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Introductory Concepts

1. Chemistry is best defined as the study of matter and the changes it undergoes.
2. Chemistry is called the central science because it connects the other physical sciences conceptually.
3. The factor being tested in an experiment is called the independent variable.
4. A statement that describes how nature behaves but does not explain why it behaves that way is called a(n): natural law.
5. An experimental control does what? responds in a predictable way and provides a comparison (alternatively experimental control is defined as conditions limiting the variables that impact the outcomes of an experiment).

SI System						
Measure	Length	Time	Mass	Current	Amount	Temperature
Base Unit	meter	second	kilogram	ampere	mole	kelvin

Prefix	kilo-	hecto-	deka-	deci-	centi-	milli-
Value	1000	100	10	1/10	1/100	1/1000

6. Which instrument used to measure liquid volume precisely in the lab? graduated cylinder (or pipet or buret).
7. How many joules are in 1kJ? 1000.
8. How many millimeters are in 1m? 1000.
9. Density = mass/volume.
10. Matter is anything that occupies space and has mass.
11. The first step in the scientific method is making an observation that inspires a question.
12. Compare accuracy and precision of a measurement. Accuracy relates to the comparability of a measure to a known standard. Precision relates to the ability to replicate the measure and is favored by smaller graduations on an instrument.
13. Why is it necessary to perform multiple trials of a scientific experiment? To confirm that no uncontrolled variables have affected the outcome and to increase the statistical probability that the outcome is valid.
14. What type of instrument is best for measuring mass, volume, and length? Mass – balance, volume – graduated cylinder, buret or pipet, length - ruler
15. How many significant figures are in the following measurements?
    a. 0.000653 g — 3
    b. 1200. m — 4
    c. 8.50×10^{-9} m — 3
    d. 0.00200 mL — 3
    e. 24 000 mL — 2
    f. 0.025 060 s — 5

Matter

16. A blend of two or more pure substances is called a(n) mixture.
17. liquid is the state of matter characterized by definite volume, but lack of definite shape.
18. A substance that contains two or more elements chemically combined in a fixed proportion is called a(n) compound.
19. Elements and compounds are the two types of pure substances.
20. A substance that cannot be separated into simpler substances by a chemical change is called a(n) element.
21. The temperature at which all motion stops in a sample of matter is referred to as absolute zero.
22. A combination of oil and vinegar is a(n) heterogeneous mixture.
23. Salt dissolved in water is an example of a(n) homogeneous mixture.
24. A(n) mixture is a combination of two or more substances in which each substance retains its properties.
25. The rusting of a nail is an example of a(n) chemical change.
26. The fact that water boils at 100°C is an example of a(n) physical (or intensive) property.
27. A substance that is yellow at room temperature is heated over a flame and turns a bright orange. Upon cooling the substance returns to its original yellow color. This type of change is a(n) physical change.
28. A material that contains only one type of atom is referred to as a(n) element.
29. The compound $C_6H_{12}O_6$ has a total of 24 atoms.

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30. Liquid mixtures may be separated by using distillation which takes advantage of difference in the boiling points of mixed substances.
31. When dissolved solids are separated from a solution by boiling the solution away it is called ___crystallization .

Atomic Structure

32. The 3 primary parts of the atom are the ___protons, neutrons & electrons___.
33. The negative region of the atom is the ____electron cloud__.
34. The positive region of the atom is the ____nucleus___.
35. Describe the gold foil experiment. What was the important discovery in that experiment? __Ernest Rutherford bombarded a sheet of gold foil with alpha particles testing the Plum-pudding model. Some alpha particles were deflected, indicating a dense core (which he called the nucleus). This led to the planetary model of the atom.
36. The _nucleus (protons and neutrons)_ of the atom makes up the mass of the atom.
37. The _electrons cloud_ of the atom makes up the volume of the atom.
38. What is atomic number? __The number of protons in the nucleus
39. What is atomic mass? What two subatomic particles make up the majority of an atom's mass? __Atomic mass is the weighted average of all isotopes of a particular element. Since protons and neutrons make up almost all of the atom's mass, atoms of an element with different numbers of neutrons will have different masses, therefore the average is required.__
40. A(n) ____ion__ is an atom that has gained or lost electrons.
41. ____Isotopes__ are atoms of the same element that have different numbers of neutrons.
42. Zn^{2+} has _30_ protons, _35_ neutrons and _28_ electrons.*
43. O^{2-} has _8_ protons, _8_ neutrons and _10_ electrons.*
44. Au has _79_ protons, _118_ neutrons and _79_ electrons.*

