**Quiz 0.1 – Atoms, Ions, and Molecules**
1. Please identify each of the following as an *atom*, an *element*, an *ion*, a *molecule*, or a *compound*:
   
   A. H₂O  
   B. NH₄⁺  
   C. NH₃  
   D. H₂  
   E. C₆H₁₂O₆

   A. H₂O⁺  
   B. F⁻  
   C. Na  
   D. Ar  
   E. O₂²⁻

**Quiz 0.2 – Naming Ionic Compounds**
1. Please provide a correct name for each of the following compounds:
   
   MgF₂  
   FeCl₃  
   SnF₂  
   SnF₄

   Fe(NO₃)₂  
   NH₄Cl  
   Al₂O₃  
   FeCl₂

2. Please give the *empirical formula* for the following molecules:
   
   Ammonium chloride
   Iron (III) cyanide
   Magnesium sulfate
   Ammonium sulfate

3. Please name the following ionic compounds:
   
   FeCO₃  
   Fe(CN)₃  
   MgSO₄  
   Al(OH)₃

   CsNO₃  
   Zn₃(PO₄)₂  
   (NH₄)₂SO₄

**Quiz 0.3 – Naming Molecular Compounds and Acids**
1. Please give the name or chemical formula for each of the following molecular compounds:

- Tetraphosphorus hexasulfide
  \( \text{PCl}_3 \)
- Dinitrogen tetroxide
  \( \text{Cl}_2\text{O}_3 \)
- Sulfur hexafluoride
  \( \text{SF}_6 \)

2. Please give the name or chemical formula (whichever one isn’t provided) for each of the following molecular acids:

- Hydroiodic acid
  \( \text{HI} \)
- Hydrochloric acid
  \( \text{HClO}_3 \)
- Hypochlorous acid
  \( \text{HClO}_4 \)
- Sulfurous acid
  \( \text{H}_2\text{SO}_3 \)
- Sulfuric acid
  \( \text{H}_2\text{SO}_4 \)
- Phosphoric acid
  \( \text{H}_3\text{PO}_4 \)
- Nitric acid
  \( \text{HNO}_3 \)
- Carbonic acid
  \( \text{H}_2\text{CO}_3 \)

3. Please give the name or chemical formula (whichever one isn’t provided) for each of the following molecular acids:

- Hydrobromic acid
  \( \text{HBr} \)
- Hydrofluoric acid
  \( \text{HF} \)
- Nitrous acid
  \( \text{HNO}_2 \)
- Bicarbonate
  \( \text{H}_2\text{CO}_3 \)

**Quiz 0.4 – Metric Units and Dimensional Analysis**

- How many seconds are in a day?
- How many hours are in a year?
- How many kg are there in \( 7 \times 10^{25} \) \( \mu \text{g} \)?
- How many sec are there in \( 5.2 \times 10^{18} \) ns?
- Please convert 20 km to cm.
- If you’re going 50 miles/hour, how many feet/second are you traveling?

**Quiz 0.4 – Metric Units and Dimensional Analysis (Continued)**

- On average, hummingbirds fly at 30 miles/hour. Convert this speed to m/s. (**Note:** 1 m = 3.28 feet and 1 mile = 1.61 km.)
• The Brooklyn Bridge is 277 feet tall. Convert this to meters. (Note: 1 m = 3.28 feet.)

• In a vacuum, light travels at a speed of 2.998 \times 10^8 \text{ m/s}. Convert this to miles per hour. (Note: 1 mile = 1.61 \text{ km}.)

• What is the mass in kilograms of 18.5 gallons of gasoline? (Note: the density of gasoline is 0.70 g/mL)

• What is the density, in g/cm\(^3\), of a 1.5 kg cube whose edges are each 5 cm?

• The density of aluminum is 2.7 g/mL. What volume (in L) would a 50.2 g sample of aluminum occupy?
Quiz 0.5 – Significant Figures and Percent Error

1. After carrying out the following operation, the reported value should have how many significant figures?

\[(6.943 \text{ cm} - 5.81 \text{ cm}) \times 3.23 \text{ cm}\]

A. 2  
B. 3  
C. 4  
D. 5  
E. 6

2. Which measurement below contain 3 significant figures?

I. 0.02 cm  
II. 1.32 mL  
III. 0.000500 kg

A. I and II  
B. I and III  
C. II and III  
D. II only  
E. I, II, and III

3. Which of the following has the same number of significant figures as the number 1.00310?

A. \(5 \times 10^6\)  
B. 299.782  
C. 7.92  
D. 9.234  
E. 300
**Quiz 1.1 – The Mole and Molecular Weights**

1. Arrange the following in order of increasing number of moles of Cl.

I. 1 g of KCl (74.6 g/mol)
II. 10 g of MgCl$_2$ (95.2 g/mol)
III. 0.1 mol of AlCl$_3$

A. II < I < III  
B. III < I < II  
C. I < II < III  
D. II < III < I  
E. III < II < I

2. What is the mass in grams of $2.0 \times 10^5$ atoms of naturally occurring potassium?

3. Hydrogenated vegetable oil is prepared by adding H$_2$ to double bonds of unsaturated fats (one H$_2$ per double bond). How many double bonds are present in a molecule of adrenic acid (M = 332.5 g/mol) if 16.16 g of H$_2$ is required to fully hydrogenate 665 g of the fat?

   A. 8  
   B. 4  
   C. 2  
   D. 6  
   E. 10

**Quiz 1.2 – Balancing Reactions**

1. Please balance the following equations:

   \[\_\_\_\ NH_4NO_3 \rightarrow \_\_\_ N_2 + \_\_\_ O_2 + \_\_\_ H_2O\]

   \[\_\_\_ Al(OH)_3 + \_\_\_ H_2SO_4 \rightarrow \_\_\_ Al_2(SO_4)_3 + \_\_\_ H_2O\]

2. Balance the following equations and indicate whether they are combination, decomposition, or combustion reactions:

   \[\_\_\_ C_3H_6 + \_\_\_ O_2 \rightarrow \_\_\_ CO_2 + \_\_\_ H_2O\]

   \[\_\_\_ NH_4NO_3 \rightarrow \_\_\_ N_2O + \_\_\_ H_2O\]

   \[\_\_\_ PbCO_3 \rightarrow \_\_\_ PbO + \_\_\_ CO_2\]

**Quiz 1.3 – Molecular Formulas and Percent Composition**
1. The empirical formula of a compound with molecules containing 12 carbon atoms, 14 hydrogen atoms, and 6 oxygen atoms is _________.
   A. C₁₂H₁₄O₆
   B. C₆H₇O₃
   C. CHO
   D. CH₂O
   E. C₂H₄O

2. Calculate the percent composition, by mass, of carbon in the following compounds:
   - Cocaine, C₁₇H₂₁NO₄
   - Vancomycin, C₅₅H₇₅Cl₂N₉O₂₄

3. What is the molecular formula of each of the following compounds?
   - Empirical formula CH₂, molar mass = 84 g/mol
   - Empirical formula HCO₂, molar mass = 90.0 g/mol

4. What is the empirical formula of a compound that contains 12.1% C, 16.2% O, and 71.7% Cl?
   A. COCl
   B. COCl₂
   C. C₂OCl
   D. CO₂Cl
   E. ClCO₄

5. Which of the following ores containing Cu (63.5 g/mol) has the highest percentage of Cu by mass? (The molar masses are given inside the parenthesis.)
   A. Chalcopyrite, CuFeS₂ (183.5 g/mol)
   B. Chalcocite, Cu₄S (159 g/mol)
   C. Covellite, CuS (95.6 g/mol)
   D. Cuprite, Cu₂O (143 g/mol)
   E. Tenorite, CuO (79.5 g/mol)

6. An ore containing gold was processed so that all of the gold is converted to AuCl₃ (303 g/mol). If a 4.0 g sample of an ore produced 3.03 g of AuCl₃, what is the percentage of Au (197 g/mol) in the ore?
   A. 10.0 %
   B. 49.3 %
   C. 75.8 %
   D. 98.5 %
   E. 24.6 %
Quiz 1.4 – Product and Reactant Amounts
1. How many grams of CO₂ (44.01 g/mol) are produced from the complete combustion of 0.25 mol of butane, C₄H₁₀ (58 g/mol)?
   A. 11 g  
   B. 44 g  
   C. 22 g  
   D. 88 g  
   E. 15 g

2. How many grams of Mg (24 g/mol) are needed to react with excess HCl (36.5 g/mol) to obtain 0.90 mol of MgCl₂ (95.2 g/mol) if the reaction has 90% yield?
   Mg + 2 HCl → MgCl₂ + H₂
   A. 22 g  
   B. 2.4 g  
   C. 24 g  
   D. 12 g  
   E. 48 g

Quiz 1.5 – Limiting Reactants
1. Hydrogen gas reacts with oxygen gas to form water: H₂ + O₂ → H₂O
   When the molecules in the diagram react to form the maximum amount of water, what is the limiting reactant and how many molecules of water can be formed?
   A. O₂ is limiting and 8 molecules of water will be formed  
   B. O₂ is limiting and 4 molecules of water will be formed  
   C. H₂ is limiting and 6 molecules of water will be formed  
   D. H₂ is limiting and 3 molecules of water will be formed  
   E. H₂ is limiting and 12 molecules of water will be formed

2. Ammonia is a major component of household cleaner. It is synthesized industrially via the Haber-Bosch process according to the reaction: N₂ + 3 H₂ → 2 NH₃
   If a mixture of 14 g of N₂ and 10 g of H₂ produced 0.40 moles of ammonia, what is the percent yield of the reaction?
   A. 24 %  
   B. 20 %  
   C. 80 %  
   D. 40 %  
   E. 60 %

Quiz 1.6 – Theoretical and Percent Yields
1. Iron (Fe, 55.8 g/mol) is an abundant element on Earth. An iron ore called hematite contains iron in the form of \( \text{Fe}_2\text{O}_3 \) (160 g/mol). To extract the iron from hematite, the \( \text{Fe}_2\text{O}_3 \) is reacted with carbon monoxide according to the reaction:

\[
\text{Fe}_2\text{O}_3 + 3 \text{CO} \rightarrow 2 \text{Fe} + 3 \text{CO}_2
\]

How much hematite ore, which contains 64% \( \text{Fe}_2\text{O}_3 \) by mass, is required to produce 0.20 mol of elemental iron?

A. 17.4 g  
B. 34.9 g  
C. 8.72 g  
D. 25.0 g  
E. 50.0 g

2. Pentane is a component of gasoline. How many liters of air, which contains 20% \( \text{O}_2 \) by volume, is required to completely combust 2.0 L of pentane?

\[
\text{C}_5\text{H}_{12} (g) + 8 \text{O}_2 (g) \rightarrow 5 \text{CO}_2 (g) + 6 \text{H}_2\text{O} (g)
\]

A. 20 L  
B. 40 L  
C. 60 L  
D. 80 L  
E. 100 L

**Quiz 1.6 – Theoretical and Percent Yields (Part 2)**

3. A 6.55-gram sample of an unknown element Q reacts with excess fluorine gas to form 10.35 grams of QF\(_4\). The unknown element is most likely which of the following?

A. P  
B. S  
C. Xe  
D. C  
E. Te
Quiz 2.1 – Atomic Structure and the Bohr Model

1. There are four isotopes of calcium. Their mass numbers are 40, 42, 43, and 44.
   - Please write the complete chemical symbol (superscript and subscript) for each isotope.
   - How many neutrons are in an atom of each isotope?
   - How many neutrons, protons, and electrons are present in a $^{42}\text{Ca}^{2+}$ ion?

2. Give the chemical symbol, with its superscripted mass number, for:
   - The element with 42 protons and 54 neutrons
   - The ion with 15 protons, 16 neutrons, and 18 electrons
   - The isotope of silver that has 61 neutrons

3. How many protons, neutrons and electrons does the $^{52}\text{Cr}^{3+}$ ion have, respectively?
   A. 24, 52, 26
   B. 52, 52, 55
   C. 24, 28, 26
   D. 24, 28, 21
   E. 24, 28, 25

4. The symbol of an ion is $^{238}_{92}X^{4+}$. Which of the following is true regarding the ion?
   I. It has 92 protons.
   II. It has 238 neutrons.
   III. It has 88 electrons.
   IV. The element X is uranium.
   A. I and II only
   B. I, III and IV only
   C. II and IV only
   D. I and IV only
   E. I, II, III, and IV
Quiz 2.2 – Atomic Orbitals
No quiz :)

Quiz 2.3 – Quantum Numbers
No quiz :)

Quiz 2.4 – Electron Configuration Basics
1. Which of the following sets of quantum numbers represents a valence electron in phosphorous?
   A. \( n = 4, l = 2, m_l = -2, m_s = -\frac{1}{2} \)
   B. \( n = 2, l = 0, m_l = 0, m_s = +\frac{1}{2} \)
   C. \( n = 3, l = 3, m_l = -2, m_s = -\frac{1}{2} \)
   D. \( n = 3, l = 1, m_l = -1, m_s = +\frac{1}{2} \)

2. Please answer each of the following:
   a. For \( n = 4 \), what are the possible values of \( l \) ?

   b. For \( l = 2 \), what are the possible values of \( m_l \) ?

   c. If \( m_l \) is 2, what are the possible values for \( l \) ?

3. Which of the following is a possible set of quantum numbers?
   A. \( n = 5, l = 3, m_l = -2, m_s = -\frac{1}{2} \)
   B. \( n = 3, l = 3, m_l = 0, m_s = +\frac{1}{2} \)
   C. \( n = 6, l = 5, m_l = -2, m_s = -\frac{1}{2} \)
   D. \( n = 0, l = 0, m_l = -1, m_s = +\frac{1}{2} \)

Quiz 2.5 – Noble Gas Configuration, Valence, and Energy Diagrams
*Comprehensive quiz after Lesson 2.7

Quiz 2.6 – Electron Configuration Exceptions (Cr and Cu)
*Comprehensive quiz after Lesson 2.7
**Quiz 2.5 to 2.7 – Electron Configuration**

1. Which of the following *electron configurations* represents an element in an excited state?
   - A. $1s^22s^22p^5$
   - B. $1s^22s^22p^63s^23p^64s^23d^8$
   - C. $1s^22s^22p^63s^23p^64s^13d^5$
   - D. $1s^22s^22p^63s^23p^64s^33d^5$
   - E. $1s^22s^22p^63s^23p^24s^1$

2. How many unpaired electrons are present in the $+2$ cobalt ion, Co$^{2+}$, in its ground state?
   - A. 2
   - B. 3
   - C. 4
   - D. 5
   - E. 6

3. Please indicate which element corresponds to each of the following *electron configurations*.
   a. $1s^22s^22p^6$
   b. $1s^22s^22p^5$
   c. $[\text{Kr}]5s^14d^6$
   d. $[\text{Ar}]4s^23d^{10}4p^4$

4. Why is each of the following *electron configurations* incorrect for atoms in their ground states? How could you correct each of these?
   a. $1s^22s^23s^2$
   b. $[\text{Ar}]3s^23p^3$
   c. $[\text{Ar}]4s^24d^3$

*Continue to next page...*
Quiz 2.4 to 2.7 – Electron Configuration (Continued)

5. Please write the electron configuration of each of the following elements:
   a. Oxygen
   b. Potassium
   c. Arsenic
   d. Manganese
   e. Copper

6. Please write the condensed electron configuration of each of the following elements:
   a. Fr
   b. Br
   c. Ca
   d. Zn
Quiz 2.8 – Paramagnetic vs. Diamagnetic

1. Which of the following statements is true regarding the Mn$^+$ ion?
   I. It is paramagnetic.
   II. It has five d-electrons.
   III. It has 2 valence electrons.

   A. I only
   B. II and III only
   C. II only
   D. I and III only
   E. I and II only

2. Which of the following are paramagnetic? (Circle all that apply.)
   A. Cl
   B. Pt
   C. Cr$^{3+}$
   D. Br$^-$
   E. O

3. Which of the following would interact with a magnetic field? (Circle all that apply.)
   A. Cl
   B. Pt
   C. Cr$^{3+}$
   D. Br$^-$
   E. O
Quiz 2.9 – Emission Spectra and the Photoelectric Effect

1. Electronic transitions are accompanied by absorption or emission of light. Which electronic transition corresponds to emission of light with the shortest wavelength?
   A. $n = 1 \rightarrow n = 5$
   B. $n = 3 \rightarrow n = 1$
   C. $n = 5 \rightarrow n = 4$
   D. $n = 2 \rightarrow n = 3$
   E. $n = 4 \rightarrow n = 3$

2. You have a photocell whose work function value equals 3 eV. If this photocell is struck with a photon that has an energy of 2 eV, will it produce a current?

   • In the previous question, if your photocell is struck by a photon with an energy of 5 eV, what will the resulting energy of the expelled electron be?
Chapter 3 Quiz: Molecular Structure and Geometry

1. Aspirin, shown below, is used to treat pain, fever, and inflammation. How many sigma (σ) and pi (π) bonds are present in a molecule of aspirin?

   ![Aspirin molecule diagram]

   A. 21 σ and 5 π  
   B. 16 σ and 10 π  
   C. 13 σ and 5 π  
   D. 13 σ and 10 π  
   E. 20 σ and 6 π

2. The central atom of a molecule has 3 bonding pairs and 1 lone pair. What is the molecular geometry of the molecule?

   A. trigonal planar  
   B. square  
   C. trigonal pyramid  
   D. tetrahedral  
   E. bent
3. Arrange the following in increasing order of the indicated bond angle.

