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Atoms, Periodicity & the Octet Rule

Study Guide — Periodic Table & Trends

Foundations of chemistry for beginners: atomic structure, isotopes and ions, periodic table groups/periods, key periodic trends, and the octet rule (including common exceptions).

40 items — Study Guide with Answers

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1 **What determines which element an atom is (e.g., carbon vs nitrogen)?**

A The number of neutrons

B **The number of protons** ✓

C The number of electrons only

D The mass number

E The number of electron shells

▶ **Explanation:** An element is defined by its atomic number, which equals the number of protons. Neutrons can vary (isotopes) and electrons can vary (ions) without changing the element.

2 **The atomic number of an element is equal to the number of:**

A Neutrons

B **Protons** ✓

C Protons + neutrons

D Electrons + neutrons

E Electron shells

▶ **Explanation:** Atomic number = number of protons. Mass number is protons + neutrons.

3 **The mass number of an atom is equal to the number of:**

A Protons only

B Neutrons only

C Electrons only

D **Protons + neutrons** ✓

E Protons + electrons

▶ **Explanation:** Mass number counts the particles in the nucleus: protons + neutrons. Electrons have very small mass and are not included in mass number.

4 **Two atoms are isotopes of the same element if they have the same number of protons but different numbers of:**

A Electrons

B **Neutrons** ✓

C Shells

D Valence electrons

E Protons

▶ **Explanation:** Isotopes differ in neutron number (and therefore mass number) but keep the same proton number, so they remain the same element.

5 **Which statement correctly describes a cation?**

A A cation has gained electrons and is negatively charged

B **A cation has lost electrons and is positively charged** ✓

C A cation has gained protons and is positively charged

D A cation has lost neutrons and is positively charged

E A cation is always neutral

▶ **Explanation:** A cation is a positively charged ion formed when an atom loses one or more electrons (protons do not change in normal chemical reactions).

6 **Chlorine has atomic number 17. How many electrons does a chloride ion (Cl^-) have?**

A 16

B 17

C **18** ✓

D 34

E 35

▶ **Explanation:** Neutral Cl has 17 electrons. Cl^- has gained 1 electron, so it has 18 electrons.

7 **A neutral atom has equal numbers of:**

A Neutrons and electrons

B Protons and neutrons

C **Protons and electrons** ✓

D Shells and protons

E Mass number and atomic number

▶ **Explanation:** Neutral means total positive charge equals total negative charge, so number of protons = number of electrons.

8 **Calcium has atomic number 20. How many electrons does Ca^{2+} have?**

A 20

B 22

C **18** ✓

D 40

E 2

▶ **Explanation:** Neutral Ca has 20 electrons. Ca^{2+} has lost 2 electrons, so it has 18 electrons.

9 **What is the maximum number of electrons in the first energy level (first shell)?**

A 1

B **2** ✓

C 8

D 10

E 18

▶ **Explanation:** The first shell (n = 1) can hold a maximum of 2 electrons.

10 **Valence electrons are the electrons:**

A In the nucleus

B In the innermost shell only

C **In the outermost occupied shell** ✓

D That determine the mass number

E That are always paired

▶ **Explanation:** Valence electrons are in the outermost occupied shell and are mainly responsible for bonding and chemical reactivity.

11 **For main-group elements (groups 1–2 and 13–18), the group number is most useful for predicting:**

A The number of neutrons

B **The number of valence electrons** ✓

C The melting point

D The atomic mass

E The number of electron shells

▶ **Explanation:** In the main groups, elements in the same group share the same number of valence electrons, so they have similar chemical behavior.

12 **The period number (row) of an element in the periodic table is most closely related to the number of:**

A Protons

B Valence electrons

C **Occupied electron shells (energy levels)** ✓

D Neutrons

E Isotopes

▶ **Explanation:** Elements in the same period have the same number of occupied electron shells (energy levels).

13 **Which group contains the alkali metals (very reactive metals like Li, Na, K)?**

A **Group 1** ✓
B Group 2
C Group 17
D Group 18
E Group 14

▶ **Explanation:** Alkali metals are in group 1 and typically form +1 ions by losing one valence electron.

14 **Which ion is most commonly formed by magnesium (Mg), a group 2 metal?**

A Mg−
B Mg^{+}
C **Mg^{2+}** ✓
D Mg^{3+}
E Mg2−

▶ **Explanation:** Group 2 metals have 2 valence electrons and commonly lose both to form 2+ cations.

15 **Noble gases (group 18) are generally unreactive mainly because they:**

A Have the largest atoms

B **Have a full valence shell** ✓

C Have no protons

D Have no electrons

E Always form 3− ions

▶ **Explanation:** A filled outer shell is especially stable, so noble gases have little tendency to gain, lose, or share electrons (with rare exceptions for heavier noble gases).

16 **Metalloids (like silicon) are typically found:**

A Only in group 1

B Only in group 18

C **Along the "staircase" boundary between metals and nonmetals** ✓

D Only in the center transition metal block

E Only in period 1

▶ **Explanation:** Metalloids lie near the zig-zag staircase dividing metals (left) and nonmetals (right) and often show mixed properties.

17 **Which general trend is correct for atomic radius (size) in the periodic table?**

A Increases left to right across a period and decreases down a group

B **Decreases left to right across a period and increases down a group** ✓

C Increases left to right across a period and increases down a group

D Decreases left to right across a period and decreases down a group

E Has no pattern

▶ **Explanation:** Across a period, increasing nuclear charge pulls electrons closer (smaller radius). Down a group, new shells are added, increasing radius.

18 Which general trend is correct for first ionization energy (energy required to remove one electron)?

A Decreases left to right and increases down a group

B **Increases left to right and decreases down a group** ✓

C Increases left to right and increases down a group

D Decreases left to right and decreases down a group

E Has no relationship to periodic position

▶ **Explanation:** Across a period, electrons are held more tightly (higher ionization energy). Down a group, outer electrons are farther from the nucleus and more shielded (lower ionization energy).

19 Which element is the most electronegative (strongest tendency to attract bonding electrons) in the periodic table?

A Sodium (Na)

B Oxygen (O)

C **Fluorine (F)** ✓

D Chlorine (Cl)

E Neon (Ne)

▶ **Explanation:** Electronegativity generally increases up and to the right, and fluorine is the highest because it is small and has a strong effective nuclear charge on valence electrons.

20 **Which element is most likely to have the greatest tendency to gain an electron (highest electron affinity, conceptually)?**

A Sodium (Na)

B Magnesium (Mg)

C Argon (Ar)

D **Chlorine (Cl)** ✓

E Neon (Ne)

▶ **Explanation:** Halogens (group 17) are one electron away from a full valence shell, so they strongly tend to gain an electron. Noble gases (Ar, Ne) already have full shells and do not tend to gain electrons.

21 **Which statement best describes the trend in metallic character?**

A Metallic character increases up a group and increases left to right

B Metallic character decreases down a group and increases left to right

C **Metallic character increases down a group and increases right to left across a period** ✓

D Metallic character is highest at the top-right of the table

E Metallic character does not relate to electron behavior

▶ **Explanation:** Metals tend to lose electrons. Down a group, electrons are held less tightly (more metallic). Moving left across a period, elements are more likely to lose electrons (more metallic).

22 **Which trend is correct for the reactivity of alkali metals (group 1) as you go DOWN the group?**

A Reactivity decreases because they hold electrons more tightly

B **Reactivity increases because it becomes easier to lose the outer electron** ✓

C Reactivity stays the same because they all have 1 valence electron

D Reactivity is random

E Reactivity increases because electronegativity increases down the group

▶ **Explanation:** Down group 1, the outer electron is farther from the nucleus and more shielded, so it is lost more easily—making alkali metals more reactive.

23 **Which trend is correct for the reactivity of halogens (group 17) as you go DOWN the group?**

A Reactivity increases because they gain electrons more easily

B **Reactivity decreases because their attraction for an added electron becomes weaker** ✓

C Reactivity is highest at the bottom because atomic radius is largest

D Reactivity stays constant because they all form -1 ions

E Reactivity decreases because they lose electrons more easily

▶ **Explanation:** Halogens react by gaining an electron. Down the group, atoms are larger and less electronegative, so gaining an electron becomes harder—reactivity decreases.

24 **The octet rule is the idea that many atoms tend to:**

A Have exactly 8 protons

B **Gain, lose, or share electrons to achieve 8 electrons in their valence shell** ✓

C Always form 8 bonds

D Have exactly 8 neutrons

E Always become noble gases

▶ **Explanation:** Many main-group atoms are most stable with 8 valence electrons (like noble gases), so they bond to reach that configuration—though there are important exceptions.

25 **Why is hydrogen often an exception to the octet rule?**

A Hydrogen prefers 8 electrons but cannot hold any electrons

B **Hydrogen's first shell is full with 2 electrons (duet rule)** ✓

C Hydrogen has no nucleus

D Hydrogen always forms ionic bonds only

E Hydrogen usually wants 18 electrons

▶ **Explanation:** The first energy level can hold only 2 electrons. Hydrogen becomes stable by achieving a duet (2 valence electrons), not an octet.

26 **Which molecule can be drawn with ALL atoms obeying the octet rule (basic Lewis structure idea)?**

A BF_3

B $BeCl_2$

C **CO_2** ✓

D NO

E SF_6

▶ **Explanation:** In CO2, carbon forms two double bonds so carbon and both oxygens can reach octets. BF3 and BeCl2 have incomplete octets on the central atom, NO has an odd electron, and SF6 has an expanded octet on sulfur.

27 **In the molecule BF3, the boron atom has how many electrons around it in a typical Lewis structure?**

A 4

B **6** ✓

C 8

D 10

E 12

▶ **Explanation:** Boron forms three single bonds to fluorine, giving it 6 electrons around it (an incomplete octet). This is a classic octet-rule exception for boron compounds.

28 **Which molecule is a classic example of an "expanded octet" on the central atom?**

A CH4

B NH3

C H2O

D **SF6** ✓

E BF3

▶ **Explanation:** In SF6, sulfur (period 3) is shown with 12 electrons around it in a Lewis structure (expanded octet). Second-period elements like C, N, O, F do not show expanded octets in basic models.

29 **Which element is NOT able to show an expanded octet in a basic Lewis structure model because it is in period 2?**

A Phosphorus (P)

B Sulfur (S)

C Chlorine (Cl)

D **Nitrogen (N)** ✓

E Xenon (Xe)

▶ **Explanation:** Second-period elements (like C, N, O, F) are typically limited to an octet in basic Lewis models. Expanded octets are commonly discussed for period 3 and beyond (P, S, Cl, Xe).

30 **Which molecule has an odd number of total valence electrons and therefore cannot satisfy the octet rule for all atoms at once?**

A CO_2

B CH_4

C NH_3

D **NO** ✓

E H_2O

▶ **Explanation:** NO has 11 total valence electrons (N has 5, O has 6). Odd-electron species (radicals) cannot give every atom a complete octet simultaneously.

31 **Which description best matches an ionic bond?**

A Sharing electrons equally between two nonmetals

B Sharing electrons unequally between two nonmetals

✓ **C** **Electrostatic attraction between oppositely charged ions formed by electron transfer**

D Attraction between neutrons and protons in the nucleus

E A bond formed by overlapping electron shells only in noble gases

▶ **Explanation:** Ionic bonding involves transfer of electrons (typically metal → nonmetal) creating ions that attract each other due to opposite charges.

32 **Magnesium forms Mg^{2+} and chlorine forms Cl—. What is the correct formula for the compound formed from these ions?**

A MgCl

B **MgCl2** ✓

C Mg2Cl

D Mg2Cl2

E MgCl3

▶ **Explanation:** Charge must balance to zero. One Mg^{2+} requires two Cl— ions to balance: MgCl2.

33 **Which pair is most likely to form a covalent bond (basic rule)?**

A Na and Cl

B Mg and O

C **C and O** ✓

D K and Br

E Ca and F

▶ **Explanation:** Covalent bonds commonly form between two nonmetals by sharing electrons (e.g., C and O). Metal + nonmetal combinations often form ionic compounds.

34 **Aluminum commonly forms Al^{3+} and oxygen commonly forms O2−. What is the simplest neutral formula of aluminum oxide?**

A AlO

B AlO2

C **Al2O3** ✓

D Al3O2

E Al6O6

▶ **Explanation:** To balance charges: $2 \times (+3) = +6$ and $3 \times (-2) = -6$, so Al2O3 is neutral. (Al6O6 reduces to Al2O3.)

35 **Which element would you expect to have the highest first ionization energy?**

A Sodium (Na)

B Magnesium (Mg)

C Neon (Ne)

D **Helium (He)** ✓

E Potassium (K)

▶ **Explanation:** Ionization energy is highest for small atoms with a strong effective nuclear charge; helium is very small and holds its electrons extremely tightly.

36 **Which is generally larger: a sodium atom (Na) or a sodium ion (Na^+)?**

A Na^+ is larger because it has a positive charge

B Na is larger because Na^+ has lost an electron shell (and electrons are pulled in more tightly) ✓

C They are the same size because the number of protons is unchanged

D Na^+ is larger because it has fewer electrons

E You cannot compare atom and ion sizes

▶ **Explanation:** Cations are smaller than their neutral atoms because they have fewer electrons (often losing an outer shell) and the remaining electrons feel a stronger pull from the nucleus.

37 **Which is generally larger: a chlorine atom (Cl) or a chloride ion (Cl^-)?**

A Cl is larger because atoms are always bigger than ions

B Cl^- is larger because adding an electron increases electron–electron repulsion and expands the electron cloud ✓

C They are equal in size because the nucleus is unchanged

D Cl is larger because Cl^- has more protons

E Cl^- is smaller because negative charge pulls electrons inward

▶ **Explanation:** Anions are larger than their neutral atoms because adding electrons increases repulsion and reduces the effective pull per electron, expanding the radius.

38 **$O2^-$, F^-, Ne, Na^+, and Mg^{2+} are all isoelectronic (same number of electrons). Which has the smallest radius?**

A $O2^-$

B F^-

C Ne

D Na^+

E **Mg^{2+}** ✓

▶ **Explanation:** In an isoelectronic series, the species with the most protons pulls the same electron cloud in most strongly, making it smallest. Mg^{2+} has the highest nuclear charge here (12 protons).

39 **Atomic radius generally decreases from left to right across a period mainly because:**

A Atoms lose neutrons across a period

B The number of shells decreases across a period

C **The effective nuclear charge increases, pulling electrons closer while shielding stays similar** ✓

D Atoms gain many new electron shells across a period

E Electrons become heavier across a period

▶ **Explanation:** Across a period, proton number increases, but electrons are added to the same main shell, so shielding doesn't rise much. The stronger pull (higher effective nuclear charge) draws electrons closer, shrinking the atom.

40 **Helium is in group 18 but has only 2 valence electrons. Why is helium still very unreactive?**

A Because helium has no valence electrons

B **Because helium's first shell is full with 2 electrons (a stable duet)** ✓

C Because helium always forms four bonds to reach an octet

D Because helium is a metal and metals are unreactive

E Because helium has the largest atomic radius

▶ **Explanation:** The first energy level can hold only 2 electrons, and helium already has those 2. So it already has a complete outer shell and does not need to gain/lose/share electrons.