Test bank chapter (1)

Choose the correct answer

1. The SI unit of time is the

   a) hour
   b) second
   c) minute
   d) ampere

2. The diameter of an atom is approximately 1×10^{-7} mm. What is this diameter when expressed in nanometers?

   a) 1×10^{-18} nm
   b) 1×10^{-15} nm
   c) 1×10^{-9} nm
   d) 1×10^{-1} nm

3. 6.0 km is how many micrometers?

   a) 6.0×10^{6} µm
   b) 1.7×10^{-7} µm
   c) 6.0×10^{9} µm
   d) 1.7×10^{-4} µm

4. The SI prefixes *giga* and *micro* represent, respectively:

   a) 10^{-9} and 10^{-6}.
   b) 10^{6} and 10^{-3}.
   c) 10^{3} and 10^{-3}.
   d) 10^{9} and 10^{-6}.

5. Which of these quantities represents the largest mass?

   a) 2.0×10^{2} mg
   b) 0.0010 kg
   c) 1.0×10^{5} µg
   d) 2.0×10^{2} cg

6. How many cubic centimeters are there in exactly one cubic meter?

   a) 1×10^{-6} cm^3
   b) 1×10^{-3} cm^3
   c) 1×10^{-2} cm^3
   d) 1×10^{6} cm^3

7. Ammonia boils at -33.4°C. What temperature is this in °F?

a) -60.1°F
b) -92.1°F
c) -28.1°F
d) +13.5°F

8. Which of the following is not an SI base unit?

a) **Kilometer**
b) Kilogram
c) Second
d) Kelvin

9. Which of the following SI base units is not commonly used in chemistry?

a) kilogram
b) kelvin
c) **candela**
d) mole

10. Which of the following prefixes means 1/1000?

a) kilo
b) deci
c) centi
d) milli

11. Which of the following prefixes means 1000?

a) kilo
b) deci
c) centi
d) milli

12. Convert -77°F to kalvin ?

a) 212.6 K
b) -212.6 K
c) -28.1 K
d) +13.5 K

13. The number 0.0005678 expressed in scientific notation is:

a) 5.678×10^4
b) 5.67×10^{-7}
c) 5.678×10^{-4}
d) $5.678\ 10^{-3}$

Explanation: Since this number is less than one star moving the decimal point to the right until there is ONE non-zero number to the left of the decimal point. Write the rest of the number as is. Write the exponent as the number of places the decimal point was moved.

14. Which of the following is the smallest distance?

a) 21 m
b) 2.1×10^2 cm
c) 21 mm
d) **2.1×10^4 pm**

Explanation: Even though 2.1×10^4 is the largest number in this question, the units of pm (picometers) are the smallest units here, making it the smallest distance.

15. What temperature is 95 °F when converted to degrees Celsius?

a) 63 °C
b) **35 °C**
c) 127 °C
d) 15 °C

16. What temperature is 37 °C when converted to kelvin?

a) **310.15**
b) 99 k
c) 236 k
d) 67.15

17. What temperature is 77 K when converted to degrees Celsius?

a) –296 °C
b) 105 °C
c) **–196 °C**
d) 25 °C

18. Express 75 Tg as pg

a) 0.75 pg
b) **$75 \text{ X}10^{24}$ pg**
c) 0.75 pg
d) $75 \text{ X}10^{-24}$ pg

19. The SI prefixes *Tara* and *nano* represent, respectively:

a) 10^{-9} and 10^{-6}.
b) 10^{6} and 10^{-3}.
c) 10^{3} and 10^{-3}
d) 10^{12} and 10^{-9}

20. Which of these quantities represents the smallest mass?

a) 2.0×10^2 mg
b) 0.0010 kg
c) **1×10^5 µg**
d) 2.0×10^2 cg

21. Express 7.5 ng as Tg

a) **7.5 X10^{-21} Tg**
b) 75 X10^{24} Tg
c) 0.75 Tg
d) 7.5 X10^{21} Tg

28. At what temperature does the numerical reading on a Fahrenheit thermometer equal that on a Celsius thermometer?

a) 0 °F
b) **– 40°F**
c) 100 °F
d) –32 °F

Explanation: since the temperature reading is the same so that mean °F = °C

? F = [°C × 9/5] + 32 °F

Let temperature = t

t = [t × 9/5] + 32 °F

t - 9/5 t = 32 °F

-4/5 t = 32 °F

t = - 40 °F = -40 °C

Test bank chapter (2)

Choose the correct answer

NOTE: A periodic table is required to work many of the problems in this chapter.

1. Which of these elements is most likely to be a good conductor of electricity?

   a) N
   b) S
   c) He
   d) Fe

2. An atom of the isotope sulfur-31 consists of how many protons, neutrons, and electrons? (p = proton, n = neutron, e = electron)

   a) 15 p, 16 n, 15 e
   b) 16 p, 15 n, 16 e
   c) 16 p, 31 n, 16 e
   d) 32 p, 31 n, 32 e

3. A magnesium ion, Mg^{2+}, has

   a) 12 protons and 13 electrons.
   b) 24 protons and 26 electrons.
   c) 12 protons and 10 electrons.
   d) 24 protons and 22 electrons.

4. Which of these pairs of elements would be most likely to form an ionic compound?

   a) P and Br
   b) Cu and K
   c) C and O
   d) O and Zn

5. The elements in a column of the periodic table are known as
   a) metalloids.
   b) a period.
   c) noble gases.
   d) a group.

6. Which is the correct formula for copper (II) phosphate?

   a) Cu_2PO_4
   b) $Cu_3(PO_4)_2$
   c) Cu_2PO_3
   d) $Cu(PO_4)_2$

7. The correct name for NH_4NO_3 is

   a) ammonium nitrate.
   b) ammonium nitrogen trioxide.
   c) ammonia nitrogen oxide.
   d) hydrogen nitrogen oxide.

8. What is the formula for the ionic compound formed by calcium ions and nitrate ions?

   a) Ca_3N_2
   b) $Ca(NO_3)_2$
   c) Ca_2NO_3
   d) Ca_2NO_2

9. The Stock system name for Mn_2O_7 is

   a) dimanganese heptaoxide.
   b) magnesium oxide.
   c) manganese(VII) oxide.
   d) manganese(II) oxide.

10. Which of these elements is chemically similar to oxygen?

   a) sulfur
   b) calcium
   c) iron
   d) nickel

11. The formula of stannic oxide is SnO_2. The valence of Sn is:

   a) +1
   b) +2
   c) +3
   d) **+4**

Explanation: to know the charge on Sn atom, make this simple calculation remember that the charge on oxygen atom is -2 , let X is the charge on Sn atom

X+ (-2 (charge on O) × 2 (number of O atoms) = 0 (equal zero because the compound is neutral)

X - 4 = 0 >>>>>> x = + 4

12. Which pair of atoms constitutes a pair of isotopes of the same element?

(a). ${}^{14}_{6}X$ ${}^{14}_{7}X$
(b). ${}^{14}_{6}X$ ${}^{12}_{6}X$
(c). ${}^{17}_{9}X$ ${}^{17}_{8}X$
(d). ${}^{19}_{10}X$ ${}^{19}_{9}X$

Explanation: Isotopes of an element are atoms of the same element with same number of protons but different number of neutrons. Only choice (b) has 2 atoms of X with 6 protons and 8 and 6 neutrons respectively.

13. Elements in Group 8A are known as the ___________.

a) chalcogens
b) alkali metals
c) **noble gases**
d) alkaline earth metals

14. ____________ typically forms ions with a 2+ charge.

a) Transition metals
b) Halogens
c) **Alkaline earth metals**
d) Alkali metals

Explanation: The alkaline earth metals are in group 2A of the periodic table and lose 2 electrons to form cations with 2 positive charges.

15. An *anion* is defined as
a) **a charged atom or group of atoms with a net negative charge.**
b) a stable atom.
c) a group of stable atoms.
d) an atom or group of atoms with a net positive charge.

16. A cation is defined as

a) a charged atom or group of atoms with a net negative charge.
b) a stable atom.
c) a group of stable atoms.
d) **an atom or group of atoms with a net positive charge.**

17. Atoms of the same element with different mass numbers (or number of neutrones) are called

a) ions.
b) neutrons.
c) chemical families.
d) **isotopes.**

18. How many neutrons are there in an atom of lead $_{82}Pb$ whose mass number is 208?

a) 82
b) **126**
c) 208
d) 290

19. Molecules consist of the same element with different numbers of atoms and chemical structure are called …

a) Ions
b) Neutrons
c) **Allotropes**
d) Isotopes

20. An atom of the isotope ^{16}S-31 consists of how many protons, neutrons, and electrons?

a) 15 p, 16 n, 15 e
b) **16 p, 15 n, 16 e**
c) 16 p, 31 n, 16 e
d) 32 p, 31 n, 32 e

21. A magnesium ion, $_{20}Ca^{2+}$, has

a) 20 protons and 22 electrons.
b) 20 protons and 20 electrons.
c) **20 protons and 18 electrons.**
d) 22 protons and 20 electrons.

22. A sulfide ion, $_{16}S^{2-}$, has:

a) 16 protons and 16 electrons
b) 32 protons and 16 electrons
c) 16 protons and 14 electrons
d) **16 protons and 18 electrons**

23. Which of these pairs of elements would be most likely to form a molecular compound?

a) Na and Br
b) Ca and O
c) **C and O**
d) Zn and O

24. What is the formula for the ionic compound formed by calcium ions and nitrate ions?

a) Ca_3N_2
b) **$Ca(NO_3)_2$**
c) Ca_2NO_3
d) Ca_2NO_2

25. Which is the correct formula for copper(II) phosphate?

a) Cu_2PO_4
b) **$Cu_3(PO_4)_2$**
c) Cu_2PO_3
d) $Cu(PO_4)_2$

26. The correct name for NH_4NO_3 is

a) **ammonium nitrate.**
b) ammonium nitrogen trioxide.
c) ammonia nitrogen oxide.
d) hydrogen nitrogen oxide.

27. The correct name for PCl_5 is

a) monophosphate pentachloride
b) phosphorus chloride
c) monophosphate tetrachloride
d) **phosphorus pentachloride**

28. Which of the following expressions represents two molecules of water?

a) H_2O
b) H_2O_2
c) **$2\ H_2O$**
d) $2\ HO_2$

29. The empirical formula of a compound with molecules containing 12 carbon atoms, 14 hydrogen atoms, and 6 oxygen atoms is ___________.

a) $C_{12}H_{14}O_6$
b) C_2H_4O
c) CH_2O
d) **$C_6H_7O_3$**

Explanation: The empirical formula is always the simplest possible whole number ratio between the atoms of the molecules.

30. The charge on the manganese in the salt MnF_3 is ______________.

a) +1
b) -1
c) **+3**
d) -2

Explanation: Since every F has one negative charge, the Mn can have only 3 positive charges.

30. Magnesium reacts with a certain element to form a compound with the general formula MgX. What would the most likely formula be for the compound formed between potassium and element X?

a) KX
b) K_2X_2
c) K_2X_3
d) **None of the above**

Explanation: In the compound MgX, X must have 2 negative charges since Mg will always have 2 positive charges. The element K will always form an ion with 1 positive charge and hence the only combination of K and X could be K_2X, which is not one of the options.

31. Barium forms an ion with a charge of _________.

a) +1
b) -2
c) +3
d) None of the above.

Explanation: Barium is in group 2A of the periodic table and forms ions with only 2 positive charges.

