18.E: Representative Metals, Metalloids, and Nonmetals (Exercises)

18.E.1: 18.1: Periodicity

How do alkali metals differ from alkaline earth metals in atomic structure and general properties?

The alkali metals all have a single *s* electron in their outermost shell. In contrast, the alkaline earth metals have a completed *s* subshell in their outermost shell. In general, the alkali metals react faster and are more reactive than the corresponding alkaline earth metals in the same period.

Why does the reactivity of the alkali metals decrease from cesium to lithium?

Predict the formulas for the nine compounds that may form when each species in column 1 of Table reacts with each species in column 2.

1	2
Na	I
Sr	Se
Al	O

$$\begin{aligned} \mathrm{Na + I_2 \longrightarrow 2\,NaI} \\ \mathrm{2\,Na + Se \longrightarrow Na_2Se} \\ \mathrm{2\,Na + O_2 \longrightarrow Na_2O_2} \end{aligned} \tag{18.E.1}$$

$$\begin{aligned} \mathrm{Sr + I_2 \longrightarrow SrI_2} \\ \mathrm{Sr + Se \longrightarrow SeSe} \\ \mathrm{2\,Sr + O_2 \longrightarrow 2\,SrO} \end{aligned} \tag{18.E.2}$$

$$\begin{aligned} \mathrm{2\,Al + 3\,I_2 \longrightarrow 2\,AlI_3} \\ \mathrm{2\,Al + 3\,Se \longrightarrow Al_2Se_3} \\ \mathrm{4\,Al + 3\,O_2 \longrightarrow 2\,Al_2O_3} \end{aligned} \tag{18.E.3}$$

Predict the best choice in each of the following. You may wish to review the chapter on electronic structure for relevant examples.

1. (a) the most metallic of the elements Al, Be, and Ba
2. (b) the most covalent of the compounds NaCl, $CaCl_2$, and $BeCl_2$
3. (c) the lowest first ionization energy among the elements Rb, K, and Li
4. (d) the smallest among Al, Al^+, and Al^{3+}
5. (e) the largest among Cs^+, Ba^{2+}, and Xe

Sodium chloride and strontium chloride are both white solids. How could you distinguish one from the other?

The possible ways of distinguishing between the two include infrared spectroscopy by comparison of known compounds, a flame test that gives the characteristic yellow color for sodium (strontium has a red flame), or comparison of their solubilities in water. At 20 °C, NaCl dissolves to the extent of $\frac{35.7\text{ g}}{100\text{ mL}}$ compared with $\frac{53.8\text{ g}}{100\text{ mL}}$ for $SrCl_2$. Heating to 100 °C provides an easy test, since the solubility of NaCl is $\frac{39.12\text{ g}}{100\text{ mL}}$, but that of $SrCl_2$ is $\frac{100.8\text{ g}}{100\text{ mL}}$. Density determination on a solid is sometimes difficult, but there is enough difference (2.165 g/mL NaCl and 3.052 g/mL $SrCl_2$) that this method would be viable and perhaps the easiest and least expensive test to perform.

The reaction of quicklime, CaO, with water produces slaked lime, $Ca(OH)_2$, which is widely used in the construction industry to make mortar and plaster. The reaction of quicklime and water is highly exothermic:

$$\mathrm{CaO}(s) + \mathrm{H_2O}(l) \longrightarrow \mathrm{Ca(OH)_2}(s) \quad \Delta H = -350\text{ kJ mol}^{-1} \tag{18.E.4}$$

1. (a) What is the enthalpy of reaction per gram of quicklime that reacts?
2. (b) How much heat, in kilojoules, is associated with the production of 1 ton of slaked lime?

Write a balanced equation for the reaction of elemental strontium with each of the following:

1. (a) oxygen
2. (b) hydrogen bromide
3. (c) hydrogen
4. (d) phosphorus
5. (e) water

(a) $2\,Sr(s) + O_2(g) \longrightarrow 2\,SrO(s)$; (b) $Sr(s) + 2\,HBr(g) \longrightarrow SrBr_2(s) + H_2(g)$; (c) $Sr(s) + H_2(g) \longrightarrow SrH_2(s)$; (d) $6\,Sr(s) + P_4(s) \longrightarrow 2\,Sr_3P_2(s)$; (e) $Sr(s) + 2\,H_2O(l) \longrightarrow Sr(OH)_2(aq) + H_2(g)$

How many moles of ionic species are present in 1.0 L of a solution marked 1.0 *M* mercury(I) nitrate?

What is the mass of fish, in kilograms, that one would have to consume to obtain a fatal dose of mercury, if the fish contains 30 parts per million of mercury by weight? (Assume that all the mercury from the fish ends up as mercury(II) chloride in the body and that a fatal dose is 0.20 g of $HgCl_2$.) How many pounds of fish is this?

11 lb

The elements sodium, aluminum, and chlorine are in the same period.

1. (a) Which has the greatest electronegativity?
2. (b) Which of the atoms is smallest?
3. (c) Write the Lewis structure for the simplest covalent compound that can form between aluminum and chlorine.
4. (d) Will the oxide of each element be acidic, basic, or amphoteric?

Does metallic tin react with HCl?

Yes, tin reacts with hydrochloric acid to produce hydrogen gas.

What is tin pest, also known as tin disease?

Compare the nature of the bonds in $PbCl_2$ to that of the bonds in $PbCl_4$.