    * Really depends on the isotope. These are based on average atomic mass.
45. Determine the number of protons, neutrons, and electrons in each atom or ion:

    a. $^{91}_{40}Zr$ Protons: __40__ Neutrons: __51__ Electrons: __40__

    b. $^{122}_{51}Sb$ Protons: __51__ Neutrons: __71__ Electrons: __51__

    c. $^{200}_{80}Hg^{2+}$ Protons: __80__ Neutrons: __120__ Electrons: __78__

    d. $^{31}_{15}P^{3-}$ Protons: __15__ Neutrons: __16__ Electrons: __18__

46. The force that holds the nucleus of an atom together is called ___strong nuclear force__.
47. The force that causes the nucleus to be unstable is called __electromagnetic repulsion between positive protons___.
48. List the two types of radiation that are particles. ____alpha____ & __beta_
49. The smallest particle of radiation is the __beta__.
50. List the three types of radiation in order of ability to penetrate. (Least to Most)

    Least: ___alpha____ ____beta_ Most: ___gamma___
51. Gamma radiation can best be described as a ___high energy electromagnetic waves___.
52. When the nucleus of an atom is split, it is called ____fission____
53. Write the formula for an isotope that has 92 protons, 143 neutrons, and 90 electrons

    $^{235}_{92}U^{2+}$
54. If element X has a half-life of 20 years and you have 2g now, how much would have been present 140 years ago?

    20ya = 4g, 40ya = 8g, 60ya = 16g, 80ya = 32g, 100ya = 64g, 120ya = 128g, 140ya = 256g
55. If a 1.0g sample of ancient wood has $1/8^{th}$ the radioactivity of a 1.0g modern sample of the same wood, how old is the ancient wood. Half-life of carbon 14 = 5730 years.

    1HL = ½, 2HL = ¼, 3HL = ⅛, thus 3(5730) = 17190yrs.
56. An __electron orbital_ specifies a volume of space where an electron is likely to be found 90% of the time.

