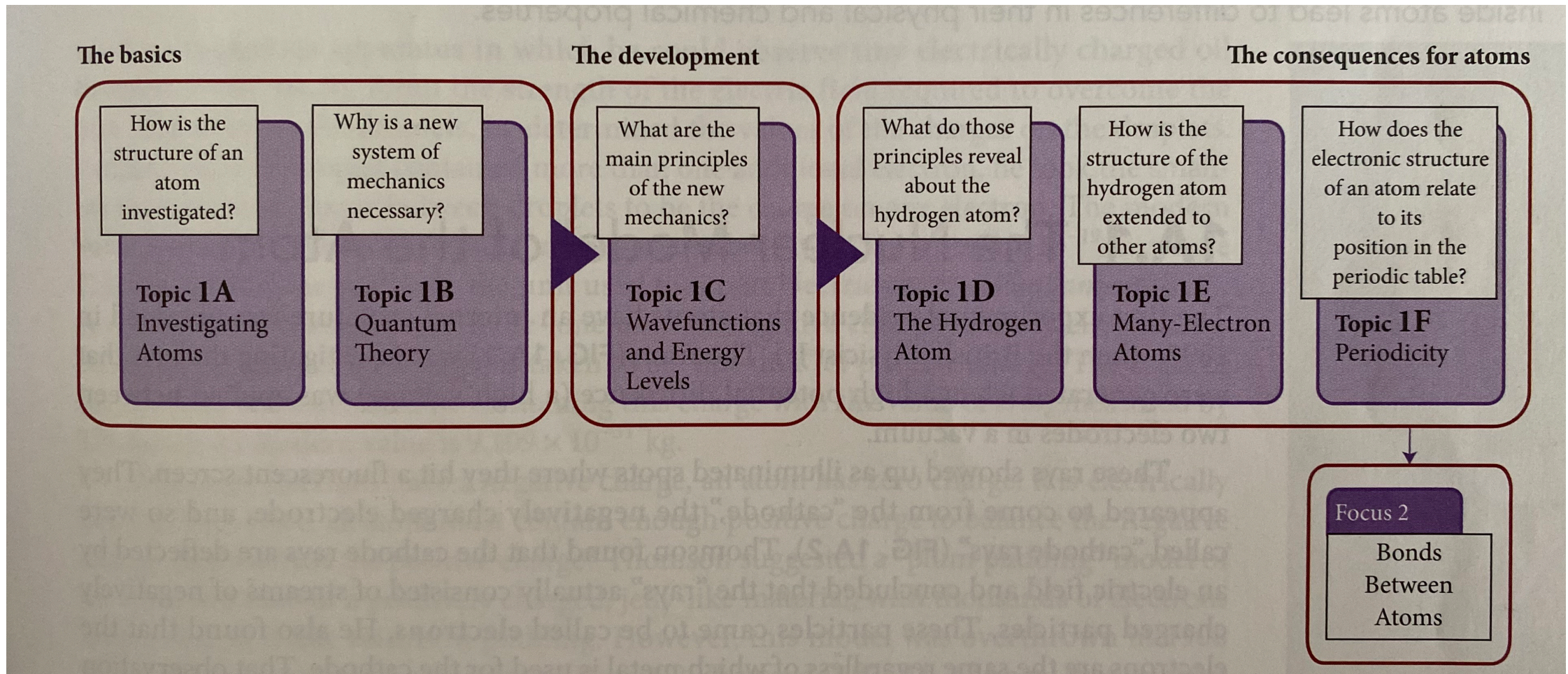




# CH-110 Advanced General Chemistry I

Dr. Milena Schuhmacher

# Overview Chapter 1 (Focus 1: Atoms)



# Periodicity

Topic 1F



Topic 1F.1 Atomic radius

Topic 1F.2 Ionic radius

Topic 1F.3 Ionization energy

Topic 1F.4 Electron affinity

Topic 1F.5 Electronegativity

Topic 1F.6 The general properties of the elements

WHY DO YOU NEED TO KNOW THIS MATERIAL?

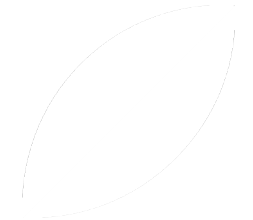
- The periodic table summarizes trends in the properties of the elements. The ability to **predict the properties of an element from its location** in the periodic table is a central skill of a chemist.

WHAT DO YOU NEED TO KNOW ALREADY?

- The structure of the **periodic table**, the **building-up principle** (Topic 1E)
- **Oxidation state** (Fundamentals K)
- Definition of **ionization energy** (Topic 1D)

# Atomic Radius

Topic 1F.1



## 1F.1 Atomic radius

# Effective nuclear charge, $Z_{eff}$ and Shielding

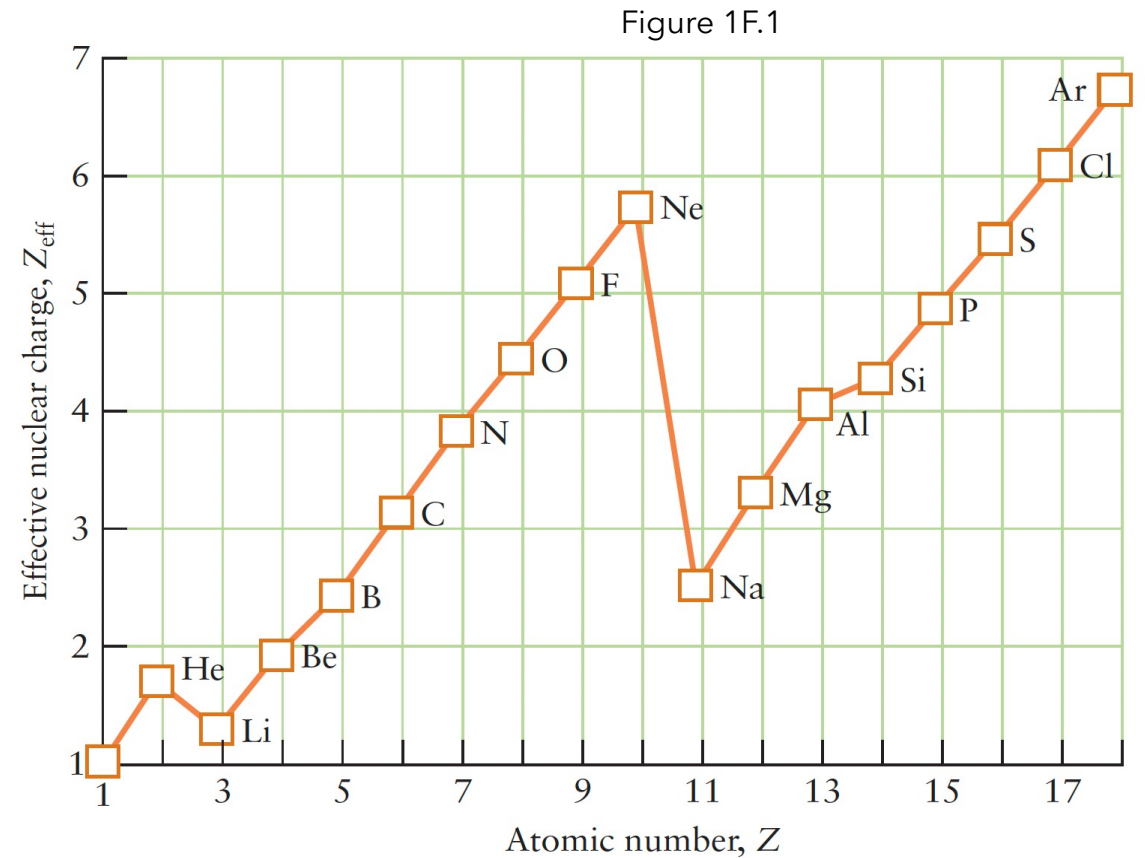
The **effective nuclear charge** is the net positive charge that a valence electron feels after accounting for the shielding effect of inner electrons.