In $PbCl_2$, the bonding is ionic, as indicated by its melting point of 501 °C. In $PbCl_4$, the bonding is covalent, as evidenced by it being an unstable liquid at room temperature.

Is the reaction of rubidium with water more or less vigorous than that of sodium? How does the rate of reaction of magnesium compare?

18.E.2: 18.2: Occurrence and Preparation of the Representative Metals

Write an equation for the reduction of cesium chloride by elemental calcium at high temperature.

$$2\,CsCl(l) + Ca(g) \xrightarrow{\substack{\text{countercurrent}\\\text{fractionating}\\\text{tower}}} 2\,Cs(g) + CaCl_2(l) \tag{18.E.5}$$

Why is it necessary to keep the chlorine and sodium, resulting from the electrolysis of sodium chloride, separate during the production of sodium metal?

Give balanced equations for the overall reaction in the electrolysis of molten lithium chloride and for the reactions occurring at the electrodes. You may wish to review the chapter on electrochemistry for relevant examples.

Cathode (reduction): $2\,Li^+ + 2\,e^- \longrightarrow 2\,Li(l)$; Anode (oxidation): $2\,Cl^- \longrightarrow Cl_2(g) + 2\,e^-$; Overall reaction: $2\,Li^+ + 2\,Cl^- \longrightarrow 2\,Li(l) + Cl_2(g)$

The electrolysis of molten sodium chloride or of aqueous sodium chloride produces chlorine.

Calculate the mass of chlorine produced from 3.00 kg sodium chloride in each case. You may wish to review the chapter on electrochemistry for relevant examples.

What mass, in grams, of hydrogen gas forms during the complete reaction of 10.01 g of calcium with water?

0.5035 g H_2

How many grams of oxygen gas are necessary to react completely with 3.01×10^{21} atoms of magnesium to yield magnesium oxide?

Magnesium is an active metal; it burns in the form of powder, ribbons, and filaments to provide flashes of brilliant light. Why is it possible to use magnesium in construction?

Despite its reactivity, magnesium can be used in construction even when the magnesium is going to come in contact with a flame because a protective oxide coating is formed, preventing gross oxidation. Only if the metal is finely subdivided or present in a thin sheet will a high-intensity flame cause its rapid burning.

Why is it possible for an active metal like aluminum to be useful as a structural metal?

Describe the production of metallic aluminum by electrolytic reduction.

Extract from ore: $\mathrm{AlO(OH)}(s) + \mathrm{NaOH}(aq) + \mathrm{H_2O}(l) \longrightarrow \mathrm{Na[Al(OH)_4]}(aq)$

Recover: $2\,\mathrm{Na[Al(OH)_4]}(s) + \mathrm{H_2SO_4}(aq) \longrightarrow 2\,\mathrm{Al(OH)_3}(s) + \mathrm{Na_2SO_4}(aq) + 2\,\mathrm{H_2O}(l)$

Sinter: $2\,\mathrm{Al(OH)_3}(s) \longrightarrow \mathrm{Al_2O_3}(s) + 3\,\mathrm{H_2O}(g)$

Dissolve in $Na_3AlF_6(l)$ and electrolyze: $\mathrm{Al^{3+}} + 3\,\mathrm{e^-} \longrightarrow \mathrm{Al}(s)$

What is the common ore of tin and how is tin separated from it?

A chemist dissolves a 1.497-g sample of a type of metal (an alloy of Sn, Pb, Sb, and Cu) in nitric acid, and metastannic acid, H_2SnO_3, is precipitated. She heats the precipitate to drive off the water, which leaves 0.4909 g of tin(IV) oxide. What was the percentage of tin in the original sample?

25.83%

Consider the production of 100 kg of sodium metal using a current of 50,000 A, assuming a 100% yield.

(a) How long will it take to produce the 100 kg of sodium metal?

(b) What volume of chlorine at 25 °C and 1.00 atm forms?

What mass of magnesium forms when 100,000 A is passed through a $MgCl_2$ melt for 1.00 h if the yield of magnesium is 85% of the theoretical yield?

39 kg

18.E.3: 18.3: Structure and General Properties of the Metalloids

Give the hybridization of the metalloid and the molecular geometry for each of the following compounds or ions. You may wish to review the chapters on chemical bonding and advanced covalent bonding for relevant examples.

1. (a) GeH_4
2. (b) SbF_3
3. (c) $Te(OH)_6$
4. (d) H_2Te
5. (e) GeF_2
6. (f) $TeCl_4$
7. (g) SiF_6^{2-}
8. (h) $SbCl_5$
9. (i) TeF_6

Write a Lewis structure for each of the following molecules or ions. You may wish to review the chapter on chemical bonding.

1. (a) H_3BPH_3
2. (b) BF_4^-
3. (c) BBr_3
4. (d) $B(CH_3)_3$
5. (e) $B(OH)_3$

(a) H_3BPH_3:

This Lewis structure is composed of a boron atom single bonded to a phosphorus atom. Each of these atoms is single bonded to three hydrogen atoms. ;

(b) BF_4^-:

This Lewis structure is composed of a boron atom single bonded to four fluorine atoms, each of which has three lone pairs of electrons. The structure is surrounded by brackets, and a negative sign appears as a superscript outside the brackets.

