

CHAPTER 3: CHEMICAL PERIODICITY

INTRODUCTION

Chapter 3 emphasizes chemical principles and structures, with very few calculations. You will be studying the similar and not-so-similar chemical behavior of groups of elements arranged according to the Periodic Table. The concept of electronegativity will enable you to distinguish between the two principal kinds of chemical bonds – ionic and covalent. You will learn how to draw Lewis diagrams to show the structure of covalent molecules, and how to use the VSEPR theory to predict the shapes of molecules. Finally, you will learn the systematic ways of naming various sorts of compounds.

REVIEW

You should review the elementary atomic structural principles covered in Chapter 1, as these are essential in our study of chemical bonding. The material on chemical equations and stoichiometric calculations that you covered in Chapters 1 and 2 is not emphasized in this chapter. Instead, the properties and bonding in individual compounds and molecules are studied. However, you will have to use chemical calculations all through the book, so keep your skills current.

3-1 Groups of Elements

Chapter 3 begins with a descriptive survey of seven different groups of elements, to show the kind of evidence that led to the development of the Periodic Table in the late 19th century. Read this section, noting the types of overall behavior that distinguish each group. Learn the names of the important groups of elements discussed: the alkali and alkaline earth metals, the chalcogens and halogens, and which elements are included in each. Then go on to read Section 3-2 on the periodic table. Look at the periodic table inside the front cover of the text or of the Study Guide, and fix in your mind the general location of the **metals**, **nonmetals**, and **metalloids** or **semi-metals**. Check with your instructor about how much detail from Section 3-1 you are expected to learn at this time.

3-2 The Periodic Table

During the first half of the 19th century more than 60 elements were identified and their chemical and physical characteristics were described. It was discovered that by listing the known

CHAPTER 3: CHEMICAL PERIODICITY

elements in order of increasing atomic mass, the chemical and physical properties of the elements could be described as periodic functions of their atomic masses.

In such a listing the chemical and physical properties tended to repeat every seventh element. (This appeared to be the case in 1860, before the discovery of the noble gas elements. If we include the **noble gases**, the properties repeat every eighth element.) Mendeleev and Meyer independently produced the first **periodic table** in which the list of elements begins a new **row** every seventh element (every eighth element today) as the properties repeat. The elements in a **column** of the periodic table belong the same **group**, and share very similar properties.

Where the periodicity broke down, Mendeleev left gaps in the periodic table which he predicted would eventually be filled by newly discovered elements. The atomic masses and other properties of these new elements turned out to be very close to the values predicted by Mendeleev.

The modern periodic table is listed not by increasing atomic mass, but by increasing atomic number. Recall from Chapter 1 that the atomic number equals the number of positively-charged **protons** in the nucleus, and also the number of negatively-charged **electrons** surrounding the nucleus in an electrically neutral atom. In this and later chapters we will see that the number of electrons is the main factor determining an element's chemical properties.

The relative atomic mass is approximately equal to sum of the masses of the protons and the neutrons in the nucleus. Ordering elements by atomic mass and ordering by atomic number (number of protons) lead to approximately the same Periodic Table, but there are a few exceptions that establish atomic number as the correct basis for the Table.

At this point in your study, go back to Section 3-1, and using the periodic table inside the cover of the text, locate the named groups: **alkali metals**, **alkaline earth metals**, **chalcogens**, **halogens**, and also the noble or inert gases. Each occupies a column of the table. These names are important, and you should learn the group number associated with each. These elements are called **main** group or **representative** elements. We study them early on, because their chemistry follows regular, predictable patterns. The **transition elements**, in the center section of the table, and the **lanthanide** and **actinide** elements in the two strips at the bottom, have more complicated behavior that is covered later.

Electronegativity, an Important Periodic Property

Electronegativity, the relative ability of an atom to attract electrons in a bonding situation, is used in Chapter 3 and particularly in Chapter 17, to predict the type of chemical bonds formed between any pair of elements. For now you should fix the basic principles:

- Metals are relatively **electropositive** (have *low* electronegativities). They easily give up one or more electrons in forming bonds with non-metals.

- Non-metals are relatively **electronegative** (have high electronegativities). They tend to gain one or more electrons in forming bonds with metals.
- The lowest electronegativities belong to the alkali and alkaline-earth elements. The highest electronegativities are those of the elements in the upper right corner of the Periodic Table.

3-3 Ions and Ionic Compounds

Two elements in the same group of the periodic table have different numbers of electrons, but have similar chemical properties because they have the same number of outer electrons arranged in similar spatial patterns. The **Lewis electron-dot model** of the atom divides an element's electrons into two classes: **core electrons**, which are held close to the nucleus and are *not* significantly involved in chemical bonding, and **valence electrons**, which occupy the outer layer of the atom and *do* participate in chemical bonds. Among main-group elements, the number of valence electrons is the same as the group number, which accounts for why elements in a single group exhibit similar chemical behavior.

The Lewis electron-dot model best accounts for the behavior of main-group elements. The electron structures of the transition metals, lanthanides, and actinides are too complicated to be successfully described by the simple Lewis model.