# 1F.1 Atomic radius

## Effective nuclear charge, $Z_{eff}$

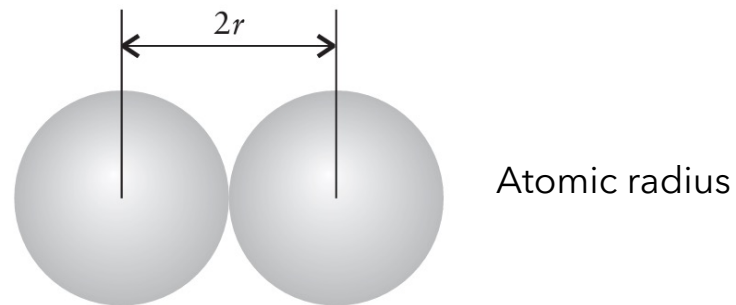
- The **effective nuclear charge** is the net positive charge that a valence electron feels after accounting for the shielding effect of inner electrons.
- The **effective nuclear charge**,  $Z_{eff}e$ , is always smaller than  $Ze$ .



# 1F.1 Atomic radius

## Different types of radii

- Electron clouds have no sharp boundaries → exact radius is technically undefined
- Atomic radius: half the distance between centers of neighbouring atoms
- **For metals:** defined as half of the distance between neighbouring nuclei in the solid sample
- Example: Cu-Cu distance = 256 pm → atomic radius = 128 pm



# 1F.1 Atomic radius

## Different types of radii

- **For nonmetals/metalloids:** covalent radius = half the bond length of atoms joined together (see Topic 2D)

*Example:  $Cl_2$  bond = 198 pm  $\rightarrow$  radius(Cl) = 99 pm.*

- **Noble gases:** van der Waals radius = half the distance between neighbouring atoms in solidified gas

*Larger than covalent radius; not included in discussion of periodic trends*

# 1F.1 Atomic radius

## Atomic radii of main-group elements

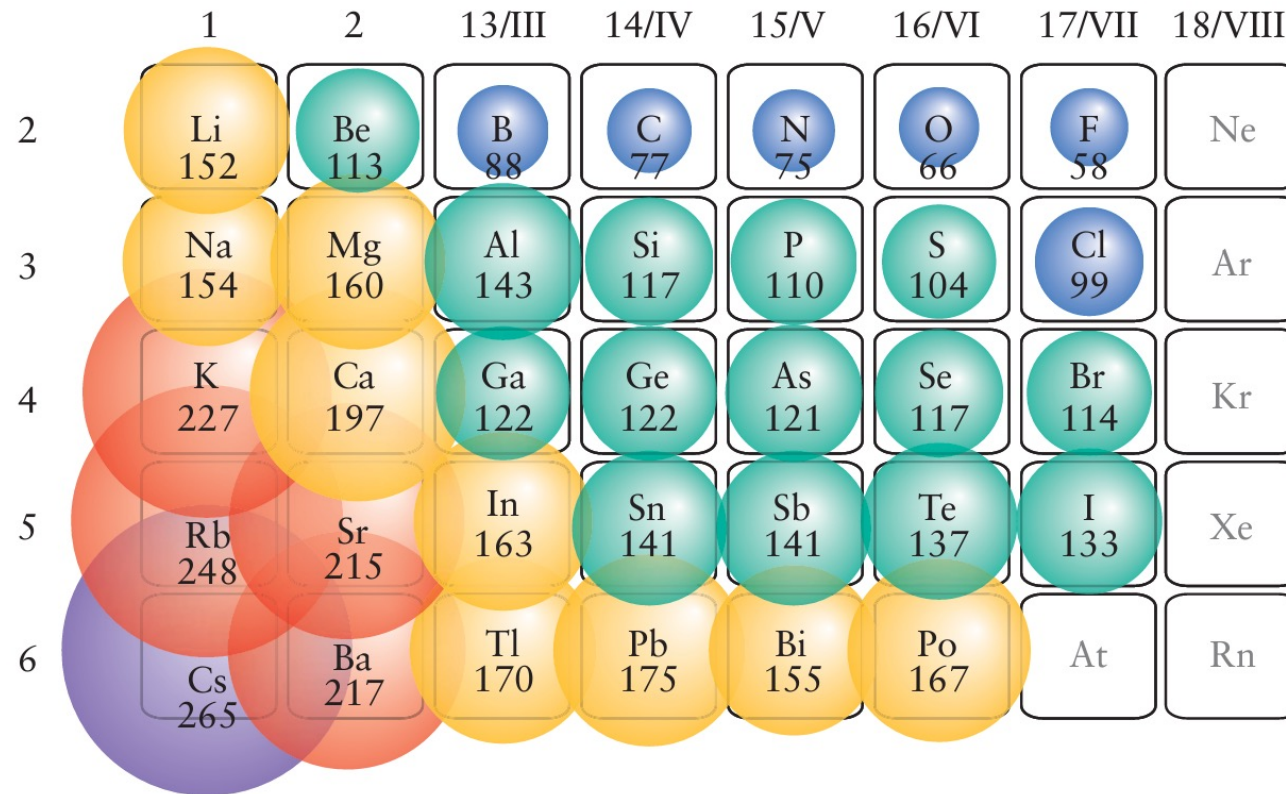


Figure 1F.2

**TREND:** Atomic radius generally decreases from left to right across a period and increases down a group.

## 1F.1 Atomic radius

### Summary: Atomic radii of main-group elements

**TREND:** Atomic radius generally decreases from left to right across a period and increases down a group.

↓ **Group (Li → Cs):** radius ↑

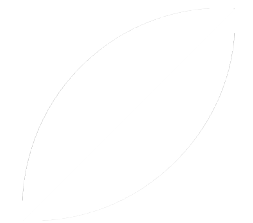
- Electrons added to higher shells (larger  $n$ ) → farther from nucleus

→ **Period (Li → Ne):** radius ↓

- Electrons added to same shell.
- Poor shielding → effective nuclear charge ↑
- Stronger attraction pulls electrons closer → atom more compact

# Ionic Radius

Topic 1F.2



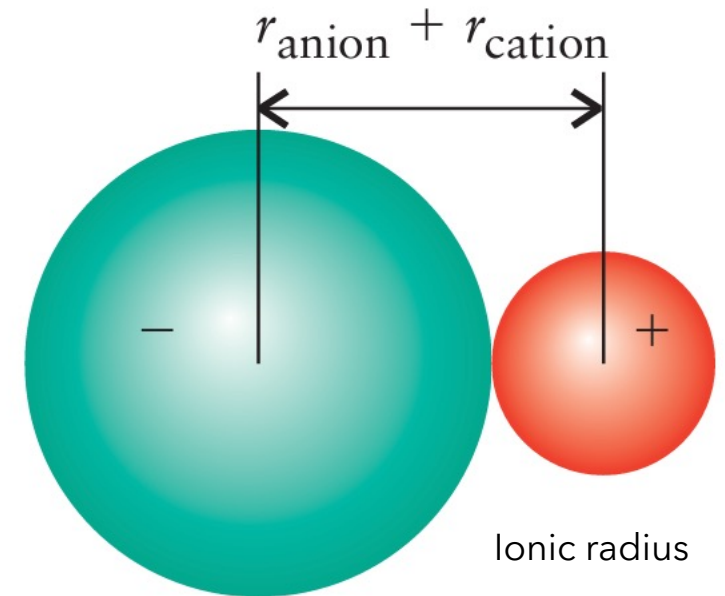
## 1F.2 Ionic radius

### Another type of radius: ionic radius

- The radii of ions are **very different** to the radii of their parent atoms.

#### How is it measured:

- Each ion in an ionic solid is surrounded by ions of opposite charge.
- The ionic radius of an element is its share of the distance between neighboring ions in an ionic solid.
- In practice, the radius of the oxide ion is taken to be 140 pm, and the radii of other atoms are then calculated on the basis of that value.
- For example, the distance of  $\text{Mg}^{2+}$  and  $\text{O}^{2-}$  ions in magnesium oxide is 212 pm, the radius of the  $\text{Mg}^{2+}$  ion is reported as  $212 - 140 \text{ pm} = 72 \text{ pm}$ .



