Hund's rule: The most stable configuration of electrons in a subshell is to maximize the total spin. This means that electrons will fill orbitals within each subshell one at a time, until each orbital has one electron, all with the same spin. For example, in filling the 7 electrons of N, you would fill orbitals as follows:

Rather than

\[
\begin{array}{c}
1s \\
2s \\
2p
\end{array}
\]

\[
\begin{array}{c}
1s \\
2s \\
2p
\end{array}
\]

\[
\begin{array}{c}
1s \\
2s \\
2p
\end{array}
\]

Or

\[
\begin{array}{c}
1s \\
2s \\
2p
\end{array}
\]

Ge: This location (Period 4, Group 4A) specifies its electronic configuration completely.

\[
[\text{Ar}]^{18}3\text{~d}^34\text{~s}^24\text{~p}^2
\]

(by Hund) Group 4A

Mg: 

\[
[\text{Ne}]^{10}\text{~3s}^2
\]

\[
\begin{array}{c}
1s \\
2s \\
2p \\
3s
\end{array}
\]

Ar: 

\[
[\text{Ar}]^{18}
\]

\[
\begin{array}{c}
1s \\
2s \\
2p \\
3s \\
3p
\end{array}
\]

The group # tells you how many e- are in the outmost shell, while the period specifies which shell forms the outmost one (eg: n = 3)
10. Ta

\[ \text{[Xe]}^{54} \text{4f}^{14} \text{5d}^3 \text{6s}^2 \]

(Hint: Ta has 73 e\(^-\), so use this to help you remember the 4f electrons.)

Pt

\[ \text{[Xe]}^{54} \text{4f}^{14} \text{5d}^8 \text{6s}^2 \]

24. Mn

(a) \[ \text{[Ar]}^{18} \text{3d}^5 \text{4s}^2 \]

(b) \[ \text{Mn}^{2+} \to \text{[Ar]}^{18} \text{3d}^3 \text{4s}^2 \]

(c) Yes. 5 \text{np} unpaired electrons, \to \text{paramagnetic}

(d) \[ \text{Cu}^{1+} \text{[Ar]}^{18} \text{3d} \text{4s} \]

[Cu\(^{2+}\) is not paramagnetic, Cu\(^{1+}\) is paramagnetic]
33. Ti  
\[ [Ar]^{18} 4s^2 3d^1 \]  
\[ n = 3 \]
\[ l = 2 \]
\[ m_l = \pm 2, \pm 1, 0, -1, -2 \]
\[ m_s = \pm \frac{1}{2}, \pm \frac{1}{2} \]  
\[ 4s \]
\[ n = 4 \]
\[ l = 0 \]
\[ m_l = 0 \]
\[ m_s = \pm \frac{1}{2}, \pm \frac{1}{2} \]  
\[ 7L \]

36. (b) Cl < Cl⁻  
(atomic radius decreases as you go down the periodic table, and to the left)

(b) O < Al  
(atomic radius increases as you go down the periodic table, and to the left)

(c) In > I  
(atomic radius increases as you go to the left)

(46) Ce  
\[ [Xe]^{54} 5d^1 \]  
\[ Ce^{3+} [Xe]^{54} 4f^1 \]  
\[ \text{\rightarrow} [Xe]^{54} \text{ } 4f \]

(6) Ho  
\[ [Xe]^{54} 10 \text{ } 5d^1 \]  
\[ \text{Ho}^{3+} [Xe]^{54} 4f^{10} \]  
\[ \text{\rightarrow} [Xe]^{54} \text{ } 4f \]
**Atomic Radius**
- Na > Mg > P > O

**EA (Electron Affinity)**
- The most negative EA is O
- (O + e⁻ → O⁻) releases more energy

**IE (Ionization Energy)**
- Na < Mg < P < O
- easiest to remove
- (Na → Na⁺ + e⁻ requires least input of energy)

**S > O > F**
- Atomic radius reflects the effective core charge felt by the outer shell of e⁻
- Thus, O is bigger than F, as O has a lower effective core charge
- Atomic radius also reflects the fact that electron shells are bigger for higher n principle quantum numbers
- Thus, S is bigger than O, as S has n = 3, while O has n = 2

**IE (M → M⁺ + e⁻, IE = ΔE_(ion))**
- Moving to the right, through a period, each successive element has an increasing Z_{eff} (or Z*), making it harder to remove an electron (so IE becomes more positive)
- The twist to IE turns in that e⁻-e⁻ repulsions make it a little easier to remove one e⁻ from a doubly-occupied orbital
- And finally, bigger principle quantum numbers (increasing n) make it easier to remove one e⁻, causing the IEs to become less positive
- So, the easy part is to set the period 3 elements in or by (Si < P < S)
- Hence the e⁻-e⁻ repulsions make it easier than expected to ionize S
- Where to put Se? It will be easier to ionize Se than S
- So, the final answer is: **S > Se > P > Si**

**Ionic Radius**
- N³⁻ > O²⁻ > F⁻
- The more charges added, the bigger the ion
- Particularly true when Z_{eff} also decrease (Z_{eff} for N < Z_{eff} F)

**Sr > Ba > Cs**
- IEs increase as you move up a group, and the increase again move to the right across a period
(54) \[ Cl^- < Cl < Ca^{2+} \] IEs are the energy to remove one from a species, and reflect the attraction between the e^- & the nucleus. This series here is set due to the relative charge of each species: Cl^- has an extra electron, so it is more easily removed than in the case of Cl. The Ca^{2+} ion would need to undergo a third ionization, which is very unfavorable.

(55) a. metal
   b. Element B
   c. Element C
   d. element A

(60) isoelectronic (same # of orbital shells) of e^- \[ Cl^- = K^+ = Ca^{2+} \]
   a. size \[ Cl^- > K^+ > Ca^{2+} \]
   b. IEs \[ Ca^{2+} > K^+ > Cl^- \]
   c. EAs \[ Ca^{2+} > K^+ > Cl^- \]

(63) IEs decrease as you move down a group due to the larger separation between the nucleus and the outermost e^- (remember, Coulomb's Law is \( Ex^{-1} \))

(64) a. Co
   b. paramagnetic
   c. Co^{3+} has 4 unpaired e^- in the gas phase.
This is due to the 2p orbitals being at higher energy than 1s.

Thus, Be (1s² 2s² 2p²) is more easily ionized than Be (1s² 2s²).

This 'dip' in 1Es is found as each new subshell begins, and as each new subshell begins to have paired e⁻s in orbitals (see the 1Es for N vs O).

This trend reflects the e-e repulsion that begins at 5.

ionization makes it easier than expected to ionize the element.

Thus, S is easier to ionize than expected.

Na(s) + ½ Cl₂ (g) → NaCl (s)

Cl⁻ Na⁺

(1) The Na₂ Ion is the reducing agent because it has a relatively low IE (easy to donate an e⁻).

(2) Cl is the reducing agent. Cl₂ has a more negative EA, meaning that it 'picks' to accept an e⁻.

(3) NaCl has Na⁺ & Cl⁻ ions, each of which exhibit a noble gas 'full shell' electron configuration. Na₂Cl or NaCl₂ would require that either Cl₂ or Na⁺ be stable, by having each does not have the 'full shell'.

Do the modelling exercise. But, yes, there is reasonably good agreement between calculated bond lengths \( (r_a + r_b = \text{Bond length for } A-B) \) and observed bond lengths.