The Lewis model ignores the core electrons, and represents the valence electrons by dots around the four sides of the symbol for an element. You should be able to apply the rules for writing Lewis dot symbols to any main-group element:

- The first four electrons are written as lone dots on the four sides of the symbol.
- The next four electrons are paired with electrons already present.

The heart of the Lewis model is the **octet rule**: main-group atoms are most stable in chemical combination when they have eight valence electrons. The noble gas elements are chemically inert because they already contain a stable octet. Other elements have a strong drive to form that stable octet, either by forming ions (Section 3-3) or by sharing electron pairs with other atoms (Section 3-4).

- Atoms in Groups I and II (those with lowest electronegativities) tend to achieve a stable valence electron *octet* by losing one (Group I) or two (Group II) electrons to form a **cation** (a positively charged ion).
- Atoms in Groups VI and VII (those with highest electronegativities) tend to achieve a stable valence electron *octet* by gaining one (Group VII) or two (Group VI) electrons to form an **anion** (a negatively charged ion).

CHAPTER 3: CHEMICAL PERIODICITY

- Group VI and VII nonmetals tend to form ionic bonds in reaction with elements in Groups I and II metals. This comes about because of their very different *electronegativities*.
- The very smallest atoms, H, Li, Be, B, cannot crowd a full octet of electrons into the limited space around themselves, and are satisfied with 0 or 2, 2, 4, and 6, respectively.

Formation, Names, and Formulas of Ionic Compounds

Follow through the examples of how to use the charges on the ions to deduce the formulas of compounds. The basic idea is to end up with equal numbers of positive and negative charges, so the compound has a net zero charge (**charge neutrality**). Compounds are named by naming the cation (positive ion) first, then the negative ion. Notice how the endings of the names of the nonmetals change to *-ide*, when they are in ionic form.

Two key facts about ionic compounds:

- (1) They are solids at room temperature and do not contain individual molecules.
- (2) They are called *salts* if they do not contain oxide (O^{2-}) or hydroxide (OH^-) ions.

Table 3-2 lists the common monatomic negative ions (anions) and positive ions (cations). By now you should be able to name these elements and give their symbols. Now you should learn their ionic charges and the relation of these to the periodic group. Practice writing formulas of the compounds formed between any of these cations and anions.

3-4 Covalent Bonding and Lewis Structures

In reactions involving only nonmetals and semimetals, atoms tend to share electrons, forming **covalent bonds**. This is because the nonmetals and semimetals of Groups IV to VII do not differ enough in their power to attract electrons (in their electronegativities) to transfer electrons completely, so they instead form stable valence *octets* by sharing electrons with each other. Compounds formed in this way are called **covalent compounds**. Compounds which contain only covalent bonds (with no ionic bonds) are usually molecular in nature.

Lewis structures

Numerous molecular structures are shown, called **Lewis structures**, in which pairs of electrons around each atom are shown as pairs of dots. In the drawings of ball-and-stick models the sticks represent electron pairs in a bond. Atoms other than H usually share enough electrons so that each has a full octet. Such simple pictures of molecules are very helpful in describing shapes of molecules and predicting reactions.

Electron pairs that are not involved in a bond are called **lone pairs**. Those pairs which form a bond are **bonding pairs**. They can be shown as a single dash rather than two dots between the atoms. Remember, the octet of electrons counted with an atom includes *both* electrons in a bond to that atom. Hydrogen, the smallest atom, can only share in *one* pair of electrons. Other atoms acquire an octet, with the exceptions noted in Section 3-5.

Notice that **multiple bonding** occurs, when there are too few electrons to fill out octets by sharing single pairs. A single pair forms a **single bond**, two pairs form a **double bond**, and three pairs form a **triple bond**. There are no quadruple bonds. Table 3-3 shows average *bond lengths* in simple molecules. The bond length is the distance between the centers of the bonded atoms. Notice that multiple bonds are shorter than single bonds.

Formal Charges

We now tackle a method of counting electrons, **formal charge**. This will be useful in Section 3-5, where we deal with how to draw Lewis structures. Remember: the formal charge is a bookkeeping device in a covalent molecule, not an actual ionic (electric) charge. Study the authors' formula and Example 3-4 on pp.108-109 to learn how to calculate formal charge on atoms in molecules. Simply put, the formal charge counts the net number of electrons an isolated atom gives up or gains when it forms covalent bonds. When choosing between possible Lewis structures, we choose the ones with the smallest formal charges on the atoms, as demonstrated in Section 3-5.

3-5 Drawing Lewis Structures

Lewis structures of simple molecules can be drawn by trial and error, but if several atoms from groups V and VI are present, a more formal method sometimes works better. The rules on p. 110 work very well. You count all the valence electrons available on the atoms (A), and subtract this from the number (N) needed to give each atom individually an octet. (Hydrogen only needs two, remember.) The result (S) is the number available to be shared in bonds.

$$S = N - A$$

Arrange the atoms on your paper. Place single bonds between the atoms, then use up any of the remaining S electrons to form double or triple bonds. Then add in the rest of the available (total A) electrons to form lone pairs and fill out the octets, always counting all the electrons in the bonds with both bonded atoms.