31. Aluminum forms an ion with a charge of___________.

a) +2
b) -3
c) **+3**
d) +1

32. Iodine forms an ion with a charge of _____________.

a) -7
b) +1
c) **-1**
d) +2

33. The chemical symbol for the ion with 11 protons and 10 electrons.

a) Na
b) F^-
c) Ne
d) **Na^+**

34. Which of these compounds is a binary compound?

a) **NaCl**
b) $MgSO_4$
c) NaOH
d) HCN

35. Atoms with the same number of electrons and number of protons are called…

a) ions
b) isotopes
c) **neutral atoms**
d) different atoms

36. Atoms which have different number of electrons are called…

a) **ions**
b) isotopes
c) neutral atoms
d) different atoms

37. Use the following table and choose which of the species are positively charged?

Atom or ion element	I	II	III	IV	V	VI
Atom or ion electrons (e)	6	10	18	10	28	7
Atom or ion protons (p)	6	8	17	11	30	7
Atom or ion neutrons (n)	6	8	18	11	36	6

A. III and V
C. II and III
B. **IV and V**
D. I and VI

38. Which isotope has 45 neutrons?

(a). ${}^{80}_{36}Kr$
(b). ${}^{78}_{34}Se$
(c). ${}^{80}_{35}\mathbf{Br}$
(d). ${}^{34}_{17}Cl$

39. In the periodic table, the elements are arranged in __________.

a) alphabetical order
b) **order of increasing atomic number**
c) order of increasing metallic properties
d) order of increasing neutron content

40. An element in the upper right corner of the periodic table is _________.

a) either a metal or metalloid
b) definitely a metal
c) **definitely a non-metal**
d) either a metalloid or a non-metal

41. An element that appears in the lower left corner of a periodic table is _________.

a) either a metal or metalloid
b) **definitely a metal**
c) either a metalloid or a non-metal
d) definitely a non-metal

42. A molecular formula always indicates __________.

a) **how many of each atom are in a molecule**
b) the simplest whole-number ratio of different atoms in a compound
c) which atoms are attached to which in a molecule
d) the isotope of each element in a compound
e)

43. An empirical formula always indicates___________.

a) which atoms are attached to which in a molecule
b) how many of each atom are in a molecule
c) **the simplest whole-number ratio of different atoms in a compound**
d) the geometry of a molecule

44. There are ________ protons, ________ neutrons, and ________ electrons in $^{131}I^{-}$.

a) 131, 53, and 54
b) 131, 53 and 52
c) **53, 78, and 54**
d) 53, 131, and 52

45. Which species has 48 electrons?

(a). $^{118}_{50}\mathbf{Sn}^{+2}$
(b). $^{116}_{50}Sn^{+4}$
(c). $^{112}_{48}Cd^{+2}$
(d). $^{68}_{31}Ga$

Test bank chapter (3)

Choose the correct answer

1. What is the mass, in grams, of one copper atom?

a) 1.055×10^{-22} g
b) 63.55 g
c) 1 amu
d) 1.66×10^{-24} g

2. Determine the number of moles of aluminum in 96.7 g of Al.

a) 0.279 mol
b) 3.58 mol
c) 7.43 mol
d) 4.21 mol

3. Which of the following samples contains the greatest number of atoms?

a) 100 g of Pb
b) 2.0 mole of Ar
c) mole of Fe
d) 5 g of He

4. Formaldehyde has the formula CH_2O. How many molecules are there in 0.11 g of formaldehyde?

a) 6.1×10^{-27}
b) 3.7×10^{-3}
c) 4
d) 2.2×10^{21}

5. How many sulfur atoms are present in 25.6 g of $Al_2(S_2O_3)_3$?

a) 0.393
b) 6
c) 3.95×10^{22}
d) 2.37×10^{23}

6. The percent composition by mass of a compound is 76.0% C, 12.8% H, and 11.2% O. The molar mass of this compound is 284.5 g/mol. What is the molecular formula of the compound?

a) $C_{10}H_6O$
b) $C_9H_{18}O$
c) $C_{16}H_{28}O_4$
d) $C_{18}H_{36}O_2$

7. What is the coefficient of H_2O when the following equation is properly balanced with the smallest set of whole numbers?

___ Al_4C_3 + ___ H_2O → ___ $Al(OH)_3$ + ___ CH_4

a) 3
b) 4
c) 6
d) 12

8. When 22.0 g NaCl and 21.0 g H_2SO_4 are mixed and react according to the equation below, which is the limiting reagent? **$2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$**

a) H_2SO_4
b) Na_2SO_4
c) HCl
d) NaCl

9. When the following equation is balanced, the coefficients are ____________.

NH_3 (g) + O_2 (g) → NO2 (g) + H2O (g)

(a). 1, 1, 1, 1

(b). 2, 3, 2, 3

(c). 4, 7, 4, 6

(d). 1, 3, 1, 2

10. How many moles of carbon atoms are in 4 mol of dimethylsulfoxide (C_2H_6SO)?

a) 2
b) 6
c) 8
d) 4

Explanation: This is based on reading the formula and correctly extracting information from it. The formula C_2H_6SO indicates that every mole of this compound has 2 moles of carbon atoms in it. Thus 4 moles of the compound would have 4 x 2 = 8 moles of C atoms.

11. There are ______ sulfur atoms in 25 molecules of $C_4H_4S_2$.

a) 1.5×10^{25}
b) 4.8×10^{25}
c) 3.0×10^{23}
d) **50**

Explanation: The molecular formula indicates that every molecule of $C_4H_4S_2$ has 2 sulfur atoms per molecule and hence 25 molecules of this compound will have 25 x 2 = 50 atoms of sulfur.

12. There are _______ hydrogen atoms in 25 molecules of $C_4H_4S_2$.

a) 25
b) 3.8×10^{24}
c) 6.0×10^{25}
d) **100**

Explanation: The formula of $C_4H_4S_2$ indicates that there are 4 hydrogen atoms per molecule and hence 100 hydrogen atoms in 25 molecules of $C_4H_4S_2$.

13. How many grams of oxygen are in 65.0 g of $C_2H_2O_2$?

a) 18
b) 29
c) 9.5
d) **35.8**

Explanation: This question uses the mole to mole ratio between oxygen and $C_2H_2O_2$ and needs the following steps. $\frac{65.0\text{ g } C_2H_2O_2}{58.0\text{ g}\cdot\text{mol}^{-1}} \times \frac{2\text{ moles O}}{1\text{ mole } C_2H_2O_2} \times \frac{15.99\text{ g O}}{1\text{ mole of O}} = 35.8\text{ g of O}$

17. How many moles of carbon dioxide are there in 52.06 g of carbon dioxide?

a) 0.8452
b) **1.183**
c) 1.183×10^{23}
d) 8.648×10^{2}

Explanation: This is a straight-forward conversion from grams to moles of CO_2 which is done as follows:

$$52.06\text{ g } CO_2 \times \frac{1\text{ mole } CO_2}{43.99\text{ g } CO_2} = 1.183\text{ moles of } CO_2$$

18. How many moles of the compound magnesium nitrate, Mg(NO3)2, are in a2.35 g sample of this compound?

a) 38.4
b) 65.8
c) 0.0158
d) 0.0261

Explanation: This is a straight-forward conversion from grams to moles of Mg(NO3)2 which is done as follows:

$$2.35 \text{ g } Mg(NO_3)_2 \times \frac{1 \text{ mole } Mg(NO_3)_2}{148.3148 \text{ g}} = 0.0158 \text{moles}$$

19. A 25.5-g sample of ammonium carbonate contains _______ mol of ammonium ions.

a) 0.468
b) 0.288
c) 0.531
d) 2.00

Explanation: Realize that the formula for ammonium carbonate is (NH4)2CO3 and calculate the molar mass (96.0856 g/mol). Convert grams to moles and then using the stoichiometric ratio find the # of moles of ammonium ions.

$$25.5 \text{ g } (NH_4)_2CO_3 \times \frac{1 \text{ mol } (NH_4)_2CO_3}{96.0856 \text{ g}} \times \frac{2 \text{ moles } NH_4^+}{1 \text{ mol } (NH_4)_2CO_3} = 0.531 \text{ moles}$$

20. Magnesium and nitrogen react in a combination reaction to produce magnesium nitride:

$$3Mg + N_2 \rightarrow Mg_3N_2$$

In a particular experiment, a 5.47-g sample of N_2 reacts completely. How many grams of Mg are needed for this reaction?

a) 14.2 g
b) 24.1 g
c) 16.1 g
d) 0.92 g

Explanation: Ensure that the equation is balanced. The grams of N_2 must be converted to moles of N_2 and then using the stoichiometric ratio between the Mg and N_2, the grams of Mg can be calculated.

$$5.47 \text{ g } N_2 \times \frac{1 \text{ mole } N_2}{28.0134 \text{ g}} \times \frac{3 \text{ mole Mg}}{1 \text{ mole } N_2} \times \frac{24.3050 \text{ g Mg}}{1 \text{mole Mg}} = 14.2 \text{ g Mg}$$

21. What information would you need to calculate the average atomic mass of an element?

a) The number of neutrons in the element.
b) The atomic number of the element.
c) The mass and abundance of each isotope of the element.
d) The position in the periodic table of the element.

22. The atomic masses of Cl (75.53 %) and Cl (24.47 %) are 34.968 amu and 36.956 amu, respectively. Calculate the average atomic mass of chlorine.

a) 35.96 amu
b) 35.45 amu
c) 36.47 amu
d) 71.92 amu

23. How many atoms are there in 5.10 moles of sulfur (16S=32 amu)?

a) 3.07×10^{24}
b) 9.59×10^{22}
c) 6.02×10^{23}
d) 9.82×10^{25}

24. Iodine has two isotopes 126I and 127I, with the equal abundance.Calculate the average atomic mass of Iodine (53I).

a) 126.5 amu
b) 35.45 amu
c) 1.265 amu
d) 71.92 amu

25. The atomic masses of 6Li and 7Li are 6.0151 amu and 7.0160 amu, respectively. Calculate the natural abundance of these two isotopes. The average atomic mass of Lithium (Li=6.941 amu).

a) ^{6}Li= 7.49% , ^{7}Li= 92.51%
b) ^{7}Li= 7.49% , ^{6}Li= 92.51%
c) ^{6}Li= 8.49% , ^{7}Li= 95.51%
d) ^{7}Li= 7.22% , ^{6}Li= 82.51%

26. How many atoms are present in 3.14 g of copper (Cu)?

a) 2.98×10^{22}
b) 1.92×10^{23}
c) 1.89×10^{24}
d) 6.02×10^{23}

27. Nitric oxide (NO) reacts with oxygen gas to form nitrogen dioxide (NO_2), a dark-brown gas:

$$2NO\,(g) + O_2(g) \rightarrow 2NO_2(g)$$

In one experiment 0.886 mole of NO is mixed with 0.503 mole of O_2. Calculate the number of moles of NO_2 produced (note: first determine which is the limiting reagent).

a) 0.886 mol

b) 0.503 mol

c) 1.01 mol

d) 1.77 mol

28. The fertilizer ammonium sulfate [$(NH_4)_2SO_4$] is prepared by the reaction between ammonia (NH_3) and sulfuric acid:

$$2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$$

How many kilograms of NH_3 are needed to produce 1.00 ´ 10^5 kg of $(NH_4)_2SO_4$?

a) 1.70×10^4 kg

b) 3.22×10^3 kg

c) 2.58×10^4 kg

d) 7.42×10^4 kg

29. Consider the combustion of carbon monoxide (CO) in oxygen gas:

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

Starting with 3.60 moles of CO, calculate the number of moles of CO_2 produced if there is enough oxygen gas to react with all of the CO.

a) 7.20 mol
b) 44.0 mol
c) 3.60 mol
d) 1.80 mol

30. Nitrous oxide (N_2O) is also called "laughing gas." It can be prepared by the thermal decomposition of ammonium Nitrate (NH_4NO_3). The other product is H_2O. The balanced equation for this reaction is:

$$NH_4NO_3 \rightarrow N_2O + 2H_2O$$

How many grams of N_2O are formed if 0.46 mole of NH_4NO_3 is used in the reaction?

a) 2.0 g

b) 3.7×10^1 g

c) 2.0×10^1 g

d) 4.6×10^{-1} g

31. What is the theoretical yield of chromium that can be produced by the reaction of 40.0 g of Cr_2O_3 with 8.00 g of aluminum according to the chemical equation below?

$$2Al + Cr_2O_3 \rightarrow Al_2O_3 + 2Cr$$

a) 7.7 g
b) **15.4 g**
c) 27.3 g
d) 30.8 g

32. Hydrogen fluoride is used in the manufacture of Freons (which destroy ozone in the stratosphere) and in the production of aluminum metal. It is prepared by the reaction

$$CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$$

In one process 6.00 kg of CaF_2 are treated with an excess of H_2SO_4 and yield 2.86 kg of HF. Calculate the percent yield of HF.

a) 93.0 %
b) 95.3 %
c) 47.6 %
d) 62.5 %

33. When 22.0 g NaCl and 21.0 g H_2SO_4 are mixed and react according to the equation below, which is the limiting reagent?

$$2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$

a) NaCl
b) H_2SO_4
c) Na_2SO_4
d) No reagent is limiting.