;

(c) BBr_3:

This Lewis structure is composed of a boron atom single bonded to three bromine atoms, each of which has three lone pairs of electrons. ;

(d) $B(CH_3)_3$:

This Lewis structure is composed of a boron atom that is single bonded to three carbon atoms, each of which is single bonded to three hydrogen atoms. ;

(e) $B(OH)_3$:

This Lewis structure is composed of a boron atom that is single bonded to three oxygen atoms, each of which has two lone pairs of electrons. Each oxygen atom is single bonded to a hydrogen atom.

Describe the hybridization of boron and the molecular structure about the boron in each of the following:

1. (a) H_3BPH_3
2. (b) BF_4^-
3. (c) BBr_3
4. (d) $B(CH_3)_3$
5. (e) $B(OH)_3$

Using only the periodic table, write the complete electron configuration for silicon, including any empty orbitals in the valence shell. You may wish to review the chapter on electronic structure.

$1s^2 2s^2 2p^6 3s^2 3p^2 3d^0$.

Write a Lewis structure for each of the following molecules and ions:

1. (a) $(CH_3)_3SiH$
2. (b) SiO_4^{4-}
3. (c) Si_2H_6
4. (d) $Si(OH)_4$
5. (e) SiF_6^{2-}

Describe the hybridization of silicon and the molecular structure of the following molecules and ions:

1. (a) $(CH_3)_3SiH$
2. (b) SiO_4^{4-}
3. (c) Si_2H_6
4. (d) $Si(OH)_4$
5. (e) SiF_6^{2-}

(a) $(CH_3)_3SiH$: sp^3 bonding about Si; the structure is tetrahedral; (b) SiO_4^{4-}: sp^3 bonding about Si; the structure is tetrahedral; (c) Si_2H_6: sp^3 bonding about each Si; the structure is linear along the Si-Si bond; (d) $Si(OH)_4$: sp^3 bonding about Si; the structure is tetrahedral; (e) SiF_6^{2-}: sp^3d^2 bonding about Si; the structure is octahedral

Describe the hybridization and the bonding of a silicon atom in elemental silicon.

Classify each of the following molecules as polar or nonpolar. You may wish to review the chapter on chemical bonding.

(a) SiH_4

(b) Si_2H_6

(c) $SiCl_3H$

(d) SiF_4

(e) $SiCl_2F_2$

(a) nonpolar; (b) nonpolar; (c) polar; (d) nonpolar; (e) polar

Silicon reacts with sulfur at elevated temperatures. If 0.0923 g of silicon reacts with sulfur to give 0.3030 g of silicon sulfide, determine the empirical formula of silicon sulfide.

Name each of the following compounds:

1. (a) TeO_2
2. (b) Sb_2S_3
3. (c) GeF_4
4. (d) SiH_4
5. (e) GeH_4

(a) tellurium dioxide or tellurium(IV) oxide; (b) antimony(III) sulfide; (c) germanium(IV) fluoride; (d) silane or silicon(IV) hydride; (e) germanium(IV) hydride

Write a balanced equation for the reaction of elemental boron with each of the following (most of these reactions require high temperature):

1. (a) F_2
2. (b) O_2
3. (c) S
4. (d) Se
5. (e) Br_2

Why is boron limited to a maximum coordination number of four in its compounds?

Boron has only *s* and *p* orbitals available, which can accommodate a maximum of four electron pairs. Unlike silicon, no *d* orbitals are available in boron.

Write a formula for each of the following compounds:

1. (a) silicon dioxide
2. (b) silicon tetraiodide
3. (c) silane
4. (d) silicon carbide
5. (e) magnesium silicide

From the data given in Appendix I , determine the standard enthalpy change and the standard free energy change for each of the following reactions:

1. (a) $\mathrm{BF_3}(g) + 3\,\mathrm{H_2O}(l) \longrightarrow \mathrm{B(OH)_3}(s) + 3\,\mathrm{HF}(g)$
2. (b) $\mathrm{BCl_3}(g) + 3\,\mathrm{H_2O}(l) \longrightarrow \mathrm{B(OH)_3}(s) + 3\,\mathrm{HCl}(g)$
3. (c) $\mathrm{B_2H_6}(g) + 6\,\mathrm{H_2O}(l) \longrightarrow 2\,\mathrm{B(OH)_3}(s) + 6\,\mathrm{H_2}(g)$

(a) $\Delta H°$ = 87 kJ; $\Delta G°$ = 44 kJ; (b) $\Delta H°$ = −109.9 kJ; $\Delta G°$ = −154.7 kJ; (c) $\Delta H°$ = −510 kJ; $\Delta G°$ = −601.5 kJ

A hydride of silicon prepared by the reaction of Mg_2Si with acid exerted a pressure of 306 torr at 26 °C in a bulb with a volume of 57.0 mL. If the mass of the hydride was 0.0861 g, what is its molecular mass? What is the molecular formula for the hydride?

Suppose you discovered a diamond completely encased in a silicate rock. How would you chemically free the diamond without harming it?

A mild solution of hydrofluoric acid would dissolve the silicate and would not harm the diamond.

18.E.4: 18.4: Structure and General Properties of the Nonmetals

Carbon forms a number of allotropes, two of which are graphite and diamond. Silicon has a diamond structure. Why is there no allotrope of silicon with a graphite structure?

Nitrogen in the atmosphere exists as very stable diatomic molecules. Why does phosphorus form less stable P_4 molecules instead of P_2 molecules?