Finally, if you have found more than one **isomer**, i.e. more than one possible structure, which frequently happens, add up the formal charges on all atoms, and choose the structure that has smallest values. Reject structures that give large positive formal charges to electronegative atoms. The formal charges must add up to give zero for a neutral molecule, or to give the ionic charge if you are working with an ion.

CHAPTER 3: CHEMICAL PERIODICITY

Follow through Examples 3-5, 6, and 7, and practice with problems 31-54 at the end of the chapter.

Resonance Structures and Breakdown of the Octet Rule

There are many cases in which the simple Lewis octet rule is not adequate to describe the molecules that nature gives us. **Resonance structures** are different but equally correct Lewis structures for a molecule, and no single one of them really pictures the molecule accurately. In the example of SO_2 , both S-O bonds are known experimentally to be identical, and the true structure (for now, at least) may be thought of as an average or **resonance hybrid** between the two Lewis structures. Typical cases of resonance are the sulfur oxides, ozone, and the ions NO_3^- and CO_3^{2-} . In Chapters 16 and 17 we find that electrons are not correctly described as point charges, and in Section 18-3 we learn more about how the electron density is distributed throughout real molecules, and more about the limitations of Lewis structures. The three cases of *exceptions to the octet rule* should seem reasonable to you, and you should be able to pick out examples if given a list of molecular formulas.

- Case 1: If there are is odd number of electrons available, then all electrons cannot be paired up. The molecule is then called an **odd-electron molecule**.
- Case 2: Very small atoms (H, Li, Be, and B) form **octet-deficient molecules**. They don't have enough room to permit eight electrons to surround them.
- Case 3: Larger atoms can accommodate one or two extra electron pairs beyond the octet, a so-called **valence-shell expansion**. Reasons for this will become evident in later chapters.

3-6 Naming Compounds in Which Covalent Bonding Occurs

We have to deal with two kinds of compounds here: binary molecular compounds and ionic compounds in which one or both of the ions is itself polyatomic – a group of atoms covalently bonded together.

Naming Binary Molecular Compounds

If a pair of elements forms only one compound, the naming is simple: HBr is hydrogen bromide, BeCl_2 is beryllium chloride, etc. At this stage of your study it may not be obvious which pairs form only one compound, so if you use the systematic name (beryllium dichloride), that is correct also. Note the greek prefixes used to state the numbers of atoms, in Table 3-4, and follow the examples in the text. Thus, P_4O_{10} is tetraphosphorus decaoxide, and so on. Memorize the common (unsystematic) names of the eight compounds on p. 120.

Naming Compounds that Contain Polyatomic Ions

The **polyatomic** or **molecular** ions are listed in Table 3-5. Only two are cations, hydronium, H_3O^+ and ammonium, NH_4^+ . Of the many anions, most are **oxoanions**, in which a central atom like S, P, N, or Cl is covalently bonded to several O atoms. These oxoanions stay together as a single ion in chemical reactions. You should start learning these names and formulas now, even though you may be confused as to how the names are arrived at. Writing formulas of compounds containing polyatomic ions is done by balancing the positive and negative charges, just like with binary ionic compounds. Naming the compounds also follows the rules used with binary compounds, except for the endings of the anion names. Notice that only the first three anions containing oxygen have names that end in *-ide*. The others end in *-ate* or *-ite*.

3-7 The Shapes of Molecules

The VSEPR Theory

This section describes a common-sense way to figure out the shape of a polyatomic molecule by considering the mutual repulsion of the electron pairs around the central atom in its Lewis diagram. It is called the **Valence Shell Electron Pair Repulsion** or **VSEPR** theory. The repulsion between electron pairs forces them to locate as far apart as possible (like hostile bugs on a billiard ball). The repulsion occurs both for bonding pairs and for lone pairs, with somewhat more repulsion where lone pairs are involved. Two pairs grouped in a double bond have about the same effect as the single pair in a single bond. Notice that it is the *electron pairs or groups* that assume the resulting positions. To a good approximation the atoms at the other end of each bond just go along for the ride.

Figures 3-17, 18, 19, 20, and 21 show the result. We define the **steric number (SN)** to be the number of electron *groups* around the central atom - pairs in bonds or in lone pairs and groups of electrons in double or in triple bonds. (Note: count only one group for a double or triple bond.) You should learn the shapes of the figures corresponding to SNs from two to six: 2 linear; 3 trigonal planar; 4 tetrahedral; 5 trigonal bipyramidal; 6 octahedral (see Figure 3-17). Several of these shapes require a central atom with an expanded octet, i.e. more than 4 electron pairs around it.

Take note of the bond angles around the central atom in each of these cases. Be sure to state whether you are describing the arrangement of electron *groups* (or *pairs*) in the molecule or the arrangement of *atoms*. The arrangement of *atoms* is usually called the geometry of the molecule, but the arrangement of electron pairs determines the location of the atoms.

