**Chapter 0**

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| 1. | What is the proper formula for the ionic compound formed between Fe3+ and the hydroxide ion?     |  |  | | --- | --- | |  | (A) FeO3 | |  | (B) Fe3O | |  | (C) FeO | |  | (D) FeOH3 | |  | (E) Fe(OH)3 | |
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| 2. | What mass of K2Cr2O7 is needed to make 2.50 L of a 0.200 *M* K2Cr2O7 solution?     |  |  | | --- | --- | |  | (A) 17.0 g | |  | (B) 23.5 g | |  | (C) 53.5 g | |  | (D) 147 g | |  | (E) 588 g | |
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| 3. | A compound was found to contain 43.64% P and 56.36% O by mass. The compound has a molar mass of 283.88 g/mol. What is the compound's molecular formula?   |  |  | | --- | --- | |  | (A) P2O5 | |  | (B) P4O5 | |  | (C) P4O10 | |  | (D) PO2 | |  | (E) P2O | |
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| 4. | Which compound is NOT correctly named?   |  |  | | --- | --- | |  | (A) (NH4)2SO4  ammonium sulfate | |  | (B) O2F2 dioxygen difluoride | |  | (C) RbCl rubidium chloride | |  | (D) Cu3(PO4)2 copper phosphate | |  | (E) None of the above. | |
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| 5. | What are the spectator ions in the reaction of aqueous solutions of PbCl2 and Li2SO4 ?   |  |  | | --- | --- | |  | (A) Li+ and Cl- | |  | (B) Pb2+ and SO42- | |  | (C) Li+ and Pb2+ | |  | (D) SO42- and Cl- | |  | (E) Li+ and SO42- | |
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| 6. | The active ingredient in baking soda is NaHCO3. How many grams of NaHCO3 are needed to provide 2.15 x 1023 atoms of oxygen?   |  |  | | --- | --- | |  | (A) 6.19 g | |  | (B) 10.0 g | |  | (C) 18.6 g | |  | (D) 30.0 g | |  | (E) 90.0 g | |
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| 7. | A 20.0-mL solution of 1.10 M KCl is diluted. How much water must be added to dilute the concentration to 0.750 M KCl?     |  |  | | --- | --- | |  | (A) 9.3 mL | |  | (B) 21.3 mL | |  | (C) 29.3 mL | |  | (D) 41.3 mL | |  | (E) 82.5 mL | |
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| 8. | Nitrogen gas can be formed from the reaction of NH3­­(g) with CuO(s) according to the following equation. How much nitrogen gas is formed when 18.1 g of NH3 are reacted with 90.4 g of CuO?  2NH3(g) + 3CuO­(s) 🡪 N2(g) + 3Cu(s) + 3H2O(g)     |  |  | | --- | --- | |  | (A) 4.30 g | |  | (B) 10.6 g | |  | (C) 14.9 g | |  | (D) 29.8 g | |  | (E) 31.8 g | |
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| 9. | Which statement regarding images a and b below is *false*? Assume that each sphere represents one mole of substance. |

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|  | (A) In image (a), there are equal masses but different numbers of moles of substances. |
|  | (B) In image (b), there are equal numbers of moles but different masses of each substance. |
|  | (C) In image (b), the substance on the right has a smaller molar mass than the substance on the left. |
|  | (D) In image (b), both substances have the same molar mass. |
|  | (E) None of the above. |

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| 10. | What is the coefficient for O2(g) when this equation is balanced?  C2H6(g) + O2(g) 🡪 H2O(g) + CO2(g) |
|  | A) 7  B) 4  C) 3.5  D) 1.75  E) 2 |

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| 11. | Which ionic compound below is *insoluble* in aqueous solution?     |  |  | | --- | --- | |  | (A) LiOH | |  | (B) CaCO3 | |  | (C) K3PO4 | |  | (D) (NH4)2SO4 | |  | (E) FeCl2 | |

13. What is the chemical formula for the compound containing the iron(III) cation and the sulfate anion?

A) Fe3SO4

B) FeS

C) Fe3(SO4)2

D) Fe2(SO4)3

E) Fe2S3

14. Which chemical species is *incorrectly* named?

A) CaBr2 calcium bromide

B) TiCl2 titanium(II) chloride

C) (NH4)2CO3 ammonium carbonate

D) PF5 phosphorus pentafluoride

E) SO32- sulfur trioxide

15. The metal X forms the chloride XCl4 containing 74.76% Cl by mass. What is element X?

A) B

B) Ti

C) Mg

D) C

E) There is not enough information to answer the question.

16. What are the coefficients represented by A and B, respectively, in the balanced chemical reaction:

\_\_\_ SrO(s) + \_\_\_Al(s) 🡪 \_\_Sr(s) + \_\_\_Al2O3(s)

A B C D

A) 1 and 1

B) 3 and 2

C) 2 and 3

D) 1 and 2

E) 2 and 2

17. How many atoms of nitrogen are in 20.85 g Ba(NO3)2 ? The molar mass of Ba(NO3)2 is 261.337 g/mol.

A) 8.84 x 1024 atoms

B) 4.80 x 1022 atoms

C) 9.60 x 1022 atoms  
 D) 1.92 x 1023 atoms

E) 8.13 x 1026 atoms

18. What is the solid produced when Li3PO4 reacts with calcium nitrate?

A) LiNO3

B) Ca3(PO4)2

C) Ca2+

D) NO3-

E) Li

19. How many moles of HCl are required to completely react with 1.87 g Al according to the reaction:

6 HCl(aq) + 2Al(s) 🡪 2AlCl3(aq) + 3H2(g)

A) 0.0347 mol

B) 0.0693 mol

C) 0.208 mol

D) 0.419 mol

E) 7.40 mol

20. A 10.0 mL sample of an unknown concentration of K2CrO­4­ was diluted with water to a total volume of 250.0 mL. The concentration of the resulting solution was 0.0200 *M*. What was the concentration of the original 10.0 mL sample?

A) 0.0800 *M*

B) 0.172 *M*

C) 0.240 *M*

D) 0.250 *M*  
 E) 0.500 *M*

21. What is the maximum mass (in g) of PCl3 that can be produced from the reaction of 125 g P4 with 323 g Cl2?

P4(s) + 6Cl2 🡪 4PCl3

A) 555 g

B) 418 g

C) 626 g

D) 139 g

E) There is not enough information to answer the question.

22. The percent yield for a given reaction is 79.4%. If the theoretical yield is 60.1g, then how much product was actually made?

A) 0.760 g

B) 47.7 g

C) 75.7 g

D) 79.4 g

E) There is not enough information to answer this question.

23. What is the *net* ionic equation for the following chemical reaction?

Na2CO3(aq) + Ca(NO3)2(aq) 🡪 CaCO3(s) + NaNO3(aq)

A) Na2CO3(aq) + Ca(NO3)2(aq) 🡪 CaCO3(s) + NaNO3(aq)

B) 2Na+(aq)+ CO32-(aq)+ Ca2+(aq)+ 2NO3-(aq) 🡪 CaCO3(s) + Na+(aq)+ NO3-(aq)

C) CO32-(aq) + Ca2+(aq) 🡪 CaCO3(s)

D) Na2CO3(aq) + Ca2+(aq)+ 2NO3-(aq) 🡪 CaCO3(s) + Na+(aq)+ NO3-(aq)

E) 2Na+(aq)+ CO32-(aq)+ Ca2+(aq)+ 2NO3-(aq) 🡪 CaCO3(s) + 2Na+(aq)+ 2NO3

24. Which compound is *correctly* named?

A) HCl (aq) hydrogen chloride

B) H2S sulfuric acid

C) CoCl2 cobalt chloride

D) VPO4 vanadium(III) phosphate

E) NaClO sodium chlorite

25. Which statement is *true*?

A) A 1.0 L sample of 0.200 *M* Ca(OH)2 contains 1.20 x 1023 hydroxide ions.

B) In a dilution, the volume of solution remains constant while the moles of solute decreases.

C) There are 124 moles of solute in 145.6 mL of 0.850 *M* sodium cyanide.

D) To prepare 0.80 L of 0.15 *M* NaCl, 0.020L of 5.0 *M* NaCl must be diluted to a total volume of 800. mL.

E) It takes 21.1 g of KI to make 500.0 mL of 0.254 *M* KI solution.

26. Boron has two isotopes; one of these is boron-10 with a mass of 10.0129 amu and a percent abundance of 19.91%. If the universal mass of boron is 10.811, what is the mass of the other isotope of boron.

A) 11.01 amu

B) 44.29 amu

C) 14.02 amu

D) 10.90 amu

E) There is not enough information to answer this question.

27. All of the substances below are fertilizers. Which is the richest source of nitrogen on a mass percentage basis?

A) Urea, (NH­2)2CO

B) Ammonium nitrate, NH4NO3

C) Guanidine, HNC(NH2)2

D) Ammonia, NH3

E) There is not enough information to answer this question

28. Depicted below is a reaction vessel containing a mixture of H2(g) and N2(g) before the reaction begins. The filled spheres represent nitrogen and the open spheres represent hydrogen. Given the equation below, which image best represents the vessel *after* reaction occurs?

N2(g) + 3H2(g) 🡪 2NH3(g)

ANS: B

29. How many moles of carbon monoxide are required to react with one mole of O2(g) to produce 2 moles of carbon dioxide gas?

A) ½

B) 1

C) 2

D) 3/2

E) 3

30. What are the spectator ions in the reaction of sulfuric acid with barium hydroxide?

A) Ba2+ and SO42-

B) H+ and OH-

C) OH- and Ba2+

D) H+ and SO4 2-

E) There are no spectator ions in this reaction.

31. What is the name of Cr(PO4)2?

A) chromium phosphide

B) chromium phosphate

C) chromium(II) phosphate

D) chromium diphosphorus octoxide

E) chromium (VI) phosphate

33. An element has two naturally occurring isotopes. Isotope 1 has a mass of 129.9038 amu and a relative abundance of 57.4%. Isotope 2 has a mass of 122.9042 amu. What is the atomic weight of the element?

A) 126.4

B) 126.9

C) 129.9

D) 122.9

E) There is not enough information to answer this question.

34. What is the mass percent of chlorine in C2Cl4F2?

A) 69.6%

B) 17.4%

C) 53.3%

D) 30.4%

E) 34.8%

35. The empirical formula of butanedione is C2H3O. If 11.19 g of butanedione contains 0.130 moles, then what is the molecular formula?