In the N_2 molecule, the nitrogen atoms have an σ bond and two π bonds holding the two atoms together. The presence of three strong bonds makes N_2 a very stable molecule. Phosphorus is a third-period element, and as such, does not form π bonds efficiently; therefore, it must fulfill its bonding requirement by forming three σ bonds.

Write balanced chemical equations for the reaction of the following acid anhydrides with water:

1. (a) SO_3
2. (b) N_2O_3
3. (c) Cl_2O_7
4. (d) P_4O_{10}
5. (e) NO_2

Determine the oxidation number of each element in each of the following compounds:

1. (a) HCN
2. (b) OF_2
3. (c) $AsCl_3$

(a) H = 1+, C = 2+, and N = 3−; (b) O = 2+ and F = 1−; (c) As = 3+ and Cl = 1−

Determine the oxidation state of sulfur in each of the following:

1. (a) SO_3
2. (b) SO_2
3. (c) SO_3^{2-}

Arrange the following in order of increasing electronegativity: F; Cl; O; and S.

S < Cl < O < F

Why does white phosphorus consist of tetrahedral P_4 molecules while nitrogen consists of diatomic N_2 molecules?

18.E.5: 18.5: Occurrence, Preparation, and Compounds of Hydrogen

Why does hydrogen not exhibit an oxidation state of 1− when bonded to nonmetals?

The electronegativity of the nonmetals is greater than that of hydrogen. Thus, the negative charge is better represented on the nonmetal, which has the greater tendency to attract electrons in the bond to itself.

The reaction of calcium hydride, CaH_2, with water can be characterized as a Lewis acid-base reaction:

$$\mathrm{CaH_2}(s) + 2\,\mathrm{H_2O}(l) \longrightarrow \mathrm{Ca(OH)_2}(aq) + 2\,\mathrm{H_2}(g) \tag{18.E.6}$$

Identify the Lewis acid and the Lewis base among the reactants. The reaction is also an oxidation-reduction reaction. Identify the oxidizing agent, the reducing agent, and the changes in oxidation number that occur in the reaction.

In drawing Lewis structures, we learn that a hydrogen atom forms only one bond in a covalent compound. Why?

Hydrogen has only one orbital with which to bond to other atoms. Consequently, only one two-electron bond can form.

What mass of CaH_2 is necessary to react with water to provide enough hydrogen gas to fill a balloon at 20 °C and 0.8 atm pressure with a volume of 4.5 L? The balanced equation is:

$$\mathrm{CaH_2}(s) + 2\,\mathrm{H_2O}(l) \longrightarrow \mathrm{Ca(OH)_2}(aq) + 2\,\mathrm{H_2}(g) \tag{18.E.7}$$

What mass of hydrogen gas results from the reaction of 8.5 g of KH with water?

$$\mathrm{KH} + \mathrm{H_2O} \longrightarrow \mathrm{KOH} + \mathrm{H_2} \tag{18.E.8}$$

0.43 g H_2

18.E.6: 18.6: Occurrence, Preparation, and Properties of Carbonates

Carbon forms the $\mathrm{CO_3^{2-}}$ ion, yet silicon does not form an analogous $\mathrm{SiO_3^{2-}}$ ion. Why?

Complete and balance the following chemical equations:

(a) hardening of plaster containing slaked lime

$$\mathrm{Ca(OH)_2} + \mathrm{CO_2} \longrightarrow \tag{18.E.9}$$

(b) removal of sulfur dioxide from the flue gas of power plants

$$CaO + SO_2 \longrightarrow \tag{18.E.10}$$

(c) the reaction of baking powder that produces carbon dioxide gas and causes bread to rise

$$NaHCO_3 + NaH_2PO_4 \longrightarrow \tag{18.E.11}$$

(a) $Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$; (b) $CaO(s) + SO_2(g) \longrightarrow CaSO_3(s)$;

(c) $2\,NaHCO_3(s) + NaH_2PO_4(aq) \longrightarrow Na_3PO_4(aq) + 2\,CO_2(g) + 2\,H_2O(l)$

Heating a sample of $Na_2CO_3 \cdot xH_2O$ weighing 4.640 g until the removal of the water of hydration leaves 1.720 g of anhydrous Na_2CO_3. What is the formula of the hydrated compound?

18.E.7: 18.7: Occurrence, Preparation, and Properties of Nitrogen

Write the Lewis structures for each of the following:

1. (a) NH^{2-}
2. (b) N_2F_4
3. (c) NH_2^-
4. (d) NF_3
5. (e) N_3^-

(a) NH^{2-}:

This Lewis structure shows a nitrogen atom with three lone pairs of electrons single bonded to a hydrogen atom. The structure is surrounded by brackets. Outside and superscript to the brackets is a two negative sign.
; (b) N_2F_4:

This Lewis structure shows two nitrogen atoms, each with one lone pair of electrons, single bonded to one another and each single bonded to two fluorine atoms. Each fluorine atom has three lone pairs of electrons. ;
(c) NH_2^- :

This Lewis structure shows a nitrogen atom with two lone pairs of electrons single bonded to two hydrogen atoms. The structure is surrounded by brackets. Outside and superscript to the brackets is a negative sign. ;
(d) NF_3: This Lewis structure shows a nitrogen atom, with one lone pair of electrons, single bonded to three fluorine atoms. Each fluorine atom has three lone pairs of electrons. ; (e) N_3^- :
Three Lewis structures are shown and connected by double-headed arrows in between. The left structure shows a nitrogen atom with a lone pair of electrons triple bonded to a second nitrogen which is single bonded to a third nitrogen. The third nitrogen has three lone pairs of electrons. The entire structure is surrounded by brackets, and outside and superscript to the brackets is a negative sign. The middle structure shows a nitrogen atom with three lone pair of electrons single bonded to a second nitrogen which is triple bonded to a third nitrogen. The third nitrogen which has one lone pair of electrons. The entire structure is surrounded by brackets, and outside and superscript to the brackets is a negative sign. The right structure shows a nitrogen atom with two lone pairs of electrons double bonded to a second nitrogen which is double bonded to a third nitrogen. The third nitrogen atom has two lone pairs of electrons. The entire structure is surrounded by brackets, and outside and superscript to the brackets is a negative sign.

