1. If a 40.1 g piece of iron at 652 °C is dropped into a sample of 328 g of water at 32.4 °C, what will the final temperature be after thermal equilibrium is established? Assume that no heat is lost during the process.

**Answer:** 40. °C

**Analysis of the situation:** What happened: 328 g of water went from 32.4 °C to some (as yet unknown) final temperature \((T_{f, \text{water}})\) as 40.1 g of iron (Fe) went from 652 °C to some (as yet unknown) final temperature \((T_{f, \text{Fe}})\).

**Conclusions from the above:**
1) Since the iron was initially hotter than the water, some heat energy \(q\) must have transferred from the iron to the water during the described process.
2) The value of \(q\) that flowed into the water must equal the value of \(q\) that flowed out from the iron (since there was only one “amount” of heat that flowed). In mathematical terms, this can be stated as: \(q_{\text{water}} = -q_{\text{Fe}}\) (or \(q_{\text{Fe}} = -q_{\text{water}}\); either is fine.)
3) Since the two samples end up at thermal equilibrium, the two final temperatures noted above must be the SAME temperature. That is, \(T_{f, \text{water}} = T_{f, \text{Au}}\), and so I will refer to these both as simply \(T_f\) from here on out in this problem.

**Strategy:**
1) Use the knowledge (equation) that the amount of heat transferred from (or to) a sample is equal to the product of its specific heat, mass, and change in temperature. Write this product for both \(q_{\text{water}}\) and \(q_{\text{Fe}}\) (obtaining the specific heat capacity of iron from a table in your text)
2) If you then set \(q_{\text{water}}\) equal to the opposite of \(q_{\text{Fe}}\) (see (2) above in conclusions), you will see that there is only one unknown left to solve for: \(T_f\).
3) Use algebra to solve for the value of \(T_f\).

**Execution of Strategy:**

\[
q_{\text{water}} = -q_{\text{Fe}}
\]

\[
C_{s,\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}} = -\left( C_{s,\text{Fe}} \times m_{\text{Fe}} \times \Delta T_{\text{Fe}} \right)
\]

\[
\Rightarrow \quad \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ \text{C}} \times 328 \text{ g} \times (T_f - 32.4 \ ^\circ \text{C}) = -\left[ \frac{0.45 \text{ J}}{\text{g} \cdot ^\circ \text{C}} \times 40.1 \text{ g} \times (T_f - 652) \right]
\]

\[
\Rightarrow \quad \frac{1372.4 \text{ J}}{^\circ \text{C}} \times (T_f - 32.4 \ ^\circ \text{C}) = -\frac{18.05 \text{ J}}{^\circ \text{C}} \times (T_f - 652)
\]

Dividing through by 18.05 J/^\circ C yields (you could do the algebra other ways as well [i.e., distribute on both sides]):

\[
\frac{(1372.4 \text{ J}/^\circ \text{C})}{(18.05 \text{ J}/^\circ \text{C})} (T_f - 32.4 \ ^\circ \text{C}) = -(T_f - 652 \ ^\circ \text{C})
\]

\[
\Rightarrow \quad 76.0(T_f - 32.4 \ ^\circ \text{C}) = -T_f + 652 \ ^\circ \text{C}
\]

\[
\Rightarrow \quad 76.0T_f - 2463 \ ^\circ \text{C} = -T_f + 652 \ ^\circ \text{C}
\]

\[
\Rightarrow \quad 77.0T_f = 652 \ ^\circ \text{C} + 2463 \ ^\circ \text{C} = 3115 \ ^\circ \text{C}
\]
2. How much heat (in kJ) is evolved or absorbed in the reaction of 233.0 g of carbon with enough CaO to produce calcium carbide? (b) Is the process exothermic or endothermic?

\[ \text{CaO(s)} + 3 \text{C(s)} \rightarrow \text{CaC}_2(s) + \text{CO(g)}; \quad \Delta H^\circ = 464.8 \text{ kJ} \]

(a) **Answer:** 3006 kJ

**Strategy:**
1) Determine the number of moles of C in 233.0 g of it (by using the molar mass of C).
2) Recognize that the “464.8 kJ” here means that “464.8 kJ is absorbed when three moles of C(s) reacts with 1 mol of CaO(s) to produce.....” and use that ratio to figure out how much energy is absorbed when \(x\) moles of C(s) are reacted (where \(x\) is the number determined in (1)).