A) C2H3O

B) C6H9O3

C) C3H2O3

D) C4H6O2

E) There is not enough information.

36. Given the *unbalanced* equation H2(g) + O2(g) 🡪 H2O(g) and the image below, which statement is true?

A) O2 is the limiting reagent in this reaction.

B) H2 is the excess reagent in this reaction.

C) This reaction has no limiting reagent.

D) To completely consume all the O2 present in the reactants image, 12 H2 are required.

E) To completely consume all the H2 present in the reactants image, 12 O2 are

required.

37. Which ions correlate to the image shown? Correct stoichiometry is NOT shown.

A B C D

A) Na+ NO3- PO43- Ba2+

B) Ba2+ Na+ NO3- PO43-

C) Na+ Ba2+ Cl- NO3-

D) Na+ Cl- Ba2+ NO3-

E) NH4+ Na+ Cl- SO42-

38. What is the coefficient for CH2O when the following equation is balanced?

\_\_\_\_\_\_ CH2O + \_\_\_\_\_\_ O2 🡪 \_\_\_\_\_\_ CO + \_\_\_\_\_\_ H2O

A) 4

B) 3

C) 0

D) 1

E) 2

39. What is the concentration of a 20.00 mL sample of H2SO4 if it takes 25.00 mL of 0.500 M NaOH to completely react with the H2SO4?

2NaOH(aq) + H2SO4(aq) 🡪 2H2O(l) + Na2SO4(aq)

A) 0.400 *M*

B) 0.157 *M*

C) 0.500 *M*

D) 0.313 *M*

E) 0.625 *M*

40. Which reaction does NOT occur?

A) K2CO3(aq) + NiCl2(aq) 🡪

B) NaNO3(aq) + Li2SO4(aq) 🡪

C) KI(aq) + Pb(NO3)2 🡪

D) AgNO3(aq) + NaCl(aq) 🡪

E) CaBr2(aq) + Na2CO3(aq) 🡪

**Chapter 1**

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| 1. | Estimate the frequency of the wave shown below. | |
| A) | 1.5 × 1017 s−1 |
| B) | 3.0 × 1017 s−1 |
| C) | 1.0 × 1017 s−1 |
| D) | 2.0 × 1017 s−1 |
| E) | Not enough information is given to permit calculation of the frequency. |

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| 2. | Which of the following emission lines corresponds to the longest wavelength?   |  |  | | --- | --- | |  | (A) *n*2  *n*1 | |  | (B) *n*4  *n*2 | |  | (C) *n*4  *n*1 | |  | (D) *n*3  *n*2 | |  | (E) *n*4  *n*3 | |
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| 4. | The total number of orbitals in a shell with principal quantum number 5 is |
|  | A) 32  B) 50  C) 25  D) 40  E) 5 |

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| 5. | How many unpaired electrons are in a ground state vanadium (V) atom? |
|  | A) 0  B) 1  C) 3  D) 4  E) 5 |

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| 6. | Which set of quantum numbers could correspond to a 4f-orbital? | | |
| A) | *n* = 4, *l* = 4, *ml* = +3 |
| B) | *n* = 4, *l* = 3, *ml* = +4 |
| C) | *n* = 4, *l* = 3, *ml* = 3 |
| D) | *n* = 3, *l* = 2, *ml* = +1 |
| E) | *n* = 3, *l* = 2, *ml* = 0 |

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| 7. | Calculate the wavelength of a motorcycle of mass 275 kg traveling at a speed of 125 kmhr1. | |
| A) 3.44 x 10-19 m |
| B) 1.93 x 10-35 m  C) 1.93 x 10-38 m |
| D) 6.94 x 10-38 m |
| E) 3.44 x 10-35 m |

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| 8. | Which statement *best* explains why the electron affinity of B is ***less positive*** than that of Li? |
|  | A) B has electrons in p orbitals, so it is lower in energy.  B) Li does not want to gain an electron because it already has a half-full s orbital.  C) B is a smaller atom, so its valence electrons are more tightly held.  D) B is farther to the right in the same row as Li.  E) The gain of an electron fills the s orbital for Li while it results in neither a full nor half-full subshell for B. |

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| 9. | Arrange the elements in order of *increasing* first ionization energy: Cl, Ga, Si |
|  | A) Si < Ga < Cl  B) Ga < Si < Cl  C) Cl < Si < Ga  D) Cl < Si < Ga  E) Si < Cl <Ga |

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| 10. | Which of the following statements regarding electromagnetic radiation is true? | |
| A) | Electromagnetic radiation with a wavelength of 400 nm travels faster than that with a wavelength of 600 nm. |
| B) | The frequency of electromagnetic radiation determines how fast it travels. |
| C) | Electromagnetic radiation with a wavelength of 400 nm has a frequency that is smaller than that with a wavelength of 600 nm. |
| D) | Electromagnetic radiation with a wavelength of 600 nm travels faster than that with a wavelength of 400 nm. |
| E) | Electromagnetic radiation with a wavelength of 600 nm has a frequency that is smaller than that with a wavelength of 400 nm. |
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| 11. | Estimate the energy of the wave shown below. | |
| A) | 1 x 10-9 J |
| B) | 2 x 10-9 J |
| C) | 2 x 10-15 J |
| D) | 1 x 10-16 J |
| E) | Not enough information is given to permit calculation of the frequency. |
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| 12. | Which of the following emission lines corresponds to the greatest frequency?   |  |  | | --- | --- | |  | (A) *n*2  *n*1 | |  | (B) *n*4  *n*2 | |  | (C) *n*4  *n*1 | |  | (D) *n*3  *n*2 | |  | (E) *n*4  *n*3 | |

13. Which image could correlate to the following set of quantum numbers?

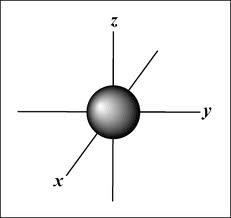
n = 2

l = 1

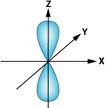
ml = -1

ms = +1/2

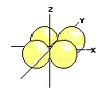
A)



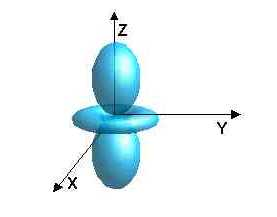
B)



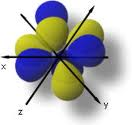
C)



D)



E)



14. What it the maximum number of electrons in an atom than can have the quantum numbers n = 3 and ml = +1?

A) 1

B) 2

C) 4

D) 9

E) 18

16. Which of the following statements is true?

I: Photons of ultraviolet radiation have less energy than photons of infrared radiation.

II: The kinetic energy of an electron ejected from a metal surface when the metal is irradiated with ultraviolet radiation is independent of the frequency of the radiation.

III: The energy of a photon is inversely proportional to the wavelength of the radiation.

A) I only

B) III only

C) I and III only

D) II and III only

E) I, II, and III

17. Which set of quantum numbers is NOT allowed?

n l ml ms

A) 6 4 -3 +1/2

B) 3 2 +1 -1/2

C) 1 0 0 +1/2

D) 2 1 +2 +1/2

E) 4 1 -1 -1/2

18. What is the best ground state electron configuration for Cu?

A) [Ar]3s23d9

B) [Ar]3s24d9

C) [Ar]3s13d10

D) [Ar]4s13d10

E) [Ar]4s23d9

19. How many unpaired electrons are in the electron configuration for Ge?

A) 0

B) 1

C) 2

D) 4

E) 3

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| 20. Which statement best explains why the ionization energy of O is less positive than that of N? |
| A) N has fewer p electrons.  B) N does not want to gain an electron because it already has a half-full p orbital.  C) O is a smaller atom, so its valence electrons are less tightly held.  D) O is farther to the right in the same row as N.  E) The loss of an electron from O results in a more stable electron configuration whereas  the loss of an electron from N results in a less stable electron configuration. |
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21. A radio wave has a frequency of 3.8 x 1010 s-1. How many photons of this radiation are needed to produce 60.7 J of energy?

A) 2.52 x 10-23 photons

B) 2.41 x 1024 photons

C) 1.74 x 10-44 photons

D) 1.88 x 1017 photons

E) Not enough information

22. Rank the emissions in terms of increasing energy:

I. *n* = 5 🡪 *n* = 4

II. *n* = 3 🡪 *n* = 2

III. *n* = 2 🡪 *n* = 1

A) I < II < III

B) III < II < I

C) II < I < III

D) III < I < II

E) The energies of emissions I, II, and III are all the same

23. Which statement is *false*?

A) The principle quantum number indicates the energy level and size of an orbital.

B) The only allowed values for the spin quantum number are +½ and – ½.

C) The values for the magnetic quantum number indicate the

orientation of the orbitals in space.

D) The value *l* = 2 indicates a peanut shaped orbital.

E) With an increase in energy level, a new type of orbital becomes available.

24. How many *orbitals* in an atom can have the designation 5f?

A) 5

B) 7

C) 14

D) 25

E) 50

25. Which statement best explains why the following ionic radii trend is *true*?

S2- > Cl -  > K+

A) Zeff increases with the number of electrons.

B) The potassium ion has the least amount of shielding, so Zeff is greatest for it.

C) Zeff increases with the number of protons.

D) The amount of shielding is equal, but the nucleus of sulfur is least positive

resulting in smallest Zeff.

E) The radii trend above is not correct.

26. Which periodic trend refers to the ability of an atom in a molecule to attract electrons to itself?

A) electron affinity

B) ionization energy

C) polarizability

D) polarizing power

E) electronegativity

27. Which experiment or phenomenon contributed to the evidence that light could be treated as a particle?

A) emission of light when current is passed through H2 gas

B) photoelectric effect

C) Hund's rule

D) aufbau principle

E) diffraction

28. A laser pulse with wavelength 532 nm contains 0.00467 J of energy. How many photons are in the laser pulse?

A) 1.50 x 1024

B) 1.25 x 1016

C) 3.74 x 10-19

D) 1.25 x 1025

E) 3.74x 10-28

29. Which statement is *true*?

A) Energy is emitted as an electron goes from *n* = 1 to *n* = 2.

B) *n* = 0 represents the ground state, or the lowest energy level an electron

can occupy.