34. Hydrochloric acid can be prepared by the following reaction:

$$2NaCl(s) + H_2SO_4(aq) \rightarrow 2HCl(g) + Na_2SO_4(s)$$

How many grams of HCl can be prepared from 2.00 mol H_2SO_4 and 150 g NaCl?

a) 7.30 g
b) 93.5 g
c) 146 g
d) 150 g

35. Calculate the molar mass of Li_2CO_3.

a) 73.89 g
b) 66.95 g
c) 41.89 g
d) 96.02 g

36. How many molecules of ethane (C_2H_6) are present in 0.334 g of C_2H_6?

a) 2.01×10^{23}
b) 6.69×10^{21}
c) 4.96×10^{22}
d) 8.89×10^{20}

37. All of the substances listed below are fertilizers that contribute nitrogen to the soil. Which of these is the richest Source of nitrogen on a mass percentage basis?

a) Urea, $(NH_2)_2CO$
b) Ammonium nitrate, NH_4NO_3
c) Guanidine, $HNC(NH_2)_2$
d) Ammonia, NH_3

38. Allicin is the compound responsible for the characteristic smell of garlic. An analysis of the compound gives the following percent composition by mass: C: 44.4 percent; H: 6.21 percent; S: 39.5 percent; O: 9.86 percent. What is its molecular formula given that its molar mass is about 162 g?

a) $C_{12}H_{20}S_4O_2$
b) $C_7H_{14}SO$
c) $C_6H_{10}S_2O$
d) $C_5H_{12}S_2O_2$

39. The formula for rust can be represented by Fe_2O_3. How many moles of Fe are present in 24.6 g of the compound?

a) 2.13 mol
b) 0.456 mol
c) 0.154 mol
d) 0.308 mol

40. What is the mass, in grams, of one copper atom?

a) 1.055×10^{-22} g
b) 63.55 g
c) 1 amu
d) 1.66×10^{-24} g

41. How many grams of sulfur (S) are needed to react completely with 246 g of mercury (Hg) to form HgS?

a) 39.3 g
b) 24.6 g
c) 9.66×10^{3} g
d) 201 g

42. Tin(II) fluoride (SnF_2) is often added to toothpaste as an ingredient to prevent tooth decay. What is the mass of F in grams in 24.6 g of the compound?

a) 18.6 g
b) 24.3 g
c) 5.97 g
d) 75.7 g

43. What is the empirical formula of the compound with the following composition? 2.1 percent H, 65.3 percent O, 32.6 percent S.

a) H_2SO_4
b) H_2SO_3
c) $H_2S_2O_3$
d) HSO_3

44. Which of the following samples contains the greatest number of atoms?

a) 100 g of Pb
b) 2.0 mole of Ar
c) mole of Fe
d) 5 g of He

45. Formaldehyde has the formula CH_2O. How many molecules are there in 0.11 g of formaldehyde?

a) 6.1×10^{-27} molecule
b) 3.7×10^{-3} molecule
c) 4×10^{21} molecule
d) 2.2×10^{21} molecule

46. Determine the number of moles of aluminum in 96.7 g of Al.

a) 0.279 mol
b) 3.58 mol
c) 7.43 mol
d) 4.21 mol

47. How many sulfur atoms are present in 25.6 g of $Al_2(S_2O_3)_3$?

a) 0.393
b) 3.95×10^{22}
c) 7.90×10^{22}
d) 2.37×10^{23}

48. What is the coefficient of H_2O when the following equation is properly balanced with the smallest set of whole numbers?

$$\_\_\_\ \mathbf{Al_4C_3} + \_\_\_\ \mathbf{H_2O} \rightarrow \_\_\_\ \mathbf{Al(OH)_3} + \_\_\_\ \mathbf{CH_4}$$

a) 3
b) 4
c) 6
d) 12

50. Which of the following equations is balanced?

A) $2C + O_2 \rightarrow CO$
B) $2CO + O_2 \rightarrow 2CO_2$
C) $H_2 + Br_2 \rightarrow HBr$
D) $2K + H_2O \rightarrow 2KOH + H_2$

51. Determine the number of moles of aluminum in 96.7 g of Al

a) 0.279 mol
b) 3.58 mol
c) 7.43 mol
d) 4.21 mol

Test bank chapter (4)

Choose the correct answer

1. A 50.0 mL sample of 0.436 M NH_4NO_3 is diluted with water to a total volume of 250.0 mL. What is the ammonium nitrate concentration in the resulting solution?

a) 21.8 M
b) 0.459 M
c) 2.18×10^{-2} M
d) 8.72×10^{-2} M

2. How many milliliters would you need to prepare 60.0 mL of 0.200 *M* HNO_3 from a stock solution of 4.00 *M* HNO_3?

a) 3 mL
b) 240 mL
c) 24 mL
d) 1000 mL

3. What is the concentration (M) of KC1 in a solution made by mixing 25.0 mL of 0.100 M KC1 with 50.0 mL of 0.100 M KC1?

a) 0.0500
b) **0.100**
c) 0.0333
d) 0.0250

4. What is the concentration (M) of CH_3OH in a solution prepared by dissolving 11.7 g of CH_3OH in sufficient water to give exactly 230 mL of solution?

a) 11.7
b) 2.30 x 10-2
c) 0.0841
d) **1.59**

Explanation: Need to convert the grams of CH3OH to moles and then find the molarity of the solution by using the molarity formula. Do not forget to convert the ml to L.

$$11.7\,\text{g}\,CH_3OH \times \frac{1\,\text{mole}\,CH_3OH}{32.042\,\text{g}} \times \frac{1}{0.230\,\text{L}} = 1.59\,\text{M}$$

5. How many grams of H_3PO_4 are in 35.1 mL of a 2.75 M solution of H_3PO_4?

a) 0.61
b) **9.46**
c) 20
d) 4.9

Explanation: Need to convert the ml of H_3PO_4 to liters and then find the # of moles of phosphoric acid. The moles of phosphoric acid can then be converted to grams of phosphoric acid.

$$35.1\ \text{ml} \times \frac{1\ \text{L}}{1000\ \text{ml}} \times 2.75\ \text{M} \times \frac{97.99\ \text{g}\ H_3PO_4}{1\ \text{mole}} = 9.46\ \text{g}\ H_3PO_4$$

7. What is the concentration (M) of a Na_2SO_4 solution prepared by dissolving 5.35 g of Na_2SO_4 in sufficient water to give 330 mL of solution?

a) 1.14 x 102
b) 0.016
c) 61.7
d) **0.114**

Explanation: Convert grams of Na_2SO4 to moles of Na2SO4, ml of water to liters of water and then find the molarity of the solution by using the molarity formula.

$$5.35\ \text{g}\ Na_2SO_4 \times \frac{1\ \text{mole}\ Na_2SO_4}{142.035\ \text{g}\ Na_2SO_4} \times \frac{1}{0.330\ \text{L}} = 0.114\ \text{M}\ Na_2SO_4$$

8. How many grams of LiOH are there in 750.0 mL of a 0.0158 M LiOH solution?

a) 2.11 x 10-5
b) 11.3
c) **0.284**
d) 3.50

Explanation: Calculate the number of moles of LiOH present in this solution using the molarity formula and the convert the number of moles to grams of LiOH.

$$7.50 \times 10^{-1}\ \text{L} \times 0.0158\ \text{M} \times \frac{23.948\ \text{g}}{1\ \text{mole LiOH}} = 0.284\ \text{g LiOH}$$

10. A 50.0 mL sample of 0.436 M NH_4NO_3 is diluted with water to a total volume of 250.0 mL. What is the ammonium nitrate concentration in the resulting solution?

a) 21.8 M
b) 0.459 M
c) 2.18×10^{-2} M
d) 8.72×10^{-2} M

12. A 3.682 g sample of potassium chlorate $KClO_3$ is dissolved in enough water to give 375 mL of solution. What is the chlorate ion concentration in this solution?

a) 3.00×10^{-2} M
b) 4.41×10^{-2} M
c) 0.118 M
d) 8.01×10^{-2} M

Test bank chapter (5)

Choose the most correct answer

1. A sample of oxygen occupies 47.2 liters under a pressure of 1240 torr at 25°C. What volume would it occupy at 25°C if the pressure were decreased to 730 torr?

a) 27.8 L
b) 29.3 L
c) 32.3 L
d) **80.2 L**

2. Under conditions of fixed temperature and amount of gas, Boyle's law requires that

I. $P_1V_1 = P_2V_2$
II. PV = constant
III. $P_1/P_2 = V_2/V_1$

a) I only
b) II only
c) III only
d) **I, II, and III**

3. The volume of a sample of nitrogen is 6.00 liters at 35°C and 740 torr. What volume will it occupy at STP?

a) 6.59 L
b) 5.46 L
c) 6.95 L
d) **5.18 L**

4. The density of chlorine gas at STP, in grams per liter, is approximately:

a) 6.2
b) **3.2**
c) 3.9
d) 4.5

Explanation: d = molar mass × p/ RT = 70 × 1/ 0.082× 273 = 3.17 g/L

5. What pressure (in atm) would be exerted by 76 g of fluorine gas in a 1.50 liter vessel at -37°C?

a) **26 atm**
b) 4.1 atm
c) 19,600 atm
d) 84 atm

6. What is the density of ammonia gas at 2.00 atm pressure and a temperature of 25.0°C?

a) 0.720 g/L
b) 0.980 g/L
c) **1.39 g/L**
d) 16.6 g/L

7. Convert 2.0 atm to mmHg

a) 150 mmHg
b) 0.27 mmHg
c) 150 mmHg
d) **1520 mmHg**

8. A container with volume 71.9 mL contains water vapor at a pressure of 10.4 atm and a temperature of 465^{o}C. How many grams of the gas are in the container?