For each of the following, indicate the hybridization of the nitrogen atom (for N_3^-, the central nitrogen).

1. (a) N_2F_4
2. (b) NH_2^-
3. (c) NF_3
4. (d) N_3^-

Explain how ammonia can function both as a Brønsted base and as a Lewis base.

Ammonia acts as a Brønsted base because it readily accepts protons and as a Lewis base in that it has an electron pair to donate.

Brønsted base: $NH_3 + H_3O^+ \longrightarrow NH_4^+ + H_2O$ Lewis base: $2\,NH_3 + Ag^+ \longrightarrow [H_3N - Ag - NH_3]^+$

Determine the oxidation state of nitrogen in each of the following. You may wish to review the chapter on chemical bonding for relevant examples.

1. (a) NCl_3
2. (b) ClNO
3. (c) N_2O_5
4. (d) N_2O_3
5. (e) NO_2^-
6. (f) N_2O_4
7. (g) N_2O
8. (h) NO_3^-
9. (i) HNO_2
10. (j) HNO_3

For each of the following, draw the Lewis structure, predict the ONO bond angle, and give the hybridization of the nitrogen. You may wish to review the chapters on chemical bonding and advanced theories of covalent bonding for relevant examples.

(a) NO_2

(b) NO_2^-

(c) NO_2^+

(a) NO_2:

Two Lewis structures are shown and connected by double-headed arrows in between. The left structure shows a nitrogen atom with a single electron double bonded to an oxygen atom which has two lone pairs of electrons. The nitrogen atom is also single bonded to an oxygen atom with three lone pairs of electrons. The right structure is a mirror image of the left structure.

Nitrogen is sp^2 hybridized. The molecule has a bent geometry with an ONO bond angle of approximately 120°. (b) NO_2^- :

Two Lewis structures are shown and connected by double-headed arrows in between. Each structure is surrounded by brackets, and outside and superscript to the brackets is a negative sign. The left structure shows a nitrogen atom with a lone pair of electrons double bonded to an oxygen atom which has two lone pairs of electrons. The nitrogen atom is also single bonded to an oxygen atom with three lone pair of electrons. The right structure is a mirror image of the left structure.

Nitrogen is sp^2 hybridized. The molecule has a bent geometry with an ONO bond angle slightly less than 120°. (c) NO_2^+ :

This Lewis structure shows a nitrogen atom double bonded on both sides to an oxygen atom which has two lone pairs of electrons each. The structure is surrounded by brackets and outside and superscript to the brackets is a negative sign.

Nitrogen is *sp* hybridized. The molecule has a linear geometry with an ONO bond angle of 180°.

How many grams of gaseous ammonia will the reaction of 3.0 g hydrogen gas and 3.0 g of nitrogen gas produce?

Although PF_5 and AsF_5 are stable, nitrogen does not form NF_5 molecules. Explain this difference among members of the same group.

Nitrogen cannot form a NF_5 molecule because it does not have *d* orbitals to bond with the additional two fluorine atoms.

The equivalence point for the titration of a 25.00-mL sample of CsOH solution with 0.1062 *M* HNO_3 is at 35.27 mL. What is the concentration of the CsOH solution?

18.E.8: 18.8: Occurrence, Preparation, and Properties of Phosphorus

Write the Lewis structure for each of the following. You may wish to review the chapter on chemical bonding and molecular geometry.

1. (a) PH_3
2. (b) PH_4^+
3. (c) P_2H_4
4. (d) PO_4^{3-}
5. (e) PF_5

(a)

This Lewis structure shows a phosphorus atom with a lone pair of electrons single bonded to three hydrogen atoms. ;

(b)

This Lewis structure shows a phosphorus atom single bonded to four hydrogen atoms. The structure is surrounded by brackets and has a superscript positive sign outside the brackets. ;

(c)

This Lewis structure shows two phosphorus atoms, each with a lone pair of electrons, single bonded to one another. Each phosphorus atom is also single bonded to two hydrogen atoms. ;

(d)

This Lewis structure shows a phosphorus atom single bonded to four oxygen atoms, each with three lone pairs of electrons. The structure is surrounded by brackets and has a superscript 3 negative sign outside the brackets.

;

(e)

This Lewis structure shows a phosphorus atom single bonded to five fluorine atoms, each with three lone pairs of electrons.

Describe the molecular structure of each of the following molecules or ions listed. You may wish to review the chapter on chemical bonding and molecular geometry.

1. (a) PH_3
2. (b) PH_4^+
3. (c) P_2H_4
4. (d) PO_4^{3-}

Complete and balance each of the following chemical equations. (In some cases, there may be more than one correct answer.)

