# Midterm Review Short Answer

Base your answers to questions **1** and **2** on the information below.

The accepted values for the atomic mass and percent natural abundance of each naturally occurring isotope of silicon are given in the data table below.

**Naturally Occuring Isotopes of Silicon**

| **Isotope** | **Atomic Mass** (atomic mass unit) | **Percent Natural Abundance** (%) |
|---|---|---|
| $Si-28$ | 27.98 | 92.22 |
| $Si-29$ | 28.98 | 4.69 |
| $Si-30$ | 29.97 | 3.09 |

1. Show a correct numerical setup for calculating the atomic mass of Si.

2. Determine the total number of neutrons in an atom of Si-29.

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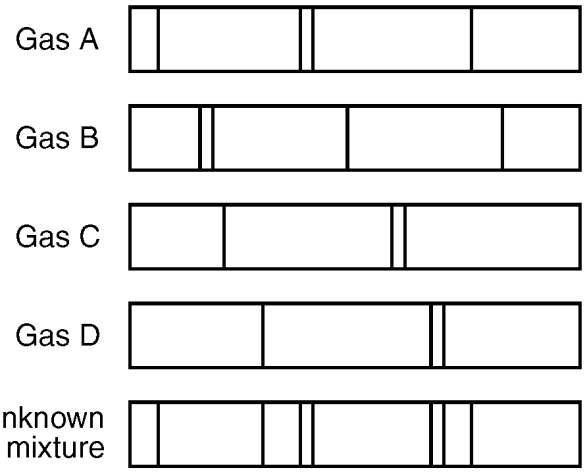
3. Write an electron configuration for an atom of aluminum-27 in an excited state.

4. Draw a Lewis electron-dot diagram for a sulfur atom in the ground state.

5. Base your answer to the following question on the information and the bright-line spectra represented below.

Many advertising signs depend on the production of light emissions from gas-filled glass tubes that are subjected to a high-voltage source. When light emissions are passed through a spectroscope, bright-line spectra are produced.




Explain the production of an emission spectrum in terms of the *energy states of an electron*.

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Base your answers to questions **6** through **8** on the information below.

A metal, *M*, was obtained from a compound in a rock sample. Experiments have determined that the element is a member of Group 2 on the Periodic Table of the Elements.

6. Explain why the radius of a positive ion of element *M* is *smaller* than the radius of an atom of element *M*.

7. Explain, in terms of electrons, why element *M* is a good conductor of electricity.

8. What is the phase of element *M* at STP?

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Base your answers to questions **9** through **11** on the information below.

Rust on an automobile door contains $Fe_2O_3(s)$. The balanced equation representing one of the reactions between iron in the door of the automobile and oxygen in the atmosphere is given below.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

9. Write the IUPAC name for $Fe_2O_3$.

10. Determine the gram-formula mass of the product of this reaction.

11. Identify the type of chemical reaction represented by this equation.

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12. Given the balanced equation:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

What is the total number of moles of $O_2(g)$ that must react completely with 8.0 moles of Al(s) in order to form $Al_2O_3(s)$?

13. A student heats a 243 gram sample of $BaCl_2 \cdot 2H_2O$ hydrated crystals?

*a* What was the percentage by mass of water in the hydrated crystals?[Show all work]

*b* If all the water was driven out of the crystal, what would be the expected weight of the $BaCl_2$ sample remaining?[Show all work]

14. Explain, in terms of electronegativity, why a P–Cl bond in a molecule of $PCl_5$ is more polar than a P–S bond in a molecule of $P_2S_5$.

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15. Base your answer to the following question on the table below.

**Physical Properties of Four Gasses**

| **Name of Gas** | hydrogen | hydrogen chloride | hydrogen bromide | hydrogen iodide |
|---|---|---|---|---|
| **Molecular Structure** | H-H | H-Cl | H-Br | H-I |
| **Boiling Point (K) at 1 Atm** | 20. | 188 | 207 | 237 |
| **Density (g/L) at STP** | 0.0899 | 1.64 | ? | 5.66 |

Explain, in terms of intermolecular forces, why hydrogen has a *lower* boiling point than hydrogen bromide.

16. Base your answer to the following question on the information below.

Each molecule listed below is formed by sharing electrons between atoms when the atoms within the molecule are bonded together.

Molecule *A*: $Cl_2$ Molecule *B*: $CCl_4$ Molecule *C*: $NH_3$

Explain why $NH_3$ has stronger intermolecular forces of attraction than $Cl_2$.

17. Draw the electron-dot (Lewis) structure of calcium chloride.

# Answer Key
# Midterm Short answer review

1. (27.98)(0.9222) + (28.98)(0.0469) + (29.97)(0.0309)
2. 15
3. *Examples:* – 2-7-4 – 1-8-4 – 2-6-2-3
4. 

5. Acceptable responses: Energy is released when an electron falls from a high state (excited) to a low state (ground), excited state to ground state, high energy to low energy.
6. Examples: – The ionic radius is smaller because the atom loses two electrons. – The ion has one less occupied energy level.
7. Examples: – Metals have freely moving valence electrons. – mobile valence electrons – sea of mobile electrons – Electrons are delocalized.
8. solid
9. iron(III) oxide
10. 160.g/mol.
11. *Examples:* – synthesis – redox – oxidation
12. 6
13. a) 36/ (137 + 71 + 36) × 100 = 15%  b) 207 grams
14. A P–Cl bond is more polar than a P–S bond because the electronegativity difference for P–Cl is 1.0 and the electronegativity difference for P–S is 0.4.
15. Examples: – Hydrogen has weaker intermolecular forces than HBr. – hydrogen – weaker forces.
16. Examples:
– $NH_3$ has polar molecules that attract each other.
– $NH_3$ has an unshared pair of electrons around the center atom.
– $NH_3$ is capable of hydrogen bonding.
– unequal distribution of electrons — in strong attraction
17. examples: