

CHAPTER 6 STUDY GUIDE FOR CONTENT MASTERY

The Periodic Table and Periodic Law

Section 6.1 Development of the Modern Periodic Table

In your textbook, read about the history of the periodic table's development.

Use each of the terms below just once to complete the passage.

octaves	atomic mass	atomic number	nine
elements	properties	Henry Moseley	eight
protons	periodic law	Dmitri Mendeleev	accepted

The table below was developed by John Newlands and is based on a relationship called the law of (1) octaves. According to this law, the properties of the elements repeated every (2) eight elements. Thus, for example, element two and element (3) nine have similar properties. The law of octaves did not work for all the known elements and was not generally (4) accepted.

1	2	3	4	5	6	7
H	Li	G	Bo	C	N	O
8	9	10	11	12	13	14
F	Na	Mg	Al	Si	P	S

The first periodic table is mostly credited to (5) Mendeleev. In his table, the elements were arranged according to increasing (6) mass. One important result of this table was that the existence and properties of undiscovered (7) elements could be predicted.

The element in the modern periodic table are arranged according to increasing (8) atomic number, as a result of the work of (9) Moseley. This arrangement is based on number of (10) protons in the nucleus of an atom of the element. The modern form of the periodic table results in the (11) periodic law, which states that when elements are arranged according to increasing atomic number, there is a periodic repetition of their chemical and physical (12) properties.

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Section 6.1 *continued*

In your textbook, read about the modern periodic table.

Use the information in the box on the left taken from the periodic table to complete the table on the right.

7
N
Nitrogen
14.007
[He]2s ² 2p ³

Atomic Mass	13. 14.007
Atomic Number	14. 7
Electron Configuration	15. [He]2s ² 2p ³
Chemical Name	16. Nitrogen
Chemical Symbol	17. N

For each item in Column A, write the letter of the matching item in Column B.

- | Column A | Column B |
|--|----------------------------|
| <u>B</u> 18. A column on the periodic table | a. metals |
| <u>C</u> 19. A row on the periodic table | b. group |
| <u>D</u> 20. Group A elements | c. period |
| <u>A</u> 21. Elements that are shiny and conduct electricity | d. representative elements |
| <u>E</u> 22. Group B elements | e. transition elements |

In the space at the left, write *true* if the statement is true; if the statement is false, change the italicized word or phrase to make it true.

- false, three 23. There are *two* main classifications of elements.
- false, metals 24. More than three-fourths of the elements in the periodic table are *nonmetals*.
- true 25. Group 1A elements (except for hydrogen) are known as the *alkali metals*.
- false, 2A 26. *Group 3A* elements are the alkaline earth metals.
- true 27. Group 7A elements are highly reactive nonmetals known as *halogens*.
- false, noble gases 28. Group 8A elements are very unreactive elements known as *transition metals*.
- false, nonmetals 29. Metalloids have properties of both metals and *inner transition metals*.

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Section 6.2 Classification of the Elements

In your textbook, read about organizing the elements by electron configuration.

Use the periodic table on pages 156–157 in your textbook to match each element in Column A with the element in Column B that has the most similar chemical properties.

Column A	Column B
<u>H</u> 1. arsenic (As) 5A	a. boron (B)
<u>F</u> 2. bromine (Br) 7A	b. cesium (Cs)
<u>N</u> 3. cadmium (Cd) 2B	c. chromium (Cr)
<u>A</u> 4. gallium (Ga) 3A	d. cobalt (Co)
<u>K</u> 5. germanium (Ge) 4A	e. hafnium (Hf)
<u>D</u> 6. iridium (Ir) 9	f. iodine (I)
<u>L</u> 7. magnesium (Mg) 2A	g. iron (Fe)
<u>O</u> 8. neon (Ne) 8A	h. nitrogen (N)
<u>I</u> 9. nickel (Ni) 10	i. platinum (Pt)
<u>G</u> 10. osmium (Os) 8	j. scandium (Sc)
<u>B</u> 11. sodium (Na) 1A	k. silicon (Si)
<u>M</u> 12. tellurium (Te) 6A	l. strontium (Sr)
<u>C</u> 13. tungsten (W) 6B	m. sulfur (S)
<u>J</u> 14. yttrium (Y) 3B	n. zinc (Z)
<u>E</u> 15. zirconium (Zr) 4B	o. xenon (Xe)

Answer the following questions.

16. Why do sodium and potassium, which belong to the same group in the periodic table, have similar chemical properties?

They have the same number of valence electrons

17. How is the energy level of an element's valence electrons related to its period on the periodic table? Give an example.

The number of the period is the same as the number of the highest energy level for a given element found there
Na: [Ne]3s¹ → in Period 3 energy level 3 is where valence e⁻ is.

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Section 6.2 *continued*

In your textbook, read about s-, p-, d-, and f-block elements.

Use the periodic table on pages 156–157 in your textbook and the periodic table below to answer the following questions.