**Execution of Strategy:**
1) The molar mass of C is 12.01 g/mol. So 233.0 g is roughly 20 moles (233 is about 20 times bigger than 12). More precisely, it is 233.0 g / (12.01 g/mol) = 19.40 mol or:

\[ 233.0 \text{ g} \times \frac{1 \text{ mol C}}{12.01 \text{ g}} = 19.400 \text{ mol C} \]

2) \[ 19.400 \text{ mol C} \times \frac{464.8 \text{ kJ (of heat absorbed)}}{3 \text{ mol C (reacted)}} = 3005.7 \text{ kJ absorbed} \]

(b) The process is **endothermic.** A positive value of \(\Delta H_{sys}\) indicates that H increases for the system, and thus energy flows in as the process occurs. That describes endothermic.

3. 85.8 kJ of energy is evolved (i.e., released) at constant pressure when 3.56 g of P₄ is burned according to: \( \text{P}_4(s) + 5 \text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(s) \) What is the \(\Delta H\) for the (thermo)chemical equation?

**Answer:** \(-2.99 \times 10^3 \text{ kJ}\)

**Reasoning:**
Since the coefficient of $P_4$ is “1” in the balanced equation, you need to find the amount of energy released when ONE MOLE of $P_4$ is burned to get the magnitude of the $\Delta H$ for the (thermo)chemical equation. How many moles is 3.56 g of $P_4$? Molar mass of $P_4 = 4(30.97\text{ g/mol}) = 123.9\text{ g/mol }P_4$. So:

$$\frac{3.56\text{ g}}{123.9\text{ g/mol}} = 0.02873\text{ mol }P_4$$

Thus the energy released PER MOLE is just (remember the “How many numbers do you see in the fraction on the left below? Just one….you just haven’t done the division yet!” idea):

$$\frac{85.8\text{ kJ released}}{0.02873\text{ mol }P_4\text{ burned}} = \frac{2986\text{ kJ released}}{\text{mol }P_4\text{ burned}} = 2.99 \times 10^3\text{ kJ (released per mol }P_4)$$

Since energy is released, the sign of $\Delta H$ is negative. Thus, $\Delta H =\boxed{-2.99 \times 10^3\text{kJ}}$.

4. Instant cold packs contain solid $\text{NH}_4\text{NO}_3$ and a pouch of water. When the pack is squeezed, the pouch breaks and the solid dissolves, lowering the temperature because of the endothermic process:

$$\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq});\quad \Delta H = +25.7\text{ kJ}$$

What is the final temperature in a squeezed cold pack that contains 50.0 g of $\text{NH}_4\text{NO}_3$ dissolved in 125 mL of water? Assume a specific heat of 4.18 J/(g·°C) for the solution, an initial temperature of 25.0 °C, and no heat transfer between the cold pack and the environment. (Recall, $d_{\text{water}} \approx 1.0\text{ g/mL}$)

**Answer:** $-5.73\text{ °C}$ (or 3.1 °C; see comment below*)

**Strategy:** Recognize that this is a “heat flow” kind of problem, which means:

$$q_{\text{sys}} = -q_{\text{surr}}$$

Since a process is involved, let the system be the atoms/molecules/ions that are involved in the process. Here that would be the $\text{NH}_4\text{NO}_3$ (and the few water molecules that solvate the ions, once dissolved). Thus $q_{\text{sys}}$ becomes $q_{\text{process}} = \Delta H_{\text{sys}}$. The surroundings is thus the (bulk) water (or the whole “cold pack”; see comment*). Thus the above equation simplifies to:

$$\Delta H_{\text{sys}} = -q_{\text{water}} \quad (or \quad -q_{\text{cold pack}}; \text{ see comment*}) \quad \text{which becomes}$$

$$\Delta H_{\text{sys}} = -(C_{\text{s,water}} \times m_{\text{water}} \times \Delta T_{\text{water}}) \quad (where \Delta T_{\text{water}} = T_{\text{final}} - T_{\text{initial}} \text{ for the water})$$

Knowing the mass of $\text{NH}_4\text{NO}_3$ that “reacts” (strictly speaking this is actually a physical change, but we can view the solid as a “reactant” here) and knowing the enthalpy change for when one MOLE of it “reacts” (i.e., understanding the meaning of a $\Delta H$ that comes after a chemical equation) allows you to figure out $\Delta H_{\text{sys}}$ (i.e., how much energy was absorbed by the system as this specific amount of dissolution (50 g) takes place). $m_{\text{water}}$ can be obtained from $V_{\text{water}}$ and the density of water. $s_{\text{water}}$ is given (see comment*) and $T_{\text{initial}}$ is given. Thus the only unknown is $T_{\text{final}}$.