Note particularly that in the molecules NH_3 and H_2O , the four electron pairs are arranged approximately tetrahedrally (Figures 2-19 and 20). However, the atoms are said to be in a

CHAPTER 3: CHEMICAL PERIODICITY

trigonal or triangular pyramid in NH_3 and in a bent structure in H_2O . Because of the larger repulsion (or larger volume occupied) by lone pairs, the H-atoms are pushed closer together, and the H-N-H and H-O-H bond angles are somewhat smaller than the standard tetrahedral value of 109.5° . Figures 3- 21 and 22 show how the shapes of SF_4 , ClF_3 , XeF_2 , and IF_5 reflect the presence of lone pairs. Learn to use the terms **axial** and **equatorial** in describing locations of electron pairs and atoms.

Dipole Moments

The shapes of molecules strongly influence the forces of attraction between them. This is the major reason we learn to assign shapes to molecules. Because of differences in electronegativity, atoms attract the electron pairs in covalent bonds unequally. This gives rise to a *bond dipole*, that is, a partial separation of charge in any bond between unlike atoms. Adding up all the bond dipoles in the molecule gives the resulting **dipole moment** of the whole molecule. The larger the dipole moment, the more strongly the molecule attracts others. See Section 6-1 for more details.

To summarize, we may state that linear and nonlinear, symmetrical molecules (examples N_2 , O_2 , CO_2 , CCl_4 , and SO_3) have zero dipole moment. Even with identical bonds, in asymmetrical molecules like NH_3 and H_2O , the bond dipoles do not cancel each other out, so they have non-zero dipole moments. Molecules with net non-zero dipole moments are called *polar molecules*. You should analyze the molecules pictured in Figure 3-23, and be able to predict whether other molecules will be polar or not. Note that the CCl_4 molecule is tetrahedral, even though on first glance Figure 3-23 (d) may not appear tetrahedral.

In the Chemistry in Your Life essay entitled **Dry Ice Versus Wet Ice** one effect of the zero dipole moment of carbon dioxide is explored. Solid CO_2 vaporizes away while remaining at -78°C . It's all explained here! When under very high pressure, superfluid CO_2 is used to decaffeinate coffee, as described in another Chemistry in Your Life essay in Chapter 6.

3-8 Elements Forming More Than One Ion

Transition metals and the heavier elements in Groups III, IV, and V have more complex chemistry than the metals of Groups I and II. They form more than one ion with different charges, and they do not follow the octet rule, so Lewis structures are not ordinarily used for their compounds. Their compounds are named differently, in a way that includes their ionic charge as a Roman numeral in the name; for example, FeCl_3 is iron(III)chloride. Familiarize yourself with how this is done. The charge on any monoatomic ion, whether positive or negative, is also called its **oxidation number** or **oxidation state**.

In later chapters oxidation numbers will be found useful in analyzing some kinds of reactions and in systemizing chemical facts. We shall even assign oxidation numbers to atoms in

covalent and partially covalent compounds. In those cases the oxidation numbers are a sort of “fake ionic charge” that we use for bookkeeping. As an example lead(II)chloride, PbCl_2 , contains Pb^{2+} ions, while lead(IV)chloride, PbCl_4 , does not contain Pb^{4+} ions. For now learn how to name these compounds; the chemical details will come later.

Coordination complexes are ions in which a transition metal ion is covalently bonded to from 2 to 6 anions or neutral molecules, which are called **ligands**. The electron pairs used in the coordinate covalent bonds are contributed by the ligands, which use their lone pairs for the purpose. The authors describe some of the colors and other phenomena associated with coordination complexes, but put off extensive discussion until Chapter 19.

Charge, Formal Charge, and Oxidation State

This last section should be read now and *remembered throughout the course*. The distinctions between these ways of assigning numbers to atoms, and when they are used, are clearly stated. Actual *ionic charge* is due to a *real* difference between the number of electrons in the ion and the number the same neutral atom or atoms would have. *Formal charge* is artificially assigned to atoms in molecules that are treated as though they were truly *covalent*. It is useful only for deciding on the best Lewis structure of a molecule. *Oxidation state* is artificially assigned to atoms in compounds by treating them as though they were ionic. It is useful for naming compounds and for balancing oxidation-reduction reactions (Chapter 12).

LEARNING GOALS . . . CHAPTER 3

KEY TERMS - Define or explain the following terms in your own words.

metals, nonmetals
 metalloids, semimetals
 periodic law
 alkali metals
 alkaline earth metals
 chalcogens
 halogens
 alloys
 periodic table, periodic law
 noble gases
 groups, periods
 representative elements, main group elements
 transition elements
 lanthanide elements
 actinides
 electronegativity

CHAPTER 3: CHEMICAL PERIODICITY

electropositive, electronegative elements
Lewis electron-dot model
core electrons, valence electrons, valence shell
Lewis dot symbol, Lewis structure
ion, cation, anion
octets, octet rule
monatomic, polyatomic ions, molecular ions
oxoanions
charge neutrality
covalent bonds, covalent compounds
lone pairs, bonding pairs
double bond, triple bond
formal charge
isomers
resonance hybrid
valence-shell expansion
Valence Shell Electron Pair Repulsion (VSEPR) Theory
steric number (SN)
axial and equatorial sites
polar molecules
bond dipole, dipole moment
oxidation state, oxidation number
coordination complexes, ligands

STUDY QUESTIONS AND PROBLEM TYPES

Section 3-1

1. Name the seven groups of elements described in Section 3-1. For each group, describe some properties the elements belonging to the group have in common. How do these properties differ from the properties of other groups? Name some elements belonging to each group.