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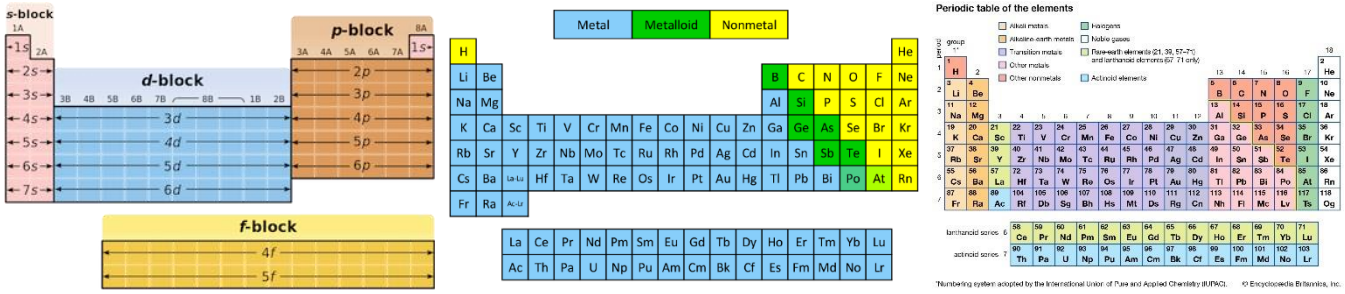
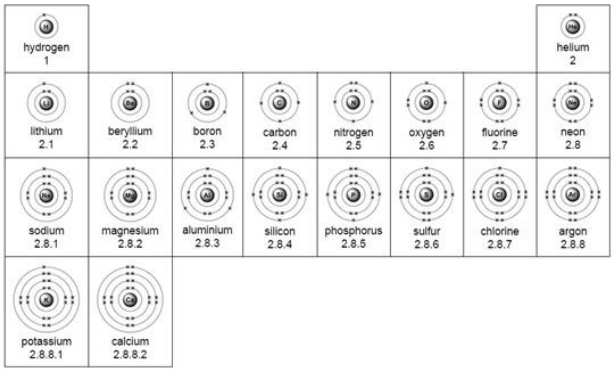
57. Describe the energy level with the principal quantum number $n = 1$. It's the one closest to the nucleus, the one with the lowest energy and can only hold two electrons because it is so small.
58. Why can't electrons exist between energy levels in an atom? The energy of electrons is quantized and limited to certain values that correspond to wave functions of particular energies.
59. Compare the wavelengths and frequencies of gamma rays, visible light and radio waves. Which has the most the energy? Gamma is high frequency (energy), low wavelength. Radio is low frequency, high wavelength. Visible light is somewhere between these two extremes.
60. What happens to electrons in an atom when they gain energy? What happens to electrons in an atom when they lose energy? Electrons move to excited states (farther away from the nucleus) when they gain energy and move to lower energy levels when they lose energy. The energy is given off as photons of electromagnetic energy (light, perhaps).
61. Identify the s,p,d and f blocks of the periodic table.
62. Where are the metals, nonmetals, and semimetals on the periodic table?
63. Identify the groups for alkali metals, alkaline earth metals, halogens, and transition metals?



64. Why do elements in the same group (family) have similar properties? They all have the same number of valence electrons.
65. Write electron configurations for the following box orbital diagrams.

1s	2s	2p(x)	2p(y)	2p(z)	3s	3p(x)	3p(y)	3p(z)	Configuration
↑↓ / ↑↓	↑↓ / ↑↓	↑↓ / ↑↓	↑↓ / ↑↓	↑↓ / ↑↓	↑↓ / ↑↓				$1s^2 2s^2 p^6 3s^2$
						↑	↑	↑	$1s^2 2s^2 p^6 3s^2 p^3$

66. Draw a Bohr model for the first three atoms of the alkali metals, alkaline earth metals and boron families. Alkali metals form ions with a 1+ charge, alkaline-earth metals form ions with 2+ charge and boron family ions have 3+ charge. What does the charge have to do with the models?



The charge is created by the emptying of the outermost (valence) shell.

67. Why do halogens all have a charge of 1- and oxygen family elements have a charge of 2-? Halogens must gain 1 electron to achieve a full valence shell. Oxygen family elements must gain two to achieve a full valence shell.

