minus the number of electrons assigned to the atom in the Lewis structure.
- For example: consider CN^- (cyanide ion):
 - For carbon:
 - There are four valence electrons (from the periodic table).
 - In the Lewis structure there are two nonbonding electrons and three electrons from the triple bond.
 - There are five electrons from the Lewis structure.
 - Formal charge: $4 - 5 = -1$.
 - For nitrogen:
 - There are five valence electrons.
 - In the Lewis structure there are two nonbonding electrons and three from the triple bond.
 - There are five electrons from the Lewis structure.
 - Formal charge = $5 - 5 = 0$.
- Use formal charge calculations to distinguish between alternative Lewis structures:
 - the most stable structure has the smallest formal charge on each atom and
 - the most negative formal charge on the most electronegative atoms.
- It is important to keep in mind that formal charges do NOT represent REAL charges on atoms!



8.6 Resonance Structures

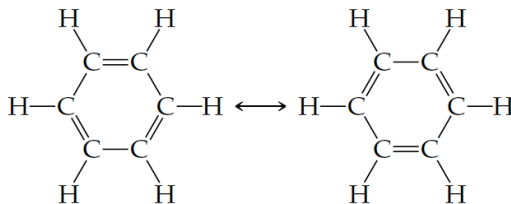
- Some molecules are not adequately described by a single Lewis structure.
 - Typically, structures with multiple bonds can have similar structures with the multiple bonds between different pairs of atoms.
 - Example: Experimentally, ozone has two identical bonds whereas the Lewis structure requires one single (longer) and one double bond (shorter).
- **Resonance structures** are attempts to represent a real structure that is a mix between several extreme possibilities.
 - Resonance structures are Lewis structures that differ only with respect to placement of the electrons.
 - The “true” arrangement is a blend or hybrid of the resonance structures.
- Example: In ozone the extreme possibilities have one double and one single bond.
 - The resonance structure has two identical bonds of intermediate character.
 - We use a double headed arrow (\leftrightarrow) to indicate resonance.

- Common examples are O_3 , NO_3^- , SO_3 , NO_2 , and benzene.



Resonance in Benzene

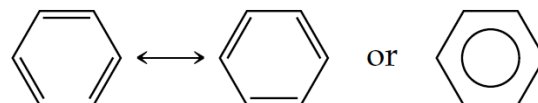
- Benzene belongs to an important category of organic molecules called *aromatic* compounds.



- Benzene (C_6H_6) is a cyclic structure.

- It consists of six carbon atoms in a hexagon.
- Each carbon atom is attached to two other carbon atoms and one hydrogen atom.
- There are alternating double and single bonds between the carbon atoms.
- Experimentally, the C–C bonds in benzene are all the same length.
- Experimentally, benzene is planar.

- To emphasize the resonance between the two Lewis structures (hexagons with alternating single and double bonds), we often represent benzene as a hexagon with a circle in it.



8.7 Exceptions to the Octet Rule

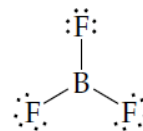
- There are three classes of exceptions to the octet rule:
 - molecules with an odd number of electrons,
 - molecules in which one atom has less than an octet, and
 - molecules in which one atom has more than an octet.

Odd Number of Electrons

- Most molecules have an even number of electrons and complete pairing of electrons occurs although some molecules have an odd number of electrons.
 - Examples are ClO_2 , NO , and NO_2 .

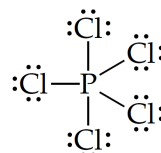
Less than an Octet of Valence Electrons

- Molecules with less than an octet are also relatively rare.
- They are most often encountered in compounds containing boron or beryllium.
 - A typical example is BF_3 .



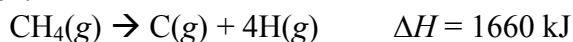
More than an Octet of Valence Electrons

- This is the largest class of exceptions.
- Atoms from the third period and beyond can accommodate more than an octet.
 - Examples are PCl_5 , SF_6 , AsF_6^- , and ICl_4^- .
- Elements from the third period and beyond have unfilled *d* orbitals that can be used to accommodate the additional electrons.
- Size also plays a role.
 - The larger the central atom, the larger the number of atoms that can surround it.
 - The size of the surrounding atoms is also important.
 - Expanded octets occur often when the atoms bound to the central atom are the smallest and most electronegative (e.g., F, Cl, O).