a) 0.421 g
b) **0.222 g**
c) 0.183 g
d) 0.129 g

Explanation: n = PV/RT = 0.0719× 10.4 = 0.0821 × (465 +273) = 0.012 mole

Mass = n × **molar mass** = 0.012 × 18 = 0.222 g

9. What is the molar mass of a pure gaseous compound having a density of 4.95 g/L at -35 ^{o}C and 1020 torr?

a) 24 g/mole
b) 11 g/mole
c) **72 g/mole**
d) 120 g/mole

10. A 0.580 g sample of a compound containing only carbon and hydrogen contains 0.480 g of carbon and 0.100 g of hydrogen. At STP, 33.6 mL of the gas has a mass of 0.087 g. What is the molecular (true) formula for the compound?

a) CH_3
b) C_2H_6
c) C_2H_5
d) **C_4H_{10}**

11. Gas occupy 6L at 37^{0}C what will be its volume when its temperature is doubled?

a) 12 L
b) 6L
c) 3.2 L
d) 2L

12. A mixture of 90.0 grams of CH_4 and 10.0 grams of argon has a pressure of 250 torr under conditions of constant temperature and volume. The partial pressure of CH_4 in torr is:

(a) 143
(b) 100
(c) 10.7
(d) 239

Explanation: from Dalton law >>> $P_{CH4} = X_{CH4} P_{total}$, n_{CH4} = 90/ 16 =5.625 mole , n_{Ar} = 10/39.95 = 0.250 mole

$X_{CH4} = n_{CH4} / n_{CH4} + n_{Ar}$ = 5.625 / 5.625 + 0.250 = 0.96 >>>> P_{CH4} = 0.96 ×250 = 239.3 torr

13. What pressure (in atm) would be exerted by a mixture of 1.4 g of nitrogen gas and 4.8 g of oxygen gas in a 200 mL container at 57^{o}C?

a) 4.7
b) 34
c) 47
d) **27**

Explanation: P = n_{total} RT/V , n_{N2} = 1.4/ 2×14 = 0.05 mole , n_{O2} = 4.8/2×16 = 0.15 mole

P = (0.05 +0.15) 0.0821×(57+273) /0.2 = 27 atm

14. A sample of hydrogen gas collected by displacement of water occupied 30.0 mL at 24°C and pressure 736 torr. What volume would the hydrogen occupy if it were dry and at STP? The vapor pressure of water at 24.0°C is 22.4 torr.

a) 32.4 mL
b) 21.6 mL
c) 36.8 mL
d) **25.9 mL**

Explanation: from Dalton law >>> $P_{H2} = P_{total} - P_{H2O}$, $P_{H2} = 736 - 22.4 = 713.6$ torr

n = PV /RT >>>>> n =(713.6/760) × 0.03/ 0.0821× (24+273) = 0.00115 mle

at STP >>>>> V = nRT/P = 0.00115× 0.0821×273 /1 = 0.026 L × 1000 = 25.89 mL

15. Ammonia burns in oxygen gas to form nitric oxide (NO) and water vapor. How many volumes of NO are obtained from one volume of ammonia at the same temperature and pressure?

$$4\,NH_3\,(g) + 5\,O_2\,(g) \rightarrow 4NO\,(g) + 6\,H_2O\,(g)$$

a) **One**
b) (b) Two
c) (c) Three
d) (d) Four

16. The pressure of 6.0 L of an ideal gas in a flexible container is decreased to one-third of its original value, and its absolute temperature is decreased by one-half. What is the final volume of the gas?

a) **9.0 L**
b) 6.0 L
c) 4.0 L
d) 1 L

Explanation: let $V_1 = 6$ & $V_2 = ?$, $T_1 = T$ & $T_2 = ½\,T$, $P_1 = P$ & $P_2 = 1/3\,P$

From combined gas law **$P_1V_1/T_1 = P_2V_2/T_2$** >>>>> $\frac{P\times 6}{T} = \frac{(\frac{1}{3})P\times V2}{(\frac{1}{2})T}$ >>> $V2 = \frac{P\times 6\times T\times 3}{T\times 2\times P} = 9\,L$

17. Gas A is at 30°C and gas B is at 20°C. Both gases are at 1atmosphere. What is the ratio of the volume of 1 mole gas A to 1 mole of gas B

a) 606:303
b) 3 : 2
c) 2 : 3
d) **303 :293**

Explanation: $\frac{V_A}{T_A} = \frac{V_B}{T_B}$ >>>>> $\frac{V_A}{30+273} = \frac{V_B}{20+273}$ >>>> $\frac{V_A}{303} = \frac{V_B}{293}$

18. The sample of argon occupies 50L at standard temperature. Assuming constant pressure, what volume with the gas occupy if the temperature is doubled.

a) 25L
b) 50L
c) **100L**
d) 100 mL

19. What total gas volume (in liters) at 520°C and 880 torr would result from the decomposition of 33 g of potassium bicarbonate according to the equation:

$$2KHCO_3(s) \longrightarrow K_2CO_3(s) + CO_2(g) + H_2O(g)$$

(a) 56 L
(b) 37 L
(c) 10 L
(d) 19 L

20. Calculate the weight of $KClO_3$ that would be required to produce 29.5 L of oxygen measured at 127°C and 760 torr.
$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$

(a) **73.5 g**
(b) 12.2 g
(c) 14.6 g
(d) 24.4 g

21. The ideal gas law predicts that the molar volume (volume of one mole) of gas equals:

(a) mRT/PV
(b) (MM)P/RT
(c) $1/2ms^{-2}$
(d) RT/P

22. For a gas, which pair of variables are inversely proportional to each other (if all other conditions remain constant)?

a) **P, V**
b) V, T
c) n, V
d) n, P

23. Convert 562 mmHg to atm

a) **0.739 atm**
b) 4.27×10^5 atm
c) 1.05 atm
d) 0.562 atm

24. What is the volume of one mole of an ideal gas at STP?

a) 24.5 L
b) **22.4 L**
c) 1.0 L
d) 10.0 L

25. What are standard temperature and pressure (STP)?

a) 0 C, 1 torr
b) 25 C, 1 torr
c) **0 C, 1 atm**
d) 25 °C, 1 atm

26. What is the unit of mole fraction

a) mol
b) mol^{-1}
c) **unitless**

27. Refer to Dalton's law of partial pressures and explain what mole fraction is

a) The number of moles of one component
b) **The ratio of the number of moles of one component to the number of moles of all components present.**
c) The number of moles of one component divided by 100
d) The ratio of the number of moles of all components present to the number of moles of one component.

28. Write the ideal gas equation. Give the units for each term in the equation

a) $PV = nRT$; P in torr, V in L, n in mol, R in Latm/Kmol, T in °C.
b) $PV = nRT$; P in torr, V in L, n in mol, R in Latm/Kmol, T in K.
c) **(c) $PV = nRT$; P in atm, V in L, n in mol, R in Latm/Kmol, T in K.**
d) $PV = nRT$; P in atm, V in L, n in mol, R in Latm/Kmol, T in °C .

29. What is the difference between a gas and a vapor?

a) **A gas is a substance normally in the gaseous state at normal atmospheric conditions (25C, 1 atm); a vapor is the gaseous form of any substance that is a liquid or a solid at normal temperatures and pressures.**
b) A gas is the gaseous form of any substance; a vapor refers to a gas over a water surface.
c) A gas is a substance normally in the gaseous state at normal atmospheric conditions (25C, 1 atm); a vapor is a gas over a water surface.
d) A gas and a vapor are two interchangeable nomenclatures; they are identical.

30. What volume is occupied by 19.6 g of methane (CH_4) at 27°C and 1.59 atm?

a) 1.71 L
b) **18.9 L**
c) 27.7 L
d) 302 L

31. A 4.37 gram sample of a certain diatomic gas occupies a volume of 3.00 L at 1.00 atm and a temperature of 45°C. Identify this gas.

a) **F_2**
b) N_2
c) H_2
d) O_2

Explanation: MM = mRT/ PV >>>> MM = 4.37 × 0.0821× (45+273) /1×3 = 37.77/2 = 18.88 g/mole~ F_2

32. A sample of hydrogen gas was collected over water at 21°C and 685 mmHg. The volume of the container was 7.80 L. Calculate the mass of H_2(g) collected. (Vapor pressure of water = 18.6 mmHg at 21°C.)

a) 0.283 g
b) **0.572 g**
c) 0.589 g
d) 7.14

33.Which of the following is/are characteristic(s) of gases?

a) High compressibility
b) Relatively large distances between molecules
c) Formation of homogeneous mixtures regardless of the nature of gases
d) High compressibility, relatively large distances between molecules AND formation of homogeneous mixtures regardless of the nature of gases

34. A small bubble rises from the bottom of a lake, where the temperature and pressure are 4°C and 3.0 atm, to the water's surface, where the temperature is 25°C and the pressure is 0.95 atm. Calculate the final volume of the bubble if its initial volume was 2.1 mL.

a) 0.72 mL
b) 6.2 mL
c) 41.4 mL
d) 7.1 mL

35.Calculate the mass, in grams, of 2.74 L of CO gas measured at 33°C and 945 mmHg.

a) 0.263 g
b) 2.46 g
c) 3.80 g
d) 35.2 g

36. Which of the following gases will have the greatest density at the same specified temperature and pressure?

a) H_2
b) $CClF_3$
c) CO_2
d) C_2H_6

37. Determine the molar mass of chloroform gas if a sample weighing 0.389 g is collected in a flask with a volume of 102 cm^3 at 97°C. The pressure of the chloroform is 728 mmHg.

a) 187 g/mol
b) 121 g/mol
c) 112 g/mol
d) 31.6 g/mol

38. What is the molar mass of Freon-11 gas if its density is 6.13 g/L at STP?

a) 0.274 g/mol
b) 3.64 g/mol
c) 78.2 g/mol
d) 137 g/mol

40. A mixture of three gases has a total pressure of 1,380 mmHg at 298 K. The mixture is analyzed and is found to contain 1.27 mol CO_2, 3.04 mol CO, and 1.50 mol Ar. What is the partial pressure of Ar?

a) 0.258 atm
b) 301 mmHg
c) 356 mmHg
d) 5,345 mmHg

41. A sample of hydrogen gas was collected over water at 21°C and 685 mmHg. The volume of the container was 7.80 L. Calculate the mass of $H_2(g)$ collected. (Vapor pressure of water = 18.6 mmHg at 21°C.)

a) 0.283 g
b) 0.572 g
c) 0.589 g
d) 7.14 g

42. A 0.271 g sample of an unknown vapor occupies 294 mL at 140°C and 847 mmHg. The empirical formula of the compound is CH_2. What is the molecular formula of the compound?

a) CH_2
b) C_2H_4
c) C_3H_6
d) C_4H_8

43. How many liters of chlorine gas at 25°C and 0.950 atm can be produced by the reaction of 12.0 g of MnO_2?

$$MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

a) 5.36×10^{-3} L
b) 0.138 L
c) 0.282 L
d) 3.55 L

Test bank chapter (7)

Choose the most correct answer

1. The lowest energy state of an atom is referred to as its
   a) bottom state.
   b) **ground state.**
   c) fundamental state.
   d) original state.

