# Periodic Properties Practice Items

1. Upon what basis did Mendeleev construct his periodic table?

   **A.** the electron configuration of atoms as determined by emission spectra
   **B.** the atomic radii of the elements as determined by $\alpha$ particle collision data
   **C.** the variation of chemical properties with atomic weight
   **D.** the periodic nature of electronegativity

2. Why do non-metals have high ionization energies?

   **A.** They are very large atoms and thus have a stronger hold on their electrons.
   **B.** They are on the right side of the periodic table.
   **C.** High effective nuclear charge provides a stronger hold on outer shell electrons.
   **D.** Their outer shell electrons are shielded from the nuclear charge.

3. Which has higher first ionization energy?

   **A.** lithium
   **B.** aluminum
   **C.** boron
   **D.** carbon

4. Which of the following statements is **true**?

   **A.** Fluorine is less electronegative than chlorine.
   **B.** Potassium has greater ionization energy than sodium.
   **C.** Nitrogen has a greater electron affinity than carbon.
   **D.** Carbon has a greater atomic radius than oxygen.

5. Which of the following statements is consistent with the table below?

   **A.** The more electropositive the central atom of an oxide the more basic the oxide.
   **B.** The larger the central atom of the oxide the more acidic the oxide.
   **C.** The oxides of non-metals are acidic.
   **D.** Nonmetal oxides react with acids to produce salts

Acidity or basicity of oxides of *s*- and *p*- block elements when dissolved in water

| | | | | | | |
|---|---|---|---|---|---|---|
| Li | Be | B | C | N | O | F |
| Na | Mg | Al | Si | P | S | Cl |
| K | Ca | Ga | Ge | As | Se | Br |
| Rb | Sr | In | Sn | Sb | Te | I |
| Cs | Ba | Tl | Pb | Bi | Po | At |
| Basic Oxides | | | Amphoteric Oxides | | Acidic Oxides | |

6. Which of the following elements has an electronegativity value of 1.0 on the Pauling scale?

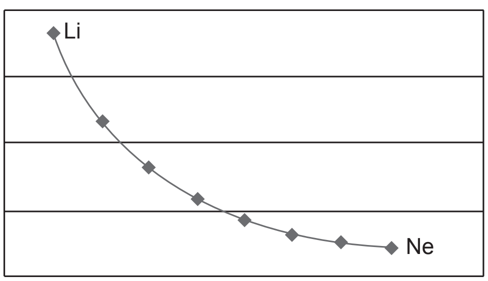
   **A.** calcium
   **B.** carbon
   **C.** hydrogen
   **D.** oxygen

7. All of the following can be measured **except**

A. the ionization energy of krypton.
B. the atomic radius of hydrogen.
C. the Pauling electronegativity of helium.
D. the electron affinity of lithium.

8. The graph below shows the variation across the 2nd period of the periodic table of this property

A. ionization energy
B. electron affinity
C. atomic radius
D. electronegativity




9. Which ion has the smallest radius?

A. $Be^{+}$
B. $Li^{+}$
C. $F^{-}$
D. $O^{2-}$

10. For which group is the difference between the first and second ionization energy the greatest?

A. alkali metals
B. alkaline earth metals
C. halogens
D. noble gases

11. The following elements appear prominently in organic chemistry – carbon, hydrogen, oxygen, nitrogen, and sulfur. Which two are closest to each other in Pauling electronegativity?

A. carbon and sulfur
B. oxygen and nitrogen
C. carbon and hydrogen
D. sulfur and nitrogen

12. Which of the following elements possesses an electron affinity that is actually endothermic?

A. carbon
B. nitrogen
C. oxygen
D. fluorine

13. For which of these salts does the bond between the anion and cation possess the greatest covalent character?

A. $CaBr_2$
B. AgCl
C. KI
D. NaOH

14. In comparing the standard state forms of various elements, which of the following properties is more positive the greater the Pauling electronegativity of the element?

A. standard reduction potential
B. electron affinity
C. brittleness
D. electric conductivity

The following passage pertains to questions 15 - 17.

Linus Pauling first proposed the concept of electronegativity in 1932. The concept explained the additional stabilization of a heteronuclear bond in terms of the contribution of ionic forms to an underlying covalent bond. The difference in electronegativity between atoms A and B is given by:

$$\chi_A - \chi_B = (eV)^{-1/2}\sqrt{E_d(AB) - [E_d(AA) + E_d(BB)]/2}$$

where the dissociation energies, $E_d$, of the A–B, A–A and B–B bonds are expressed in electron volts. Hence, for hydrogen and bromine, the dissociation energies H–Br, 3.79 eV; H–H, 4.52 eV; Br–Br 2.00 *eV* entail a difference in Pauling electronegativity of 0.73 between hydrogen and bromine.

One consequence of the Pauling electronegativity scale is that it provides a semi-empirical formula to determine dissociation energies. This energy estimate can be only used for single, not multiple bonds.

$$E_d(AB) = [E_d(AA) + E_d(BB)]/2 + (\chi_A - \chi_B)^2 eV$$

As only differences in electronegativity are defined, it is necessary to choose an arbitrary reference point in order to construct a scale. Because it forms covalent bonds with a large variety of elements, hydrogen was chosen as the reference, fixed at 2.1 (later revised to 2.2).

**15.** A heteronuclear covalent bond in which there is a strong contribution of ionic forms to the underlying covalent bond is a

**A.** π bond
**B.** σ bond
**C.** polar covalent bond
**D.** coordinate covalent bond

**16.** Which statement below follows from the method of determining bond dissociation energies described in the passage?

**A.** Polar bonds tend to be stronger than nonpolar bonds.
**B.** Bond formation between two atoms with high electronegativity tends to be very exothermic.
**C.** The electronegativity difference between bonded atoms determines whether the bond is a single or double bond.
**D.** Bond dissociation energy increases with ionization energy.

**17.** According to the theory underlying the electronegativity scale, which of the following is closest to the contribution made by the polar character of the bond to the bond dissociation energy of H–Br?

**A.** 0.53 *eV*
**B.** 0.63 *eV*
**C.** 0.73 *eV*
**D.** 2.52 *eV*

## Periodic Properties

Answers and Explanations

**1. C**

The fundamental observation that led to the creation of the periodic table by Mendeleev was that the elements, if arranged according to their atomic weight, exhibit a periodicity of properties.

**2. C**

The nonmetal elements occupy the upper right-hand corner of the period table. Across a period from left to right, effective nuclear charge increases. With each step to the right the nucleus gains a proton, but the shielding of the nucleus by inner shell electrons remains constant. The greater effective nuclear charge pulls the electron cloud closer to the nucleus, strengthening the nuclear attraction to the outer-most electron so the general trend with movement towards the right is that removing an electron will require greater energy.

**3. D**

Moving left to right within a period, the first ionization energy generally increases. Moving downward with a group, ionization energy generally decreases.

**4. D**

Moving left to right within a period, atomic radius *decreases*. Even though the atoms are becoming more massive, the greater effective nuclear charge pulls the electron cloud closer to the nucleus.

Note that choice '**C**' is untrue. Electron affinity of carbon is greater than nitrogen even though the general trend is for electron affinity to increase left to right. The electron affinity of nitrogen is actually slightly endothermic, because nitrogen has a half filled *p* orbital. Three single electrons with parallel spin in the three *p* orbitals is particularly stable. Nitrogen is an important exception both in electron affinity, which is much lower than you would expect and in ionization energy which is higher than you would expect.

**5. C**

Oxides are chemical compounds with one or more oxygen atoms combined with another element. Based on their acid-base characteristics oxides are classified as acidic, basic, amphoteric or neutral. An oxide that combines with water to give an acid is termed as an acidic oxide. As the table accompanying the question shows, the oxides of nonmetals are acidic.

**6. A**

Electronegativity on the Pauling scale runs from 0.79 (cesium) to 3.98 (fluorine). A good rule of thumb for alkali and alkaline earth metals is that their electronegativity is close to 1. For example, sodium is 0.93. Magnesium is 1.3. Calcium is 1.0.

**7. C**

Pauling electronegativity describes the tendency of an atom to pull a shared pair of electrons towards itself. Helium is not listed on the electronegativity scale because it does not form covalent bonds.

**8. C**

Although ionization energy and electron affinity, both exhibit a few exceptions to the general trend, moving left to right within a period, atomic radius is the only trend that *decreases*. Even though the atoms are becoming more massive, the greater effective nuclear charge pulls the electron cloud closer to the nucleus.

**9. B**

Lithium ion has lost its only L shell electron, so it now has an atomic radius nearer to helium than a 2nd period element. All of the other choices possess outer shell electrons in the L shell.

**10. A**

Alkali metals have one outer shell electron. The first ionization energy is low because after losing one electron the cationic species will achieve noble gas configuration. However, if we try to remove a second electron that will require a very large amount of energy because it will disturb the noble gas configuration. That's why first ionization energy of alkali metals is low and second ionisation energy is high.

**11. A**

It really is an important facet of a general sciences knowledge base to know the number values of the electronegativities of the elements that play a prominent role in organic chemistry. Along with valence, electronegativity is the clearest window into the 'personality' of an element. As you move from organic chemistry into biochemistry, having a good sense of the personality of sulfur becomes increasingly important. The electronegativity of sulfur is 2.58, just a hair higher than the electronegativity of carbon at 2.55. While this does mean that carbon-sulfur bonds are *nonpolar*, it also means that sulfur oxidizes carbon. Both of those facts are very important in biochemistry.

**12. B**

Nitrogen is an exception to the general trend of electron affinity increasing from left to right on the periodic table. Nitrogen has a half-filled 2p subshell, so that there is one electron in each orbital. This creates an unusually stable atom because of half-shell stability. In other words the electron affinity of nitrogen is actually slightly endothermic because the additional electron would disturb the stable arrangement of three electrons with parallel spin in the 2p subshell.

**13. B**

For all of the salts among the answer choices except AgCl, the cation species is the ion of an alkali metal or an alkaline earth metal. Alkali metals and alkaline earth metals have an electronegativity of approximately 1 (a bit less in the case of alkali metals) so paired with oxygen or a halogen, the electronegativity difference reflects an ionic bond. Silver, on the other hand, is a much more electronegative transition metal. Its electronegativity on the Pauling scale is 1.93. Its bond to chlorine in AgCl has significant covalent character. Greater covalent character is one of the reasons a salt may be insoluble or sparingly soluble.

**14. A**

In the twentieth century, the methods of problem solving in oxidation-reduction grew out of the fundamental discoveries behind Pauling's electronegativity scale. When a very electronegative element forms a covalent bond, it pulls the electrons in the bond in towards its powerful nucleus. This makes the bond stronger. In electrochemistry, we conceptualize the electron falling towards a positive potential. A powerful oxidizing agent is being reduced.

**15. C**

Linus Pauling proposed an empirical relationship which relates the percent ionic character in a bond to the electronegativity difference. Bonds for which the electronegativity difference is greater than 1.7 are said to have greater than 50% ionic character. Bonds with significant ionic character, though less than 1.7 electronegativity difference, are polar covalent bonds. In polar covalent bonding the pair of electrons is unequally shared between two atoms.

**16. A**

The formula below presents a picture in which the greater the electronegativity difference between elements A and B, the greater the additional bond dissociation energy due to the partial ionic character of the bond, ie. the polar character.

$$E_d(AB) = [E_d(AA) + E_d(BB)]/2 \ + \ (\chi_A - \chi_B)^2 eV$$

The greater the bond dissociation energy, the more energy is required to break the bond. Polar covalent bonds tend to be stronger.

**17. A**

$$E_d(AB) = [E_d(AA) + E_d(BB)]/2 \ + \ (\chi_A - \chi_B)^2 eV$$

In the formula above, the contribution of the polar character of the bond to the bond dissociation is represented by the term: $(\chi_A - \chi_B)^2$eV. As described in the passage, the electronegativity difference of hydrogen and bromine is 0.73. Therefore the contribution of the polar character is $(0.73)^2$ cV = 0.53 eV.