Section 3-2

2. What is the *periodic law*? Why is the periodic table called periodic?
3. What are the exceptions to the periodic law? Can you explain these exceptions?
(Hint: review mass spectrometry in Section 1-6.)
4. Explain how Mendeleev was able to predict the atomic mass and chemical properties of several previously undiscovered elements.

5. Be able to identify the parts making up the periodic table: groups, periods, etc. On the periodic table in Figure 3-14, identify the general regions corresponding to metals, semi-metals and metalloids. What are the properties that distinguish metals, semi-metals and metalloids? Locate the alkali and alkaline earth metals, the chalcogens, the inert gases.

6. What is electronegativity and how does it help predict the kinds of bonds an element forms?

Section 3-3

7. What are *valence electrons* and how do they differ from *core electrons*? From memory, draw the Lewis dot symbols for elements in the first three periods of the periodic table. For *any* main group element, given its atomic number and group number, calculate how many core and valence electrons it has and draw its Lewis structure.

8. How do elements from groups I, II, VI, and VII form a valence octet? Choose one or two elements from each of these groups. For each such element write a balanced chemical equation showing how the valence octet is formed, writing the Lewis structures on both sides of the equation.

9. For any cation from group I or II and for any anion from group VI or VII, write their Lewis symbols and predict the formula of a compound formed between them.

10. Describe the naming conventions for binary ionic compounds. For any pair of anions and cations in Table 3-2, write the formula of the compound they form and write its name.

Section 3-4

11. Describe the *octet rule* for *covalent* and *ionic* compounds. How are the octets different for covalent and ionic compounds? Which periodic groups tend to form covalent bonds, and which tend to form ionic bonds?

12. Describe *single*, *double*, and *triple bonds* in terms of valence electrons and Lewis symbols. What are *lone pairs*? Write Lewis structures for Cl_2 , N_2 , S_2 .

13. What is a *formal charge*? How is it calculated? What is it used for? Is the formal charge on an atom the same as the actual charge on that atom in a compound?

14. For several covalent molecules, compute the formal charges on the atoms. Do not just apply the algebraic expression for formal charge in a mechanical way! Explain *how* the expression for formal charge is derived. At each step in your calculation, express in words what the values you compute represent.

Section 3-5

15. Use pencil and paper to follow through the drawing of Lewis structures in Examples 3-5, 3-6, and 3-7, and calculate the formal charges on all the atoms. Then do the Exercise following each of these examples.
16. What is a *resonance hybrid*? Give some examples. How does a resonance hybrid structure remedy a shortcoming of the Lewis structure?
17. Describe the three kinds of *exceptions* to the octet rule for Lewis structures. Give an example of each.

Section 3-6

18. Write the systematic name of any binary compound whose formula is given. Write the formula of any binary compound whose name is given.
19. Give the common names and formulas of several compounds commonly called by nonsystematic names.
20. Work on memorizing the names and formulas of the common molecular ions in Table 3-5. All will be used during the full course. Ask your instructor for the ones you should learn immediately. We suggest starting with H_3O^+ , NH_4^+ , OH^- , CO_3^{2-} , NO_3^- , PO_4^{3-} , SO_4^{2-} , ClO_3^- , and CN^- . Then work on the anions related to each of these ions that ends in *-ate*.
21. Write formulas for many compounds containing polyatomic ions, given their names. Write names of such compounds, given their formulas.

Section 3-7

22. Explain why repulsion between electron pairs enables you to predict the three-dimensional geometry of a molecule? What do the letters V, S, E, P, R stand for in the name of the VSEPR theory.
23. What is the steric number SN of an atom in a molecule? How can you get the SN from the atom's Lewis structure? Draw the molecular geometries associated with SN = 2 to 5.
24. Why do lone pair electrons have greater steric demands than bonding pairs? What consequences does this have for the molecular geometry of H_2O and NH_3 ? Where do you tend to find lone pairs in SN = 5 structures?
25. Use the VSEPR model to predict the geometry of any molecule whose central atom belongs to one of the main group elements of the first three rows of the periodic table.

26. Draw the Lewis structure and predict the molecular shape for the polyatomic ions SO_4^{2-} , NH_4^+ , and CN^- . Be sure to include the ionic charge when you count the electrons.