C) More energy is emitted when an electron transitions from the *n* = 1 to the

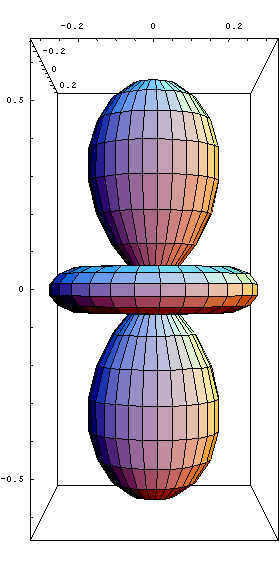
*n* = 2 energy level than when an electron transitions from the *n* = 2 to the *n* = 3 energy level.

D) The wavelength of light that must be absorbed for an electron to be promoted from *n* = 2 to *n* = 3 is less than that required to promote an electron from *n* = 3 to *n* = 4*.*

E) Once an electron is promoted from one energy level to another, it is

"stuck" there and cannot return to the ground state.

30. What is the value of *l* for this orbital?



A) 1

B) 0

C) 3

D) 2

E) The value of *l* cannot be determined without more information.

31. Which set of quantum numbers is not possible?

*n* *l ml ms*

A) 4 3 -3 +1/2

B) 1 0 0 -1/2

C) 6 4 +2 -1/2

D) 2 2 -1 +1/2

E) 3 1 +1 -1/2

32. Which quantum number describes the size of an orbital?

A) n

B) ml

C) l

D) ms

E) It depends on the orbital.

33. How many unpaired electrons are in the ground state electron configuration for ruthenium (Ru)?

A) 6

B) 0

C) 4

D) 8

E) 2

34. Which statement is *true?*

A) O2- and F- have the same ionic radii because they are isoelectronic.

B) Al3+ has a greater ionic radius than Ca2+ because Al has more protons.

C) P has smaller atomic radius than S because its half-full p orbital leads to greater stability.

D) Na has smaller atomic radius than K because K has more core electrons than Na.

E) None of these statements are true.

35. Which statement best explains why nitrogen has more positive electron affinity than oxygen?

A) Oxygen is more electronegative.

B) Nitrogen has a smaller atomic radius.

C) Nitrogen has greater ionic radius.

D) Nitrogen has more stable ground state electron configuration.

E) Nitrogen does NOT have more positive electron affinity than oxygen.

36. Which period 2 (row 2) element could have the following ionization energies?

IE1 = 800.6 kJ/mol

IE2 = 2427.1 kJ/mol

IE3 = 3659.7 kJ/mol

IE4 = 12,547.2 kJ/mol

A) Li

B) Be

C) B

D) C

E) N

**Chapter 2**

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| --- | --- |
| 1. | Which of the compounds below has bonds with the most ionic character? |
|  | A) CaO  B) K2O  C) MgO  D) MgS  E) CaS |

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| 2. | Which statement regarding the resonance structures below is *true?*  resonance 2.JPG |
|  | A) Structure I has the lowest magnitudes of formal charge.  B) Structure III is the greatest contributor to the resonance hybrid.  C) Structure II is the greatest contributor to the resonance hybrid.  D) The three resonance structures are equivalent.  E) Structures I and II are equal contributors to the resonance hybrid. |

|  |  |
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| 3. | How many lone *pairs* of electrons are found on the central atom in the Lewis structure of the compound ICl3? |
|  | A) 2  B) 3  C) 4  D) 8  E) 11 |

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| 4. | Which best describes the number and type of bonds in the HCN molecule?  A) 2 double bonds  B) 2 single bonds  C) 1 double bond and 1 single bond  D) 1 single bond and 1 triple bond  E) 1 double bond and 1 triple bond |

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| 5. | What is the formal charge on the Xe atom in XeF4? |
|  | A) 0  B) - 4  C) +2  D) +4  E) -1 |

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| 6. | What is the ground-state electron configuration of Mn4+? |
|  | A) [Ar]4s23d1  B) [Ar]3d3  C) [Ar]4d3  D) [Ar]4s24d1  E) [Ar]4s23d5 |

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| 7. | Which of the following has the largest atomic or ionic radius? |
|  | A) S2-  B) Cl  C) Cl-  D) K+  E) S |

9. What is the bestelectron configuration for Ni4+ ?

A) [Ar]4s2d4

B) [Ar]3d6

C) [Ar]4s2d6

D) [Ar]4s23d104p2

E) [Ar]4s03d104p4

10. Which best describes the number and type of bonds in SiO2?

A) 1 Si to O single bond and 1 Si to O double bond

B) 1 O to O double bond and 1 Si to O single bond

C) 1 O to O single bond and 1 Si to O double bond

D) 2 Si to O double bonds

E) 1 Si to O triple bond and 1 Si to O single bond

11. What is the formal charge the O atom in the following structure?

A) 0

B) -1

C) -2

D) +1

E) +2

12. The following Lewis structure was drawn for a Period 3 element. Identify the element.

A) Se

B) S

C) Br

D) Cl

E) Si

13. Which resonance structure is the greatest contributor to the resonance hybrid?

ANS: A

14. Which bond is most polar?

A) C—C

B) C—H

C) C—O

D) C—F

E) C—Br

15. Which molecule has the *strongest* C to O bond?

A) CO2

B) CO

C) CH3OH

D) CH3COOH

E) They all have the same strength

16. Which of these compounds does not exist (use Lewis structures to answer this question)?

A) C2H2

B) C2H4

C) C2H6

D) C2H8

E) All of these can exist.

17. Which atom or ion is *smallest*?

A) F –

B) Mg2+

C) Na

D) Al3+

E) Ga

18. Which transition metal ion has the electron configuration [Ar]3d3?

A) V3+

B) Mn4+

C) Cr3+

D) Fe3+

E) Ti2+

20. Which statement regarding the following Lewis structure is *true*?

A) The structure has too many electrons

B) The structure has no resonance structures.

C) The formal charge distribution prevents it from contributing significantly to the resonance hybrid.

D) The structure is wrong because boron cannot have a complete octet.

E) The structure is missing an overall charge.

21. How many unpaired electrons are in the ground state electron configuration for Cu2+?

A) 0

B) 1

C) 2

D) 3

E) 4

22. The Lewis structure of which molecule contains at least one double bond?

A) AsF5

B) PCl3

C) CS2

D) H2S

E) Br3-

23. Which statement is *true*?

A) The stronger an ionic bond, the more ionic character it has.

B) The larger the difference in electronegativity between two atoms, the less ionic character there is in a bond containing them.

C) The more highly charged two ions are, the stronger the ionic bond between them.

D) The more covalent character a bond has, the weaker it is.

E) All ionic bonds have covalent character, but covalent bonds do not have

ionic character.

24. Which of these *best* explains why B has an incomplete octet in BF3?

A) formal charge

B) bond order

C) electronegativity

D) polarity

E) ionic charge

25. Use the figures below to determine which statement is *true.*

A) Structure A has the longest N—O bond because double bonds are longer than single bonds.

B) You cannot tell which structure has the longest N—O bond because B and C both have equivalent resonance structures.

C) Structure B has weaker N—O bonds than structure C because the average

bond length is longer than in structure C.

D) Structure A has the strongest N—O bonds because the average bond length is shortest.

E) The N—O bond lengths are all equal in these structures.

26. Which statement is *true*?

A) Polarizability accounts for the ionic character of covalent bonds.

B) Electronegativity accounts for the covalent character of ionic bonds.

C) Na+ has greater polarizing power than Cs+.

D) KCl is a purely ionic compound.

E) In H2O, the O atom has a f*ull* negative charge.

27. Which resonance structure for the perchlorate ion is the greater contributor to the resonance hybrid and why?

ANS: E

A) Structure A is the greater contributor because structure B has too many

bonds to Cl.

B) Neither structure is a contributor because they both have an expanded octet on Cl.

C) StructureB is the greater contributor because all of the oxygen atoms have zero formal charge.

D) Structure B is the greater contributor because it is more symmetrical.

E) Structure A is the greater contributor because it has low magnitude of charge and the most electronegative atom has the most negative formal charge.

**Chapter 3**

|  |  |  |
| --- | --- | --- |
| 1. | What is the *shape* of AsF3? | |
| A) tetrahedral |
| B) trigonal planar |
| C) trigonal pyramidal |
| D) bent (V-shaped or angular) |
| E) trigonal bipyramidal |

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| 2. | In which molecule would you expect deviations in the ideal bond angles? |
|  | A) NH4+  B) SF6  C) TeBr4  D) PCl5  E) BH4+ |

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| 3. | Which molecule is polar? |
|  | A) CCl4  B) PCl5  C) ICl4-  D) SF6  E) AsF5 |

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| 4. | Which molecule is nonpolar but contains polar bonds? | |
| A) N2 |
| B) SiH2Br2 |
| C) CF4 |
| D) CH3OH |
| E) NH3 |

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| 5. | Which theory is *most* appropriate for determining whether a molecule is polar? | |
| A) Valence bond theory |
| B) Lewis theory |
| C) VSEPR theory |
| D) Molecular orbital theory |
| E) None of these theories tell us about polarity. |

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| 6. | The hybrid orbitals used in carbons 2 and 4, respectively, in the molecule below are: | | | |
| A) | *sp*3 and *sp* |  |  |
| B) | *sp*2 and *sp* |  |  |
| C) | *sp*3 and *sp*3 |  |  |
| D) | *sp*2 and *s2* |  |  |
| E) | *spd*and *sp* |  |  |

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| 7. | What is the ground-state electron configuration of F2+ ? |
|  | A) (2*s*)2(2*s*\*)2(2*p*)4(2*p*)2(2*p*\*)3 |
|  | B) (2*s*)2(2*s*\*)2(2*p*)2(2*p*)4(2*p*\*)4(2*p*\*)1 |
|  | C) (2*s*)2(2*s*\*)2(2*p*)2(2*p*)4(2*p*\*)4 |
|  | D) (2*s*)2(2*s*\*)2(2*p*)2(2*p*)4(2*p*\*)3 |
|  | E) (2*s*)2(2*s*\*)2(2*p*)4(2*p*)2(2*p*\*)4(2*p*\*)1 |

|  |  |
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| 8. | Which statement regarding the images below is *true*?  Exam Two sigma an dbond.JPG |
|  | A) Images I and III depict bonds |
|  | B) Image II depicts a  bond |
|  | C) Images I and II depict  bonds |
|  | D) Images II and III depict bonds |
|  | E) Images I, II, and III all depict bonds |

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| 9. | What is the bond order for N22+ ? |
|  | A) 2.5  B) 1  C) 2  D) 1.5  E) 3 |

12. For which molecule is the electron group arrangement the same as the shape?

A) NF3

B) H2S

C) I3-

D) SeBr4

E) HCN

13. Which of these molecules is *nonpolar*?

A) PH3

B) BH3

C) SiCl2F2

D) XeF5+

E) OCl2

14. Which statement regarding sigma and pi bonds is *true*?

A) All sigma bonds are composed of s atomic orbitals.