1. (a) $P_4 + Al \longrightarrow$
2. (b) $P_4 + Na \longrightarrow$
3. (c) $P_4 + F_2 \longrightarrow$
4. (d) $P_4 + Cl_2 \longrightarrow$
5. (e) $P_4 + O_2 \longrightarrow$
6. (f) $P_4O_6 + O_2 \longrightarrow$

(a) $P_4(s) + 4\,Al(s) \longrightarrow 4\,AlP(s)$; (b) $P_4(s) + 12\,Na(s) \longrightarrow 4\,Na_3P(s)$; (c) $P_4(s) + 10\,F_2(g) \longrightarrow 4\,PF_5(l)$; (d) $P_4(s) + 6\,Cl_2(g) \longrightarrow 4\,PCl_3(l)$ or $P_4(s) + 10\,Cl_2(g) \longrightarrow 4\,PCl_5(l)$; (e) $P_4(s) + 3\,O_2(g) \longrightarrow P_4O_6(s)$ or $P_4(s) + 5\,O_2(g) \longrightarrow P_4O_{10}(s)$; (f) $P_4O_6(s) + 2\,O_2(g) \longrightarrow P_4O_{10}(s)$

Describe the hybridization of phosphorus in each of the following compounds: P_4O_{10}, P_4O_6, PH_4I (an ionic compound), PBr_3, H_3PO_4, H_3PO_3, PH_3, and P_2H_4. You may wish to review the chapter on advanced theories of covalent bonding.

What volume of 0.200 *M* NaOH is necessary to neutralize the solution produced by dissolving 2.00 g of PCl_3 is an excess of water? Note that when H_3PO_3 is titrated under these conditions, only one proton of the acid molecule reacts.

291 mL

How much $POCl_3$ can form from 25.0 g of PCl_5 and the appropriate amount of H_2O?

How many tons of $Ca_3(PO_4)_2$ are necessary to prepare 5.0 tons of phosphorus if the yield is 90%?

28 tons

Write equations showing the stepwise ionization of phosphorous acid.

Draw the Lewis structures and describe the geometry for the following:

1. (a) PF_4^+
2. (b) PF_5
3. (c) PF_6^-
4. (d) POF_3

(a)

This Lewis structure shows a phosphorus atom single bonded to four fluorine atoms, each with three lone pairs of electrons. The structure is surrounded by brackets and has a superscript positive sign outside the brackets. The label, “Tetrahedral,” is written under the structure.

;

(b)

This Lewis structure shows a phosphorus atom single bonded to five fluorine atoms, each with three lone pairs of electrons. The label, “Trigonal bipyramidal,” is written under the structure. ;

(c)

A Lewis structure shows a phosphorus atom single bonded to six fluorine atoms, each with three lone pairs of electrons. The structure is surrounded by brackets and has a superscript negative sign outside the brackets. The label, “Octahedral,” is written under the structure.

;

(d)

This Lewis structure shows a phosphorus atom single bonded to three fluorine atoms, each with three lone pairs of electrons. The phosphorus atom is also double bonded to an oxygen atom with two lone pairs of electrons. The label, “Tetrahedral,” is written under the structure.

Why does phosphorous acid form only two series of salts, even though the molecule contains three hydrogen atoms?

Assign an oxidation state to phosphorus in each of the following:

1. (a) NaH_2PO_3
2. (b) PF_5
3. (c) P_4O_6
4. (d) K_3PO_4
5. (e) Na_3P
6. (f) $Na_4P_2O_7$

(a) P = 3+; (b) P = 5+; (c) P = 3+; (d) P = 5+; (e) P = 3−; (f) P = 5+

Phosphoric acid, one of the acids used in some cola drinks, is produced by the reaction of phosphorus(V) oxide, an acidic oxide, with water. Phosphorus(V) oxide is prepared by the combustion of phosphorus.

1. (a) Write the empirical formula of phosphorus(V) oxide.
2. (b) What is the molecular formula of phosphorus(V) oxide if the molar mass is about 280.
3. (c) Write balanced equations for the production of phosphorus(V) oxide and phosphoric acid.
4. (d) Determine the mass of phosphorus required to make 1.00×10^4 kg of phosphoric acid, assuming a yield of 98.85%.

18.E.9: 18.9: Occurrence, Preparation, and Compounds of Oxygen

Predict the product of burning francium in air.

FrO_2

Using equations, describe the reaction of water with potassium and with potassium oxide.

Write balanced chemical equations for the following reactions:

1. (a) zinc metal heated in a stream of oxygen gas
2. (b) zinc carbonate heated until loss of mass stops
3. (c) zinc carbonate added to a solution of acetic acid, CH_3CO_2H
4. (d) zinc added to a solution of hydrobromic acid

(a) $2\,Zn(s) + O_2(g) \longrightarrow 2\,ZnO(s)$; (b) $ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$; (c) $ZnCO_3(s) + 2\,CH_3COOH(aq) \longrightarrow Zn(CH_3COO)_2(aq) + CO_2(g) + H_2O(l)$; (d) $Zn(s) + 2\,HBr(aq) \longrightarrow ZnBr_2(aq) + H_2(g)$

Write balanced chemical equations for the following reactions:

1. (a) cadmium burned in air
2. (b) elemental cadmium added to a solution of hydrochloric acid
3. (c) cadmium hydroxide added to a solution of acetic acid, CH_3CO_2H

Illustrate the amphoteric nature of aluminum hydroxide by citing suitable equations.