## 1F.2 Ionic radius

### Ionic radii of the ions of the main-group elements

	1	2	13/III	14/IV	15/V	16/VI	17/VII	18/VIII
2	$\text{Li}^+$ 76	$\text{Be}^{2+}$ 45	$\text{B}^{3+}$ 23	C	$\text{N}^{3-}$ 171	$\text{O}^{2-}$ 140	$\text{F}^-$ 133	Ne
3	$\text{Na}^+$ 102	$\text{Mg}^{2+}$ 72	$\text{Al}^{3+}$ 54	Si	$\text{P}^{3-}$ 212	$\text{S}^{2-}$ 184	$\text{Cl}^-$ 181	Ar
4	$\text{K}^+$ 138	$\text{Ca}^{2+}$ 100	$\text{Ga}^{3+}$ 62	Ge	$\text{As}^{3-}$ 222	$\text{Se}^{2-}$ 198	$\text{Br}^-$ 196	Kr
5	$\text{Rb}^+$ 152	$\text{Sr}^{2+}$ 118	$\text{In}^{3+}$ 80	Sn	$\text{Sb}^{3-}$ 221	$\text{Te}^{2-}$ 221	$\text{I}^-$ 220	Xe
6	$\text{Cs}^+$ 167	$\text{Ba}^{2+}$ 135	$\text{Tl}^{3+}$ 89	Pb	Bi	Po	At	Rn

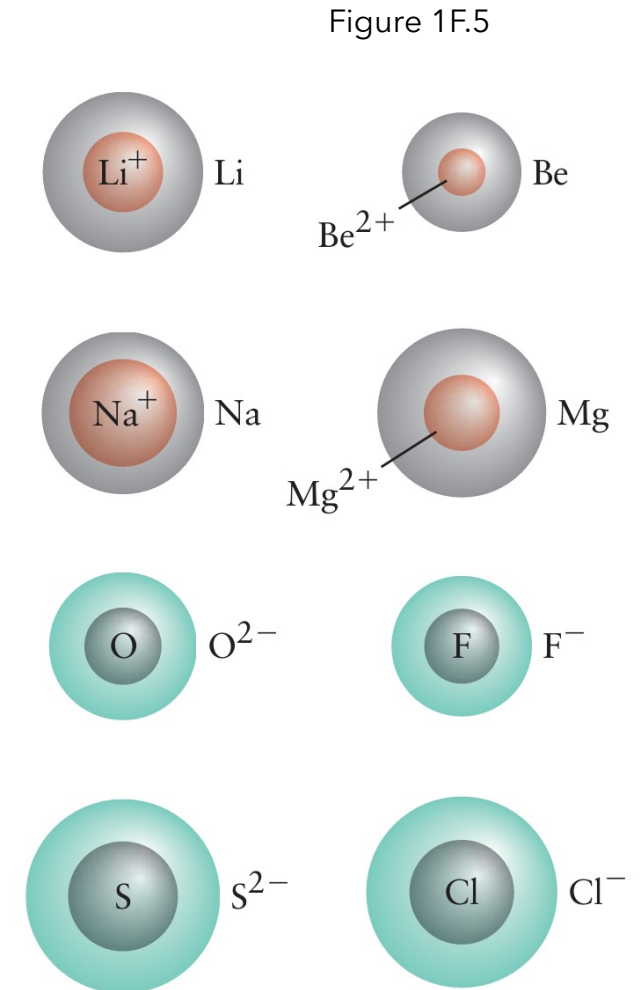
Figure 1F.4

**TREND:** Cations are smaller than their parent atoms, whereas anions are larger.

## 1F.2 Ionic radius

### Ionic radii of the ions of the main-group elements

- **Anions are larger than their parent atom:** increased number of electrons in the valence shell  $\rightarrow$  repulsive effects exerted by electrons on one another.
- **Isoelectronic:** atoms and ions with the same number of electrons.
- $\text{Na}^+$ ,  $\text{F}^-$ , and  $\text{Mg}^{2+}$  are isoelectronic with their electronic configuration  $[\text{He}]2s^22p^6$ . Radii differ, because of different nuclear charge.



**TREND:** Cations are smaller than their parent atoms, whereas anions are larger.

## 1F.2 Ionic radius

### Example 1F.1 Predicting the relative sizes of ions

- Arrange each of the following pairs of ions in order of increasing ionic radius

(a)  $\text{Mg}^{2+}$  and  $\text{Ca}^{2+}$

(b)  $\text{O}^{2-}$  and  $\text{F}^-$

## 1F.2 Ionic radius

### Example 1F.1 Predicting the relative sizes of ions

#### EXAMPLE 1.11 Deciding the relative sizes of ions

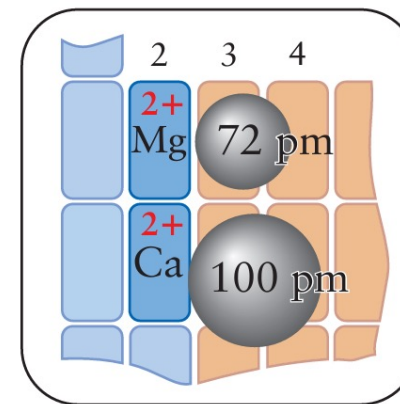
Arrange each of the following pairs of ions in order of increasing ionic radius: (a)  $\text{Mg}^{2+}$  and  $\text{Ca}^{2+}$ ; (b)  $\text{O}^{2-}$  and  $\text{F}^{-}$ .

**PLAN** The smaller member of a pair of isoelectronic ions in the same period will be an ion of an element that lies farther to the right in a period, because that ion has the greater effective nuclear charge. If the two ions are in the same group, the smaller ion will be the one that lies higher in the group, because its outermost electrons are closer to the nucleus.

#### SOLVE

(a) Mg lies above Ca in Group 2.

$\text{Mg}^{2+}$  has the smaller ionic radius.

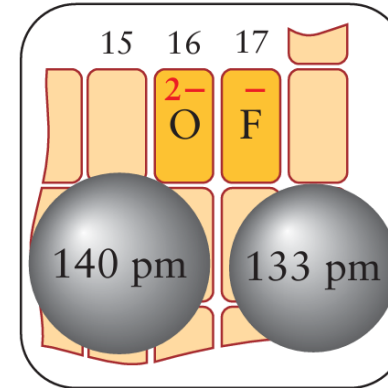


## 1F.2 Ionic radius

### Example 1F.1 Predicting the relative sizes of ions

(b)  $F^-$  lies to the right of O in Period 2.

$F^-$  has the smaller ionic radius.



**Evaluate** Appendix 2C shows that the actual values are (a) 72 pm for  $Mg^{2+}$  and 100 pm for  $Ca^{2+}$ ; (b) 133 pm for  $F^-$  and 140 pm for  $O^{2-}$ .

**Self-Test 1.13A** Arrange each of the following pairs of ions in order of increasing ionic radius: (a)  $Mg^{2+}$  and  $Al^{3+}$ ; (b)  $O^{2-}$  and  $S^{2-}$ .

[*Answer:* (a)  $r(Al^{3+}) < r(Mg^{2+})$ ; (b)  $r(O^{2-}) < r(S^{2-})$ ]

**Self-Test 1.13B** Arrange each of the following pairs of ions in order of increasing ionic radius: (a)  $Ca^{2+}$  and  $K^+$ ; (b)  $S^{2-}$  and  $Cl^-$ .

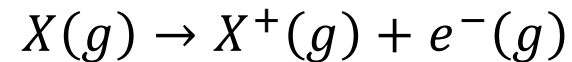
## 1F.2 Ionic radius

### Summary

**TREND:** Ionic radii generally increase down a group and decrease from left to right across a period. Cations are smaller than their parent atoms, and anions are larger.