B) All pi bonds contain a sigma bond.

C) All triple bonds contain three pi bonds.

D) A sigma bond can be composed of p orbitals, but only if they are oriented along the internuclear axis.

E) All statements are true.

15. What is the molecular structure of ClF3 and why?

A) Trigonal planar. The lone pairs in ClF3 occupy the axial positions.

B) Trigonal bipyramidal. The lone pairs in ClF3 occupy the equatorial positions.

C) T-shaped. The lone pairs in ClF3 occupy the equatorial positions.

D) T-shaped. The lone pairs in ClF3 occupy the axial positions.

E) Trigonal pyramidal. The lone pair in ClF3 is neither axial nor equatorial.

16. What is the *shape* of SO42-?

A) tetrahedral

B) trigonal planar

C) trigonal bipyramidal

D) trigonal pyramidal

E) T-shaped

17. What is the bond order for F2- according to molecular orbital theory?

A) 1

B) 2.5

C) 0.5

D) 0

E) 2

18. According to molecular orbital theory, how many unpaired electrons does a molecule of O2 have?

A) 0

B) 1

C) 2

D) 3

E) 4

19. Which molecule or ion has an sp2 hybridized central atom?

A) HCN

B) CO32-

C) NH3

D) H2O

E) SiO2

20. Which molecule is trigonal pyramidal?

A) SO2

B) AsBr3

C) BrF3

D) BH3

E) CH2O

21. Which molecule or ion is polar?

A) XeF4

B) PCl3

C) CCl4

D) I3 –

E) BH3

22. Arrange the following in order of increasing net dipole (dipole moment).

A) I < II < III

B) III < I < II

C) II < III < I

D) II < I < III

E) I < III < II

24. How many and what type of atomic orbitals must be combined to account for the tetrahedral shape of CH4?

A) 2 s orbitals

B) 4 p orbitals

C) 8 p orbitals

D) 1 s and 3 p orbitals

E) 2 s and 6 p orbitals

25. Which of the following species has a longer bond length and why?

N2 or N2+

A) N2 has a longer bond because its bond order is smaller

B) N2 has a longer bond because its bond order is greater

C) Both species have the same bond length

D) N2+ has a longer bond because its bond order is smaller

E) N2+ has a longer bond because its bond order is greater

26. What is the molecular orbital electron configuration for F2 - ?

A) (σ2s) 2( σ2s\*)2(σ2p)2(π2p)4(π2p\*)4(σ2p\*)1

B) (σ2s) 2( σ2s\*)2(σ2p)2(π2p)4(π2p\*)3

C) (σ2s) 2( σ2s\*)2(σ2p)2(π2p)4(π2p\*)4

D) (σ2s) 2( σ2s\*)2(π2p)4(σ2p)2(π2p\*)4(σ2p\*)1

E) (σ2s) 2( σ2s\*)2(π2p)4(σ2p)2(π2p\*)3

27. What is the electron group arrangement for XeF2?

A) linear

B) trigonal planar

C) see-saw

D) trigonal bipyramidal

E) t-shaped

28. Which molecule *must* be polar?

A) C4H8

B) ICl4-

C) PCl2Br3

D) GeF2Cl2

E) I3-

29. What is the value of x in the formula O2x if its ground state molecular orbital configuration is (2s)2(2s)2(2p)2(2p)4?

A) 1-

B) 1+

C) 2-

D) 2+

E) There is not enough information to answer this question.

30. According to molecular orbital theory, what is the bond order for N2- ?

A) 1

B) 2

C) 2.5

D) 3

E) 3.5

31. Which of the following *best* and uniquely characterizes pi bonds?

A) p orbitals

B) s orbitals

C) head on overlap

D) overlap without any nodes

E) overlap above and below the internuclear axis

32. Which statement is *true*?

A) Anti-bonding molecular orbitals are lower in energy than the atomic orbitals of which they are composed.

B) Anti-bonding molecular orbitals are created from constructive interference of atomic orbitals.

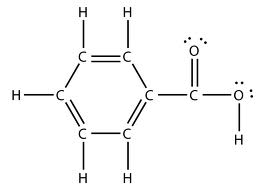
C) Electrons are never found in anti-bonding orbitals.

D) Bonding molecular orbitals have electron density primarily inside the

space between the two nuclei.

E) In the molecular orbital diagram for a heteronuclear diatomic molecule, the bonding orbitals are more similar to the atomic orbitals of the less electronegative atom.

33. How many sigma and pi bonds are in the following structure?



A) 15 sigma, 4 pi

B) 11 sigma, 4 pi

C) 11 sigma, 8 pi

D) 19 sigma, 4 pi

E) 15 sigma, 8 pi

**Chapter 4**

|  |  |  |
| --- | --- | --- |
| 1. | If 250.0 mL of a gas at STP weighs 2.00 g, what is the molar mass of the gas? | |
| A) 55.8 g/mol |
| B) 179 g/mol |
| C) 32.8 g/ mol |
| D) 8.00 g/mol |
| E) There is not enough information to answer this question. |

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| 2. | A cylinder containing a gas of volume 30.0 L, at a pressure of 110. kPa and a temperature of 420. K. Find the temperature of the gas that has a volume 40.0 L at a pressure of 120. kPa. |
|  | A) 312 K  B) 41.2 K  C) 344 K  D) 164 K  E) 611 K |

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| 3. | A gas mixture contains CO, Ar, and H2. What is the total pressure of the mixture if the mole fraction of H2 is 0.35 and the partial pressure of H2 is 0.58 atm? |
|  | A) 0.603 atm  B) 0.203 atm  C) 1.66 atm  D) 0.0552 atm  E) There is not enough information to answer this question. |

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| 4. | |  | | --- | | How many moles of NaO2 are needed to react with 75.0 L of carbon dioxide at STP?  4NaO2(s) + 2CO2(g)  2Na2CO3(s) + 3O2(g) | |
|  | A) 0.15 moles  B) 1.67 moles  C) 3.35 moles  D) 6.70 moles  E) 13.4 moles |

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| 6. | The graph below represents the gases He, Ar, N2, and CH4 all at the same temperature. Which statement regarding this graph is *true*? |
|  | A) Curve 3 represents N2.  B) Curve 1 represents N2  C) Curve 4 represents Ar.  D) Curve 2 represents He.  E) Curve 3 represents CH4. |

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| 7. | An ideal gas differs from a real gas in that the molecules of an ideal gas\_\_\_\_\_\_\_\_\_. |
|  | A) have an average molecular mass.  B) have no kinetic energy.  C) have a molecular weight of zero.  D) have no attraction for one another.  E) take up a significant amount of volume in comparison to the volume of the sample. |

|  |  |  |
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| 9. | A plot of the Maxwell distribution against speed for different molecules shows that | |
| A) | heavy molecules have a higher average speed. |
| B) | light molecules have a very narrow range of speeds. |
| C) | molecules at low temperatures have a wide range of speeds. |
| D) | molecules at high temperatures have a lower average speed. |
| E) | heavy molecules travel with speeds close to their average values. |

10. Which homonuclear diatomic molecule has a root mean square speed (rms) of 1920 m/s?

A) H2

B) N2

C) O­2

D) F2

E) There is not enough information to solve this problem.

11. Which visualization best represents the distribution of O2 and SO2 molecules near an orifice some time after effusion occurs in the direction indicated by the arrows?

ANS: C

14. What is the partial pressure of HCl if the total pressure inside the container shown below is 0.70 atm?

A) 0.41 atm

B) 0.70 atm

C) 0.29 atm

D) 3.5 atm

E) 0.14 atm

15. The volume of a sample of ideal gas is 24.8 mL at 1.12 atm. If the pressure of the sample is increased, at constant temperature, to 2.64 atm, then what is the new volume of the sample in liters?

A) 10.5 L

B) 0.119 L

C) 0.0566 L

D) 0.0105 L

E) 0.00660 L

16. If it takes 1.25 min for 0.010 mol of He to effuse, how long will it take for the same amount of C2H6 to effuse under the same conditions?

A) 3.43 min

B) 2.19 min

C) 0.456min

D) 2.74 min

E) 0.0228 min

17. What mass of potassium chloride forms when 5.25 L of chlorine gas at 0.950 atm and 293 K reacts with 17.0 g of potassium?

A) 30.9 g KCl

B) 32.4 g KCl

C) 681 g KCl

D) 15.4 g KCl

E) 341 g KCl

18. A sample of SO2 gas has a volume of 575 mL at 22.0oC and 0.989 atm. How many moles of SO2 are in the sample?

A) 0.0235 moles

B) 42.6 moles

C) 0.315 moles

D) 1.03 moles

E) 3. 17 moles

19. A sample of CO2 gas has a pressure of 75.4 torr in a 125 mL flask at a given temperature. The sample is transferred to a new flask, and the pressure is measured at 62.3 torr at the same temperature. Which statement is true?

A) The volume of the gas decreases.

B) The volume of the gas stays the same.

C) The volume of the gas increases.

D) The number of moles of gas increases.

E) None of these statements are true.

20. A newly discovered substance in the gas phase has density of 2.39 g/L at 25.0oC and 0.941 atm. What is the molar mass of the substance?

A) 16.1 g/mol

B) 62.1 g/mol

C) 5.21 g/mol

D) 0.0919 g/mol

E) This problem cannot be solved without the volume of sample.

21. Which statement is *true*?

A) Partial pressures of gases can be added to find the total pressure of a system for all non-reactive ideal gases.

B) Partial pressures of gases can be added to find the total pressure of a system only when the gases are Noble gases.

C) Partial pressures of gases cannot be added to find the total pressure of a system because pressure increases exponentially with the addition of gases.

D) The total pressure of a gas mixture is the average of the partial pressures of the gases in the mixture.

E) None of these statements are true.

22. An 11.2 L sample of C2H6 reacts with 44.8 L of O2 at STP. By the end of the reaction, the temperature has increased to 318.5 K and the pressure increased to 1.10 atm. How many liters of water forms?

