

Chapter 7. Periodic Properties of the Elements

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Other Resources**Further Readings:***Chemical and Engineering News*, **September 8, 2003**

Using the Learning Cycle to Introduce Periodicity

The Nuts and Bolts of Chemistry

Mendeleev and Moseley: The Principal

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Mendeleev's Other Predictions

Atomic Numbers Before Moseley

D. I. Mendeleev and the English Chemists

The Evolution of the Periodic System

Periodic Table of Elemental Abundance

A Different Approach to a 3-D Periodic System

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 Aqueous Hydrogen Peroxide: Its Household Uses
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Halogens Compete for Electrons
 Acidic and Basic Properties of Oxides
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 Simple Flame Test Techniques Using Cotton Swabs
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Chapter 7. Periodic Properties of the Elements

Common Student Misconceptions

- Students have difficulty with the concepts of shielding and effective nuclear charge. As you move to the right in a period, shielding does not increase appreciably but the nuclear charge does. Therefore, effective nuclear charge increases steadily as you move to the right along the period.
- Students are confused why, within a period, atomic radii decrease with increasing atomic number.
- Students often do not understand slight irregularities in periodic trends for elements in each row after each *ns* subshell becomes filled, and after *np* and $(n - 1)d$ subshells become half-filled.
- Students often have problems with the signs of electron affinities; in particular, why group 1A metals have negative (exothermic) electronegativities.
- Students are often confused regarding the placement of hydrogen on the periodic table; despite its common placement in column 1A, hydrogen is a nonmetal.
- Students often confuse behavior of elements in *aqueous* phase with periodic properties determined in gas phase (*ionization energy*, *electron affinity*) or in solid phase (*ionic radius*).
- Students often confuse *isoelectronic* species with those with the same number of *valence electrons*.

Teaching Tips

- Students need to be shown how position on the periodic table and electron configurations can be used to highlight periodic properties.
- Emphasize the periodic table as an organizational tool; it will help students recall chemical facts.
- Students find the descriptive chemistry/group trends a bit overwhelming at first.
- Live demonstrations, CD videos, and web-based animations are very helpful in stimulating student interest in the group trends.

Lecture Outline

7.1 Development of the Periodic Table^{1,2,3,4,5,6,7,8,9,10,11,12}

- The periodic table is the most significant tool that chemists use for organizing and recalling chemical facts.
- Elements in the same column contain the same number of outer-shell electrons or **valence electrons**.
- How do we organize the different elements in a meaningful way that will allow us to make predictions about undiscovered elements?

¹ “Periodic Properties” Animation from Instructor’s Resource CD/DVD

² “Periodic Table” Activity from Instructor’s Resource CD/DVD

³ **September 8, 2003** issue of *Chemical and Engineering News* from Further Readings

⁴ “Using the Learning Cycle to Introduce Periodicity” from Further Readings

⁵ “The Nuts and Bolts of Chemistry” from Further Readings

⁶ “Mendeleev and Moseley: The Principal Discoverers of the Periodic Law” from Further Readings

⁷ “The Evolution of the Periodic System” from Further Readings

⁸ “Mendeleev’s Other Prediction” from Further Readings

⁹ “Atomic Number Before Moseley” from Further Readings

¹⁰ “D. I. Mendeleev and the English Chemists” from Further Readings

¹¹ “Periodic Tables of Elemental Abundance” from Further Readings

¹² “A Different Approach to a 3-D Periodic System Including Stable Isotopes” from Further Readings

- Arrange elements to reflect the trends in chemical and physical properties.
- The periodic table arises from the periodic patterns in the electronic configurations of the elements.
 - Elements in the same column contain the same number of valence electrons.
 - The trends within a row or column form patterns that help us make predictions about chemical properties and reactivity.
- In the first attempt Mendeleev and Meyer arranged the elements in order of increasing atomic weight.
 - Certain elements were missing from this scheme.
 - For example, in 1871 Mendeleev noted that As properly belonged underneath P and not Si, which left a missing element underneath Si. He predicted a number of properties for this element.
 - In 1886 Ge was discovered; the properties of Ge match Mendeleev's predictions well.
- In the modern periodic table, elements are arranged in order of *increasing atomic number*.

FORWARD REFERENCES

- Periodic trends and chemical properties of nonmetals will be further discussed in Chapter 22 (section 22.1).