## 1F.3 Ionization energy

- The formation of an **ionic compound** depends on the **removal of one or more electrons from one atom** (Na to Na<sup>+</sup>) and the transfer of those electrons to another atom (Cl to Cl<sup>-</sup>).
- The **ionization energy**,  $I$ , is the minimum energy needed to remove an electron from an atom in the gas phase:

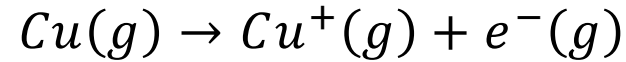


$$I = E(X^+) - E(X)$$

- Where  $E(X)$  is the energy of species X.
- And  $E(X^+) - E(X)$  is the energy difference between the ionized and the neutral atom.
- Reported as molar quantities in kilojoules per mole (kJ/mol) or in electronvolts (eV).

## 1F.3 Ionization energy

- **First ionization energy  $I_1$** : minimum energy needed to remove an electron from a neutral atom in the gas phase



$$\text{energy required} = I_1 \left( 746 \frac{\text{kJ}}{\text{mol}}, 7.73 \text{ eV} \right)$$

- Ionization energy is a measure of how difficult it is to remove an electron:
  - Elements with **low ionization energies** can be expected to form cations readily, conduct electricity (requires movement of free electrons)
  - Elements with **high ionization energies** are unlikely to form cations and are unlikely to conduct electricity

# 1F.3 Ionization energy

Figure 1F.7

	1	2	13/III	14/IV	15/V	16/VI	17/VII	18/VIII
2	Li 519	Be 900	B 799	C 1090	N 1400	O 1310	F 1680	He 2370
3	Na 494	Mg 736	Al 577	Si 786	P 1011	S 1000	Cl 1255	Ar 1520
4	K 418	Ca 590	Ga 577	Ge 784	As 947	Se 941	Br 1140	Kr 1350
5	Rb 402	Sr 548	In 556	Sn 707	Sb 834	Te 870	I 1008	Xe 1170
6	Cs 376	Ba 502	Tl 590	Pb 716	Bi 703	Po 812	At 1037	Rn 1036

- **Decrease down group:** outermost electrons are farther and farther away from nucleus, experiencing less nuclear charge.

**TREND:** First ionization energies typically decrease down a group.

First ionization energies generally increase across a period.

# 1F.3 Ionization energy

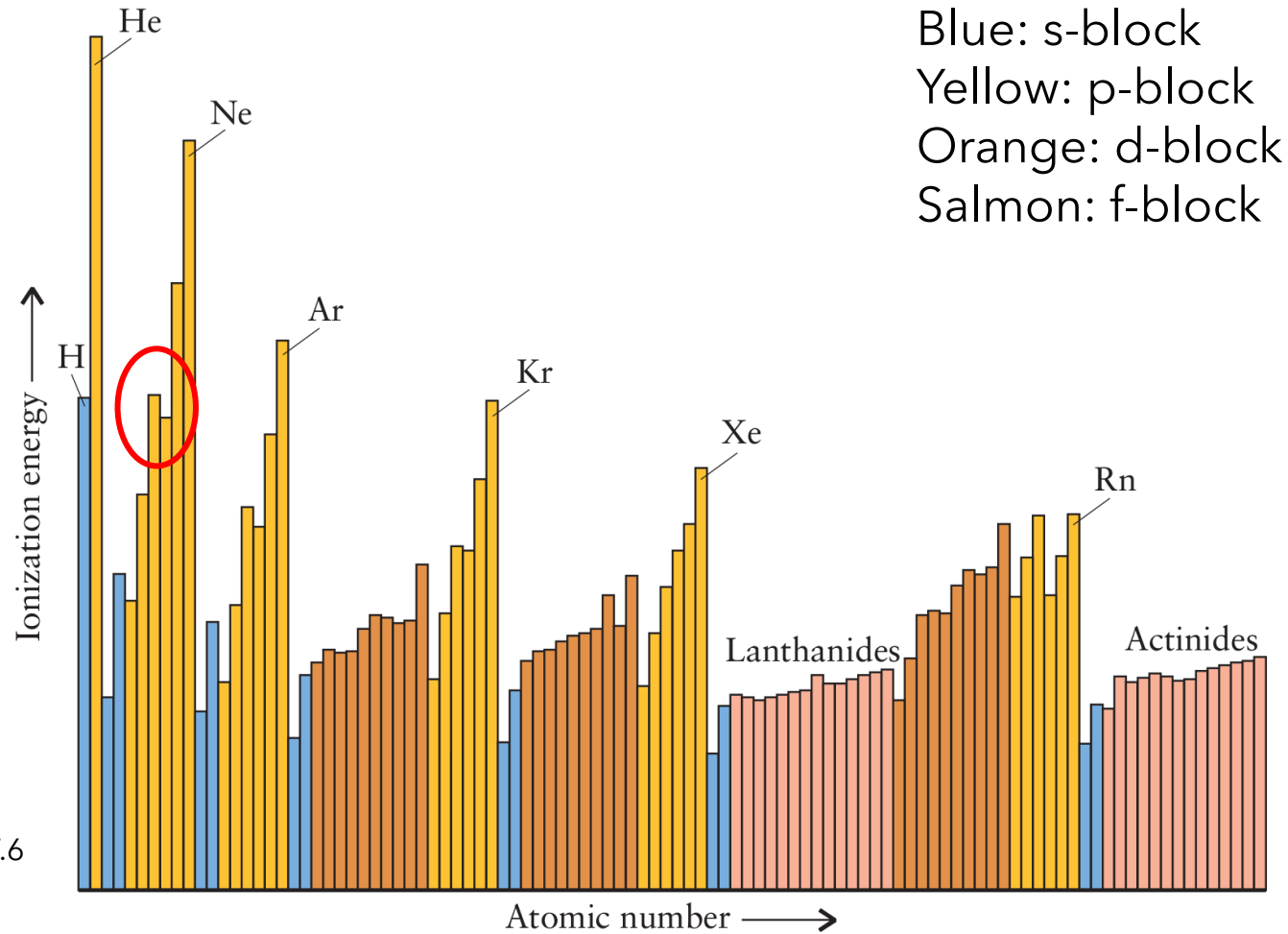


Figure 1F.6

**TRENDS:** First ionization energies typically decrease down a group.

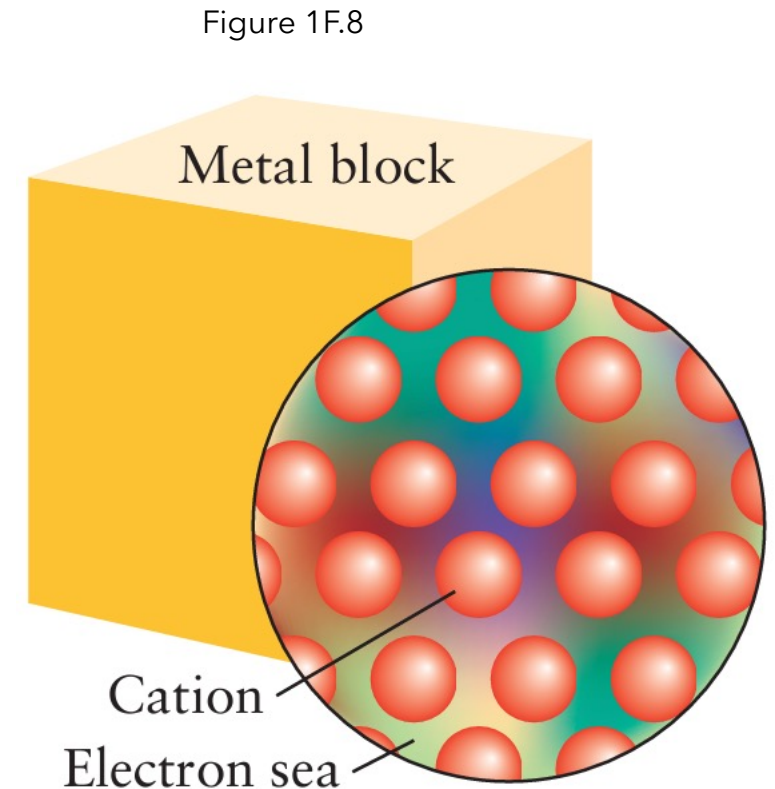
First ionization energies generally increase across a period.