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68. Why can't you necessarily use an electron configuration to identify an element? Many particles may share an electron configuration as ions (isoelectronic), particularly if it's the configuration of a noble gas.
69. Why do electrons initially spread out into separate orbitals when filling in orbitals of equal energy? Electrons repel one another and cause maximum spreading in space to become stable.
70. State the definition and trend for ionization energy. Ionization energy is the energy required to remove the highest energy electron from an atom. Ionization energy increases with atomic number across a period (due to higher effective nuclear charge) and decreases with radius down a family (due to increased distance between the nucleus and the electrons). Successive ionization energies are the energies for additional electrons to be removed. It always increases as electrons are removed because the remaining electrons will be held tighter by the nucleus.
71. State the definition and trend for electronegativity. Electronegativity is the attraction of an atom to electrons in a chemical bond. Electronegativity increases with atomic number across a period (due to increasing effective nuclear charge) and decreases down a family as radius increases (due to additional layers of electrons and greater distance between the nucleus and the attracted electrons).
72. State the definition and trend for the size of atoms. Atomic radius is the size of an atom and it decreases moving left to right across a period due to increasing effective nuclear charge and increases down a family as additional levels of electrons are added to the atom.
73. State the trends for the relative sizes of ions to atoms. Atoms get larger when they gain electrons to form anions (this applies mainly to non-metals) and get smaller when they lose electrons (this applies mainly to metals and is pronounced when an energy level is emptied).
74. Which family of elements has the highest ionization energies? noble gases Why? They have the greatest effective nuclear charge of any atom in a period.
75. Which family of elements has the lowest ionization energies? alkali metals Why? They have the lowest effective nuclear charge of any atom in a given period.
76. Complete the following nuclear decay reactions:
    a. $^{218}_{88}\text{Ra} \rightarrow {}^{218}_{89}\text{Ac} + {}^{0}_{-1}\text{e}$
    b. $^{238}_{92}\text{U} \rightarrow {}^{4}_{2}\text{He} + {}^{234}_{90}\text{Th}$
    c. $^{14}_{6}\text{C} \rightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e}$
    d. $^{218}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{214}_{82}\text{Pb}$
77. What is the strong nuclear force? One of the fundamental forces of the universe that holds nucleons together in the nucleus. Both protons and neutrons exhibit this short range force.
78. What is the equation that allows us to calculate the energy released in nuclear reactions? $E=mc^2$
79. What are the three most common forms of radioactive decay? Describe each type of decay. How does the nucleus change with each type of decay?
    a. Alpha – a helium nucleus with relatively low energy and little penetrative ability. The nucleus gets lighter by 4 *u* and the nuclear charge drops by 2.
    b. Beta – a high speed electrons with slightly more penetrative ability than alpha particles. This particles is emitted when a neutron turns into a proton causing the charge number of the nucleus to increase by 1.
    c. Gamma – a high energy burst of electromagnetic radiation that is emitted during various nuclear processes. Since gamma radiation has no mass and no charge, the main change in the nucleus resulting from gamma emission is an increase in stability.
80. How does the energy release in a nuclear reaction compare to the energy release in a chemical reaction? Energy release in nuclear reactions is significantly greater then in chemical reactions.

Bonding

81. The first valence shell holds 2 electrons.
82. The second valence shell holds 8 electrons.
83. Which family of elements has the lowest ionization energies? Alkali metals

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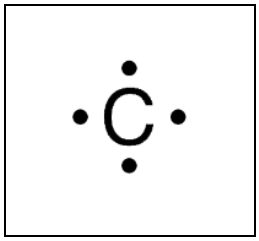
84. Non-metals have high / low ionization energies and high / low electron affinities.

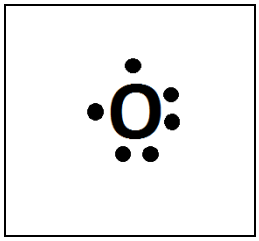
85. Metals have high / low ionization energies and high / low electron affinities.

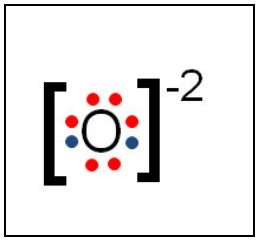
86. What are the general rules that apply to determining electronegativity of elements using the periodic table? The closer to fluorine an atom is, the higher its electronegativity. Higher period – higher EN and higher group number = higher EN.

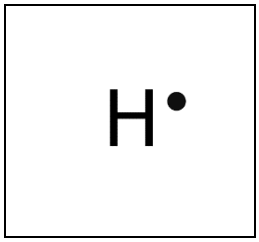
87. Up to how many electrons are drawn on each side of the atomic symbol in a Lewis structure? 2

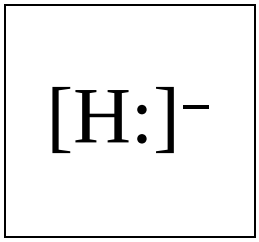
88. Draw a Lewis structure for the following atoms or ions.