2. All s orbitals are
   a) shaped like four-leaf clovers.
   b) dumbbell-shaped.
   c) spherical.
   d) triangular.

3. $[He]2s^2 2p^2$ is the electron configuration of which element?
   a) Beryllium Be
   b) Boron B
   c) carbon C
   d) nitrogen N

4. What are the valence electrons of vanadium (V)?
   a) $4s^2$
   b) $3d^3$
   c) $4s^2 3d^3$
   d) $3d^5$

5. What are the valence electrons of gallium Ga?
   a) $4s^2$
   b) $3d^3$
   c) $4s^2 4p^1$
   d) $3d^5$

6. The electron configuration of a neutral atom is [Ne] $3s^2 3p^1$. The four quantum numbers of the last electron are:

   a) (2, 1, -1, +1/2)
   b) (3, 3, -1 , +1/2)
   c) (3, 0, -1, +1/2)
   d) (3, 1, -1, +1/2)

7. How many unpaired electrons does chromium (Cr) have?
   a) 0
   b) 2
   c) 4
   d) **6**

8. How many unpaired electrons does selenium (Se) have?
   a) 0
   b) **2**
   c) 4
   d) 6

9. What is the maximum number of orbitals described by the quantum numbers: $n = 3$ $l = 2$
a) 1
b) 3
c) **5**
d) 9

10. What is the maximum number of orbitals described by the quantum numbers: $n = 4$
a) 7
b) 14
c) **16**
d) 48

11. The maximum number of electrons that can occupy an energy level described by the principal quantum number, *n*, is
a) $n + 1$
b) $2n$
c) $2n^2$
d) n^2

12. A possible set of quantum numbers for the last electron added to complete an atom of sodium Na in its ground state is

a) $n = 3$, $l = 1$, $m_l = 0$, $m_s = ½$
b) $n = 3$, $l = 0$, $m_l = 0$, $m_s = ½$
c) $n = 2$, $l = 1$, $m_l = -1$, $m_s = ½$
d) $n = 2$, $l = 0$, $m_l = -1$, $m_s = ½$

13. The ground-state electron configuration of a calcium atom is

a) $[Ne]3s^2$
b) $[Ne]3s^23p^6$
c) $[Ar]4s^13d^1$
d) $[Ar]4s^2$

14. Which one of the following sets of quantum numbers is not possible?

	n	l	m_l	m_s
Row 1	4	3	−2	+1/2
Row 2	3	2	−3	−1/2
Row 3	3	0	0	+1/2
Row 4	4	1	1	−1/2
Row 5	2	0	0	+1/2

a) Row 1
b) Row 2
c) Row 3
d) Row 4

15. The number of orbitals in a *d* subshell is

a) 1
b) 3
c) **5**
d) 7

16. Which ground-state atom has an electron configuration described by the following *orbital diagram*?

[Ar] ↑↓   ↑↓ ↑↓ ↑↓ ↑↓ ↑↓   ↑↓ ↑ ↑

4s   3d   4p

a) phosphorus
b) germanium
c) selenium
d) tellurium

17. A ground-state atom of nickel has ___ unpaired electrons and is _____.

a) 0, diamagnetic
b) 6, diamagnetic
c) 3, paramagnetic
d) 2, paramagnetic

18. What is the frequency (s^{-1}) of electromagnetic radiation that has a wavelength of 0.53 m?

a) **5.7×10^8**
b) 1.8×10^{-9}
c) 1.6×10^8
d) 1.3×10^{-33}

Explanation: The frequency and wavelength of electromagnetic radiation are related by the equation $c = \lambda\nu$, where c is the speed of light (=3.00 x 108 m/s), λ is the wavelength in m and ν is the frequency is s^{-1} or Hz. The frequency can be calculated by rearranging the above formula to get $\nu = c/\lambda = 3 \times 10^8 / 0.53 = 5.7 \times 10^8\ s^{-1}$

19. The energy of a photon of light is____ proportional to its frequency and _____ proportional to its wavelength.

a) directly, directly
b) inversely, inversely
c) inversely, directly
d) directly, inversely

20. The wavelength of a photon of energy 5.25×10^{-19} J is __________m.

a) 2.64×10^6
b) **3.79×10^{-7}**
c) 2.38×10^{23}
d) 4.21×10^{-24}

Explanation: The wavelength and energy are related by the formula $E = hc/\lambda$, where h (6.626×10^{-34} Js) is Planck's constant , c is the speed of light (3.00×10^8 m/s) and λ is the wavelength in meters. The wavelength can then be calculated by rearranging the above formula as follows: $\lambda = hc/E = 6.63\times10^{-34} \times 3\times10^8 / 5.25\times10^{-19} = 3.79 \times 10^{-7}$ m

21. What is the frequency (s^{-1}) of a photon of energy 4.38×10^{-18}J?

a) 438
b) 1.45×10^{-16}
c) **6.61×10^{15}**
d) 2.30×10^{7}

Explanation: The frequency ν of this photon can be calculated by rearranging the equation E = h ν where E is the energy, h = Planck's constant and ν = frequency in s^{-1}. $\nu = E/h = 4.38 \times 10^{-18} / 6.63 \times 10^{-34} = 6.61 \times 10^{15}$

22. An electron is a Bohr hydrogen atom has energy of -1.362×10^{-19}J. The value of n for this electron is __________.

a) 1
b) 2
c) 3
d) **4**

Explanation: The energy of an electron in a particular energy state in the hydrogen atom can be calculated by using the formula $E = (-2.18 \times 10^{-18}\ J)/n^2$, where n is the principal quantum number for the energy state. The value of n can be found by rearranging the above formula as follows:

$$n = \sqrt{\frac{-2.18\times10^{-18}\ J}{-1.362\times10^{-19}\ J}} = 4$$

19. The n = 2 to n = 6 transition in the Bohr hydrogen atom corresponds to the __ of a photon with a wavelength of__ nm.

a) emission, 411
b) **absorption, 411**
c) absorption, 657
d) emission, 389

Explanation: There are 2 parts to this question. Since the electron is moving from a smaller value of n (ni) to a larger value of n (n_f), it must be absorbing energy. The wavelength responsible for this transition can be calculated by using the formula: $E = R_H (1/n_i^2 - 1/n_f^2)$ & $E = hc/\lambda$

20. How many quantum numbers are necessary to designate a particular electron in an atom __________?

a) 3
b) **4**
c) 2
d) 1

21. The __________ quantum number defines the shape of an orbital.

a) spin
b) magnetic
c) principal
d) **angular**

22. There are __________ orbitals in the third shell

a) 25
b) 4
c) **9**
d) 16

Explanation: The number of orbitals in a shell is easily calculated by the formula # of orbitals = n^2 where n = principal quantum number, which is 3 in this case.

23. The angular quantum number is 3 in ___________ orbitals.

a) s
b) p
c) **d**
d) f

24. The n = 1 shell contains ___________p orbitals. All the other shells contain ___________p orbitals.

a) 3, 6
b) **0, 3**
c) 6, 2
d) 3, 3

Explanation: If n = 1, then the only possible value of ℓ is 0 which means that n = 1 can contain only s orbitals. When n > 1, the value of ℓ = 1 is possible making the existence of 3 p orbitals possible.

25. The principal quantum number of the first d subshell is ___________.

a) 1
b) 2
c) **3**
d) 4

26. The total number of orbitals in a shell is given by ___________.

a) L^2
b) $\mathbf{n^2}$
c) 2n
d) 2n + l

28. Each p-subshell can accommodate a maximum of ___________electrons.

a) **6**
b) 2
c) 10
d) 3

Explanation: There are 3 different p orbitals: px, py and pz. Each of these can contain 2 electrons leading to the maximum number of electrons as 6.

29. Each p-subshell can accommodate a maximum of __________electrons.

a) **6**
b) 2
c) 10
d) 3

Explanation: There are 3 different p orbitals: p_x, p_y and p_z. Each of these can contain 2 electrons leading to the maximum number of electrons as 6.

30. The 3p subshell in the ground state of atomic xenon contains __________ electrons.

a) 2
b) **6**
c) 36
d) 10

Explanation: Since Xe is a noble gas, its subshells will be completely filled regardless of their principal quantum number. Thus the 3p subshell will contain 6 electrons.

31. $[Ar]4s^2 3d^{10} 4p^3$ is the electron configuration of a(n) __________ atom.

a) **As**
b) V
c) P
d) Sb

Explanation: The easiest way to answer this question is to count the total number of electrons and find which element that number corresponds to. The total number of electrons is = 18 (for the Ar) + 2 +10 + 3 = 33 which corresponds to As.

32. The principal quantum number for the outermost electrons in a Br atom in the ground state is __________.

a) 2
b) 3
c) **4**
d) 5

Explanation: The electronic configuration of bromine is $[Ar]3d^{10}4s^2 4p^5$ shows that the outermost electrons are in the s and p orbitals in the 4th energy level making the principal quantum number = 4.

33. All of the __________ have a valence shell electron configuration ns^1.

a) noble gases
b) halogens
c) chalcogens
d) **alkali metals**

34. Which one of the following is correct?

a) $\nu + \lambda = c$
b) $\nu/\lambda = c$
c) $\lambda = c\nu$
d) $\nu\lambda = c$

35. In the Bohr model of the atom, __________.

a) electrons travel in circular paths called orbitals
b) electrons can have any energy
c) **electron energies are quantized**
d) electron paths are controlled by probability

36. Which one of the following is not a valid value for the magnetic quantum number of an electron in a 5d subshell?

a) 2
b) 3
c) 0
d) 1

Explanation: For an electron in the 5d subshell the value of $\ell = 2$ and the magnetic quantum number $m\ell$ can have values from –l,…0,…+l, meaning $m\ell$ could not have a value = 3.

37. Which of the subshells below do not exist due to the constraints upon the angular quantum number?

a) 2s
b) **2d**
c) 2p
d) none of the above

Explanation: The values of the azimuthal quantum number "l" are decided by the values of the principal quantum number "n". The values of l will only be from 0…n – 1. Thus for n = 2, only the values of 0 and 1 will be possible for ℓ, which means that only the 2s and 2p orbitals will be possible.

38. An electron cannot have the quantum numbers n = ________, l = _________, $m\ell$ = _________.

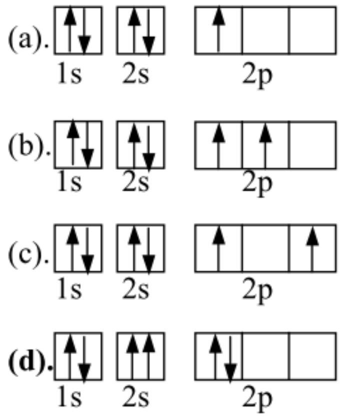
a) 2, 0, 0
b) 2, 1, -1
c) 3, 1, -1
d) 1, 1, 1

Explanation: The values of 1, 1, 1 would be impossible since if n = 1, the only value of ℓ would be = 0.

39. Which quantum number determines the energy of an electron in a hydrogen atom?

a) n
b) n and ℓ
c) ml
d) ℓ

39. Which electron configuration represents a violation of the Pauli exclusion principle?



Explanation: According to the Pauli Exclusion Principle no two electrons in an atom cannot have the same 4 quantum numbers. The2 electrons in the 2s orbital have the same value for their ms which is not allowed. **(d)**

40. Which of the following is a valid set of four quantum numbers? (n, ℓ, mℓ, ms)

a) **2, 0, 0, + ½**
b) 2, 2, 1, - ½
c) 1, 0, 1, + ½
d) 2, 1, +2, + ½

Explanation: Here is why only option (a) is the correct answer: In option (b), ℓ =2 which is not allowed, in (c) mℓ ≠ 1 since l = 0 and in (d) m ℓ > l which are all not allowed.

41. Which of the following is not a valid set of four quantum numbers? (n, ℓ, mℓ, ms)

a) 2, 0, 0, + ½
b) 2, 1, 0, - ½
c) 1, 1, 0, + ½
d) 1, 0, 0, + ½

Explanation: Since n can never be equal to ℓ, option c is the only set that is not valid.

Test bank chapter (8)

Choose the most correct answer

1. Elements in the modern version of the periodic table are arranged in order of increasing __________.

a) oxidation number
b) atomic mass
c) average atomic mass
d) **atomic number**

Explanation: The older version of the periodic table had the elements arranged in order of increasing atomic mass, but the modern version of the periodic table is based on the increasing order of atomic number.