$Al(OH)_3(s) + 3\,H^+(aq) \longrightarrow Al^{3+} + 3\,H_2O(l)$; $Al(OH)_3(s) + OH^- \longrightarrow [Al(OH)_4]^-(aq)$

Write balanced chemical equations for the following reactions:

1. (a) metallic aluminum burned in air
2. (b) elemental aluminum heated in an atmosphere of chlorine
3. (c) aluminum heated in hydrogen bromide gas
4. (d) aluminum hydroxide added to a solution of nitric acid

Write balanced chemical equations for the following reactions:

1. (a) sodium oxide added to water
2. (b) cesium carbonate added to an excess of an aqueous solution of HF
3. (c) aluminum oxide added to an aqueous solution of $HClO_4$
4. (d) a solution of sodium carbonate added to solution of barium nitrate
5. (e) titanium metal produced from the reaction of titanium tetrachloride with elemental sodium

(a) $Na_2O(s) + H_2O(l) \longrightarrow 2\,NaOH(aq)$; (b) $Cs_2CO_3(s) + 2\,HF(aq) \longrightarrow 2\,CsF(aq) + CO_2(g) + H_2O(l)$; (c) $Al_2O_3(s) + 6\,HClO_4(aq) \longrightarrow 2\,Al(ClO_4)_3(aq) + 3\,H_2O(l)$; (d) $Na_2CO_3(aq) + Ba(NO_3)_2(aq) \longrightarrow 2\,NaNO_3(aq) + BaCO_3(s)$; (e) $TiCl_4(l) + 4\,Na(s) \longrightarrow Ti(s) + 4\,NaCl(s)$

What volume of 0.250 M H_2SO_4 solution is required to neutralize a solution that contains 5.00 g of $CaCO_3$?

Which is the stronger acid, $HClO_4$ or $HBrO_4$? Why?

$HClO_4$ is the stronger acid because, in a series of oxyacids with similar formulas, the higher the electronegativity of the central atom, the stronger is the attraction of the central atom for the electrons of the oxygen(s). The stronger attraction of the oxygen electron results in a stronger attraction of oxygen for the electrons in the O-H bond, making the hydrogen more easily released. The weaker this bond, the stronger the acid.

Write a balanced chemical equation for the reaction of an excess of oxygen with each of the following. Remember that oxygen is a strong oxidizing agent and tends to oxidize an element to its maximum oxidation state.

1. (a) Mg
2. (b) Rb
3. (c) Ga
4. (d) C_2H_2
5. (e) CO

Which is the stronger acid, H_2SO_4 or H_2SeO_4? Why? You may wish to review the chapter on acid-base equilibria.

As H_2SO_4 and H_2SeO_4 are both oxyacids and their central atoms both have the same oxidation number, the acid strength depends on the relative electronegativity of the central atom. As sulfur is more electronegative than selenium, H_2SO_4 is the stronger acid.

18.E.10: 18.10: Occurrence, Preparation, and Properties of Sulfur

Explain why hydrogen sulfide is a gas at room temperature, whereas water, which has a lower molecular mass, is a liquid.

Give the hybridization and oxidation state for sulfur in SO_2, in SO_3, and in H_2SO_4.

SO_2, sp^2 4+; SO_3, sp^2, 6+; H_2SO_4, sp^3, 6+

Which is the stronger acid, $NaHSO_3$ or $NaHSO_4$?

Determine the oxidation state of sulfur in SF_6, SO_2F_2, and KHS.

SF_6: S = 6+; SO_2F_2: S = 6+; KHS: S = 2−

Which is a stronger acid, sulfurous acid or sulfuric acid? Why?

Oxygen forms double bonds in O_2, but sulfur forms single bonds in S_8. Why?

Sulfur is able to form double bonds only at high temperatures (substantially endothermic conditions), which is not the case for oxygen.

Give the Lewis structure of each of the following:

1. (a) SF_4
2. (b) K_2SO_4
3. (c) SO_2Cl_2
4. (d) H_2SO_3
5. (e) SO_3

Write two balanced chemical equations in which sulfuric acid acts as an oxidizing agent.

There are many possible answers including:

$$\mathrm{Cu}(s)+2\,\mathrm{H_2SO_4}(l)\longrightarrow \mathrm{CuSO_4}(aq)+\mathrm{SO_2}(g)+2\,\mathrm{H_2O}(l) \tag{18.E.12}$$

$$\mathrm{C}(s)+2\,\mathrm{H_2SO_4}(l)\longrightarrow \mathrm{CO_2}(g)+2\,\mathrm{SO_2}(g)+2\,\mathrm{H_2O}(l) \tag{18.E.13}$$

Explain why sulfuric acid, H_2SO_4, which is a covalent molecule, dissolves in water and produces a solution that contains ions.

How many grams of Epsom salts ($MgSO_4 \cdot 7H_2O$) will form from 5.0 kg of magnesium?

5.1×10^4 g

18.E.11: 18.11: Occurrence, Preparation, and Properties of Halogens

What does it mean to say that mercury(II) halides are weak electrolytes?

Why is SnCl4 not classified as a salt?

SnCl4 is not a salt because it is covalently bonded. A salt must have ionic bonds.