7.2 Effective Nuclear Charge^{13,14,15,16,17}

- **Effective nuclear charge** (Z_{eff}) is the charge experienced by an electron on a many-electron atom.
- The effective nuclear charge is not the same as the charge on the nucleus because of the effect of the inner electrons.
- The electron is attracted to the nucleus, but repelled by electrons that *shield* or *screen* it from the full nuclear charge.
- The nuclear charge experienced by an electron depends on its distance from the nucleus and the number of electrons in the spherical volume out to the electron in question.
- As the average number of screening electrons (S) increases, the effective nuclear charge (Z_{eff}) decreases.

$$Z_{\text{eff}} = Z - S$$

- As the distance from the nucleus increases, S increases and Z_{eff} decreases.
 - S is called the *screening constant*, which represents the portion of the nuclear charge that is screened from the valence electron by other electrons in the atom.
 - The value of S is usually close to the number of core electrons in an atom.

7.3 Sizes of Atoms and Ions^{18,19,20}

- Consider a collection of argon atoms in the gas phase.
 - When they undergo collisions, they ricochet apart because electron clouds cannot penetrate each other to a significant extent.
 - The *apparent* radius is determined by the closest distances separating the nuclei during such collisions.
 - This radius is the *nonbonding radius*.
 - Nonbonding atomic radii are also called *van der Waals radii*.
 - These are used in space-filling models to represent the sizes of different elements.

¹³ “Effective Nuclear Charge” Animation from Instructor’s Resource CD/DVD

¹⁴ Figure 7.2 from Transparency Pack

¹⁵ Figure 7.3 from Transparency Pack

¹⁶ “Screen Percentages Based on Slater Effective Nuclear Charge as a Versatile Tool for Teaching Periodic Trends” from Further Readings

¹⁷ Figure 7.4 from Transparency Pack

¹⁸ Figure 7.6 from Transparency Pack

¹⁹ “Pictorial Analogies: VI: Radial and Angular Wave Function Plots” from Further Readings

²⁰ “Methanethiol (methyl mercaptan)” 3-D Model from Instructor’s Resource CD/DVD

- Now consider a simple diatomic molecule.
 - The distance between the two nuclei is called the **bonding atomic radius**.
 - It is shorter than the nonbonding radius.
 - If the two atoms that make up the molecule are the same, then half the bond distance is called the covalent radius of the atom.

Periodic Trends in Atomic Radii^{21,22}

- Atomic size varies consistently through the periodic table.
 - As we move down a group, the atoms become larger.
 - As we move across a period, atoms become smaller.
 - There are two factors at work:
 - the principal quantum number, n , and
 - the effective nuclear charge, Z_{eff} .
 - As the principal quantum number increases (i.e., we move down a group), the distance of the outermost electron from the nucleus becomes larger. Hence, the atomic radius increases.
 - As we move across the periodic table, the number of core electrons remains constant; however, the nuclear charge increases. Therefore, there is an increased attraction between the nucleus and the outermost electrons. This attraction causes the atomic radius to decrease.

Periodic Trends in Ionic Radii^{23,24,25}

- Ionic size is important:
 - in predicting lattice energy and
 - in determining the way in which ions pack in a solid.
- Just as atomic size is periodic, ionic size is also periodic.
- In general:
 - Cations are smaller than their parent atoms.
 - Electrons have been removed from the most spatially extended orbital.
 - The effective nuclear charge has increased.
 - Therefore, the cation is smaller than the parent atom.
 - Anions are larger than their parent atoms.
 - Electrons have been added to the most spatially extended orbital.
 - This means total electron-electron repulsion has increased.
 - Therefore, anions are larger than their parent atoms.
- For ions with the same charge, ionic size increases down a group.
- All the members of an **isoelectronic series** have the same number of electrons.
 - As nuclear charge increases, in an isoelectronic series the ions become smaller:

$$\text{O}^{2-} > \text{F}^{-} > \text{Na}^{+} > \text{Mg}^{2+} > \text{Al}^{3+}$$

FORWARD REFERENCES

- Sizes and charges of ions will be instrumental in determining lattice energies (Chapter 8).
- Structures of ionic solids in Chapter 12 (section 12.2).
- Atomic radii will affect relative strengths of binary acids of nonmetals from a given group, as discussed in Chapter 16 (section 16.10).
- Periodic properties of nonmetals in groups 4A-8A are tabulated throughout Chapter 22.
- Periodic properties for the first transition-series elements are shown in Chapter 23 (section 23.1).

²¹ “Periodic Trends: Atomic Radii” Animation from Instructor’s Resource CD/DVD

²² “Using Balls from Different Sports to Model the Variation of Atomic Sizes” from Further Readings

²³ “Gain and Loss of Electrons” Animation from Instructor’s Resource CD/DVD

²⁴ Figure 7.7 from Transparency Pack

²⁵ “Periodic Contractions Among the Elements; Or, On Being the Right Size” from Further Readings

7.4 Ionization Energy^{26,27}

- The **ionization energy** of an atom or ion is the minimum energy required to remove an electron from the ground state of the isolated gaseous atom or ion.
- The *first ionization energy*, I_1 , is the amount of energy required to remove an electron from a gaseous atom:



- The *second ionization energy*, I_2 , is the energy required to remove the second electron from a gaseous ion:



- The larger the ionization energy, the more difficult it is to remove the electron.
- There is a sharp increase in ionization energy when a core electron is removed.

Variations in Successive Ionization Energies^{28,29}

- Ionization energies for an element increase in magnitude as successive electrons are removed.
 - As each successive electron is removed, more energy is required to pull an electron away from an increasingly more positive ion.
- A sharp increase in ionization energy occurs when an inner-shell electron is removed.

Periodic Trends in First Ionization Energies^{30,31,32}

- Ionization energy generally increases across a period.
 - As we move across a period, Z_{eff} increases, making it more difficult to remove an electron.
 - Two exceptions are removing the first p electron and removing the fourth p electron.
 - The s electrons are more effective at shielding than p electrons. So, forming the s^2p^0 configuration is more favorable.
 - When a second electron is placed in a p orbital, the electron-electron repulsion increases. When this electron is removed, the resulting s^2p^3 configuration is more stable than the starting s^2p^4 configuration. Therefore, there is a decrease in ionization energy.
- Ionization energy decreases down a group.
 - This means that the outermost electron is more readily removed as we go down a group.
 - As the atom gets bigger, it becomes easier to remove an electron from the most spatially extended orbital.
 - Example: For the noble gases, the ionization energies follow the order:
 $\text{He} > \text{Ne} > \text{Ar} > \text{Kr} > \text{Xe}$
- The representative elements exhibit a larger range of values for I_1 than transition metals.

Electron Configurations of Ions

- These are derived from the electron configurations of elements with the required number of electrons added or removed from the most accessible orbital.
 - Li: $[\text{He}]2s^1$ becomes Li^+ : $[\text{He}]$
 - F: $[\text{He}]2s^22p^5$ becomes F^- : $[\text{He}]2s^22p^6 = [\text{Ar}]$

²⁶ “Ionization Energy” Animation from Instructor’s Resource CD/DVD

²⁷ “Ionization Energies of Atoms and Atomic Ions” from Further Readings

²⁸ “Ionization Energies” Activity from Instructor’s Resource CD/DVD

²⁹ Table 7.2 from Transparency Pack

³⁰ Figure 7.9 from Transparency Pack

³¹ “Periodic Trends: Ionization Energies” Animation from Instructor’s Resource CD/DVD

³² “Trends in Ionization Energy of Transition-Metal Elements” from Further Readings

- 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 4*s* before the 3*d*, etc.
 - In other words, when writing electron configurations of transition metal cations, the order of removal of electrons is **not** exactly opposite to the order in which subshells were occupied when an electron configuration of the parent atom was written.

FORWARD REFERENCES

- Octet rule will be introduced in Chapter 8.
- Discussion of electron configurations of the representative elements and transition metals will continue in Chapter 8.
- Photoionization processes and ionization energies will be linked together in Chapter 18. (section 18.1).

7.5 Electron Affinities^{33,34,35,36}

- **Electron affinity** is the energy change when a gaseous atom gains an electron to form a gaseous ion.
- Electron affinity and ionization energy measure the energy changes of opposite processes.
 - Electron affinity: $\text{Cl}(g) + e^- \rightarrow \text{Cl}^-(g)$ $\Delta E = -349 \text{ kJ/mol}$
 - Ionization energy: $\text{Cl}(g) \rightarrow \text{Cl}^+(g) + e^-$ $\Delta E = 1251 \text{ kJ/mol}$
- Electron affinity can either be exothermic (as the above example) or endothermic:

$$\text{Ar}(g) + e^- \rightarrow \text{Ar}^-(g) \quad \Delta E > 0$$
- Look at electron configurations to determine whether electron affinity is positive or negative.
 - The extra electron in Ar needs to be placed in the 4*s* orbital, which is significantly higher in energy than the 3*p* orbital.
 - The added electron in Cl is placed in the 3*p* orbital to form the stable 3*p*⁶ electron configuration.
 - Electron affinities do not change greatly as we move down in a group.

7.6 Metals, Nonmetals, and Metalloids^{37,38,39}

- **Metallic character** refers to the extent to which the element exhibits the physical and chemical properties of metals.
 - Metallic character increases down a group.
 - Metallic character decreases from left to right across a period.

Metals^{40,41,42,43,44,45}

- Metals are shiny and lustrous, malleable and ductile.
- Metals are solids at room temperature (exception: mercury is liquid at room temperature; gallium and cesium melt just above room temperature) and have very high melting temperatures.

³³ “Electron Affinity” Animation from Instructor’s Resource CD/DVD

³⁴ “Periodic Trends: Electron Affinity” Animation from Instructor’s Resource CD/DVD

³⁵ Figure 7.11 from Transparency Pack

³⁶ “Halogens Compete for Electrons” from Live Demonstrations

³⁷ Figure 7.12 from Transparency Pack

³⁸ “Acidic and Basic Properties of Oxides” from Live Demonstrations

³⁹ “Water” 3-D Model from Instructor’s Resource CD/DVD

⁴⁰ Figure 7.14 from Transparency Pack

⁴¹ “Periodicity in the Acid-Base Behavior of Oxides and Hydroxides” from Further Readings

⁴² “Periodic Trends: Acid-Base Behavior of Oxides” Animation from Instructor’s Resource CD/DVD

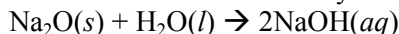
⁴³ “Acidic and Basic Properties of Oxides” from Live Demonstrations

⁴⁴ “Disappearing Ink” from Live Demonstrations

⁴⁵ “Hydroxide Ion” 3-D Model from Instructor’s Resource CD/DVD

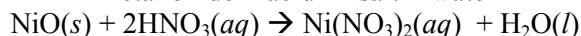
- Metals tend to have low ionization energies and tend to form cations easily.
- Metals tend to be oxidized when they react.
- Compounds of metals with nonmetals tend to be ionic substances.
- Metal oxides form basic ionic solids.
 - Most metal oxides are basic:

Metal oxide + water \rightarrow metal hydroxide



- Metal oxides are able to react with acids to form salts and water:

Metal oxide + acid \rightarrow salt + water



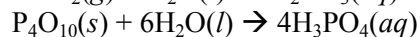
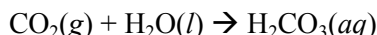
Nonmetals⁴⁶

- Nonmetals are more diverse in their behavior than metals.
- In general, nonmetals are nonlustrous, are poor conductors of heat and electricity, and exhibit lower melting points than metals.
- Seven nonmetallic elements exist as diatomic molecules under ordinary conditions:
 - $\text{H}_2(g)$, $\text{N}_2(g)$, $\text{O}_2(g)$, $\text{F}_2(g)$, $\text{Cl}_2(g)$, $\text{Br}_2(l)$, $\text{I}_2(s)$
- When nonmetals react with metals, nonmetals tend to gain electrons:

Metal + nonmetal \rightarrow salt

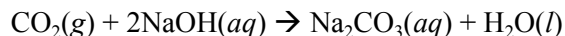
$$2\text{Al}(s) + 3\text{Br}_2(l) \rightarrow 2\text{AlBr}_3(s)$$
- Compounds composed entirely of nonmetals are molecular substances.
- Most nonmetal oxides are acidic:

Nonmetal oxide + water \rightarrow acid



- Nonmetal oxides react with bases to form salts and water:

Nonmetal oxide + base \rightarrow salt + water



Metalloids⁴⁷

- Metalloids have properties that are intermediate between those of metals and nonmetals.
 - For example, Si has a metallic luster but it is brittle.
- Metalloids have found fame in the semiconductor industry.

FORWARD REFERENCES

- The role of metals and metalloids in semiconductors will be discussed in Chapter 12 (section 12.7).
- Arrhenius, Brønsted-Lowry, and Lewis acids and bases will be discussed in Chapter 16.
- Acids and bases as well as reactions between them will be discussed in Chapter 16.
- An in-depth discussion of nonmetals will be provided in Chapter 22.
- Physical properties for the first transition-series elements are tabulated in Chapter 23 (section 23.1)

7.7 Trends for Group 1A and Group 2A Metals

- The **alkali metals** (group 1A) and the **alkaline earth metals** (group 2A) are often called the active metals.

⁴⁶ “Square Planar Structure” 3-D Model from Instructor’s Resource CD/DVD

⁴⁷ “Metalloids” from Further Readings

Group 1A: The Alkali Metals^{48,49,50,51,52,53,54,55}

- The alkali metals are in Group 1A.
- Alkali metals are all soft.
- Their chemistry is dominated by the loss of their single *s* electron:

$$M \rightarrow M^+ + e^-$$
- Reactivity increases as we move down the group.
- Alkali metals react with hydrogen to form hydrides.
 - In hydrides, the hydrogen is present as H^- , called the **hydride ion**.

$$2M(s) + H_2(g) \rightarrow 2MH(s)$$
- Alkali metals react with water to form MOH and hydrogen gas:

$$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$$
- Alkali metals produce different oxides when reacting with O_2 :
 - $4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$ (*oxide*)
 - $2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$ (*peroxide*)
 - $K(s) + O_2(g) \rightarrow KO_2(s)$ (*superoxide*)
- Alkali metals emit characteristic colors when placed in a high-temperature flame.
 - The *s* electron is excited by the flame and emits energy when it returns to the ground state.
 - The Na line occurs at 589 nm (yellow), characteristic of the $3p \rightarrow 3s$ transition.
 - The Li line is crimson red.
 - The K line is lilac.

Group 2A: The Alkaline Earth Metals^{56,57,58,59}

- Alkaline earth metals are harder and more dense than the alkali metals.
- Their chemistry is dominated by the loss of two *s* electrons:

$$M \rightarrow M^{2+} + 2e^-$$

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$
- Reactivity increases down the group.
 - Be does not react with water.
 - Mg will only react with steam.
 - Ca and the elements below it react with water at room temperature as follows:

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

7.8 Trends for Selected Nonmetals**Hydrogen**

- Hydrogen is a unique element.

⁴⁸ Table 7.4 from Transparency Pack⁴⁹ “Sodium and Potassium in Water” Movie from Instructor’s Resource CD/DVD⁵⁰ “Flame Tests for Metals” Movie from Instructor’s Resource CD/DVD⁵¹ “A Dramatic Flame Test Demonstration” from Live Demonstrations⁵² “Simple Flame Test Techniques Using Cotton Swabs” from Live Demonstrations⁵³ “The Legend of Dr. Pepper/ Seven-Up” from Further Readings⁵⁴ “A Little Lithium May Be Just What the Doctor Ordered” from Further Readings⁵⁵ “A Variation on the Determination of the Properties of the Alkali Metals” from Further Readings⁵⁶ Table 7.5 from Transparency Pack⁵⁷ “Producing Hydrogen Gas from Calcium Metal” from Live Demonstrations⁵⁸ “Life, Death, and Calcium” from Further Readings⁵⁹ “Update on Intake: Calcium Consumption Low” from Further Readings

- It most often occurs as a colorless diatomic gas, H_2 .
- Reactions between hydrogen and nonmetals can be very exothermic:

$$2\text{H}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) \quad \Delta H^\circ = -571.7 \text{ kJ}$$
- It can either gain another electron to form the hydride ion, H^- , or lose its electron to become H^+ :

$$2\text{Na}(\text{s}) + \text{H}_2(\text{g}) \rightarrow 2\text{NaH}(\text{s})$$

$$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$$
- H^+ is a proton.
- The aqueous chemistry of hydrogen is dominated by $\text{H}^+(\text{aq})$.

Group 6A: The Oxygen Group^{60,61,62,63,64,65,66,67,68,69,70,71}

- As we move down the group, the metallic character increases.
 - O_2 is a gas, Te is a metalloid, Po is a metal.
- Two of the important forms of oxygen are O_2 and **ozone**, O_3 .
 - O_2 and O_3 are allotropes.
 - Allotropes are different forms of the same element in the same state (in this case, gaseous).
 - Ozone can be prepared from oxygen:

$$3\text{O}_2(\text{g}) \rightarrow 2\text{O}_3(\text{g}) \quad \Delta H^\circ = +284.6 \text{ kJ}$$
 - Ozone is pungent and toxic.
 - Oxygen (or dioxygen, O_2) is a potent oxidizing agent since the O^{2-} ion has a noble gas configuration.
 - There are two oxidation states for oxygen: -2 (e.g., H_2O) and -1 (e.g., H_2O_2).
- Sulfur is another important member of this group.
 - The most common form of sulfur is yellow S_8 .
 - Sulfur tends to form S^{2-} in compounds (sulfides).

Group 7A: The Halogens^{72,73,74,75,76,77}

- Group 7A elements are known as the **halogens** ("salt formers").
- The chemistry of the halogens is dominated by gaining an electron to form an anion:

$$\text{X}_2 + 2\text{e}^- \rightarrow 2\text{X}^-$$
- Fluorine is one of the most reactive substances known:

$$2\text{F}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{HF}(\text{aq}) + \text{O}_2(\text{g}) \quad \Delta H = -758.9 \text{ kJ}$$

⁶⁰ Table 7.6 from Transparency Pack

⁶¹ "A Second Note on the Term Chalcogen" from Further Readings

⁶² "Preparation and Properties of Oxygen" from Live Demonstrations

⁶³ "Plastic Sulfur" from Live Demonstrations

⁶⁴ "Allotropes and Polymorphs" from Further Readings

⁶⁵ "The Origin of the Term Allotrope" from Further Readings

⁶⁶ "Aqueous Hydrogen Peroxide: Its Household Uses and Concentration Units" from Further Readings

⁶⁷ "Preparation and Properties of Oxygen" from Live Demonstrations

⁶⁸ "Plastic Sulfur" from Live Demonstrations

⁶⁹ "Reactions with Oxygen" Movie from Instructor's Resource CD/DVD

⁷⁰ "Oxygen" 3-D Model from Instructor's Resource CD/DVD

⁷¹ "Sulfur" 3-D Model from Instructor's Resource CD/DVD

⁷² Table 7.7 from Transparency Pack

⁷³ "Physical Properties of the Halogens" Movie from Instructor's Resource CD/DVD

⁷⁴ "The Chemistry of Swimming Pool Maintenance" from Further Readings

⁷⁵ "Chlorine" 3-D Model from Instructor's Resource CD/DVD

⁷⁶ "Bromine" 3-D Model from Instructor's Resource CD/DVD

⁷⁷ "Iodine" 3-D Model from Instructor's Resource CD/DVD

- All halogens consist of diatomic molecules, X_2 .
- Chlorine is the most industrially useful halogen.
 - In 2008 total production of chlorine was 21 billion pounds, making it one of the top ten most produced chemicals in the US.
 - The reaction between chlorine and water produces hypochlorous acid (HOCl), which is used to disinfect swimming pool water:



- Halogens react with hydrogen to form gaseous hydrogen halide compounds:

$$\text{H}_2(g) + \text{X}_2 \rightarrow 2\text{HX}(g)$$
- Hydrogen compounds of the halogens are all strong acids with the exception of HF.

Group 8A: The Noble Gases⁷⁸

- The group 8A elements are known as the **noble gases**.
 - These are all nonmetals and monoatomic.
 - They are notoriously unreactive because they have completely filled *s* and *p* subshells.
- In 1962 the first compounds of the noble gases were prepared: XeF_2 , XeF_4 , and XeF_6 .
- In 2000 Finnish scientists reported the first neutral molecule that contains argon: the HArF molecule, which is stable only at low temperatures.

FORWARD REFERENCES

- Electron configurations of noble gases and the octet rule will be introduced in Chapter 8 (section 8.1).
- The role of the expanded octet in the formation of compounds involving heavier noble gas atoms will be discussed in Chapter 8 (section 8.7).

⁷⁸ Table 7.8 from Transparency Pack

Further Readings:

1. The **September 8, 2003** issue of *Chemical and Engineering News* is a special issue celebrating the periodic table of the elements on C&EN's 80th anniversary. Approximately 90 short articles, each featuring a different element, are found in this issue.
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5. Eric R. Scerri, "The Evolution of the Periodic System," *Scientific American*, **September 1998**, 78–83.
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8. Yu. I. Solov'ev, "D. I. Mendeleev and the English Chemists," *J. Chem. Educ.*, Vol. 61, **1984**, 1069–1071.
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11. Kimberley A. Waldron, Eric M. Fehringer, Amy E. Streeb, Jennifer E. Trosky, and Joshua J. Pearson, "Screen Percentages Based on Slater Effective Nuclear Charge as a Versatile Tool for Teaching Periodic Trends," *J. Chem. Educ.*, Vol. 78, **2001**, 635–639.
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20. Michael W. Miller, "A Little Lithium May Be Just What the Doctor Ordered," *Wall Street Journal*, **September 23, 1994**.
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Live Demonstrations:

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