I. H–P–H in PH₃:

II. H–S–H in H₂S

III. H–Si–H in SiH₄

A. II < I < III  
B. III < I < II  
C. I < III < II  
D. II < III < I  
E. I < II < III

4. Arrange the following diatomic species in order of increasing bond strength.

I. O₂  
II. N₂  
III. F₂

A. III < II < I  
B. III < I < II  
C. I < II < III  
D. II < I < III  
E. II < III < I
5. Which of the following is polar?
   I. CH₄
   II. H₂S
   III. BH₃
   IV. NH₃
   A. I and II only
   B. I and III only
   C. I, II, and III only
   D. I, II, and IV only
   E. II and IV only

6. Arrange the following in order of increasing C–O bond strength.
   I. CO₂
   II. CO₃²⁻
   III. CO
   A. I < III < II
   B. II < III < I
   C. III < I < II
   D. II < I < III
   E. II < III < I

7. The central atom of the most stable Lewis structure of SCl₂ has ______ bonding pairs (bp) and _______ lone pairs (lp).
   A. 2 bp and 1 lp
   B. 2 bp and 4 lp
   C. 4 bp and 0 lp
   D. 2 bp and 2 lp
   E. 3 bp and 1 lp
1. Answer: A

Single bonds are made of sigma bonds. Double bonds are made of one sigma bond and one pi bond; triple bonds are made of one sigma bond and two pi bonds. So for aspirin, we count all the single bonds (which each contain one sigma bond) and all of the double bonds (which each contain one sigma bond), for a total of 21 sigma bonds. Each of the double bonds contains one pi bond, which gives us a total of 5 pi bonds.

2. Answer: C

The valence shell electron pair repulsion theory predicts the geometry (shape) of molecules. It is based on the idea that atoms in a molecule arrange themselves so as to minimize electronic repulsions. In this case, the molecule resists repulsion from its lone electrons, which take up more space than bonded electrons, by compressing its bonding electrons closer together, forming a trigonal pyramidal shape.

A molecule with a central atom having 3 bonding pairs and 1 lone pair adopts a trigonal pyramid shape:

3. Answer: A

All three molecules contain four electron pairs (bonding pairs + lone pairs) around the central atoms. For SiH₄, which contains all four bonding pairs, the bond angle is 109.5° (tetrahedral shape). In PH₃, stronger repulsion by the lone pair make the H–P–H bond angle less than the ideal 109.5° angle. The lone pair on PH₃ forces the P–H bonding pairs closer to each other, resulting in a smaller bond angle. The bond angle in H₂S is much smaller, since the repulsion from two lone pairs is greater than from one.

4. Answer: B

Bond strength increases with increasing number of bonds: triple bonds are stronger than double bonds, which in turn are stronger than single bonds.

The bond in O₂ is a double bond: O=O, in N₂ a triple bond, N≡N, and in F₂ a single bond, F–F. Therefore, F₂ has the weakest bond and N₂ has the strongest.
5. Answer: E

Polar molecules are those that have nonzero dipole moments. These are molecules that contain polar bonds (bonds between atoms of different electronegativities) and are asymmetrical. In order to determine the polarity of molecules, we have to know its Lewis structure and geometry, and assess the bond polarity based on the structure.

CH$_4$ is tetrahedral, which is symmetrical, and its bonds (C–H) are considered nonpolar. Therefore, CH$_4$ is nonpolar.

\[ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \]

H$_2$S is bent, and the two H–S bonds are polar. The molecule is asymmetrical, therefore it is polar.

\[ \text{H} \quad : \text{S} \quad : \text{H} \]

The shape of BH$_3$ is trigonal planar. The molecule is symmetrical, with all three bond equivalent. Therefore, it is nonpolar.

\[ \text{H} \quad \text{B} \quad \text{H} \]

The N–H bonds on NH$_3$ are polar and the molecule is asymmetrical (trigonal pyramidal in shape). Therefore, NH$_3$ is polar.

\[ \text{H} \quad \text{N} \quad \text{H} \]

6. Answer: D

The Lewis structures of the three species given is shown below:

\[ \text{O}=\text{C}=\text{O} \]

\[ \text{O} \quad \text{C} \quad \text{O} \quad \text{O} \quad \text{O} \quad : \text{C}=\text{O} : \]

Double bonds are stronger than single bonds, and triple bonds are stronger than double bonds. The C–O bond in CO$_3^{2-}$ is the weakest. This is followed by CO$_2$, which has a C=O bond. CO has the strongest C–O bond, since it’s a triple bond.
7. Answer: D

There are 20 valence electron in SCl₂ (6 from S, 7 from each of the Cl). The S atom is the central atom, with two Cl bonded to it. We put 8 electrons on each of the Cl (two of which are bonding electrons) as shown:

```
Cl-S-Cl
```