s block																s ²																									
s ¹																2																									
1	H															He																									
s ²																																									
3	Li	4	Be																																						
		p block																																							
		p ¹ p ² p ³ p ⁴ p ⁵ p ⁶																																							
11	Na	12	Mg	d block																																					
		d ¹ d ² d ³ d ⁴ d ⁵ d ⁶ d ⁷ d ⁸ d ⁹ d ¹⁰																																							
19	K	20	Ca	21	Sc	22	Ti	23	V	24	Cr	25	Mn	26	Fe	27	Co	28	Ni	29	Cu	30	Zn	31	Ga	32	Ge	33	As	34	Se	35	Br	36	Kr						
37	Rb	38	Sr	39	Y	40	Zr	41	Nb	42	Mo	43	Tc	44	Ru	45	Rh	46	Pd	47	Ag	48	Cd	49	In	50	Sn	51	Sb	52	Te	53	I	54	Xe						
55	Cs	56	Ba	71	Lu	72	Hf	73	Ta	74	W	75	Re	76	Os	77	Ir	78	Pt	79	Au	80	Hg	81	Tl	82	Pb	83	Bi	84	Po	85	At	86	Rn						
87	Fr	88	Ra	103	Lr	104	Rf	105	Db	106	Sg	107	Bh	108	Hs	109	Mt	110	Uun	111	Uuv	112	Uub																		
		f block																																							
		f ¹ f ² f ³ f ⁴ f ⁵ f ⁶ f ⁷ f ⁸ f ⁹ f ¹⁰ f ¹¹ f ¹² f ¹³ f ¹⁴																																							
57	La	58	Ce	59	Pr	60	Nd	61	Pm	62	Sm	63	Eu	64	Gd	65	Tb	66	Dy	67	Ho	68	Er	69	Tm	70	Yb														
89	Ac	90	Th	91	Pa	92	U	93	Np	94	Pu	95	Am	96	Cm	97	Bk	98	Cf	99	Es	100	Fm	101	Md	102	No														

18. Into how many blocks is the periodic table divided? 4: s, p, d, f
19. What groups of elements does the s-block contain? alkali + alkaline earth
20. Why does the s-block portion of the periodic table span two groups?
s orbital holds 2 valence electrons
21. What groups of elements does the p-block contain? boron family → noble gases
22. Why are members of group 8A virtually unreactive?
They have a complete octet for their valence shell so they are stable and don't want to gain/lose electrons by reacting.
23. How many d-block elements are there? 40 (10 across, 4 periods)
24. What groups of elements does the d-block contain? transition metals
25. Why does the f-block portion of the periodic table span 14 groups?
The f orbital can hold 14 electrons
26. What is the electron configuration of the element in period 3, group 6A? [Ne]3s²3p⁴

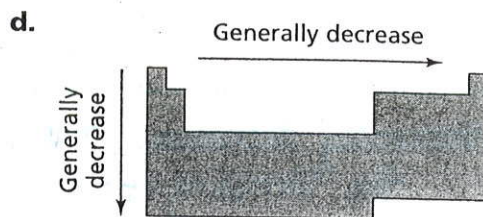
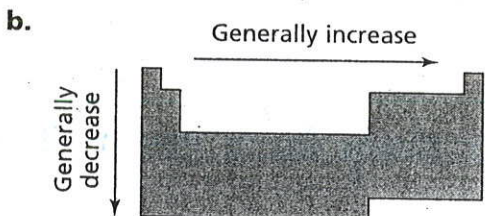
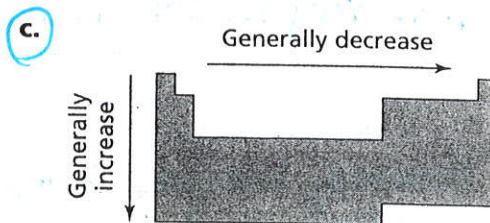
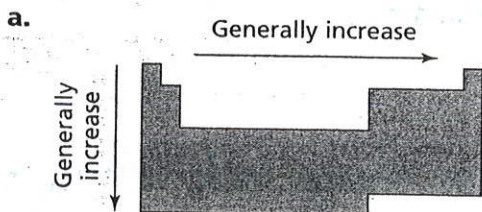
Section 6.3 Periodic Trends

In your textbook, read about atomic radius and ionic radius.

Circle the letter of the choice that best completes the statement or answers the question.

1. Atomic radii cannot be measured directly because the electron cloud surrounding the nucleus does not have a clearly defined _____.
- a. charge. b. mass. **c. outer edge.** d. probability.

2. Which diagram best represents the group and period trends in atomic radii in the periodic table?

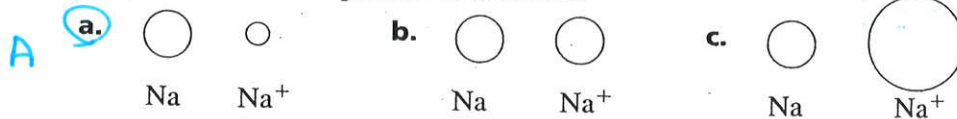


3. The general trend in the radius of an atom moving down a group is partially accounted for by the _____.
- a. decrease in the mass of the nucleus. **c. increase in the charge of the nucleus.**
d. shielding of the outer electrons by inner electrons.

4. A(n) _____ is an atom, or bonded group of atoms, that has a positive or negative charge.
- a. halogen **b. ion** c. isotope d. molecule

5. An atom becomes negatively charged by _____.
- a. gaining an electron.** b. gaining a proton. c. losing an electron. d. losing a neutron.

6. Which diagram best represents the relationship between the diameter of a sodium atom and the diameter of a positive sodium ion?



CHAPTER

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Section 6.3 *continued*

In your textbook, read about ionization energy and electronegativity.

Answer the following questions.

7. What is ionization energy?

Energy required to remove the outermost electron

8. Explain why an atom with a high ionization-energy value is not likely to form a positive ion.

It is very hard to remove the electron

9. What is the period trend in the first ionization energies? Why?

Generally increases → higher nuclear charge, holds electrons tighter
→ there are some exceptions though

10. What is the group trend in the first ionization energies? Why?

Generally decreases → lower nuclear charge due to more PEL's and the shielding effect

11. State the octet rule.

Eight electrons filling the valence shell of an atom

12. What does the electronegativity of an element indicate?

The ability of an atom to attract an electron to itself

13. What are the period and group trends in electronegativities?

period → increases
group → decreases