- First ionization energies **increase across period:** increase in effective nuclear charge across period.
- **Ionization energy of oxygen < nitrogen:** nitrogen has one electron per p-orbital, oxygen is pairing up the 8th electron with other p-electrons, energy goes up

# 1F.3 Ionization energy

## Metallic character

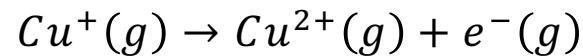
- **Low ionization energies**
- Lower left of periodic table
- A block of metal: **collection of cations of the element surrounded by a sea of valence electrons that the atoms have lost**
- s-block, d-block, f-block, lower left of the p-block can form metallic solids because they can lose electrons easily.
- Elements with high ionization energies are on the upper right of the periodic table



# 1F.3 Ionization energy

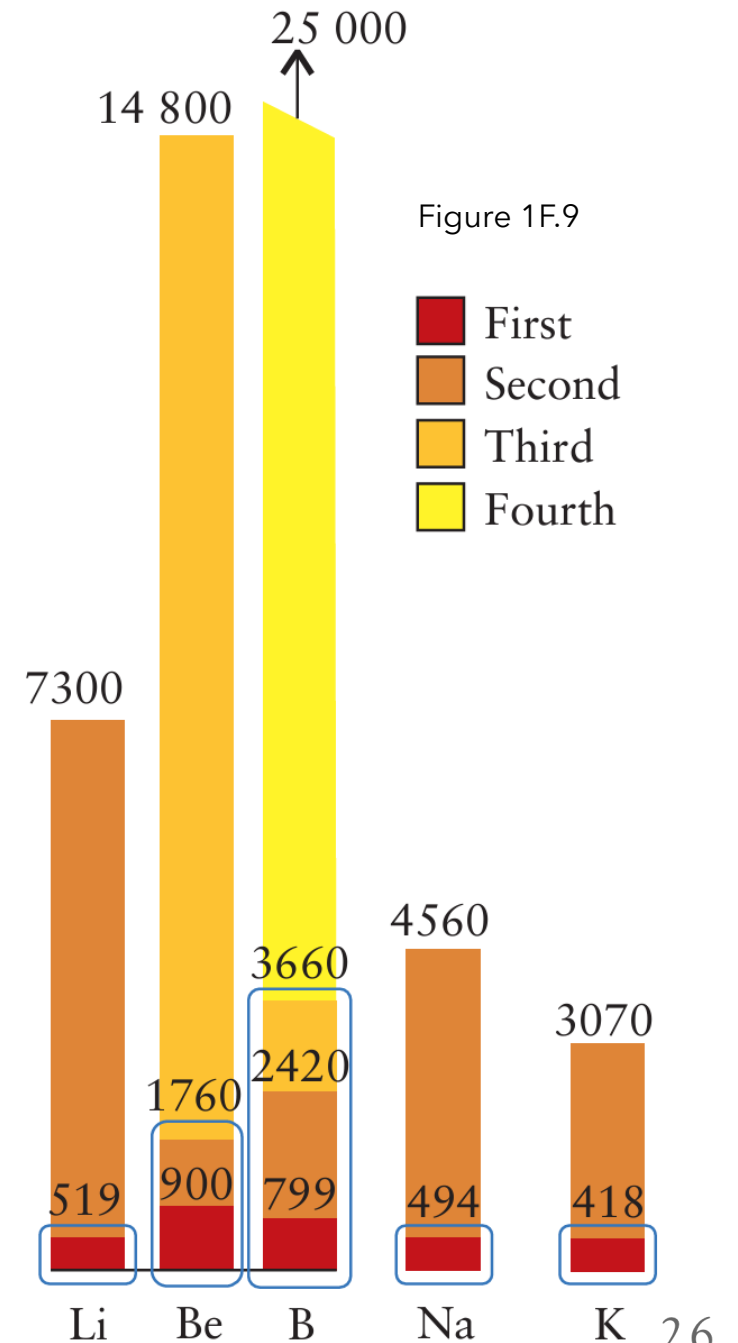
## The second ionization energy, $I_2$

- The minimum energy needed to remove an electron from a singly charged gas-phase cation.



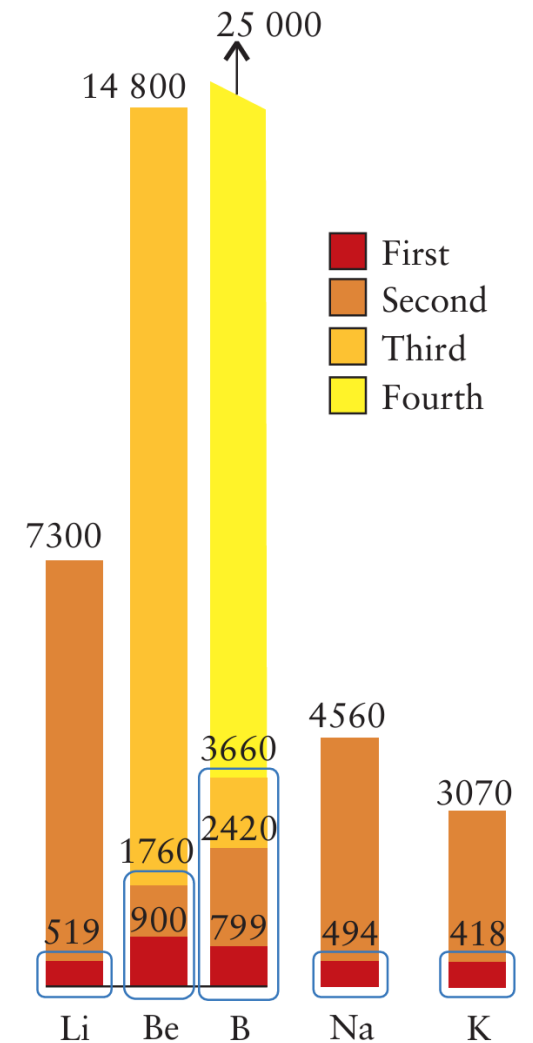
$$\text{energy required} = I_2 \left( 1958 \frac{\text{kJ}}{\text{mol}}, 20.29 \text{ eV} \right)$$

- The second ionization energy is **always higher than the first**  
→ it takes more energy to remove an electron from a positively charged atom than from a neutral one
- Group 1:  $I_2 \gg I_1$
- Group 2:  $I_2 > I_1$



# Why is there a large decrease in third ionization energy between beryllium and boron?

There is a large decrease in third ionization energy between beryllium and boron because, after removing two electrons from beryllium, the third ionization involves removing an electron from a stable, fully-filled noble gas core, which requires significantly more energy, while boron's third electron is from a higher-energy p-orbital, making it easier to remove.



# 1F.3 Ionization energy

## Summary

### TRENDS:

The **first ionization energy** is highest for elements close to helium and is lowest for elements close to cesium.

**Second ionization energies** are higher than first ionization energies (of the same element) and very much higher if the electron is to be removed from a closed shell.

**Metals are found toward the lower left of the periodic table** because these elements have low ionization energies and can readily lose their electrons.

# Electron Affinity

Topic 1F.4



## 1F.4 Electron affinity

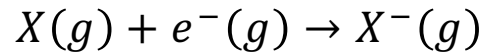
### Ionization energy vs. electron affinity

- Ionization energy indicates how difficult it is to **remove** an electron from an atom.
- How does the energy change when an electron **attaches** to an atom to form a negative ion (Cl to Cl<sup>-</sup>)?
- **The electron affinity,  $E_{ea}$ , of an element is the energy released when an electron is added to a gas-phase atom.**
- **Pos.  $E_{ea}$ :** energy released when an electron attaches to atom
- **Neg.  $E_{ea}$ :** energy must be supplied to push an electron onto an atom

## 1F.4 Electron affinity

### Ionization energy vs. electron affinity

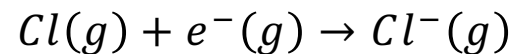
The electron affinity of an element  $X$  is defined as



$$E_{ea}(X) = E(X) - E(X^{-})$$

Where  $E(X)$  is the energy of a gas-phase atom  $X$  and  $E(X^{-})$  is the energy of the gas-phase anion.

For instance, the electron affinity of chlorine is the energy released in the process



$$\text{energy released} = E_{ea} \left( 349 \frac{\text{kJ}}{\text{mol}}, 3.63 \text{ eV} \right)$$

Because the electron has a lower energy when it occupies one of the atom's orbitals, the difference  $E(Cl) - E(Cl^{-})$  is positive, and the electron affinity of chlorine is positive.

## 1F.4 Electron affinity

### Variation in electron affinity in kJ/mol of main-group elements

			H +73				18/VIII He <0	
	1	2	13/III	14/IV	15/V	16/VI	17/VII	
2	Li +60	Be ≤0	B +27	C +122	N -7	O +141 -844	F +328	Ne <0
3	Na +53	Mg ≤0	Al +43	Si +134	P +72	S +200, -532	Cl +349	Ar <0
4	K +48	Ca +2	Ga +29	Ge +116	As +78	Se +195	Br +325	Kr <0
5	Rb +47	Sr +5	In +29	Sn +116	Sb +103	Te +190	I +295	Xe <0
6	Cs +46	Ba +14	Tl +19	Pb +35	Bi +91	Po +174	At +270	Rn <0

**TREND:** Electron affinities are highest toward the right of the periodic table.

**Pos.  $E_{ea}$ :** energy released when an electron attaches to atom

**Neg.  $E_{ea}$ :** energy must be supplied to push an electron onto an atom

Figure 1F.10

## 1F.4 Electron affinity

### TRENDS

**TREND:** Elements with the highest electron affinities are those in Groups 16 and 17.

- Particularly true in the upper right of the periodic table: **oxygen, sulfur, halogens**
- **Noble gases have negative electron affinities** because any electron added to them must occupy an orbital outside a closed shell and far from the nucleus: this process requires energy, the electron affinity is negative
- **Halide (group 17) plus electron builds a closed shell** (noble-gas configuration). Second electron affinity for halides is strongly negative.
- Group 16 (O or S): two vacancies in its valence shell p-orbitals, can accommodate two additional electrons. First electron affinity positive.

# Electronegativity

Topic 1F.5



## 1F.5 Electronegativity

### A combination of ionization energy and electron affinity

- Introduced by American chemist Robert Mulliken in 1934:
- Electronegativity,  $X$  (Greek letter chi)
- **“The power of an atom to attract electrons to itself when part of a compound.”**  
(see Topic 2D)

#### **Definition:**

$$X = \frac{1}{2}(I + E_a)/eV$$

With the ionization energy and the electron affinity expressed in electronvolts (so electronegativity is **unitless**).

# 1F.5 Electronegativity

## Variation in the electronegativities of the main-group elements

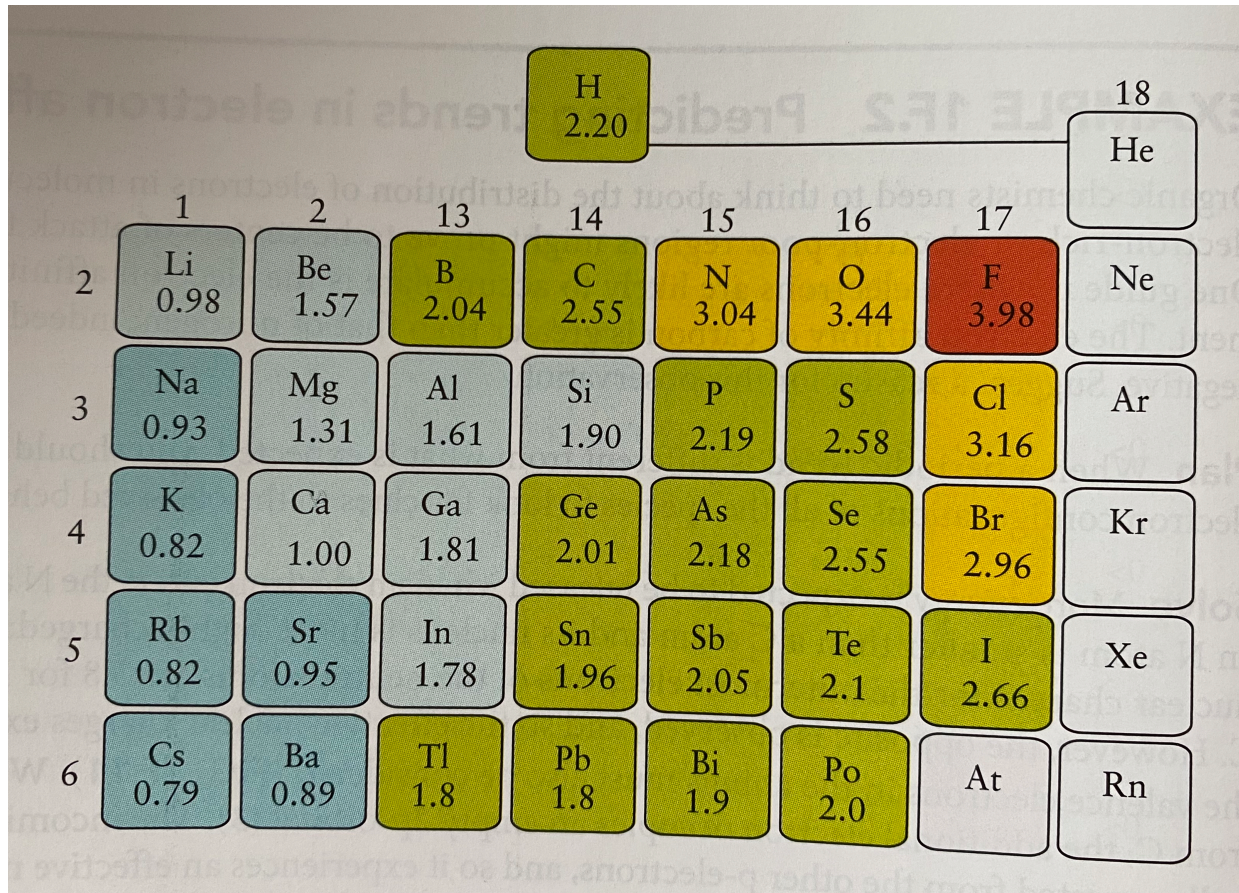


Figure 1F.12

- Elements with **high ionization energies and high electron affinities**: highly electronegative
- Elements **with low ionization energies and low electron affinities**: low electronegativity.
- Careful: electropositive is used for a different concept and is NOT the opposite of electronegative.

# 1F.5 Electronegativity

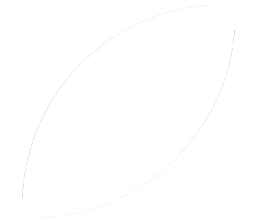
## Summary

**TREND:** The electronegativity of an element, its power of an atom to attract electrons to itself when part of a compound, is highest close to the top right of the periodic table.

$F > O > Cl > N > Br > I > S > C/Se > H > \dots$

# The General Properties of the Elements

Topic 1F.6



# 1F.6 The general properties of the elements

**TABLE 1F.1** Characteristics of Metallic and Nonmetallic Elements

<b>Metallic</b>	<b>Nonmetallic</b>
<b>Physical properties</b> good conductors of electricity malleable ductile lustrous typically: solid; high melting point; good conductors of heat	poor conductors of electricity not malleable not ductile not lustrous typically: solid, liquid, or gas; low melting point; poor conductors of heat
<b>Chemical properties</b> react with acids form basic oxides form cations form ionic halides	do not react with acids form acidic oxides form anions form covalent halides

# 1F.6 The general properties of the elements

## Alkali metals (Group 1)

- **s-block** elements
- Low ionization energy: valence electrons easily lost
- Soft, silvery, reactive metals
- Low melting points
- Produce hydrogen when they come in contact with water.
- Sodium is kept under mineral oil to protect it from air (**Fig. 1F.13**)
- A freshly cut surface becomes quickly covered with oxide.

PERIODIC TABLE OF THE ELEMENTS

Group 1 2 13 14 15 16 17 18  
I II IIIA IVA V VIA VIIA VIIIA  
1 2 10 11 12 13 14 15 16 17 18  
Period 1 H He  
Period 2 Li Be B C N O F Ne  
Period 3 Na Mg Al Si P S Cl Ar  
Period 4 K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr  
Period 5 Rb Sr Y Zr Nb Mo Ru Rh Pd Ag Cd In Sn Sb Te I Xe  
Period 6 Cs Ba La Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu  
Period 7 Fr Ra Ac Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

Major masses (atomic weights) listed for the majority of elements. Significant figures given here can be regarded as typical of most naturally occurring samples.

\*The names of the elements 112 and higher have not yet been determined; both 112 and 114 have been confirmed.



Figure 1F.13

# 1F.6 The general properties of the elements

## Group 14 elements

- Elements on the left of the **p-block**
- As you move down the group, they have low enough ionization energies to have metallic character.
- From left to right (**Fig. 1F.14**): carbon (as graphite), silicon, germanium, tin, and lead

PERIODIC TABLE OF THE ELEMENTS

The periodic table shows elements grouped by periods and groups. Group 14 elements are highlighted in red: Carbon (6C), Silicon (14Si), Germanium (32Ge), Tin (50Sn), and Lead (82Pb). The table includes atomic numbers, symbols, names, and electron configurations for each element. Lanthanoids and Actinoids are shown at the bottom.



Figure 1F.14

# 1F.6 The general properties of the elements

## Group 16 to 18 elements

- Elements at the right of the **p-block**
- **High electron affinities:** they tend to gain electrons to complete shells.
- Except for the metalloids tellurium and polonium, the members of group 16 and 17 are **non-metals**.
- They typically form molecular **compounds** with one another.
- From left to right (**Fig. 1F.15**): oxygen, sulfur, selenium, and tellurium (Group 16). Note the trend from nonmetal to metalloid moving down the group.
- Group 18: Noble gases, completed shells, unreactive (inert gases).

PERIODIC TABLE OF THE ELEMENTS

Group 1 2 ... 13 14 15 16 17 18  
I II III IV V VI VII VIII  
IA IIA IIIA IVA VA VIA VIIA VIIIA  
1 H hydrogen 1.0079 1s<sup>1</sup> 2 He helium 4.00 1s<sup>2</sup>  
3 Li lithium 6.94 2s<sup>2</sup> 4 Be beryllium 9.01 2s<sup>2</sup> 5 B boron 10.81 2s<sup>2</sup>2p<sup>1</sup> 6 C carbon 12.01 2s<sup>2</sup>2p<sup>2</sup> 7 N nitrogen 14.01 2s<sup>2</sup>2p<sup>3</sup> 8 O oxygen 16.00 2s<sup>2</sup>2p<sup>4</sup> 9 F fluorine 18.99 2s<sup>2</sup>2p<sup>5</sup> 10 Ne neon 20.18 2s<sup>2</sup>2p<sup>6</sup>  
11 Na sodium 22.99 3s<sup>2</sup> 12 Mg magnesium 24.31 3s<sup>2</sup> 13 Al aluminum 26.98 3s<sup>2</sup>3p<sup>1</sup> 14 Si silicon 28.09 3s<sup>2</sup>3p<sup>2</sup> 15 P phosphorus 30.97 3s<sup>2</sup>3p<sup>3</sup> 16 S sulfur 32.06 3s<sup>2</sup>3p<sup>4</sup> 17 Cl chlorine 35.45 3s<sup>2</sup>3p<sup>5</sup> 18 Ar argon 39.95 3s<sup>2</sup>3p<sup>6</sup>  
19 K potassium 39.10 4s<sup>2</sup> 20 Ca calcium 40.08 4s<sup>2</sup> 21 Sc scandium 44.96 3d<sup>1</sup>4s<sup>2</sup> 22 Ti titanium 47.87 3d<sup>2</sup>4s<sup>2</sup> 23 V vanadium 50.94 3d<sup>3</sup>4s<sup>2</sup> 24 Cr chromium 51.99 3d<sup>5</sup>4s<sup>1</sup> 25 Mn manganese 54.94 3d<sup>5</sup>4s<sup>2</sup> 26 Fe iron 55.84 3d<sup>6</sup>4s<sup>2</sup> 27 Co cobalt 58.93 3d<sup>7</sup>4s<sup>2</sup> 28 Ni nickel 58.69 3d<sup>8</sup>4s<sup>2</sup> 29 Cu copper 63.55 3d<sup>10</sup>4s<sup>1</sup> 30 Zn zinc 65.38 3d<sup>10</sup>4s<sup>2</sup> 31 Ga gallium 69.72 4s<sup>2</sup>4p<sup>1</sup> 32 Ge germanium 72.64 4s<sup>2</sup>4p<sup>2</sup> 33 As arsenic 74.92 4s<sup>2</sup>4p<sup>3</sup> 34 Se selenium 78.96 4s<sup>2</sup>4p<sup>4</sup> 35 Br bromine 79.90 4s<sup>2</sup>4p<sup>5</sup> 36 Kr krypton 83.80 4s<sup>2</sup>4p<sup>6</sup>  
37 Rb rubidium 85.47 5s<sup>2</sup> 38 Sr strontium 87.62 5s<sup>2</sup> 39 Y yttrium 88.91 4d<sup>1</sup>5s<sup>2</sup> 40 Zr zirconium 91.22 4d<sup>2</sup>5s<sup>2</sup> 41 Nb niobium 92.91 4d<sup>4</sup>5s<sup>1</sup> 42 Mo molybdenum 95.94 4d<sup>5</sup>5s<sup>1</sup> 43 Tc technetium (98) 4d<sup>5</sup>5s<sup>2</sup> 44 Ru ruthenium 101.07 4d<sup>7</sup>5s<sup>1</sup> 45 Rh rhodium 102.90 4d<sup>8</sup>5s<sup>1</sup> 46 Pd palladium 106.42 4d<sup>10</sup> 47 Ag silver 107.87 4d<sup>10</sup>5s<sup>1</sup> 48 Cd cadmium 112.41 4d<sup>10</sup>5s<sup>2</sup> 49 In indium 114.82 5s<sup>2</sup>5p<sup>1</sup> 50 Sn tin 118.71 5s<sup>2</sup>5p<sup>2</sup> 51 Sb antimony 121.76 5s<sup>2</sup>5p<sup>3</sup> 52 Te tellurium 127.60 5s<sup>2</sup>5p<sup>4</sup> 53 I iodine 126.90 5s<sup>2</sup>5p<sup>5</sup> 54 Xe xenon 132.91 5s<sup>2</sup>5p<sup>6</sup>  
55 Cs cesium 132.91 6s<sup>2</sup> 56 Ba barium 137.33 6s<sup>2</sup> 57 La lanthanum 138.91 5d<sup>1</sup>6s<sup>2</sup> 58 Ce cerium 140.12 5d<sup>1</sup>6s<sup>2</sup> 59 Pr praseodymium 140.91 4f<sup>3</sup>6s<sup>2</sup> 60 Nd neodymium 144.24 4f<sup>4</sup>6s<sup>2</sup> 61 Pm promethium (145) 4f<sup>5</sup>6s<sup>2</sup> 62 Sm samarium 150.36 4f<sup>6</sup>6s<sup>2</sup> 63 Eu europium 151.96 4f<sup>7</sup>6s<sup>2</sup> 64 Gd gadolinium 157.25 4f<sup>7</sup>5d<sup>1</sup>6s<sup>2</sup> 65 Tb terbium 158.93 4f<sup>9</sup>6s<sup>2</sup> 66 Dy dysprosium 162.50 4f<sup>10</sup>6s<sup>2</sup> 67 Ho holmium 164.93 4f<sup>11</sup>6s<sup>2</sup> 68 Er erbium 167.26 4f<sup>12</sup>6s<sup>2</sup> 69 Tm thulium 168.93 4f<sup>13</sup>6s<sup>2</sup> 70 Yb ytterbium 173.04 4f<sup>14</sup>6s<sup>2</sup> 71 Lu lutetium 174.97 4f<sup>14</sup>6s<sup>2</sup>  
72 Hf hafnium 178.49 5d<sup>2</sup>6s<sup>2</sup> 73 Ta tantalum 180.95 5d<sup>3</sup>6s<sup>2</sup> 74 W tungsten 183.84 5d<sup>4</sup>6s<sup>2</sup> 75 Re rhenium 186.21 5d<sup>5</sup>6s<sup>2</sup> 76 Os osmium 190.23 5d<sup>6</sup>6s<sup>2</sup> 77 Ir iridium 192.22 5d<sup>7</sup>6s<sup>2</sup> 78 Pt platinum 195.08 5d<sup>9</sup>6s<sup>1</sup> 79 Au gold 196.97 5d<sup>10</sup>6s<sup>1</sup> 80 Hg mercury 200.59 5d<sup>10</sup>6s<sup>2</sup> 81 Tl thallium 204.38 6s<sup>2</sup>6p<sup>1</sup> 82 Pb lead 207.2 6s<sup>2</sup>6p<sup>2</sup> 83 Bi bismuth 208.98 6s<sup>2</sup>6p<sup>3</sup> 84 Po polonium (209) 6s<sup>2</sup>6p<sup>4</sup> 85 At astatine (210) 6s<sup>2</sup>6p<sup>5</sup> 86 Rn radon (222) 6s<sup>2</sup>6p<sup>6</sup>  
87 Fr francium (223) 7s<sup>2</sup> 88 Ra radium (226) 7s<sup>2</sup> 89 Ac actinium (227) 6d<sup>1</sup>7s<sup>2</sup> 90 Th thorium 232.04 6d<sup>2</sup>7s<sup>2</sup> 91 Pa protactinium 231.04 5f<sup>2</sup>6d<sup>1</sup>7s<sup>2</sup> 92 U uranium 238.03 5f<sup>3</sup>6d<sup>1</sup>7s<sup>2</sup> 93 Np neptunium (237) 5f<sup>4</sup>6d<sup>1</sup>7s<sup>2</sup> 94 Pu plutonium (244) 5f<sup>6</sup>7s<sup>2</sup> 95 Am americium (243) 5f<sup>7</sup>7s<sup>2</sup> 96 Cm curium (247) 5f<sup>7</sup>7s<sup>2</sup> 97 Bk berkelium (247) 5f<sup>9</sup>7s<sup>2</sup> 98 Cf californium (251) 5f<sup>10</sup>7s<sup>2</sup> 99 Es einsteinium (252) 5f<sup>11</sup>7s<sup>2</sup> 100 Fm fermium (257) 5f<sup>12</sup>7s<sup>2</sup> 101 Md mendelevium (258) 5f<sup>13</sup>7s<sup>2</sup> 102 No nobelium (259) 5f<sup>14</sup>7s<sup>2</sup> 103 Lr lawrencium (262) 5f<sup>14</sup>7s<sup>2</sup>  
\*The names of the elements 112 and higher have not yet been determined, both 112 and 114 have been confirmed.