Now, we have placed 16 total electrons, which means that we still have 4 more electrons to add into the structure for a total of 20 valence electrons. We place these 4 electrons as lone pairs on S:

```
Cl-S-Cl
```

```
**Chapter 4 Quiz: Periodic Trends**

1. The compound formed between an unknown metallic element, M, and chlorine has a formula $\text{MCl}_2$. What is the formula of the compound formed between M and oxygen?

   A. MO$_2$
   B. M$_2$O
   C. MO
   D. M$_2$O$_2$
   E. MO$_3$

2. The electrons in the 2p orbitals of which of the following is closest to the nucleus?

   A. Al
   B. Si
   C. P
   D. S
   E. The 2p electrons of these atoms have the same distance from the nucleus

3. The first electron affinity of chlorine is correctly represented by which equation?

   A. $\text{Cl}_2 + 2e^- \rightarrow 2 \text{Cl}^-$
   B. $\text{Cl} + e^- \rightarrow \text{Cl}^-$
   C. $\text{Cl} \rightarrow \text{Cl}^+ + e^-$
   D. $\text{Cl}_2 + e^- \rightarrow \text{Cl} + \text{Cl}^-$
   E. $\text{Cl}_2 \rightarrow 2 \text{Cl}^+ + 2e^-$

4. Rank the following species in order of increasing size.

   I. $\text{Ca}^{2+}$
   II. $\text{S}^{2-}$
   III. $\text{Cl}^-$
   IV. Ar

   A. II $<$ III $<$ I $<$ IV
   B. I $<$ IV $<$ III $<$ II
   C. III $<$ I $<$ II $<$ IV
   D. IV $<$ III $<$ I $<$ II
   E. They all have the same radius since they are isoelectronic
5. Consider the electron configuration of 5 species given below. Which electron configuration has the smallest radius?
   A. 1s²2s²2p⁶3s²3p²
   B. 1s²2s²2p⁶3s¹
   C. 1s²2s²2p⁶3s²3p⁶
   D. 1s²2s²2p⁶3s²3p⁶4s²
   E. 1s²2s²2p⁶

6. Which of the following statements is true?
   I. Ca²⁺ and S²⁻ are isoelectronic.
   II. An atom with electron configuration 1s²2s²2p⁶3s² will form a stable –2 ion.
   III. An atom, X, with electron configuration 1s²2s²2p⁶3s²3p⁴ forms a compound with sodium with the formula Na₂X.

   A. I and II only
   B. I and III only
   C. II and III only
   D. I, II, and III
   E. I only

7. Which of the following is most likely the first, second, and third ionization energies of Be?

   A. 520, 7300, 11800 kJ/mol
   B. 1680, 3370, 6050 kJ/mol
   C. 500, 4560, 6910 kJ/mol
   D. 900, 1760, 14800 kJ/mol
   E. 800, 2430, 3660 kJ/mol
8. Arrange the following in order of INCREASING first ionization energy.
   I. Ne 
   II. Se 
   III. Ca 
   IV. Cl 

   A. II < III < I < IV 
   B. IV < II < I < III 
   C. I < IV < II < III 
   D. III < II < IV < I 
   E. III < II < I < IV 

9. Which of the following statements is true regarding metals?

   A. Metals form basic oxides. 
   B. Alkali metals form stable +1 ions. 
   C. Alkaline earth metals, M, react with oxygen to form a compound with a formula MO. 
   D. Transition metals form brightly colored compounds. 
   E. All of these statements are true. 

10. Arrange the following species in order of increasing radius.
    I. P 
    II. Na 
    III. Ca 
    IV. K 

    A. I < II < III < IV 
    B. II < III < I < IV 
    C. II < I < IV < III 
    D. I < II < IV < III 
    E. III < IV < II < I 

11. Arrange the following in order of increasing bond dissociation energy.
    I. H–Cl 
    II. H–Br 
    III. H–F 

    A. II < III < I 
    B. III < I < II 
    C. II < I < III 
    D. I < III < II 
    E. I < II < III 

ANSWER KEY
1. Answer: C

A metal reacts with a nonmetal to form an ionic compound. If M reacts with Cl₂ to form MCl₂, this means that M forms a stable +2 ion (M²⁺) since chlorine forms Cl⁻ (there are two Cl atoms in MCl₂). When M reacts with oxygen, the product formed would have the formula MO since oxygen forms O²⁻ and we’ve determined that M is M²⁺.

2. Answer: D

The three atoms are in period 3 of the periodic table. The electrons in S experience greater attraction towards the nucleus since it has the greatest number of protons, and the same number of inner shell electrons. Thus, its electrons will be closest to the nucleus. This is the reason why atomic size decreases from left to right across a period, due to an increase in the effective nuclear charge.

3. Answer: B

Electron affinity of an atom is the energy change associated with when an electron is added to a gaseous atom in the ground state. The first electron affinity denotes the addition of one electron, the second electron affinity for the second electron, and so on.

4. Answer: B

Isoelectronic species contain the same number of electrons, thus they have the same electron configuration. The size of isoelectronic series is determined by the effective nuclear charge, which in turn is related to the number of protons in the nucleus. The more protons, the greater the effective nuclear charge is, and the smaller the atom/ion will be. Thus, Ca²⁺ is the smallest (20 protons), followed by Ar (18 protons), then Cl⁻ (17 protons), and lastly by S²⁻ (16 protons).

5. Answer: E

The electron configuration with the smallest radius is the one with electron configuration of smallest n (principal quantum number). The size of atoms/ions increase with increasing n, as observed when you go down a group in the periodic table (e.g. in group IA: the trend in atomic size is H < Li < Na < K < Rb)
6. Answer: B

I. \( \text{Ca}^{2+} \) and \( \text{S}^{2-} \) are isoelectronic.

Isoelectronic species are those that have the same number of electrons and thus have the same electron configuration. \( \text{Ca}^{2+} \) and \( \text{S}^{2-} \) are isoelectronic because they both have 18 electrons.

II. An atom with electron configuration \( 1s^22s^22p^63s^2 \) will form a stable \(-2\) ion.

Atoms lose or gain electrons forming ions so that they will have an electron configuration of a noble gas. An atom with the electron configuration \( 1s^22s^22p^63s^2 \) can give up 2 electrons to have an electron configuration of \( 1s^22s^22p^6 = \text{[Ne]} \). Therefore, it will form a \(+2\) ion.

III. An atom, \( X \), with electron configuration \( 1s^22s^22p^63s^23p^4 \) forms a compound with sodium with the formula \( \text{Na}_2X \).

Following II, an atom with electron configuration \( 1s^22s^22p^63s^23p^4 \) will gain 2 electrons to form a \(-2\) ion with electron configuration \( 1s^22s^22p^63s^23p^6 = \text{[Ar]} \). It will react with sodium, which forms a \(+1\) ion, to form an ionic compound with formula \( \text{Na}_2X \).

7. Answer: D

Be is in group IIA of the periodic table and has the electron configuration of \( 1s^22s^2 \). It has 2 electrons in the valence shell, and 2 core electrons. Removal of the first 2 valence electrons requires less energy than removal of the core electrons. Thus, we should expect to see a huge jump in the ionization energy after the removal of the second electron, or between second and third ionization energy. From the given choices, we can see that in D, there is a huge gap between the second and third ionization energy, from 1760 to 14,800 kJ/mol.

8. Answer: D

Ionization energy increases across a period left to right due to increase in effective nuclear charge, and increases up a group due to decrease in principal quantum number, \( n \). That is, ionization energy increases as you to UP and to the RIGHT. According to their locations on the periodic table, Ca will have the lowest first ionization energy, followed by Se, then Cl, then Ne.
9. Answer: E

Generally, metals form bases called base anhydride or basic oxide. When soluble, it reacts with water to form a basic solution:

$$\text{CaO} + 2 \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$$

Alkali metals are those that belong in group IA of the periodic table. They have a valence electron configuration of ns\(^1\). They tend to lose an electron to form a stable +1 cation.

Alkaline earth metals belong in group IIA of the periodic table. They form stable +2 cations. They react with oxygen to form a basic oxide of the formula MO, e.g. CaO, MgO.

10. Answer: A

Atomic radius decreases across a period (left to right) due to increase in effective nuclear charge, and increases down a group. P and Na are on the third row of periodic table while Ca and K are on the forth row. So P and Na have smaller radii than Ca and K. P is on the right of Na, so P is smaller than Na. Similarly, Ca is smaller than K.

11. Answer: C

The larger the atoms, the farther they will be from each other, and the weaker the bond between the atoms will be. Thus, bond strength increases with decreasing size of the halogen. Atomic size increases down a group. So, H–Br is the weakest bond and H–F is the strongest.
Chapter 5 Quiz: Gases

1. Two ideal gases, X and Y, are at the same temperature and have the same number of molecules. The pressure of gas X is twice that of gas Y. The volume of gas X will be _________
   _______.
   
   A. twice that of gas Y  
   B. the same as that of gas Y  
   C. one-half that of gas Y  
   D. four times that of gas Y  
   E. one-fourth that of gas Y

2. A mixture of He and Ne gases exerts a total pressure of 1.50 atm. If the mole fraction of He in the mixture is 0.40, what is the partial pressure of Ne?
   
   A. 0.30 atm  
   B. 0.70 atm  
   C. 0.90 atm  
   D. 0.60 atm  
   E. 1.20 atm

3. Which two of the following gases will exhibit approximately the same rate of diffusion?
   
   I. Ethylene, C₂H₄  
   II. Ethane, C₂H₆  
   III. Carbon Monoxide, CO  
   IV. Carbon Dioxide, CO₂  
   
   A. I and III  
   B. III and IV  
   C. I and II  
   D. I and IV  
   E. II and III
4. Arrange the following gases in order of increasing density at STP, assuming ideal behavior.
   I. Methane, CH₄
   II. Fluorine, F₂
   III. Carbon Dioxide, CO₂

   A. I < II < III  
   B. I < III < II  
   C. II < I < III  
   D. II < III < I  
   E. III < I < II

5. Consider two gases: 1.0 mol of He in a 1.0 L container at 25 °C, and 2.0 mol of Ne in a 1.0 L container at 25 °C. Which of the following statements is true regarding these two gases? Assume both gases obey ideal gas behaviors.
   I. Molecules of He travel faster than molecules of Ne.
   II. The gas molecules of Ne have a higher average kinetic energy than those of He.
   III. The two gases have the same pressure.

   A. I and III only  
   B. I only  
   C. II and III only  
   D. I and II only  
   E. I, II, and III

6. An unknown gas was found to have a density of 2.88 g/L at 27 °C and 1 atm. Which of the following could be the unknown gas?
   A. Ar  
   B. CH₄  
   C. Cl₂  
   D. NH₃  
   E. CO₂
7. A gas cylinder containing an unknown gas and methane, CH₄, developed a leak. The unknown gas was found to effuse half as fast as methane. Which of the following could be the unknown gas?

A. He  
B. NH₃  
C. SO₂  
D. N₂  
E. NO₂

8. Consider the gaseous reaction

\[ X(g) + Y(g) \rightarrow XY(g) \]

A mixture of 0.8 mole of gas X and 1 mole of gas Y was allowed to react at 300 K in a 1 L sealed container. At the end of the reaction, what is the total pressure of the mixture? The temperature and volume remains constant during the reaction.

A. 24 atm  
B. 1.3 atm  
C. 15 atm  
D. 45 atm  
E. 2.4 atm

9. A flask contains 5.5 mol of O₂ gas at 25 °C and 2.00 atm. How much O₂ gas, in moles, must be added to increase the pressure to 8.00 atm at constant temperature and volume?

A. 2.0 mol  
B. 5.5 mol  
C. 11.5 mol  
D. 9.0 mol  
E. 16.5 mol

10. A sample of an ideal gas at 298K and 3 atm has a volume of 2.5 L. What will be the volume of the same sample of gas at 1 atm pressure and 298K?

A. 0.8 L  
B. 5.0 L  
C. 2.5 L  
D. 7.5 L  
E. 12 L
ANSWER KEY

1. Answer: C

At constant temperature and number of moles (moles is a representation of number of molecules), the pressure, \( P \) of an ideal gas is inversely proportional to its volume, \( V \) (Boyle’s Law):

\[
P \propto \frac{1}{V}
\]

Thus, in order for \( X \) to have a pressure twice that of \( Y \), its volume must be one-half that of \( Y \).

2. Answer: C

Dalton’s Law of partial pressure states that the pressure of gas mixtures is the sum of the partial pressures of each of the gases in the mixture. The partial pressure of a gas in a mixture is also equal to the mole fraction of that gas times the total pressure of the mixture. Given that the mole fraction of He is 0.40, and that the only other gas in the mixture is Ne, it follows that the mole fraction of Ne is 0.60 (the mole fraction of all the components in a mixture adds up to 1). Therefore, the partial pressure of Ne is 0.60 \( \times \) 1.50 atm = 0.90 atm.

3. Answer: A

The rate of diffusion of a gas is directly proportional to its molecular weight: lighter molecules travel faster than heavier ones. Comparing the molecular weight of the given gases, we see that ethylene and carbon monoxide have the same molecular weight of 28 g/mol. Thus, we expect them to have the same rate of diffusion.
4. Answer: A

The density of an ideal gas is given by the equation

\[
\text{density} = \frac{PM}{RT}
\]

where P is the pressure, M is the molar mass, R is the ideal gas constant, and T is the temperature. At constant pressure and temperature, the larger the molar mass of a gas, the higher its density will be. Therefore, methane has the lowest density as it has the lowest M, while carbon dioxide has the highest density as the highest M.

<table>
<thead>
<tr>
<th></th>
<th>Mass (g/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I.</td>
<td>Methane, CH₄</td>
</tr>
<tr>
<td>II.</td>
<td>Fluorine, F₂</td>
</tr>
<tr>
<td>III.</td>
<td>Carbon Dioxide, CO₂</td>
</tr>
</tbody>
</table>

5. Answer: B

The speed of gas molecules depends on the temperature and the molecular weight of the gas. At the same temperature, lighter molecules travel faster than heavier ones. Thus, statement I is correct because He (M = 4 g/mol) is lighter than Ne (M = 20 g/mol). The average kinetic energy is dependent only on temperature. Since the two gases are at the same temperature, they have the same average kinetic energy. Thus, statement II is incorrect. Statement III is also incorrect. The pressure of Ne will be twice that of He, since there is twice the amount of Ne compared to He (double the moles), at the same temperature and volume.
6. Answer: C

We can calculate the molar mass of the unknown gas using the given density in order to identify it. The molar mass of a gas is related to its density via the equation:

\[ \frac{\rho M}{RT} = \frac{P M}{RT} \]

where \( \rho \) is the density, \( P \) is the pressure of the gas, \( M \) is the molar mass of the gas, \( R \) is the ideal gas constant, and \( T \) is the absolute temperature. Substituting the given values and solving for \( M \)

\[
M = \frac{RT}{P} \left( \frac{2.88 \text{ g/L}}{1 \text{ atm}} \times \frac{0.0821 \text{ L-atm/mol-K}}{300 \text{ K}} \right)
\]

\[
M \approx 3 \times 0.08 \times 300 \text{ g/mol} \approx 72 \text{ g/mol}
\]

The molar mass of the unknown gas is approximately 72 g/mol, and the gas that has the closest molar mass is Cl₂ (70.9 g/mol).
7. Answer: **C**

We can use Graham’s law of effusion to determine the molar mass of the unknown gas. Graham’s law of effusion is written mathematically as

\[
\frac{Rate_A}{Rate_B} = \sqrt{\frac{Molar \ Mass_B}{Molar \ Mass_A}}
\]

where \( r_1 \) and \( r_2 \) are the rate of effusion of gas 1 and 2, respectively, and \( M_1 \) and \( M_2 \) are the molar mass of gas 1 and 2, respectively.

Since the unknown gas effuses at a rate that is half as fast as methane (\( CH_4, \ 16 \ \text{g/mol} \)), then we can write

\[
\frac{r_x}{r_{CH_4}} = 2
\]

Using Graham’s law

\[
\frac{r_{CH_4}}{r_x} = \sqrt{\frac{M_x}{M_{CH_4}}}
\]

\[
2 = \sqrt{\frac{M_x}{16}}
\]

\[
4 = \frac{M_x}{16}
\]

\[
M_x = 64
\]

Of the given gases, \( SO_2 \) has a molar mass of 64. Therefore, the unknown gas could be \( SO_2 \).
8. Answer: A

We are given the initial amount of reactants and we need to determine the amount of reactants and products after the reaction. Then we use the ideal gas equation to solve for the total pressure.

If we start with 0.8 mol of X and 1 mol of Y, then at the end of the reaction, we would have no more X left (all of X would have reacted). The amount of Y that reacts with 0.8 mol of X is 0.8 mol, which means that we have 0.2 mol of Y left of the reactants. We also have 0.8 mol XY formed. The total mol of gas at the end of the reaction would be 1 mol (0.2 mol of Y and 0.8 mol of XY). Using the ideal gas equation, the total pressure is

\[
P = \frac{nRT}{V}
\]

\[
P = \frac{(1 \text{ mol}) \times \left(0.0821 \frac{\text{L-atm}}{\text{mol-K}}\right) \times (300 \text{ K})}{1 \text{ L}}
\]

\[
P \approx \frac{1 \times 0.08 \times 300}{1} \text{ atm} \approx 24 \text{ atm}
\]

Therefore, the total pressure of the mixture is approximately 24 atm and choice A is the closest.

9. Answer: E

At constant temperature and volume, the pressure of a gas is directly proportional to the number of moles of gas.

\[
PV = nRT
\]

\[
\frac{P_1}{n_1} = \frac{P_2}{n_2}
\]

\[
\frac{2 \text{ atm}}{5.5 \text{ mol}} = \frac{8 \text{ atm}}{n_2}
\]

\[
n_2 = 22 \text{ mol}
\]

\[
22 \text{ mol} - 5.5 \text{ mol} = 16.5 \text{ mol}
\]
Increasing the pressure from 2.00 atm to 8.00 atm requires an increase in the number of moles of O₂ from 5.5 mol to 22 mol, or an additional 16.5 mol of O.

10. Answer D

At constant temperature and amount of gas, pressure is inversely proportional to volume. If pressure decreases to 1/3 its initial amount (3 atm → 1 atm), the volume will increase by 3 times. Therefore, the new volume will be 7.5 L.
Chapter 6 Quiz: Liquids and solids

1. Molecules of acetone (shown below) exhibit which intermolecular force(s)?

   \[
   \begin{array}{c}
   O \\
   \text{CH}_3 - C - \text{CH}_3
   \end{array}
   \]

   I. dipole-dipole  
   II. London dispersion  
   III. Hydrogen bonding

   A. II only  
   B. I and II only  
   C. I and III only  
   D. II and III only  
   E. I, II, and III

2. Rank the following liquids in order of increasing boiling point.

   I. Butane, CH\textsubscript{3}CH\textsubscript{2}CH\textsubscript{2}CH\textsubscript{3} (58 g/mol)  
   II. n-propanol, CH\textsubscript{3}CH\textsubscript{2}CH\textsubscript{2}OH (60 g/mol)  
   III. acetone, CH\textsubscript{3} - C - \text{CH}_3 (58 g/mol)

   A. I < II < III  
   B. I < III < II  
   C. III < I < II  
   D. II < I < III  
   E. III < II < I

3. Arrange the following in order of increasing vapor pressure at 25 °C.

   I. Butane, CH\textsubscript{3}CH\textsubscript{2}CH\textsubscript{2}CH\textsubscript{3}  
   II. Ethanol, CH\textsubscript{3}CH\textsubscript{2}OH  
   III. Dimethyl ether, CH\textsubscript{3}OCH\textsubscript{3}

   A. II < I < III  
   B. I < III < I  
   C. II < III < I  
   D. I < II < III  
   E. III < II < I
4. Which of the following substances exhibits a dipole-dipole attractive force between its molecules?
   I. Br₂
   II. CH₂Cl₂
   III. CH₃CN
   A. I only
   B. I and II only
   C. II and III only
   D. I and III only
   E. I, II, and III

5. Which of the following molecules is/are polar?
   I. 
   II. 
   III. 
   IV. 
   A. III and IV only
   B. I and II only
   C. I, III, and IV only
   D. IV only
   E. I, II, III, and IV

6. Molecules of which of the following molecules would have intermolecular hydrogen bonding to other molecules of the same type?
7. Consider a substance that has a triple point of 1.5 atm and 250 K and a critical point of 10 atm and 700 K. Its solid-liquid line has a positive slope (leans to the right). At which of the following pressures and temperatures will the substance exist as a solid?

A. 1.0 atm and 250 K
B. 2.0 atm and 250 K
C. 1.5 atm and 300 K
D. 3.0 atm and 500 K
E. 0.5 atm and 600 K

8. Given the phase diagram below, which of the following statement(s) is/are true?

A. I and II only
B. I, II, and IV only
C. III and IV only
D. I only
E. II only
I. Starting from A, application of pressure at constant temperature will cause the substance to melt.

II. Increasing the temperature of the substance at point B under constant pressure will cause the substance to vaporize.

III. Starting from C, decreasing the temperature at constant pressure will cause the gaseous substance to condense.

A. I and II only
B. I and III only
C. II and III only
D. I only
E. III only

9. Which of the following statements is true regarding the vapor pressure of a liquid?

I. Vapor pressure of a liquid increases with increasing temperature.

II. Liquids that have higher vapor pressure boil at a lower temperature.

III. Liquids that have high vapor pressure are volatile.

A. I only
B. III only
C. I and II only
D. II and III only
E. I, II, and III
10. The vapor pressure of A and B are 200 and 50 torr, respectively. A and B form a miscible solution. Which of the following statements is true?

A. The vapor pressure of a solution of A and B is greater than 200 torr.
B. Pure A boils at a higher temperature than pure B.
C. A solution containing both A and B boils at a lower temperature than either pure A or pure B.
D. A is more volatile than B.
E. The vapor above a solution of A and B will be richer in B.

11. What type of intermolecular forces dominate when NaCl(s) dissolves in water?

A. dipole-dipole
B. H-bonding
C. dispersion
D. ion-dipole
E. ionic bond

12. Arrange the following substances in order of decreasing strength of intermolecular forces of attraction:

   a. Ar
   b. CH₃CH₂OH
   c. CH₂Cl₂

A. a > b > c
B. b > a > c
C. c > a > b
D. b > c > a
E. a > c > b
ANSWER KEY

1. Answer: B

Polar molecules exhibit dipole-dipole intermolecular forces. London dispersion forces are present in all molecules and are thus the only intermolecular force in nonpolar molecules. Hydrogen bonding is a special case of dipole-dipole and is found in molecules that have an H atom covalently bonded to O, N, or F atom. We do not have hydrogen bonding in this molecule.

Acetone is a polar molecule (due to polar C=O bond), so the intermolecular forces present in acetone are London dispersion and dipole-dipole.

2. Answer: B

The boiling point depends on the strength of the attractive forces in each substance. We need to assess the intermolecular force of attraction for each substance.

Molecules of butane experience the weakest attractive force, London dispersion. London dispersion is present in all molecules and is the only attractive force in nonpolar molecules. So butane has the lowest boiling point.

Molecules of n-propanol exhibit London dispersion and H-bonding. H-bonding is a strong type of intermolecular attractive force. So, n-propanol has the highest boiling point.

Acetone is a polar molecule. Thus, its molecules exhibit London dispersion and dipole-dipole. It has a boiling point between butane and n-propanol.

3. Answer: C

The vapor pressure is dependent on the strength of intermolecular forces in each liquid. At a given temperature, the weaker the intermolecular force in the liquid, the higher its vapor pressure.

Butane exhibits London dispersion between molecules, which is the weakest type of intermolecular force. So, butane should have the highest vapor pressure.

Ethanol molecules are held together by strong hydrogen bonding so it has the lowest vapor pressure.

Dimethyl ether has dipole-dipole interactions as a result of the difference in the electronegativity of its C and O atoms. Dipole-dipole interactions are weaker than hydrogen
bonds, but stronger than London dispersion forces. Thus, the vapor pressure of dimethyl ether will be between ethanol and butane.
4. Answer: C

Dipole-dipole intermolecular forces are found in molecules that are polar (net nonzero dipole moment). Bromine, Br₂, is a nonpolar molecule so it does not exhibit dipole-dipole. Dichloromethane, CH₂Cl₂, and acetonitrile, CH₃CN, both have polar groups and are polar. These two substances are held together by dipole-dipole intermolecular forces.

5. Answer: A

The shape and the type of bonds in a molecule determine polarity. Polar molecules have a nonzero dipole moment, and are those that contain polar bonds and have an asymmetrical shape. Lewis structures will help us assess the polarity of molecules.

BH₃ has polar B–H bonds but the bond dipoles cancel out so the molecule is nonpolar.

CH₄ is basically nonpolar. C–H bonds are considered nonpolar. Hydrocarbons (those that contains C’s and H’s only) are nonpolar molecules.

NH₃ has a polar N–H bond. It has a trigonal pyramidal geometry. Therefore, it is polar.

H₂O has polar O–H bonds, and has a bent or V-shape geometry. It is polar.

6. Answer: D

Hydrogen bonding is exhibited by molecules that contain an H atom covalently bonded to highly electronegative atoms F, N, or O. The H has to be bonded to one of these atoms directly. Although all the molecules given have O and H in their formula, only molecule A has an H that is bonded to an O and therefore exhibits hydrogen bonding.
7. **Answer: B**

From the given information, we can sketch the following phase diagram:

![Phase Diagram](image)

From the given pressures and temperatures, we can see that only B is where the substance exists as a solid.

8. **Answer: A**

The phase diagram below shows the regions where solid, liquid, and gas phases exist. The arrows indicate the changes indicated in each statement. Statement I is correct: increasing the pressure at constant temperature will cause the solid substance to melt (solid to liquid). Statement II is also correct, as increasing the temperature of B under constant pressure will cause the liquid to turn into gas. Statement III is incorrect, since decreasing the temperature at point C under constant pressure will convert the gaseous substance into a solid (deposition).
9. Answer: E

Statement I is correct. Increasing the temperature increases the average kinetic energy of molecules, allowing more molecules to escape the liquid phase and go to the gas phase increasing the vapor pressure.

The boiling point is the temperature at which the vapor pressure equals the external pressure. The higher the vapor pressure of a liquid, the less energy it requires to increase its vapor pressure to the external pressure. Thus, the liquid will have a relatively lower boiling point.

Liquids with high vapor pressure boil at a lower temperature. Thus, they are said to be volatile (boil easily).

10. Answer: D

The vapor pressure of a liquid is a measure of the amount of vapor that can exist above the liquid. The higher the vapor pressure, the easier it is to boil the liquid (lower boiling point) and the more volatile the liquid is (it evaporates easily). Thus, choice B is incorrect and choice D is correct.

A solution containing two liquids will have a vapor pressure intermediate of the vapor pressure of the pure liquids and will have a boiling point between that of the pure liquids. The more volatile component (higher vapor pressure) will be relatively richer in the vapor.

11. Answer: D

Sodium chloride (NaCl, table salt) dissolves readily in water because the interaction between water and Na\(^+\) or Cl\(^-\) ions are strong enough to overcome the attractive forces between the ions in NaCl. The oxygen-hydrogen bond in water is polar, and the O atom has a higher electron density than the H atom. Thus, O has a partial negative charge and H has a partial positive charge.

\[
\begin{align*}
\text{H} & \quad \delta^+ \\
\text{O} & \quad \delta^- \\
\end{align*}
\]

When NaCl is added to water, the water molecules attract the Na\(^+\) with its partial-negatively charged O atoms and the Cl\(^-\) with its partial-positively charged H atoms, forming a strong ion-dipole interaction.
12. Answer: D

The strength of intermolecular forces of attraction follows the trend: H-bond > dipole-dipole > dispersion. Ethanol, CH₃CH₂OH, has an H bonded to O, so it can hydrogen-bond. Dichloromethane, CH₂Cl₂, is polar and therefore has dipole-dipole. Argon is nonpolar and only has dispersion.
Chapter 7 Quiz: Chemical Solutions

1. The solubility of BaSO$_4$ in water is 2 mg/L at 25 °C. A 2 mg BaSO$_4$ sample was added to 100 mL water. Some of the solid dissolved while 1.8 mg of it remained as precipitate. The solution is said to be __________.
   - A. saturated
   - B. supersaturated
   - C. unsaturated
   - D. dilute
   - E. pure

2. Why is methanol more soluble than isopentyl alcohol (which is also known as isoamyl alcohol)?
   - A. Methanol has less steric hindrance
   - B. Water can more directly solvate methanol
   - C. Methanol has fewer hydrocarbons
   - D. Branched alkanes aggregate to form precipitates
   - E. None of the above. Methanol is less soluble than isopentyl alcohol.

3. In which of the following solvents is carbon tetrachloride, CCl$_4$, most soluble, and what intermolecular force between solute and solvent is the most important?
   - A. water, H-bonding
   - B. hexane (C$_6$H$_{14}$), dispersion
   - C. acetone (CH$_3$C(O)CH$_3$), dispersion
   - D. ethanol (C$_2$H$_5$OH), dipole-dipole
   - E. chloroform (CHCl$_3$), dispersion
4. The plot below depicts the solubility of solid X in water as a function of temperature. A 3-g sample of X was added to 100 mL of water at 25 °C. Which of the following statements is true?

A. A supersaturated solution of X is formed.
B. The 3-g sample of X will dissolve completely at 25 °C.
C. Some X will dissolve, and heating the resulting solution to 40 °C, avoiding loss of solvent through evaporation, would result in complete dissolution of A.
D. An unsaturated solution of X is formed.
E. All 2-g of X will not dissolve.

![Solubility Plot](image)

5. Rank the following aqueous solutions in order of increasing boiling point.
   I. 0.20 M CaCl₂
   II. 0.30 M sucrose
   III. 0.40 M NaCl

A. I < III < II
B. III < I < II
C. I < II < III
D. II < III < I
E. II < I < III
6. Of the following aqueous solutions, ________ has the lowest freezing point.
   A. 0.20 m acetic acid
   B. 0.20 m sucrose
   C. 0.20 m NaCl
   D. 0.20 m NH₄OH
   E. they all freeze at the same temperature

7. Chloroform dichloromethane boils at 60 °C and 40 °C, respectively. Equimolar amounts of chloroform and dichloromethane was mixed. The resulting solution will boil ______.
   A. above 61 °C
   B. below 40 °C
   C. at 60 °C
   D. at 50 °C
   E. at 40 °C

8. Determine the vapor pressure of a solution of chloroform and dichloromethane, where the mole fraction of chloroform is 0.40. The vapor pressure of chloroform is 160 mmHg, while that of dichloromethane is 540 mmHg.
   A. 160 mmHg
   B. 350 mmHg
   C. 380 mmHg
   D. 388 mmHg
   E. 540 mmHg

9. All of the following would triple the osmotic pressure EXCEPT one. Which one is the EXCEPTION?
   A. Increasing the temperature from 298K to 894K
   B. Changing the solute from 0.2M acetic acid to 0.2M CaCl₂
   C. Changing the solute from 0.2M NaCl to 0.6M Na₃PO₄
   D. Tripling the amount of NaCl solute in a solution
10. The following diagram depicts three different mixtures containing He, Ne, and Ar gas. Rank them in order of increasing mole fraction of He.

```
I
II
III
```

A. I < III < II
B. II < III < I
C. II < I < III
D. III < I < II
E. III < I = I

11. Which one of the following is most likely to dissolve in toluene, C₇H₈?

A. HBr
B. NaCl
C. C₆H₁₄
D. CH₃CN
E. CHCl₃

12. A concentrated solution of nitric acid contains 90% HNO₃ by mass and has a density of 1.50 g/mL. Which of the following statements is true?

A. 1 mL of this solution contains 1.50 g of HNO₃
B. 100 g of this solution contains 150 g of HNO₃
C. 1 L of this solution contains 90 g of HNO₃
D. 100 g of this solution contains 90 g of HNO₃
E. 1.50 mL of this solution contains 90 g of HNO₃
13. Consider a solution that contains 0.1 mol of benzene and 0.2 mol of toluene. What is the vapor pressure of this solution at 25 °C? The vapor pressure of pure benzene and pure toluene is 75 torr and 21 torr, respectively.

A. 12 torr  
B. 39 torr  
C. 48 torr  
D. 54 torr  
E. 96 torr
ANSWER KEY

1. Answer: A

A saturated solution is one that contains the maximum amount of solute that a solvent can dissolve. In such a solution, the rate of dissolution is equal to the rate of crystallization. If the amount of solute dissolved is greater than the solubility, then we have a supersaturated solution. On the other hand, if the solute dissolved is less than the solubility, then an unsaturated solution is formed.

Thus, the solution being described is a saturated solution. When 2 mg BaSO₄ is added to 100 mL only 0.2 mg of it dissolved (corresponding to a solubility of 0.2 mg/100 mL = 2 mg/L) and the rest remained as solid.

2. Answer: C

Alcohols, having an –OH group, are capable of hydrogen bonding with water, facilitating their solubility in water. Their hydrocarbon portions are nonpolar thereby decreasing their solubility. Thus, the solubility of alcohols decreases as we add more hydrocarbons.

3. Answer: B

The solubility of a solute in a solvent is determined by the relative strength of attractive forces between solute-solute, solvent-solvent, and solute-solvent. A solute is more soluble in a solvent with similar intermolecular forces to its own (like dissolves like). Carbon tetrachloride is nonpolar and its molecules are held by dispersion forces. So we expect it to dissolve through dispersion forces. All other solvents given, except hexane, have polar groups. The nonpolar CCl₄ will interact weakly with these polar solvents.

4. Answer: C

From the plot, we see that the solubility of X at 25 °C is 20 g/L or 2 g/100 mL. If we add 3 g of X into 100 mL of water, only 2 g of it will dissolve, forming a saturated solution of X. Increasing the temperature increases the solubility of X. At 40 °C, its solubility is 35 g/L or 3.5 g/100 mL. Heating the resulting solution to 40 °C should result to complete dissolution of X.
5. Answer: E

The boiling point of a solvent is increased in the presence of a nonvolatile solute. The increase in boiling point is determined by the product of the concentration, $C$ and the number of particles in solution, $i$.

When $\text{CaCl}_2$ dissolves in water, it dissociates into 1 $\text{Ca}^{2+}$ and 2 $\text{Cl}^{-}$ ions per formula unit. Similarly, NaCl dissociates into 1 $\text{Na}^{+}$ and 1 $\text{Cl}^{-}$ ion. On the other hand, sucrose does not dissociate into ions when dissolved in water. Considering both concentration and number of particles in solution, a 0.40 M NaCl will have the highest boiling point ($0.40 \text{ M} \times 2 \text{ particles} = 0.80$) while sucrose will have the lowest boiling point ($0.30 \text{ M} \times 1 \text{ particle} = 0.30$).

6. Answer: C

The freezing point of solvent is decreased when added with a nonvolatile solute. This is due to the fact the the vapor pressure of the solution is always less than that of the solvent. The freezing point depression, $\Delta T_f$, can be expressed mathematically as

$$T_f = imK_f$$

where $i$ is the van’t Hoff factor (related to how many particles the solute dissociates into in solution, 1 for nonelectrolytes) $m$ is the molality, and $K_f$ is the freezing-point depression constant characteristic of the solvent.

We are given four solutions with the same concentration. We now compare the van’t Hoff factors for each solute. Sucrose is a nonelectrolyte, i.e. it does not dissociate into ions in solution, so its $i = 1$. Acetic acid, NaCl, and NH$_4$OH all dissociates into ions in solution. Acetic acid and NH$_4$OH are weak electrolytes, they are only partially dissociated into ions in solution. NaCl on the other hand is a strong electrolyte, i.e. it completely dissociates into ions in solution. So NaCl has the highest $i$ value and thus the lowest freezing point.
7. Answer: D

From Raoult's law, the vapor pressure of a solution of two volatile liquids will be between the vapor pressure of the pure liquids, and this is given mathematically as

\[ P_{\text{solution}} = c_A P_A^o + c_B P_B^o \]

where \( P_{\text{solution}} \) is the vapor pressure of the solution, \( c_A \) and \( c_B \) are the mole fractions of A and B, respectively, and \( P_A^o \) and \( P_B^o \) are the vapor pressure of pure A and pure B, respectively.

Boiling point is defined as the temperature at which the vapor pressure equals the atmospheric pressure. Since the vapor pressure of the solution is somewhere in between the vapor pressure of the pure liquids, then we expect the boiling point of the solution to be in between the boiling points of the pure liquids. If equimolar amounts were not mixed (ie. more of one of the solutions) the boiling point of the solution will be closer to the boiling point of the component with the higher mole fraction.

8. Answer: D

We use Raoult's law to determine the vapor pressure of a solution containing two volatile liquids:

\[ P_{\text{solution}} = c_A P_A^o + c_B P_B^o \]

where \( P_{\text{solution}} \) is the vapor pressure of the solution, \( c_A \) and \( c_B \) are the mole fractions of A and B, respectively, and \( P_A^o \) and \( P_B^o \) are the vapor pressure of pure A and pure B, respectively.

We are given the vapor pressure of chloroform and dichloromethane. The mole fraction of chloroform in solution is given as 0.40, which means that the mole fraction of dichloromethane is 0.60. Substituting these into the equation above gives

\[ P_{\text{solution}} = 0.40 (160 \text{ mmHg}) + 0.60 (540 \text{ mmHg}) = 64 + 324 \text{ mmHg} = 388 \text{ mmHg} \]
9. Answer: C

Osmotic pressure is the pressure that must be applied to prevent net movement of water from lower concentration to higher concentration. The osmotic pressure, \( \Pi \), of a solution is given by

\[
\Pi = iMRT
\]

where \( i \) is the van’t Hoff factor, \( M \) is the molarity of the solution, \( R \) is the ideal gas constant, and \( T \) is the absolute temperature. The van’t Hoff factor is determined by the number of particles that the solute dissociates into in solution (\( i = 1 \) for nonelectrolyte).

Tripling the temperature (A), tripling the van’t Hoff factor (B), or tripling the molarity (D) would triple the osmotic pressure. In C), we double the van’t Hoff factor and also triple the molarity, which would result in six times the osmotic pressure.

10. Answer: D

We are given three mixtures containing different amounts of gases. We count the number of atoms of He gas and the total number of atoms of gas in each mixture and use the equation for the mole fraction,

\[
c = \frac{\text{number of mole of a component}}{\text{total mole in the mixture}}
\]

For I, there are 3 He and 9 total gas molecules, so the mole fraction of He is \( \frac{3}{9} = \frac{1}{3} \).

For II, 2 He and 4 total gas molecules are in the mixture, so the mole fraction of He is \( \frac{2}{4} = \frac{1}{2} \).

For III, there are 2 He and 8 total gas molecules, so He has a mole fraction of \( \frac{2}{8} = \frac{1}{4} \).

11. Answer: C

A substance is expected to dissolve in a solvent if the strength of their intermolecular forces are similar. Toluene, like other hydrocarbons, is nonpolar and is held by dispersion forces. Other nonpolar substances, such as \( \text{C}_6\text{H}_{14} \), are thus expected to dissolve in toluene. HBr, \( \text{CH}_3\text{CN} \) and \( \text{CHCl}_3 \) are polar substances and are held by dipole-dipole forces, and NaCl are held by ionic bonds.
12. Answer: **D**

The percent by mass of a solution is the mass of solute dissolved in 100 g of solution (mass of solute + mass of solvent).

\[
\% \text{mass} = \frac{\text{g solute}}{\text{g solution}} \times 100\%
\]

Thus, a 90% HNO\(_3\) mass means that there are 90 g of HNO\(_3\) dissolved per 100 g of solution.

The density of a solution is the mass of solution per mL of solution. A 1.50 g/mL solution means the solution weighs 1.50 g for every mL of solution.

13. Answer: **B**

The partial pressure above the solution are given by Raoult’s law:

\[
P_{\text{benzene}} = X_{\text{benzene}} P^o_{\text{benzene}} \quad \text{and} \quad P_{\text{toluene}} = X_{\text{toluene}} P^o_{\text{toluene}}
\]

where \( P \) is the partial pressure of each component above the solution, respectively, \( X \) is the mole fraction of each component in the solution and \( P^o \) is the vapor pressure of pure liquid component. The total vapor pressure above the solution is

\[
P_{\text{total}} = P_{\text{benzene}} + P_{\text{toluene}}
\]

So,

\[
P_{\text{total}} = X_{\text{benzene}} P^o_{\text{benzene}} + X_{\text{toluene}} P^o_{\text{toluene}} = \frac{n_{\text{benzene}}}{n_{\text{total}}} P^o_{\text{benzene}} + \frac{n_{\text{toluene}}}{n_{\text{total}}} P^o_{\text{toluene}}
\]

\[
P_{\text{total}} = \left(\frac{0.1}{0.1 + 0.2}\right) \times (75 \text{ torr}) + \left(\frac{0.2}{0.1 + 0.2}\right) \times (21 \text{ torr}) = 25 + 14 \text{ torr} = 39 \text{ torr}
\]
Chapter 8 Quiz: Chemical Kinetics

1. Consider the following reaction: 2 A + B → 3 C

What is the rate of formation of C at an instant when A is reacting at a rate of 0.60 M/s?

A. 0.10 M/s  
B. 0.20 M/s  
C. 0.40 M/s  
D. 0.90 M/s  
E. 1.80 M/s

2. The reaction CH$_4$(g) + 2 O$_2$(g) → CO$_2$(g) + 2 H$_2$O(l) was studied by monitoring the concentration of CH$_4$(g) as the reaction progresses. The data from this monitoring is shown in the table. What is the average rate of formation of H$_2$O(l) from 0 to 200 s?

<table>
<thead>
<tr>
<th>Time, s</th>
<th>0</th>
<th>50</th>
<th>100</th>
<th>200</th>
</tr>
</thead>
<tbody>
<tr>
<td>[CH$_4$(g)], M</td>
<td>1.00</td>
<td>0.841</td>
<td>0.707</td>
<td>0.600</td>
</tr>
</tbody>
</table>

A. $1.00 \times 10^{-3}$ M/s  
B. $2.00 \times 10^{-3}$ M/s  
C. $3.00 \times 10^{-3}$ M/s  
D. $4.00 \times 10^{-3}$ M/s  
E. $6.00 \times 10^{-3}$ M/s

3. The reaction 2 A + B → A$_2$B is zero order with respect to A and second order with respect to B. Doubling the concentration of A and increasing the concentration of B by a factor of three would __________.

A. double the rate  
B. increase the rate by a factor of six  
C. decrease the rate by half  
D. decrease the rate by 1/6  
E. increase the rate by a factor of nine
4. Consider the reaction \(2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)\)

Suppose equimolar amounts of NO and O\(_2\) were placed in a sealed container and allowed to react. Which of the following statements regarding the reaction is correct?

I. The rate of the reaction increases as the reaction progresses.
II. Decreasing the volume of the container increases the rate.
III. The rate constant increases with time.

A. I and II only  
B. I and III only  
C. II and III only  
D. I, II, and III  
E. II only

5. The reaction \(\text{NO}_2(g) + \text{CO}(g) \rightarrow \text{NO}(g) + \text{CO}_2(g)\) has a rate law that is second order overall. What are the units of the rate constant?

A. M s\(^{-1}\)  
B. M\(^{-1}\) s\(^{-1}\)  
C. s\(^{-1}\)  
D. M s  
E. M\(^{-2}\) s\(^{-1}\)

6. Consider the reaction \(\text{S}_2\text{O}_8^{2-} + 3\text{I}^- \rightarrow 2\text{SO}_4^{2-} + \text{I}_3^-\)

Determine the rate law for this reaction given the following experimental rate data.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>([\text{S}_2\text{O}_8]^{2-}, \text{M})</th>
<th>([\text{I}^-], \text{M})</th>
<th>Instantaneous Rate, (x 10^3) M/s</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.500</td>
<td>0.200</td>
<td>2.0</td>
</tr>
<tr>
<td>2</td>
<td>0.500</td>
<td>0.400</td>
<td>4.0</td>
</tr>
<tr>
<td>3</td>
<td>1.00</td>
<td>0.200</td>
<td>4.0</td>
</tr>
</tbody>
</table>

A. Rate = \(k[\text{S}_2\text{O}_8][\text{I}^-]^3\)  
B. Rate = \(k[\text{S}_2\text{O}_8][\text{I}^-]\)  
C. Rate = \(k[\text{S}_2\text{O}_8]^2[\text{I}^-]\)  
D. Rate = \(k[\text{S}_2\text{O}_8][\text{I}^-]^2\)  
E. Rate = \(k[\text{I}^-]\)
7. For the reaction

$$A \rightarrow B + C$$

the concentration of A was monitored as a function of time and the following plots were generated. What is the rate law and the value of the rate constant for this reaction?

The equation below each graph is the equation of the best fit line.

$$y = -1.79 \times 10^{-3} x + 0.642$$

$$y = -0.01 x - 1.84 \times 10^{-8}$$

$$y = 0.161 x - 14.75$$

A. Rate = $k[A]$, $k = 0.01$ s$^{-1}$
B. Rate = $k[A]^2$, $k = 0.161$ M$^{-1}$ s$^{-1}$
C. Rate = $k$, $k = 1.79 \times 10^{-3}$ M s$^{-1}$
D. Rate = $k[A]$, $k = 1.84 \times 10^{-8}$ s$^{-1}$
E. Rate = $k$, $k = 0.642$ M s$^{-1}$

8. Which of the following statements is true regarding a first order reaction?

A. The half-life of a first order reaction increases with increasing initial concentration of the reactant.
B. The half-life is dependent of the initial concentration of the reactant.
C. The half-life is independent of the initial concentration of the reactant.
D. The rate constant of a first order reaction increases with time.
E. The rate of a first order reaction increases with time.
9. Consider the reaction \( C(s) + CO_2(g) \rightarrow 2 \text{ CO}(g) \)

This is a first order reaction with respect to \( CO_2 \) and zero order with respect to \( C \). The following diagram depicts a reaction mixture in a sealed container as a function of time. What is the half-life of the reaction?

A. 20 min  
B. 25 min  
C. 50 min  
D. 100 min  
E. 200 min
ANSWER KEY

1. Answer: D

The rate of a chemical reaction is related to the stoichiometry of the balanced chemical formula. This is seen in the following equation:

\[
\text{Rate of reaction} = \frac{1}{2} \frac{[\text{A}]}{t} = \frac{1}{1} \frac{[\text{B}]}{t} = \frac{1}{3} \frac{[\text{C}]}{t}
\]

We are given the rate of A reacting is \( \frac{[\text{A}]}{t} = -0.60 \text{ M/s} \), and we need to solve \( \frac{[\text{C}]}{t} \).

Looking at the balanced formula, we see that for every 2 moles of A that react, 3 moles of C are formed.

From the equation above, we see that \( \frac{[\text{C}]}{t} = \frac{3}{2} \frac{[\text{A}]}{t} = \frac{3}{2} \times 0.60 \text{ M/s} = 0.90 \text{ M/s} \).

2. Answer: D

The rate of formation of \( \text{H}_2\text{O} \) is related to the rate of consumption of \( \text{CH}_4 \) by the following equation:

\[
\frac{1}{1} \frac{[\text{CH}_4]}{t} = \frac{1}{2} \frac{[\text{H}_2\text{O}]}{t}
\]

We can calculate for \( \frac{[\text{CH}_4]}{t} \) from 0 to 200 s, using the date in the table given

\[
\frac{[\text{CH}_4]}{t} = \frac{[\text{CH}_4]_{200 \text{ s}}}{200 \text{ s}} = \frac{0.600 \text{ M}}{200 \text{ s}} = 3 \times 10^{-3} \text{ M/s}
\]

It follows that \( \frac{[\text{H}_2\text{O}]}{t} = \frac{2}{1} \frac{[\text{CH}_4]}{t} = 4.0 \times 10^{-3} \text{ M/s} \).
3. Answer: E

The rate law for the reaction can be written as

\[ \text{Rate} = k[A]^0[B]^2 = k[B]^2 \]

Thus, doubling the concentration of A will not affect the rate, while increasing the concentration of B by a factor of three will increase the rate by a factor of nine.

\[ \text{Rate}_{\text{new}} = k[3B]^2 \]
\[ \text{Rate}_{\text{new}} = 9 \times \text{Rate}_{\text{old}} \]

4. Answer: E

The rate of a chemical reaction depends on the concentration of reactants. As the reaction progresses, reactants are converted into products, so their concentration decreases. Therefore, the rate decreases with time. Decreasing the volume of the container increases the pressure of gaseous reactants (pressure and volume are inversely related), thereby increasing the reaction rate. The rate constant does not change with time. Only by changing the temperature does the value of the rate constant change.

5. Answer: B

The units for rate is M/s, and the unit for concentration is M. The rate law for the reaction can be written as

\[ \text{Rate} = k[\text{reactant}]^2 \]

Now let’s substitute in our units

\[ \frac{M}{s} = k[M]^2 \]

\[ \frac{M}{s} = kM^2 \]