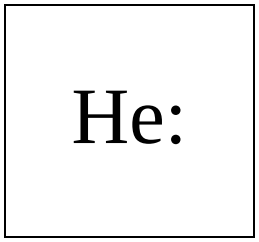
a. C 

b. O 

c. O^{2-} 

d. H 

e. H^- 

f. He 

89. How many electrons do the following atoms have available for bonding*:

a. I		c. Ca	2	e. Sb	3	g. S	2
b. H	1	d. Al	3	f. Rb	1	h. C	4

* This is unpaired electrons only. In reality they can use all of their valence electrons from bonding.

90. How many covalent bonds does C form? 4

91. How many valence electrons does:
    a. the halogen family have? 7
    b. the boron family? 3
    c. the oxygen family? 6
    d. the carbon family? 4

92. How many bonding sites do the halogens have? 1 What about members of the nitrogen family? 3

93. Why is an ion of sodium (Na^+) safe to eat, but an atom of sodium (Na) isn't? Sodium ions are stable because they have already achieved a full valence shell by losing an electron. Na atoms are highly reactive.

94. What will be the most likely ion that the following atoms form:

a. Na	Na^+	c. Mg	Mg^{2+}	e. Al	Al^{3+}
b. P	P^{3-}	d. Br	Br^-	f. S	S^{2-}

95. Define the octet rule. Atoms tend to form bonds in order to achieve a full valence shell which is generally 8 electrons (though may only be two for smaller atoms).

96. What is a covalent bond? When two elements of similar electronegativity fight over (share) electrons.
    a. What types of elements form covalent bonds? Non-metals
    b. How is the octet rule satisfied? By the elements sharing electrons to achieve filled valence shells
    c. What is the difference between polar and nonpolar covalent bonds? In non-polar covalent bonds the electrons are shared equally because the atoms are very similar in electronegativity. In polar bonds, the electrons spend more time near one atoms as opposed to the other.

97. What is an ionic bond? An electrostatic attraction between cations and anions formed when the atoms exchange electrons.
    a. What types of elements form ionic bonds? Typically a metal and a non-metal
    b. How is the octet rule satisfied? By one element removing electrons from another. (Transfer of electrons)

98. What is a metallic bond? The collective sharing of delocalized electrons by metallic cations.

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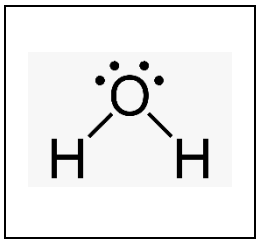
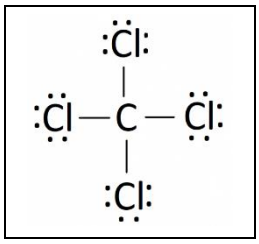
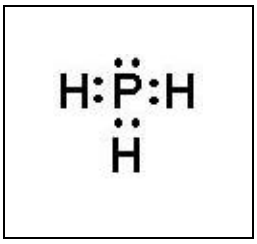
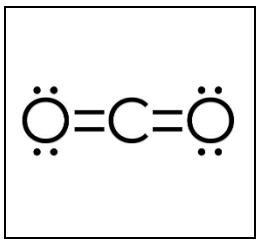
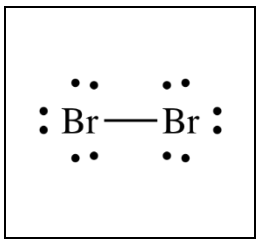
99. A(n) ____ionic____ bond is one in which two oppositely charged particles are drawn together by electrostatic forces.
100. A(n) covalent__ bond occurs when two atoms of similar electron affinity and ionization energy fight over electrons.
101. Electrons are easily stolen from metals because their _____ionization energy___ is relatively low when compared to non-metals.
102. Covalent bonds occur because atoms need to become stable by filling their ____valence shell___.
103. Non-metals are able to steal a metal's electrons because non-metals have high _____electron affinity____.
104. A __double___ covalent bond occurs when two atoms "share" two pairs of electrons.
105. An ionic bond occurs when electrons are stolen from _____metals___ by non-metals.
106. What type of bond does a molecule have? _____covalent_____
107. What type of bond does a salt crystal have? ____ionic_____
108. What type of bond has a repeating pattern of positive and negative ions? __Ionic_____ How are these ions held together? ____Electrostatic forces hold these particles in a crystalline structure_
109. Draw Lewis dot structures for the following:

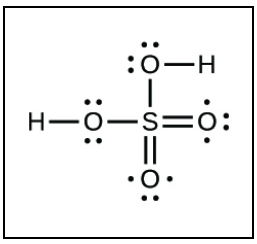
a. H_2O 

b. CCl_4 

c. PH_3 

d. CO_2 

e. Br_2 

f. H_2SO_4 

110. The octet rule states that elements form bonds in order to form a(n) __full valence shell____.
111. When 3 pairs of electrons are shared by two atoms, you have a(n) ___triple covalent bond__.
112. What is VSEPR? What does it cause electron pairs to do in the valence shell? __VSEPR = valence shell electron pair repulsion theory. VSEPR explains how electron pairs around the central atom of a molecule spread out due to electrostatic repulsion and cause predictable geometric arrangements of atoms.___
113. Rank the repulsion of the following electron pairs.

Pair Types	Shared – Shared	Shared – Unshared	Unshared – Unshared
Rank	least	in between	most

114. What 2 things help determine if a molecule is a polar molecule? _____Polarity of bonds and geometry_____
115. Rank the following bonds from most polar (5) to least polar (1).

Bond	B-C	C-F	N-F	B-N	B-F
Polarity Rank	1	4	3	2	5

116. Which of the following would have the highest boiling point? (remember polar compounds have higher boiling points because they are "stickier") N_2, CO_2, H_2O, BF_3. _H_2O_
117. Label the ends of the following bonds as positive or negative.

+ → –	+ → –	– ← +	– ← +
C – O	N – O	F – Cl	O – P

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118. Why is an ionic compound like NaCl more soluble in water than a covalent compound like CO_2? ___The ions in the ionic substance will be attracted to the polar water molecules through ion-dipole attractions whereas CO_2 is non-polar and can only form ion-induced dipole attractions.___
119. What type of substance has a high melting point, conducts electricity in solution, but does not conduct electricity as a solid? ___Ionic____
120. Write the formula for the following substances:

    a. calcium hydroxide ___$Ca(OH)_2$_____
    b. potassium hydrogen carbonate __$KHCO_3$_____
    c. silicon hexafluoride _____SiF_6_____
    d. dinitrogen tetroxide _____N_2O_4____
    e. sodium oxalate _____$Na_2C_2O_4$__
    f. mercury (I) bromide __Hg_2Br_2_____
    g. manganese (III) oxide _____Mn_2O_3___
    h. boron trifluoride _____BF_3_____

121. The force that holds water molecules together is a type of __intermolecular__ force, called a __hydrogen_ bond.
122. The phase change from solid to liquid is called __melting_. The phase change between liquid and solid is called _freezing_. These two phase changes occur at the __same_ temperature.
123. To cause something to freeze, ___energy__ must be removed.
124. The phase change from liquid to gas is called _vaporization_. The phase change between gas and liquid is called _condensation_.
125. When molecules lose _energy_ they get closer together allowing the hydrogen bonds to attract them together.
126. The boiling point of water at standard atmospheric pressure is __100_ ºC.
127. Whenever a phase change occurs between solid to liquid or liquid to gas it requires extra _energy_ to break the attractions between molecules.
128. Whenever a phase change occurs between gas to liquid, or liquid to solid, it gives off extra __energy_ to the surrounding environment.
129. Rank the following intermolecular forces from weakest (1) to strongest (4).

IMF	dipole-dipole	ion-dipole	induced dipole-induced dipole	induced dipole-permanent dipole
Rank	3	4	1	2