Figure 1F.15

# 1F.6 The general properties of the elements

## d-block elements

- All **d-block** elements are metals: d-metals or **transition metals**
- Their properties are transitional between the s- and the p-block elements (except group 12)
- When a d-block element loses electrons to form a cation, it first loses its s-electrons.**
- From left to right (**Fig. 1F.16**). Top row: scandium, titanium, vanadium, chromium, and manganese. Bottom row: iron, cobalt, nickel, copper, and zinc.

PERIODIC TABLE OF THE ELEMENTS

Molar masses (atomic weights) quoted to the number of significant figures given here can be regarded as typical of most naturally occurring samples.

\*The names of the elements 112 and higher have not yet been determined; both 112 and 114 have been confirmed.

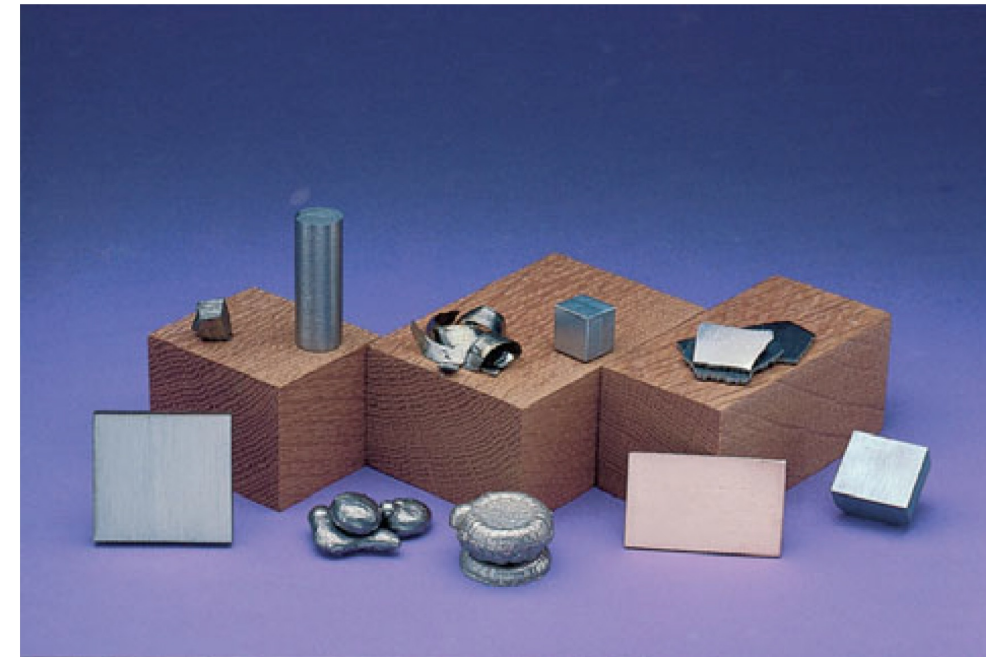


Figure 1F.16

## 1F.6 The general properties of the elements

### d-block elements are (bio-)catalysts

- Availability of d-orbitals and the similarity of atomic radii of the d-block elements are important for many areas:
- They can accelerate reactions as **catalysts** in industry.

*Catalyst: a substance that accelerates a reaction but is not itself consumed*

- Ability to form ions with different charges is important **in nature**

# 1F.6 The general properties of the elements

## f-block elements

- f-block elements of same period **have very similar radii** and chemical properties
- Studied intensely for **superconducting** materials
- Actinoids ("actinides") are radioactive
- None of the elements following plutonium occurs naturally on earth in any significant amount

# 1F.6 The general properties of the elements

## Summary

All elements in the s-block are reactive metals.

The p-block elements tend to gain electrons to complete closed shells; they range from metals through metalloids to nonmetals.

All d-block elements are metals with properties between those of s-block and p-block metals. Many d-block elements are capable of forming cations with several different charges.

## The skills you have mastered are the ability to

- ❑ Account for periodic trends in atomic radii, ionization energies, electron affinities, and electronegativities.
- ❑ Summarize in a general way the properties of the elements in relation to their location in the periodic table.

**Summary:** Many properties of the elements, especially their periodic variation, can be predicted through the periodic table and the concept of effective nuclear charge.