Therefore, in order to have the units we require so that the rate will have units M/s, the rate constant must have units of M\(^{-1}\)s\(^{-1}\).

When M\(^{-1}\)s\(^{-1}\) is multiplied by M\(^2\) on the right hand side of the equation, we end up with M/s on the left side of the equation.
6. Answer: B

The rate of this reaction is given by: \( \text{Rate} = k[S_2O_8^2][I^-]^y \)

where \( x \) and \( y \) are the order of the reaction with respect to each of the reactant.

From experiment 1 and 2, we see that doubling the concentration of \( I^- \) while keeping the concentration of \( S_2O_8^{2-} \) constant also doubles the rate. This means that the reaction is first order with respect to \( I^- \).

Similarly, from experiment 1 and 3, doubling the \( [S_2O_8^{2-}] \) while keeping \( [I^-] \) constant also doubles the rate. This means that the reaction is first order with respect to \( S_2O_8^{2-} \) as well.

Therefore, our rate law is

\[ \text{Rate} = k[S_2O_8][I^-] \]

7. Answer: A

The plot of ln \([A]\) vs. time is linear for a first order reaction.

For second order reaction, \( 1/[A] \) vs. time is linear.

For a zero order reaction, \([A]\) varies linearly with time.

The rate constant is the absolute value of the slope of the best fit line of the graph.

Since the given graph of ln \([A]\) vs. time is linear, then the reaction is first order with respect to \( A \).

The equation of a line of best fit is \( y = mx + b \), with \( m \) representing the slope. We see in our data that \( y = -0.01x - 1.84 \times 10^{-8} \)

The absolute value of the slope tells us our rate constant = 0.01.
8. Answer: C

Remember, half life represents the time it takes for a substance to decrease to one half the initial concentration.

For a first order reaction with reactant A, rate = $k[A]$.

The half-life for a first order reaction is given as

$$t_{1/2} = \frac{\ln 2}{k}$$

Since there is no concentration term in the equation above, the half life is independent on the concentration of reactant.

The rate constant only varies with temperature. It doesn’t matter what the concentration of the reactant is, the half life time is constant.

Since the concentration of $[A]$ decreases with time, so does the rate.

9. Answer: C

The half-life of a reaction is the time it takes for the concentration of a reactant to reach half its initial value. We can count the molecules in each mixture and determine at what point does the amount of CO$_2$ reach half its original concentration. In A, we start with 8 CO$_2$ molecules. In B, we have 6 CO$_2$ and 4 CO, which means that only 2 CO$_2$ has reacted after 20 min. In C, we have 4 CO$_2$ and 8 CO; half of the initial CO$_2$ has been reacted. Thus, the half-life of the reaction is 50 min.
Chapter 9 Quiz: Chemical Equilibria

1. Which of the following statements is true regarding chemical equilibrium?
   I. The concentrations of reactants and products at equilibrium are constant, which means that all chemical processes have ceased at equilibrium.
   II. The concentrations of reactants and products are always equal at equilibrium.
   III. The equilibrium constant will change as temperature changes.

   A. I and III only
   B. II and III only
   C. I and II only
   D. II only
   E. III only

2. Consider the reaction
   \[ A(g) + 2B(g) \rightleftharpoons AB_2(g) \]
   which has an equilibrium constant of 0.50. Which of the following diagrams represents a mixture of the reaction at equilibrium?

   A. III only
   B. I only
   C. I and III only
   D. II and III only
   E. I, II, and III
3. The reaction

\[ \text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \]

has an equilibrium constant \( K_c = 40 \) at 298 K. Which diagram below represents a reaction mixture that proceeds to the right to achieve equilibrium?

A. I only  
B. III only  
C. II and III only  
D. I and III only  
E. I, II, and III

4. A mixture of \( \text{Br}_2 \) and \( \text{Cl}_2 \) was allowed to react in a closed container and attain equilibrium according to the reaction

\[ \text{Br}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{BrCl}(\text{g}) \]

Which of the following changes would increase the yield of the reaction?

I. increase the container volume  
II. add more \( \text{Br}_2(\text{g}) \)  
III. remove some \( \text{Cl}_2(\text{g}) \)

A. I and II only  
B. II only  
C. I and III only  
D. II and III only  
E. I only
5. Consider the reaction at equilibrium

\[ 3 \text{Fe(s)} + 4 \text{H}_2\text{O(g)} \rightleftharpoons \text{Fe}_3\text{O}_4(s) + 4 \text{H}_2(g) \]

Which of the following would happen when more \( \text{H}_2\text{O(g)} \) is added to the system at equilibrium?

A. The system will adjust by shifting the equilibrium to the reactant side.
B. The amount of Fe(s) will increase
C. The equilibrium constant, \( \text{K}_{eq} \) will increase.
D. The amount of \( \text{H}_2(g) \) will increase
E. Adding \( \text{H}_2\text{O} \) will have no net effect.

6. Ammonia, \( \text{NH}_3(g) \) and oxygen, \( \text{O}_2(g) \) were allowed to react until equilibrium is established according to the reaction

\[ 4 \text{NH}_3(g) + 5 \text{O}_2(g) \rightleftharpoons 4 \text{NO(g)} + 6 \text{H}_2\text{O(g)} \quad \Delta H^\circ = -904 \text{ kJ} \]

Of the following, _____________ will cause the equilibrium to shift to the left, decreasing product yield.

A. increasing the temperature
B. increasing the container volume
C. adding a catalyst
D. adding more \( \text{O}_2(g) \)
E. removing some \( \text{NO(g)} \)

7. Of which of the following salts is the solubility higher in 0.10 M HCl solution than in pure water?

I. \( \text{AgCl} \)
II. \( \text{BaF}_2 \)
III. \( \text{PbI}_2 \)

A. I only
B. II and III only
C. I and II only
D. II only
E. I, II, and III
8. Of the following sulfide salts, ________ has the lowest molar solubility in water.

A. CuS, \( K_{sp} \ 8 \times 10^{-34} \)
B. CdS, \( K_{sp} \ 1 \times 10^{-24} \)
C. NiS, \( K_{sp} \ 3 \times 10^{-16} \)
D. ZnS, \( K_{sp} \ 2 \times 10^{-22} \)
E. PbS, \( K_{sp} \ 3 \times 10^{-25} \)

9. Which of the following will increase the solubility of FeS?
   I. increasing the temperature
   II. decreasing the pH
   III. adding \( Na_2S \)

A. I only
B. II and III only
C. I and II only
D. I and III only
E. I, II, and III

10. Exactly 4.5 moles of \( N_2 \) and 2.5 moles of \( H_2 \) were added to a closed 1 L container at 278 K and allowed to react to form ammonia gas, \( NH_3 \). At equilibrium, there is 1 mole of \( NH_3 \). What is the equilibrium constant for this reaction at this temperature?

A. \( 1.4 \times 10^{-2} \)
B. \( 2.78 \times 10^{-2} \)
C. \( 3.56 \times 10^{-2} \)
D. \( 8.47 \times 10^{-2} \)
E. \( 2.5 \times 10^{-1} \)

11. A closed container contains 0.5 M of \( H_2 \) and 1 M of \( I_2 \), and the two are allowed to react according to the following reaction until equilibrium.

\[ H_2 + I_2 \rightarrow 2 \text{ HI} \]

The equilibrium constant, \( K_{eq} \), for this reaction is \( 3.2 \times 10^{-3} \). What is the equilibrium concentration of \( HI \)?

A. \( 4 \times 10^{-4} \) mol
B. \( 2 \times 10^{-2} \) mol
C. \( 4 \times 10^{-2} \) mol
D. \( 8 \times 10^{-4} \) mol
E. \( 16 \times 10^{-4} \) mol
12. In a 5 L container, 15 moles of SO\(_2\) and 10 moles of O\(_2\) are mixed and allowed to react according to the following process:

\[
2 \text{SO}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{SO}_3 (g)
\]

The equilibrium constant, \(K_{eq}\), is \(2 \times 10^{-6}\). Determine the equilibrium concentration of the product, SO\(_3\).

A. \(3 \times 10^{-6}\)  
B. \(6 \times 10^{-6}\)  
C. \(3 \times 10^{-3}\)  
D. \(6 \times 10^{-3}\)  
E. \(15 \times 10^{-3}\)
ANSWER KEY

1. Answer: E

At equilibrium, the concentrations of reactants and products remains unchanged. However, this does not mean that the reaction has stopped. Reactants are still converted into products, and products are converted back to reactants. The concentrations of reactants and products remains constant because the rate at which the reactant is consumed is the same as the rate at which it is being produced. Thus, statement I is incorrect.

Statement II is also incorrect, since it is the rate of forward and reverse reactions, and not the concentrations, that are equal at equilibrium.

The equilibrium constant gives the ratio of the composition of the products and reactants at equilibrium. It does not depend on the initial composition of the reaction. It does depend on temperature. Therefore, statement III is correct.

2. Answer: B

We count the number of reactants and products from each mixture and substitute these into the reaction quotient expression, \( Q_c \). If \( Q_c = K_c \), then the reaction is at equilibrium.

\[
Q_c = \frac{[AB_2]^2}{[A][B]^3}
\]

For I, there are 2 A, 2 B, and 4 \( AB_2 \) molecules, so

\[
Q_c = \frac{[AB_2]^2}{[A][B]^3} = \frac{4}{2 \times (2)^2} = \frac{4}{8} = 0.50
\]

Thus, I is at equilibrium as \( Q_c = K_c \).

For II, there are 1 A, 3 B, and 3 \( AB_2 \) molecules, so

\[
Q_c = \frac{[AB_2]^2}{[A][B]^3} = \frac{3}{1 \times (3)^2} = 0.33
\]

\( Q_c \) does not equal \( K_c \) so II is not at equilibrium.

For III, there are 2 A, 2 B, and 8 \( AB_2 \) molecules, so
\[ Q_c = \frac{[AB_2]}{[A][B]^2} = \frac{8}{2 \times (2)^2} = 1 \]

\( Q_c \) does not equal \( K_c \) \( \text{III} \) is not at equilibrium.

3. Answer: \( C \)

We are given the equilibrium constant. We can calculate for the reaction quotient, \( Q_c \), by counting the reactants and products in each diagram. If \( Q_c < K_c \), then the reaction will proceed to the right to achieve equilibrium.

For \( \text{I} \), there are 1 \( \text{N}_2 \), 1 \( \text{H}_2 \), and 7 \( \text{NH}_3 \). Substituting these into the reaction quotient expression gives

\[ Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{7^2}{1 \times 1^3} = 49 > K_c \]

Thus, this reaction mixture will proceed to the left until equilibrium is established.

For \( \text{II} \), \( Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{2^2}{4 \times 1^3} = 1 < K_c \), so the reaction proceeds to the right.

For \( \text{III} \), \( Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{0^2}{3 \times 5^3} = 0 < K_c \), so the reaction proceeds to the right to achieve equilibrium.
4. Answer: B

We use Le Chatelier’s principle to predict the effect of the changes in the equilibrium position of the reaction.

Increasing the container volume has an effect of decreasing the total pressure, so the system will adjust and shift the equilibrium position towards the side with more moles of gas to increase the pressure. Since both sides have equal number of moles of gas, then changing the volume will have no effect on the equilibrium position.

When more Br$_2$(g) is added, the system will adjust to decrease the concentration of Br$_2$(g) by shifting the equilibrium position to the right, forming more products.

If [Cl$_2$] is decreased, the system will adjust to increase the [Cl$_2$], so the equilibrium shifts to the left, decreasing the amount of products.

5. Answer: D

Le Chatelier’s principle predicts that when additional H$_2$O(g) is added, the system will adjust to decrease the concentration of H$_2$O(g), so the equilibrium will shift towards the product side, increasing the amounts of Fe$_3$O$_4$ and H$_2$. The equilibrium constant will not be affected by changes in concentration. Only temperature can affect the value of the equilibrium constant.

6. Answer: A

The reaction given is exothermic, so we can think of heat as one of the products. Increasing the temperature increases the heat on the product side, which will shift the equilibrium position towards the reactants. Increasing container volume has an effect of decreasing the total pressure. The system will adjust to increase the pressure, so the equilibrium will shift towards the side with more moles of gas, which in this case is the product side. Adding a catalyst will have no net effect on the position of the equilibrium since it affects the rate of both the forward and reverse reactions equally. A catalyst simply helps us reach equilibrium more quickly, but it does not change the position of equilibrium. Adding more O$_2$(g) and removing some NO(g) will shift the equilibrium towards the product side.
7. Answer: D

Salts whose anions are basic (anions that are the conjugate base of a weak acid are basic) will be more soluble in acidic solution than in a neutral/basic solution. Of the given choices, only BaF$_2$ will be more soluble in 0.10 M HCl, because F$^-$ is basic (F$^-$ is the conjugate base of the weak acid HF).

\[
\text{BaF}_2(s) \rightleftharpoons \text{Ba}^{2+}(aq) + 2 \text{F}^-(aq)
\]

In acidic solution, F$^-$ will react with the acid (0.10 M HCl solution), thereby decreasing the F$^-$ concentration. The system will adjust to increase the concentration of F$^-$, so the equilibrium will shift to the right, increasing the solubility of BaF$_2$.

HCl and HI are both very strong acids, and therefore the anion salts are very weakly basic.

8. Answer: A

The solubility of salt is determined by its solubility product constant, $K_{sp}$. For salts with the same number of cations and anions in its formula, we can use $K_{sp}$ to determine the relative solubility. The higher the $K_{sp}$, the greater the amount of ions in solution which means the higher the solubility of the salt is.

For the type of salts given, the solubility can be calculated from the $K_{sp}$ expression. For example, for CuS, the pertinent equilibrium reaction is

\[
\text{CuS}(s) \rightleftharpoons \text{Cu}^{2+}(aq) + \text{S}^{2-}(aq)
\]

and the $K_{sp}$ expression is

\[
K_{sp} = [\text{Cu}^{2+}] [\text{S}^{2-}] = s^2
\]

where $s$ is the solubility. From this equation, we see that $s = \sqrt{K_{sp}}$.

Looking at our answer choices, [A] has the smallest $K_{sp}$ and therefore the lowest solubility.

9. Answer: C
Increasing the temperature increases the solubility of salts. This is due to an increase in the kinetic energy as temperature is increased.

Decreasing the pH, i.e. making the solution acidic, will increase the solubility of FeS. This is because $S^{2-}$ is weakly basic, and reacts with an acid, thereby decreasing its concentration. Le Chatelier’s predicts that the equilibrium shift to the right, increasing solubility.

Adding more Na$_2$S will decrease the solubility of FeS, due to the presence of a common ion ($S^{2-}$), which shifts the equilibrium to the reactant side.

10. Answer: E

Before beginning our calculations, it is helpful to make a list of the information that we have and the information that we need. From the problem, we have the following information:

The initial concentration of the reactant N$_2$ = 4.5 mol
The initial concentration of the reactant H$_2$ = 2.5 mol
Volume of the container = 1 L
Temperature of the reaction = 278 K
Equilibrium concentration of the product NH$_3$ = 1 mol
And we are asked to determine the value of $K_{eq}$.

Before we can solve for the equilibrium constant, $K_{eq}$ however, we need to know the chemical equation for this reaction. N$_2$ and H$_2$ are the reactants and they form NH$_3$, so the balanced reaction is:

N$_2$ + 3 H$_2$ $\leftrightarrow$ 2 NH$_3$

Therefore, the $K_{eq}$ is:

$$K_{eq} = \frac{[NH_3]_{eq}^2}{[N_2]_{eq} \times [H_2]_{eq}^3}$$

We know the equilibrium concentration of NH$_3$, so let's now find the equilibrium concentrations of the reactants and solve for $K_{eq}$. To do that, we use an ICE table:

<table>
<thead>
<tr>
<th></th>
<th>N$_2$</th>
<th>+ 3 H$_2$</th>
<th>$\leftrightarrow$ 2 NH$_3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>I:</td>
<td>4.5 M</td>
<td>2.5 M</td>
<td>0</td>
</tr>
<tr>
<td>C:</td>
<td>- x</td>
<td>- 3x</td>
<td>+ 2x</td>
</tr>
<tr>
<td>E:</td>
<td>4.5 - x</td>
<td>2.5 - 3x</td>
<td>2x</td>
</tr>
</tbody>
</table>
In the "I" row, we plug in the initial molarity of each reactant or product. Since this reaction is occurring in a 1 L container, the concentration given in moles is equal to the molarity for each reactant or product. In the "C" row, we plug in the change in concentration that will occur over the course of the reaction. Reactants decrease in concentration and products increase in concentration as the reaction proceeds and product is formed, so the change is negative for the reactants and positive for the products. Remember to include the coefficients from the balanced reaction in this value for each reactant or product. In the "E" row, we combine our initial concentration and change in concentration to create an expression for the concentration at equilibrium.