2 C2H6(g) + 7 O2(g) 🡪 4 CO2(g) + 6 H2O(g)

A) 33.8 L

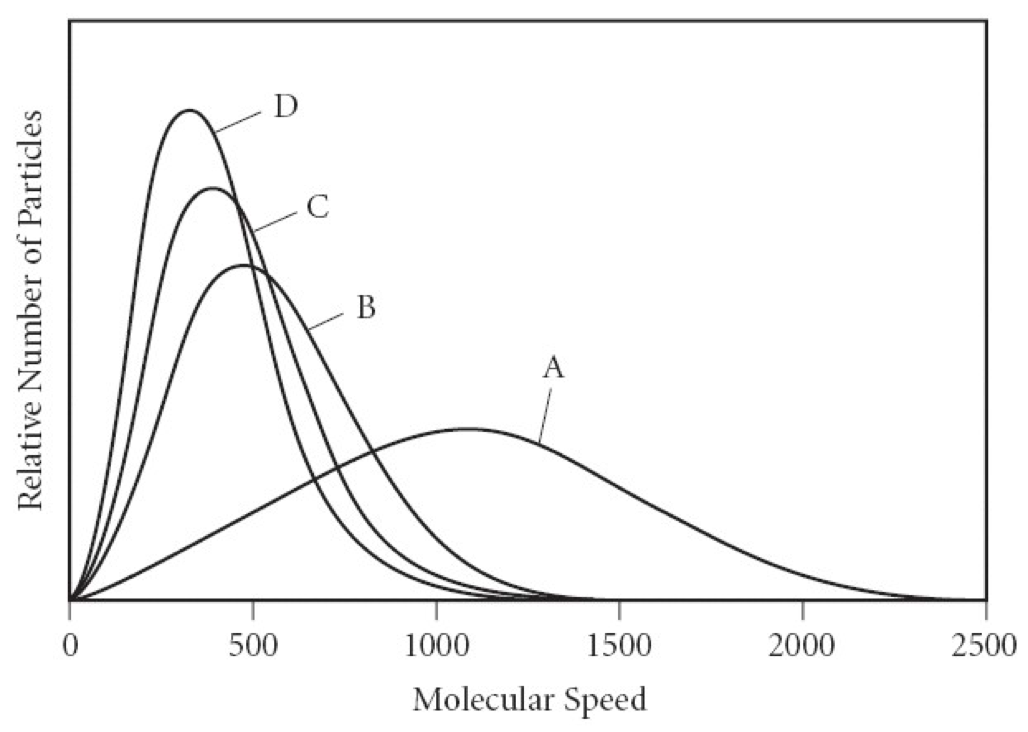
B) 67.2 L

C) 71.9 L

D) 35.6 L

E) 38.3 L

24. The graph below represents the gases Cl2, Ne, O2, and CH4 all at the same temperature. Which curve represents O2?



**1**

**2**

**3**

**4**

|  |  |
| --- | --- |
|  | A) Curve 1  B) Curve 2  C) Curve 3  D) Curve 4  E) None of the curves can represent O2 |

25. To what temperature must a sample of O2 gas be heated to have the same root mean square speed as a sample of H2 gas at 385 K?

A) 6.10x103 K

B) 24.1 K

C) 12300 K

D) 12.0 K

E) Not enough information.

**Chapter 5**

|  |  |
| --- | --- |
| 1. | Which of the following can form intermolecular hydrogen bonds (with other molecules of itself)? |
|  | A)NH2CH2COOH  B) SiH4  C) CH3COCH3  D) CH2F2  E) CH3Cl |

|  |  |
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| 2. | Which of the following has the *second lowest* boiling point? |
|  | A) HF  B) GeH4  C) SiH4  D) PH3  E) H2Se |

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| 3. | Which property explains why water beads on the hood of your car after a rainstorm? |
|  | A) capillary action  B) surface tension  C) viscosity  D) intramolecular attractions  E) All of the above. |

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| 4. | Which of the following cations is likely to be *most strongly* hydrated in compounds? |
|  | A) Rb+  B) Li+  C) K+  D) Cs+  E) NH4+ |

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| 5. | Which statement *best* explains why neopentane has a lower boiling point than n-pentane?    **neopentane n-pentane** |
|  | A) Neopentane's primary intermolecular force is London interactions (London  dispersion forces) while n-pentane's is dipole-dipole interactions.  B) The shape of n-pentane results in more London interactions than occur in  neopentane.  C) n-pentane has greater molar mass than neopentane.  D) They have the same molar mass.  E) Neopentane is less viscose than n-pentane. |

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| 6. | The atomic radius of magnesium is 160 pm. Estimate its density, given that the metal has a face centered cubic structure. | | |
| A) | 1.74 gcm3 |
| B) | 0.435 gcm3 |
| C) | 10.5 gcm3 |
| D) | 4.45 gcm3 |
| E) | 2.78 gcm3 |

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| 7. | What is the coordination number for each atom in the structure below? | | |
| A) | 1 |
| B) | 3 |
| C) | 6 |
| D) | 7 |
| E) | 12 |

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| --- | --- | --- |
| 8. | What are all the intermolecular forces that are responsible for the existence of ice? | |
| A) | hydrogen bonding and London forces |
| B) | London forces |
| C) | dipole-dipole, London forces, and hydrogen bonding |
| D) | dipole-dipole and ion-ion |
| E) | hydrogen bonding and dipole-dipole |

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| --- | --- | --- | --- |
| 9. | The mass of a face-centered cubic unit cell is | | |
| A) | two times the mass of one atom. |
| B) | five times the mass of one atom. |
| C) | equal to the mass of one atom. |
| D) | six times the mass of one atom. |
| E) | four times the mass of one atom. |

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| 10. | Select the *true* statement(s).  I. H2O has a lower boiling point than diamond because H2O is a molecular solid held together only by intermolecular forces.  II. Diamond has a higher boiling point than H2O because network solids are held together by intramolecular forces.  III. H2O has a higher boiling point than KCl because its intermolecular forces are stronger. |
|  | A) I only  B) I and II  C) I, II, and III  D) III only  E) II and III |

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| 11. | Identify the number of atoms per cell. |
|  | A) 1  B) 2  C) 3  D) 5  E) 9 |

12. Which statement is false?

A) Capillary action is responsible for the meniscus in a test tube or graduated cylinder.

B) Viscosity is the property that makes room temperature syrup pour more

slowly than water.

C) Surface tension is the property that explains why a bug can walk on water.

D) Capillary action and viscosity are dependent on intermolecular

attractions, but surface tension is not.

E) All of these are true.

13. The melting point of O2 is -218oC. Based on this data point, at what temperature do you predict N2 will melt and why?

A) -210 oC; N2 has stronger intermolecular forces than O2

B) -210 oC; N2 has weaker intermolecular forces than O2

C) -218 oC; N2 and O2 are both nonpolar and their intermolecular forces are the same strength

D) -226 oC; N2 has stronger intermolecular forces than O2

E) -226oC; N2 has weaker intermolecular forces than O2

14. What is the mass of a body centered cubic unit cell of gold (Au)?

A) 6.54 x 10-22 g

B) 3.27 x 10-22 g

C) 1.31 x 10-21 g

D) 1.65 x 10-22 g

E) There is not enough information to solve this problem.

15. The Li+ cation is strongly hydrated in aqueous solution. Which intermolecular force best accounts for this?

A) hydrogen bonding

B) ion-ion interaction

C) ionic bonding

D) ion-dipole interaction

E) dipole-dipole interaction

|  |  |
| --- | --- |
| 16. | Which molecular solid has the highest melting point?  ANS: A |

17. The radius of aluminum is 1.43 x 10-10 m. What is the edge length given that the metal has face-centered cubic unit cell structure?

A) 3.30 x 10-10 m

B) 4.04 x 10-10 m

C) 2.86 x 10-10 m

D) 3.38 x 10-5 m

E) There is not enough information to solve this problem.

18. What is the primary intermolecular force in a sample of H2S?

A) ion-dipole interaction

B) London forces

C) dipole-dipole interaction

D) hydrogen bonding

E) covalent bonding

19. Which statement best explains why CH4 has a lower boiling point than C2H6?

A) CH4 does not have a lower boiling point than C2H6; CH4 is polar while C2H6 is not.

B) C2H6 is polar while CH4 is not.

C) C2H6 is more polarizable due to its greater number of electrons.

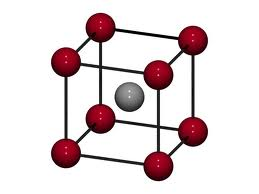
D) CH4 has London forces while C2H6 does not.

E) CH4 does not have a lower boiling point than C2H6; CH4 can hydrogen bond while C2H6 cannot.

|  |  |  |
| --- | --- | --- |
| 20. | Which of the following statements is true? | |
| A) | has a higher boiling point than |
| B) | has a lower boiling point than |
| C) | has a higher boiling point than |
| D) | has a higher boiling point than |
| E) | H—F has a lower boiling point than H—Br |

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| 21. | If a unit cell of an ionic compound has A cations at the corners and X anions on the face centers, what is the empirical formula of the compound? | |
| A) | AX |
| B) | AX3 |
| C) | AX2 |
| D) | A2X3 |
| E) | A2X |

22. How many atoms are in one unit cell of the structure below?



A) 1

B) 2

C) 3

D) 4

E) 9

23. Which substance would you expect to have the lowest boiling point?

A) CH4

B) SiH4

C) CH3Cl

D) SiH2Cl2  
 E) CH3F

24. If a unit cell of an ionic compound has A cations at the corners and the face centers and X anions in the centers of the edges of the unit cells, what is the empirical formula of the compound?

A) A4X3

B) A6X3

C) A4X2

D) A3X4

E) A3X3

25. An element crystallizes in a face-centered cubic unit cell has a density of

1.45 g/cm3. The edge length of its unit cell is 4.52 x 10-8 cm. What is the approximate atomic mass of the element in g/mol?

A) 3.35 g/mol

B) 9.87 g/mol

C) 20.2 g/mol

D) 80.7 g/mol  
 E) Not enough information

26. What is the *primary* intermolecular force in a sample of CH2Cl2?

A) London dispersion forces

B) dipole – dipole interactions

C) hydrogen bonding

D) dipole induced-dipole interactions

E) ion—dipole interactions

27. Which statement is *true*?

A) Two molecules of CH2F2 can hydrogen bond to one another.

B) Intermolecular forces are stronger than intramolecular forces.

C) A molecule of CH3OH can hydrogen bond to a molecule of CF4.