27. For a polyatomic molecule, explain how knowing the *bond dipole* of each bond allows you to predict the total molecular dipole moment. For any given molecule draw the Lewis dot structure, use the VSEPR theory to get the molecular shape, and predict whether the net dipole moment is zero or non-zero.

Section 3-8

28. Write the formula for any binary transition metal compound, given its systematic name, including the oxidation number. Given the formula, write the systematic name, including the oxidation number.

29. Discuss whether (and when) oxidation number is the same thing as ionic charge.

30. Describe the type of bonds formed between the central atom or ion and the ligands in a coordination complex.

31. Why does a $\text{Cu}(\text{NO}_3)_2$ (*aq*) solution change color, when aqueous ammonia is added?

32. Explain the difference between formal charge and oxidation number. Which one can be the same as an actual ionic charge?

CHAPTER 3: PRACTICE EXAM

Section 3-1

1. Identify the elements on the left with the classification on the right.

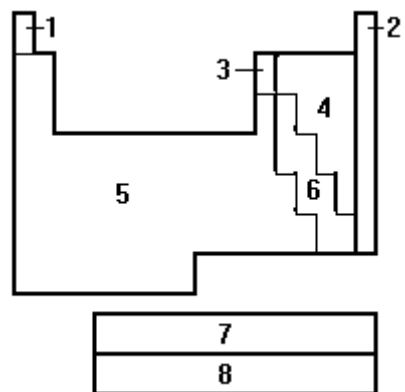
- | | |
|-----------------------|--------------------|
| A. chlorine, bromine | I chalcogens |
| B. diamond, graphite | II alkaline earths |
| C. magnesium, calcium | III. allotropes |
| D. oxygen, sulfur | IV. halogens |

- | | | | |
|-----------|-------|-------|------|
| (a) A I | B II | C III | D IV |
| (b) A IV | B III | C I | D II |
| (c) A III | B I | C IV | D II |
| (d) A IV | B III | C II | D I |

Section 3-2

2. Certain regions are numbered on the outline periodic table at the right. Identify by number or numbers the region or regions in which the following kinds of elements are found.

- Inert gases
- Metals
- Semi-metals
- Non-metals



3. The symbol for the element with an isotope of atomic number 27 and isotopic mass 56 is:

- | | | |
|--------|--------|--------|
| (a) Co | (c) Cu | (e) Al |
| (b) Ba | (d) Bi | |

Section 3-3

4. Write formulas for the following compounds.

- Calcium phosphide
- Aluminum chloride
- Sodium nitride
- Potassium sulfide

5. Name the following:

- | | | | |
|---------------------|--------------------|--------------------|-----------------------------|
| (a) SiBr_4 | (b) PCl_5 | (c) ZrO_2 | (d) Cl_2O_7 |
|---------------------|--------------------|--------------------|-----------------------------|

Section 3-4

6. Write Lewis symbols for the element N and the ion N^{3-} .

(a) symbol for N _____ (b) symbol for N^{3-} _____

7. Determine the formal charge on all the atoms in the following Lewis structures:

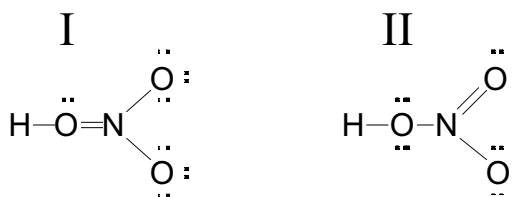


Which one would probably be favored as the structure of the molecule Cl_2O ?

Section 3-5

8. Draw the Lewis dot structure of the sulfuric acid molecule H_2SO_4 .

9. Two possible structures of the nitric acid molecule, HNO_3 , are given on the right. Find the formal charge (FC) on the nitrogen atom and of the oxygen bonded to the hydrogen in each possibility.



	F.C. on N	F.C. on O
Structure I	_____	_____
Structure II	_____	_____

10. Draw Lewis structures for the nitrate ion NO_3^- . Show any resonance structures.

11. All of the following are stable covalent molecules. However, all but one of them are exceptions to the octet rule. Which molecule *obeys* the octet rule?

(a) CO_2 (b) BCl_3 (c) XeF_4 (d) PCl_5 (e) SF_6

12. If one uses the formal charges (FCs) as a basis of choice between the structures of HNO_3 , one selects the possibility that fulfills all of the following criteria except one. Which one is either unnecessary, irrelevant, or wrong?

- (a) The sum of the FCs on all atoms must be zero.
- (b) The most electronegative atoms should have negative rather than positive values of FC.
- (c) All oxygen atoms must have the same formal charge.
- (d) The structure with smaller formal charges should be chosen.

CHAPTER 3: CHEMICAL PERIODICITY

13. Some cities have begun using chlorine dioxide, ClO_2 , to replace chlorine in water treatment. Which of the following statements about ClO_2 is clearly true?
- ClO_2 is an odd-electron molecule.
 - ClO_2 is probably an ionic compound.
 - ClO_2 has a triple bond between the two oxygen atoms.
 - ClO_2 has no lone pair electrons.
14. Which of the following molecules has *more than one* resonance structure?
- A. CH_2OH B. SO_2 C. PCl_5 D. CCl_4 E. NH_3
- A and B only
 - C and D only
 - A and B only
 - B only
 - All of them.