8.8 Strengths of Covalent Bonds

- The energy required to break a covalent bond is called the **bond enthalpy**, D .
 - That is, for the Cl_2 molecule, $D(Cl-Cl)$ is given by ΔH for the reaction: $Cl_2(g) \rightarrow 2Cl(g)$.
- When more than one bond is broken:



- The bond enthalpy is a fraction of ΔH for the atomization reaction:

$$D(C-H) = \frac{1}{4} \Delta H = \frac{1}{4}(1660 \text{ kJ}) = 415 \text{ kJ}$$

- Bond enthalpy is always a positive quantity.

Bond Enthalpies and the Enthalpies of Reactions

- We can use bond enthalpies to calculate the enthalpy for a chemical reaction.
- We recognize that in any chemical reaction bonds need to be broken and then new bonds form.
- The enthalpy of the reaction is given by:
 - the sum of bond enthalpies for bonds broken less the sum of bond enthalpies for bonds formed.
- Where ΔH_{rxn} is the enthalpy for a reaction,

$$\Delta H_{\text{rxn}} = \sum D(\text{bonds broken}) - \sum D(\text{bonds formed})$$
- We illustrate the concept with the reaction between methane, CH_4 , and chlorine:

$$\text{CH}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{CH}_3\text{Cl}(\text{g}) + \text{HCl}(\text{g})$$
 - In this reaction one C–H bond and one Cl–Cl bond are broken while one C–Cl bond and one H–Cl bond are formed.
 - So $\Delta H_{\text{rxn}} = [D(\text{C–H}) + D(\text{Cl–Cl})] - [D(\text{C–Cl}) + D(\text{H–Cl})] = -104 \text{ kJ}$.
 - The overall reaction is exothermic which means that the bonds formed are stronger than the bonds broken.
 - The above result is consistent with Hess's law.

Bond Enthalpy and Bond Length

- The distance between the nuclei of the atoms involved in a bond is called the **bond length**.
- Multiple bonds are shorter than single bonds.
 - We can show that multiple bonds are stronger than single bonds.
 - As the number of bonds between atoms increases, the atoms are held closer and more tightly together.
 - ❖ Multiple bonds are shorter than single bonds

TABLE 8.4 Average Bond Enthalpies (kJ/mol)

Single Bonds

C–H	413	N–H	391	O–H	463	F–F	155
C–C	348	N–N	163	O–O	146		
C–N	293	N–O	201	O–F	190	Cl–F	253
C–O	358	N–F	272	O–Cl	203	Cl–Cl	242
C–F	485	N–Cl	200	O–I	234		
C–Cl	328	N–Br	243			Br–F	237
C–Br	276			S–H	339	Br–Cl	218
C–I	240	H–H	436	S–F	327	Br–Br	193
C–S	259	H–F	567	S–Cl	253		
		H–Cl	431	S–Br	218	I–Cl	208
Si–H	323	H–Br	366	S–S	266	I–Br	175
Si–Si	226	H–I	299			I–I	151
Si–C	301						
Si–O	368						
Si–Cl	464						

Multiple Bonds

C=C	614	N=N	418	O ₂	495
C≡C	839	N≡N	941		
C=N	615	N=O	607	S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

Practice Quiz Answer Key

- | | |
|-------|-----------------|
| 1. C | 18. E |
| 2. A | 19. D |
| 3. B | 20. D |
| 4. A | 21. A |
| 5. E | 22. A |
| 6. D | 23. D |
| 7. A | 24. B |
| 8. B | 25. C |
| 9. D | 26. E |
| 10. B | 27. C |
| 11. A | 28. C |
| 12. B | 29. C |
| 13. A | 30. A |
| 14. B | 31. A |
| 15. B | 32. D |
| 16. B | 33. E |
| 17. B | 34. endothermic |