D) H2S has a higher boiling point than H2O.

E) Covalent bonds are stronger than ionic bonds.

28. Which answer lists *all* of the intermolecular forces present in an a sample of NH3

A) hydrogen bonding only

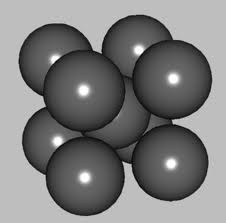
B) dipole dipole interactions and London forces

C) hydrogen bonding, dipole dipole interactions, and London forces

D) dipole dipole interactions only

E) London forces only

30. How many atoms are contained in the cubic unit cell depicted below?



A) 1

B) 2

C) 4

D) 8

E) 9

31. Which of these is *least* likely to exist as discrete molecules?

A) P2O5(g)

B) C(s)

C) H2O(s)

D) C6H12O6(s)

E) SO2(g)

32. Which of these is NOT a type of liquid crystal?

A) Smectic

B) Cholesteric

C) Nematic

D) Anisotropic

E) All of these are types of liquid crystals.

33. Which of the following has the *second highest* boiling point?

A) HF

B) GeH4

C) SiH4

D) PH3

E) CCl4

**Chapter 7**

|  |  |  |  |
| --- | --- | --- | --- |
| 1. | If an isolated system contained +5 kJ of energy, after 100 years *U* = | | |
| A) | The answer is impossible to determine |
| B) | slightly less than +5 kJ |
| C) | +5 kJ |
| D) | 0 kJ |
| E) | 5 kJ |

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| 2. | A CD player and its battery together do 500 kJ of work. The battery also releases 250 kJ of energy as heat and the CD player releases 50 kJ as heat due to friction from spinning. Assume that the system is the battery *and* the CD player together. For the system, what are the signs of heat and work? |
|  | A) q and w are both positive  B) q and w are both negative  C) q is positive and w is negative  D) q is negative and w is positive  E) Cannot determine without the internal energy of the system |

|  |  |
| --- | --- |
| 3. | You hold a gram of aluminum in one hand and a gram of copper in the other; each started at the same temperature. To/from which metal will more heat have been transferred when both have reached body temperature? Given C­s,Cu= 0.387 J/g°C, dCu= 8.93 g/cm3, Cs,Al= 0.900 J/g°C, and dAl= 2.70 g/cm3. |
|  | A) Copper, because it’s specific heat capacity is lower  B) Copper, because it’s density is higher  C) Aluminum, because it’s specific heat capacity is higher  D) Aluminum, because it’s density is lower  E) Copper and aluminum will reach body temperature at the same time |

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 5. | Calculate the standard reaction enthalpy for the following reaction.   |  |  | | --- | --- | |  | CH4(g) + H2O(g)  CO(g) + 3H2(g) | | Given: | 2H2(g) + CO(g)  CH3OH(l) | *H*° = 128.3 kJ | |  | CH4(g) + ½ O2(g)  CH3OH(l) | *H*° = 164.1 kJ | |  | 2H2O(g) 2H2(g) + O2(g) | *H*° = 483.6 kJ | | | |
| A) | +155.5 kJ |
| B) | +191.2 kJ |
| C) | +447.8 kJ |
| D) | -50.6 kJ |
| E) | +206.0 kJ |

|  |  |
| --- | --- |
| 6. | The standard enthalpy of formation of ammonium perchlorate at 298 K is  295.31 kJmol1. Which equation represents the enthalpy of formation?  A) ½N2(g) + ½Cl2(g) + 2O2(g) + 2H2(g) + 295.31 kJ NH4ClO4(s)  B) ½N2(g) + ½Cl2(g) + 2O2(l) + 2H2(g)  NH4ClO4(s) + 295.31 kJ  C) N2(g) + Cl2(g) + 4O2(g) + 4H2(g)  2NH4ClO4(s) + 590.62 kJ  D) ½N2(g) + ½Cl2(g) + 2O2(g) + 2H2(g)  NH4ClO4(s) + 295.31 kJ  E) ½N2(g) + ½Cl2(g) + 2O2(g) + 2H2(g)  NH4ClO4(s) + 147.66 kJ |

7. How many grams of aluminum are formed when 1005 kJ of heat is absorbed?

Al2O3(s) → 2Al(s) + (g) Hrxn = 1676 kJ

A) 32.36 g aluminum

B) 16.18 g aluminum

C) 0.600 g aluminum

D) 1676 g aluminum

E) 3236 g aluminum

8. Consider the following:

|  |  |  |
| --- | --- | --- |
| **Metal** | **Al** | **Cu** |
| Mass (g) | 30 | 10 |
| Specific heat capacity (J/g°C) | 0.900 | 0.385 |
| Temperature ( °C) | 40 | 60 |

When these two metals are placed in contact, which will occur?

A) Heat will flow from Al to Cu because Al has the larger specific heat.

B) Heat will flow from Al to Cu because Al has the larger mass.

C) Heat will flow from Cu to Al because Cu has the smaller heat capacity.

D) Heat will flow from Cu to Al because Cu is at a higher temperature.

E) No heat will flow in either direction.

Use this figure to answer questions 9-10:

**D**

**B**

**A**

**Heat Removed 🡪**

**Temperature (oC)**

9. For the region marked C, calculate the heat lost by 2.50 mol of H2O during the transition at 1.00 atm. The molar heat capacities are: H2O(l) = 75.4 J/mol•K; H2O(g) = 33.1 J/mol•K; H2O(s) = 37.6 J/mol•K. The enthalpy of fusion for water is 6.02 kJ/mol.

A) -8.3 kJ

B) -18.9 kJ

C) -15.0 kJ

D) -102 kJ

E) -3.76 kJ

10. The line in Stage B is longer than the line in Stage D. This is *best* explained because:

A) Two physical states are present in Stage B.

B) Liquid water has a greater heat capacity than solid water (ice).

C) The enthalpy of vaporization of water is greater than its enthalpy of fusion.

D) The average kinetic energy of the molecules is changing in Stage D, but only the average potential energy is changing in Stage B.

E) Temperature is changing more rapidly in stage B than in Stage D.

11. Which is ***not*** a state function?

A) enthalpy

B) entropy

C) Gibbs free energy

D) internal energy

E) work

|  |  |
| --- | --- |
| 12. | Calculate the final temperature when 2.50 kJ of energy is transferred as heat to 1.50 mol Rn (radon) at 298 K and 1 atm at constant volume. |
|  | A) 134C  B) 225C  C) 159C  D) 432C  E) 105C |
|  |  |

|  |  |
| --- | --- |
| 13. | Calculate the enthalpy change that occurs when 1 lb (454 g) of mercury freezes at its freezing point (234.3 K). The standard enthalpy of fusion of mercury is 2.29 kJmol1. |
|  | A) 2.29 kJ  B) 1.04  103 kJ  C) +5.18 kJ  D) +2.29 kJ  E) 5.18 kJ |

14. Which are appropriate methods for calculating Horxn?

I. Using Hess's law with given reaction enthalpies

II. Using Hof

III. Using bond enthalpies

A) I only

B) II only

C) III only

D) I and II only

E) I, II, and III

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| 15. | A system had 150 kJ of work done on it and its internal energy increased by 60 kJ. How much energy did the system gain or lose as heat? | | |
| A) | The system lost 90 kJ of energy as heat. | |
| B) | The system lost 210 kJ of energy as heat. | |
| C) | The system gained 60 kJ of energy as heat. | |
| D) | The system gained 90 kJ of energy as heat. | |
| E) | The system gained 210 kJ of energy as heat. | |

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| 17. | Which statement is true? | | |
| A) | A Styrofoam cup of hot chocolate with a lid is an example of an isolated system. | |
| B) | A bowl of soup is an example of a closed system. | |
| C) | A vacuum flask (thermos) of coffee is an example of an open system. | |
| D) | Only matter can be exchanged between system and surroundings in closed system. | |
| E) | Neither energy nor matter can be exchanged in an open system. | |
|  | |  | | |

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| --- | --- | --- | --- |
| 18. | Which value is dependent on the pathway? | | |
| A) | temperature | |
| B) | heat | |
| C) | enthalpy | |
| D) | pressure | |
| E) | Internal energy | |
|  | |  | | |

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| --- | --- | --- | --- | --- | --- |
| 19. | A piece of a newly synthesized material of mass 25.0 g at 80.0C is placed in a calorimeter containing 100.0 g of water at 20.0C. Assume the calorimeter absorbs no heat. If the final temperature of the system is 24.0C, what is the specific heat capacity of this material? The specific heat capacity for water is 4.184 JgC. | | | | |
| A) | 14.6 JgC |
| B) | 234 JgC |
| C) | 1.19 JgC |  |  | |
| D) | 0.0747 JgC |  |  | |
| E) | 16.7 JgC |  |  | |

20. The Hsub for a substance is 62.4 kJ/mol. If 153 kJ of energy are required to sublime 623 g of the substance, then what is the molar mass of the substance?

A) 9.98 g/mol

B) 95.5 g/mol

C) 1530 g/mol

D) 254 g/mol

E) 4.07 g/mol

21. Calculate the heat that must be supplied to raise the temperature of 0.250 g of N2 gas from 22.0oC to 40.5oC at constant volume. Assume that N2 behaves as an ideal gas.

A) 6.87 J

B) 4.80 J

C) 3.43 J

D) 2.06 J

E) 5.49 J

22. Calculate the standard reaction enthalpy for the following reaction:

2C3H5(NO3)3 (l) 🡪 3N2(g) + ½ O2(g) + 6CO2(g) + 5H2O(g)

**Compound Standard Molar Enthalpy of Formation (kJ/mol)**

C3H5(NO3)3(l) -364

CO2(g) -393.5

H2O(g) -241.8

H2O(l) -285.8

A) -1285 kJ

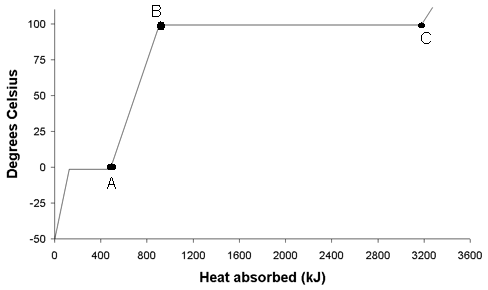
B) -3062 kJ

C) -2842 kJ

D) -4398 kJ

E) -271.3 kJ

Use this figure to answer questions 23 and 24. Assume the substance has three phases (solid, liquid, and vapor) under atmospheric pressure.