Section 3-6

15. Give either common or systematic names for the following compounds.
- (a) PbSO_4 (b) $\text{K}_2\text{Cr}_2\text{O}_7$ (c) NaHCO_3 (d) FeO
16. Balance the following reactions, for which all reactants and products are given.
- (a) Precipitation of lead nitrate from solution:
 $\underline{\hspace{1cm}} \text{NaCl}(aq) + \underline{\hspace{1cm}} \text{Pb}(\text{NO}_3)_2(aq) \rightarrow \underline{\hspace{1cm}} \text{PbCl}_2(s) + \underline{\hspace{1cm}} \text{NaNO}_3(aq)$
- (b) Neutralization of phosphoric acid:
 $\underline{\hspace{1cm}} \text{Ca}(\text{OH})_2(aq) + \underline{\hspace{1cm}} \text{H}_3\text{PO}_4(aq) \rightarrow \underline{\hspace{1cm}} \text{Ca}_3(\text{PO}_4)_2(aq) + \underline{\hspace{1cm}} \text{H}_2\text{O}(aq)$

Section 3-7

17. Which of the following species should be linear, that is, all of the atoms are centered on a line?
- HCN
 - H_2O
 - SO_2
 - H_2CO
 - None of these is linear.
18. Which of the following would be expected to have tetrahedral atomic geometry?
- NH_3
 - SF_4
 - XeF_4
 - SF_6
 - AlCl_4^{1-}

19. The Cl atoms in PCl_5 form a trigonal bipyramid around the central atom, while the F atoms in BrF_5 are in a square pyramidal arrangement. Which of the following statements best accounts for this difference?
- F is more electronegative than Cl.
 - Cl is larger in diameter than F.
 - There are 5 electron pairs around P in PCl_5 , but 6 pairs around Br in BrF_5 .
 - PCl_5 is ionic while BrF_5 is covalent.
 - PCl_5 obeys the octet rule while BrF_5 does not.
20. In ammonia, NH_3 , the electron pairs around N are in an approximately _____ arrangement, and the atoms are in a(n) _____ arrangement.
- linear
 - trigonal planar
 - trigonal pyramidal
 - tetrahedral
 - octahedral
21. Dipole moments of several molecules are given below. Pick the **false** statement regarding these data.
- | molecule | HF | HCl | HBr | H_2O | N_2 | CO_2 | SO_2 |
|-----------------------------|------|------|------|----------------------|--------------|---------------|---------------|
| dipole moment (debye units) | 1.92 | 1.08 | 0.78 | 1.87 | 0.00 | 0.00 | 1.60 |
- SO_2 is bent, while CO_2 is linear.
 - The bond in HF has more ionic character than the HBr bond.
 - H_2O has a dipole moment because it is non-linear, even though the O-H bond is non-polar.
 - All homonuclear diatomic molecules have zero dipole moment.
22. Which of the following molecules would be expected to have the largest dipole moment?
- NF_3
 - BF_3
 - SF_6
 - CCl_4
 - $\text{CH}_3\text{-CH}_3$
23. In which of these pairs do the molecules or ions differ most in shape and bonding?
- SiCl_4 and CCl_4
 - CN^- and CO
 - BrF_5 and PF_5
 - HCl and HBr
 - SF_6 and FeF_6^{3-}

Section 3-8

CHAPTER 3: CHEMICAL PERIODICITY

24. Which of the following is INCORRECTLY paired?

- (a) CuCl copper(I) chloride
- (b) BiCl₃ bismuth(III) chloride
- (c) FeCl₂ ferrous chloride
- (d) MnO₂ manganese(II) oxide
- (e) FeBr₃ iron(III) bromide

25. Which of the following statements about the compound CaCO₃ is FALSE? If no answer is false, choose (e) as correct.

- (a) CaCO₃ contains both ionic and covalent bonds.
- (b) CaCO₃ contains carbonate ions with charges of (2-).
- (c) The carbonate ion is an example of a resonance structure.
- (d) CaCO₃(s) is an insulator, not a conductor of electricity.
- (e) All of the above (a-d) are true statements.

CHAPTER 3: PRACTICE EXAM ANSWER KEY

1. (d)

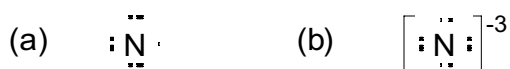
2. (a) 2 (b) 5, 7, 8 (c) 3, 6 (d) 1, 4

3. (a) Co

4. (a) Ca₃P₂ (b) AlCl₃ (c) Na₃N (d) K₂S

5. (a) silicon tetrabromide. (b) vanadium pentachloride. (c) zirconium dioxide.
(d) dichlorine heptaoxide

6.

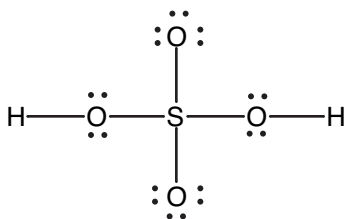


7.