From the question, we know that the equilibrium concentration of NH₃ is 1 M, and from our ICE table, we know that the equilibrium concentration of NH₃ is also equal to 2x. We can use this information to determine what x equals:

\[ 2x = 1 \text{ mol} \]
\[ x = 0.5 \text{ mol} \]

We can use the value of x to solve for the equilibrium concentrations of the reactants:

\[ [\text{N}_2]_{eq} = 4.5 \text{ mol} - 0.5 \text{ mol} = 4 \text{ mol} \]
\[ [\text{H}_2]_{eq} = 2.5 \text{ mol} - 3(0.5 \text{ mol}) = 2.5 \text{ mol} - 1.5 \text{ mol} = 1 \text{ mol} \]

Now, we can plug all three equilibrium values into the equation for the equilibrium constant and calculate the value of \( K_{eq} \).

\[ K_{eq} = \frac{[\text{NH}_3]_{eq}^2}{[\text{N}_2]_{eq} \times [\text{H}_2]_{eq}^3} \]
\[ K_{eq} = \frac{(1)^2}{4 \times (1)^3} \]
\[ K_{eq} = 0.25 \]

Note that for this problem, we are NOT able to disregard the value of x, the change in concentration of the reactants, in the equation for \( K_{eq} \). We are able to disregard this x-value when \( K_{eq} \) is equal to or smaller than a number times \( 10^{-3} \). In this problem, however, we are not given \( K_{eq} \), so we cannot be sure that \( K_{eq} \) is small enough to make the x-value negligible. Keep this in mind when choosing whether to ignore "x" on I-C-E problems on the DAT!

11. Answer: C

In order to solve this problem, we need to set up an I-C-E table:

<table>
<thead>
<tr>
<th></th>
<th>( \text{H}_2 )</th>
<th>+ ( \text{I}_2 )</th>
<th>( \rightarrow \text{2 HI} )</th>
</tr>
</thead>
</table>

© DAT Bootcamp
<table>
<thead>
<tr>
<th>I:</th>
<th>0.5 M</th>
<th>1 M</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>C:</td>
<td>- x</td>
<td>- x</td>
<td>+ 2x</td>
</tr>
<tr>
<td>E:</td>
<td>0.5 - x</td>
<td>1 - x</td>
<td>2x</td>
</tr>
</tbody>
</table>

Note that the problem gave us the initial concentrations of the reactants in the form of molarity, so we were able to plug those given values straight into our “I” row.

Next, we plug the equilibrium concentrations from our I-C-E table into the equation for equilibrium constant, $K_{eq}$. We will use this equation to solve for $x$ and ultimately solve for the equilibrium concentration of HI.

$$K_{eq} = \frac{[HI]_{eq}^2}{[H_2]_{eq} \times [I_2]_{eq}}$$

$$K_{eq} = \frac{(2x)^2}{(0.5-x) \times (1-x)}$$

Since $K_{eq}$ is small (the $K_{eq}$ value is given in the problem and it is less than or equal to a number times $10^{-3}$), we can assume the change in concentration of the reactants is negligible, and approximate the equation as follows:

$$K_{eq} = \frac{(2x)^2}{(0.5)(1)}$$

$$K_{eq} = \frac{4x^2}{0.5}$$

Now, we plug in the given value for $K_{eq}$:

$$K_{eq} = 3.2 \times 10^{-3} = \frac{4x^2}{0.5}$$

To aid our calculation, we can convert $3.2 \times 10^{-3}$ to $32 \times 10^{-4}$. Remember, when converting the scientific notation of a number, if we make the number bigger, we must make the exponent smaller. So, to change $3.2$ to $32$ (making the number bigger), we must make the exponent smaller ($10^{-3}$ to $10^{-4}$).

$$K_{eq} = \frac{32 \times 10^{-4}}{0.5}$$

$$16 \times 10^{-4} = 4x^2$$

$$x^2 = 4 \times 10^{-4}$$

$$x = 2 \times 10^{-2}$$

Since the equilibrium concentration of HI is equal to $2x$, we must double this value, making $4 \times 10^{-2}$. Thus, C is the correct answer.

12. Answer: D
In this problem, we are given the following:

The initial concentration of the reactant SO\(_2\) = 15 mol
The initial concentration of the reactant O\(_2\) = 10 mol
The volume of the container = 5 L
K\(_{eq}\) = 2 \times 10^{-6}
And we are asked to find the equilibrium concentration of the product, SO\(_3\).

To find the equilibrium concentration of SO\(_3\) we will need to set up an I-C-E table. However, before we can do this, we need to convert the given initial concentrations of the reactants from moles to molarity, using the given volume of the container, 5 L.

Molarity is calculated by dividing moles of substance by liters of the container:
Molarity = moles / liter
[SO\(_2\)] = 15 mol / 5 L = 3 M
[O\(_2\)] = 10 mol / 5 L = 2 M

<table>
<thead>
<tr>
<th></th>
<th>2 SO(_2)</th>
<th>+ O(_2)</th>
<th>——&gt; 2 SO(_3)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>3 M</td>
<td>2 M</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>- 2x</td>
<td>- x</td>
<td>+ 2x</td>
</tr>
<tr>
<td>E</td>
<td>3 - 2x</td>
<td>2 - x</td>
<td>2x</td>
</tr>
</tbody>
</table>

Now we plug the expressions for equilibrium concentration for each reactant or product into our equation for the equilibrium constant, K\(_{eq}\).

K\(_{eq}\) = [SO\(_3\)]\(_{eq}\)^{2} / [SO\(_2\)]\(_{eq}\)^{2} * [O\(_2\)]\(_{eq}\)
K\(_{eq}\) = (2x)^{2} / (3-2x)^{2} * (2-x)

Since K\(_{eq}\) is small (the value of K\(_{eq}\) is given in the problem and it is less than or equal to a number times 10\(^{-4}\)), we can assume the change in concentration of the reactants is negligible, and approximate the equation as follows:

K\(_{eq}\) = 4x\(^2\) / (3)^{2} * 2 = 4x\(^2\) / 18

We plug in the given value for K\(_{eq}\) and solve for x:

K\(_{eq}\) = 2 \times 10^{-6} = 4x\(^2\) / 18
36 \times 10^{-6} = 4x\(^2\)
9 \times 10^{-6} = x\(^2\)
x = 3 \times 10^{-3}
Finally, we use the value of $x$ to calculate the equilibrium concentration of $\text{SO}_3$.

\[
[\text{SO}_3] = 2x \\
[\text{SO}_3] = 2(3 \times 10^{-3}) = 6 \times 10^{-3} \text{ M}
\]
Chapter 10 Quiz: Acid-Base Equilibria

1. Of the following acids, ________ is the strongest.

   A. HCl  
   B. HBr  
   C. HI  
   D. HF  
   E. H₂O

2. Which of the following statements regarding pH is true?
   
   I. A 0.1 M solution of a strong acid has a higher pH than a 0.1 M solution of a weak acid.  
   II. The pH of a neutral solution is 7 at all temperatures.  
   III. The pH of a 0.10 M strong base is higher than a 0.10 M weak base.

   A. II only  
   B. II and III only  
   C. I and II only  
   D. III only  
   E. I, II, and III

3. Which of the following bases has the strongest conjugate acid?

   A. NH₃, \( K_b = 1.8 \times 10^{-5} \)  
   B. C₆H₅NH₂, \( K_b = 4.3 \times 10^{-10} \)  
   C. (CH₃)₂NH, \( K_b = 5.4 \times 10^{-4} \)  
   D. HONH₃, \( K_b = 1.1 \times 10^{-8} \)  
   E. CH₃NH₂, \( K_b = 4.4 \times 10^{-4} \)
4. The diagram below depicts a solution of three acids at the same temperature and concentration. Which acid is the strongest?

![Diagram of acid solutions]

A. I  
B. II  
C. III  
D. IV  
E. V

5. Which of the following 0.1 M solutions has the highest pH?

A. HCl  
B. HF ($K_a = 6.8 \times 10^{-4}$)  
C. HIO ($K_a = 2.3 \times 10^{-11}$)  
D. H$_2$CO$_3$ ($K_a = 4.5 \times 10^{-7}$)  
E. HC$_2$H$_3$O$_2$ ($K_a = 1.8 \times 10^{-5}$)

6. Which of the following expressions correctly gives the pH of a 1.0 M solution of NH$_3$ ($K_b = 1.8 \times 10^{-5}$)

A. pH = 14 – $pK_a$  
B. pH = $pK_b$  
C. pH = 14 – $\frac{1}{2}$ $pK_b$  
D. pH = $\frac{1}{2}$ $pK_b$  
E. pH = 7
7. Which of the following solutions has a pH > 7 at 25°C?
   I. 0.10 M NaCN
   II. 0.05 M \((\text{CH}_3)_2\text{NH}\)
   III. 0.25 M \(\text{NH}_4\text{Cl}\)

   A. II only
   B. I and II only
   C. II and III only
   D. I and III only
   E. I, II, and III

8. Which of the following conjugate acid-base pairs is best to use to prepare a buffer of pH 3.50?
   A. \(\text{HCl}–\text{NaCl}\)
   B. \(\text{HCHO}_2–\text{NaCHO}_2\) \((K_a \text{ of HCOOH } = 1.8 \times 10^{-4})\)
   C. \(\text{HC}_2\text{H}_3\text{O}_2–\text{NaC}_2\text{H}_3\text{O}_2\) \((K_a \text{ of CH}_3\text{COOH } = 1.8 \times 10^{-5})\)
   D. \(\text{HC}_2\text{H}_3\text{O}_2–\text{KClO}_2\) \((K_a \text{ of HC}_3\text{H}_5\text{O}_2 = 1.3 \times 10^{-5})\)
   E. \(\text{HClO}_2–\text{NaC}_2\text{H}_3\text{O}_2\) \((K_a \text{ of HClO}_2 \text{ } 1.1 \times = 10^{-2})\)

9. Arrange the following substances in order of weakest to strongest base.
   I. \(\text{NO}_2^-\) \((K_a \text{ for } \text{HNO}_2 \text{ is } 7.1 \times 10^{-4})\)
   II. \(\text{C}_6\text{H}_{10}\text{NH}\) \((K_b \text{ of } \text{HC}_6\text{H}_5\text{NH}_2 = 1.3 \times 10^{-3})\)
   III. \(\text{C}_6\text{H}_5\text{COO}^-\) \((K_a \text{ for } \text{C}_6\text{H}_5\text{COOH} \text{ is } 6.3 \times 10^{-5})\)

   A. II < I < III
   B. III < I < II
   C. I < III < II
   D. II < III < I
   E. I < II < III

10. Arrange the following salt solutions from lowest to highest pH.
    I. \(\text{NaBr}\)
    II. \(\text{KF}\)
    III. \(\text{NH}_4\text{I}\)

    A. II < I < III
    B. III < II < I
    C. III < I < II
    D. I < II < III
    E. II = I = III (all have pH = 7)
1. Answer: C

The strength of binary acids (acids that are composed of H and a nonmetal) is determined by the electronegativity and size of the nonmetal. For acids whose nonmetals are in the same column in the periodic table (and thus have the same electronegativity), we look at the size of the nonmetal to determine the strength of the acid. Larger atoms can better handle and disperse the negative charge that results from dissociation. So, larger nonmetals in binary acids make stronger acids. Size increases down a column in the periodic table, so HI is the strongest among the acids given.

2. Answer: D

The pH of a solution is defined as the negative log of the $\text{H}^+$ concentration. The higher the $[\text{H}^+]$, the lower the pH is, and the more acidic. The relationship between $[\text{H}^+]$ and $[\text{OH}^-]$ is given by

$$K_w = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14} \text{ @ } 25 \degree \text{C}$$

A neutral solution is one where $[\text{H}^+] = [\text{OH}^-]$. At 25 °C, the $[\text{H}^+] = 1.0 \times 10^{-7}$ M for a neutral solution. Like any other equilibrium constant, the value of $K_w$ varies with temperature, so the $[\text{H}^+]$ for a neutral solution and its pH also varies with temperature. Thus, statement II is incorrect.

Strong acids completely dissociate into ions in solution, while weak acids partially dissociate. So, a strong acid has a higher $[\text{H}^+]$ concentration (lower pH) than a weak acid. Thus, statement I is incorrect.

Similarly, strong bases completely ionize, while weak bases partially ionize in solution. A base increases the $[\text{OH}^-]$ concentration, and strong bases have higher $[\text{OH}^-]$ than weak bases. From the equation above, we see that the higher the $[\text{OH}^-]$, the lower the $[\text{H}^+]$. The lower the $[\text{H}^+]$, the higher the pH is. So, a strong base has a higher pH than a weak base.
3. Answer: B

The base dissociation constant, $K_b$, is a measure of the degree of ionization (strength) of a base. The stronger the base, the higher its $K_b$. Similarly, the acid dissociation constant $K_a$ is a measure of the strength of an acid. The stronger the acid, the higher the $K_a$ is. In a Bronsted-Lowry acid-base conjugate pair, the relative strength of the acid and the base is given as

$$K_w = K_aK_b$$

From this equation, we see that the lower the $K_b$, the higher the $K_a$ (the weaker the base, the stronger is its conjugate acid).

4. Answer: B

The degree of dissociation of acids determines the strength of the acid. The greater the acid is dissociated into ions, the stronger the acid is. We are given five diagrams of acid solution, and we see that II has all but one acid molecule dissociated into ions.

5. Answer: C

The solution with the highest pH is the one with the lowest $[H^+]$. We are given five acids of varying strength. One is a strong acid, HCl, and the rest are weak acids. The amount of $H^+$ in solution will depend on the strength of the acid. The stronger the acid, the more dissociated it is, so the higher the $[H^+]$ will be in solution. Conversely, the weaker the acid, the lower the $[H^+]$ in solution. We can eliminate A from our choices since HCl is a strong acid and will have the highest $[H^+]$ (lowest pH) among the five choices given.

For weak acids, the value of the $K_a$ is a measure of the strength of the acid. The higher the $K_a$, the stronger the acid is. So, HIO (which has the lowest $K_a$) will have the lowest $[H^+]$ and highest pH.
6. Answer: C

NH₃ is a weak base, and to determine the pH of a weak base, we set-up an ICE table and use the equilibrium constant expression for NH₃ in H₂O:

<table>
<thead>
<tr>
<th></th>
<th>NH₃</th>
<th>⇌</th>
<th>NH₄⁺</th>
<th>+</th>
<th>OH⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.0</td>
<td></td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>− x</td>
<td>+ x</td>
<td>+ x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1 − x</td>
<td>x</td>
<td>x</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Using the equilibrium constant expression,

\[ K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \]

\[ K_b = \frac{x^2}{1-x} \approx \frac{x^2}{1} \]

\[ x = [OH^-] = \sqrt{K_b} \]

Note we can assume 1-x ~ 1 because this is a weak base (there is not much dissociation; we were told this in the question stem with \( K_b = 1.8 \times 10^{-5} \)).

Taking the \( -\log [OH^-] \) to get pOH:

\[ pOH = \log \sqrt{K_b} = \log (K_b)^{1/2} = \frac{1}{2} \log K_b = \frac{1}{2} pK_b \]

and the pH is

\[ pH = 14 \quad pOH = 14 \quad \left( \frac{1}{2} pK_b \right) \]

7. Answer: B

A solution with a pH > 7 is a basic solution. We need to analyze each of the given substances to see if it’s acidic or basic.

NaCN is basic, because CN⁻ is the conjugate base of the acid HCN.

(CH₃)₂NH is a base. A nitrogen bonded to three other groups (in this case 2 CH₃ and an H) is a weak base.
NH₄Cl is acidic, since NH₄⁺ is the conjugate acid of the base NH₃.

8. Answer: B

A buffer is a solution of a weak acid and its conjugate base or a weak base and its conjugate acid. The best conjugate acid-base pair to use to prepare a buffer is the one whose acid has a pKₐ at or near the desired pH. Since we want a buffer of pH 3.5, we would pick a conjugate acid-base pair with a pKₐ close to this pH. By approximation, we see that HCHO₂ has a pKₐ (~ log (1.8 × 10⁻⁴)) of somewhere between 3 and 4.

9. Answer: C

The base dissociation constant, K₏, is a measure of the strength of a base: the higher the Kₖ, the stronger the base. For I and III, we are given the Kₐ of the conjugate acid. To get the Kₖ, we use

\[ K_w = K_a \times K_b \]

where \( K_w = 1.0 \times 10^{-14} \) is the autoionization constant of water. Thus, for I, \( K_b \sim 10^{-10} \), and for II, \( K_b \sim 10^{-9} \). Therefore, II is the strongest base and I is the weakest.

10. Answer: C

We determine whether a solution of salt is acidic, basic, or neutral by assessing each ion.

For NaBr, the ions are Na⁺ and Br⁻. Na⁺ is the cation of an alkali metal and therefore will not affect the pH. Br⁻ will also not affect the pH since it is the conjugate base of HBr, a strong acid. Therefore, the pH of NaBr solution is the same as that of water (pH 7).

The ions in KF are K⁺ and F⁻. K⁺, like Na⁺, will not affect the pH of the solution. F⁻ is the conjugate base of the weak acid HF, so it is basic. Thus, a solution of KF will have a pH > 7.

In NH₄I, the ions are NH₄⁺ and I⁻. Like Br⁻, I⁻ will not affect the pH of the solution since it is the conjugate base of the strong acid HI. But NH₄⁺ will make the pH < 7 because it is weakly acidic (conjugate acid of the base NH₃).
Chapter 11 Quiz: Thermodynamics and Thermochemistry

1. Which of the following reactions involve an increase in entropy?
   I. \( S(s) + O_2(g) \rightarrow SO_2(g) \)
   II. \( 2 \text{ KClO}_3(s) \rightarrow 2 \text{ KCl}(s) + 3 \text{ O}_2(g) \)
   III. \( \text{Mg}(s) + \text{O}_2(g) \rightarrow \text{MgO}(s) \)

   A. I only
   B. I and III only
   C. II only
   D. II and III only
   E. I, II, and III

2. What can be said about an endothermic reaction that is accompanied by an increase in entropy?

   A. The Gibbs free energy change of the reaction has a negative value at all temperatures.
   B. The reaction will be spontaneous at all temperatures.
   C. The reaction will be spontaneous above a certain temperature.
   D. The reaction will be spontaneous below a certain temperature.
   E. The reaction is spontaneous only at 25°C.

3. Arrange the following in order of increasing entropy.

   I. \( \text{N}_2\text{O}_4(g) \)
   II. \( \text{CO}(g) \)
   III. \( \text{SO}_2(g) \)

   A. I < II < III
   B. II < I < III
   C. III < II < I
   D. III < I < II
   E. II < III < I
4. Of the following processes, __________ is exothermic.
   I. boiling of water
   II. burning of fuel
   III. formation of water droplets on a cold early morning

   A. I and III only
   B. II and III only
   C. I only