2. The first ionization energies of the elements __________ as you go from left to right across a period of the periodic table, and __________ as you go from the bottom to the top of a group in the table. Shaped like four-leaf clovers.

a) **increase, increase**
b) increase, decrease
c) decrease, increase
d) decrease, decrease

Explanation: The ionization energies (IE s) of elements increase to the right in a row since larger amounts of energy need to be supplied to remove an electron. The elements become more non-metallic making it harder to remove an electron.

3. The __________ have the most negative electron affinities

a) alkaline earth metals
b) alkali metals
c) **halogens**
d) transition metals

Explanation: The electron affinity of an element is defined as the energy change that occurs when an electron is added to a gaseous atom. The halogens have the most negative electron affinities indicating that they are most comfortable accepting an electron. The formation of an anion essentially gives the halogen atom the electron configuration of the nearest noble gas. The negative sign here indicates that the addition of an electron to the halogens results in energy being released by the halogen atom.

4. Na reacts with element X to form an ionic compound with the formula Na_3X. Ca will react with X to form _____.

a) CaX_2
b) CaX
c) Ca_2X_3
d) **Ca_3X_2**

Explanation: The element X must have 3 negative charges for it to form the compound Na_3X, since each Ca has 2 positive charges, the formula of the compound formed by the reaction of Ca and X would have to be Ca_3X_2.

5. 16. Atomic radius generally increases as we move __________.

a) **down a group and from right to left across a period**
b) up a group and from left to right across a period
c) down a group and from left to right across a period
d) there is no trend

6. 18. The atomic radius of main-group elements generally increases down a group because __________.

a) effective nuclear charge increases down a group
b) effective nuclear charge decreases down a group
c) both effective nuclear charge increases down a group and the principal quantum number of the valence orbitals increases

d) **the principal quantum number of the valence orbitals increases**

7. Which of the following correctly lists the five atoms in order of increasing size (smallest to largest)?

a) O < F < S < Mg < Ba
b) **F < O < S < Mg < Ba**
c) F < O < S < Ba < Mg
d) O < F < S < Ba < Mg

Explanation: Fluorine and oxygen are in the same period (#2) and next to each other with F being the smallest of these 5 atoms. Ba is in group 2A and in row 6 (farthest "down" a group) and is the largest of the atoms. Mg is in group 2A and in the third period and hence will be bigger than F, O and S. Even though S is in the same period as Mg it is in group 6A making it smaller than Mg.

8. Which of the following correctly lists the five atoms in order of increasing size (smallest to largest)? 0

a) F < K < Ge < Br < Rb
b) F < Ge < Br < K < Rb
c) F < K < Br < Ge < Rb
d) **F < Br < Ge < K < Rb**

Explanation: Fluorine is in group 7A and period 2 making it the smallest of the 5 atoms here. Br is also in group 7A but is in period 4 making it larger than F, Ge is in group 4A and also in period 4 but is to the left of Br making it larger than Br. K and Rb are both in group 1A but K is in period 4 and Rb is in period 5, making the Rb atom the largest of all the 5 atoms.

9. Of the following atoms, which has the largest first ionization energy?

a) Br
b) **O**
c) C
d) P

Explanation: The ionization energy (IE) typically increases from left to right in a period and decreases from top to bottom in a group. Thus for C and O which are in the same period, O will have the larger IE. Br is in period #4 and will have the lowest IE out of these 4 elements.

10. Of the following elements, which has the largest first ionization energy?

a) Na
b) A1
c) Se
d) **C1**

Explanation: The ionization energy (IE) typically increases from left to right in a period and decreases from top to bottom in a group. Na, Al and Cl are all in period 3 with the chlorine atom to the farthest right and will have the highest IE.

11. Which ion below has the largest radius?

a) $C1^-$
b) K^+
c) Br^-
d) F^-

Explanation: Typically cations are smaller than their parent atoms while anions are larger than the parent atoms. Of the atoms here, the Br atom would be the largest as it is farthest down the group and hence its anion also will be the largest ion.

12. The ion with the smallest radius is __________.

a) Br^-
b) $C1^-$
c) O^{2-}
d) F^-

Explanation: Typically anions are larger than the parent atoms. Of the atoms here, the F atom would be the smallest as it is farthest down the group and hence its anion also will be the smallest ion

13.Which of the following is an isoelectronic series?

a) B^{5-}, Si^{4-}, As^{3-}, Te^{2-}
b) **O^{2-}, F^-, Ne, Na^+**
c) S, C1, Ar, K
d) None of the above

Explanation: Isoelectronic series contain a combination of atoms and ions or only ions with the same number of electrons. Here the series containing the O^{2-}, F^-, Ne and Na^+ is the only one where all atoms/ions contain 10 electrons.

14. __________ is isoelectronic with argon and __________ is isoelectronic with neon.

a) **$C1^-$, F^-**
b) $C1^-$, $C1^+$
c) F^+, F^-
d) Ne, Kr^+

Explanation: The Cl^- ion has 18 electrons and is isoelectronic with argon while the F^- ion has 10 electrons making it isoelectronic with neon.

15. Chlorine is much more apt to exist as an anion than is sodium. This is because __________.

a) chlorine is bigger than sodium
b) chlorine has a greater ionization energy than sodium does
c) **chlorine has a greater electron affinity than sodium does**
d) chlorine is a gas and sodium is a solid

16. The alkaline earth metals are found in _______ of the periodic table.
a) Group 1A
b) Group 2A
c) Group 7A
d) Group 8A

17.How many *valence electrons* does a boron atom (B) have?

a) 1
b) 3
c) 5
d) 7

18. Which ion is *isoelectronic* with Ar?

a) Ni^{2+}
b) F^-
c) Br^-
d) K^+

19. Which of these choices is the electron configuration of the iron (III) ion (Fe^{3+})?

a) $[Ar]4s^2 3d^3$
b) $[Ar]4s^1 3d^5$
c) $[Ar]3d^5$
d) $[Ar]3d^6$

20. In what group of the periodic table is the element with the electron configuration $[Ar]4s^2 3d^{10} 4p^1$?

a) 1A
b) 2A
c) 3A
d) 4A

21. How many *valence electrons* does a tin (Sn) atom have?

a) 2
b) 4
c) 14
d) 36

22. Which of these ground-state ions has unpaired electrons?

a) P^{3-}
b) V^{5+}
c) S^{2-}
d) Sc^{2+}

23. Consider the element with the electron configuration $[Xe]6s^2 4f^7$. This element is

a) a representative element.
b) a lanthanide element.
c) a nonmetal.
d) an actinide element

24. If the radius of atom X is greater than the radius of atom Y, then it is also likely that

a) X has a larger electron affinity than Y does.
b) X has a larger first ionization energy than Y does.
c) X has greater metallic character than Y does.

25. Arrange these ions in order of increasing ionic radius: K^+, P^{3-}, S^{2-}, Cl^-.

	Increasing radius →
Row 1	$K^+ < Cl^- < S^{2-} < P^{3-}$
Row 2	$K^+ < P^{3-} < S^{2-} < Cl^-$
Row 3	$P^{3-} < S^{2-} < Cl^- < K^+$
Row 4	$Cl^- < S^{2-} < P^{3-} < K^+$
Row 5	$Cl^- < S^{2-} < K^+ < P^{3-}$

a) Row 1
b) Row 2
c) Row 3
d) Row 4

26. - Selenium ($_{34}Se$) element is

a) a nonmetal
b) found in group 6A
c) found in period 2
d) **both a and b**

28. The outer electron configuration of the noble gases is

a) **$ns^2\ np^6$**
b) $ns^2\ nd^{10}$
c) $ns^2\ np^4$
d) $ns^2\ np^8$

29. Which of the following species is isoelectronic with Cl^-

a) **K^+**
b) Na^+
c) O F^-
d) $^{2-}$

30. Gallium (Ga) element is found in the periodic table in

a) period 3, group 1B
b) period 3A, group 4
c) period 4, group 1A
d) **period 4, group 3A**

31. Titanium (Ti) element is found in the periodic table in

a) s-block
b) P-block
c) **d-block**
d) f-block

32. Write the electronic configuration for Co^{+2}

a) [Ar] $4s^2\ 3d^5$
b) **[Ar] $4s^2 3d^7$**
c) [Ar] $4s^1\ 3d^6$
d) [Ar] $4s^1\ 3d^5$

33. Select the correct order of radius of the two ions

a) $A^+ > A^-$
b) $A^- < A$
c) $A^{2+} > A^+$
d) **$A^{2+} < A^+$**

34. Two ions are referred to as isoelectronic if they have the same number of

a) **electrons.**
b) protons.
c) atoms.
d) neutrons.

35. The energy required to remove an electron from an atom in its ground state is known as the

a) potential energy.
b) activation energy.
c) electron affinity.
d) **ionization energy.**

36. Which will have the highest ionization energy?

a) C
b) **N**
c) O
d) B

37. Order the following (N^{-3} , Li^+ , C , O^{2-})according to increasing atomic/ionic radius.

a) $C < Li^+ < O^{2-} < N^{3-}$
b) $N^{3-} < O^{2-} < C < Li^+$
c) $Li^+ < C < N^{3-} < O^{2-}$
d) **$Li^+ < C < O^{2-} < N^{3-}$**

Test bank chapter (9)

Choose the most correct answer

1. The two types of chemical bonds commonly found in compounds are:

a) doric and covalent.
b) ionic and electrolytic.
c) **ionic and covalent.**
d) electrolytic and compound.

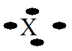
2. The electrons used by atoms to form chemical bonds are the:
a) core electrons.
b) valence electrons.
c) lone pair electrons.
d) unpaired electrons.

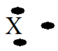
3. “Atoms tend to gain, lose, or share electrons until they are surrounded by eight valence electrons” is a statement of:
a) the rule of octaves.
b) the double quartet rule.
c) the eight electron rule.
d) the octet rule.

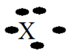
4. When a transition metal atom becomes a +1 ion, the electron lost usually comes from what type of orbital?
a) p
b) f
c) d
d) s

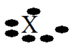
5. A molecule of CS_2 contains
a) two single bonds.
b) two double bonds.
c) one single bond and one double bond.
d) one single bond and one triple bond.

6. An atom in the ground state has atomic number Z=5. Choose the correct electron-dot structure which represents this atom? **ANS. B**

(A) 

(B) 

(C) 

(D) 

7. Which compound below contains an atom that is surrounded by more than an octet of electrons?

a) PF_5
b) CH_4
c) NBr_3
d) OF_2

8. Which choice below correctly lists the elements in order of increasing electronegativity?

a) **C < N < O < F**
b) N < C < O < F
c) N < C < F < O
d) C < N < F < O

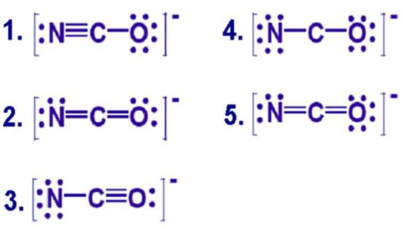
9. Which atom sometimes violates the octet rule?

a) C
b) N
c) O
d) S

10. How many resonance structures can be drawn for NO^{3-}?

a) 1
b) 2
c) 3
d) 4

11. Considering formal charge, what is the preferred Lewis structure of NCO^-? ANS.1



12. In Lewis structure of $(SO4)^{-2}$structure the correct formal charge on sulfur (S) is:

a) +2
b) -2
c) +1
d) 0

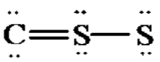
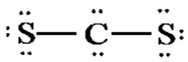
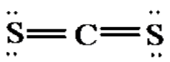
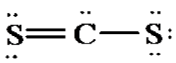
13. Which of these pairs of elements would be most likely to form an ionic compound?

a) Cl and I
b) Al and K
c) **Cl and Mg**
d) C and S

14. Which of these covalent bonds is the most polar (i.e., highest percent ionic character)?

a) Al — I
b) Si — I
c) Al — Cl
d) Si — Cl

15. The Lewis structure for CS_2 is: ANS. c

a) 
b) 
c) 
d) 

16. The number of lone electron pairs in the N_2 molecule is ___.
a) 1
b) 2
c) 3
d) 4

17. Classify the O-H bond in CH_3OH as ionic, polar covalent, or nonpolar covalent.
a) Ionic
b) polar covalent
c) nonpolar covalent
d) none of the above

18. The Lewis structure for a chlorate ion, ClO_3^{-1}, should show ____ single bond(s), ____ double bond(s), and ____ lone pair(s).

a) 2, 1, 10
b) 3, 0, 9
c) 2, 1, 8
d) **3, 0, 10**

19. The number of resonance structures for the sulfur dioxide (SO_2) molecule that satisfy the octet rule is

a) 1
b) 2
c) 3
d) None of these.