The following reactions are all similar to those of the industrial chemicals. Complete and balance the equations for these reactions:

(a) reaction of a weak base and a strong acid

$$NH_3 + HClO_4 \longrightarrow \tag{18.E.14}$$

(b) preparation of a soluble silver salt for silver plating

$$Ag_2CO_3 + HNO_3 \longrightarrow \tag{18.E.15}$$

(c) preparation of strontium hydroxide by electrolysis of a solution of strontium chloride

$$SrCl_2(aq) + H_2O(l) \xrightarrow{\text{electrolysis}} \tag{18.E.16}$$

Which is the stronger acid, HClO3 or HBrO3? Why?

In oxyacids with similar formulas, the acid strength increases as the electronegativity of the central atom increases. HClO3 is stronger than HBrO3; Cl is more electronegative than Br.

What is the hybridization of iodine in IF3 and IF5?

Predict the molecular geometries and draw Lewis structures for each of the following. You may wish to review the chapter on chemical bonding and molecular geometry.

(a) IF_5

(b) I_3^-

(c) PCl_5

(d) SeF_4

(e) ClF_3

(a)

This Lewis structure shows an iodine atom with one lone pair of electrons single bonded to five fluorine atoms, each of which has three lone pairs of electrons. The image is labeled, “Square pyramidal.” ;

(b)

This Lewis structure shows an iodine atom with three lone pairs of electrons single bonded to two iodine atoms, each of which has three lone pairs of electrons. The image is surrounded by brackets. A superscript negative sign appears outside the brackets. The image is labeled, “Linear.”

;

(c)

This Lewis structure shows a phosphorus atom single bonded to five chlorine atoms, each of which has three lone pairs of electrons. The image is labeled, “Trigonal bipyramidal.” ;

(d)

This Lewis structure shows a selenium atom with one lone pair of electrons single bonded to four fluorine atoms, each of which has three lone pairs of electrons. The image is labeled “Seesaw.” ;

(e)

This Lewis structure shows a chlorine atom with two lone pairs of electrons single bonded to three fluorine atoms, each of which has three lone pairs of electrons. The image is labeled, “T-shaped.”

Which halogen has the highest ionization energy? Is this what you would predict based on what you have learned about periodic properties?

Name each of the following compounds:

(a) BrF_3

(b) $NaBrO_3$

(c) PBr_5

(d) $NaClO_4$

(e) KClO

(a) bromine trifluoride; (b) sodium bromate; (c) phosphorus pentabromide; (d) sodium perchlorate; (e) potassium hypochlorite

Explain why, at room temperature, fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid.

What is the oxidation state of the halogen in each of the following?

(a) H_5IO_6

(b) IO_4^-

(c) ClO_2

(d) ICl_3

(e) F_2

(a) I: 7+; (b) I: 7+; (c) Cl: 4+; (d) I: 3+; Cl: 1−; (e) F: 0

Physiological saline concentration—that is, the sodium chloride concentration in our bodies—is approximately 0.16 *M*. A saline solution for contact lenses is prepared to match the physiological concentration. If you purchase 25 mL of contact lens saline solution, how many grams of sodium chloride have you bought?

18.E.12: 18.12: Occurrence, Preparation, and Properties of the Noble Gases

Give the hybridization of xenon in each of the following. You may wish to review the chapter on the advanced theories of covalent bonding.

1. (a) XeF_2
2. (b) XeF_4
3. (c) XeO_3
4. (d) XeO_4
5. (e) $XeOF_4$

(a) sp^3d hybridized; (b) sp^3d^2 hybridized; (c) sp^3 hybridized; (d) sp^3 hybridized; (e) sp^3d^2 hybridized;

What is the molecular structure of each of the following molecules? You may wish to review the chapter on chemical bonding and molecular geometry.

1. (a) XeF_2
2. (b) XeF_4
3. (c) XeO_3
4. (d) XeO_4
5. (e) $XeOF_4$

Indicate whether each of the following molecules is polar or nonpolar. You may wish to review the chapter on chemical bonding and molecular geometry.

1. (a) XeF_2
2. (b) XeF_4
3. (c) XeO_3
4. (d) XeO_4
5. (e) $XeOF_4$

(a) nonpolar; (b) nonpolar; (c) polar; (d) nonpolar; (e) polar

What is the oxidation state of the noble gas in each of the following? You may wish to review the chapter on chemical bonding and molecular geometry.

1. (a) XeO_2F_2
2. (b) KrF_2
3. (c) XeF_3^+
4. (d) XeO_6^{4-}
5. (e) XeO_3

A mixture of xenon and fluorine was heated. A sample of the white solid that formed reacted with hydrogen to yield 81 mL of xenon (at STP) and hydrogen fluoride, which was collected in water, giving a solution of hydrofluoric acid. The hydrofluoric acid solution was titrated, and 68.43 mL of 0.3172 *M* sodium hydroxide was required to reach the equivalence point. Determine the empirical formula for the white solid and write balanced chemical equations for the reactions involving xenon.

The empirical formula is XeF_6, and the balanced reactions are:

$$Xe(g) + 3\,F_2(g) \xrightarrow{\Delta} XeF_6(s) \tag{18.E.17}$$

$$XeF_6(s) + 3\,H_2(g) \longrightarrow 6\,HF(g) + Xe(g) \tag{18.E.18}$$

Basic solutions of Na_4XeO_6 are powerful oxidants. What mass of $Mn(NO_3)_2•6H_2O$ reacts with 125.0 mL of a 0.1717 *M* basic solution of Na_4XeO_6 that contains an excess of sodium hydroxide if the products include Xe and solution of sodium permanganate?