   D. II only
   E. I, II, and III

5. Which of the following diatomic molecules has a nonzero standard enthalpy of formation?

   A. O₂(g)
   B. H₂(g)
   C. N₂(g)
   D. Br₂(g)
   E. F₂(g)

6. 5.0 g of food was combusted in a bomb calorimeter. The temperature increased from 25 °C to 50 °C. How much energy can be obtained per gram of the food? The heat capacity of the bomb calorimeter is 10.0 kJ/°C.

   A. 100 kJ
   B. 50 kJ
   C. 500 kJ
   D. 25 kJ
   E. 20 kJ

7. How many of the following transformations has \( \Delta S < 0 \)?

   I. \( \text{CO}_2(g, 25 \, ^\circ\text{C}, 1 \, \text{atm}) \rightarrow \text{CO}_2 (g, 25 \, ^\circ\text{C}, 2 \, \text{atm}) \)
   II. \( \text{CO}_2(s, -78.5 \, ^\circ\text{C}, 1 \, \text{atm}) \rightarrow \text{CO}_2(g, -78.5 \, ^\circ\text{C}, 1 \, \text{atm}) \)
   III. \( \text{CO}_2(s, -100 \, ^\circ\text{C}, 1 \, \text{atm}) \rightarrow \text{CO}_2(s, -200 \, ^\circ\text{C}, 1 \, \text{atm}) \)
   IV. \( \text{CO}_2(l, -23 \, ^\circ\text{C}, 6 \, \text{atm}) \rightarrow \text{CO}_2(g, -23 \, ^\circ\text{C}, 6 \, \text{atm}) \)
   V. \( \text{CO}_2(g, 30 \, ^\circ\text{C}, 1 \, \text{atm}) \rightarrow \text{CO}_2(g, 40 \, ^\circ\text{C}, 1 \, \text{atm}) \)

   A. one
   B. two
   C. three
   D. four
   E. five
8. Arrange the following in order of increasing final temperature when the same amount of heat is added to 1 g of each substance at the same initial temperature, assuming no heat is lost during the process.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat Capacity, J/g °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>I. Mg</td>
<td>1.020</td>
</tr>
<tr>
<td>II. Al</td>
<td>0.900</td>
</tr>
<tr>
<td>III. Fe</td>
<td>0.444</td>
</tr>
<tr>
<td>IV. Ag</td>
<td>0.240</td>
</tr>
</tbody>
</table>

A. I < II < III < IV  
B. II < I < IV < III  
C. IV < III < II < I  
D. I < III < II < IV  
E. III < IV < I < II

9. Which of the following reactions is exothermic?

<table>
<thead>
<tr>
<th>Bond Energy, kJ</th>
<th>A₂</th>
<th>B₂</th>
<th>C₂</th>
<th>AB</th>
<th>BC</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bond</td>
<td>100</td>
<td>50</td>
<td>250</td>
<td>300</td>
<td>40</td>
</tr>
</tbody>
</table>

I. A₂ + B₂ → 2 AB  
II. 2 BC → B₂ + C₂  
III. A₂ + 2 BC → 2 AB + C₂

A. I and II only  
B. I and III only  
C. II and III only  
D. I, II, and III
10. What is the sign of enthalpy change and entropy change for the process shown depicted below, in which a gas becomes a solid?

A. $\Delta H > 0$, $\Delta S > 0$
B. $\Delta H > 0$, $\Delta S < 0$
C. $\Delta H < 0$, $\Delta S > 0$
D. $\Delta H < 0$, $\Delta S < 0$
E. $\Delta H = 0$, $\Delta S = 0$

11. Given the heating curve for 1 mole of a substance below, estimate the amount of heat needed to completely power the phase change of vaporization of the liquid substance.

A. 50 J
B. 100 J
C. 150 J
D. 200 J
E. 250 J

12. Consider the heating curve shown below. Which of the following statement(s) is/are true?
I. The heat required to vaporize the substance is greater than that required to melt the substance.

II. The heat capacity of the solid is less than the heat capacity of the liquid.

III. There are two phases that exist on segment DE.

A. I only
B. III only
C. I and III only
D. I and II only
E. I, II, and III
13. The slope of which segment in the heating curve shown corresponds to the heat capacity of the liquid.

A. **AB**  
B. **BC**  
C. **CD**  
D. **DE**  
E. **EF**

14. Arrange the following in order of increasing $\Delta G^\circ$ (most negative to least negative) at 298 K.

**I.** $N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$ \hspace{1cm} $K = 1.45 \times 10^{-5}$

**II.** $H_2(g) + I_2(g) \rightarrow 2 HI(g)$ \hspace{1cm} $K = 50.5$

**III.** $PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$ \hspace{1cm} $K = 0.50$

A. **I < III < II**  
B. **II < III < I**  
C. **III < I < II**  
D. **I < II < III**  
E. **III < II < I**
**ANSWER KEY**

1. **Answer: C**

   Entropy is related to dispersing kinetic energy.

   Gases have more freedom of motion and have higher entropy than liquids or solids.

   A reaction that produces more moles of gas has a positive (increasing) entropy change.

   For **I**, we have 1 mole of gas on both the reactant and product side, so entropy has not changed. For **II**, there are 3 moles of gas formed on the product side, and only a solid on the reactant side, so entropy has increased. For **III**, we go from 1 mole of gas on the reactant side to only solid on the product side, so entropy has decreased.

2. **Answer: C**

   A reaction is spontaneous if it has a negative Gibbs free energy change, $\Delta G^\circ$. The free energy change ($\Delta G$) is related to the enthalpy ($\Delta H$) and entropy ($\Delta S$) by the following equation:

   $$\Delta G = \Delta H - T\Delta S$$

   For an endothermic reaction ($\Delta H > 0$) with increasing entropy ($\Delta S > 0$), we see from the above equation that $\Delta G$ will be negative (a spontaneous reaction) only if $T\Delta S > \Delta H$, i.e. at high $T$.

3. **Answer: E**

   Entropy increases with increasing molecular complexity. The more atoms in a molecule, the more types of motion available, and the higher the entropy. Thus, $N_2O_4(g)$ has the highest entropy and CO(g) has the lowest.

4. **Answer: B**

   Exothermic processes are those that release heat to the surroundings, while endothermic processes are those that absorb heat.

   We must supply heat to water in order to bring it to boil, thus boiling is an endothermic process. It requires energy for the liquid water molecules to vaporize into a gas.

   The burning of fuel is an exothermic process; during combustion, heat is released.

   Formation of water droplets (condensation) is the opposite process of boiling (vaporization), thus it is an exothermic process. We take gaseous water (present in the atmosphere) and
heat is released into the surroundings as the water cools down and changes into its liquid form.

5. Answer: D

The standard enthalpy of formation, $\Delta H^\circ_f$, is zero for elements in their standard states (i.e. the most stable form of the element at 1 atm, and 298 K).

All of the given molecules are elements in their standard states except $\text{Br}_2(g)$; $\text{Br}_2$ exists as a liquid.

H, F, O, N, Cl and the noble gases all exist in the gaseous form in their standard state. Mercury ($\text{Hg}$) and bromine ($\text{Br}_2$) are in the liquid form in their standard state. All other elements exist as solids in their standard state.

6. Answer: B

The heat of a reaction in a bomb calorimeter is given by

$$q_{\text{rxn}} = C \times \Delta T$$

where $C$ is the heat capacity of the bomb calorimeter and $\Delta T$ is the change in temperature.

So, for the reaction given:

$$q_{\text{rxn}} = -10.0 \text{ kJ/}^\circ\text{C} \times (50 - 25 ^\circ\text{C}) = -250 \text{ kJ.}$$

This means that 250 kJ of heat is released (the negative sign indicates that heat is released) for 5 grams of food, or 50 kJ of heat per gram.
7. Answer: B

Entropy decreases ($\Delta S < 0$) with a decrease in temperature, decrease in volume (which is an increase in pressure) or decrease in the number of gas particles.

**I.** $\text{CO}_2(\text{g}, \ 25^\circ \text{C}, \ 1 \ \text{atm}) \rightarrow \text{CO}_2(\text{g}, \ 25^\circ \text{C}, \ 2 \ \text{atm})$ – pressure has increased, volume has decreased: $\Delta S < 0$

**II.** $\text{CO}_2(\text{s}, \ -78.5^\circ \text{C}, \ 1 \ \text{atm}) \rightarrow \text{CO}_2(\text{g}, \ -78.5^\circ \text{C}, \ 1 \ \text{atm})$ – phase change from solid to gas: $\Delta S > 0$

**III.** $\text{CO}_2(\text{s}, \ -100^\circ \text{C}, \ 1 \ \text{atm}) \rightarrow \text{CO}_2(\text{s}, \ -200^\circ \text{C}, \ 1 \ \text{atm})$ – temperature has decreased: $\Delta S < 0$

**IV.** $\text{CO}_2(\text{l}, \ -23^\circ \text{C}, \ 6 \ \text{atm}) \rightarrow \text{CO}_2(\text{g}, \ -23^\circ \text{C}, \ 6 \ \text{atm})$ – phase change from liquid to gas: $\Delta S > 0$

**V.** $\text{CO}_2(\text{g}, \ 30^\circ \text{C}, \ 1 \ \text{atm}) \rightarrow \text{CO}_2(\text{g}, \ 40^\circ \text{C}, \ 1 \ \text{atm})$ – temperature has increased: $\Delta S > 0$

8. Answer: A

The amount of heat absorbed by a substance, $q$, is related to the increase in temperature, $\Delta T$, of the substance by the equation

$$q = mC \ \Delta T$$

where $C$ is the specific heat capacity of the substance and $m$ is the mass of the substance.

We see from this equation that for a given amount of heat ($q$), temperature change ($\Delta T$) increases as heat capacity ($C$) decreases.

Thus the final temperature increases with decreasing $C$. Mg has the largest heat capacity, and therefore will have the lowest final temperature. Ag has the smallest heat capacity, and therefore will have the highest final temperature.
9. Answer: D

The enthalpy of a reaction is related to its bond energy (BE) by the following equation:

\[ \Delta H_{\text{rxn}} = \text{sum of BE of the bonds broken} - \text{sum of BE of the bonds formed} \]

<table>
<thead>
<tr>
<th>Bond</th>
<th>A₂</th>
<th>B₂</th>
<th>C₂</th>
<th>AB</th>
<th>BC</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bond Energy, kJ</td>
<td>100</td>
<td>50</td>
<td>250</td>
<td>300</td>
<td>40</td>
</tr>
</tbody>
</table>

**I.** \( A₂ + B₂ \rightarrow 2 \ AB \)

\[ \Delta H_{\text{rxn}} = [100 + 50] - [2(300)] \]
\[ \Delta H_{\text{rxn}} = 150 - 600 \]
\[ \Delta H_{\text{rxn}} = -450 \text{ kJ} \]

**II.** \( 2 \ BC \rightarrow B₂ + C₂ \)

\[ \Delta H_{\text{rxn}} = [2(40)] - [50 + 250] \]
\[ \Delta H_{\text{rxn}} = 80 - 300 \]
\[ \Delta H_{\text{rxn}} = -220 \text{ kJ} \]

**III.** \( A₂ + 2 \ BC \rightarrow 2 \ AB + C₂ \)

\[ \Delta H_{\text{rxn}} = [100 + 2(40)] - [2(300) + 250] \]
\[ \Delta H_{\text{rxn}} = 180 - 850 \]
\[ \Delta H_{\text{rxn}} = -670 \text{ kJ} \]

Exothermic reactions (\( \Delta H = - \)) are those reactions where the sum of the BE of the reactants are weaker (less positive) than the bond energies of the products. All three reactions have stronger bonds on the product side, and negative \( \Delta H \). They are all exothermic reactions.

10. Answer: D

The diagram depicts the conversion of a gas to a solid (deposition). During this process, the entropy change and enthalpy change are negative (entropy decreases and the process is exothermic).
11. Answer: B

Segment D–E represents the vaporization of the substance. To completely vaporize the liquid substance initially at its boiling point, we are bringing the substance from point D to point E. To bring about this change, heat corresponding to the difference in the heat added from point D to point E, i.e. 250 - 150 = 100 J, must be added.

12. Answer: E

The length of segments BC and DE represents the heat required to completely melt and vaporize the substance, respectively. Since segment DE is longer than BC, it takes more heat to completely vaporize the substance that it does to melt it.

The slope of segments AB and CD can be used to determine the heat capacities of the solid and the liquid, respectively. Segment AB has a greater slope (steeper line), than that of segment CD, and therefore the solid’s temperature increases more rapidly with less input of energy. Therefore, the solid has a lower heat capacity than the liquid.

Horizontal lines on a heating curve indicate a phase change, as T remains constant as heat is added. Thus, there are two phases that exist on segment DE.

13. Answer: C

Heat capacity is the amount of heat needed to raise the temperature of a substance by 1 degree. This means that the slope of segment AB is representative of the heat capacity of the solid, while the slope of segment CD represents the heat capacity of the liquid. A greater slope indicates a lower heat capacity.

Segments BC and DE correspond to phase change (melting and vaporization, respectively).

14. Answer: B

The relationship between $\Delta G^\circ$ and $K$ is given as

$$G^\circ = RT \ln(K)$$

From the above equation, we see that the larger the value of $K$, the smaller (more negative) the value of $\Delta G^\circ$ is. Also note that for $K > 1$, $\ln(K) > 0$, so $\Delta G^\circ$ is negative, which implies a spontaneous (product-favored) reaction.
Chapter 12 Quiz: Oxidation and Reduction

1. For the redox conversion $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$, which of the following statements is TRUE?
   A. There are 5 electrons involved per mole of $\text{MnO}_4^-$.
   B. $\text{MnO}_4^-$ is reduced to form $\text{Mn}^{2+}$.
   C. In acidic medium, there are 8 mol of $\text{H}^+$ involved per mol of $\text{MnO}_4^-$.
   D. two of the above
   E. all of the above

2. Which of the following reduction reactions involve the greatest number of electrons per mol of oxidant?
   A. $\text{CO}_2 \rightarrow \text{C}_2\text{O}_4^{2-}$
   B. $\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}$
   C. $\text{NO}_3^- \rightarrow \text{NO}$
   D. $\text{N}_2 \rightarrow \text{N}_2\text{H}_5^+$
   E. $\text{Al}^{3+} \rightarrow \text{Al}$

3. When the following half-reaction is balanced in acidic medium, the number of mol of $\text{H}^+$ per mol of $\text{HClO}$ is ________.
   $$\text{HClO} \rightarrow \text{Cl}_2$$
   A. 1
   B. 2
   C. 3
   D. 4
   E. 5

4. In which of the following compounds does nitrogen have the most positive oxidation state?
   A. $\text{NH}_3$
   B. $\text{N}_2\text{O}_4$
   C. $\text{NO}_3^-$
   D. $\text{N}_2\text{H}_5^+$
   E. $\text{N}_2$
5. Given the following standard reduction potentials, which of the following statements is correct?

\[
\begin{align*}
&\text{Ni}^{2+} + 2 \text{e}^- \rightarrow \text{Ni} & -0.25 \text{ V} \\
&\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu} & +0.521 \text{ V} \\
&\text{Co}^{3+} + 3\text{e}^- \rightarrow \text{Co} & +1.83 \text{ V}
\end{align*}
\]

A. When Co^{3+}/Co and Ni^{2+}/ Ni are used in a voltaic cell, Ni^{2+}/ Ni is the cathode.
B. Co^{3+}/Co can reduce Ni^{2+}/ Ni
C. Cu^+/Cu is a stronger oxidizing agent than Co^{3+}/Co.
D. Ni^{2+}/ Ni is a stronger reducing agent than Cu^+/Cu.
E. Ni^{2+}/ Ni can oxidize Cu^+/Cu.

6. Given the following standard electron potentials, which metal cation can oxidize Ni but not Ag?

\[
\begin{align*}
&\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au} & +1.50 \text{ V} \\
&\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag} & +0.80 \text{ V} \\
&\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} & +0.34 \text{ V} \\
&\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni} & -0.25 \text{ V} \\
&\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe} & -0.44 \text{ V}
\end{align*}
\]

A. Au^{3+}
B. Fe^{2+}
C. Cu^{2+}
D. Fe^{2+} and Cu^{2+}
E. Au^{3+} and Cu^{2+}

7. The following are the redox reactions in a voltaic cell containing a 1.0 M solution of cations.

\[
\begin{align*}
&\text{cathode: } \text{Ag}^+ + \text{e}^- \rightarrow \text{Ag} \\
&\text{anode: } \text{Cu} \rightarrow \text{Cu}^{2+} + 2 \text{ e}^-
\end{align*}
\]

Which of the following choices would result in an increase in the cell potential?

A. Using a bigger Ag electrode
B. Increasing the volume of a 1.0 M Cu^{2+} solution
C. Increasing the concentration of Ag^+
D. Decreasing the size of the Cu electrode
E. Increasing the concentration of Cu^{2+}

8. Which of the following statements is true regarding the redox reaction
2 Al^{3+} + 3 Zn → 3 Zn^{2+} + 2 Al \quad E_{\text{cell}} = -0.9 \text{ V}

A. Al^{3+} is the reducing agent in the forward reaction.
B. Zn^{2+} is the reductant in the reverse reaction.
C. Zn reduces Al^{3+} spontaneously.
D. Al^{3+}/Al has a more positive reduction potential than Zn^{2+}/Zn.
E. Al is a stronger reducing agent than Zn.

9. A voltaic cell was constructed using the following half-reactions (written as reduction half-reactions). What is the overall reaction and the standard cell potential of this voltaic cell?

\begin{align*}
\text{Cd}^{2+} + 2e^- & \rightarrow \text{Cd} \quad -0.40 \text{ V} \\
\text{Pb}^{2+} + 2e^- & \rightarrow \text{Pb} \quad -0.13 \text{ V}
\end{align*}

A. \text{Cd}^{2+} + \text{Pb} → \text{Cd} + \text{Pb}^{2+}, \ E_{\text{cell}} = +0.27 \text{ V}
B. \text{Cd}^{2+} + \text{Pb} → \text{Cd} + \text{Pb}^{2+}, \ E_{\text{cell}} = -0.27 \text{ V}
C. \text{Pb}^{2+} + \text{Cd} → \text{Pb} + \text{Cd}^{2+}, \ E_{\text{cell}} = +0.27 \text{ V}
D. \text{Pb}^{2+} + \text{Cd} → \text{Pb} + \text{Cd}^{2+}, \ E_{\text{cell}} = -0.27 \text{ V}
E. \text{Pb}^{2+} + \text{Cd} → \text{Pb} + \text{Cd}^{2+}, \ E_{\text{cell}} = -0.54 \text{ V}