**Execution of Strategy:**

1) $50.0\text{ g/(80.0}\text{ g/mol}) = 0.625\text{ mol }\text{NH}_4\text{NO}_3$
2) $(0.625 \text{ mol NH}_4\text{NO}_3)(25.7 \text{ kJ/mol}) = 16.06 \text{ kJ} = 16.06 \times 10^3 \text{ J absorbed} = \Delta H_{\text{sys}}$

3) \[ \Delta H_{\text{sys}} = -q_{\text{water}} \]

\[ \frac{16.06 \times 10^3 \text{ J}}{4.18 \text{ J/}\text{g }\cdot{\degree}\text{C}} \left( \frac{125 \text{ mL} \times 1.00 \text{ g}{\text{mL}}}{125 \text{ g}} \right) \left( T_f - 25.0 \degree{\text{C}} \right) \]

\[ \Rightarrow \quad \frac{16.06 \times 10^3 \text{ J}}{4.18 \text{ J/}\text{g }\cdot{\degree}\text{C}} \times T_f - 25.0 \degree{\text{C}} = 30.73 \degree{\text{C}} = T_f - 25.0 \degree{\text{C}} \]

\[ \Rightarrow \quad T_f = -30.73 + 25.0 \degree{\text{C}} = -5.73 \degree{\text{C}} \] (SEE NOTE BELOW*)

*NOTE: If you assume the specific heat capacity of the solution is 4.18 J/(g.\degree{\text{C}}) [in effect, if you assume that the specific heat capacity of the dissolved salt is the same as that of water, which it clearly isn't] and use a solution mass of 175 g (i.e., including the mass of the NH$_4^+$ and NO$_3^-$ ions), the answer obtained is about 3.1 \degree{\text{C}} for the final temperature. The actual "answer" is likely some value in between—I think it would be closer to -5.7 than +3.1 because the specific heat capacity of the ions is likely to much less than that of water, whose C$_s$ value is quite high compared to just about any other substance. In lab, we IGNORED the mass of the dissolved solutes when doing this sort of calculation which is why I did the current calculation that way in this key.

5. **8.45.** Assume that a particular reaction produces 244 kJ of heat and that 35 kJ of PV (expansion/contraction) work is done on the system. What are the values of $\Delta E$ and $\Delta H$ for the system? For the surroundings?

**Answers:** $\Delta E_{\text{sys}} = -244 + 35 = -209 \text{ kJ}$; $\Delta H_{\text{sys}} = -244 \text{ kJ}$; $\Delta E_{\text{surr}} = +209 \text{ kJ}$; $\Delta H_{\text{surr}} = +244 \text{ kJ}$

**Reasoning:** The energy of a system can only be changed by two things: addition (or loss) of heat (energy) or by work being done on (or by) the system. Another way to say this is that $E_{\text{sys}}$ can go down if heat energy is lost (to the surroundings) and/or if work is done **by** the system (on the surroundings); $E_{\text{sys}}$ can go up if heat energy is gained (from the surroundings) and/or if work is done **on** the system (by the surroundings). So if 244 kJ of heat is “evolved” by the system, this means that 244 kJ of heat energy is lost, and so $E$ should go down by 244 kJ as a result of this change. If 35 kJ of work is gained by the system, $E$ should go up by 35 kJ as a result of this change. If both changes occur, the result is the sum of the two: -244 kJ + 35 kJ = -209; a net loss of 209 kJ (by the system to the surroundings) results. So $\Delta E_{\text{sys}} = -209 \text{ kJ}$ and $\Delta E_{\text{surr}} = +209 \text{ kJ}$. Since $\Delta H = q_p$, $\Delta H_{\text{sys}}$ is simply -244 kJ here and $\Delta H_{\text{surr}}$ is the opposite of this: +244 kJ.

Using an equation rather than verbal reasoning: $\Delta E_{\text{sys}} = q + w = -244 \text{ kJ} + 35 \text{ kJ} = -209 \text{ kJ}$ (the other results follow)