	(a) $\ddot{\text{Cl}}-\ddot{\text{Cl}}-\ddot{\text{O}}\cdot$			(b) $\ddot{\text{Cl}}-\ddot{\text{O}}-\ddot{\text{Cl}}\cdot$		
Group number	(a) Cl	Cl	O	(b) Cl	O	Cl
- number of lone pair electrons	+ 7	+ 7	+ 6	+ 7	+ 6	+ 7
- 1/2 (number of bonding pair electrons)	- 6	- 4	- 6	- 6	- 4	- 6
	- 1	- 2	- 1	- 1	- 2	- 1
	-----			-----		
Formal charge	0	+ 1	- 1	0	0	0

(b) is favored.

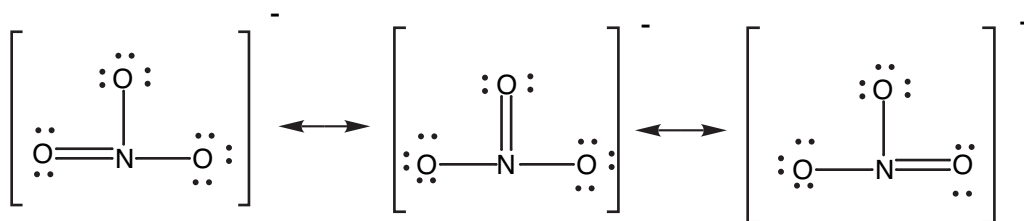
8.



9.

	(I)	N	O	(II)	N	O
Group number		+ 5	+ 6		+ 5	+ 6
- number of lone pair electrons		- 0	- 2		- 0	- 4
- 1/2 (number of bonding pair electrons)		- 4	- 3		- 4	- 2
Formal charge		+ 1	+ 1		+ 1	0

10.



11. (a) CO_2 has two double bonds, but has an octet around C. (b) BCl_3 has only 3 pairs around B, while (c), (d), (e) all have expanded octets around the central atom.

12. (c) The different FC s on different O atoms help choose the most probable structure.

13. (a) Available electrons $A = 6 + 6 + 7 = 19$.

(b) is false, since Cl is a Group VII element, and O is a group VI element.

(c) is false, by Rule 5 on p. 90.

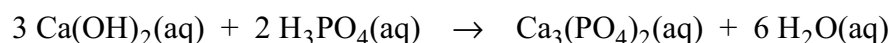
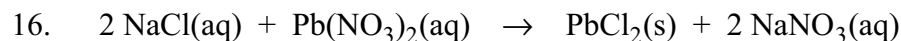
(d) is false; there are many lone pairs.

14. (d)

15. (a) lead(II) sulfate. (b) potassium dichromate.

(c) sodium hydrogen carbonate (common nonsystematic name: sodium bicarbonate)

(d) ferrous oxide or iron(II) oxide



CHAPTER 3: CHEMICAL PERIODICITY

17. (a) $\text{H}-\text{C}\equiv\text{N}$: is linear. The groups of electrons on the central carbon will lie on opposite sides of it. (b) and (c) are bent or angular because the central atom has one (SO_2) or two (H_2O) lone pairs in addition to two bonded atoms. (d) $\text{H}_2\text{C}=\text{O}$ has 3 electron groups (2 single and one double bond) so it is trigonal planar.
18. (e) AlCl_4^- ion is the only one that has both 4 electron pairs and 4 bonded atoms around the central atom. (a-c) have 4 bonded atoms but also the following numbers of lone electron pairs: NH_3 one, SF_4 one, XeF_4 two. (d) SF_6 is octahedral.
19. (c)
20. (d) and (c) There are four electron pairs (tetrahedral) and the atoms form a trigonal pyramid, with the lone, nonbonded pair projecting above the pyramid.
21. (c) The $\text{O}-\text{H}$ bond is polar. H_2O has two lone pairs that force the $\text{H}-\text{O}-\text{H}$ atoms to a bent configuration. (a) is true. The linear molecule $\text{O}=\text{C}=\text{O}$ has zero dipole moment. (b) is true because the more polar HF bond is more ionic than the HBr bond. (d) is true.
22. (a) NF_3 has the same general shape as NH_3 . See item 17. The other molecules are symmetrical so their bond dipoles cancel each other. BF_3 is trigonal planar with only three electron pairs on boron, and so has zero dipole moment.
23. (c) PF_5 and BrF_5 differ in the steric number of the central atom. See question 19 above. (b) CN^- and CO are isoelectronic and have triple bonds. In (a) both have $\text{SN} = 4$ and are tetrahedral. In (e) the SN is 6 and both are octahedral. In (d) both have a single bond and Cl and Br are in the same periodic group.
24. (d) MnO_2 is manganese(IV) oxide. The oxide ion is O^{2-} , so M^{4+} is the cation.
25. (e) Note (a) is true; the $\text{C}-\text{O}$ bonds in carbonate ion CO_3^{2-} are covalent.