10. A voltaic cell based on the reaction 3Cu^{2+} + 2Cr → 2Cr^{3+} + 3Cu generates \( E^o_{\text{cell}} = +1.08 \text{ V} \).

Given the \( E^o \) for \( \text{Cu}^{2+} + 2e^- → \text{Cu} \) is +0.34 V, what is the reduction potential for \( \text{Cr}^{3+} + 3 \text{e}^- → \text{Cr} \)?

A. +0.74 V
B. −0.74 V
C. +1.42 V
D. −1.42 V
E. +1.08 V
ANSWER KEY

1. Answer: E

The following are the rules for balancing half-reactions:

a) Balance the atoms other than hydrogen and oxygen
   \[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \]

b) Balance the oxygen by adding water.
   \[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O} \]

c) Balance the hydrogen by adding H⁺.
   \[ \text{MnO}_4^- + 8 \text{H}^+ \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O} \]

d) Balance the charges by adding electrons.
   \[ \text{MnO}_4^- + 8 \text{H}^+ + 5 \text{e}^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O} \]

From the balanced equation above, we see that MnO₄⁻ is reduced by 5 electrons/mol MnO₄⁻ to Mn²⁺ (the charge on the Mn atom goes from +7 in the MnO₄⁻ state to +2 in the Mn²⁺ state), and that there are 8 mol of H⁺ involved per mol of MnO₄⁻.

2. Answer: B

The following are the rules for assigning oxidation number:

a) The oxidation number of an atom in its elemental form is zero (e.g. N₂, Ag).

b) The oxidation number of a monatomic ion is equal to its charge (e.g. Ca²⁺ has an oxidation number of +2).

c) The sum of oxidation numbers in a molecule or formula is zero. The sum of oxidation numbers in a polyatomic ion is equal to the charge of the ion.

d) The oxidation number of Group IIA is +2 in all of its compounds. Group IIA has an oxidation number of +2 in all of its compounds.

e) Hydrogen has an oxidation number of +1 if combined with nonmetals and −1 if combined with a metal or boron.

f) Fluorine has an oxidation number of −1 in all of its compounds.

g) Oxygen has an oxidation number of −2 in all of its compounds except in peroxide (−1) and when combined with F (e.g. in OF₂, because F has an oxidation number of −1 in all of its compounds, it follows that O has an oxidation number of +2).

h) Halogens have an oxidation number of −1 if combined with metals, nonmetals (except O) and other halogens lower in the group.

We assign the oxidation number and determine the number of electrons involved in each reduction reaction.
A:  \[ \text{CO}_2 \rightarrow \text{C}_2\text{O}_4^{2-} \]

\[ \text{CO}_2: \quad \text{O} = -2, \text{C} = +4 \]
\[ \text{C}_2\text{O}_4^{2-}: \quad \text{O} = -2, \text{C} = +3 \]

Carbon undergoes a change in oxidation from +4 to +3, corresponding to a one-electron reduction per mol of \( \text{CO}_2 \).

B:  \[ \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+} \]

\[ \text{Cr}_2\text{O}_7^{2-}: \quad \text{O} = -2, \text{Cr} = +6 \]
\[ \text{Cr}^{3+}: \quad \text{Cr} = +3 \]

Chromium changes oxidation number from +6 to +3, corresponding to a 3-electron reduction per mol of \( \text{Cr} \), or 6 e\(^{-}\) reduction per mol of \( \text{Cr}_2\text{O}_7^{2-} \).

C:  \[ \text{NO}_3^- \rightarrow \text{NO} \]

\[ \text{NO}_3^-: \quad \text{O} = -2, \text{N} = +5 \]
\[ \text{NO}: \quad \text{O} = -2, \text{N} = +2 \]

Nitrogen undergoes a +5 to +2 change in oxidation number, which means that \( \text{NO}_3^- \) is reduced by 3 electrons to \( \text{NO} \).

D:  \[ \text{N}_2 \rightarrow \text{N}_2\text{H}_5^+ \]

\[ \text{N}_2: \quad \text{N} = 0 \]
\[ \text{N}_2\text{H}_5^+: \quad \text{H} = +1, \text{N} = -2 \]

Nitrogen changes its oxidation number by 2, so \( \text{N}_2 \) is reduced to \( \text{N}_2\text{H}_5^+ \) by 4 electrons.

E:  \[ \text{Al}^{3+} \rightarrow \text{Al} \]

\[ \text{Al}^{3+}: \quad \text{Al} = +3 \]
\[ \text{Al}: \quad \text{Al} = 0 \]

\( \text{Al}^{3+} \) is reduced to \( \text{Al} \) by 3 electrons.
3. Answer: A

To balance the redox reaction, we use the following steps:

a) Balance the atoms other than hydrogen and oxygen
   
   \[2 \text{HClO} \rightarrow \text{Cl}_2\]

b) Balance the oxygen by adding water.
   
   \[2 \text{HClO} \rightarrow \text{Cl}_2 + 2 \text{H}_2\text{O}\]

c) Balance the hydrogen by adding H⁺.
   
   \[2 \text{HClO} + 2 \text{H}^+ \rightarrow \text{Cl}_2 + 2 \text{H}_2\text{O}\]

d) Balance the charges by adding electrons.
   
   \[2 \text{HClO} + 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{Cl}_2 + 2 \text{H}_2\text{O}\]

There are 2 mol of H⁺ involved per 2 mol of HClO, or 1 mol H⁺/mol HClO.

4. Answer: C

The following are the rules for assigning oxidation number:

a) The oxidation number of an atom in its elemental form is zero (e.g. N₂, Ag)

b) The oxidation number of a monatomic ion is equal to its charge (e.g. Ca²⁺ has an oxidation number of +2)

c) The sum of oxidation numbers in a molecule or formula is zero. The sum of oxidation numbers in a polyatomic ion is equal to the charge of the ion.

d) The oxidation number of Group IIA is +1 in all of its compounds. Group IIA has an oxidation number of +2 in all of its compounds.

e) Hydrogen has an oxidation number of +1 if combined with nonmetals and −1 if combined with a metal or boron.

f) Fluorine has an oxidation number of −1 in all of its compounds.

g) Oxygen has an oxidation number of −2 in all of its compounds except in peroxide (−1) and when combined with F (e.g. in OF₂, because F has an oxidation number of −1 in all of its compounds, it follows that O has an oxidation number of +2).

h) Halogens have an oxidation number of −1 if combined with metals, nonmetals (except O) and other halogens lower in the group.

Following these rules, we assign the oxidation numbers on each compound given.

\[\text{NH}_3: \text{H has an ON of } +1, \text{ so N has ON of } -3\]
\( \text{N}_2\text{O}_4 \): O has an ON of \(-2\) with a total of \(-8\) for 4 O, which means that 2 N should have a total of \(+8\), or ON of \(+4\) for each N

\( \text{NO}_3 \): O has an ON of \(-2\), so the ON of N is \(+6\)

\( \text{N}_2\text{H}_5^+ \): H has an ON of +1, so each N has an ON of \(-2\)

\( \text{N}_2 \): ON = 0

5. Answer: D

The reaction in a voltaic cell is one that has a positive cell potential, which means that the cathode reaction (where reduction occurs) must have a more positive reduction potential than the anode. Thus, A is incorrect. In a voltaic cell containing \( \text{Co/Co}^{3+} \) and \( \text{Ni/Ni}^{2+} \), cobalt is the cathode since it has a more positive electrode potential.

The relative strength of oxidizing/reducing agent is determined by the standard reduction potential. The higher the potential, the greater the tendency for reduction to occur, and therefore, the stronger the oxidizing agent is (or weaker the reducing agent). Thus, C is incorrect since cobalt has a more positive reduction potential, is more likely to be reduced, and cobalt a stronger oxidizing agent than copper.

A substance with a more positive reduction potential can oxidize those that have a less positive reduction potential. With cobalt and nickel, cobalt is more likely to be reduced and therefore will oxidize nickel. Therefore, B is incorrect. With nickel and copper, nickel is more likely to be oxidized and therefore would reduce copper. Thus, E is incorrect.

D is correct since nickel has a less positive reduction potential, is more likely to be oxidized, and therefore is a stronger reducing agent, than copper.

6. Answer: C

Substances that have a more positive electrode potential can oxidize those that have a less positive electrode potential. \( \text{Au}^{3+} \) can oxidize both Ni And Ag, \( \text{Fe}^{2+} \) can oxidize neither Ni nor Ag, and \( \text{Cu}^{2+} \) can oxidize Ni but not Ag.
7. Answer: **C**

First, we write the overall reaction for the voltaic cell. We multiply the first equation by 2 to balance the number of electrons and add the two equations to get the overall equation:

\[ 2 \text{Ag}^+ + \text{Cu} \rightarrow 2 \text{Ag} + \text{Cu}^{2+} \]

We can use Le Chatelier’s principle to predict the effect of each of the changes to the cell potential. Any change that favors the forward reaction will increase the cell potential.

Using a bigger or smaller electrode will not affect the position of the equilibrium in the overall reaction since the electrodes are solid, and solids are not included in equilibrium equations.

Increasing the volume of 1.0 M Cu\(^{2+}\) solution will not affect the position of the equilibrium, as the concentration remains 1.0 M, there is just more solution.

Increasing the concentration of Cu\(^{2+}\) would shift the equilibrium to the left, thus decreasing the cell potential.

Increasing the concentration of Ag\(^+\) would shift the equilibrium to the right, thus increasing the cell potential.

8. Answer: **E**

First, let’s assign oxidation numbers to all atoms in the reaction to determine which substance is oxidized or reduced.

<table>
<thead>
<tr>
<th>Oxidation Number</th>
<th>2Al(^{3+})</th>
<th>+3</th>
<th>3Zn</th>
<th>+2</th>
<th>3Zn(^{2+})</th>
<th>+0</th>
<th>2Al</th>
</tr>
</thead>
</table>

The ON of Al decreased from +3 to 0 in going from Al\(^{3+}\) to Al, therefore Al\(^{3+}\) is reduced. The ON of Zn increased in going from Zn to Zn\(^{2+}\) by 2 so Zn is oxidized.

Looking at the forward reaction, Al\(^{3+}\) oxidizes Zn to Zn\(^{2+}\), and Zn reduces Al\(^{3+}\) to Al. Therefore, Al\(^{3+}\) is the oxidizing agent (it is reduced).

In the reverse reaction, Al reduces Zn\(^{2+}\) to Zn and Zn\(^{2+}\) oxidizes Al to Al\(^{3+}\). Therefore, in the reverse reaction Zn\(^{2+}\) is the oxidant (it is reduced).

Both A and B are incorrect.
A negative cell potential indicates a nonspontaneous process (or a spontaneous process in the reverse reaction). Therefore, the reduction of $\text{Al}^{3+}$ by $\text{Zn}$ is nonspontaneous.

The cell potential of an electrochemical cell is given by

$$E_{\text{cell}} = E_{\text{reduction}} + E_{\text{oxidation}}$$

The overall cell potential is negative, which means that the standard electrode potential of $\text{Al}$ is less positive (or more negative) than that of $\text{Zn}$, and statement D is incorrect.

In the reverse reaction, $\text{Al}$ is oxidized to $\text{Al}^{3+}$ and this reverse reaction is spontaneous. $\text{Al}$ is more likely to be oxidized than $\text{Zn}$. This means that $\text{Al}$ is a stronger reducing agent (the smaller the electrode potential, the stronger the reducing agent is) than $\text{Zn}$. $\text{E}$ is your answer.

9. Answer: C

A voltaic cell is one that uses a spontaneous redox reaction to perform electrical work. A spontaneous redox reaction is one with a positive cell potential.

We are given two half reactions and asked to predict the overall reaction and the cell potential for the voltaic cell. The cell potential for an electrochemical cell is given by

$$E_{\text{cell}} = E_{\text{reduction}} + E_{\text{oxidation}}$$

In order for $E_{\text{cell}}$ to be positive, the half-reaction with a more positive potential must be the cathode (the site of reduction). Thus, $\text{Pb}^{2+}/\text{Pb}$ must be the cathode (reduction) and $\text{Cd}^{2+}/\text{Cd}$ must be the anode (oxidation). So we flip the $\text{Cd}$ half-reaction (and flip the sign of the $\text{Cd}$ half reaction so we have our oxidation potential) and add the two equations to get the overall reaction:

$$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb} \quad -0.13 \text{ V}$$

+ 

$$\text{Cd} \rightarrow \text{Cd}^{2+} + 2e^- \quad 0.40 \text{ V}$$

$$\text{Pb}^{2+} + \text{Cd} \rightarrow \text{Pb} + \text{Cd}^{2+} \quad E_{\text{cell}} = -0.13 \text{ V} + 0.40 \text{ V} = +0.27 \text{ V}$$

10. Answer: B

The cell potential for an electrochemical cell is given by
We are given the overall cell potential and the potential of Cu$^{2+}$. From the provided equation, we see that Cu$^{2+}$ is reduced to Cu while Cr is oxidized to Cr$^{3+}$, so copper is the cathode and chromium is the anode. Substituting the given into the equation above:

$$E_{cell} = E_{reduction} + E_{oxidation}$$

$$+1.08 \, V = +0.34 \, V + E_{ox}$$

$$E_{ox} = 0.74 \, V$$

The question asks us to find the $E_{red}$ and we have calculated the $E_{ox}$, therefore we must flip the sign to give us $E_{red} = -0.74 \, V$. 
Chapter 13 Quiz: Nuclear Reactions

1. The mass number of a radioactive element decrease by 4 and its atomic number decreases by 2 if it undergoes __________.
   A. alpha decay
   B. beta emission
   C. positron capture
   D. electron capture
   E. positron emission

2. Which of the following radioactive processes results in an increase in atomic number?
   I. alpha decay
   II. positron emission
   III. beta decay
   A. I and II only
   B. I only
   C. III only
   D. II and III only
   E. I, II, and III

3. What is formed when $^{145}\text{Pm}$ decomposes by emitting a beta particle?
   A. $^{145}\text{Nd}$
   B. $^{145}\text{Sm}$
   C. $^{146}\text{Pm}$
   D. $^{146}\text{Sm}$
   E. $^{146}\text{Nd}$
4. When $^{137}$Ba is bombarded with a neutron particle, a certain particle is emitted and $^{137}$Cs is formed. What is the emitted particle?

A. alpha particle  
B. beta particle  
C. proton  
D. neutron  
E. positron

5. $^{210}$Pb undergoes a beta decay to form $^{210}$Bi with a half-life of 22 years. How long will it take for a 2 g sample of $^{210}$Pb to decay to 0.25 g?

A. 22 years  
B. 44 years  
C. 11 years  
D. 66 years  
E. 88 year

6. The reaction $A \rightarrow B$ follows first order kinetics and has a half life of 100 s. How long would it take for the concentration of A to decrease from 1.0 M to 0.125 M?

A. 100 s  
B. 200 s  
C. 300 s  
D. 400 s  
E. 500 s
ANSWER KEY

1. Answer: A

Alpha particles contains 2 protons and 2 neutrons. It is identical to a helium-4 nuclei and is symbolized as \( _2^4\text{He}^{2+} \). Alpha decay involves the release of an alpha particle from the nucleus, and results in a decrease in atomic number by 2 and a decrease in the mass number by 4. For example, uranium-238 decays to thorium-234 by alpha decay:

\[
^{238}_{92}\text{U} \rightarrow _2^4\text{He}^{2+} + ^{234}_{90}\text{Th}
\]

2. Answer: C

Alpha decay involves the release of an alpha particle, \( _2^4\text{He}^{2+} \), from the nucleus resulting in a decrease in atomic number by 2 of the parent nuclide.

Positron emission involves the loss of a positron, or \( _0^+\text{e} \) or \( ^0_1\text{b} \) particle. The atomic number of a nuclide that undergoes positron emission decreases by 1.

Beta decay occurs through the loss of a beta particle, \( _{-1}^0\text{e} \). During a beta decay, a neutron in the nucleus is converted into a proton and an electron. The proton remains in the nucleus and the electron is ejected as a beta particle. The atomic number therefore increases by 1.

3. Answer: B

Beta decay occurs through the loss of a beta particle, \( _{-1}^0\text{e} \). During beta decay, a neutron in the nucleus is converted into a proton and an electron. The proton remains in the nucleus and the electron is ejected as a beta particle.

When \(^{145}_{61}\text{Pm}\) undergoes beta decay, one of the product is a beta particle. To find the nuclide that is formed, we simply balance the mass number and atomic number on both sides of the equation. The atomic number of Pm is 61, and beta particle is \( _{-1}^0\text{e} \), so we write:

\[
^{145}_{61}\text{Pm} \rightarrow _{-1}^0\text{e} + __
\]

We can see that the missing product will have an atomic number of 62 and a mass number of 145, or \(^{145}_{62}\text{Sm}\).
4. Answer: C

We write the nuclear transmutation reaction:

\[ ^{137}_{56}\text{Ba} + ^0_1n \rightarrow ^{137}_{55}\text{Cs} + ? \]

We balance the mass number and the atomic number on both sides of the equation to determine the emitted particle. We see that the emitted particle has Z = 1 and A = 1, or a proton, \(^1p\):

\[ ^{137}_{56}\text{Ba} + ^0_1n \rightarrow ^{137}_{55}\text{Cs} + ^1_1p \]

5. Answer: D

Radioactive decay follows first-order kinetics, so its half life is a constant. The half-life is the time it takes for half of a sample to decompose. If we start with 2 g sample, it will take 3 half-lives (or \(2 \times 3 = 66\) years) for the sample to decay to 0.25 g (2 g → 1 g → 0.5 g → 0.25 g).

6. Answer: C

The half-life of a reaction is the time it takes to decrease the amount of reactants to half its original value. We are told the half life is 100 s.

After 100s, [A] decreases from 1.0 M to 0.5 M; after another 100 s, [A] decreases from 0.5 M to 0.25 M, another 100 s, [A] decreases from 0.25 M to 0.125 M, or 3 half-lives.

Decreasing the concentration of A from 1.0 M to 0.125 M will therefore take 300 s.