20. What is the formal charge on the oxygen atom in N_2O (the atomic order is N-N-O)?

a) 0
b) +1
c) **-1**
d) -2

21. Which of these substances will display an incomplete octet in its Lewis structure?

a) CO_2
b) Cl_2
c) ICl
d) NO

22. There are __________ paired and __________ unpaired electrons in the Lewis symbol for a phosphorus atom (P).

a) 4, 2
b) 2, 4
c) 4, 3
d) 2, 3

Explanation: Read the question carefully here, you are being asked for how many valence electrons are paired and how many are unpaired. The abbreviated electron configuration of the P atom is given by [Ne] $3s^2 3p^3$. The outermost electrons would be arranged as 2 electrons paired and 3 electrons unpaired as shown below:

23. Based on the octet rule, magnesium (Mg) most likely forms a __________ ion.

a) Mg^{2-}
b) Mg^{2+}
c) Mg^{6-}
d) Mg^{6+}

Explanation: According to the octet rule the Mg atom will achieve an octet by losing its 2 outermost electrons and thus gaining 2+ charges. Since Mg is located in the alkali metal group it will lose electrons rather than gaining them.

24. Based on the octet rule, phosphorus (P) most likely forms a __________ ion.

a) P^{3+}
b) P^{5-}
c) P^{5+}
d) P^{3-}

Explanation: According to the octet rule the phosphorus atom should gain 3electrons, thus gaining 3 negative charges and forming the phosphide ion.

25- The only noble gas without eight valence electrons is __________.

a) Ar
b) Ne
c) He
d) Kr

Explanation: The noble gases are characterized by the presence of eight electrons in their outermost shell with one notable exception of Helium. Since He has only 2 electrons it can never have 8 in its outermost shell.

26- What is the maximum number of double bonds that a hydrogen atom (H) can form?

a) 0
b) 1
c) 2
d) 3

Explanation: Each hydrogen atom has a single electron in its valence shell and as a result can form only one bond. It cannot form a double bond as it does not have the necessary electrons to share.

28. What is the maximum number of double bonds that a carbon atom (C) can form?

a) 4
b) 1
c) 2
d) 0

Explanation: Each carbon atom has 4 valence electrons that it can share with other atoms. Since each double bond corresponds to a pair of electrons, the carbon atom can form only 2 double bonds.

29. Given the electronegativities below, which covalent single bond is most polar?

Atom	H	C	N	O
Electronegativity	2.1	2.5	3.0	3.5

a) C-H
b) N-H
c) O-H
d) O-N

Explanation: Bond polarity can be judged based on the differences between the electronegativities of the atoms involved. Of the available choices, the bond between O and H will have the largest electronegativity difference making it the most polar bond in this group.

30. The ion ICL_4^- has __________ valence electrons.

a) 34
b) 36
c) 35
d) 28

Explanation: valence electrons A = (7×1) +(7×1) +1 = 36

31- Electronegativity __________ from left to right within a period and _____from top to bottom within a group.

a) decreases, increases
b) increases, increases
c) stays the same, increases
d) **increases, decreases**

Explanation: Atomic size decreases from the left to the right in a period thus making it easier for the nuclei to attract electrons towards themselves resulting in an increase in the electronegativity. On the other hand atomic size increases down a group making it harder for the nuclei to attract the valence electrons towards themselves resulting in a decrease in electronegativity.

32. The Lewis structure of PF_3 shows that the central phosphorus atom has ___nonbonding and__ bonding electron pairs.

a) 2, 2
b) 1, 3
c) 3, 1
d) 1, 2

33. Which of the following molecules contains both ionic and covalent bonds?

a) C_5H_{12}
b) **$NaClO_4$**
c) $CaCl_2$
d) H_2O

34. The ability of an atom in a molecule to attract electron density to itself is termed

a) Electronegativity
b) Electron affinity
c) Diamagnetism
d) Ionization energy

35- the most polar bond is

a) Br-H
b) I-H
c) Cl-H
d) H-H

Test bank chapter (14)

Choose the most correct answer

1. Which is the correct equilibrium constant expression for the following reaction?
$Fe_2O_3(s) + 3H_2(g) \leftrightarrow 2Fe(s) + 3H_2O(g)$

a) $K_c = [Fe_2O_3]\,[H_2]^3/[Fe]^2[H_2O]^3$
b) $K_c = [H_2]/[H_2O]$
c) $K_c = [H_2O]^3/[H_2]^3$
d) $K_c = [Fe]^2[H_2O]^3/[Fe_2O_3]\,[H_2]^3$

2. The following reactions occur at 500 K. Arrange them in order of increasing tendency to proceed to completion (least →greatest tendency).

1. $2NOCl \leftrightharpoons 2NO + Cl_2$ $K_p = 1.7 \times 10^{-2}$
2. $2SO_3 \leftrightharpoons 2SO_2 + O_2$ $K_p = 1.3 \times 10^{-5}$
3. $2NO_2 \leftrightharpoons 2NO + O_2$ $K_p = 5.9 \times 10^{-5}$

a) $2 < 1 < 3$
b) $1 < 2 < 3$
c) $2 < 3 < 1$
d) $3 < 2 < 1$

3.Calculate K_p for the reaction $2NOCl(g) \leftrightarrow 2NO(g) + Cl_2(g)$ at 400°C if K_c at 400°C for this reaction is 2.1×10^{-2}.

a) 2.1×10^{-2}
b) 1.7×10^{-3}
c) 0.70
d) 1.2

4. For the reaction $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$, $K_c = 50.2$ at 445ºC. If $[H_2] = [I_2] = [HI] = 1.75 \times 10^{-3}$ M at 445ºC, which one of these statements is *true*?

a) The system is at equilibrium, thus no concentration changes will occur.
b) The concentrations of HI and I_2 will increase as the system approaches equilibrium.
c) The concentration of HI will increase as the system approaches equilibrium.
d) The concentrations of H_2 and HI will fall as the system moves toward equilibrium.

5.For the following reaction at equilibrium, which choice gives a change that will shift the position of equilibrium to favor formation of more products?
$2NOBr(g) \leftrightarrow 2NO(g) + Br_2(g)$, $\Delta H^\circ_{rxn} = 30$ kJ/mol

a) Increase the total pressure by decreasing the volume.
b) Add more NO.
c) Remove Br_2.
d) Lower the temperature.

6 - For the following reaction at equilibrium in a reaction vessel, which one of these changes would cause the Br_2 concentration to *decrease*?

$2NOBr(g) \leftrightarrow 2NO(g) + Br_2(g)$, $\Delta H^o_{rxn} = 30$ kJ/mol

a) Increase the temperature.
b) Remove some NO.
c) Add more NOBr.
d) Compress the gas mixture into a smaller volume.

7. For the reaction at equilibrium $2SO_3 \leftrightarrow 2SO_2 + O_2$ ($\Delta H^o_{rxn} = 198$ kJ/mol), if we increase the reaction temperature, the equilibrium will

a) shift to the right.
b) shift to the left.
c) not shift.
d) The question cannot be answered because the equilibrium constant is not given.

8. For the equilibrium reaction $2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$, $\Delta H^o_{rxn} = -198$ kJ/mol. Which one of these factors would cause the equilibrium constant to *increase*?

a) Decrease the temperature.
b) Add SO_2 gas.
c) Remove O_2 gas.
d) Add a catalyst.

9. The reaction $2SO_3(g)$ $2SO_2(g) + O_2(g)$ is endothermic. If the temperature is increased

a) more SO_3 will be produced.
b) K_c will decrease.
c) no change will occur in K_c .
d) K_c will increase.

10.If a catalyst is added to a chemical reaction, the equilibrium yield of a product will be … , and the time taken to come to equilibrium will be ……than before.

a) higher; less
b) lower; the same
c) higher; the same
d) the same; less

11- For the reaction , $N_2(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g)$

Kc = 0.0600 at a certain temperature. In an equilibrium mixture of the three gases, $[NH_3] = 0.242$ M and $[H_2] = 1.03$ M. What is the concentration of N2 in this system?

a) 3.9 M
b) 3.2×10^{-3} M
c) 0.89 M
d) 1.4×10^{-2} M

11. Consider the reaction ,$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$.

If an equilibrium mixture of these three substances is compressed, equilibrium will _____, because _____.

a) shift to the right; higher pressure favors fewer moles of gas
b) shift top the right; higher pressure favors more moles of gas
c) **shift to the left; higher pressure favors fewer moles of gas**
d) shift to the left; higher pressure favors more moles of gas

12- Consider the equilibrium system; $C(s) + CO2(g) \rightleftharpoons 2CO(g)$.

If more C(s) is added, the equilibrium will ____; if CO is removed the equilibrium will ____.

a) shift to the left; shift to the left
b) shift to the right; shift to the right
c) **shift to the right; shift to the left**
d) be unchanged; shift to the left

8. Consider the exothermic reaction at equilibrium: $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$, If the system is cooled, the equilibrium will _____, because _____.

a) shift to the left; decreased temperature favors an exothermic reaction
b) shift to the right; decreased temperature favors an exothermic reaction
c) shift to the right; decreased temperature favors an endothermic reaction
d) shift to the left; decreased temperature favors an endothermic reaction

9. A large value of the equilibrium constant indicates that when the reaction reaches equilibrium, mostly ______ will be present.

a) reactants
b) products
c) catalysts
d) shrapnel

10. When equilibrium is achieved

a) $Q > K$
b) $Q < K$
c) $Q = K$
d) $Q2 = K$

11. CO_2 + $H_2 \leftrightarrow$ CO + H_2O

If all species are gases and H_2 is added, the amount of CO present at equilibrium will:

a) increase.
b) decrease.
c) remain unchanged.
d) disappear.

12. CO_2 + $H_2 \leftrightarrow$ CO + H_2O

If all species are gases and H_2O is added, the amount of CO present at equilibrium will:

a) increase.
b) decrease.
c) remain unchanged.
d) disappear.

13. CO_2 + $H_2 \leftrightarrow$ CO + H_2O

If the reaction is endothermic and the temperature is raised, the amount of CO present will:

a) increase.
b) decrease.
c) remain unchanged.
d) disappear.

14. CO_2 + $H_2 \leftrightarrow$ CO + H_2O

If all species are gases and the container is compressed, the amount of CO present will:

a) increase.
b) decrease.
c) remain unchanged.
d) disappear.

15. CO_2 + $H_2 \leftrightarrow$ CO + H_2O

If all species are gases and the container is compressed, the amount of CO present will:

a) increase.
b) decrease.
c) remain unchanged.
d) disappear.