23. Which statements regarding the heating curve are ***true***?

I: Molecules at points B and C have the same kinetic energy.

II: Molecules at points A and C are the same phase.

III: Molecules at point B have greater kinetic energy than molecules at point A.

A) I only.

B) II only

C) III only

D) I and III

E) II and III

24. The slope between points A and B corresponds with:

A) Heat capacity of the substance

B) Temperature change of the substance

C) Heat added to the substance

D) Heat capacity of the liquid phase of the substance

E) None of these

25. What quantity of heat energy is required to decompose 12.6 g of liquid water to the elements?

2 H2O(l) 🡪 2 H2(g) + O2(g) H = +571.6 kJ

A) 200. kJ

B) 3.60 x 103 kJ

C) 400. kJ

D) 7.20 x 103 kJ

E) 572 kJ

26. What is the value of HB given the overall reaction:

C(s) + O2(g) 🡪 CO2(g) Hrxn = -393.5 kJ

A) CO2(g) 🡪 CO(g) + ½ O2(g) HA = +283.0 kJ

B) C(s) + ½ O2(g) 🡪 CO(g) HB = ?

A) 676.5 kJ

B) 110.5 kJ

C) -110.5 kJ

D) -676.5 kJ

E) Not enough information.

28. Which statement is *true*?

A) At constant pressure, expansion work cannot be done.

B) Heat cannot be transferred at constant volume.

C) At constant volume, *q* = H

D) At constant pressure, there can be no change in internal energy.

E) None of these statements are true.

29. Which of these processes is *endothermic*?

A) fusion (melting)

B) deposition

C) freezing

D) condensation

E) cooling a liquid

30. If expanding gases in a car engine do 451 J of work on the pistons, and the system loses 325 J to the surroundings as heat, calculate the change in internal energy (U) in J.

A)+136 J

B) -776 J

C) -96 J

D) -126 J

E) +802 J

31. A layer of copper lining a skillet weighs 125 g. What is the specific heat capacity of the copper if 13.3 kJ heat is needed to raise the temperature of the copper from 25oC to 300.oC ?

A) 9.40 J/goC

B) 2.56 J/goC

C) 0.387 J/goC

D) 0.000387 J/goC

E) 29.26 J/goC

32. For which species is the enthalpy of formation NOT zero?

A) Na(s)

B) Cl2(g)

C) N2(g)

D) H2O(g)

E) None of the above

33. Consider two metals A and B. The specific heat capacity of A has a greater value than that of B. Which statement is *true*?

A) More heat is required to raise the same mass of metal B by 1oC than metal A.

B) Given the same amount of heat energy and the same mass of both metals,

the temperature of A will rise more than the temperature of B.

C) If the same mass of both metals, beginning at the same initial temperature,

are each placed on a table, then metal B will reach room temperature

faster.

D) If the metals, both at different temperatures, are placed in contact with

each other, metal A will exhibit a greater temperature change than metal B.

E) None of these statements are true.

36. Determine the standard enthalpy of formation for C2H2 from its elements given the information below.

2C(graphite) + H2(g) 🡪 C2H2(g) Ho = ?

Given:

C(graphite) + O2(g) 🡪 CO2(g) Ho = -383.5 kJ/mol

H2(g) + ½ O2(g) 🡪 H2O(l) Ho = -285.8 kJ/mol

2C2H2(g) + 5 O2(g) 🡪 4CO2(g) + 2H2O(l) Ho = -2598.8 kJ/mol

A) +226.6 kJ/mol

B) +246.6 kJ/mol

C) +1930 kJ/mol  
 D) -3650 kJ/mol

E) -1830 kJ/mol

37. Which statement regarding the graph is *true*?



A) Regions A, C, and E represent phase changes.

B) No heat is exchanged in regions B and D.

C) Hvapis much larger than Hfus for this substance.

D) The freezing point of the substance is -25oC.

E) None of these statements are true.

38. Which process(es) can be represented by a perfectly horizontal line on a heating curve?

Melting ice

Heating CO2

Freezing ethanol

Both A & C

All of the above

39. A 25.0 g block of metal initially at 95.0oC is placed in a calorimeter containing 121 g of water initially at 27.2oC. If the final temperature of the water plus metal is 28.8oC, then what is the heat capacity of the metal? The heat capacity of water is 4.184 J/goC. Assume no heat is lost to the calorimeter or the surroundings.

A) 2.0 J J/g°C

B) 0.12 J/g°C

C) 8.4x102 J/g°C

D) 0.49 J/g°C

E) 0.21 J/g°C

40. How much heat is released when 32.0 g of O2(g) is reacted with excess copper?

(1/2)O2(*g*) + Cu(*s*)  CuO(*s*) *Ho*rxn ­= –156 kJ/mol

A) 78.0 kJ

B) 156 kJ

C) 312 kJ

D) 2500 kJ

E) Heat is not released in this reaction.

41. Which of these is NOT a state function?

A) enthalpy

B) internal energy

C) heat

D) temperature

E) pressure

42. The change in internal energy for a system that does 172 J of work on its surroundings is +311 J. Which statement about this system is *true*?

A) heat is lost by the system

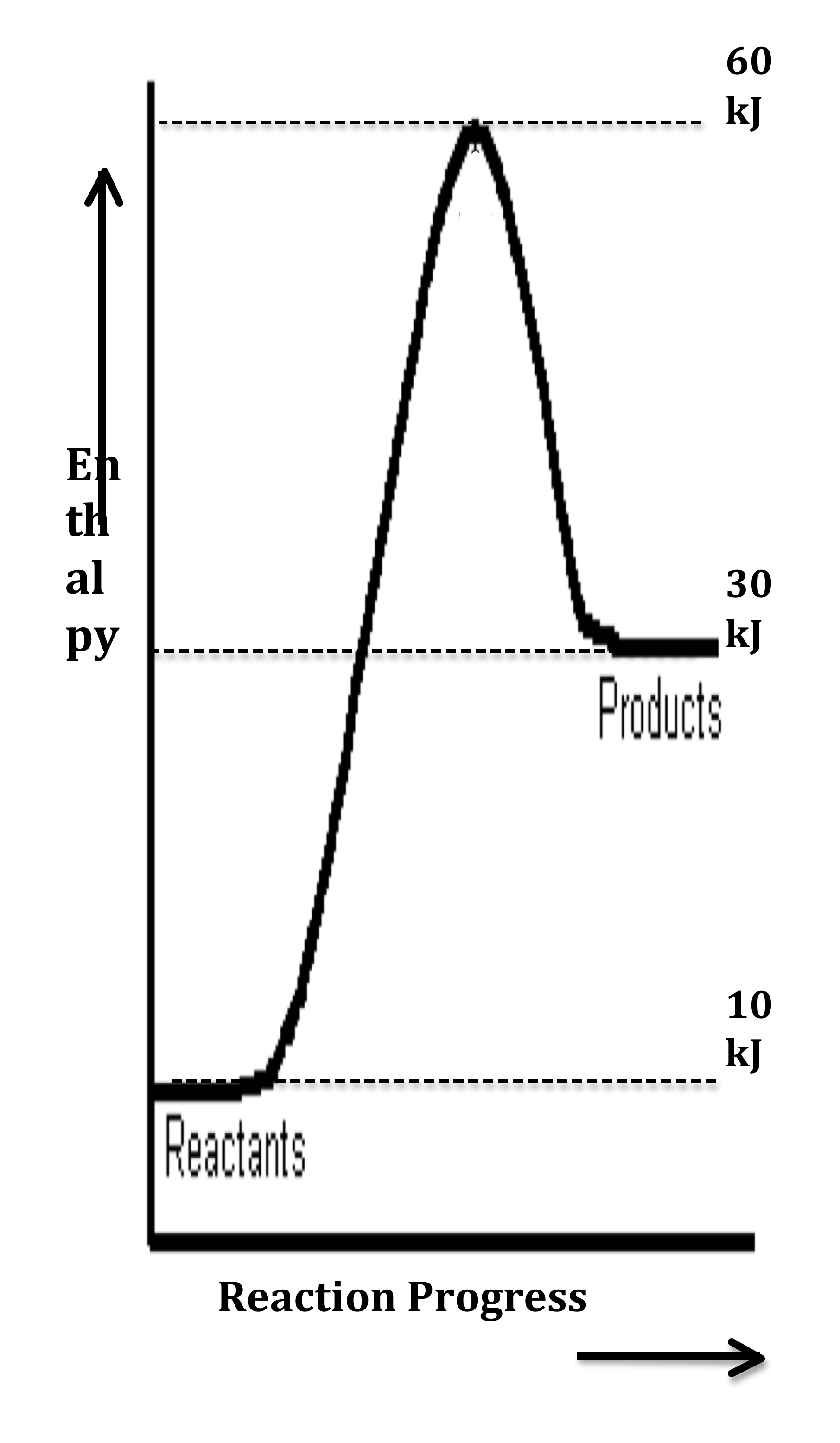
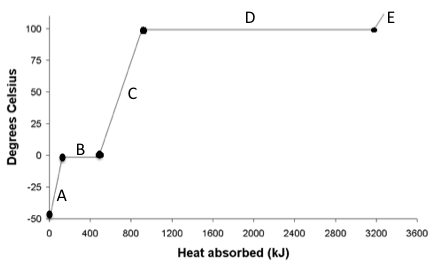
B) heat is gained by the system

C) heat is not transferred

D) q < 0

E) w > 0

43. The diagram below could represent which equation?



**Enthalpy**

A) A(g) + B(g) 🡪 AB(g) H = - 20 kJ

B) A(g) + B(g) 🡪 AB(g) H = + 30 kJ

C) A(g) + B(g) 🡪 AB(g) H = + 20 kJ

D) A(g) + B(g) 🡪 AB(g) H = - 30 kJ

E) A(g) + B(g) 🡪 AB(g) H = + 50 kJ

44. Which process is *exothermic*?

A) freezing

B) sublimation

C) fusion (melting)

D) evaporation

E) heating a liquid

**Chapter 8**

1. Which of the following reactions is spontaneous only ***below*** the equilibrium temperature?

A) NH4Br(s) + 188 kJ 🡪 NH3(g) + Br2(l)

B) NH3(g) + HCl(g) 🡪 NH4Cl(s) + 176 kJ

C) 2H2O2(l) 🡪 2H2O(l) + O2(g) + 196 kJ

D) both a and b

E) both b and c

|  |  |
| --- | --- |
| 3. | Which should have the greatest molar entropy at 298 K? |

|  |  |
| --- | --- |
| A) | CH4(*g*) |

|  |  |
| --- | --- |
| B) | H2O(*l*) |

|  |  |
| --- | --- |
| C) | NaCl(*s*) |

|  |  |
| --- | --- |
| D) | CH3CH2NH2(*g*) |

|  |  |
| --- | --- |
| E) | H2(*g*) |

4. Which statement is true?

A) All spontaneous reactions occur quickly.