Chapter 8 Homework Assignments

Sections 8.1- 8.4 pg 336 #8, 12, 16, 19, 20, 22, 25, 29, 30, 31, 33, 37, 39

Sections 8.5- 8.7 pg 337 #46, 47, 48, 50, 53, 55, 58, 60, 61, 64

Section 8.8 pg 338 #65(a&b), 69, 88(a&b), 90, 99

AP Chemistry Chapter Eight Practice Test
Basic Concepts of Chemical Bonding

- 1) There are _____ paired and _____ unpaired electrons in the Lewis symbol for a phosphorus atom.
A) 4, 2 B) 2, 4 C) 2, 3
D) 4, 3 E) 0, 3
- 2) Based on the octet rule, Mg most likely forms a _____ ion.
A) Mg^{2+} B) Mg^{2-} C) Mg^{6-}
D) Mg^{6+} E) Mg^{-}
- 3) Phosphorus most likely forms a _____ ion.
A) P^{3+} B) P^{3-} C) P^{5+}
D) P^{5-} E) P^{+}
- 4) How many unpaired electrons are there in the Lewis structures of a N^{3-} ion?
A) 0 B) 1 C) 2
D) 3 E) This cannot be predicted.
- 5) The electron configuration of the phosphide ion (P^{3-}) is _____.
A) $[Ne]3s^2$ B) $[Ne]3s^2 3p^1$ C) $[Ne]3s^2 3p^3$
D) $[Ne]3p^2$ E) $[Ne]3s^2 3p^6$
- 6) The halogens, alkali metals, and alkaline earth metals have _____ valence electrons, respectively.
A) 7, 4, and 6 B) 1, 5, and 7
C) 8, 2, and 3 D) 7, 1, and 2
E) 2, 7, and 4
- 7) The principal quantum number of the electrons that are lost when tungsten forms a cation is _____.
A) 6 B) 5 C) 4
D) 3 E) 2
- 8) What is the electron configuration for the Co^{2+} ion?
A) $[Ar]4s^1 3d^6$ B) $[Ar]4s^0 3d^7$
C) $[Ar]4s^0 3d^5$ D) $[Ar]4s^2 3d^9$
E) $[Ne]3s^2 3p^{10}$
- 9) The formula of palladium(IV) sulfide is _____.
A) Pd_2S_4 B) PdS_4 C) Pd_4S
D) PdS_2 E) Pd_2S_2
- 10) How many single covalent bonds must a silicon atom form to have a complete octet in its valence shell?
A) 3 B) 4 C) 1
D) 2 E) 0

11) A _____ covalent bond between the same two atoms is the longest.

- A) single B) double C) triple
D) They are all the same length. E) strong

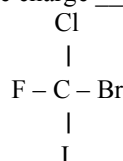
12) A double bond consists of _____ pairs of electrons shared between two atoms.

- A) 1 B) 2 C) 3
D) 4 E) 6

13) What is the maximum number of double bonds that a hydrogen atom can form?

- A) 0 B) 1 C) 2
D) 3 E) 4

14) In the molecule below, which atom has the largest partial negative charge _____?

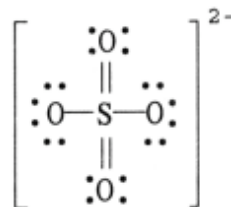


- A) Cl B) F C) Br
D) I E) C

15) The Lewis structure of AsH_3 shows _____ nonbonding electron pair(s) on As.

- A) 0 B) 1 C) 2
D) 3 E) This cannot be determined from the data given.

16) The formal charge on sulfur in SO_4^{2-} is _____, where the Lewis structure of the ion is:



- A) -2 B) 0 C) +2
D) +4 E) -4

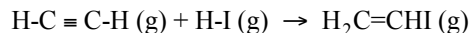
17) In the Lewis structure of ClF , the formal charge on Cl is _____ and the formal charge on F is _____.

- A) -1, -1 B) 0, 0 C) 0, -1
D) +1, -1 E) -1, +1

18) How many equivalent resonance structures can be drawn for the molecule of SO_3 without having to violate the octet rule on the sulfur atom?

- A) 5 B) 2 C) 1
D) 4 E) 3

19) Using the table of average bond energies below, the ΔH for the reaction is _____ kJ.



Bond:	C \equiv C	C=C	H-I	C-I	C-H
D (kJ/mol):	839	614	299	240	413

- A) +506 B) -931 C) -506
 D) -129 E) +129

20) The chloride of which of the following metals should have the greatest lattice energy?

- A) potassium B) rubidium C) sodium
 D) lithium E) cesium

21) Lattice energy is _____.