16. What is K_P in terms of K_c *for the following reaction* ?2 NO(g) + O_2(g) $\leftrightarrow$2 NO_2(g)

a) Kp = KcRT
b) Kp = Kc/RT
c) Kp = KcR/T
d) Kp = $Kc/(RT)^2$

17. What is the correct equilibrium constant expression for the reaction: **ANS. 3**

$$P_4(s) + 6\,Cl_2(g) \longrightarrow 4\,PCl_3(l)$$

1. $\dfrac{[PCl_3]^4}{[P_4][Cl_2]^6}$

2. $\dfrac{[PCl_3]^4}{[Cl_2]^6}$

3. $\dfrac{1}{[Cl_2]^6}$

4. $\dfrac{[Cl_2]^6}{[PCl_3]^4}$

5. $\dfrac{[4\,PCl_3]^4}{[P_4][6\,Cl_2]^6}$

18. The equation relating Kp and Kc is

a) **$K_p = K_c (RT)^{\Delta n}$**
b) $K_p = K_c RT^{\Delta n}$
c) $K_c = K_p RT^{\Delta n}$
d) $K_c = K_p (RT)^{\Delta n}$

19. K_p will be equal to K_c if

a) $\Delta n = 1$
b) **$\Delta n = 0$**
c) $RT = 0$
d) $\Delta n = -1$

20. Consider the reversible reaction at equilibrium at 392 ^oC:

$$\mathbf{2A(g) + B(g) \rightleftarrows C(g)}$$

The partial pressures are found to be: A: 6.70 atm, B: 10.1 atm, C: 3.60 atm. Evaluate K_p for this reaction.

a) 7.94×10^{-3}
b) 0.146
c) 0.0532
d) 54.5

20. Which of the following will result in an equilibrium shift to the right?

$$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g) \quad DH° = –87.9 \text{ kJ/mol}$$

a) Increase temperature/increase volume
b) Increase temperature/decrease volume
c) Decrease temperature/increase volume
d) Decrease temperature/decrease volume

21.Which accurately reflects the changes in concentration that will occur if O_2 is added to disturb the equilibrium?

$$2\,NO(g) + O_2(g) \leftrightarrow 2\,NO_2(g)$$

	[NO]	[O_2]	[NO]
a)	Increase	Increase	Increase
b)	Increase	Increase	Decrease
c)	Decrease	Decrease	Decrease
d)	Decrease	Increase	Increase

Test bank chapter (15)

Choose the most correct answer

1-What is the concentration of H^+ in a 2.5 M HCl solution?

a) 0
b) 1.3 M
c) 2.5 M
d) 5.0 M

2. What is the OH^- ion concentration in a 5.2×10^{-4} M HNO_3 solution?

a) 1.9×10^{-11} M
b) 1.0×10^{-7} M
c) 5.2×10^{-4} M
d) Zero

3. Calculate the H^+ ion concentration in lemon juice having a pH of 2.4.

a) 4.0×10^{-2} M
b) 250 M
c) 0.38 M
d) 4.0×10^{-3} M

4. Calculate the pH of a 6.71×10^{-2} M NaOH solution.

a) 12.83
b) 2.17
c) 11.82
d) 6.71

5. What is the pH of 0.0200 M aqueous solution of HBr ?

a) 1.00
b) 1.70
c) 2.30
d) 12.30

6. The pOH of a solution of NaOH is 11.30. What is the $[H^+]$ for this solution?

a) 2.0×10^{-3}
b) 2.5×10^{-3}
c) 5.0×10^{-12}
d) 4.0×10^{-12}

7. What is the pH of a 0.0400 M aqueous solution of KOH?

a) 12.60
b) 10.30
c) 4.00
d) 1.40

8. What is the approximate pH of a solution labeled 6×10^{-5} M HBr?

a) **4.2**
b) 4.5
c) 5.8
d) 9.8

9. If the pH = 2 for an HNO_3 solution, what is the concentration of HNO_3?

a) 0.10
b) 0.20
c) **0.010**
d) 0.020

10. A solution in which $[H^+] = 10^{-8}$ M has a pH of ___ and is ___.

a) 8, acidic
b) 6, basic
c) -6, basic
d) 8, basic

11. Which of the following solutions has the lowest pH at 25°C? (No calculations required.)

a) 0.2 M NaOH
b) 0.2 M NH_3
c) 0.2 M HCl
d) pure water

12. Calculate the pH of a 3.5×10^{-3} M HNO_3 solution.

a) -2.46
b) 0.54
c) 2.46
d) 3.00

13. The pH of 2.6×10^{-2} M KOH is

a) 12.41
b) 15.59
c) 2.06
d) 7.00

14. What is the $[H^+]$ ion in a 4.8×10^{-2} M KOH solution?

a) 2.8×10^{-13} M
b) 1×10^{-7} M
c) 4.8×10^{-11} M
d) 4.8×10^{-2} M

15. What is the $[OH^-]$ ion in a 5.2×10^{-4} M HNO_3 solution?

a) 1.9×10^{-11} M
b) 1.0×10^{-7} M
c) 5.2×10^{-4} M
d) zero

Test bank chapters (24 & 25)

Choose the most correct answer

1. $C_{10}H_{22}$ is the formula of an ….

a) alkane.
b) alkene.
c) alkyne.
d) aromatic hydrocarbon.

2. A molecule with the formula C_3H_8 is a

a) hexane
b) propane
c) decane
d) butane

3. Which compound below does not have geometric isomers (cis-trans isomers)?

a) 1-butene
b) 2-butene
c) 2-pentene
d) 3-hexene

4. The hybridization of carbon atoms in alkanes is

a) sp
b) sp^2
c) sp^3
d) sp^3d

5. Select the correct IUPAC name for $\boxed{CH_3CH_2\overset{CH_3}{\overset{|}{C}}HCH_2\overset{CH_3}{\overset{|}{C}}HCH_3}$

a) 1,1,3-trimethylpentane
b) 1-ethyl-1,3-dimethylbutane
c) 2,4-dimethylhexane
d) 3,5-dimethylhexane.

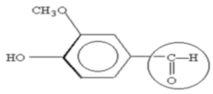
6 - An alkane with seven carbon atoms in a linear configuration is called a

a) hexene
b) heptene
c) hepylane
d) **heptane**

7. Which type of functional group does not include a carbonyl group in its structure?

a) carboxylic acid
b) ether
c) ketone
d) aldehyde

8. Vanillin is used as a flavoring agent. Identify the functional group circled.



a) aldehyde
b) ketone
c) carboxylic acid
d) Alcohol

9. The formula ($CH_3CH_2CH_2CH_2CH_2CH_2CH_2OH$) represents:

a) an alcohol
b) an alkene
c) an alkyne
d) an unsaturated hydrocarbon

10. Which of the following is a ketone?

a) $CH_3CH_2COCH_3$
b) CH_3CH_2CHO
c) CH_3OCH_3
d) CH_3CH_2COOH

11. The general formula for *alkenes* is

a) C_nH_{2n+2}
b) $C_{2n}H_{2n}$
c) C_nH_{n+2}
d) C_nH_{2n}

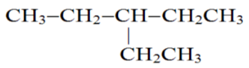
12- Which of these molecules is *unsaturated*?

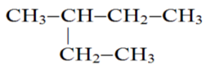
a) C_3H_8
b) CH_3OH
c) C_5H_{10}
d) CH_4

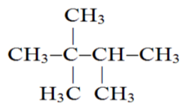
8. Which of these species is an aromatic compound?

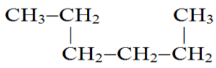
a) C_2H_2
b) C_6H_{12}
c) $C_6H_4Br_2$
d) C_5H_{10}

9. Which of these species are *structural isomers* of C_6H_{14}?

I. 

II. 

III. 

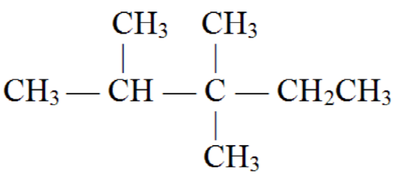
IV. 

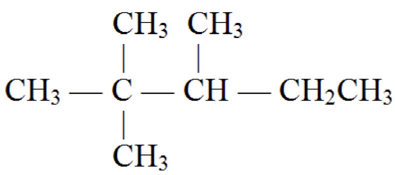
a) I and II
b) I and III
c) II and III
d) II and IV

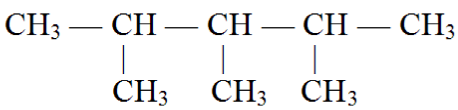
10. The compound that has a triple bond between one pair of carbon atoms is called

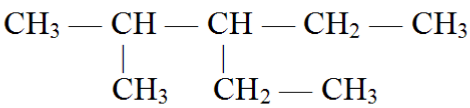
a) an alkane.
b) a chlorofluorocarbon.
c) an alkyne.
d) an alkene.

11. The correct structure for 2,3,3-trimethylpentane is

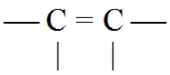
A. 

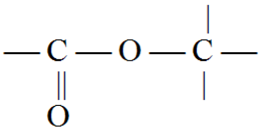
B. 

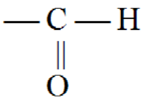
C. 

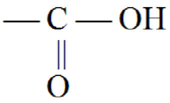
D. 

12. Which one of these structures represents an *ester* functional group?

A. 

B. 

C. 

D. 

13. The functional group (RCOR) is characteristic of organic ______

a) ketones
b) acids
c) aldehydes
d) esters

14. Which of the following hydrocarbons does not have isomers?

a) C_7H_{16}
b) C_6H_{14}
c) C_5H_{10}
d) C_3H_8

15. Which of the following does NOT exhibit geometric isomerism? (Hint: draw them!)

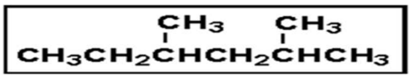
a) 4-octene
b) 2-pentene
c) 3-hexene
d) 1-hexene

16. For which of the compounds below are cis-trans isomers possible?

$CH_3CH{=}CH_2$	$CH_3CH{=}CHCH_2CH_3$	$CH_3CH{=}CHCH_3$
(1)	(2)	(3)

a) only 2
b) both 1 and 2
c) **both 2 and 3**
d) all three

17. Select the correct IUPAC name for



a) 1,1,3-trimethylpentane
b) 1-ethyl-1,3-dimethylbutane
c) 2,4-dimethylhexane
d) 3,5-dimethylhexane

18. A protein is:

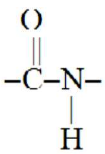
a) a polymer of amino acids
b) a fatty acid ester of glycerol
c) a polysaccharide
d) an addition polymer

19. A peptide bond (also called an amide bond) joins two amino acids together. What atoms are linked by this bond?

a) C — O
b) C — H
c) **C — N**
d) N — S

20.An amino acid is a compound that contains at least

a) one amino group and one amide group.
b) two amino groups and one carboxylic acid group.
c) one hydroxyl group and one methyl group.
d) one carboxylic acid group and one amino group

21. The functional group  found in proteins is called a (an)

a) amide.
b) carboxylic acid.
c) amine.
d) amino acid.

22. Which one of these choices is the general structural formula of an amino acid?

A. $R{-}CH_2{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}NH_2$

B. $R{-}CH_2{-}NH_2$

C. $\begin{array}{l} R{-}CH{-}OH \\ \quad\;| \\ \quad NH_2 \end{array}$

D. $\begin{array}{l} R{-}CH{-}NH_2 \\ \quad\;| \\ \quad COOH \end{array}$