B) All spontaneous processes release heat.

C) The boiling of water at 100oC and 1atm is spontaneous

D) If a process increases the freedom of motion of the particles of a system, the system’s entropy decreases.

E) The energy of the universe is constant; the entropy of the universe decreases toward a minimum.

5. Which is true for pure oxygen gas, O2(g), at 1.00 bar pressure?

A) ΔH°f > 0

B) ΔH°f < 0

C) ΔG°f > 0

D) ΔG°f < 0

E) S° > 0

|  |  |
| --- | --- |
| 6. | Calculate Δ*S*°for the reaction  SiCl4(*g*) + 2Mg(*s*) → 2MgCl2(*s*) + Si(*s*)  Substance: SiCl4(*g*) Mg(*s*) MgCl2(*s*) Si(*s*)  *S*°(J/K·mol): 330.73 32.68 89.62 18.83 |

|  |  |
| --- | --- |
| A) | -254.96 J/K |

|  |  |
| --- | --- |
| B) | -198.02 J/K |

|  |  |
| --- | --- |
| C) | 198.02 J/K |

|  |  |
| --- | --- |
| D) | 254.96 J/K |

|  |  |
| --- | --- |
| E) | 471.86 J/K |

|  |  |
| --- | --- |
| 7. | Hydrogen sulfide decomposes according to the following reaction  2H2S(*g*) → 2H2(*g*) + S2(*g*)  For this reaction Δ*S*° = 78.1 J/K and Δ*H*° = 169.4 kJ. What is the value of Δ*G*° at 298 K? |

|  |  |
| --- | --- |
| A) | -69881 kJ |

|  |  |
| --- | --- |
| B) | 48.4 kJ |

|  |  |
| --- | --- |
| C) | 99.1 kJ |

|  |  |
| --- | --- |
| D) | 146 kJ |

|  |  |
| --- | --- |
| E) | -23104.4 kJ |

8. Consider the following reaction:

A + B 🡪 C + D ΔGrxn= -200 kJ

Based on this information, which statement can be made about the reaction?

I. The reaction is spontaneous.

II. The products are more stable than the reactants.

III. The reaction is exothermic.

IV. The reactants are labile.

A) I only

B) III only

C) I and II only

D) I, II, and IV only

E) I, II, III, and IV

10. For which phase change(s) is S less than zero?

I. sublimation

II. condensation

III. freezing

IV. vaporization

A) I and IV

B) II and III

C) I only

D) II

E) I and III

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| 11. | Calculate the change in molar entropy when 2.00 mol of ozone are compressed isothermally to one quarter of its original volume. Treat ozone as an ideal gas. | | | | |
|  |  |  |  | |
| A) | 23.1 J |  |  | |
| B) | 10.0 J |  |  | |
| C) | 1.39 J |  |  | |
| D) | +10.0 J |  |  | |
| E) | +23.1 J |  |  | |

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 12. | Consider the following processes (treat all gases as ideal).   |  |  | | --- | --- | | I. | The pressure of 1 mole of oxygen gas is allowed to double isothermally. | | II. | Carbon dioxide is allowed to expand isothermally to 10 times its original volume. | | III. | The temperature of 1 mol of helium is increased 25C at constant pressure. | | IV. | Nitrogen gas is compressed isothermally to half its original volume. |   Which of these processes lead(s) to a decrease in entropy? |
|  | A) I and IV  B) I only  C) III and IV  D) II and III  E) I and II |

13. What is the change in entropy of the surroundings for the following process at 25oC?

N2(g) + 3H2(g) 🡪 2NH3(g) Ho = -92.6 kJ/mol

A) +311 J/K

B) -311 J/K

C) +3220 J/K

D) -371 J/K

E) -27.4 J/K

15. Calculate the molar heat of fusion for benzene given the following data:

Sfus = 39.1 J/mol•K

Svap = 87.8 J/mol•K

Hvap = 33.9 kJ/mol

Normal melting point = 5.5oC

Normal boiling point = 80.1oC

A) 31.0 kJ/mol

B) 7.11 kJ/mol

C) 140. kJ/mol

D) 10.9 kJ/mol

E) 215 kJ/mol

16. Use the information from the question above to determine the Gibbs free energy for the evaporation of benzene.

A) 0 kJ/mol

B) 2.90 kJ/mol

C) 3.10 x 104 kJ/mol

D) 64.9 kJ/mol

E) Not enough information

17. For which of these processes is S positive?

I. Ag+(aq) + Cl – (aq) 🡪 AgCl(s)

II. sublimation

III. 2O3(g) 🡪 3O2(g)

IV. condensation

A) I and IV

B) II and III

C) II and IV

D) II only

E) III and IV

18. Which statement best explains why water freezes only at or below 0oC at atmospheric pressure.

A) Freezing is an endothermic process with an increase in entropy.

B) Freezing is an exothermic process with decrease in entropy.

C) The sign of G for the process is always positive.

D) Freezing is an endothermic process with a decrease in entropy.

E) Freezing is an exothermic process with an increase in entropy.

19. A 2.5 mole sample of H2O­(g) is heated at constant pressure from 298 K to some final temperature. If So for the H2O(g) is +18.7 J, then what is the final temperature of the gas? Assume that H2O(g) behaves ideally.

A) 522 K

B) 235 K

C) 373 K

D) 733 K

E) 2830 K

20. What is the boiling point of ammonia in oC?

Hvap = 23.35 kJ/mol Svap  = 97.41 J/mol·K

A) 4.17 oC

B) 2.27 oC

C) 20.0 oC

D) 240. oC

E) -33.3 oC

21. Which result in an *increase* in entropy (S = +) for a system?

A) condensation

B) increasing temperature

C) dissolution of a solid in a liquid

D) both A and C

E) both B and C

22. Which reaction results in a *positive* change in entropy?

A) H2O(g) 🡪 H2O(s)

B) NaCl(aq) 🡪 NaCl(s)

C) 2NOCl(g) 🡪 2NO(g) + Cl2(g)

D) 2HBr(g) 🡪 H2(g) + Br2(l)

E) K+(aq)+ NO3-(aq) 🡪 KNO3(s)

23. Based on the thermodynamic information given, which statement is *true* for the system described?

CCl4(g) 🡪 C(s, graphite) + 2Cl2(g) Hrxn = +95.7 kJ/mol

S = +142.2 J/mol\*K

A) The reaction is always spontaneous.

B) The reaction is never spontaneous.

C) The reaction is spontaneous only above 673 K.

D) The reaction is spontaneous only below 673 K.

E) The reaction is spontaneous only above some temperature that cannot be determined.

24. What are the signs for Gibbs free energy, enthalpy, and entropy for the melting of ice at 10.0oC?

Gibbs free energy Enthalpy Entropy

A) - - -

B) - - +

C) + - -

D) - + +

E) + + +

25. Which statement is *true*?

1. All spontaneous reactions are labile.
2. Gibbs free energy can be used to determine how quickly a reaction will proceed.
3. Gibbs free energy can be used to determine if a reaction is spontaneous.
4. Spontaneity of reactions is independent of temperature.
5. More than one of these statements is true.

**Chapter 9**

2. The density of a 2.45 *M* aqueous solution of methanol (CH3OH, molar mas = 32.04 g/mol) is 0.976 g/mL. What is the molality of the solution?

A) 2.45 *m*

B) 0.0314 *m*

C) 3.18 *m*

D) 2.51 *m*

E) 2.73 *m*

3. By how much does the vapor pressure change when 218 g of glucose (non-volatile solute, molar mass = 180.2 g/mol) is dissolved in 460.0 mL of water at 300C. At this temperature, the vapor pressure of pure water is 31.82 torr, and the density of water is 1.00 g/mL.

A) The vapor pressure decreases by 1.4 torr

B) The vapor pressure increases by 1.4 torr

C) The vapor pressure decreases by 30.4 torr

D) The vapor pressure increases by 30.04 torr

E) The vapor pressure does not change.

4. Which statement is *true*?

A) CH3OH has a higher vapor pressure than CH4 at a given temperature.

B) Boiling points tend to decrease with increasing strength of intermolecular forces.

C) Boiling point and melting point follow opposite trends with respect to

intermolecular forces.

D) The easier it is for molecules to escape the liquid phase to the gas phase, the greater the vapor pressure will be.

E) H2O has a lower boiling point than CH4.

5. Molecule A has *i* = 2 while molecule B has *i* =1. The molality of an aqueous solution of molecule A is exactly half that of molecule B. Which statement is *true*?

A) The solution of molecule B will result in a higher freezing point.

B) The solution of molecule A will result in a lower vapor pressure.

C) The osmotic pressure of solution A is greater than that of B.

D) The two solutions have the same boiling point.

E) The boiling point, freezing point, osmotic pressure, and vapor pressure cannot be calculated without knowing the identities of the molecules.

6. A solution is prepared by dissolving 35.0 g of hemoglobin (Hb) in enough water to make 1.00L of solution. If the osmotic pressure of the solution is 0.0132 atm at 25oC, then what is the molar mass of hemoglobin?

A) 108 g/mol

B) 65100 g/mol

C) 2270000 g/mol  
 D) 0.00926 g/mol

E) 1.54 x 10-5 g/mol

7. The freezing point depression of an aqueous 0.100 *m* solution of MgSO4 is 0.225oC. What is the van't Hoff factor? The Kf for water is 1.86oC/*m*.

A) 1

B) 2

C) 6

D) 1.2

E) Not enough information

8. For a given solution, which concentration values will change as temperature changes?

A) mass percent

B) molarity

C) mole fraction

D) molality

E) none of these choices is correct

9. The phase diagram for a pure substance that has three phases (solid, liquid, and gas) is shown. At 150 atm and 275K, the substances exists is what phase?

A) liquid and vapor in equilibrium with one another

B) liquid

C) vapor and solid in equilibrium with one another

D) vapor

E) solid