- A) the energy required to convert a mole of ionic solid into its constituent ions in the gas phase
 B) the energy given off when gaseous ions combine to form one mole of an ionic solid
 C) the energy required to produce one mole of an ionic compound from its constituent elements in their standard states
 D) the sum of ionization energies of the components in a salt
 E) the sum of electron affinities of the components in a salt

22) Of the atoms below, _____ is the least electronegative.

- A) Rb B) F C) Si
 D) Cl E) Ca

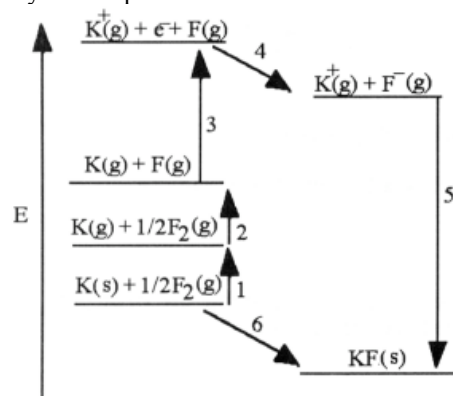
23) Of the molecules below, the bond in _____ is the most polar.

- A) HBr B) HI C) HCl
 D) HF E) H₂

24) Of the bonds below, _____ is the least polar.

- A) Na, S B) P, S C) C, F
 D) Si, Cl E) Na, Cl

The diagram below is the Born-Haber cycle for the formation of crystalline potassium fluoride.



25) Which energy change corresponds to the electron affinity of fluorine?

- A) 2 B) 5 C) 4
 D) 1 E) 6

26) Using the Born-Haber cycle, the ΔH_f° of KBr is equal to

- A) $\Delta H_f^\circ[\text{K (g)}] + \Delta H_f^\circ[\text{Br (g)}] + I_1(\text{K}) + E(\text{Br}) + \Delta H_{\text{lattice}}$
 B) $\Delta H_f^\circ[\text{K (g)}] - \Delta H_f^\circ[\text{Br (g)}] - I_1(\text{K}) - E(\text{Br}) - \Delta H_{\text{lattice}}$
 C) $\Delta H_f^\circ[\text{K (g)}] - \Delta H_f^\circ[\text{Br (g)}] + I_1(\text{K}) - E(\text{Br}) + \Delta H_{\text{lattice}}$
 D) $\Delta H_f^\circ[\text{K (g)}] + \Delta H_f^\circ[\text{Br (g)}] - I_1 - E(\text{Br}) + \Delta H_{\text{lattice}}$
 E) $\Delta H_f^\circ[\text{K (g)}] + \Delta H_f^\circ[\text{Br (g)}] + I_1(\text{K}) + E(\text{Br}) - \Delta H_{\text{lattice}}$

27) The Lewis structure of N₂H₂ shows _____.

- A) a nitrogen-nitrogen triple bond
 B) a nitrogen-nitrogen single bond
 C) each nitrogen has one nonbinding electron pair
 D) each nitrogen has two nonbinding electron pairs
 E) each hydrogen has one nonbonding electron pair

28) To convert from one resonance structure to another,

- A) only atoms can be moved
 B) electrons and atoms can both be moved
 C) only electrons can be moved
 D) neither electrons nor atoms can be moved
 E) electrons must be added

29) For resonance forms of a molecule or ion, _____.

- A) one always corresponds to the observed structure
 B) all the resonance structures are seen in various proportions
 C) the observed structure is an average of the resonance forms
 D) the same atoms need not be bonded to each other
 E) there cannot be more than two resonance structures for a given species

30) Bond enthalpy is _____.

- A) always positive B) always negative
 C) sometimes positive, sometimes negative
 D) always zero E) unpredictable

31) Given that the average bond energies for C-H and C-Br bonds are 413 and 276 kJ/mol, respectively, the heat of atomization of bromoform (CHBr₃) is _____ kJ/mol.

- A) 1241 B) 689 C) -689
 D) 1378 E) -1378

32) As the number of covalent bonds between two atoms increases, the distance between the atoms _____ and the strength of the bond between them _____.

- A) increases, increases B) decreases, decreases
 C) increases, decreases D) decreases, increases
 E) is unpredictable

33) Most explosives are compounds that decompose rapidly to produce _____ products and a great deal of _____.

- A) gaseous, gases B) liquid, heat
 C) soluble, heat D) solid, gas
 E) gaseous, heat

34) In a reaction, if the bonds in the reactants are stronger than the bonds in the product, the reaction is _____.